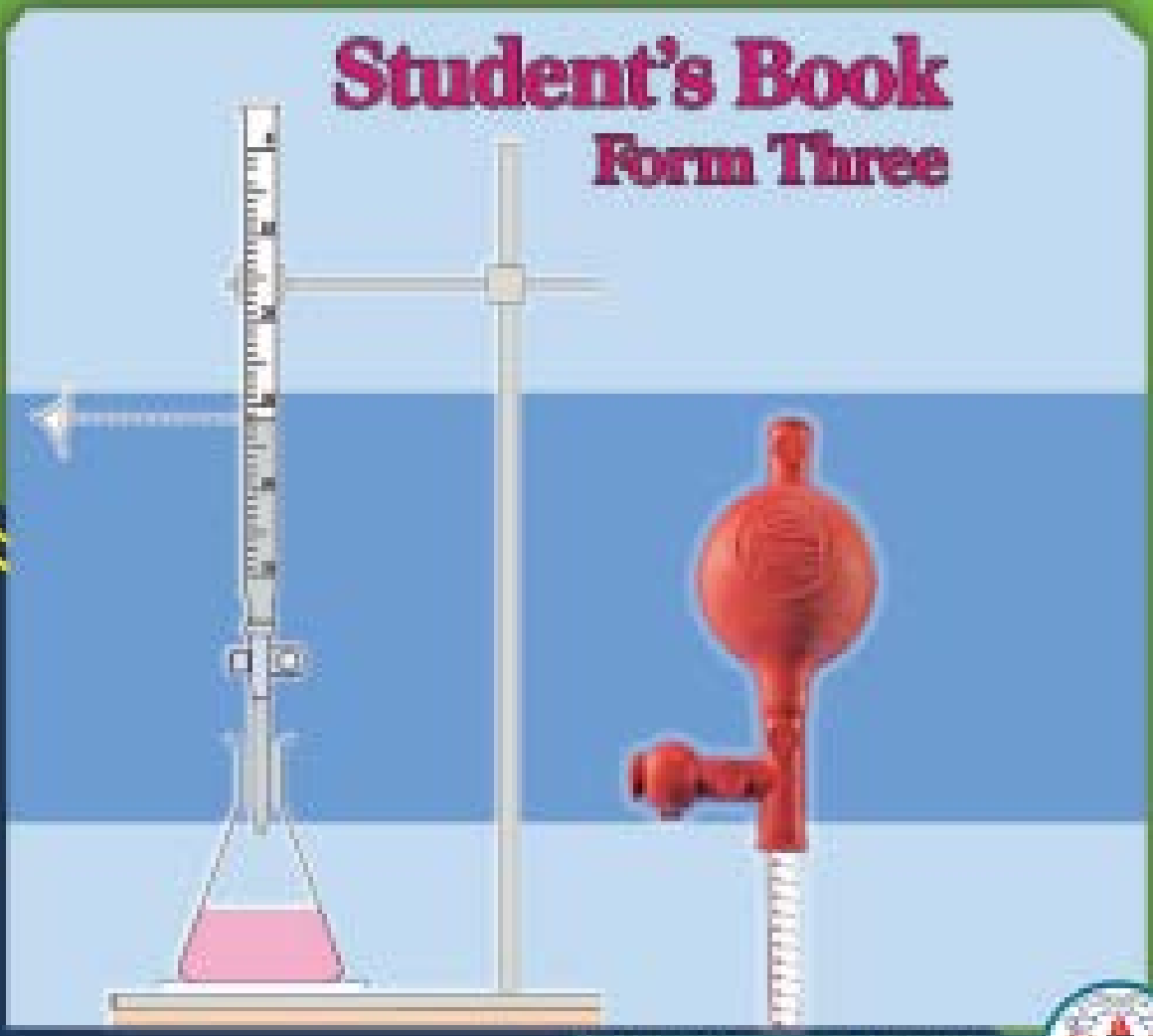
Chemistry

for Secondary Schools

Student's Book
Form Three



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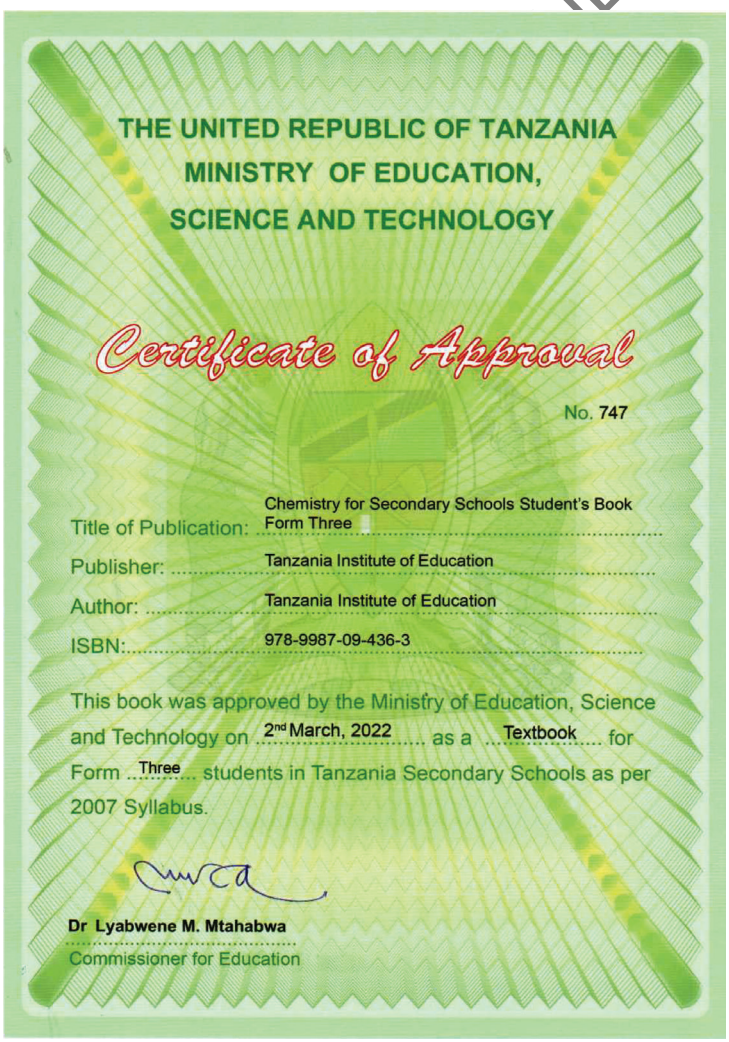
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Chemistry

for Secondary Schools

Student's Book

Form Three


THE UNITED REPUBLIC OF TANZANIA
MINISTRY OF EDUCATION,
SCIENCE AND TECHNOLOGY

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Dr Lyabwene M. Mtahabwa
Commissioner for Education

Tanzania Institute of Education

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Dr Aneth A. Komba
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Preface

This textbook, Chemistry for Secondary Schools, is written specifically for Form Three students in the United Republic of Tanzania. It is written in accordance with the 2007 Chemistry Syllabus for Ordinary Secondary Education, Form 1–IV, issued by the then, Ministry of Education and Vocational Training.

The book consists of ten chapters, namely Chemical equations, Hardness of water, Acids, bases, and salts, The mole concept and related calculations, Volumetric analysis, Ionic theory and electrolysis, Chemical kinetics, equilibrium, and energetics, Extraction and properties of metals, Compounds of metals; oxides, hydroxides, carbonates, and hydrogencarbonates, and Compounds of metals; nitrates, chlorides, and sulphates. The chapters comprise of illustrations, activities, projects, tasks, and exercises. You are encouraged to do all the activities, projects, tasks, and exercises as well as other assignments provided. Doing so, will enable you to develop the intended competencies.

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Chapter One

Chemical equations

Introduction

In Form Two, you learnt about chemical bonding and formulas. Chemical bonding gives a formula unit or molecule of a compound. These chemical species are involved in chemical reactions. Many chemical reactions take place in the environment. For example, burning fuel in car engines, burning of charcoal and wood, respiration in living organisms, and digestion of food in the body involve chemical reactions. Chemical reactions are represented by chemical equations. In this chapter, you will learn the concept of chemical reactions, concept of chemical equations, and types of chemical reactions. The competencies developed will enable you to present chemical equations in proper ways and analyse different chemical reactions that produce various products which are useful in daily life.

Concept of chemical reactions

A chemical reaction is a process in which one or more chemical substances are converted to one or more different substances. Chemical reactions take place when bonds between atoms in the reacting substance(s) are broken, the atoms rearrange, and new bonds between atoms are formed to make new substance(s). The chemicals that begin the process of a chemical reaction are called *reactants* and the new substances formed are called *products*. The products have properties that are different from their respective reactants. Features that indicate a chemical reaction has taken place include one or more of the following; evolution of a gas, formation of precipitate, and change of colour, temperature or state.

A chemical reaction is either reversible or irreversible. A reversible reaction proceeds in both forward and backward directions. In the forward reaction, the reactants are converted into products, whereas in the backward reaction, the products become the reactants. An irreversible reaction proceeds in one direction from reactants to products.

Task 1.1

Describe any five chemical reactions that occur in daily life.

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Concept of chemical equations

A chemical reaction is expressed in the form of a chemical equation. A chemical equation is a symbolic or words representation of a chemical reaction. A chemical equation written in words is referred to as a *word equation*. For example, when calcium metal reacts with chlorine gas to form solid calcium chloride, the word equation is written as:

$$\text{Calcium} + \text{Chlorine gas} \longrightarrow \text{Calcium chloride}$$

Word equations do not give details on the composition, quantities, states of the reactants and products, and the conditions for the reaction to take place.

A chemical equation written in symbols is referred to as a *formula equation*. For example, the formula equation for a reaction between calcium metal and chlorine gas is:

$$Ca(s) + Cl_2(g) \longrightarrow CaCl_2(s)$$

The formula equation is more useful than the word equation, although it is advantageous to write the equation in word form first. Formula equations provide the useful information including; compositions, amounts, formulas, and physical states of substances that are involved. Formula equations may also state the conditions for the reactions to take place. The reactants and products in the formula equations are either solids, liquids or gases. These states of matter are represented in a chemical equation by using symbols as; solid (*s*), liquid (*l*), and gas (*g*). When the substances that are involved in a reaction are dissolved in water, the word aqueous (*aq*) is used. These symbols should be included when writing the chemical equations and are stated in parentheses, ( ), after the chemical symbol.

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In a chemical equation, a headed arrow is used to separate the reactants and products. A full headed arrow is used to separate reactants and products for an irreversible reaction. For example;

$$\text{Reactants} \longrightarrow \text{Products}$$

$$\text{Reactant 1} + \text{Reactant 2} \longrightarrow \text{Product 1} + \text{Product 2}$$

Double half headed arrows pointing in the opposite directions are used to separate reactants and products for a reversible reaction. For example;

$$\text{Reactants} \underset{\text{backward reaction}}{\overset{\text{forward reaction}}{\rightleftharpoons}} \text{Products}$$

The arrows show the direction of a reaction; thus, it means 'produce' or 'yield'. Each individual substance is separated from the other by a plus sign (+). Note that, the number of reactants and products are not necessarily the same. A chemical equation

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has the following characteristics: reactants are on the left side of the arrow and products are on the right side, reactants are turned into products, and gives the relative numbers of reactant and product molecules or ions when balanced. Both the word equation and formula equation can be used to represent a molecular equation and ionic equation.

Molecular equations

A molecular equation is an equation representing a reaction in solution showing the reactants and products in undissociated form. In molecular equations, reactants and products are considered neutral regardless of their exact physical states. The following are examples of molecular equations:

$$PbO_2(s) + SO_2(g) \longrightarrow PbSO_4(s)$$

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

In principle, when writing molecular equations, the reactants and products should be balanced.

Balancing chemical equations

All chemical equations must be written in accordance with the *law of conservation of matter*. This law states that, *in a chemical reaction, the total mass of the products equals the total mass of the reactants.* This means, when balancing a chemical equation, the number of each atom on both sides of the equation must be equal because atoms do not varnish during a reaction, rather are reorganised.

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The following steps are followed when writing and balancing simple chemical equations:

1. Write the equation in a word form.
2. Write the equation by using symbols and ensure all the formulas are correct.
3. Balance the number of atoms of all the elements on each side of the equation. Balance the number of atoms by writing the appropriate number on the left side of a respective chemical formula.
4. Include the state symbols.

Example 1.1

Hydrogen chloride gas is formed when hydrogen gas burns in chlorine gas. Write a balanced chemical equation for the reaction.

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Solution

(a) Write a word equation:

Hydrogen gas + Chlorine gas ⟶ Hydrogen chloride gas

(b) Write a formula equation:

$$H_2 + Cl_2 \longrightarrow HCl$$

(c) Balance the formula equation:

The equation is not balanced because, hydrogen (H) has 2 atoms on the left and 1 atom on the right. Chlorine (Cl) has 2 atoms on the left and 1 atom on the right. Another molecule of hydrogen chloride (HCl) is needed on the right to balance the equation. Thus, 2 is put before HCl to give:

$$H_2 + Cl_2 \longrightarrow 2HCl$$

The equation is now balanced as hydrogen has 2 atoms on both sides and chlorine has 2 atoms on the left and 2 on the right hand side of the equation.

(d) Write the state symbols:

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$$

Example 1.2

Zinc dissolves in dilute hydrochloric acid to form zinc chloride solution and hydrogen gas. Write a balanced chemical equation for this reaction.

Solution

(a) Write a word equation:

Zinc + Dilute hydrochloric acid ⟶ Zinc chloride + Hydrogen gas

(b) Write a formula equation:

$$Zn + HCl \longrightarrow ZnCl_2 + H_2$$

(c) Balance the formula equation:

The equation is not balanced because, chlorine has 1 atom on the left and 2 atoms on the right and hydrogen has 1 atom on the left and 2 atoms on the right. To balance this equation, another molecule of HCl is needed on the left. Therefore, 2 is put before the HCl molecule.

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

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Zinc has 1 atom on the left and 1 atom on the right. Chlorine has 2 atoms on the left and 2 atoms on the right. Hydrogen has 2 atoms on the left and 2 atoms on the right, thus the equation is now balanced.

(d) Write the state symbols:

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

Ionic equations

An ionic equation is a chemical equation in which compounds in aqueous solutions or in molten state are written as dissociated ions. Ionic equations are commonly used in *displacement reactions* in aqueous solutions. In these equations, *spectator ions* are omitted to give a net ionic equation. *Spectator ions* are the ions that do not change their valence states in the reaction since they remain unchanged in a chemical reaction.

The following steps are followed when writing an ionic equation:

1. Write the chemical equation in words.
2. Write the balanced chemical equation in symbols; ensure all formulas are correct.
3. Split all soluble ionic compounds into individual ions to get total ionic equation.

   Note: Insoluble ionic compounds, gases, and liquids should not be split into ions.
4. Cancel out 'spectator' ions.
5. Write the net ionic equation.

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Example 1.3

Write a net ionic equation for the reaction between aqueous solutions of silver nitrate and calcium chloride.

Word equation:

Silver nitrate + Calcium chloride ⟶ Silver chloride + Calcium nitrate

Formula equation:

$$2AgNO_3(aq) + CaCl_2(aq) \longrightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$$

Total ionic equation:

$$2Ag^+(aq) + 2NO_3^-(aq) + Ca^{2+}(aq) + 2Cl^-(aq) \longrightarrow 2AgCl(s) + Ca^{2+}(aq) + 2NO_3^-(aq)$$

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Cancel out spectator ions:

$$2Ag^+(aq) + \cancel{2NO_3^-(aq)} + \cancel{Ca^{2+}(aq)} + 2Cl^-(aq) \longrightarrow 2AgCl(s) + \cancel{Ca^{2+}(aq)} + \cancel{2NO_3^-(aq)}$$

Net ionic equation:

$$2Ag^+(aq) + 2Cl^-(aq) \longrightarrow 2AgCl(s)$$

Example 1.4

Write a net ionic equation for the reaction of barium chloride with sodium sulphate aqueous solutions.

Word equation:

Barium chloride + Sodium sulphate ⟶ Barium sulphate + Sodium chloride

Formula equation:

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

Total ionic equation:

$$Ba^{2+}(aq) + 2Cl^-(aq) + 2Na^+(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s) + 2Na^+(aq) + 2Cl^-(aq$$

Cancel out spectator ions:

$$Ba^{2+}(aq) + \cancel{2Cl^-(aq)} + \cancel{2Na^+(aq)} + SO_4^{2-}(aq) \longrightarrow BaSO_4(s) + \cancel{2Na^+(aq)} + \cancel{2Cl^-(aq)}$$

Net ionic equation:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$$

Example 1.5

Write a net ionic equation for the reaction between dilute hydrochloric acid and an aqueous sodium hydroxide.

Word equation:

Dilute hydrochloric acid + Sodium hydroxide ⟶ Sodium chloride + Water

Formula equation:

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

Total ionic equation:

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

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Cancel out spectator ions:

$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \longrightarrow \cancel{Na^+(aq)} + \cancel{Cl^-(aq)} + H_2O(l)$

Net ionic equation:

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

Exercise 1.1

1. Balance the following chemical equations:
   (a) $Na(s) + Cl_2(g) \longrightarrow NaCl(s)$
   (b) $H_2(g) + I_2(g) \longrightarrow HI(g)$
   (c) $Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$
   (d) $P(s) + O_2(g) \longrightarrow P_2O_5(g)$
2. With reasons, state whether the following chemical equations are balanced:
   (a) $Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$
   (b) $CaCl_2(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + Ca(NO_3)_2(aq)$
3. Write the balanced chemical equation for the reaction involving magnesium burning in oxygen gas.
4. Balance the following chemical equations:
   (a) $CaCO_3(s) + HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
   (b) $2NH_3(g) \longrightarrow N_2(g) + H_2(g)$
   (c) $C(s) + CO_2(g) \longrightarrow CO(g)$
5. Balance each of the following chemical equations and write its ionic equation:
   (a) $Na_2CO_3(s) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$
   (b) $H_2SO_4(aq) + KOH(aq) \longrightarrow K_2SO_4(aq) + H_2O(l)$
   (c) $Zn(s) + HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$
   (d) $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow PbSO_4(s) + NaNO_3(aq)$

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Types of chemical reactions

There are many types of chemical reactions. The common ones include; combination or synthesis, decomposition, displacement, precipitation, and redox reactions.

Combination reactions

A combination reaction is a chemical reaction in which two or more chemical species combine to form a single product. Combination reactions are also referred to as synthesis reactions. This type of reaction is expressed in the form of:

$$A + B \longrightarrow AB$$

Generally, there are three types of combination reactions, namely; a reaction between two or more elements, reaction between elements and compounds, and reaction between two compounds.

An example of a combination reaction between two elements is when magnesium burns in air (oxygen) with a bright white flame to form white solids of magnesium oxide. Figure 1.1 shows the burning of magnesium in oxygen.

$$\text{Magnesium} + \text{Oxygen gas} \longrightarrow \text{Magnesium oxide}$$

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

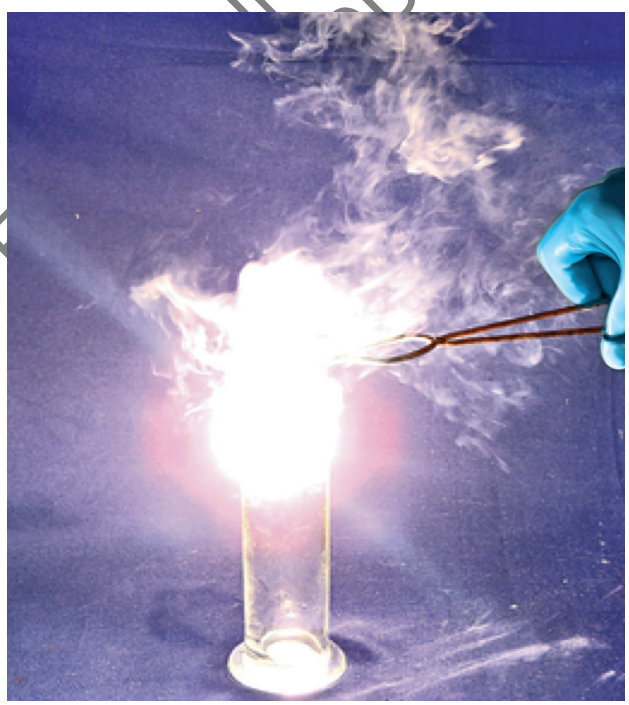


Figure. 1.1: *Magnesium burning in oxygen*

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Another example of a combination reaction between two elements is an industrial synthesis of ammonia gas from nitrogen and hydrogen gases.

Nitrogen gas + Hydrogen gas ⟶ Ammonia gas

$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

An element reacts with a compound to form another compound. For example, carbon monoxide gas reacts with oxygen gas to form carbon dioxide gas.

Carbon monoxide gas + Oxygen gas ⟶ Carbon dioxide gas

$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$

Two compounds may react with each other to form a new compound. For example, calcium oxide (quicklime) reacts with carbon dioxide gas to form calcium carbonate (limestone).

Calcium oxide + Carbon dioxide gas ⟶ Calcium carbonate

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

Another example of a combination reaction is the reaction between lead(IV) oxide and sulphur dioxide. If lead(IV) oxide is slightly heated and then lowered into a gas jar of sulphur dioxide, the two compounds combine to form one new compound, lead(II) sulphate.

Lead(IV) oxide + Sulphur dioxide gas ⟶ Lead(II) sulphate

$PbO_2(s) + SO_2(g) \longrightarrow PbSO_4(s)$

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Activity 1.1



Aim: To burn magnesium in oxygen

Requirements: A pair of tongs, Bunsen burner, gas jar, lighter, magnesium ribbon, oxygen source, and steel wool

Procedure

1. Prepare a sample of oxygen gas in a gas jar.
2. Clean about 0.1 g of a piece of magnesium ribbon by using a steel wool.
3. Hold the ribbon by using a pair of tongs and heat it over a Bunsen burner flame.
4. When it starts to burn, lower it into the gas jar of oxygen. Do not drop it into the jar.

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Questions

1. How does magnesium burn in oxygen?
2. Give the name(s) of the product(s) formed.
3. Write a balanced chemical equation for this reaction.

Decomposition reactions

A decomposition reaction is a chemical reaction in which a single compound breaks down (decomposes) into its component parts. This reaction is expressed in the form of:

$$AB \longrightarrow A + B$$

The decomposition reaction is the opposite of a combination reaction. Generally, decomposition reactions are classified into three main types, namely catalytic, electrolytic and thermal reactions. In a *catalytic decomposition reaction*, an agent called a *catalyst* is introduced that alters the rate of a chemical reaction but remains unchanged at the end of the reaction. For example, potassium chlorate readily decomposes when heated in presence of manganese(IV) oxide (catalyst) to produce oxygen gas and potassium chloride.

$$\text{Potassium chlorate} \xrightarrow[\text{heat}]{MnO_2} \text{Potassium chloride} + \text{Oxygen gas}$$

$$2KClO_3(s) \xrightarrow[\text{heat}]{MnO_2} 2KCl(s) + 3O_2(g)$$

An *electrolytic decomposition reaction* is achieved by exposing an aqueous solution or molten compound to an electric current. An example of an electrolytic decomposition reaction is the electrolysis of water, which is represented by the following chemical equation:

$$2H_2O(l) \longrightarrow 2H_2(g) + O_2(g)$$

Thermal decomposition occurs when a compound is exposed to direct heat or radiation. For example, when lead(II) nitrate crystals are heated, they decompose with a cracking sound to produce lead(II) oxide, a brown nitrogen dioxide gas, and oxygen gas.

$$\text{Lead(II) nitrate} \xrightarrow{\text{heat}} \text{Lead(II) oxide} + \text{Nitrogen dioxide} + \text{Oxygen gas}$$

$$2Pb(NO_3)_2(s) \xrightarrow{\text{heat}} 2PbO(s) + 4NO_2(g) + O_2(g)$$

Some decomposition reactions are caused by light. For example, white silver chloride breaks down when exposed to light to give tiny black crystals of silver and chlorine gas.

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$$\text{Silver chloride} \xrightarrow{\text{light}} \text{Silver} + \text{Chlorine gas}$$

$$2AgCl(s) \xrightarrow{\text{light}} 2Ag(s) + Cl_2(g)$$

Displacement reactions

A displacement reaction is a chemical reaction in which a more reactive element displaces a less reactive element from its compound. This reaction is expressed in the general form as:

$$A + BC \longrightarrow AC + B$$

Two reactants yield two different products. For example, when zinc reacts with hydrochloric acid, it displaces hydrogen to produce zinc chloride and hydrogen gas.

$$\text{Zinc} + \text{Hydrochloric acid} \longrightarrow \text{Zinc chloride} + \text{Hydrogen gas}$$

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

Another example of a displacement reaction is when solid iron reacts with copper(II) sulphate. In this reaction, iron displaces copper from copper(II) sulphate.

$$\text{Iron} + \text{Copper(II) sulphate} \longrightarrow \text{Iron(II) sulphate} + \text{Copper}$$

$$Fe(s) + CuSO_4(aq) \longrightarrow FeSO_4(aq) + Cu(s)$$

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Precipitation reactions

A precipitation reaction is a reaction in which two soluble compounds combine to give a soluble compound and an insoluble compound. This reaction is expressed in the general form of;

$$AB(aq) + CD(aq) \longrightarrow AD(s) + CB(aq)$$

The insoluble compound formed is called a *precipitate*. This type of reaction is also referred to as a *double displacement reaction*. An example of a precipitation reaction is the reaction between an aqueous solution of silver nitrate and an aqueous solution of sodium chloride to form white precipitates of silver chloride.

$$\text{Silver nitrate} + \text{Sodium chloride} \longrightarrow \text{Silver chloride} + \text{Sodium nitrate}$$

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$$

Another example is when an aqueous solution of sodium sulphate is mixed with an

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aqueous solution of barium chloride to form solid barium sulphate.

Sodium sulphate + Barium chloride ⟶ Barium sulphate + Sodium chloride

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

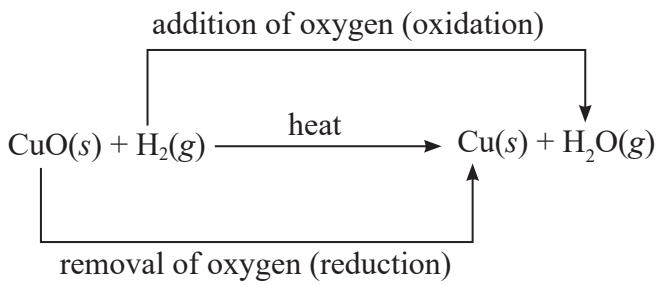
Redox reactions

A redox reaction is a reaction in which the oxidation number or oxidation state of the participating chemical species changes by losing or gaining one or more electron(s). Redox is a short form for reduction-oxidation reaction. Redox reactions are common and important to some of the basic functions of life including photosynthesis, respiration, combustion, and corrosion or rusting. In a redox reaction, oxidation and reduction reactions occur simultaneously.

A chemical reaction is said to be an oxidation reaction when its reacting substance combines with oxygen or hydrogen is removed from it, thus loss of electrons from that substance, or increase in oxidation state of that substance. On the other hand, a chemical reaction is said to be a reduction reaction when its reacting substance combines with hydrogen, or when oxygen is removed from that substance, thus a substance gains electron(s) leading to the decrease in oxidation state of that substance.

Oxidation and reduction are opposite types of reactions. During a redox reaction, if one substance is oxidised by either gaining oxygen, losing hydrogen or losing one or more electrons, another substance must at the same time be reduced by losing oxygen, gaining hydrogen or gaining one or more electrons. For example, when copper(II) oxide is heated in hydrogen gas, it is reduced to copper metal, while the hydrogen gas is oxidised to water. In this reaction, copper(II) oxide gets reduced because hydrogen takes away the oxygen. Hydrogen is thus a reducing agent. Hydrogen gas gets oxidised because copper(II) oxide gives out its oxygen to it, thus copper(II) oxide is an oxidising agent.

Copper(II) oxide + Hydrogen gas $\xrightarrow{\text{heat}}$ Copper + Steam



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Another example of a redox reaction is when iron metal is heated in chlorine gas to form iron(III) chloride.

$$2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$$

In this reaction, iron metal is oxidised, while chlorine gas is reduced.

$$Fe(s) \longrightarrow Fe^{3+}(aq) + 3e^-$$

$$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$$

Iron is the reducing agent since it loses electrons to chlorine atoms. Chlorine gas is the oxidising agent since it accepts or gains electrons from iron.

Activity 1.2



Aim: To verify the law of conservation of matter (mass) in a precipitation reaction

Requirements: Conical flask, stopper, analytical balance, 1 M barium chloride solution, and 1 M zinc sulphate solution

Procedure

1. Weigh the mass of two empty conical flasks with their stoppers on.
2. Put about 50 cm^3 of 1 M barium chloride solution in one of the flasks and another 50 cm^3 of 1 M zinc sulphate in the second flask.
3. Stopper the flasks and weigh them to get the mass of each solution.
4. Empty the weighed solution of barium chloride into a flask containing zinc sulphate. Stopper the flasks and observe the reaction.
5. Weigh the mixture after the reaction.

Questions

1. What is the total mass of the solutions before the reaction?
2. What is the mass of the mixture after the reaction?
3. What is the type of a chemical reaction that has taken place?
4. Write a balanced chemical equation for the reaction.

Task 1.2

Use any reliable sources of information to describe three common chemical reactions which occur in everyday life.

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Exercise 1.2

1. How does an oxidation reaction differ from a reduction reaction?
2. Write a balanced chemical equation for each of the following reactions:
   (a) Zinc reacts with silver nitrate solution to produce zinc nitrate and silver.
   (b) Aqueous potassium iodide reacts with aqueous solution of lead(II) nitrate to produce potassium nitrate and lead(II) iodide.
3. Identify the type of reaction in each of the following chemical equations. Explain your answer.

   (a) $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

   (b) $CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$

   (c) $Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$

   (d) $Zn(NO_3)_2(aq) + 2NaOH(aq) \longrightarrow Zn(OH)_2(s) + 2NaNO_3(aq)$

Chapter summary

1. Chemical equations are short forms of chemical reactions. The reactants are placed on the left hand side of the equation, while the products are placed on the right hand side of the equation.
2. Chemical equations are written by using the chemical symbols or formulas. The state of each of the reactants and products is indicated in parentheses and the equations are balanced.
3. An ionic equation is a chemical equation in which compounds in aqueous solutions or in molten states are written in terms of dissociated ions.
4. Spectator ions are ions that do not change their valence states in a reaction. The net ionic equation is the result of an ionic equation without the spectator ions.
5. There are different types of chemical reactions. These are combination, decomposition, displacement, precipitation, and redox reactions.

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Revision exercise 1

1. Explain the following terms:
   (a) Oxidation
   (b) Reduction
   (c) Ionic equation

2. Write the word equation for each of the following reactions:
   (a) Burning calcium in oxygen gas.
   (b) Dissolving zinc in dilute sulphuric acid.
   (c) Reacting sodium with water.

3. Balance each of the following chemical equations that involve acids and hypothetical compounds:
   (a) $MCO_3(s) + HCl(aq) \longrightarrow MCl_2(aq) + H_2O(l) + CO_2(g)$
   (b) $MOH(aq) + H_2SO_4(aq) \longrightarrow M_2SO_4(aq) + H_2O(l)$
   (c) $MO(s) + HCl(aq) \longrightarrow MCl_2(aq) + H_2O(l)$

4. Using balanced chemical equations, describe the following types of reactions:
   (a) Combination
   (b) Decomposition
   (c) Displacement
   (d) Precipitation
   (e) Redox

5. Complete and balance each of the following chemical equations:
   (a) $Zn(s) + Cl_2(g) \longrightarrow$
   (b) $CaO(s) + H_2O(l) \longrightarrow$
   (c) $MgCl_2(aq) + AgNO_3(aq) \longrightarrow$

6. Which of the following is the net ionic equation for a reaction between hydrochloric acid and potassium hydroxide?
   (a) $K^+(aq) + Cl^-(aq) \longrightarrow KCl(aq)$
   (b) $K^+(aq) + Cl^-(aq) \longrightarrow KCl(s)$
   (c) $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

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7. Differentiate between the following terms:
   (a) Ionic equation and molecular equation.
   (b) Total ionic equation and net ionic equation.

8. Study the following chemical reaction and answer the questions that follow:

$$3KOH(aq) + FeCl_3(aq) \longrightarrow 3KCl(aq) + Fe(OH)_3(s)$$

   (a) What type of a reaction is represented by the chemical equation?

   (b) Write a net ionic equation for the chemical reaction.

9. Why is the use of the balanced chemical equations so important in the chemical manufacturing industries?

10. Write the balanced chemical equations for the action of heat on:
   (a) Iron(II) sulphate
   (b) Calcium hydrogencarbonate
   (c) Ammonium nitrate

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Chapter Two

Hardness of water

Introduction

In Form Two, you learnt about water as an essential substance to all living things and for different activities. Over 70% of the Earth's surface is covered by water. When water percolates through the rock containing limestone, chalk or gypsum, the bicarbonates, sulphates or chlorides of calcium and magnesium metals dissolve in it to form hard water. In this chapter, you will learn the concept of hardness of water, types of hardness of water, treatment and purification of hard water, and the advantages and disadvantages of hard water. The competencies developed will enable you to apply various methods to treat and purify hard water for various purposes.

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Concept of hardness of water

Hardness of water is caused by dissolved salts including sulphates, chlorides, and hydrogencarbonates of calcium and magnesium. Minerals such as iron, zinc, and manganese may also contribute to the hardness of water. These salts are dissolved when rainwater falls on rocks of limestone ($CaCO_3$), gypsum ($CaSO_4$), and dolomite ($CaCO_3 \cdot MgCO_3$). In addition, as the rain falls through the air, it dissolves carbon dioxide to form a weak carbonic acidic solution. The acid solution reacts with calcium carbonate to form soluble calcium hydrogencarbonate which causes temporary hardness of water.

$$\underset{\text{Insoluble}}{CaCO_3(s)} + H_2O(l) + CO_2(g) \longrightarrow \underset{\text{Soluble}}{Ca(HCO_3)_2(aq)}$$

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The process of hard water formation in limestone rocks is summarised in Figure 2.1.

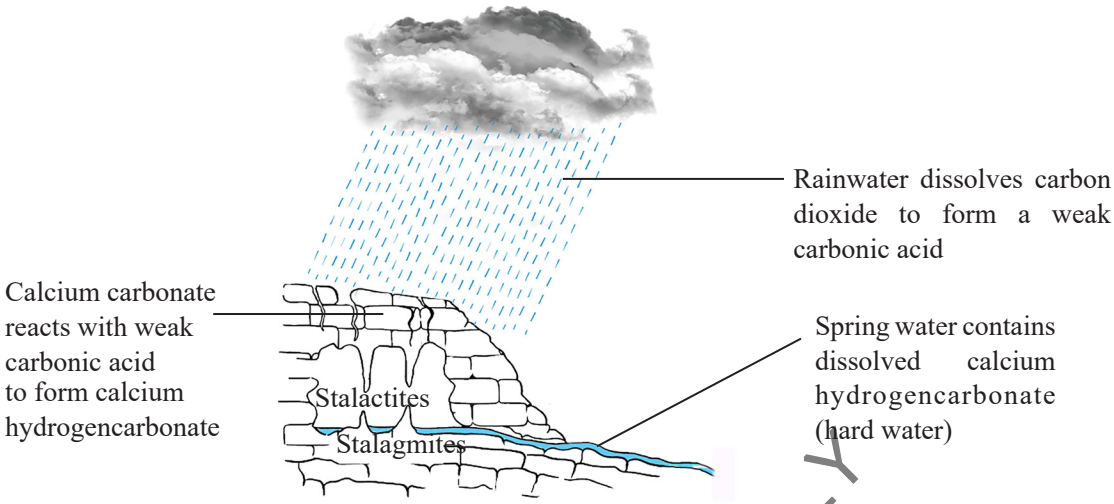


Figure 2.1: *Process of hard water formation in limestone rocks*

When natural water flows over the rocks containing chlorides and sulphates of calcium and magnesium, these salts dissolve and cause permanent hardness of water.

Action of soap on hard water

The response of a given sample of water to soap determines the hardness of water. Hard water does not readily form lather with soap. Soft water readily forms lather with soap because it has a low concentration of calcium and magnesium ions which cause hardness of water. Figure 2.2 shows lather formation in soft and hard water.



Figure 2.2: *Lather formed in soft and hard water*

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In hard water, sodium stearate (soap) is less effective because it combines with calcium or magnesium salts (hardness) to generate a white precipitate of calcium or magnesium salts of the soap. These salts are insoluble and result in soap scum. The sodium sulphate which is formed is soluble in water. This reaction is illustrated in the following word equation:

Calcium sulphate	+	Sodium stearate	⟶	Calcium stearate	+	Sodium sulphate
Hardness		Soap		Insoluble salt (scum)		Soluble salt

Activity 2.1



Aim: To distinguish between hard water and soft water

Requirements: Two 100-mL conical flasks, rainwater, salty water, distilled water, and bar soap

Procedure

1. Place the rainwater or distilled water and salty water into two different conical flasks.
2. Add an equal amount of soap in each conical flask then shake.
3. Record the results.

Questions

1. Based on the observations, which water forms more lather in the two conical flasks? Explain.
2. Based on the observations, classify the rainwater distilled water and salty water as either soft or hard water. Explain your answer.

Types of hardness of water

Hard water is categorised into temporary or permanent hard water, based on the presence of hydrogencarbonates or sulphates and chlorides of calcium and magnesium. Temporary hard water contains hydrogencarbonates of calcium and magnesium. Permanent hard water is mainly caused by the sulphates or chlorides of calcium or magnesium. Sulphates and chlorides of iron, copper, and manganese also cause permanent hardness of water. However, such salts are found in the environment in very small amounts, thus their effects on water hardness are minimum.

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Activity 2.2



Aim: To examine the effects of a soap solution on different water samples

Requirement: Filter funnel, 25-mL measuring cylinder, 100-mL measuring cylinder, conical flasks, wire gauze, spatula, Bunsen burner, soap solution, distilled water, sodium carbonate crystals, and samples of temporary and permanent hard water

Note: Prepare a soap solution by grinding a bar soap and mix with 150 cm^3 distilled water

Procedure

1. Using a 100-mL measuring cylinder, measure 50 cm^3 of distilled water into a conical flask.
2. Using a 25-mL measuring cylinder, measure 20 cm^3 of soap solution into a conical flask.
3. Mix distilled water with soap solution and shake the mixture thoroughly.
4. Leave the mixture for 20 seconds then record the observation.
5. Repeat Steps 1–4 using other water samples indicated in Table 2.1 in the place of distilled water.
6. Record the results in Table 2.1 by putting a tick (√) if the mixture forms lather or a cross (x) if the mixture does not form lather.

Table 2.1: *Effects of soap on water samples*

Water sample	Formation of lather
(a) Distilled water	
(b) Temporary hard water	
(c) Permanent hard water	
(d) Temporary hard water after boiling	
(e) Permanent hard water after boiling	
(f) Temporary hard water after adding sodium carbonate crystals	
(g) Permanent hard water after adding sodium carbonate crystals	

Questions

1. Explain the observations recorded in Table 2.1.
2. What was the purpose of adding sodium carbonate crystals into the water samples?

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Treatments and purification of hard water

Hard water is treated in order to soften so that it can be used for domestic, industrial, and other purposes. The softening is achieved either by chemicals that form precipitates or by ion exchange. Water softening is a process in which the ions of calcium, magnesium, and sometimes iron, copper, and manganese are removed from hard water. Hard water purification is a process by which undesirable chemical compounds, organic, and inorganic materials are removed from water.

Methods used for softening hard water

The methods that are used in the treatment and purification of hard water include boiling, addition of limewater, addition of aqueous ammonia, distillation, addition of washing soda, and the use of ion exchangers.

Boiling: Boiling decomposes calcium and magnesium hydrogencarbonates to form carbonates, water and carbon dioxide gas. The calcium and magnesium ions are removed from water as insoluble carbonates and then deposited as solids at the bottom of boilers. For example,

$$Ca(HCO_3)_2(aq) \xrightarrow{\text{heat}} CaCO_3(s) + H_2O(l) + CO_2(g)$$

$$Mg(HCO_3)_2(aq) \xrightarrow{\text{heat}} MgCO_3(s) + H_2O(l) + CO_2(g)$$

Boiling method is used to remove only temporary hardness of water because the sulphates or chlorides of calcium or magnesium in hard water become more soluble as the temperature increases.

Addition of aqueous ammonia: When aqueous ammonia is added to temporary hard water, the soluble hydrogencarbonates of magnesium and calcium are converted into insoluble carbonates. For example,

$$2NH_4OH(aq) + Mg(HCO_3)_2(aq) \longrightarrow MgCO_3(s) + (NH_4)_2CO_3(aq) + 2H_2O(l)$$

When aqueous ammonia is added to permanent hard water the soluble sulphates or chlorides of magnesium and calcium are converted into insoluble hydroxides. For example,

$$2NH_4OH(aq) + MgSO_4(aq) \longrightarrow Mg(OH)_2(s) + (NH_4)_2SO_4(aq)$$

The magnesium and calcium ions are removed from both temporary and permanent hard water as insoluble carbonates and hydroxides.

Distillation: This process relies on evaporation followed by condensation to purify hard water. During distillation, the hard water is heated to form steam which then cools and condenses to form soft water. The salts are left behind when water evaporates. This method is used to get rid of both temporary and permanent water hardness. It also gives the softest water that is used for various purposes such as in chemical industries, medical laboratories, and research laboratories. However, distillation process is the most expensive way to soften water.

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Treatment with calcium hydroxide (*limewater*): The soluble magnesium or calcium hydrogencarbonates react with calcium hydroxide to form insoluble magnesium carbonate or calcium carbonate and water. For example,

$$Ca(HCO_3)_2(aq) + Ca(OH)_2(aq) \longrightarrow 2CaCO_3(s) + 2H_2O(l)$$

The insoluble calcium carbonates are deposited and soft water is obtained. This method is used to remove both temporary and permanent hardness of water.

Use of washing soda (Na_2CO_3): Sodium carbonate is added into water to form calcium carbonate or magnesium carbonate precipitates. This removes both temporary and permanent hardness of water. The reactions are represented by the following equations:

$$Ca(HCO_3)_2(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + 2NaHCO_3(aq)$$

$$CaSO_4(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + Na_2SO_4(aq)$$

$$MgSO_4(aq) + Na_2CO_3(aq) \longrightarrow MgCO_3(s) + Na_2SO_4(aq)$$

Use of ion exchangers: These treat both temporary and permanent hardness by removing all the calcium and magnesium ions from water. An ion exchanger is a container full of small beads which are made of a special polymer matrix called ion exchange resin. This contains sodium ions that are weakly attached to it. When hard water is passed through an ion exchanger, calcium or magnesium ions replace sodium ions and attach themselves on the resins. Calcium or magnesium ions are therefore left behind in the resin as the soft water flows out with the sodium ions as illustrated in the following chemical equation (Y is the negatively charged ion exchanger) and Figure 2.3.

$$CaSO_4(aq) + Na_2Y(aq) \longrightarrow CaY(s) + Na_2SO_4(aq)$$

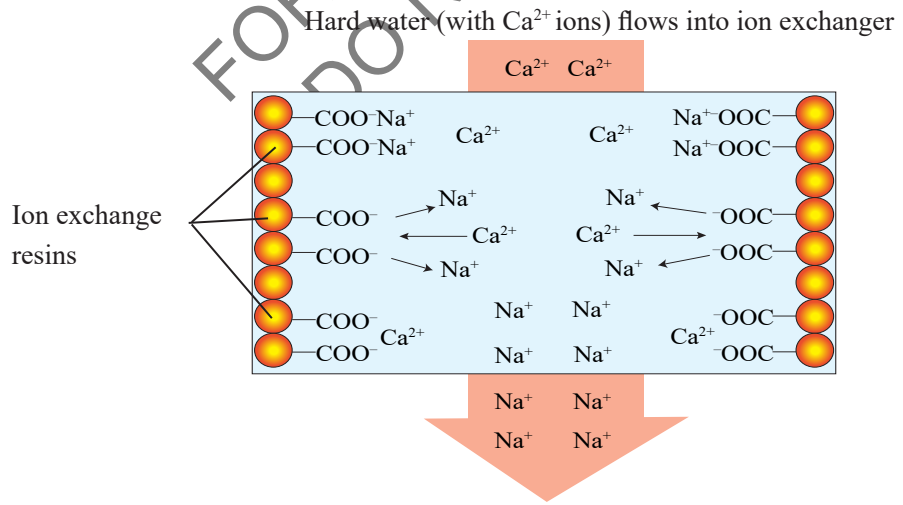


Figure 2.3: *Process of water softening by an ion exchanger*

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After some times, all the sodium ions on the resins are replaced with calcium or magnesium ions. At this point, the exchanger cannot remove hardness any more. The regeneration of the resins is accomplished by adding a concentrated solution of sodium chloride into the exchanger. Sodium ions push the calcium or magnesium ions off the resins, thus recharging the ion exchanger with sodium ions making it ready for re-use.

Importance of treatment and purification of hard water

Treatment and purification of hard water is very important due to the following reasons:

1. Hard water that has been treated is suitable for various uses in both domestic and industrial settings. Treated water minimises the quantity of soap that must be used for washing clothes, bathing, and cleaning floor or tile surfaces.
2. Hard water that has been treated is suitable for uses in laboratories and medical facilities to ensure accurate experimental findings and effective medical treatments.
3. Treated water reduces the scale build up on different containers and equipment, and therefore prolonging their lifespan.
4. Treated water saves electrical energy because there are no formations of scums or scales on appliances and utensils, making them more efficient.

Advantages of hard water

Hard water has the following advantages:

1. The presence of dissolved salts and ions such as calcium and magnesium carbonates, chlorides, and sulphates in hard water makes it taste good.
2. It provides calcium and magnesium which aid the development and strength of bones and teeth.
3. The formation of limescale inside the metal pipes due to hard water causes the inside of the pipes to be covered with insoluble carbonates. This layer of carbonates prevents water in the pipes from getting into contact with the metal surface of the pipes. This prevents the pipes from corrosion and poisonous metal salts from dissolving in water.
4. It helps in the formation of shells in some organisms.

Disadvantages of hard water

The following are the disadvantages of hard water:

1. Limescale formation in water boilers, hot water pipes, kettles, and other appliances makes them less efficient. It also causes blockage in the appliances and have to be removed from time to time, which is costly. Figure 2.4 shows limescale build up in an electric kettle.

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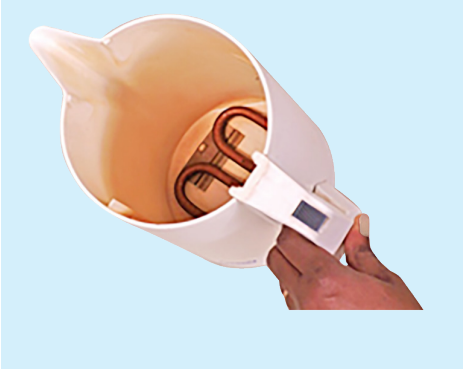


Figure 2.4: *Limescale in electric kettle*

2. Wastage of more soap: For example, when you wash clothes with hard water, instead of producing lather, the soap forms a white precipitate (scum). If there is no formation of lather, the clothes do not get cleaned, and therefore, a lot of soap should be used.
3. Stains in clothes and utensils: Soap and hard water form scum which stains clothes and obscure their colours. Hard water damages silk and nylon, stains glass, bath tiles, and bath fittings.
4. Impacts on hair and skin: When hair is cleaned with hard water, it becomes dry and rough. This is because excess minerals present in hard water form a curd-like substance which sticks on the hair. Taking a bath with hard water makes the skin dry and itchy. This is because the leftovers of soap residues stick on the body.
5. Reduction of the lifespans of appliances: The lifespans of household appliances are shortened if they are operated by using hard water because hard water wears out appliances. In addition, scales produced as a result of mineral deposits diminish appliances efficiency.

Exercise 2.1

1. Explain the meaning of the following terms:
   (a) Soft water
   (b) Hard water
   (c) Permanent hardness of water
   (d) Temporary hardness of water
2. Describe the causes of;
   (a) permanent hardness of water.
   (b) temporary hardness of water.
3. Outline any four ways used to soften hard water.
4. Explain two different chemical methods used for removal of both temporary and permanent hardness of water.
5. It is important to treat hard water. Explain.
6. Name the most economical method of treating hard water.
7. Permanent hardness of water cannot be removed by boiling. Explain.
8. Explain why soap does not form lather easily with hard water at first, but eventually forms the lather.

Chapter summary

1. Soft water easily forms lather with soap, while hard water does not easily form lather with soap.

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2. Hardness in water is mainly caused by dissolved sulphates, chlorides, and hydrogencarbonates of magnesium and calcium.
3. Temporary hardness of water is caused by dissolved calcium or magnesium hydrogencarbonates, while permanent hardness of water is caused by dissolved sulphates or chlorides of calcium or magnesium.
4. Boiling removes temporary hardness of water, but does not remove permanent hardness.
5. The temporary and permanent hardness of water can be removed by distillation, addition of sodium carbonate (washing soda), addition of limewater, and by the use of ion exchangers.

Revision exercise 2

Choose the correct answer in items 1–4.

1. Which compound among the following causes temporary hardness of water when dissolved in water?
   (a) NaCl (b) $CaCl_2$
   (c) $Mg(HCO_3)_2$ (d) $Pb(NO_3)_2$
2. One of the following substances causes permanent hardness of water when dissolved in water:
   (a) Calcium carbonate
   (b) Calcium sulphate
   (c) Magnesium carbonate
   (d) Sodium carbonate
3. Which of the following conditions of water is easily removed by boiling?
   (a) Temporary hardness
   (b) Scum
   (c) Limescale
   (d) Permanent hardness
4. Identify a rock that does not cause hardness of water among the following:
   (a) Dolomite
   (b) Gypsum
   (c) Limestone
   (d) Bauxite
5. The formation of limescale in water pipes is both advantageous and disadvantageous. Explain.
6. Identify the challenges associated with the ion exchanger method of hard water purification.
7. Samples of water (A-D) from four different areas were tested with a soap solution to establish how much soap solution was needed to form lather that lasted for 30 seconds. The experiment was repeated for the second time by using the samples that had been boiled, and then for the third time by using the samples that had been passed through an ion exchanger.

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Experimental results are presented in the following table.

Water sample	Volume of soap solution used (cm^3)		
	Untreated water	*Boiled water*	*Water passed through an ion exchanger*
A	12	1.8	1.8
B	17	17	1.7
C	26	20	1.8
D	1.6	1.6	1.6

(a) Which of the samples was the hardest water? Give reasons.

(b) Which sample behaved like distilled water? Explain.

(c) Which water samples had the permanent or temporary water hardness? Give reasons.

(d) Write an equation for the reaction of removing hardness in sample C.

8. Briefly explain the following observations about a sample of hard water:
   (a) When boils, it forms white precipitate.
   (b) After boiling, the water forms a scum with soap.
   (c) Sodium carbonate makes the water completely soft.
9. What is the economic importance of hard water in daily life?
10. Explain how an ionic exchanger works in removing hardness of water.
11. You are given a water sample which has both temporary and permanent hardness. Explain whether boiling will increase or decrease its hardness.
12. With the aid of chemical equations, briefly describe the following processes:
   (a) The removal of temporary hardness of water by boiling.
   (b) The removal of permanent hardness of water by chemical means.
13. Lack of safe and clean water for industrial and domestic uses is a serious problem in most urban areas in Tanzania. The major cause of this problem is pollution in the water sources which causes the hardness of water. State three methods that could make such water clean and safe.
14. Explain any two methods that are commonly used for treatment of hard water at home.

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Chapter Three

Acids, bases, and salts

Introduction

In Form Two, you learnt about chemical formulas of different substances including those of acids, bases, and salts. Acids, bases, and salts are important for proper functioning of our bodies and for the health of the environment. Medicines, food stuffs, and fertilisers are among the substances which contain acids, bases or salts. For example, some of the edible fruits are acidic in nature and have health importance in the body. The soaps used in daily life contain bases. In this chapter, you will learn about acids, bases, acid-base indicators, salts, their properties, and uses. The competencies developed will enable you to prepare and use acids, bases, and salts, as well as apply the concept of acid-base neutralisation in daily life.

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Acids

An *acid* is any substance which produces hydrogen ions (H^+) in water as the positively charged ions. In this case, acids are substances made of hydrogen ions, and other ions.

Some acids occur naturally, and some can be synthesised. The substances which contain natural acids include sour milk and citrus fruits like oranges, lemons, and limes. Figure 3.1 and Table 3.1 show some natural sources of acids and some natural acids with their corresponding sources, respectively.



Citrus fruits | Pineapples | Plums | Passion fruits

Figure 3.1: *Different sources of natural acids*

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Table 3.1: *Examples of natural acids and their sources*

Name of natural acid	Natural sources
Citric acid	Citrus fruits such as lemons and oranges
Lactic acid	Sour milk, cheese, and muscles
Acetic acid	Vinegar and rotten fruits such as grapes, pineapple and orange
Tartaric acid	Grapes, bananas, and tamarind
Oxalic acid	Spinach and tomatoes
Formic acid	Bee or ant stings
Ascorbic acid	Fruits and vegetables, for example, tomatoes
Folic acid	Fruits, vegetables, nuts, and grains
Malic acid	Apples, strawberries, and plums

Acids can be classified as *mineral* (*inorganic*) *acids* and *organic acids*. Mineral acids are synthesised in industries from mineral resources, while organic acids are synthesised in industries or obtained directly from natural organic materials. Examples of mineral acids include nitric acid (HNO_3), sulphuric acid (H_2SO_4), hydrochloric acid (HCl), carbonic acid (H_2CO_3), phosphoric acid (H_3PO_4), nitrous acid (HNO_2), and sulphurous acid (H_2SO_3). Examples of organic acids which are synthesised in industries include ethanoic acid/acetic acid (CH_3COOH) and methanoic acid (HCOOH). The most common acids that are found in a chemistry laboratory include hydrochloric acid, sulphuric acid, nitric acid, and acetic acid. Common acids in the laboratory are stored in various containers such as plastic or glass bottles as shown in Figure 3.2.



Glass bottle Plastic bottle

Figure 3.2: *Glass and plastic bottles for storing acids in the laboratory*

Properties of acids

Acids have certain physical and chemical properties.

Physical properties of acids

The following are the physical properties of acids:

1. Acids have a sour taste: A sour taste of vinegar, sour milk, and many unripe fruits such as lemons, is caused by the acids present in them.
2. Most acids are water soluble: They can be diluted with water to reduce the intensity of their acidity.
3. Solutions of acids turn blue litmus paper red (Figure 3.3).

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Figure 3.3: *Blue litmus paper immersed in an acid solution*

4. Acids have the ability to conduct electricity: When acids are dissolved in water, they release ions which conduct electricity. This is why acid solutions are commonly used in car and solar batteries.
5. Acids are corrosive in nature: Acids can have a corrosive action on various substances such as papers as well as clothes. They can also damage or even destroy metals. Therefore, they should be handled with extreme care, otherwise they can corrode your skin.
6. Acids are mostly present in liquid or gaseous forms, but there are some which exist in solid form such as oxalic acid.
7. Acids change the colour of phenolphthalein indicator to colourless.
8. Acids change the colour of methyl orange indicator to red.

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Chemical properties of acids

Acids have various chemical properties, including the following:

1. Acids have a pH less than 7 on a pH scale (Figure 3.4).

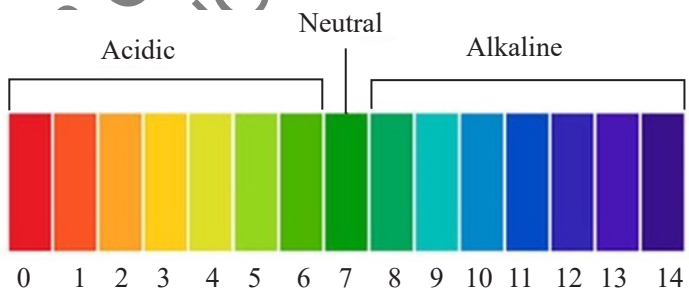


Figure 3.4: *pH scale*

2. Acids react with carbonates and hydrogencarbonates to produce salt, water, and carbon dioxide gas as shown in the following chemical equations:

$$2HCl(aq) + CaCO_3(s) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

$$H_2SO_4(aq) + Ca(HCO_3)_2(aq) \longrightarrow CaSO_4(s) + 2H_2O(l) + 2CO_2(g)$$

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One test of the produced gas (carbon dioxide) is to bubble it through limewater. The limewater will turn milky or cloudy white. The laboratory experimental set-up for the reactions of acids and carbonates is shown in Figure 3.5.

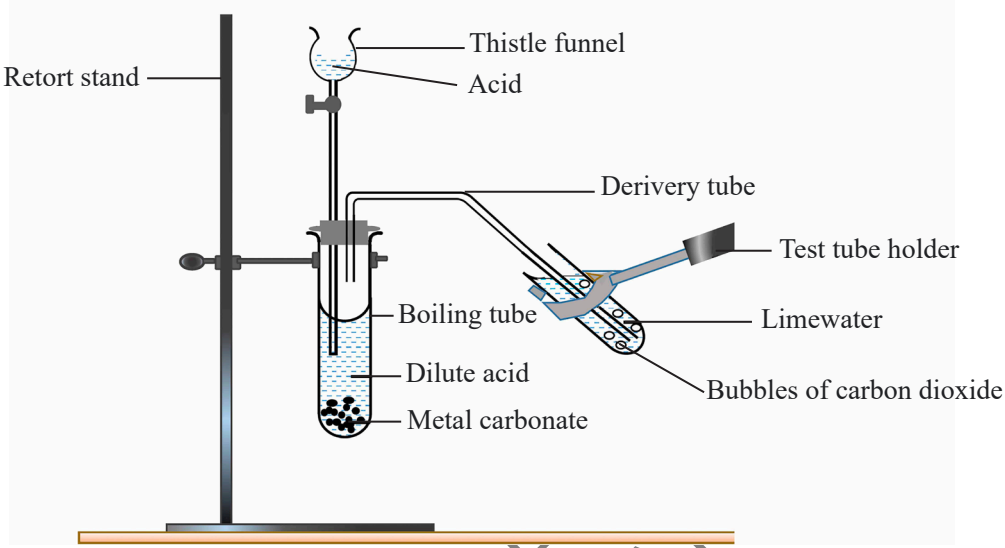


Figure 3.5: *Set-up for the reaction of acids with carbonates*

3. Acids react with metals above hydrogen in the reactivity series to form salt and hydrogen gas as shown in the following equations:

$$Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$$

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

$$Mg(s) + 2HNO_3(aq) \longrightarrow Mg(NO_3)_2(aq) + H_2(g)$$

Note: (a) The reaction between lead and dilute hydrochloric acid or sulphuric acid forms an insoluble coating of lead chloride and lead sulphate, respectively which prevents further reaction. The two reactions are very slow and it may take hours before a few small bubbles of hydrogen gas are observed. Copper does not react with dilute hydrochloric acid and sulphuric acid because it is below hydrogen in the reactivity series of metals.

(b) The reactions of dilute nitric acid and most metals produce water instead of hydrogen gas. This is because nitric acid is a strong oxidising agent. Its oxidising power depends on the concentration of the acid; the higher the concentration the higher the oxidising power of the acid. The following equations present the reaction between metals and concentrated nitric acid:

$$Pb(s) + 4HNO_3(l) \longrightarrow Pb(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$$

$$Ca(s) + 4HNO_3(l) \longrightarrow Ca(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$$

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4. Acids react with metal oxides or hydroxides to form salts and water only:

$$CuO(s) + 2HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + H_2O(l)$$

$$Ca(OH)_2(aq) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + 2H_2O(aq)$$

Activity 3.1



Aim: To investigate the reactions of dilute acids with metals

Requirements: Test tubes, test tube rack, retort stand and clamp, wooden splints, lighter, 2 M sulphuric acid, 2 M hydrochloric acid, 2 M nitric acid, zinc, magnesium ribbon, lead, and copper

Procedure

1. Transfer about 5 cm^3 of 2 M sulphuric acid, 2 M nitric acid, and 2 M hydrochloric acid in three separate test tubes.
2. Add a small piece of zinc metal to each test tube and record your observations.
3. Put a lighted splint at the mouth of each test tube one at a time as shown in Figure 3.6, and record the observations.
4. Repeat Steps 2–3 with magnesium, lead, and copper, each at a time instead of zinc metal.

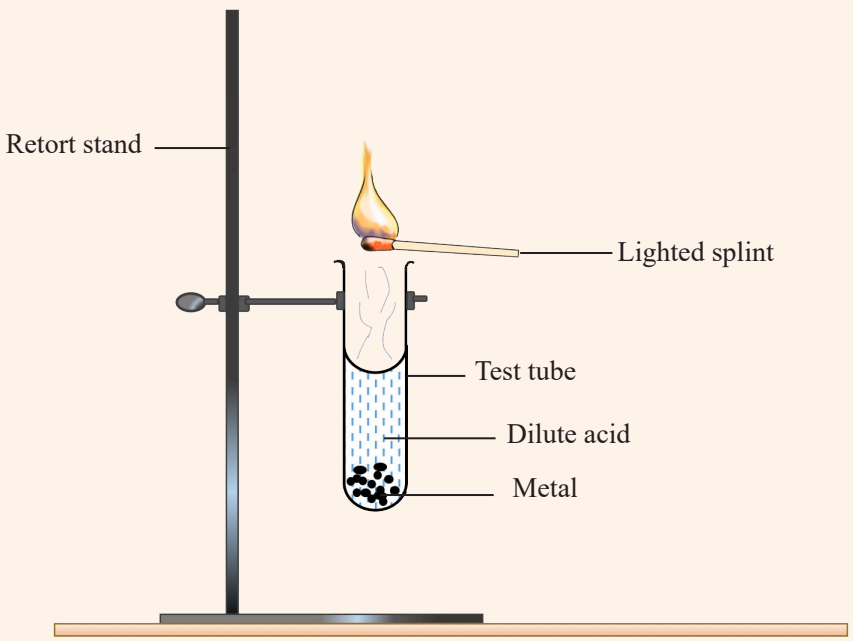


Figure 3.6: *Set-up for the reaction of dilute acids with metals*

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Questions

1. With the aid of balanced chemical equations, write the products of the reaction between:
   (a) zinc and dilute hydrochloric acid.
   (b) zinc and dilute sulphuric acid.
   (c) magnesium and dilute hydrochloric acid.
   (d) magnesium and dilute sulphuric acid.
2. Explain the cause of the observations recorded in Step 3 of the experiments.
3. (a) What happened in Steps 2 and 3 when lead and copper were used in the experiment?
   (b) Write the balanced chemical equations for the reactions in Step 2 of the experiment.

Strengths of acids

The strength of an acid is a measure of its ability to ionise (dissociate) in water to produce hydrogen ions (H^+). The more the acid ionises in water, the stronger it is. In addition, the more the H^+ ions are given out when the acid is dissolved in water, the more ionised the acid becomes. The strength of the acid solution is determined by the pH value of that particular solution. The pH is the measure of the degree of acidity or alkalinity of the solution. Based on their strengths, acids are grouped into strong acids and weak acids.

Strong acids ionise completely in water to give large amounts of H^+. Examples of strong acids are hydrochloric acid, nitric acid, and sulphuric acid. For example, hydrochloric acid ionises completely to give H^+ and Cl^-.

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

Weak acids ionise partially in water to produce small amounts of H^+. Examples of weak acids are ethanoic acid, carbonic acid, and citric acid. For example, ethanoic acid (CH_3COOH) ionises partially in water to produce few CH_3COO^- and H^+.

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

Note:

(a) The pH values of solutions increase with decrease in acidity; therefore, weak acids have higher pH values than strong acids.

(b) Solutions of strong acids are good conductors of electricity because they contain more free mobile ions to carry the charges than solutions of weak acids.

Acid concentration

The concentration of an acid is a measure of the amount of available acidic ions dissolved in a solvent. Dissolving more of the substance into a given volume of a solvent increases its concentration (Figure 3.7). On the other hand, adding more of the solvent reduces the concentration, making the solution more dilute.

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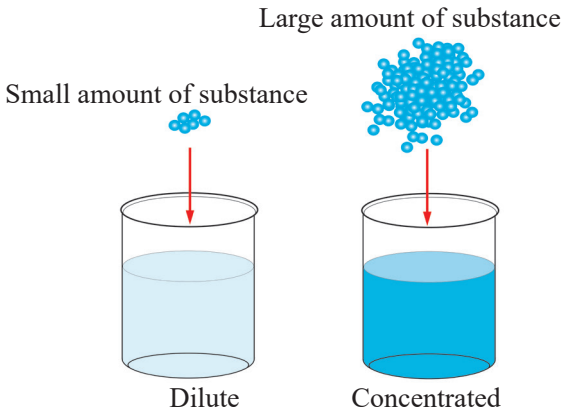


Figure 3.7: *Dilute and concentrated solutions*

Concentrated acids have large amounts of acid molecules dissolved in water, while dilute acids have small amounts of acid molecules dissolved in water. A concentrated acid is not necessarily a strong acid. Strong acids are still strong even if they are dilute, and weak acids are still weak even if they are concentrated. Concentration of an acid can be expressed in moles per litre, grams per litre, parts per million or percentage.

Basicity of an acid

Basicity of an acid is the number of ionisable hydrogen atoms per molecule of the acid that can be displaced by a metal in solution. For example, hydrochloric acid is *monobasic* as it has one ionisable hydrogen atom that can be displaced. Sulphuric acid and phosphoric acid are *dibasic* and *tribasic*, respectively. Sulphuric acid has two ionisable hydrogen atoms, while phosphoric acid has three ionisable hydrogen atoms that can be displaced (Table 3.2).

Table 3.2. *Basicity of some acids*

Acid	Basicity
$HCl \longrightarrow H^+ + Cl^-$	Monobasic
$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$	Dibasic
$H_3PO_4 \longrightarrow 3H^+ + PO_4^{3-}$	Tribasic

Bases

A *base* is a substance that neutralises an acid by reacting with hydrogen ions. Bases include the oxides, hydroxides, and carbonates of metals.

Oxides are those compounds which contain a metal and an oxygen atom. Examples of oxides include sodium oxide, Na_2O; potassium oxide, K_2O; calcium oxide, CaO; copper(II) oxide, CuO; and magnesium oxide, MgO.

Hydroxides are those compounds which contain a hydroxy group ($-OH$). Examples of hydroxides include calcium hydroxide, $Ca(OH)_2$; magnesium hydroxide, $Mg(OH)_2$; sodium hydroxide, $NaOH$; potassium hydroxide, KOH; and ammonium hydroxide, NH_4OH.

Metal carbonates are made up of a metal ion and a carbonate ion. Examples of metal carbonates include sodium carbonate, Na_2CO_3; potassium carbonate, K_2CO_3; magnesium carbonate, $MgCO_3$; and calcium carbonate, $CaCO_3$.

Bases are abundant in nature. Examples of natural substances that contain bases are ashes, banana peels, and avocado. There are also bases such as soda ash (sodium carbonate), baking soda (sodium bicarbonate) and limestone (calcium carbonate) which occur naturally on land

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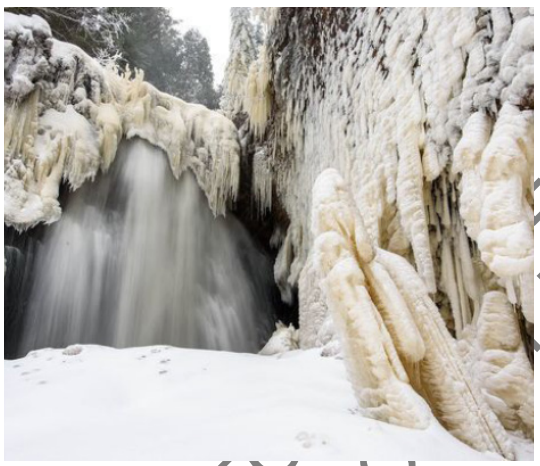
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and in water bodies. Figure 3.8 shows some sources of natural bases.



(a) Soda ash deposit



(b) Limestone deposit

Figure 3.8: *Some sources of natural bases*

Many of the common household products such as detergents, deodorants, toothpaste, and baking powder contain bases. These products contain synthesised bases.

Some bases are readily soluble in water, while others are not. Soluble bases are called alkalis. An *alkali* is a soluble base which when dissolved in water forms hydroxide ions as the only negatively charged ions. Therefore, all alkalis are bases, but not all bases are alkalis. Examples of soluble and insoluble bases are shown in Table 3.3.

Table 3.3: *Examples of soluble and insoluble bases*

Insoluble bases	Soluble bases
Copper(II) hydroxide ($Cu(OH)_2$)	Sodium hydroxide (NaOH)
Lead(II) hydroxide ($Pb(OH)_2$)	Potassium hydroxide (KOH)
Calcium hydroxide ($Ca(OH)_2$)	Sodium oxide (Na_2O)
Copper(II) oxide (CuO)	Potassium oxide (K_2O)
Iron(III) oxide (Fe_2O_3)	Sodium carbonate (Na_2CO_3)
Lead(II) oxide (PbO)	Potassium carbonate (K_2CO_3)
Lead(II) carbonate ($PbCO_3$)	
Magnesium carbonate ($MgCO_3$)	

Note: Calcium hydroxide and lead hydroxide are moderately soluble in water. Oxides and hydroxides of other metals are insoluble except those of sodium and potassium.

Properties of bases

Bases have physical and chemical properties that differ from those of acids.

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Physical properties of bases

Bases have the following physical properties:

1. Most bases have a bitter taste. For example, milk of magnesia, a common antacid.
2. Bases have a 'soapy' or slippery feel. For example, a rub of a drop of baking powder solution between fingers gives a slippery feel.
3. Most bases are insoluble in water.
4. Bases turn red litmus paper blue as shown in Figure 3.9.

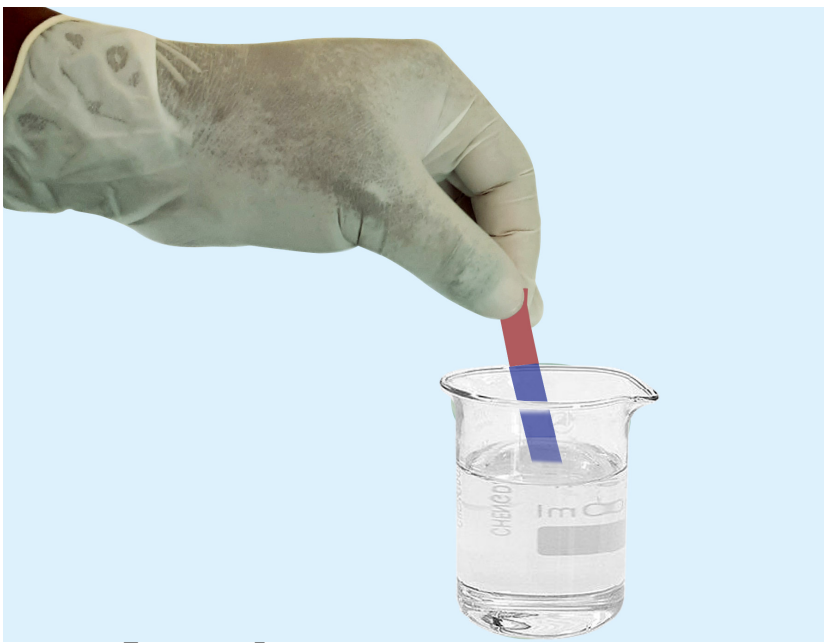


Figure 3.9: *Red litmus paper immersed in an alkali solution*

5. Bases change the colour of *phenolphthalein* (POP) indicator pink and that of *methyl orange* (MO) indicator yellow.
6. Bases generally do not have odour except for ammonia which has a pungent smell.
7. Bases are corrosive depending on their pH and concentrations.
8. Soluble bases (alkalis) conduct electricity when dissociated into ions.

Chemical properties of bases

Bases have the following chemical properties:

1. The pH values of bases are greater than 7.
2. Bases react with most cations to precipitate hydroxides. In this case, soluble bases precipitate insoluble metal hydroxides from their salt solutions. For example, potassium hydroxide reacts with iron(III) chloride to yield iron(III) hydroxide and

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potassium chloride.

$$3KOH(aq) + FeCl_3(aq) \longrightarrow Fe(OH)_3(s) + 3KCl(aq)$$

Ionically, the reaction is represented as:

$$3OH^-(aq) + Fe^{3+}(aq) \longrightarrow Fe(OH)_3(s)$$

3. Bases react with acids to form salt and water. For example, sodium hydroxide reacts with dilute hydrochloric acid to produce sodium chloride and water.

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

The ionic equation is

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

4. Alkalis react with ammonium salts to produce ammonia gas.

$$KOH(aq) + NH_4Cl(s) \longrightarrow KCl(aq) + H_2O(l) + NH_3(g)$$

Strengths of bases

The strength of a base or alkali is its ability to ionise in aqueous solution to produce hydroxide ions (OH^-). Like acids, the strength of a solution of a base is determined by the pH value of that particular solution. The more the base ionises in aqueous solution, the stronger it is. Alkalis are strong bases because they ionise completely in aqueous solution to produce free ions. Examples of strong bases are potassium hydroxide (KOH) and sodium hydroxide (NaOH).

$$KOH(aq) \longrightarrow K^+(aq) + OH^-(aq)$$

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

Weak bases are those bases which do not completely dissociate into their ions in aqueous solutions. The most common example of a weak base is ammonium hydroxide.

$$NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

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Activity 3.2



Aim: To identify acidic and basic substances

Requirements: Coconut, citrus fruits, ripe tomatoes, vinegar, sour milk, wood ash, pieces of chalk, yoghurt, blender, knife, water, red litmus papers, blue litmus papers, and beakers

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Procedure

1. Prepare juices of coconut, citrus fruits, and ripe tomatoes.
2. Use litmus papers to test the acidity or alkalinity of the juices obtained in Step 1.
3. Mix some wood ash and chalk powder, each with little water, decant to obtain a *supernatant* and use litmus papers to test the acidity or alkalinity of the supernatant.
4. Take small amounts of vinegar, sour milk, and yoghurt into separate beakers and test for acidity or alkalinity.

Questions

1. Which substances are acidic?
2. Which substances are basic?

Task

Do a search from books and internet sources to identify natural substances that contain acids or bases. Present the search results to the rest of the class members.

Acidity of a base

The *acidity* of a base is the number of hydroxide ions that a basic molecule can produce in the aqueous solution. On the basis of acidity, bases are classified into three types: monoacidic, diacidic, and triacidic bases. Table 3.4 shows some of the bases with their acidity.

Table 3.4: *Acidity of some bases*

Base	Acidity
$NaOH \longrightarrow Na^+ + OH^-$	Monoacidic
$Ca(OH)_2 \longrightarrow Ca^{2+} + 2OH^-$	Diacidic
$Al(OH)_3 \longrightarrow Al^{3+} + 3OH^-$	Triacidic

Neutralisation reaction

If an acid and a base are mixed together in correct amounts, a neutral solution is produced. A reaction of this type is called *neutralisation*. Neutralisation is therefore a reaction between an acid and a base to produce salt and water. It is essentially the reaction between the hydroxide ions found in the basic solution and the hydrogen ions found in an acidic solution. The reaction is referred to as neutralisation because the

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resulting products are neither basic nor acidic.

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water}$$

Neutralisation reactions that involve the reaction of an acid with a carbonate or hydrogencarbonate produce carbon dioxide gas in addition to salt and water.

$$\text{Acid} + \text{Carbonate/hydrogencarbonate} \longrightarrow \text{Salt} + \text{Water} + \text{Carbon dioxide}$$

The following are examples of neutralisation reactions between bases and dilute acids:

Nitric acid reacts with potassium hydroxide to yield potassium nitrate and water.

$$HNO_3(aq) + KOH(aq) \longrightarrow KNO_3(aq) + H_2O(l)$$

Sulphuric acid reacts with calcium hydroxide to yield calcium sulphate and water

$$H_2SO_4(aq) + Ca(OH)_2(aq) \longrightarrow CaSO_4(aq) + 2H_2O(l)$$

Hydrochloric acid reacts with sodium carbonate to yield sodium chloride, water, and carbon dioxide gas.

$$2HCl(aq) + Na_2CO_3(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

Hydrochloric acid reacts with sodium hydrogencarbonate to yield sodium chloride, water, and carbon dioxide gas.

$$HCl(aq) + NaHCO_3(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

To identify the end point for an acid-base reaction, an indicator must be added. For example, when phenolphthalein indicator is used in a particular *titration* (by running the acid from a burette to a base in a flask), the pink colour of the content in the flask starts to change slowly to colourless. The *end point* is marked when the reaction mixture in the flask just turns colourless. During neutralisation of a base with an acid, the volume of the base in a flask is normally known. The volume of the acid that has neutralised the alkali is called a *titre value*.

To get the exact volume (V) of the acid that neutralises a base, the initial burette reading, V_1 is subtracted from the final burette reading V_2.

$$V = V_2 - V_1$$

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Activity 3.3



Aim: To investigate the neutralisation reaction between sodium hydroxide and dilute hydrochloric acid

Requirements: Pipette, burette, retort stand and clamp, 10-mL measuring cylinder conical flasks, beakers, evaporating dish, white tile, Bunsen burner, 0.1 M sodium hydroxide, 0.1 M hydrochloric acid, and phenolphthalein indicator

Procedure

1. Transfer 25 cm^3 (or 20 cm^3) of sodium hydroxide solution into a conical flask by using a pipette.
2. Add 2 to 3 drops of phenolphthalein indicator to the solution.
3. Set the apparatus as shown in Figure 3.10.

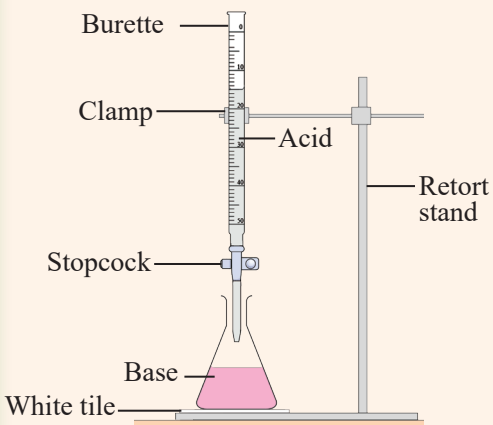


Figure 3.10: *Set-up of the acid-base titration*

4. Fill the burette with dilute hydrochloric acid and record the initial volume, V_1.
5. Carefully, run the acid into the alkali (base), a few drops at a time, until the indicator just changes colour. Swirl the components of the flask after every addition of the acid.
6. Record the new level of the remaining acid in the burette as V_2. Calculate the volume of the acid used in the reaction.
7. In a clean beaker, mix the same volume of the base used with the same volume of the acid obtained from the titration. Do not add any indicator.
8. Transfer 10 cm^3 of the resulting solution into an evaporating dish and carefully heat the solution until crystals start to form.
9. Stop heating and allow the remaining liquid to evaporate, leaving behind a white solid.

Questions

1. Explain the importance of phenolphthalein indicator in the experiment.
2. Write a balanced chemical equation for the neutralisation reaction in the experiment.
3. Write an ionic equation for the neutralisation reaction in the experiment.

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Applications of neutralisation reaction

Neutralisation reaction has many applications in daily life. Some of the applications are explained below.

Treating insect stings and bites

Insect bites or stings, and plants with stinging hair often cause small and red lumps on the skin which may be painful and itchy. Insects such as bees, have stings that inject an acidic liquid in the blood through the skin. The stings can be neutralised by rubbing baking soda on the affected area. Ant bites and nettle (a plant with stinging hair) stings contain methanoic acid which is neutralised by using baking soda or other alkaline substances such as cucumber and avocado. Wasp stings are alkaline and can be neutralised with vinegar which contains acetic acid. Figures 3.11(a) and (b) show different sources of stings or bites.

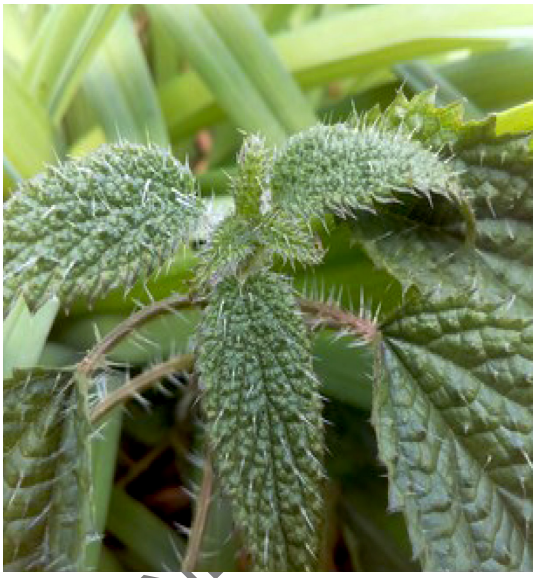


Ant



Wasp

Figure 3.11 (a) *Insects that are sources of stings and bites*



Nettle leaves

Figure 3.11 (b) *Plant with stinging hair*

Relieving indigestion

Indigestion is a discomfort in the stomach that is associated with difficulty in digesting food. It is caused by the presence of excess acid like hydrochloric acid in the stomach. The excess acid can be neutralised by taking a liquid or tablets that contain magnesium or sodium hydrogencarbonate (antacids).

Soil treatment

Most plants grow well in soils that have optimal pH values. When soils are too acidic, for example, due to applications of ammonium based fertilisers, chemicals such as calcium oxide (quicklime) and calcium hydroxide (slaked lime) are added to adjust the soil pH (Figure 3.12). Such chemicals that neutralise the soil acidity are called *liming materials.*

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(a)



(b)

Figure 3.12 (a) and (b): *Soil liming*

Treating factory wastes

Liquid wastes from factories often contain acids and bases. If the wastes get into water bodies such as lakes, ponds and rivers, they can harm aquatic organisms like fish. Acidic wastes can be controlled by adding bases such as calcium hydroxide to neutralise them.

Preventing formation of acid rain

When the pH of rainwater falls below 5.6, it is called *acid rain*. Acid rain is caused by chemical reactions between rainwater and gases such as sulphur dioxide and nitrogen dioxide which are released into the atmosphere. Acid rain increases the acidity of soils, rivers and lakes and adversely affects vegetation and aquatic organisms. To reduce this problem, air pollution devices containing bases are fitted in exhaust pipes and chimneys to neutralise the acidic compounds before reaching the atmosphere.

Neutralising accidental spills

If an acid or an alkali spills on the floor or work surface in the laboratory, it can be neutralised. For example, sulphuric acid which is very corrosive, can be neutralised by adding sodium hydroxide.

Manufacture of fertilisers

The production of ammonium fertilisers is done through the neutralisation of ammonia with a mineral acid. Ammonium nitrate for example, is produced by the reaction of ammonia with nitric acid.

$$NH_3(g) + HNO_3(aq) \longrightarrow NH_4NO_3(aq)$$

Ammonia gas also reacts with sulphuric acid to give ammonium sulphate ($(NH_4)_2SO_4$) fertiliser (Figure 3.13).

$$2NH_3(g) + H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq)$$



Figure 3.13: *Sack of ammonium sulphate fertiliser*

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Uses of acids and bases

Acids and bases have many uses in daily life. For instance, some plants grow well in acidic soil, but others grow better in a neutral or alkaline soil. Farmers need to know the pH of the soil because the soil can be too acidic causing inability of plants to grow well and reduce the yields. Bases can be used by farmers to reduce the acidity of the soil, and acids can be used to correct the pH for the basic soils.

Amino acids and some vitamins such as vitamin C are important for our health. Sulphuric acid is used in car batteries. Citric acid is used in food preservation. Bases such as sodium hypochlorite are used in bleaching to remove stains on clothes. Ammonium hydroxide is used as a glass cleaner, and magnesium hydroxide is used as a laxative to regulate bowel movements. Magnesium hydroxide is also used as an antacid to relieve indigestion, sour stomach and heartburn. Tables 3.5 and 3.6 summarise the uses of some acids, and bases, respectively.

Table 3.5: *Uses of some acids in daily life*

Acids	Uses
Hydrochloric acid	(a) Cleaning metals before electroplating (b) Household cleaning (c) Leather processing (d) Swimming pool maintenance (e) Production of inorganic compounds such as chlorides (f) Production of organic compounds, such as vinyl chloride and dichloromethane during plastic manufacturing (g) Manufacture of fertilisers, textiles, and rubber
Sulphuric acid	(a) Removal of rust from iron and steel in industries (b) Manufacture of synthetic textiles such as nylon (c) Manufacture of fertilisers (d) Used as battery electrolyte (e) Production of sulphates
Nitric acid	(a) Production of nitrogen containing fertilisers, explosives, and rocket propellants (b) Manufacture of nylon (c) Used in testing for genuine precious stones (d) Used as an oxidising agent (e) Metal cleaning (f) Synthesis of dyes

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Table 3.5: (*Continued*)

Acids	Uses
Tartaric acid	(a) Manufacturing of soft drinks (b) Provides tartness (sourness) to food (c) Used as an emetic (a substance that induces vomiting)
Carbonic acid	(a) To make bubbly drinks like soft drinks and sparkling wine (b) Used in pharmaceuticals, cosmetics, fertilisers, and food processing (c) Used as an anaesthetic
Ethanoic acid	(a) A main compound of vinegar (b) Making dyes and paints (c) Degreasing solvent and cleaning agent
Phosphoric acid	(a) Rust proofing (b) Used as a catalyst (c) Used as an ingredient in softening drinks and dental cements (d) Manufacturing of phosphates for softening water (e) Manufacturing of detergents and fertilisers (f) Used as a food additive lending acidic properties to soft drinks and other prepared foods
Benzoic acid	(a) Its salts are used to preserve food (b) Manufacture of a variety of products such as perfumes, dyes, medicines, and insect repellents

Table 3.6: *Uses of some bases in daily life*

Bases	Uses
Ammonia	(a) Production of fertilisers (b) Prevention of premature coagulation in natural or synthetic latex, and is used as a refrigerant (c) Purification of water supplies (d) Used in the manufacture of plastics, explosives, textiles, pesticides, dyes, and other chemicals
Calcium hydroxide	(a) Neutralisation of the acidity of soil (b) Preparation of ammonia (c) Used in water and sewage treatments
Sodium hydroxide	(a) Production of soaps, dyes, detergents, and cleaners (b) Metal cleaning and processing (c) Processing cotton fabrics (d) Electroplating and electrolytic extraction

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Table 3.6: (*Continued*)

Bases	Uses
Magnesium hydroxide	(a) Suspensions of magnesium hydroxide in water are used as antacids and deodorant. (b) Neutralises acidic wastewater because it is a non-hazardous alkali (c) Used as pH control in the food industry (d) Used in chip fabrication for semiconductors
Aluminium hydroxide	(a) Manufacturing of other aluminium compounds, making gastric medicine (antacid), and fire extinguishers (b) Relief of heartburn, sour stomach, and peptic ulcer pain
Potassium hydroxide	(a) Used as an electrolyte in alkaline batteries and in electroplating (b) Making soap, paint, and varnish remover

Exercise 3.1

1. Group the following substances into either acids, bases or neutral solutions:

   Ash solution, fresh milk, sour milk, soap, chalk solution, tomato juice, lemon juice, sugar solution, table salt solution, and cucumber juice.

2. Suppose you are in a village where there are no shops which sell toothpastes, identify which alternative materials available in the environment can be used as toothpastes. Explain you answer.

3. Describe five common uses of acids and bases in daily life.

Acid-base indicators

Acid-base indicators are chemicals that are used to determine the chemical nature of a substance; whether it is acidic or basic. The acid-base indicators are also known as *pH indicators* because acidity and alkalinity relate to pH range. Indicators normally change colour to indicate the presence or absence of acids or bases. The pH indicators operate efficiently over a certain pH range.

pH scale

pH scale is a scale of numbers, from 0 to 14 which is used to express acidity, neutrality, and alkalinity. The pH 7 indicates neutrality of a substance. Any substance that has a pH below 7 is acidic, while a substance with pH above 7 is basic. The smaller the number below 7 in the scale, the more acidic is the substance. Similarly, the greater the number above 7, the more basic is the substance. This means that an acid with pH 1 is more acidic than that with pH 6, and the base with pH 14 is more basic than that with

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pH 8. Figure 3.14 shows the trends in acidity, neutrality, and alkalinity in the pH scale.

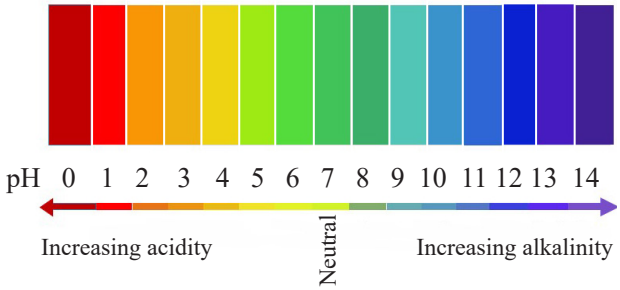


Figure 3.14: *Trends in acidity, neutrality, and alkalinity*

Occurrence of pH indicators

pH indicators may occur naturally and are also synthesised.

Natural pH indicators

A natural pH indicator is a substance which is found naturally and can be used to determine whether the substance is acidic or basic. Examples of natural pH indicators are turmeric, grape juice, red cabbage, cassava leaves, onion, beetroot, and coloured flowers like that of hibiscus. For example, red cabbage juice is pink in acids and green in bases. A hibiscus flower extract has a pink colour in an acid and green colour in a base. Onion is one of the olfactory indicators which diminishes its smell in a base and remains as it is in an acid. Vanilla essence has similar characteristics to onion. The smell of vanilla essence disappears when it is added to a base, but persists when it is added to an acid.

The most commonly used natural indicator is litmus which is extracted from lichens (Figure 3.15). Litmus has a purple, red, and blue colour in distilled water, acidic solution, and basic solution, respectively. Litmus is available either in solution, or in the form of strips of paper.



Figure 3.15: *Lichen*

Activity 3.4



Aim: To prepare a pH indicator from hibiscus flowers

Requirements: Hibiscus flowers, water, beakers, droppers, mortar and pestle, knife, 50% ethanol solution, heat source, evaporating dish, dilute hydrochloric acid, and dilute sodium hydroxide

Procedure

1. Cut the hibiscus flowers into small pieces and then crush them by using mortar and pestle.

2. Add about 20 cm^3 of ethanol solution into the crushed hibiscus flowers and continue crushing. Decant the mixture obtained and then keep the solution obtained aside.

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3. Boil the solution obtained to make sure that all the ethanol evaporates. Leave the residue to cool.
4. Add about 20 cm^3 of distilled water to the residue and stir to obtain an indicator solution.
5. Establish the colour of an indicator prepared by testing it with hydrochloric acid and sodium hydroxide, one at a time.

Questions

1. Describe the colour changes observed when you tested the hibiscus flower extract as an indicator.
2. What other substances found in your environment can be used as acid-base indicators?

Synthetic pH indicators

Synthetic pH indicators are chemical substances which are made from different substances in the laboratories used to determine the pH of substances. Examples of synthetic indicators include universal indicator, phenolphthalein (POP), and methyl orange (MO).

Universal indicator

Universal indicator is a chemical substance that displays several colour changes across a wide range of pH values as shown in Table 3.7. Basically, a universal indicator is a blend of pH indicator solutions which include phenolphthalein, methyl red, bromothymol blue, and thymol blue. When an acid or base solution is added to the universal indicator, a new colour is produced which is used to determine the pH value of the acid or the base solution by matching the colour with the colours on the pH colour chart.

Table 3.7: *pH ranges of the universal indicator*

pH value	Colour of the universal indicator
1-2	Red
3	Pink
4	Brown
5	Yellow
6-8	Green
9-10	Blue
11-12	Indigo
13-14	Violet

Other indicators

There are several other indicators such as methyl orange, methyl red, bromothymol blue, phenol red, thymol blue and phenolphthalein which are used to test the acidity or alkalinity of acids and bases. Table 3.8 shows the colour changes and the pH ranges of selected indicators.

Table 3.8: *Colour changes of common indicators in acidic and basic solutions*

Indicator	Acidic	Basic	pH range
Phenolphthalein	Colourless	Pink	8–10
Methyl orange	Red	Yellow	3–5
Methyl red	Yellow	Red	5–8
Phenol red	Yellow	Red	7–8
Bromothymol blue	Yellow	Red	6–8

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Activity 3.5



Aim: To test the acidity and alkalinity of substances by using different indicators

Requirements: Dilute sulphuric acid or dilute hydrochloric acid, sodium hydroxide solution, red and blue litmus papers, litmus solution, beakers, phenolphthalein, bottled drinking water, and methyl orange

Procedure

1. Insert a blue litmus paper into dilute sulphuric acid or hydrochloric acid. Record the observations.
2. Insert a red litmus paper into a sodium hydroxide solution, and observe any change.
3. Add two to three drops of litmus solution into a dilute acid solution, and then into sodium hydroxide solution. Record the observations.
4. Repeat Step 3 using phenolphthalein instead of litmus solution. Do the same by using methyl orange. Record the observations in each case.
5. Test the acidity and alkalinity of bottled drinking water using all the above indicators and record your observations.

Questions

1. What are the colour changes in each of the tested samples?
2. In Step 5, you used the same indicators that were used in testing acids and alkalis, but in this case you obtained quite different results.
   (a) What do the results indicate?
   (b) What other substances may give similar results?

Uses of pH indicators

pH indicators are used in a variety of ways, including the following:

1. They are used to give rough pH values of chemical solutions and drinks such as fruit juices and water.
2. They are used in acid-base titrations to show the completion of a reaction (end point).
3. They are used in agriculture to test the acidity or alkalinity of soil.
4. They are used to test the pH of water used in swimming pools to ensure that the water used is approximately neutral.
5. They are used in wastewater monitoring to avoid environmental pollution.

Exercise 3.2



1. Why is it important to use indicators in acid-base titration?
2. Why are alcohols used in the preparation of indicators from plant leaves or flowers?

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3. What will happen if water is used instead of an alcohol during the preparation of an indicator from plant leaves or flowers?

Salts

A *salt* is an ionic compound which is made up of positively charged ions (cations) and negatively charged ions (anions). Salts occur naturally and are also synthesised in the laboratory. The main natural sources of salts include sea water and rocks. Different salts are used in our homes and industries for various purposes. These include table salt (NaCl), washing soda ($Na_2CO_3 \cdot 10H_2O$), Plaster of Paris $(CaSO_4)_2 \cdot H_2O$, and bleaching powder ($Ca(OCl)_2$).

Types of salts

Salts are categorised into neutral salts, acidic salts, and basic salts depending on their pH values when dissolved in water. Salts can also be categorised as normal salts, and double salts depending on their formation and compositions.

Neutral salts

These are formed by the reactions of strong acids and strong bases. They are called neutral salts because their aqueous solutions are neutral to litmus. Neutral salts, when categorised basing on formation and composition, are termed as normal salts. A normal salt is formed by complete replacement of replaceable hydrogen atoms from an acid molecule by a metal, or a group of species acting like a metal. Examples of these salts are NaCl, KCl, K_2SO_4, $KClO_3$, $MgCl_2$, $CuSO_4$, $MgSO_4$, and $NaNO_3$.

Acidic salts

Acidic salts are those salts that produce acidic solutions when dissolved in water. They are formed when a strong acid reacts with a weak base. Their aqueous solutions turn blue litmus paper red. Acidic salts may contain one or more replaceable hydrogen atoms present in an acid molecule or positive radical. Examples of acidic salts include $FeCl_3$, $ZnCl_2$, $HgCl_2$, $Fe_2(SO_4)_3$, $NaHSO_4$, $NaHPO_4$, $(NH_4)_2SO_4$, and NH_4Cl.

Basic salts

Basic salts are a class of salts that produce alkaline solutions when dissolved in water. Their aqueous solutions turn red litmus paper blue. A basic salt is a product of a reaction between a strong base and a weak acid. Examples of basic salts include CH_3COONa, Na_2CO_3, $NaHCO_3$, and $Na_2C_2O_4$.

Double salts

A double salt is a mixture of two salts which when dissolved in water gives two different cations or anions. Examples are potash alum, $[KAl(SO_4)_2 \cdot 12H_2O]$, and ferric alum $[NH_4Fe(SO_4)_2 \cdot 12H_2O]$.

Preparation of salts

There are several methods of preparing salts. The selection of the method to be used depends on whether the salt required is soluble or insoluble in water.

Preparation of soluble salts

Soluble salts are prepared by reacting acids with metals, alkalis, insoluble bases or carbonates.

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Reactions of acids with metals (displacement method)

A displacement method involves the reaction of acids and metals above hydrogen in the electrochemical series. At the end of the reactions, the hydrogen is replaced by metals to produce salt and hydrogen gas. The displacement method is suitable for preparing salts of moderately reactive metals such as magnesium, aluminium, zinc, iron, and tin. Highly reactive metals such as K, Na, and Ca are too fast to cause explosion when reacting with the acids. The displacement reaction occurs as shown in the general equation:

Metal + Acid ⟶ Salt + Hydrogen gas

For example, zinc displaces hydrogen from sulphuric acid as shown in the following reaction equation:

$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$

Activity 3.6



Aim: To prepare zinc sulphate

Requirements: Beaker, evaporating dish, glass rod, spatula, tripod stand, wire gauze, filter papers, filter funnel, Bunsen burner, zinc powder, and 1 M dilute sulphuric acid

Procedure

1. Put about 25 cm^3 of 1 M dilute sulphuric acid in a beaker, and then add some zinc powder into it. Mix the contents by using a glass rod.
2. Continue adding more zinc to the beaker to excess until no more effervescence is observed.
3. Filter out the excess zinc and pour about 10 cm^3 of the filtrate into an evaporating dish.
4. Heat the filtrate while ensuring that not all the water is evaporated.
5. Allow the heated filtrate to cool and then filter off any excess liquid.
6. Allow the crystals to dry by rubbing them between filter papers. Record the observations.

Note: The use of excess zinc powder in this experiment is to ensure that all the hydrogen ions in the acid react to form hydrogen gas. On evaporating the filtrate, it is not allowed to dry up completely. This allows the zinc sulphate formed to maintain its *water of crystallisation,* colour, and shape. *Crystallisation* is the process of forming solid crystals from a homogeneous solution. The solution is evaporated to make it *saturated* with the solute up to the point where it cannot hold any more solute. The temperature at which crystals form as the solution evaporates is called *crystallisation point*.

Questions

1. How can you confirm whether $ZnSO_4$ is acidic, basic or neutral? Explain.

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2. Using a chemical equation, explain what will happen if magnesium is used instead of zinc powder.
3. Give a balanced chemical equation for the reaction occurring between sulphuric acid and zinc powder.

Reactions of acids with metal oxides and metal hydroxide (*neutralisation method*)

The neutralisation method is usually used to prepare soluble salts such as sodium, potassium, and ammonium salts. Metal oxides and metal hydroxides react with acids to produce salt and water only.

Metal oxide + Acid ⟶ Salt + water

Metal hydroxide + Acid ⟶ Salt + water

Activity 3.7



Aim: To prepare potassium chloride

Requirements: Pipette, burette, conical flasks, dropper, evaporating dish, tripod stand, wire gauze, Bunsen burner, 2 M potassium hydroxide solution, 2 M hydrochloric acid solution, phenolphthalein solution, and distilled water

Procedure

1. Transfer 25 cm^3 of 2 M potassium hydroxide solution into a conical flask.
2. Add two drops of phenolphthalein solution to the potassium hydroxide solution in the conical flask.
3. Fill the burette with 2 M of hydrochloric acid. Record the initial reading.
4. Slowly run the acid from the burette into the conical flask while gently swirling the flask to mix the acid and the base. Close the burette when the solution just changes its colour. Record the volume of the acid used.
5. Transfer about 10 cm^3 of the solution obtained into an evaporating dish and heat until all the water evaporates.

Questions

1. With the aid of a balanced chemical equation, explain what happens in Step 4.
2. In Activity 3.6, zinc sulphate filtrate was not allowed to dry up completely during evaporation, but in this experiment, potassium chloride solution is allowed to evaporate to dryness. Explain.

Reactions of acids with insoluble bases

Reactions of acids with insoluble bases result in the formation of soluble salts. For

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example, insoluble bases of zinc oxide, and copper(II) oxide react with acids to form soluble salts like zinc sulphate, copper(II) sulphate, copper(II) chloride, and water as products.

Insoluble base + Acid ⟶ Salt + Water

Activity 3.8



Aim: To prepare copper(II) chloride

Requirements: 100-mL measuring cylinder, tripod stand, beaker, glass rod evaporating dish, wire gauze, filter papers, funnel, Bunsen burner, 2 M hydrochloric acid solution, and copper(II) oxide

Procedure

1. Pour about 50 cm^3 of 2 M hydrochloric acid into a beaker and heat gently.
2. Dissolve some copper(II) oxide little by little in the hot acid and stir for some time until the metal oxide is in excess.
3. Filter off the solid and heat the filtrate until it approaches to dryness in order to crystallise the resulting salt.

Note: Crystallisation of copper(II) chloride may take time.

Questions

1. With the aid of a chemical equation, explain the reaction taking place in this experiment.
2. Why was it necessary to use hot hydrochloric acid in the reaction with copper(II) oxide?
3. What is the colour of the product formed when CuO is added in hot dilute HCl?

Reactions of acids with metal carbonates

Most metal carbonates such as calcium carbonate, copper(II) carbonate and potassium carbonate react with acids to produce salt, water and carbon dioxide gas as shown in the following equation:

Carbonate + Acid ⟶ Salt + Water + Carbon dioxide

Activity 3.9



Aim: To prepare copper(II) sulphate

Requirements: Beaker, evaporating dish, glass rod, wire gauze, measuring cylnder, filter paper, filter funnel, tripod stand, Bunsen burner, dilute sulphuric acid, and copper(II) carbonate

Procedure

1. Transfer about 25 cm^3 of dilute sulphuric acid into a small beaker and heat gently.
2. Add solid copper(II) carbonate to the beaker containing the hot acid and stir the mixture by using a glass rod until it is saturated.

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3. Filter the mixture to remove excess copper(II) carbonate and crystallise the filtrate by evaporation.

Note: Do not allow evaporation to its completion.

Questions

1. Write a balanced chemical equation for the reaction that occurred in this experiment.
2. How can the gas formed in this experiment be tested?

Reactions of acids with metal hydrogencarbonates

Metal hydrogencarbonates, or commonly known as bicarbonates, are compounds in the form of $MHCO_3$, where M represents a metal. These compounds react with acids in the same way as metal carbonates to form a salt, water, and carbon dioxide gas. For example, sodium hydrogencarbonate ($NaHCO_3$) reacts with dilute hydrochloric acid to produce sodium chloride, water, and carbon dioxide gas.

$$HCl(aq) + NaHCO_3(s) \longrightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

Preparation of salts by direct combination

If salts are required in anhydrous forms, they cannot be prepared by wet methods. In this case, it is possible to prepare them by direct combination. For example, in the preparation of iron(III) chloride, dry chlorine gas is passed over hot iron wire or steel wool as shown in Figure 3.16. The iron(III) chloride vapour produced is condensed to give a solid salt. This is indicated in the following equation:

$$2Fe(s) + 3Cl_2(g) \xrightarrow{\text{heat}} 2FeCl_3(s)$$

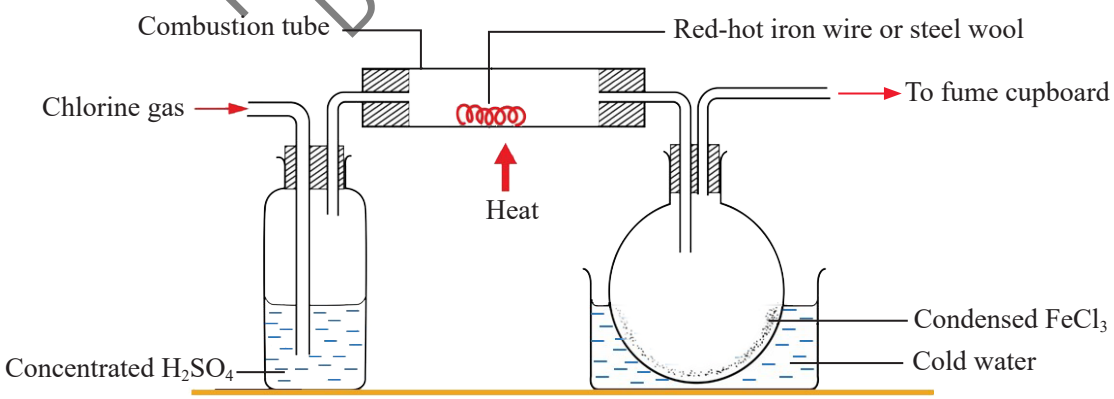


Figure 3.16: *Set-up of the direct preparation of iron(III) chloride*

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Preparation of insoluble salts

Insoluble salts are prepared by using ionic precipitation or double decomposition method. In this method, two soluble salts are used to form an insoluble salt and a soluble salt. The insoluble salt precipitates, while the soluble salt remains in solution. The precipitate is filtered and washed with distilled water and then dried. For example, to prepare insoluble lead(II) iodide, the solution containing lead ions and another containing iodide ions are mixed. Such solutions could be lead(II) nitrate and potassium iodide. A bright yellow precipitate of lead(II) iodide is formed after mixing. The chemical equation for this reaction is as follows:

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

Other common insoluble salts which can be prepared by using this method include calcium sulphate ($CaSO_4$), magnesium carbonate ($MgCO_3$), silver chloride ($AgCl$), barium carbonate ($BaCO_3$), barium sulphate ($BaSO_4$), and lead(II) sulphate ($PbSO_4$).

Project

Common salt is found in different places in Tanzania including lakes and the ocean shores. During your vacation, take time to visit one of the nearby places where common salt is produced. Study how the salt is produced and write a brief report. Present your findings to the rest of the class members when you are back to school.

Properties of salts

Salts exhibit different physical and chemical properties.

Physical properties of salts

The physical properties of salts include their appearance in colour, texture, and solubility.

Physical appearance of salts

Salts range from crystalline to powder in appearance and exhibit different colours depending on their types. For example, hydrated copper(II) sulphate is a blue crystalline salt, while sodium chloride is a white crystalline salt. Ferrous chloride is a green crystalline salt, while calcium nitrate is a white crystalline salt. Calcium carbonate is powdered white. Figure 3.17 shows some examples of crystalline and powdered salts.



Hydrated copper(II) sulphate | Sodium chloride | Iron(II) chloride | Calcium nitrate | Calcium carbonate

Figure 3.17: *Crystalline and powdered salts*

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Solubility of salts

Solubility of a solute in a given solvent is the number of the grams of the solute required to saturate 100 grams of solvent at a given temperature, or it is the amount of grams or moles of a substance that dissolve in a solvent to form a saturated solution at a given temperature. Normally, when a soluble salt is put into a glass of water, it dissolves. In this case, salt is a solute and water is a solvent. The solubility of a solute in a solvent is not fixed but varies with temperature. Some solutes are more soluble in certain solvents than others. For example, 40 g of sodium chloride can dissolve in 100 g of water at 50 °C, while 80 g of potassium nitrate can dissolve in 100 g of water at 50 °C. In this case, potassium nitrate, is more soluble than sodium chloride. Some salts such as sodium chloride and potassium nitrate are soluble in water, while others such as calcium carbonate and lead chloride are insoluble.

The solubility of some salts is summarised in Table 3.9.

Table 3.9: *Soluble and insoluble salts*

Salts	Soluble salts	Insoluble salts
Nitrates	All nitrates	None
Halides (fluorides, chlorides, bromides, and iodides)	All halides except those of Pb, Hg, and Ag	Silver, lead, and mercury halides such as AgCl and $PbCl_2$
Sulphates	All sulphates except $CaSO_4$, $BaSO_4$, and $PbSO_4$	$CaSO_4$, $BaSO_4$, and $PbSO_4$
Carbonates	All Group I metal carbonates and ammonium carbonate	All carbonates except Group I metal carbonates and ammonium carbonate
Hydrogencarbonates	All Group I metal hydrogencarbonates and ammonium hydrogencarbonate	All hydrogencarbonates except Group I metal hydrogencarbonates and ammonium hydrogencarbonate

Activity 3.10



Aim: To investigate the solubility of salts

Requirements: Seven test tubes, test tube racks, sodium carbonate, calcium nitrate, calcium carbonate, sodium sulphate, potassium chloride, lead chloride, ammonium carbonate, distilled water, and spatula

Procedure

1. Put about 2 cm^3 of distilled water into the seven test tubes.

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2. Put a pinch of each salt in a separate test tube containing an equal amount of water, then shake gently. Record the observations in each test tube.

Questions

1. Which salts dissolve readily in water?
2. Which salts do not dissolve in water?
3. Is it possible to test the solubility of all salts in the laboratory? Explain.

Chemical properties of salts

Different salts exhibit different chemical properties. For example, they decompose upon heating to produce different substances depending on the types of their metals and positions in the reactivity series.

Action of heat on salts

Different salts behave differently when heated. Some salts will decompose after slight heating, yet others have to be heated strongly to decompose. Some salts do not decompose at all, even with strong heating.

(a) Action of heat on carbonates

Upon heating, metal carbonates decompose to form metal oxides and carbon dioxide gas. The decomposition reaction can be represented by the following equation, where M represents a particular metal:

$$MCO_3(s) \xrightarrow{\text{heat}} MO(s) + CO_2(g)$$

Sodium and potassium carbonates do not decompose even when strongly heated. The carbonates of magnesium and calcium decompose to give magnesium oxide and calcium oxide, respectively, with the evolution of carbon dioxide gas.

$$MgCO_3(s) \xrightarrow{\text{heat}} MgO(s) + CO_2(g)$$

$$CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$$

Other metal carbonates decompose easily on heating. For example, green copper(II) carbonate decomposes easily upon heating to give black copper(II) oxide and carbon dioxide gas.

$$CuCO_3(s) \xrightarrow{\text{heat}} CuO(s) + CO_2(g)$$

Zinc carbonate decomposes easily on heating to form zinc oxide and carbon dioxide gas. The zinc oxide produced is yellow when hot and white when cold.

$$ZnCO_3(s) \xrightarrow{\text{heat}} ZnO(s) + CO_2(g)$$

Lead(II) carbonate decomposes when heated to give lead(II) oxide and carbon dioxide gas.

$$PbCO_3(s) \xrightarrow{\text{heat}} PbO(s) + CO_2(g)$$

Ammonium carbonate behaves so differently when heated. It decomposes slowly at room temperature to produce ammonia gas, water, and carbon dioxide.

$$(NH_4)_2CO_3(s) \xrightarrow{\text{heat}} 2NH_3(g) + H_2O(g) + CO_2(g)$$

A strong pungent smell is produced due to the presence of ammonia gas. The evolved gas turns limewater milky, indicating the presence of carbon dioxide. It also turns

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a blue cobalt(II) chloride paper pink indicating the presence of water vapour.

A red litmus paper turns blue due to the basic condition of ammonia gas in the solution. Table 3.10 summarises the behaviour of some carbonates on heating.

Table 3.10: *Effects of heat on some carbonate salts*

Carbonate salt	Behaviour on heating
Potassium carbonate and sodium carbonate	Do not decompose
Calcium carbonate, magnesium carbonate, zinc carbonate, iron(II) carbonate, lead(II) carbonate, and copper(II) carbonate	Decompose into corresponding metal oxides and carbon dioxide gas
Ammonium carbonate	Decomposes into ammonia gas, carbon dioxide gas, and water

Activity 3.11



Aim: To investigate the effects of heat on carbonate salts

Requirements: Sodium carbonate, magnesium carbonate, calcium carbonate, copper(II) carbonate, zinc carbonate, lead(II) carbonate, limewater, test tubes, spatula, Bunsen burner, red litmus paper, wire gauze, tripod stand, test tube and bent delivery tube

Procedure

1. Put a spatulaful of zinc carbonate into a dry test tube.
2. Hold the test tube with a test tube holder and heat over a non-luminous flame.
3. Pass the gas evolved into a test tube of limewater. Record the results.
4. Put a moist red litmus paper at the mouth of the test tube, and record the observations.
5. Observe the colour change of the compound when hot and when cold.
6. Repeat Steps 1–5 for the other carbonates. Record all the observations.

Questions

1. Explain the differences observed when heating zinc carbonate and lead(II) carbonate.
2. What is common between the two compounds in Question 1?

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Activity 3.12



Aim: To investigate the effect of heat on ammonium carbonate

Requirements: Spatula, test tube, test tube holder, Bunsen burner, cobalt(II) chloride paper, red litmus paper, ammonium carbonate, and limewater

Procedure

1. Put a spatulaful of ammonium carbonate salt into a dry test tube.
2. Hold the test tube with a test tube holder and heat the salt over a non-luminous flame.
3. Pass the gas evolved into a test tube containing limewater. Record the observation.
4. Put a moist red litmus paper and cobalt(II) chloride paper at the mouth of the test tube, and record the observations.

Questions

1. What differences have you observed between this activity and those from Activity 3.11?
2. Write the equation representing the action of heat on the salt in Step 2.

(b) Action of heat on nitrates

Most nitrates decompose on heating to give the metal oxide, brown fumes of nitrogen dioxide, and oxygen gas. Nitrates of alkali metals such as sodium and potassium decompose slowly on heating to give oxygen and the metal nitrite.

$$2NaNO_3(s) \xrightarrow{\text{heat}} 2NaNO_2(s) + O_2(g)$$

Nitrates of calcium and copper decompose into their respective oxides, nitrogen dioxide gas, and oxygen gas.

$$2Ca(NO_3)_2(s) \xrightarrow{\text{heat}} 2CaO(s) + 4NO_2(g) + O_2(g)$$

$$2Cu(NO_3)_2(s) \xrightarrow{\text{heat}} 2CuO(s) + 4NO_2(g) + O_2(g)$$

Exceptional behaviours are shown by ammonium nitrate and silver nitrate. Ammonium nitrate decomposes to give dinitrogen oxide gas and water, whereas silver nitrate decomposes to give silver metal, nitrogen dioxide, and oxygen gas.

$$NH_4NO_3(s) \xrightarrow{\text{heat}} N_2O(g) + 2H_2O(l)$$

$$2AgNO_3(aq) \xrightarrow{\text{heat}} 2Ag(s) + 2NO_2(g) + O_2(g)$$

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Table 3.11 summarises the effects of heat on nitrates.

Table 3.11: *Effects of heat on some nitrate salts*

Nitrate salts	Products
Nitrates of potassium and sodium	Metal nitrite + Oxygen
Nitrates of calcium, magnesium, aluminium, zinc, iron, lead, and copper	Metal oxide + Nitrogen dioxide + Oxygen
Nitrates of silver and mercury	Metal + Nitrogen dioxide + Oxygen
Ammonium nitrate	Dinitrogen oxide + water

(c) Action of heat on sulphates

Sulphates are generally more stable to heat than nitrates. Even sulphates of metals that are low in the reactivity series must be strongly heated to decompose. Sulphates of alkali metals and alkaline earth metals do not decompose when heated. However, few sulphates such as ammonium sulphate and copper(II) sulphate decompose when heated. For example, ammonium sulphate, decomposes on heating to give sulphuric acid and ammonia gas. During heating, the sulphate first decomposes to give ammonium hydrogen sulphate as shown in the following equations:

$$(NH_4)_2SO_4(s) \xrightarrow{\text{heat}} NH_4HSO_4(l) + NH_3(g)$$

On cooling the product, ammonia gas, and sulphuric acid are formed.

$$NH_4HSO_4(l) \xrightarrow{\text{cool}} NH_3(g) + H_2SO_4(l)$$

Activity 3.13



Aim: To investigate the action of heat on sodium sulphate and iron(II) sulphate

Requirements: Ignition tube, spatula, Bunsen burner, sodium sulphate, iron(II) sulphate, and red and blue litmus papers

Procedure

1. Put a spatulaful of sodium sulphate in a hard ignition tube and heat strongly.
2. Insert the wet blue and red litmus papers in the ignition tube while heating. Record the observations.
3. Repeat Steps 1 and 2 by using iron(II) sulphate in the place of sodium sulphate. Record the observations.

Question

Comment on the observations made from the two experiments.

(d) Action of heat on ammonium salts

All ammonium salts decompose on heating. For example, ammonium chloride decomposes (sublimes) into hydrogen chloride and ammonia

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gases as shown in the following chemical equation:

$$NH_4Cl(s) \overset{\text{heat}}{\rightleftharpoons} HCl(g) + NH_3(g)$$

Deliquescency, hygroscopy, and efflorescency in salts

Basing on the exposure to air, salts are categorised as deliquescent, hygroscopic or efflorescent salts.

Deliquescent salts

A salt that absorbs moisture from the air (environment) to form a solution is called a *deliquescent salt.* Examples are sodium hydroxide, phosphorus oxides, potassium hydroxide, and calcium chloride.

Hygroscopic salts

A *hygroscopic* salt is the one which absorbs water from air without necessarily forming a solution. Examples of hygroscopic salts are sodium chloride, sodium nitrate, calcium chloride, calcium oxide, and copper sulphate. The tendency of calcium chloride to absorb water vapour explains its use as a drying agent for many gases, except ammonia gas because it reacts with it. If water molecules become suspended among the substances, that can lead to physical changes of the salts (change in volume, boiling point, and viscosity).

Efflorescent salts

An efflorescent salt is the one which when left in air loses all the water of crystallisation. Examples are hydrated sodium carbonate ($Na_2CO_3 \cdot 10H_2O$), hydrated sodium sulphate ($Na_2SO_4 \cdot 10H_2O$), and hydrated iron(II) sulphate ($FeSO_4 \cdot 7H_2O$).

Uses of salts

Salts have a lot of uses in our everyday lives. Some of the uses are the control of soil pH, as antacids, fertilisers, and alleviation of health disorders. Salts are also used in food flavouring and preservation, and in production of pesticides.

Control of soil pH: Soils that are too acidic or too alkaline are not suitable for crop production. Calcium oxide (lime) is usually added to acidic soils to neutralise the acidity. When the soil is too alkaline, calcium sulphate (gypsum) is added to neutralise the condition.

Uses of salts as antacids: Antacids are substances that are in the forms of hydroxides or salts. These are used to relieve heartburn and acid build-up in the stomach. Some of the common antacids are magnesium sulphate (Epsom salt), aluminium hydroxide, and sodium hydrogencarbonate.

As inorganic fertilisers: Inorganic fertilisers are salts. Examples of these are ammonium sulphate, ammonium nitrate, and calcium phosphate.

Alleviation of health disorders: Various salts are used to alleviate problems that are associated with different health conditions. These salts include chlorides, phosphates, and sulphates of sodium, potassium, and calcium.

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Production of pesticides: Certain salts are used as pesticides to kill or control pests such as insects, rodents, weeds and fungi. For example, copper(II) sulphate, iron(II) sulphate, and sodium chlorate(V) are used as pesticides for control of fungi in crops.

Other uses of salts: There are many uses that are specific to some salts as summarised in Table 3.12.

Table 3.12: Uses of some salts in daily life

S/N	Salt	Uses
1.	Sodium chloride (NaCl)	(a) An essential ingredient in food (b) Preservation of food (c) Manufacturing of soaps (d) Making a freezing mixture which is used by ice-cream vendors
2.	Sodium carbonate (Na_2CO_3) (Washing soda)	(a) To remove stain from clothes (b) Production of detergents and soaps (c) Manufacture of glass, papers, and textiles (d) Softening hard water (e) In fire extinguishers
3.	Sodium hydrogencarbonate ($NaHCO_3$)	(a) Used as baking soda (b) In fire extinguishers (c) As an antacid in medicine
4.	Ammonium chloride (NH_4Cl)	(a) Used as an electrolyte in the manufacture of dry batteries (b) Used as an acidifier
5.	Copper(II) sulphate ($CuSO_4$)	(a) Used as a fungicide to kill certain germs (b) In electroplating (c) In dyeing (d) In correction of copper deficiency in soil and animals (e) Preparation of catalysts
6.	Silver bromide (AgBr)	Making photographic films
7.	Calcium chloride ($CaCl_2$)	(a) Used as a drying agent (b) In freezing mixtures (c) Used as food additive (d) In water treatment plants (e) As a food preservative

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Table 3.12: (*Continued*)

S/N	Salt	Uses
8.	Calcium carbonate ($CaCO_3$)	(a) Plays an important role as a construction and building material (marble) or as an ingredient in cement (b) In medicinal industries to manufacture antacids and tablets (c) Manufacture of paints, papers, and plastics
9.	Potash alum $[KAl(SO_4)_2 \cdot 12H_2O]$	(a) Purifying water; makes suspended particles in water settle down (b) Used as an antiseptic (c) In dyeing
10.	Ammonium salts	(a) Manufacture of plastics, synthetic fibres, dyes, explosives, and pharmaceuticals (b) Used as fertilisers

Exercise 3.3

1. Classify salts basing on the responses of their aqueous solutions to litmus paper.
2. Explain the uses of salts.
3. Briefly explain the following terms giving two examples in each:
   (a) Deliquescent salt
   (b) Efflorescent salt
   (c) Hygroscopic salt

Chapter summary

1. An acid is a chemical substance which produces hydrogen ions (H^+) in water as the positively charged ions.
2. A base refers to a chemical substance which when dissolved in water produces hydroxide ions (OH^-) as negatively charged ions. It is also defined as a metal oxide or hydroxide which neutralises an acid. An alkali is a soluble base.
3. A strong acid or strong base dissociates completely in water. A weak acid or weak base dissociates partially in water.
4. An acid or alkali solution is concentrated if it contains a large amount of acid or alkali in a relatively small amount of water.
5. A dilute acid or alkali has a small amount of acid or alkali dissolved in a large amount of water.
6. A pH indicator is a chemical substance that exhibits different colours in solutions of different acidities or alkalinities.

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7. The basicity of an acid is the number of hydrogen atoms per molecule of the acid that can be displaced by a metal in a solution. An acid that has only one hydrogen atom that can be displaced is said to be monobasic. An acid with a basicity of two is dibasic, while the one with basicity of three is tribasic.
8. Neutralisation is a chemical reaction in which an acid reacts with a base to form salt and water.
9. Neutralisation is applied in various useful situations such as treating insect stings, relieving indigestion, neutralising harmful substances in the environment and in the manufacture of fertilisers.
10. Acids and alkalis are used in the manufacture of various industrial products.
11. Salt is an ionic compound which is made up of the positively charged ion (cation) and negatively charged ion (anion).
12. Many salts are soluble in water. Soluble salts are made by reacting:
    (a) an acid and excess metal;
    (b) an acid and excess metal oxide;
    (c) an acid and a metal hydroxide;
    (d) an acid and a metal carbonate; or
    (e) an acid and a metal hydrogencarbonate.
13. Insoluble salts are prepared through ionic precipitation, while anhydrous salts are prepared through direct synthesis reactions.
14. The solubility of a solute in a given solvent is the number of grams of the solute that will saturate 100 grams of a solvent at a given temperature.
15. All sodium, potassium, and ammonium salts are soluble in water.
16. All nitrates are soluble in water.
17. All chlorides are soluble in water except silver, mercury, and lead chlorides.
18. All sulphates are soluble in water except barium, calcium and lead sulphates.
19. All carbonates are insoluble in water except sodium, potassium, and ammonium carbonates.
20. Salts have various uses. These include: controlling soil pH, as antacids, as inorganic fertilisers, alleviating health disorders, and in the preservation and seasoning of food.

Revision exercise 3

Choose the correct answer in items 1–6.

1. Which of the following is a natural acid?
   (a) Nitric acid
   (b) Phosphoric acid

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(c) Citric acid
(d) Sulphuric acid

2. One of the following statements best explains the meaning of an acid:
   (a) It is any substance that contains hydrogen in its formula.
   (b) It is any substance that dissolves in water to produce hydrogen ions.
   (c) It is any substance that can react with NaOH.
   (d) It is any substance that contains oxygen in its formula.

3. What name is given to the process of dissolving sodium chloride in water?
   (a) Precipitation process
   (b) Neutralisation reaction
   (c) Hydration process
   (d) Dissolution process

4. Which of the following salts is soluble in water?
   (a) Calcium carbonate
   (b) Copper carbonate
   (c) Magnesium carbonate
   (d) Sodium carbonate

5. Which one of the following substances is produced when calcium carbonate is strongly heated?
   (a) An oxide
   (b) A smoke
   (c) High heat
   (d) A bicarbonate

6. Which of the following salts decomposes to its own metal?
   (a) Sodium chloride
   (b) Potassium carbonate
   (c) Zinc sulphate
   (d) Silver nitrate

7. Choose the correct word or term from the given terms in brackets that best describes each of the following phrases:
   (a) The process by which hydrogen ion reacts with hydroxide ions to produce water (titration, neutralisation).
   (b) Dissolved ammonia in water (weak base, weak acid).
   (c) Materials that can be used to neutralise excess stomach acidity (ant-pain, antacid).
   (d) The pH of a solution (hydrogen ion, salt)
   (e) The quantity of a chemical substance in a solution (concentration, strength).
   (f) A substance that produces hydroxide ions when it dissolves in water (an acid, an alkali).

8. Explain the meaning of the following terms and give an example for each:
   (a) pH indicators
   (b) Double salt
   (c) Titration
   (d) Solubility

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9. Sort out the following salts into groups of soluble and insoluble salts:
   Sodium carbonate, lead nitrate, lead sulphate, calcium carbonate, copper(II) sulphate, sodium nitrate, zinc chloride, silver chloride, barium carbonate, sodium sulphate, sodium chloride, potassium chloride, ammonium carbonate, and barium sulphate.

10. With reasons, suggest the best methods for preparing the following compounds:
    (a) Silver chloride
    (b) Lead(II) sulphate
    (c) Copper(II) nitrate
    (d) Calcium carbonate
    (e) Anhydrous zinc chloride

11. Write a balanced chemical equation for the preparation of each of the following compounds using iron and another substance:
    (a) Iron(II) sulphate
    (b) Iron(II) chloride
    (c) Iron(III) chloride

12. A few drops of a solution of potassium iodide were added to a solution of lead nitrate.
    (a) State the observation noted.
    (b) Write an ionic equation for the observation noted.

13. (a) How do different carbonates, nitrates, chlorides, and sulphates of metals behave when heated?

    (b) Complete and balance the following chemical equations:

    (i) $PbCO_3(s) \xrightarrow{heat}$

    (ii) $Ca(NO_3)_2(s) \xrightarrow{heat}$

    (iii) $(NH_4)_2CO_3(s) \xrightarrow{heat}$

    (iv) $NH_4Cl(s) \xrightarrow{heat}$

14. How are salts useful in our life in the following aspects?
    (a) Foods
    (b) Agriculture
    (c) Medicine

15. List three physical properties of a named acid and a base.

16. Differentiate between the terms strength of an acid and concentration of an acid.

17. Write the balanced equations and net ionic equations for the reactions between sodium hydroxide and the following acids:
    (a) Sulphuric acid
    (b) Nitric acid

18. Name the acids found in each of the following substances:
    (a) Lemon juice
    (b) Vinegar

19. A chemistry student tested five solutions A, B, C, D, and E with a universal indicator solution to determine their pH values. The

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results obtained are shown in the following table:

Beaker of solution	A	B	C	D	E
pH	1	5	7	9	14

Which solution was a;

(i) strong alkali?
(ii) strong acid?
(iii) weak acid?
(iv) neutral solution?
(v) weak base?

20. During a neutralisation reaction, a student placed 25 cm^3 of sodium hydroxide in a flask and added few drops of phenolphthalein. The base required 22 cm^3 of dilute hydrochloric acid for complete neutralisation.

(a) What apparatus should have been used to measure the exact volume of sodium hydroxide solution?

(b) What was the colour of the solution at the end of the neutralisation reaction?

(c) What reagent was more concentrated than the other? Explain.

(d) Name the salt that was formed.

(e) How would you obtain pure crystals of the salt resulting from the solution?

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Chapter Four

The mole concept and related calculations

Introduction

Any chemical reaction requires specific amounts of chemicals to reach its completion. The quantities of these chemicals are estimated in different units which include the mole. The mole relates the mass of a substance with its specific molar mass. The mole concept helps to determine the concentration of the substance when dissolved in solvent or solution. In this chapter, you will learn about the mole concept, molar quantities of substances, and the applications of the mole concept in chemical reactions and different settings. The competencies developed will enable you to determine quantities in different reactions and processes occurring in daily life.

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Mole concept

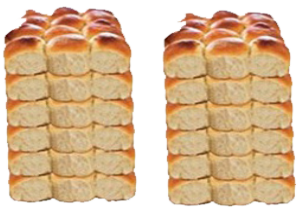
The mole is conceptually similar to other measurements such as; a *pair,* which is a set of two items, a *dozen*, which is a set of twelve items, and a *gross*, which is a set of twelve dozens with 144 items. A pair, a dozen, and a gross are units of measurements with fixed numbers of items. These measurements put the items into groups to simplify the counting. Examples of items that employ these measurements include a pair of shoes, a dozen of pencils, and a gross of scones (Figure 4.1).



A pair of shoes



A dozen of pencils



A gross of scones

Figure 4.1: *Examples of items in a pair, dozen, and gross*

In the same way, substances are made up of millions of particles which may be atoms, molecules, or ions. The particles cannot be counted one by one because they are too many and too small. Instead of counting the particles in a substance, chemists weigh

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the substance by using a standard unit of the substance. This standard unit is referred to as a *mole*.

A *mole* is the amount of the substance which contains as many particulate entities (atoms, molecules, or ions) as there are atoms in 12 g of carbon–12. Therefore, the mole is the basic unit of the amount of a substance, usually abbreviated as *mol.* The term mole is derived from the Latin word *moles*, which means a large mass or bulk, which is consistent with its use as the SI unit of the amount of a substance. To obtain the exact number of particles in any substance, the amount of that substance is multiplied by a constant factor which is called *Avogadro's constant* (*Avogadro's number*).

Avogadro's constant

In order to establish a quantitative relationship between one mole of a substance and the respective number of particulate entities, the number of atoms contained in 12 g of carbon–12 will be calculated. The measured mass of one atom of carbon–12 is 1.9927×10^{-23} g. The number of atoms in 12 g of carbon–12 is obtained by dividing the 12 g by the measured mass.

$$\text{That is: } \frac{12\text{ g}}{1.9927 \times 10^{-23}\text{ g}} = 6.02198 \times 10^{23} \approx 6.022 \times 10^{23} \text{ atoms}$$

This number 6.022×10^{23} is known as the Avogadro's constant or Avogadro's number expressed by the symbol, N_A. The constant was named after Amedeo Avogadro, an Italian physicist (1776–1856) who was honoured for his work on gases. One mole of any substance therefore contains 6.022×10^{23} particulate entities, that is, atoms, molecules or ions, just like there are two items in every pair or 12 items in every dozen. Since 6.022×10^{23} is the number of particulate entities composing a mole, the constant is properly reported with an explicit unit of *per mole*; that is, 6.022×10^{23} (atoms, molecules or ions)/mol. The Avogadro's number makes it easy to convert the number of particulate entities present in a sample to the amount of substance in moles by using the following relation:

$$\text{Amount of substance in moles (n)} = \frac{\text{number of particulate entities (N)}}{\text{number of particulates per mole } (N_A)}$$

Number of particulate entities (N), can be expressed as:

$$N = n \times N_A$$

where; n = the amount of substance in moles,

N_A = number of particulate entities per mole.

Table 4.1 shows basic units of measurements including the mole that are commonly used in chemistry from which all the other units are derived.

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Table 4.1: *Basic units of measurements*

Physical quantity	Symbol	SI unit	SI unit symbol
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Temperature	T	kelvin	K
Amount of substance	n	mole	mol
Luminous intensity	I_v	candela	cd

The physical quantities indicated in Table 4.1 are applied in different areas. Some apparatus and instruments that are used to measure the physical quantities include analytical balance, stopwatch, and thermometer as shown in Figure 4.2.



(a) Analytical balance



(b) Stopwatch



(c) Thermometer

Figure 4.2: *Examples of instruments used to measure physical quantities*

Example 4.1

How many moles of nitrogen gas are there in 2.2×10^{12} molecules of the gas?

Solution

1 mole of nitrogen gas contains 6.022×10^{23} molecules. Therefore, the number of moles of N_2 in 2.2×10^{12} molecules will be:

$$n = \frac{\text{number of particulate entities (N)}}{\text{number of particulates per mole } (N_A)}$$

$$= \frac{2.2 \times 10^{12}\ \text{molecules}}{6.022 \times 10^{23}\ \text{molecules mol}^{-1}}$$

$$= 3.65 \times 10^{-12}\ \text{mol}$$

The amount of nitrogen gas present in 2.2×10^{12} molecules of the gas = 3.65×10^{-12} mol.

Molar quantities of substances

It is often convenient to express extensive quantities such as volume and masses as the actual values divided by the amounts of the substances. The resulting quantities are called molar quantities (molar mass and molar volume).

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Molar mass

The *molar mass* of any substance is the mass of one mole of that substance. Different substances have different molar masses because the number of particles present in an atom differ from one substance to another. The molar mass of an element is expressed in gram per mole (g mol^{-1} or g/mol). It is quantitatively similar to a dimensionless physical quantity known as *Relative Atomic Mass* (RAM). For example, the molar mass of sodium is 23 g/mol, because the RAM of sodium is 23. Similarly, the molar mass of aluminium is 27 g/mol because the RAM of aluminium is 27.

The molar mass of a molecular substance is its *Relative Molecular Mass* (RMM). Consider one mole of calcium sulphate ($CaSO_4$). The relative atomic masses of the elements making up $CaSO_4$ are:

Calcium = 40, sulphur = 32, and oxygen = 16.

Therefore, the RMM of calcium sulphate is given by the sum of their atomic masses multiplied by the number of atoms as follows:

(1 × RAM (Ca)) + (1 × RAM (S)) + (4 × RAM (O)) = (1 × 40) + (1 × 32) + (4 × 16) = 136.

The molar mass of calcium sulphate is therefore 136 g/mol.

Similarly, the RMM of one mole of carbon dioxide (CO_2) is equal to (1 × RAM (C)) + (2 × RAM (O)) = (1 × 12) + (2 × 16) = 44. Therefore, the molar mass of carbon dioxide is 44 g/mol.

The relationship between the number of moles and the relative atomic mass or relative molecular mass is useful in calculating the number of moles present in a sample of a substance whose mass is known. The following formula is used:

$$n = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

From the given relationship, it is possible to calculate the mass of a substance required to make a certain number of moles of the substance.

Example 4.2

Determine the molar masses of the following compounds:

(a) Magnesium nitrate ($Mg(NO_3)_2$)

(b) Hydrated copper(II) sulphate ($CuSO_4 \cdot 5H_2O$) (RAM of Mg = 24, Cu = 63.5, S = 32, N = 14, O = 16, H = 1)

Solution

(a) Magnesium nitrate, ($Mg(NO_3)_2$)
The relative atomic masses of the elements in magnesium nitrate are:
Mg = 24 × 1 = 24
N = 14 × 2 = 28
O = 16 × 3 × 2 = 96
The RMM of $Mg(NO_3)_2$ is
24 + 28 + 96 = 148
The molar mass of $Mg(NO_3)_2$ = 148 g/mol.

(b) Hydrated copper(II) sulphate ($CuSO_4 \cdot 5H_2O$)
The relative atomic masses of the elements are:

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Cu = 63.5 × 1 = 63.5
S = 32 × 1 = 32
O = 16 × 4 = 64
H_2O = (2 + 16) × 5 = 90

The RMM of hydrated copper(II) sulphate is 63.5 + 32 + 64 + 90 = 249.5

The molar mass of hydrated copper(II) sulphate = 249.5 g/mol.

Example 4.3

The relative atomic mass of magnesium is 24:

(a) How many moles are there in 35.8 g of magnesium?

(b) What is the mass of 3.58 moles of magnesium?

Solution

(a) To calculate the number of moles in 35.8 g of magnesium.

One mole of magnesium has a mass of 24 g. The number of moles in 35.8 g of magnesium is:

$$n = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1})}$$

$$= \frac{35.8 \text{ g}}{24 \text{ g mol}^{-1}}$$

$$n = 1.492 \text{ mol}$$

(b) To calculate the mass of 3.58 moles of magnesium.

1 mole of magnesium = 24 g

3.58 moles of magnesium =? g.

Therefore, the mass of 3.58 moles is:

Mass = n × molar mass (g mol^{-1})

= 3.58 mol × 24 g mol^{-1}

= 85.92 g

Example 4.4

How many Cu atoms are there in a pure copper plate weighing 48 g?

(Cu = 63.5 g mol^{-1}).

Solution

$$N = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1})} \times N_A$$

where N_A = Avogadro's number

N = Number of particles

Since 1 mole (63.5 g) of copper has 6.022×10^{23} atoms,

Number of atoms in 48 g of copper plate

$$= \frac{48 \text{ g} \times 6.022 \times 10^{23} \text{ atoms mol}^{-1}}{63.5 \text{ g mol}^{-1}}$$

$$= 4.55 \times 10^{23} \text{ atoms}$$

Example 4.5

How many atoms are there in 14.48 g of calcium? (Ca = 40)

Solution

40 g of calcium are equivalent to 1 mole.

14.48 g are equivalent to *x* moles; where, *x* is the number of moles in 14.48 g of calcium.

$$x = \frac{14.48 \text{ g} \times 1 \text{ mol}}{40 \text{ g}}$$

$$= 0.362 \text{ mol}$$

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But, one mole of calcium contains 6.022×10^{23} atoms.

Therefore, the number of atoms in 0.362 mol is:

$$N = n \times N_A$$
$$= 0.362 \text{ mol} \times 6.022 \times 10^{23} \text{ atoms mol}^{-1}$$
$$= 2.18 \times 10^{23} \text{ atoms.}$$

Alternatively:

$$N = n \times N_A$$

But, $$n = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1})}$$

Therefore, $$N = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1})} \times N_A$$

$$N = \frac{14.48 \text{ g}}{40 \text{ g mol}^{-1}} \times 6.022 \times 10^{23} \text{ atoms mol}^{-1}$$
$$= 2.18 \times 10^{23} \text{ atoms}$$

Example 4.6

How many atoms are there in 10.6 g of sodium carbonate (Na_2CO_3)?
(Na = 23, C = 12, O = 16)

Solution

The molar mass of Na_2CO_3 is $(23 \times 2) + (12 \times 1) + (16 \times 3) = 106$ g/mol

The number of atoms (N) in 10.6 g is given by $n \times N_A$

But, $$n = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1})}$$

Therefore,

$$N = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1})} \times N_A$$

$$N = \frac{10.6 \text{ g}}{106 \text{ g mol}^{-1}} \times 6.022 \times 10^{23} \text{ atoms mol}^{-1}$$
$$= 6.022 \times 10^{22} \text{ atoms}$$

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Example 4.7

Equation 4.1 shows the dissociation of sulphuric acid into its ions.

$$H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq) \qquad (4.1)$$

From the equation, how many ions are there in 9.8 g of sulphuric acid?

Solution

Molar mass of H_2SO_4

$$= (2 \times 1) + (32 \times 1) + (16 \times 4) = 98 \text{ g/mol.}$$

The number of moles (n) in 9.8 g of H_2SO_4 is:

$$n = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1})}$$

$$= \frac{9.8 \text{ g}}{98 \text{ g mol}^{-1}}$$

$$= 0.1 \text{ mol}$$

From Equation 4.1, one mole of sulphuric acid contains two moles of hydrogen ions and one mole of sulphate ions. Therefore, there are three moles of ions in a molecule (one mole) of sulphuric acid.

But, 9.8 g of H_2SO_4 contain 0.1 mol. Therefore, the total number of moles of ions in 0.1 mol of sulphuric acid is 0.1 mol × 3 = 0.3 mol.

$$N = n \times N_A$$
$$= 0.3 \times 6.022 \times 10^{23} \text{ ions}$$
$$= 1.8066 \times 10^{23} \text{ ions}$$
$$= 1.81 \times 10^{23} \text{ ions}$$

Activity 4.1



Aim: To determine the relationship between molar masses of different substances

Requirements: Watch glass, weighing balance, carbon, iron, copper, calcium carbonate, and copper carbonate

Procedure

1. Work out on the relative atomic masses or relative molecular masses of Fe, C, $CaCO_3$, and $CuCO_3$ using the Periodic Table.
2. Calculate the masses that will give one mole of the given substances.
3. Weigh one mole of Fe, C, $CaCO_3$, and $CuCO_3$ on separate watch glasses.

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Questions

1. Which among the given substances is the heaviest?
2. Which substance is the lightest?
3. Compare and comment on the mass of one mole of each measured substance.

Exercise 4.1

1. Calculate the number of particles in each of the following amounts of substances:
   (a) 0.29 mol of iron
   (b) 0.067 mol of hydrochloric acid
   (c) 19 mol of copper(II) sulphate
   (d) 55 mol of potassium nitrate
2. How many moles are there in each of the following particulate entities?
   (a) 214 atoms (b) 7.5×10^{50} molecules (c) 2 ions (d) 0.0037 atoms
3. (a) Complete Table 4.2 by filling the relative atomic masses or relative molecular masses of the substances provided.

Table 4.2: *Relative atomic masses and relative molecular masses*

Substance	Chemical formula/ symbol	Relative atomic/molecular mass
Copper	Cu	
Carbon	C	
Iron	Fe	
Copper(II) oxide	CuO	
Sodium hydroxide	NaOH	
Calcium hydrogencarbonate	$Ca(HCO_3)_2$	
Nitric acid	HNO_3	
Iron(III) chloride	$FeCl_3$	
Potassium sulphate	K_2SO_4	

   (b) How many ions of lead and nitrate are there in 0.5 mol of lead(II) nitrate?
4. How many moles are there in each of the following substances?
   (a) 16 g of calcium
   (b) 54 g of potassium nitrate (KNO_3)
   (c) 2.115 kg of zinc carbonate ($ZnCO_3$)
   (Zn = 65.4, Ca = 40, K = 39, N = 14, O = 16, C = 12)

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5. Calculate the mass of each of the following substances:
   (a) 4 moles of iron.
   (b) 0.2 moles of hydrochloric acid.
   (c) 0.035 moles of anhydrous copper(II) sulphate.
   (d) 0.001 moles of hydrated iron(III) chloride ($FeCl_3 \cdot 6H_2O$)
   (Fe = 56, Cu = 63.5, S = 32, Cl = 35.5, O = 16, H = 1)

Molar volume

Molar volume is the volume of one mole of a substance. It is related to the density of the substance and varies with temperature and pressure. Hence, molar volume is expressed as:

$$\text{Molar volume} = \frac{\text{molar mass}}{\text{density}}$$

This relationship may be deduced into the number of moles as:

$$\text{Number of moles} = \frac{\text{volume}}{\text{molar volume}}$$

In case of gases, the volumes at standard temperature and pressure (STP) are used. For this purpose, 0 °C or 273 K temperature is taken as the standard temperature and 1 atmosphere (atm) or 1 bar or 760 mmHg or 101325 pascal (Pa) pressure is taken as the standard pressure.

One mole of any gas occupies 22.4 dm^3 at standard temperature and pressure (STP). This is according to the Avogadro's Law which states that, *equal volumes of all gases measured at the same temperature and pressure, contain the same number of particles* (*molecules or ions*). The 22.4 dm^3 is known as the *gram molecular volume* (GMV) of a gas.

For example:

One mole of oxygen gas occupies 22.4 dm^3 or 22.4 litres at STP.

One mole of hydrogen gas occupies 22.4 dm^3 or 22.4 litres at STP.

One mole of chlorine gas occupies 22.4 dm^3 or 22.4 litres at STP.

Therefore, the number of moles of a gas can be calculated by using the following equation:

$$n = \frac{\text{volume of gas at STP (litres)}}{\text{GMV}}$$

Example 4.8

How many moles of nitrogen gas are there in 5.2 dm^3 of the gas at STP?

Solution

One mole of nitrogen gas occupies 22.4 dm^3 at STP. The number of moles of nitrogen gas in 5.2 dm^3 of the gas at STP are calculated using the following formula:

$$n = \frac{\text{volume of the gas at STP}}{\text{GMV}}$$

$$n = \frac{5.2\ dm^3}{22.4\ dm^3\ mol^{-1}}$$

$$= 0.232\ mol$$

The number of moles of nitrogen gas = 0.232 mol

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Example 4.9

What is the volume of 8.0 moles of oxygen gas (O_2) at STP?

Solution

One mole of oxygen gas occupies 22.4 dm^3 at STP.

Therefore, 8.0 moles occupy x dm^3.

$$n = \frac{\text{volume of the gas at STP}}{\text{GMV}}$$

Volume of the gas at STP

$= n \times GMV$

$= 8 \text{ moles} \times 22.4 \text{ dm}^3 \text{ mol}^{-1}$

$x = 179.2 \text{ dm}^3$

Example 4.10

How many hydrogen molecules are there in 1.12 dm^3 of hydrogen gas at STP? (H = 1)

Solution

Volume of hydrogen gas at STP

$= 1.12 \text{ dm}^3$

From $N = nN_A$

$$n = \frac{\text{volume of the gas at STP}}{\text{GMV}}$$

Hence;

$$N = \frac{\text{volume of the gas at STP}}{\text{GMV}} \times N_A$$

$$N = \frac{1.12 \text{ dm}^3}{22.4 \text{ dm}^3 \text{ mol}^{-1}} \times 6.022 \times 10^{23} \text{ molecules}$$

$N = 3.011 \times 10^{22}$ molecules

Task

(a) Determine the volume that would be occupied by the following number of moles of a gas at STP: 0.00964, 0.151, and 0.0772.

(b) Construct the cubes to represent the volume that would be occupied by the gas in (a) by using construction papers, markers, metric ruler, tape or glue and a pair of scissors.

(c) Comment on the results obtained in (a) in relation to the sizes of cubes constructed.

Preparation of molar solutions

Solutions are formed by dissolving a solute in a solvent. An aqueous solution is formed when water is used as a solvent. A *molar solution* is a solution which contains one mole of the solute dissolved to make one litre (1 dm^3) of the solution. For example, the molar mass of sodium hydroxide is 40 g/mol. When 40 g of sodium hydroxide are dissolved to make one litre of solution, the resulting solution is one molar sodium hydroxide solution. Similarly, one molar sodium carbonate solution contains 106 g of sodium carbonate dissolved to make one litre of a solution.

Concentration

The concentration of a solution is the quantity of the solute present in a given quantity of solution. Concentration can be determined by measuring the mass of a solute that is dissolved per unit volume of the solution.

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$$\text{Mass concentration} = \frac{\text{mass}}{\text{unit volume of the solution}}$$

Concentration can also be determined by measuring the amount of a substance (number of moles) that is dissolved per unit volume of the solution. This concentration is known as the *molar concentration* or *molarity*. The molar concentration or molarity is the amount of a substance dissolved in a litre of the solution.

$$\text{Molarity or Molar concentration} = \frac{\text{number of moles of the solute (mol)}}{\text{unit volume of the solution (dm}^3\text{)}}$$

$$\text{Molarity} = \frac{n}{V}$$

This formula can be further linked with mass concentration by introducing mass of the substance and molar mass.

$$\text{Number of moles} = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

$$\text{Thus, molarity} = \frac{\left(\frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1}\text{)}}\right)}{\text{unit volume of solution (dm}^3\text{)}}$$

$$\text{Molarity} = \frac{\text{concentration (g dm}^{-3}\text{)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

The unit of molarity is mol/dm^3 or mol/L expressed as M.

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Example 4.11

Calculate the molarity of each of the following solutions:

(a) 4.0 g of sodium hydroxide (NaOH) in 250 cm^3 of a solution.

(b) 2 moles of sodium carbonate (Na_2CO_3) in 5 litres of a solution.

Solution

(a) 4.0 g of sodium hydroxide (NaOH) in 250 cm^3 of a solution:

Mass of NaOH in the solution = 4.0 g

Volume of the solution = 250 cm^3 equivalent to 0.25 dm^3

Molar mass of NaOH = 40 g mol^{-1}; calculated from the formula.

$$\text{Concentration (g dm}^{-3}\text{)} = \frac{\text{mass (g)}}{\text{volume (dm}^3\text{)}}$$

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$$= \frac{4.0\ \text{g}}{0.25\ \text{dm}^3}$$

$$= 16\ \text{g dm}^{-3}$$

$$\text{Molarity} = \frac{\text{concentration (g dm}^3\text{)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

$$= \frac{16\ \text{g dm}^{-3}}{40\ \text{g mol}^{-1}}$$

Molarity = 0.4 M

Alternatively,

$$\text{Molarity} = \frac{\text{number of moles}}{\text{volume in dm}^3}$$

$$\text{Moles (n)} = \frac{\text{mass of NaOH}}{\text{molar mass of NaOH}} = \frac{4.0\ \text{g}}{40\ \text{g mol}^{-1}}$$

n = 0.1 mol

V = 0.25 dm^3

$$\text{Molarity} = \frac{0.1\ \text{mol}}{0.25\ \text{dm}^3} = 0.4\ \text{M}$$

(b) 2 moles of sodium carbonate (Na_2CO_3) in 5 litres of a solution.

Number of moles of Na_2CO_3 = 2 moles

Volume = 5 litres

$$\text{Molarity} = \frac{\text{moles}}{\text{volume (dm}^3\text{)}}$$

$$= \frac{2\ \text{moles}}{5\ \text{litres}}$$

= 0.4 mol/L or 0.4 M.

Activity 4.2



Aim: To prepare 1 M of sodium hydroxide

Requirements: Analytical balance, 1000-mL volumetric flask, two 250-mL, beakers, funnel, glass rod, dropper, sodium hydroxide pellets, labels, marker pen, and distilled water.

Procedure

1. Weigh 40 g of sodium hydroxide pellets and put the pellets in a beaker.
2. Add 100 cm^3 of distilled water into the beaker containing the sodium hydroxide pellets and then stir the mixture by using a glass rod.
3. Transfer the solution into a 1000-mL volumetric flask.
4. Rinse the beaker containing sodium hydroxide by using distilled water and transfer the contents into the volumetric flask. Repeat the process four times or until there is no more solid particles of sodium hydroxide in the beaker.
5. Rinse the glass rod well and pour its content into the volumetric flask.

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6. Fill the volumetric flask containing the solution with distilled water to the mark to make one litre of solution.
7. Label the prepared solution as 1 M sodium hydroxide.

Precaution: When filling the flask, as the solution approaches the mark, use a dropper to add the distilled water until the meniscus is on the mark.

Questions

1. Why did you start dissolving the pellets with a small volume of distilled water instead of the entire volume?
2. Why did you weigh 40 g of sodium hydroxide pellets in preparing 1 M of sodium hydroxide?
3. Is it possible to prepare the same concentration using any other mass and volume? Explain with examples.

Activity 4.3



Aim: Preparing 1 M acetic acid solution

Requirements: 1000-mL volumetric flasks with their corks, two 250-mL beakers, 100-mL measuring cylinder, funnel, glass rod, acetic acid, distilled water, label, and marker pen

Procedure

1. Use a measuring cylinder to transfer a volume of 57.24 cm^3 of acetic acid into a 250-mL beaker containing distilled water.
   Note: The volume of the acid will depend on the density of stock acid solution.
2. Transfer the resulting solution into a 1000-mL volumetric flask.
3. Rinse the measuring cylinder used for measuring the solution and transfer its solution into the volumetric flask.
4. Swirl the flask gently to mix the solution properly.
5. Continue adding distilled water to the solution while swirling the flask gently.
6. Add distilled water until it is at the meniscus of the graduation mark.
7. Close the flask with a cork and invert it several times to mix the solution.
8. Label the prepared solution as 1 M acetic acid.

Questions

1. What is the number of moles in 57.24 cm^3 of acetic acid? Assume the molarity of stock acetic acid is 17.4 M.
2. Explain the significance of Step 3 of the procedure.
3. Do the concentration and number of moles of acetic acid change in Step 5? Explain.

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Dilution Law

Dilution is a process through which the concentration of a solution is lowered by adding more solvent (distilled water). In dilution, only the volume of the solvent is increased, while the number of moles of the solute in the solution remains the same.

Number of moles before dilution = Number of moles after dilution

$$n_{before} = n_{after}$$

If a dilute solution with concentration (molarity) M_2 is to be prepared from a concentrated solution with concentration M_1, the volumes of the two solutions are determined by equating the number of moles of the solutions before and after dilution as shown in the following formula:

$$M_1V_1 = M_2V_2$$

where V_1 is the volume of the concentrated solution (before dilution) and V_2 is the volume of the dilute solution.

A concentrated solution from which dilute solutions are prepared is known as a *stock solution*. The number of times the volume of the concentrated solution is diluted to obtain the less concentrated solution(s) is referred to as the *dilution factor* and is obtained by the following formula:

$$\text{Dilution factor} = \frac{V_2}{V_1}$$

$$\text{or} = \frac{M_1}{M_2}$$

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Example 4.12

(a) What volume of 10 M acetic acid is required to prepare 1.0 dm^3 of 0.50 M acetic acid?

(b) Use the information obtained in part (a) to find the dilution factor.

Solution

Data: $M_1 = 10$ M, $M_2 = 0.5$ M, $V_1 = ?$ and $V_2 = 1$ dm^3

$$M_1V_1 = M_2V_2$$

(a) Volume of acetic acid required:

$$V_1 = \frac{M_2V_2}{M_1}$$

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$$= \frac{0.5\ \text{M} \times 1\ \text{dm}^3}{10\ \text{M}}$$

$$= 0.05\ \text{dm}^3 \text{ or } 50\ \text{cm}^3$$

(b) Required dilution factor:

$$\text{Dilution factor} = \frac{M_1}{M_2}$$

$$= \frac{10\ \text{M}}{0.50\ \text{M}}$$

$$= 20$$

The dilution factor is 20; this means that the concentrated acetic acid should be diluted 20 times.

Applications of the mole concept

The mole concept is widely applied in chemistry. One of the most important application is in *stoichiometry*. Other applications of the mole concept include determination of empirical formulas, molecular formulas, molar masses, masses, concentration (molarity), and volumes of chemical substances.

Stoichiometry

Stoichiometry is a relationship between the relative quantities of substances (reactants) taking part in a chemical reaction and the products that are formed. This relationship is shown by using balanced chemical equations. Based on the balanced chemical equation, one can calculate how much of each reactant is required to produce a certain amount of product(s). Consider the following reaction between nitrogen and hydrogen to produce ammonia:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad (4.2)$$

From chemical Equation 4.2, 1 molecule of nitrogen gas requires 3 molecules of hydrogen gas to produce 2 molecules of ammonia gas. In other words, 1 mole of nitrogen gas requires 3 moles of hydrogen gas to produce 2 moles of ammonia gas.

According to the Avogadro's Law, equal volumes of gases measured at the same temperature and pressure contain the same number of particles (molecules). Based on Equation 4.2, 1 volume of nitrogen gas reacts with 3 volumes of hydrogen gas to produce 2 volumes of ammonia gas. The ratio of the gases in the reaction mixture is 1:3:2 for N_2, H_2, and NH_3, respectively.

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In any combination of the two reactants (N_2 and H_2) react to yield NH_3 (g);

The ratio of N_2: H_2 is 1: 3

The ratio of N_2: NH_3 is 1: 2

The ratio of H_2: NH_3 is 3: 2

In a reversible reaction, 2 molecules of ammonia produce 3 molecules of hydrogen and 1 molecule of nitrogen. The ratio of 1:3:2 is maintained. The Avogadro's concept can also be demonstrated by considering chemical Equation 4.3:

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g) \qquad (4.3)$$

1 volume	1 volume	2 volumes
1 molecule	1 molecule	2 molecules
1 mol	1 mol	2 mol

Example 4.13

How many moles of nitrogen gas would react with excess hydrogen gas to produce 2.24 dm^3 of ammonia gas at STP?

Solution

Write the balanced chemical equation for the reaction:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad (4.4)$$

Then convert 2.24 dm^3 of ammonia gas into moles.

1 mole of ammonia gas occupies 22.4 dm^3 at STP.

But,

$$n = \frac{\text{volume of the gas at STP}}{\text{GMV}}$$

$$\text{Number of moles of ammonia, } n = \frac{2.24\ dm^3}{22.4\ dm^3\ mol^{-1}}$$

$$= 0.1\ mol$$

The number of moles in 2.24 dm^3 of ammonia = 0.1 mol

Chemical Equation 4.4 shows that the ratio of N_2 to NH_3 is 1:2. Therefore, half the amount of ammonia is required to react with hydrogen gas. Thus the number of moles (n) of nitrogen will be:

$$= \frac{0.1\ mol}{2}$$

$$= 0.05\ mol$$

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The number of moles of nitrogen gas required = 0.05 mol

Alternatively;

1 mole of N_2 produces 2×22.4 dm^3 of NH_3 at STP

x moles of N_2 produce 2.24 dm^3 of NH_3 at STP

$$x = \frac{1\,\text{mole} \times 2.24\,\text{dm}^3}{2 \times 22.4\,\text{dm}^3}$$

$$x = 0.05\ \text{mol}$$

The number of moles of nitrogen gas required = 0.05 mol.

Example 4.14

Calculate the volume of carbon dioxide gas that will be produced at STP when a 10 g sample of calcium carbonate is completely decomposed on heating.

Solution

Write the balanced chemical equation as shown in chemical Equation 4.5.

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g) \qquad (4.5)$$

Chemical Equation 4.5, shows that 1 mole of $CaCO_3$ decomposes to give 1 mole of CaO and 1 mole of CO_2.

The molar mass of $CaCO_3$ = (40 + 12 + (16×3) = 100 g/mol

At STP, 1 mole of any gas occupies 22.4 dm^3.

100 g of $CaCO_3$ gives 1 mole of CO_2 which is equal to 22.4 dm^3 at STP.

Therefore, 10 g of $CaCO_3$ will give x dm^3

$$x = \frac{10\ \text{g} \times 22.4\,\text{dm}^3}{100\ \text{g}}$$

$$= 2.24\ \text{dm}^3 \text{ of } CO_2 \text{ at STP}$$

The volume of carbon dioxide gas that will be produced = 2.24 dm^3.

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Exercise 4.2

1. Calculate the number of moles in each of the following volumes of gases at STP:
   (a) 12 dm^3 of oxygen. (b) 0.006 L of chlorine.
   (c) 7500 L of hydrogen. (d) 0.0027 dm^3 of nitrogen.
2. Calculate the volume occupied by each of the following gases at STP:
   (a) 0.00016 mol of oxygen. (b) 17.5 mol of hydrogen.
   (c) 0.28 mol of nitrogen. (d) 21 mol of carbon dioxide.
3. Calculate the molarity of the resulting solution when 4.8 g of sodium chloride are dissolved to make 950 cm^3 of a solution.
4. What mass of sodium hydroxide is required to make 10 litres of 0.002 M sodium hydroxide solution?
5. (a) Calculate the concentration of a solution which contains 160 g of sodium hydroxide (NaOH) in 2 dm^3 of the solution.
   (b) Calculate the number of moles of sodium hydroxide in 500 cm^3 of the sodium hydroxide solution in (a).
   (c) Calculate the volume of this solution which will contain 0.1 moles of sodium hydroxide.

   Hint: Use the Periodic Table to obtain the relative atomic masses of different elements.
6. Nitric acid is available as a stock solution with a concentration of 9.8 M. What volume can be used to prepare 100 cm^3 of 1 M nitric acid solution?
7. Sodium sulphate reacts with barium nitrate as represented in chemical Equation 4.6.

   $$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow 2NaNO_3(aq) + BaSO_4(s) \quad (4.6)$$

   Calculate the mass of barium sulphate produced when 3.56 moles of sodium sulphate react completely with barium nitrate. (N = 14, O = 16, Na = 23, S = 32, Ba = 137)
8. Chemical Equation 4.7 represents the reaction between heated copper(II) oxide and ammonia gas:

   $$3CuO(s) + 2NH_3(g) \longrightarrow 3Cu(s) + N_2(g) + 3H_2O(g) \quad (4.7)$$

   If 20 g of copper(II) oxide reacted with ammonia gas at STP, calculate:
   (a) the number of moles of copper(II) oxide used.
   (b) the number of moles of ammonia gas used.
   (c) the volume of ammonia gas used.

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(d) the mass of copper formed.

(Cu = 63.5, O = 16, N = 14, H = 1, GMV = 22.4 dm^3)

9. Lead nitrate decomposes on heating as indicated in Equation 4.8.

$$2Pb(NO_3)_2(s) \longrightarrow 2PbO(s) + 4NO_2(g) + O_2(g) \qquad (4.8)$$

If a volume of 112 cm^3 of oxygen gas was collected at STP when a sample of lead nitrate was completely decomposed by heating, calculate the;

(a) mass of the lead nitrate sample.

(b) mass of lead(II) oxide produced.

(c) Volume of nitrogen dioxide gas produced at STP.

(Pb = 207, N = 14, O =16; molar volume of gas at STP = 22.4 dm^3)

Chapter summary

1. The mole is the amount of substances of a specified elementary entity, which may be an atom, molecule, ion, electron, any other particle or specified group of such particle. It is derived from mass measurements in relation to their specific molar masses.
2. A mole of any substance contains an Avogadro's number of particles which is abbreviated as N_A.
3. The molar mass of a substance is the mass of one mole of the substance.
4. The molar volume of a gas is the volume occupied by one mole of the gas at standard temperature and pressure (STP). The molar volume of the gas at STP is 22.4 dm^3 (22.4 L).
5. A molar solution is prepared by dissolving one mole of a substance in one litre of a solution.
6. The molarity of a solution is the number of moles of solute dissolved to make one litre of the solution.
7. The process through which a concentrated solution is lowered by adding a solvent is called dilution.
8. The dilution law is the quantitative relationship showing the number of moles before dilution being equal to the number of moles after dilution.
9. A balanced chemical equation gives the relative quantities of the substances that are involved in a particular reaction.
10. The relationship between the quantities of substances that are used and produced in a chemical reaction is called stoichiometry.

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11. The mole concept is widely applied in chemistry including stoichiometry, determination of empirical formulas, molecular formulas, molar masses, concentration (molarity), masses, and volumes of chemical substances.

Revision exercise 4

Useful information: Use the Periodic Table to obtain the relative atomic masses of elements.

Choose the correct answer in items 1–5.

1. How many moles are produced by 0.69 g of sodium?
   (a) 0.13 (b) 3 (c) 0.3 (d) 0.03

2. How many molecules are there in 16.8 litres of chlorine gas at STP?
   (a) 4.517×10^{23} (b) 7.5×10^{23} (c) 8.029×10^{23} (d) 9.033×10^{23}

3. What is the molarity of a solution containing 5.3 g of sodium carbonate in 100 mL of the solution?
   (a) 0.50 M (b) 0.05 M (c) 0.20 M (d) 0.01 M

4. What mass of sulphuric acid (H_2SO_4) is found in 400 cm^3 of its 0.1 M aqueous solution?
   (a) 2.45 g (b) 9.80 g (c) 4.90 g (d) 3.92 g

5. What volume of hydrogen gas is produced when 0.65 g of zinc reacts completely with excess hydrochloric acid at STP? (Zn = 65, Cl = 35, H = 1)
   (a) 224 cm^3 (b) 22.4 cm^3 (c) 22.4 dm^3 (d) 2.24 dm^3

6. Describe each of the following terms:
   (a) Mole (b) Molarity (c) Solute (d) Molar mass
   (e) Molar solution (f) Concentration

7. How many moles are there in 4.2×10^{12} molecules of nitrogen gas?

8. Calculate the number of moles in 20 g of sodium hydroxide.

9. A gas occupies 1.95 dm^3 at STP. Calculate the;
   (a) number of moles in the gas.
   (b) number of molecules in the gas.

10. What is the relative molecular mass of a hypothetical gas Y if the mass of 560 cm^3 of it measured at STP is 1.10 g?

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11. What is the molarity of a solution containing 80 g of potassium hydroxide in 2 litres?

12. Calculate the volume of oxygen gas produced at STP when 12.25 g of potassium chlorate are decomposed according to the following equation:

$$2KClO_3(s) \xrightarrow{\text{heat}} 2KCl(s) + 3O_2(g)$$

13. Calcium hydroxide reacts with ammonium chloride as indicated in the following equation:

$$Ca(OH)_2(aq) + 2NH_4Cl(aq) \longrightarrow CaCl_2(aq) + 2NH_3(g) + 2H_2O(l)$$

Calculate the mass of ammonium chloride (NH_4Cl) that will react completely with 14.8 g of calcium hydroxide.

14. Sodium sulphate reacts with barium nitrate as indicated in the following equation:

$$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow BaSO_4(s) + 2NaNO_3(aq)$$

How many moles of barium sulphate are produced when 80.5 g of sodium sulphate react completely with barium nitrate?

15. Why are the stoichiometric calculations so important in industrial production of chemical substances?

16. The reduction of copper(II) oxide can be done by using hydrogen gas.
    (a) Write a balanced chemical equation for the process.
    (b) Calculate the mass of copper produced if 4.93 g of the hydrogen gas are pumped in the reactor.
    (c) Calculate the volume of the hydrogen gas that is used when 4 g of copper(II) oxide are reduced at STP. (Cu = 63.5, O = 16, GMV = 22.4 dm^3).

17. The equation for the formation of iron(III) chloride is as follows:

$$2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$$

If the total mass of iron(III) chloride formed at STP is found to be 0.5 g. Calculate the volume of Cl_2 that reacted with iron. (Fe = 55.8, Cl = 35.5).

18. The decomposition of $Pb(NO_3)_2$ is as indicated in the following equation:

$$2Pb(NO_3)_2(s) \xrightarrow{\text{heat}} 2PbO(s) + 4NO_2(g) + O_2(g)$$

If 0.01 mole of lead nitrate is used, what will be the maximum volume of oxygen collected at room temperature and pressure? (GMV = 22.4 dm^3).

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19. A certain reaction required 20.0 cm^3 of NaOH solution containing 8.0 g dm^{-3} to neutralise 0.18 g of a dibasic acid. Calculate the relative molecular mass of the acid. (Na = 23.0, O = 16.0, H = 1).

20. Calculate the number of chloride ions in a 250 cm^3 of 1 M solution of calcium chloride.

21. A 25 cm^3 volume of 2 M NaOH is diluted to 175 cm^3.
    (a) Determine the new concentration.
    (b) Calculate the dilution factor.

22. Zinc metal was added to excess silver nitrate solution. The mass of silver formed was 5.40 g (Zn = 65.4, Ag = 108).
    (a) Write a balanced ionic equation for the reaction.
    (b) Calculate the mass of zinc used.

23. Calculate the percentage of phosphorus in one mole of $H_4P_2O_7$ (P = 31, O = 16, H = 1).

24. What mass of anhydrous sodium hydrogencarbonate would be required to complete the reaction with 20 cm^3 of 0.4 M sulphuric acid solution? (Na = 23, O = 16, H = 1, C = 12, S = 32)

25. When 32 tonnes of pure calcium carbide were used up in an industrial process, the following reaction that yielded ethyne (C_2H_2) occurred:

$$CaC_2(s) + 2H_2O(l) \longrightarrow C_2H_2(g) + Ca(OH)_2(aq)$$

    (a) Calculate:
        (i) the mass of water required for the complete reaction to occur.
        (ii) the mass of calcium hydroxide produced.
    (b) What volume of ethyne (C_2H_2) measured at STP would be evolved if CaC_2 was 50% pure.
    (c) How many molecules of ethyne were formed in (b).

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Chapter Five

Volumetric analysis

Introduction

Determinations of quantities of chemical substances in solutions are important in our daily life activities. To achieve these, volumetric analysis is widely adopted. For example, the preparation of medicines and soft drinks requires the applications of volumetric analysis to ensure proper amounts of substances are used. The skills which you learnt in Chapter Four about the mole and related calculations to determine the concentrations of substances will be applied in this chapter. In this chapter, you will learn the concept of volumetric analysis, standard volumetric apparatus, standard solutions, titration, volumetric analysis in determination of properties of chemical substances, challenges in volumetric analysis, and the applications of volumetric analysis. The competencies developed will enable you to prepare and establish different concentrations of substances that are used in daily life.

Concept of volumetric analysis

Volumetric analysis refers to the analytical method based on the measurement of the volume of a solution required to react completely with a definite volume of another solution. By comparing the volumes of the two solutions, the concentration of one substance can be calculated provided that the concentration of another solution is known. The volumes of the two solutions, their stoichiometric ratios in the chemical reaction, and their concentrations are involved in the calculations. The volume required to complete a particular reaction is easily identified through colour changes of the resulting products or pH indicators used during titration.

Thus, *titration* is a chemical analysis in which the unknown quantity or concentration of a chemical species is determined by adding to a known quantity or concentration of another substance. In titration, the two substances react in a definite known proportions. These proportions are known as *stoichiometric coefficients*. The stoichiometric coefficients are the numbers which represent the actual ratios of the quantities of the reactants and products during the occurrence of a particular chemical reaction. They are represented by numbers appearing before each reactant and product species in a chemical equation. The stoichiometric coefficients are also called *mole ratios*. The solution with a known concentration is called a *standard solution*.

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In titration, the exact volumes of solutions are measured by using apparatus such as burettes, pipettes, and volumetric flasks. The laboratory procedure involves measuring the volumes of solutions with known and unknown concentrations. The solution with known concentration is referred to as a *titrant,* and it is usually placed in the burette. The titrant is titrated against the solution with an unknown concentration called the *titrand* (*analyte)* which is kept in a conical flask. The completion of the reaction is called the *end point*. The end point is therefore the point where the analyte has completely reacted with the titrant. The completion of the reaction is indicated by colour change of an indicator used (See an example in Figure 5.1) or colour changes of the reacting species or products formed.

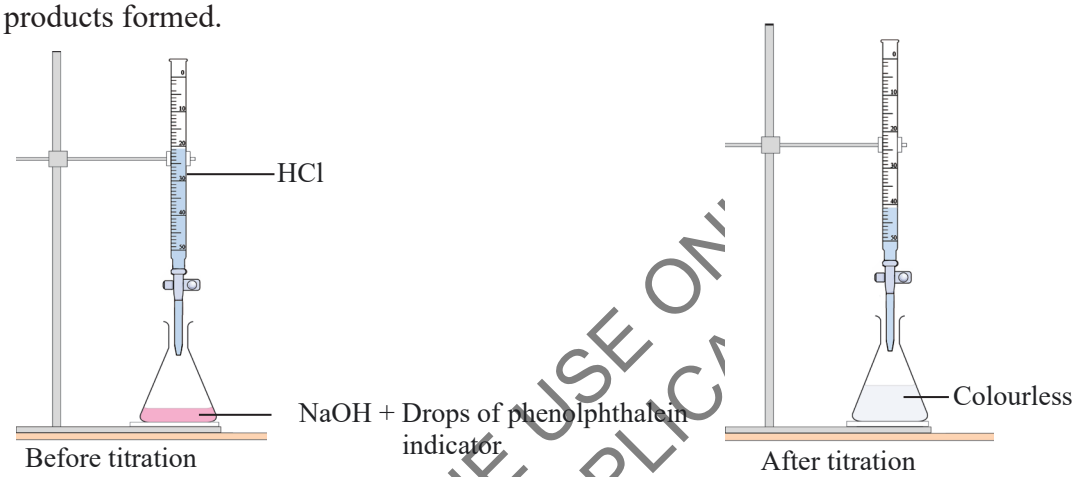


Figure 5.1: *Set-up of titration showing colour change at the end point*

Chemical reactions of the involved substances in titration are represented by using balanced chemical equations. For example, a balanced chemical equation for a titration process which involves sulphuric acid and sodium hydroxide solutions is:

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

In this balanced chemical equation, the stoichiometric coefficients are 1, 2, 1, and 2 for H_2SO_4, NaOH, Na_2SO_4, and H_2O, respectively. Thus, H_2SO_4 and NaOH react at known definite proportions of 1 to 2. These proportions will determine the reacting volumes of the acid and alkali in relation to their concentrations. If the concentration of the acid solution is known, then, by taking a fixed volume of the NaOH and adding H_2SO_4 from the burette dropwise, the end point can be identified by colour change of the pH indicator. Before the end point, the reaction mixture contains excess alkali. After the end point, the reaction mixture contains excess acid. In acid-base titration, a change in colour of the pH indicator is caused by the reaction between the pH indicator and an excess reagent when the other reagent has completely reacted.

Standard volumetric apparatus

There are various apparatus that are used in volumetric analysis. These include the apparatus used for measuring exact volumes of solutions such as burette, pipette, and

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volumetric flask. A measuring cylinder can also be used to measure approximate volumes of solutions. Exact masses of substances are measured using an analytical balance. Other apparatus such as conical flasks and beakers are used to hold solutions for titration. A funnel is used for adding liquids or powders into small openings of apparatus or containers such as burettes and bottles. A retort stand is used for supporting or holding the burette during titration. In addition, there should be a wash bottle for holding distilled water, a dropper for dispensing the pH indicator, and a white base such as a white tile or a piece of paper which is placed underneath the conical flask to aid ease spotting of the colour change at the end point. Most of these apparatus are glassware and should be handled with care. It is therefore important to learn how to use each of these apparatus before carrying out any titration reaction.

Burettes

A burette is a graduated glass tube with a tap at one end used for delivering a volume of a solution during titration. Some burettes have a rubber tube with a clip in place of the tap where liquid flow is controlled by fixing a pinch clip across the rubber tube. The most common burette has a capacity of 50 cm^3. However, burettes with different capacities of volumes such as 100 cm^3, 25 cm^3 and 10 cm^3 are available. Figure 5.2 shows examples of burettes. The smaller the volume that the burette delivers, the high the accuracy due to high precisions.

In the 50.0 cm^3 burette, the calibration marks start at 0.0 cm^3 which is near the opening and end at 50.0 cm^3 which is near the delivery point. The accuracy of the 50.0 cm^3 burette is $\pm$ 0.05 cm^3. Any volume from the burette should read in one decimal place and its accuracy is corrected at two decimal places. The second decimal place is its absolute value which is 0.05 cm^3. For example; the burette reading of exactly 21.7 cm^3 is reported in two decimal places at an accuracy value of 21.70 cm^3; and a reading of 20.8 cm^3 is 20.80 cm^3, and when the decimal measurement is uncertain, the absolute value is added. For example, if the reading scale is 30.1 but the meniscus exceeds that value and is less than 30.2 cm^3, then the absolute value is the average between the two readings, and hence it should be reported as 30.15 cm^3.

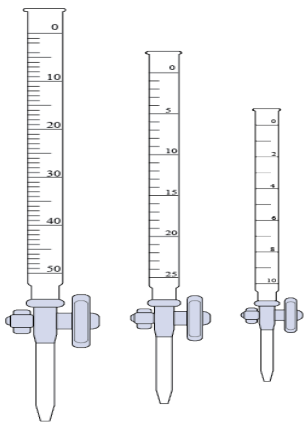


Figure 5.2: *Burettes*

Cleaning the burette

The burette that is used in titration should always be clean in order to avoid contamination that may interfere with the reaction. The process should be accomplished just before the start of a titration experiment.

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Activity 5.1

Aim: To clean a burette

Requirements: Burette, detergent, tap water, distilled water, and petroleum jelly

Procedure

1. Clean the burette by using a detergent and tap water, and rinse it by using tap water until no more foam appears in the rinsing water.
2. Fill the burette with distilled water three times, and allow the distilled water to run through the delivery tap each time.
3. Check if the burette has no leakage. If there is a leakage, apply some petroleum jelly to the leaking part to stop the leakage, or replace with a new one.
4. Finally, rinse the burette three times with 5 cm^3 to 10 cm^3 of the titration reagent and allow the reagent to run through the tap each time.

Questions

1. Why is it important to rinse the burette with the titration reagent?
2. Is cleaning of the burette before titration important? Explain.

Using the burette

It is important to know how to use a burette during titration. This is crucial for dispensing accurate volumes of the solution. During titration, correct readings are taken from the burette by viewing the meniscus level with the eye at the horizontal level. This is to eliminate parallax error. During actual titrations, the burette supported by a retort stand is placed at the centre over the conical flask so that all the titrant falls directly into the conical flask. The tip of the burette should also be dipped into the conical flask to prevent some of the titrant from being poured outside the flask. Figure 5.3 shows a simple set-up of titration apparatus and the meniscus reading.

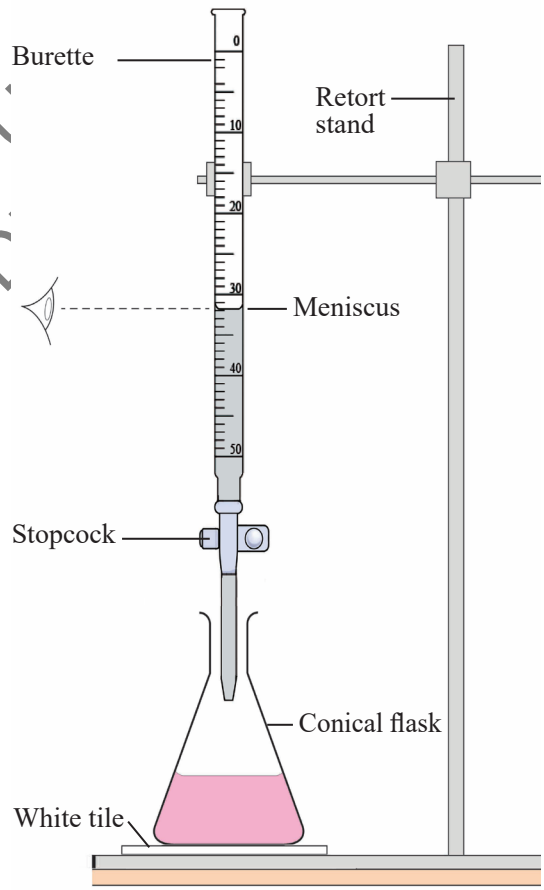


Figure 5.3: *Set-up of titration apparatus and reading the meniscus*

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Activity 5.2



Aim: To deliver a reagent by using a burette

Requirements: Retort stand, burette, conical flask, funnel, and copper(II) sulphate solution

Procedure

1. Clamp the burette in an upright position on the retort stand.
2. Clean and dry the funnel to be used in filling the copper(II) sulphate solution into the burette.
3. Fill the reagent above the zero mark by using a funnel, and then remove the funnel after pouring the titration reagent into the burette.
4. Remove the air bubbles from the tip of the burette and near the clip by running a small amount of the titration reagent from the burette.
5. Refill the burette and adjust its reading to zero or below zero. Take the initial reading.
6. Deliver the following volumes from the burette and record the final readings of the burette after every delivery in a given table:
   (a) 2.50 cm^3 (b) 2.55 cm^3
   (c) 22.55 cm^3 (d) 42.00 cm^3

Burette readings	1	2	3	4
Final burette reading (cm^3)				
Initial burette reading (cm^3)				
Volume of the reagent (cm^3)	2.50	2.55	22.55	42.00

Note:

1. It is not necessary to keep the initial reading of the burette at zero. However, the initial reading should be exactly, and not an estimated value, for example 1.00 cm^3 or 5.00 cm^3, not 1.05 cm^3 or 5.55 cm^3.
2. You can hold a black-stripped white card behind the meniscus when taking the reading to ensure the meniscus is clearly visible.
3. The volume delivered by the burette (volume of the reagent) is obtained by subtracting the initial volume from the final volume.

Question

A burette with 0.1 cm^3 calibrations cannot deliver a volume of 23.34 cm^3. Explain.

Activity 5.3



Aim: To practice how to use a burette

Requirements: Burette, funnel, beaker, retort stand, and water

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Procedure

1. Clamp the burette to a stand in a vertical position.
2. Use the funnel to fill the burette with water above the zero mark.
3. Allow water to pass the tap and fill the lower part.
4. Remove the bubble(s) if present by opening the tap to allow delivery of a small amount of water to the beaker.
5. Allow water to settle to the reading of your choice, but remember to read the meniscus.
6. Record the initial reading to two decimals.
7. Deliver the following volumes from the burette and record the initial and final readings of the burette after every delivery:

   (a) 12.50 cm^3 (b) 25.10 cm^3 (c) 35.10 cm^3 (d) 20.00 cm^3
8. Record the results in tabular form as indicated in the following table:

Burette readings	1	2	3	4
Final burette reading (cm^3)				
Initial burette reading (cm^3)				
Volume of water (cm^3)	12.50	25.10	35.10	20.00

Question

How did you manage to measure 35.10 cm^3 after measuring 25.10 cm^3? Comment on the accuracy of your measurement.

Pipettes

A *pipette* is a thin calibrated glass tube with or without a wide middle part. Pipettes are used to measure and transfer definite volumes of liquids or solutions. Common pipettes have a capacity of 20 cm^3 or 25 cm^3. Other fixed volumes which can be delivered by a pipette are 10 cm^3 and 5 cm^3. Some pipettes can deliver very small volumes of liquids. Such pipettes are known as *micropipettes*. The capacity of the pipette is usually indicated on the pipette. A mark on the upper part of the pipette indicates the level to which the liquid or solution should reach to measure the specified volume. A pipette which delivers a fixed volume of a liquid or solution is called a *volumetric pipette* or *transfer pipette*. Some pipettes have several calibrations to allow the measurements of different volumes such as 0.5 cm^3, 1 cm^3, 1.2 cm^3, and 4.1 cm^3. The calibrations start from the delivery hole and

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increase upwards. Figure 5.4 shows some different types of pipettes.



Micropipette Graduated pipette



Volumetric pipettes Digital pipette

Figure 5.4: *Different types of pipettes*

Using a pipette

Before using a pipette, it should be thoroughly cleaned in the same way as the burette. The pipette can be filled by *mouth sucking* or by using a *pipette filler*. Pipette fillers are of two types; *manual* and *electronic* pipette fillers (Figure 5.5). It is encouraged to use a pipette filler; if not available, the mouth can be used for sucking. However, make sure that the solution does not get into the mouth. Do not share a pipette with another person during mouth sucking.



Figure 5.5: *Electronic and manual pipette fillers*

Activity 5.4

Aim: To demonstrate how to use a pipette filler

Requirements: Pipette, pipette filler, conical flask, and a beaker

Procedure

1. Fix the pipette with a pipette filler as shown in Figure 5.6.

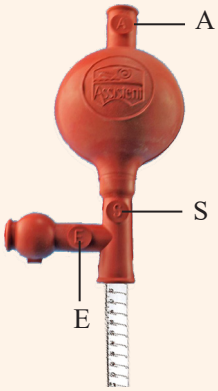


Figure 5.6: *Manual pipette filler fixed to a pipette*

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2. Press the button A while squeezing the bulb at the same time to drive out air and increase pressure in the pipette.
3. Dip the tip of the pipette into the solution.
4. Press the button E to lower the pressure in the pipette. The atmospheric pressure will force the liquid to enter into the pipette. The tip of the pipette is kept below the surface of the solution until the solution is beyond the graduation mark.
5. Adjust the level of the solution by pressing button S to allow the solution to run out of the pipette. Set the solution level to the graduation mark; align the lowest meniscus level with the graduation mark. Watch carefully on the correct level.
6. Press the button S to allow the solution to flow into a conical flask or any other container.
7. Allow the last drop of the solution in the pipette to drain out by placing the tip of the pipette against the wall of the container into which the solution is being emptied. Do not blow into the pipette to force the drop out.

Questions

1. What will happen if the bulb of the pipette filler is perforated?
2. What will you do if the solution is over-drawn into the bulb?

Filling a pipette by mouth sucking

Before sucking a solution, be sure that it is not harmful. It is not allowed to suck corrosive, poisonous or toxic substances into a pipette. Thus, the liquid should not be sucked if its safety is unknown.

Activity 5.5



Aim: To fill a pipette by mouth sucking

Requirements: Liquid solution, pipette, conical flask, and a beaker

Procedure

1. Suck a solution from the upper end of the pipette until the solution passes the graduation mark.
2. Remove the mouth immediately from the pipette opening at the top and block the opening by using an index finger as shown in Figure 5.7.

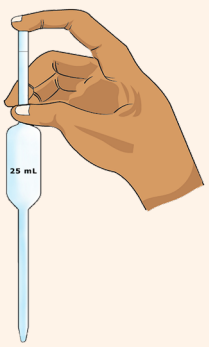


Figure 5.7: *Holding the solution in the pipette by using a finger*

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3. Adjust the level of the solution slowly by loosening the finger to let the solution run out of the pipette until the lowest meniscus level corresponds with the graduation mark.
4. Move the pipette into a conical flask or any other container into which you want to put the measured solution.
5. Remove the finger from the pipette opening to release the solution from the pipette.

Note: A small amount of the volume of the solution remaining at the bottom of the pipette should not be blown out because the required volume has already been released into the flask.

Question

The solution in the pipette cannot be emptied into the flask or beaker if the finger is still blocking the top opening. Explain.

Volumetric flasks

Volumetric flasks are used to measure actual volumes of liquids, and therefore they are used to prepare the standard solutions. Volumetric flasks are calibrated to contain a specified volume of solution. They are supplied in different capacities ranging from 5 cm^3 to 5000 cm^3. Each flask has a mark line on its neck to indicate the level to which a liquid or solution should reach as shown in Figure 5.8.

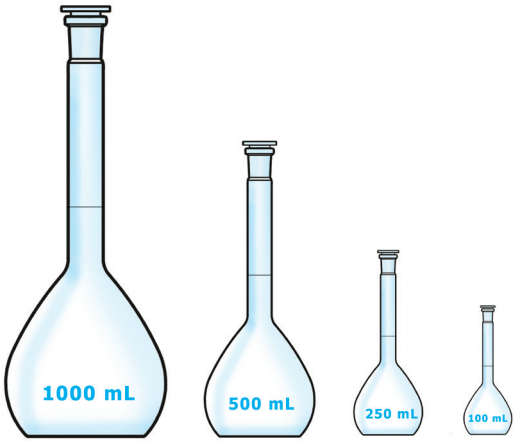


Figure 5.8: *Volumetric flasks*

Other flasks

Other flasks of different shapes and sizes are used in volumetric analysis, including conical flasks. Examples of the flasks that can handle different volumes of solutions are shown in Figure 5.9. The most commonly used one is the 250 cm^3 conical flask. It has a neck which is easily held while swirling the contents inside. It also has a flat bottom to allow it to stand on the base without extra support. All flasks should be cleaned and rinsed with distilled water before being used.

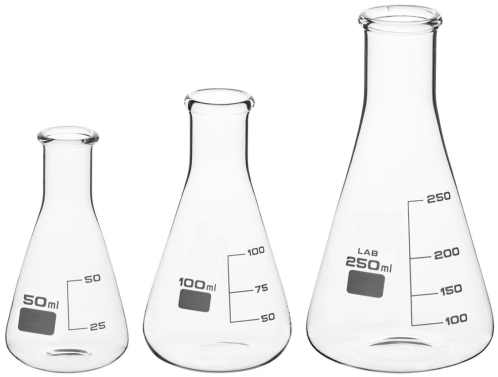


Figure 5.9: *Conical flasks*

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Funnel

A funnel is used to pour a solution into the burette without spilling it. It is made of different materials such as glass, aluminium or plastic materials. Figure 5.10 shows a funnel made up of glass.



Figure 5.10: *Funnel*

Retort stand and clamp

Retort stand and clamp (Figure 5.11) are used to hold the burette in an upright position while performing volumetric analysis experiments.



Figure 5.11: *Retort stand with a clamp*

White tile or paper

The conical flask or beaker is placed on a white tile or white paper to give a clear background for observing the colour change at the end point of the titration. Figure 5.12 shows a white tile placed on the base of the retort stand.



Figure 5.12: *White tile on a base of a retort stand*

Dropper

A dropper is an apparatus that is used to transfer an indicator into a solution, drop by drop. It is made up of glass or plastic materials (Figure 5.13). Some droppers are calibrated to allow a specific volume to be delivered.



Figure 5.13 : *Different types of droppers*

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Exercise 5.1

1. Explain the meaning of each of the following terms:
   (a) Volumetric analysis
   (b) Titration
   (c) Titrant
   (d) Titrand
2. Why is it not advisable to blow the remaining solution from the pipette?
3. Why is it necessarily to clean and rinse any volumetric apparatus with distilled water before using it?

Standard solutions

A standard solution is a solution with a precisely known concentration of a substance. The concentration of a standard solution is expressed in moles per litre (mol/L) or moles per cubic decimetre (mol/dm^3). This unit of concentration is abbreviated by a capital letter 'M' and the resulting concentration is known as the *molarity*. Thus, molarity is the concentration of a solution expressed in moles per litre of a solution.

The concentration of a standard solution is usually expressed in four decimal places, for example, 0.1000 M and 1.5000 M. In volumetric analysis, the unknown concentration of a solution is determined by using the solution of known concentration. This process is referred to as *standardisation*. The second solution, which is being standardised, is called the *titrand* or *analyte*.

Primary standard solutions

Standard solutions are usually prepared from *primary standards*. A primary standard is a substance with high purity and stability used to prepare standard solutions.

Primary standard reagents should have the following characteristics:

(a) They should have high purity to enable preparation of exact concentrations of solutions.

(b) They should be stable, that is, should not decompose with time or when exposed to air.

(c) They should have no water of hydration, that is, should not be hygroscopic or efflorescent.

(d) They should be less volatile or non-volatile so that cannot easily lose any component due to evaporation.

(e) They should be highly soluble in water.

(f) They should have high molecular mass to enable measurement of a significant amount of a substance for a specific concentration.

Examples of primary standards are sodium carbonate (Na_2CO_3) for acid-base titrations, and examples for redox titrations are potassium dichromate ($K_2Cr_2O_7$) and oxalic acid ($H_2C_2O_4$).

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Secondary standard solution

Most common acids such as hydrochloric acid (HCl), and sulphuric acid (H_2SO_4) as well as bases such as sodium hydroxide (NaOH) and potassium hydroxide (KOH) are not primary standards. This is because they absorb water or vaporise during storage or preparation. Their exact amounts cannot therefore be determined by direct measurement (weighing). However, their concentrations in solutions can be determined by standardising with primary standards. The standardised solutions are referred to as *secondary standards.*

Preparation of standard solutions

Standard solutions are prepared by dissolving a known amount of the primary standard to get a known volume of a solution by using a specific solvent such as distilled water. If a particular primary standard is not available, the solution should be standardised with a primary standard solution of another substance. For example, sodium chloride (NaCl) is a primary standard which can be used to standardise silver nitrate solution by titration. Weighing of the materials for standard solution preparation is done using an analytical balance. The analytical balance has high precision, and accuracy of more than three decimal places. Figure 5.14 is an example of an analytical balance that is used in measuring the masses to a precision of four decimal places.



Figure 5.14: *Analytical balance*

Preparation of primary standard solutions

Primary standard solutions from solid chemicals are prepared through the following procedure:

1. Identify the volume and concentration of the standard solution required.
2. Calculate the number of moles of the substance needed to make up the standard solution.
3. Work out the molar mass of the substance.
4. Using the results of the calculations in Step 2, calculate the mass of the substance needed to prepare the standard solution.
5. Carefully weigh the required mass of the substance (use an analytical balance). Do the following in order to weigh the required mass accurately:

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(a) Weigh an empty watch glass, then add the mass of the watch glass and the mass of the required substance.

(b) Repeat weighing the watch glass together with the required substance until the respective total mass is obtained to four decimal places.

Suppose you want to weigh 65.0000 g of the substance using a watch glass of 45.2551 g. The correct mass of the substance and the watch glass that should be measured on the weighing balance is 65.0000 g + 45.2551 g = 110.2551 g.

6. Transfer the primary standard into a beaker. Rinse the watch glass by using distilled water from a wash bottle. Ensure that the rinsing water goes into the beaker containing the substance being used to prepare the primary standard solution.

7. Stir the mixture in the beaker with a glass rod until all the solute is dissolved, and then transfer the solution into a volumetric flask.

8. Rinse the beaker and the glass rod by using distilled water from a wash bottle. Do this at least twice and pour the rinsing water into the volumetric flask each time.

9. Add water to the solution to just below the graduation mark on the volumetric flask.

10. Top up the solution with water up to the graduation mark by using a clean dropper. Ensure that the lowest part of the meniscus is exactly at the graduation mark.

11. Close the mouth of the flask with a stopper and invert it several times to ensure that the solution is homogeneously mixed.

12. Work out the molar concentration (molarity) of the solution.

13. Label the solution correctly by indicating its name and concentration, for example 2.5045 M NaCl. On the label, also indicate the preparation date if the solution is to be used for a long time.

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Activity 5.6



Aim: To prepare 1.0 litre standard solution of 0.1000 M sodium carbonate

Requirements: Bunsen burner, analytical balance, evaporating dish, glass rod, watch glass, 1000-mL volumetric flasks, funnel, 250-mL beaker, wash bottle, tripod stand, pipe-clay triangle, desiccator, dropper, anhydrous sodium carbonate, and distilled water

Procedure

1. Calculate the mass required to make 1.0 litre of 0.1000 M Na_2CO_3 solution. Since only 1.00 L solution

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has to be prepared, then 0.1 mole is required. The molar mass of Na_2CO_3 is 106.0 g/mol. Thus, the mass of Na_2CO_3 required in this experiment is 0.1 mol × 106.0 g/mol = 10.60 g.

2. Weigh 15 g of sodium carbonate and heat gently in an evaporating dish to obtain a constant mass. Allow it to cool to room temperature while placed in a desiccator.
3. Weigh a clean and dry watch glass and record its mass; let us assume the mass to be 10.4550 g.
4. Add the mass of the empty watch glass and that of the required sodium carbonate. That is 10.4550 g + 10.60 g = 21.0550 g.
5. Place the watch glass on the analytical balance and then add the dried sodium carbonate in Step 2, until the balance reads 21.0550 g.
6. Transfer the weighed amount of sodium carbonate into a 250-mL beaker containing at least half volume of warm distilled water. Warm water speeds up the process of dissolving the chemical. Stir the mixture by using a clean glass rod.
7. Rinse the watch glass with distilled water and transfer the rinsing water into the beaker.
8. Continue stirring the mixture until all the solute dissolves.
9. Transfer the solution into a clean 1-L volumetric flask by using a funnel. Be careful to avoid spillage of the solution outside the volumetric flask.
10. Rinse the beaker with distilled water and transfer the rinsing water into the flask.
11. Continue adding distilled water to the solution until it is almost at the 1000-mL mark.
12. Seal the volumetric flask with a stopper and invert several times to mix the solution.
13. Complete adding water to the mark using a wash bottle and invert the volumetric flask again.

Questions

1. Why is sodium carbonate heated first?
2. Why is warm water used to dissolve the solid?
3. What could happen if the heated Na_2CO_3 was placed in an open container?
4. State the possible sources of errors in this experiment.
5. A 0.1000 M sodium carbonate solution should not be prepared by adding water to the mark of a 1-L volumetric flask, then add 10.60 g of sodium carbonate. Explain.

Preparation of secondary standard solutions of common bases and mineral acids

Common bases and acids are likely to have low degree of purity. This explains why they cannot be used as primary standards.

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Preparation of standard solutions of bases and mineral acids requires standardisation with primary standard solutions. The resulting solutions obtained after standardisation are known as *secondary standard solutions*.

Solutions of mineral acids are prepared by dilution of concentrated acids. Concentrated acids are bought as stock solutions from chemical suppliers. The molar concentrations of the stock solutions are usually not indicated on the reagent bottles. Only the density or specific gravity, molar mass, and percentage purity of the acids are indicated. For example, a label on a bottle containing concentrated sulphuric acid could have the information as shown in Figure 5.15.

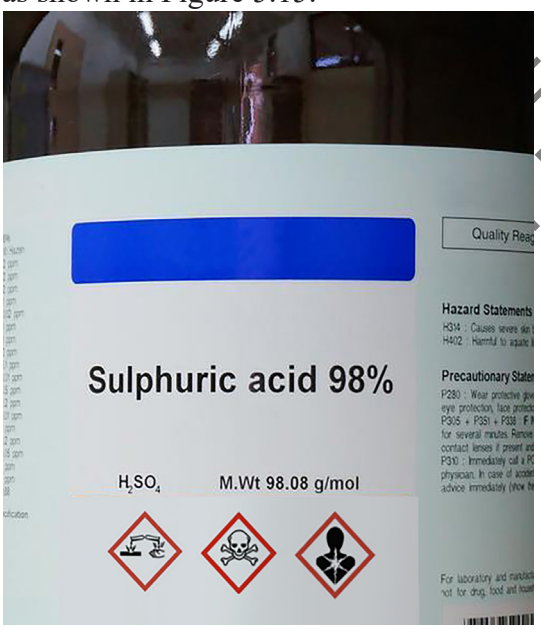


Figure 5.15: *Label on the bottle of sulphuric acid*

The molar concentration of the acids is calculated from the density and percentage purity. Stock solutions of hydrochloric acid, sulphuric acid and nitric acid usually have molarities of 11.6 M, 18.4 M, and 15.3 M, respectively. A standard solution is prepared by taking an appropriate amount of the concentrated acid and adding it to water in a volumetric flask of a known volume. The volume of the solution is then adjusted to the mark by using distilled water, stoppered and thoroughly mixed by inverting the flask several times. The molarity of the dilute acid is then indicated on the flask. This is only an approximate molar concentration of the acid. The acid is then standardised by using a standard solution of a base for accurate determination of its molarity.

To prepare 1000.00 cm^3 of 0.1000 M sulphuric acid solution from the stock solution of sulphuric acid, the dilution process is carried out as follows:

1. From the information on the bottle, 1.0000 M sulphuric acid contains 98 g in 1 dm^3 (the molar mass of sulphuric acid is 98 g/mol). Therefore, 0.1000 M sulphuric acid will contain 98 g/mol × 0.1 mol = 9.80 g in 1 dm^3.
2. From the value of the density of the concentrated sulphuric acid, 1.84 g of the sulphuric acid occupies 1 cm^3.
3. Since the concentrated sulphuric acid is 98% pure, the actual mass of sulphuric acid in 1 cm^3 of the concentrated solution is;

$$\frac{98}{100} \times 1.84 \text{ g} = 1.8032 \text{ g}$$

1.8032 g of sulphuric acid are in 1 cm^3 of concentrated sulphuric acid solution. Therefore, the volume containing 9.80 g of the acid is given by;

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$$\frac{1\ cm^3 \times 9.80\ g}{1.8032\ g} = 5.43\ cm^3$$

Alternatively, the molarity of concentrated sulphuric acid is given by:

$$= \frac{1000 \times \%\,\text{purity of acid} \times \text{density}}{100 \times \text{molar mass}}$$

$$= \frac{1000 \times 98\% \times 1.84\ g\ L^{-1}}{100 \times 98\ g\ mol^{-1}}$$

$$= 18.4\ M \text{ of concentrated acid.}$$

Then, by using the dilution law; number of moles before dilution are equally the same as the number of moles after dilution. That means $M_1V_1 = M_2V_2$

Given; $M_1 = 18.4000$ M, $M_2 = 0.1000$ M, $V_2 = 1000.00\ cm^3$

Required; V_1

From the dilution law:

$$V_1 = \frac{M_2V_2}{M_1}$$

$$V_1 = \frac{0.1000\ M \times 1000.00\ cm^3}{18.4\ M} = 5.43\ cm^3$$

Therefore, 5.43 cm^3 will be drawn from the stock solution and diluted with distilled water to the mark of 1000-mL volumetric flask. This makes a 0.1 M of sulphuric acid solution. Practically, 5.43 cm^3 cannot be measured by using the common laboratory facilities. In this case, measure 5.40 cm^3.

Caution! Never add water into acid, instead add acid into water.

Proceed with the procedure to prepare the dilute solution as described in the following steps:

4. Measure 5.40 cm^3 of the concentrated acid by using a 10-mL measuring cylinder and transfer the acid into a 1000-mL volumetric flask containing a small amount of distilled water.
5. Rinse the container containing the sulphuric acid with a small amount of distilled water and pour the rinsing water in the volumetric flask containing the solution.
6. Add more distilled water into the volumetric flask up to the mark and stopper it.
7. Invert the flask several times to obtain a uniform solution.

The solution prepared is roughly 0.1000 M sulphuric acid. If the concentration of the original solution is M_1 and that of the dilute solution is M_2, the following relationship holds:

$M_1V_1 = M_2V_2$, where, V_1 and V_2 are the volumes of the solutions of the respective concentrations. This relationship is applied when preparing solutions of dilute solutions from concentrated solutions with known molarity.

Activity 5.7



Aim: To prepare 1000.00 cm^3 of 0.5000 M nitric acid solution

Requirements: 25-mL measuring cylinder, flask with its stopper, a beaker, a glass rod, distilled water (1500 mL), 1000-mL volumetric flask, and nitric acid

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Note: Concentrated nitric acid is usually supplied from the manufacturers with the following composition or information: density = 1.41 g/cm^3, percentage purity = 70%, molecular weight = 63.0 g/mol. Different values of percentage purity are available from various manufacturers.

Procedure

1. Calculate the volume of nitric acid stock solution required for dilution.
2. Prepare a clean and dry 25-mL measuring cylinder.
3. Put a small amount of distilled water in the volumetric flask.
4. Using a 25-mL measuring cylinder, measure an appropriate amount of the calculated volume of concentrated nitric acid from the stock solution.
5. Transfer the measured volume of concentrated acid into a 1000-mL volumetric flask containing water.
6. Rinse the measuring cylinder used with distilled water and transfer its rinsing water into the volumetric flask containing the solution.
7. Add more distilled water into the volumetric flask up to the mark.
8. Stopper the volumetric flask and invert it several times to obtain a uniform solution.
9. Top-up the solution with more distilled water if the resulting solution is below the mark and mix it again.

Questions

1. What is the concentration of the nitric acid in the manufacturer's bottle?
2. What is the volume of concentrated nitric acid required to prepare 0.5000 M nitric acid? Show how you obtain the required volume.
3. Calculate the value of distilled water required in the preparation of 0.5000 M nitric acid.
4. Calculate the concentration of the concentrated and dilute nitric acid solution in g/L.
5. Calculate the number of moles of concentrated nitric acid before and after dilution.

Example 5.1

A student wants to prepare a solution whose concentration is 0.4000 M from 100.00 cm^3 of 2.0000 M hydrochloric acid. How many times should the solution be diluted?

Solution

Given; Initial concentration
$M_1 = 2.0000$ M
Initial volume $V_1 = 100.00$ cm^3
Final concentration $M_2 = 0.4000$ M
Required; Final volume $V_2 = ?$

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From the dilution law; the number of moles before dilution are equal to number of moles after dilution. Therefore,

$$M_1V_1 = M_2V_2$$

$$2.0000\ M \times 100.00\ cm^3 = 0.400\ M \times V_1$$

$$V_2 = \frac{2.0000\ M \times 100.00\ cm^3}{0.4000\ M}$$

$$V_2 = 500.00\ cm^3$$

The solution should be diluted from 100 cm^3 to 500 cm^3; therefore, the number of times the solution should be diluted is;

$$\text{Dilution factor} = \frac{\text{final volume } (cm^3)}{\text{Initial volume } (cm^3)} = \frac{500.00\ cm^3}{100.00\ cm^3} = 5$$

$$\text{or } \frac{\text{molarity of concentrated solution } (M_1)}{\text{molarity of dilute solution } (M_2)} = \frac{2.0000\ M}{0.40000\ M} = 5$$

Thus, the solution should be diluted five times.

Caution! Take care while handling concentrated acids because they are corrosive.

Task

Visit your school laboratory where chemicals are stored. Read different labels present on the bottles or other containers of chemicals and extract the following information: Percentage purity, specific gravity, density, molar mass, storage conditions, and safety symbols of the substances present in them. Present your information to other students.

Percentage purity of alkalis and acids

Many chemicals sold commercially are usually not 100% pure. Their purity in terms of compositions are based on mass or volume, and must be determined before being used in preparation of solutions. The higher the composition of the substance needed, the higher its purity. The composition can be expressed in terms of either weight (w/w%), volume (v/v%) or both weight and volume (w/v%). In many cases, the percentage weight (w/w%) is expressed as wt% or % for density. The purity is very important as impurities can interfere reactions of the pure substance involved. The percentage purity of the sample is obtained by the following formula:

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$$\text{Percentage purity} = \frac{\text{amount of pure compound in the impure sample}}{\text{amount of impure sample}} \times 100\%$$

In order to determine the mass of a pure substance from a given impure solid substance that is required to prepare a certain solution, first calculate the amount (number of moles) of a pure chemical substance. Secondly, multiply the number of moles of a pure substance by the molar mass of the substance to get the mass of the pure substance that is required. That is;

$$\text{Mass of pure substance} = \text{moles of pure substance} \times \text{molar mass}$$

To obtain the mass of an impure substance that is required to prepare a certain mass of the pure substance, divide the mass of the substance by the percentage purity. That is;

$$\text{Mass of impure} = \frac{\text{mass required}}{\text{percentage purity}}$$

For the liquid reagents such as acids, divide the mass of the impure chemical by its density to get the volume of impure substance required. That is;

$$\text{Volume needed} = \frac{\text{mass of impure substance (g)}}{\text{density of substance (g/cm}^3\text{)}}$$

Note: If an impure sample of a chemical substance of known percentage purity is used in a chemical reaction, the percentage purity has to be used in stoichiometric calculations. On the other hand, the percentage purity of an impure sample of a substance of unknown percent purity can be determined by reaction with a pure compound as in an acid-base titration.

Example 5.2

Calculate the mass of 98.5% Na_2CO_3 required for preparing 10.60 g Na_2CO_3.

Solution

$$\text{Mass of impure} = \frac{\text{mass required}}{\text{percentage purity}}$$

$$\text{Mass} = \frac{10.60\ \text{g}}{98.5/\ 100}$$

$$= \frac{10.60\ \text{g} \times 100}{98.5}$$

$$\text{Mass} = 10.76\ \text{g}$$

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Example 5.3

What mass of potassium hydroxide with 85.9% purity is needed to prepare 1 litre of a 0.2500 M solution of potassium hydroxide? The molar mass of potassium hydroxide is 56.11 g/mol.

Solution

First, calculate the mass of the pure substance required:

$$\text{Molarity} = \frac{\text{concentration (g/L)}}{\text{molar mass (g/mol)}}$$

$$\text{Concentration (g/L)} = \text{molarity} \times \text{molar mass}$$

$$= 0.25 \text{ mol/L} \times 56.11 \text{ g/L}$$

$$= 14.03 \text{ g/L}$$

The mass of pure substance in 1 litre = 14.03 g.

The mass of the impure potassium hydroxide is given by:

$$\text{Mass of an impure KOH} = \frac{\text{mass required}}{\text{percentage purity}}$$

$$= \frac{14.03 \text{ g}}{85.9/100}$$

$$= \frac{14.03\text{g} \times 100}{85.9}$$

$$= 16.33 \text{ g}$$

Therefore, 16.33 g of the impure potassium hydroxide are needed to prepare 1 litre of 0.2500 M potassium hydroxide solution.

Example 5.4

What volume of hydrochloric acid with 37.1% purity and a density of 1.2 g/cm^3 should be used to prepare 1 litre of 0.1000 M of hydrochloric acid solution? (The molar mass of hydrochloric acid is 36.5 g/mol.).

Solution

Molar mass of HCl = 36.5 g/mol.

0.1000 M hydrochloric acid solution contains 0.1 mol of the acid in 1 litre of solution.

First convert the number of moles to mass;

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$$\text{Concentration (g/L)} = \text{Molarity} \times \text{Molar mass}$$
$$= 0.1\,\text{mol/L} \times 36.5\ \text{g/mol}$$
$$= 3.65\ \text{g/L}$$

Since the acid is 37.1% pure, the mass of its impure acid required is:

$$\text{Mass of impure} = \frac{\text{Mass required}}{\text{Percentage purity}}$$
$$= \frac{3.65\ \text{g} \times 100}{37.1}$$
$$= 9.84\ \text{g}$$

The volume of the impure acid required is given by:

$$\text{Volume} = \frac{\text{mass}}{\text{density}}$$
$$= \frac{9.84\ \text{g}}{1.2\ \text{g/cm}^3}$$
$$\text{Volume} = 8.20\ \text{cm}^3$$

Note: Alternatively, the volume of impure acid can be determined by using a similar approach used in the preparation of the standard solutions of common mineral acids

Titrations

Titration is a laboratory technique which is used to determine the unknown concentration of a solution using a standard solution (whose concentration is known). The two reagents should react chemically in definite proportions. Not all chemical reactions are suitable for titrations. The following conditions are necessary for a titration technique to be applied:

1. The reaction must be fast so that the titration can be performed by using convenient time.
2. The reaction should go to completion; it should not be a reversible reaction.
3. The reaction should be free from side reactions; meaning that it can be represented by a single chemical equation.
4. The reaction should have a definite end point that can be determined accurately.

Common titration reactions include acid-base reactions, precipitation reactions, and redox reactions. In this section, only the acid-base titrations are covered.

Acid-base titrations

An acid-base titration involves an acid and a base. Acid-base titrations are the most common titrations in the laboratory.

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Usually, the *titrant* (a solution of known concentration) placed in a burette is added to a known quantity of the *titrand* (the solution of unknown concentration) until the reaction is complete. The completion of the reaction is indicated by the change in the colour of the indicator, or occurrence of certain permanent physical conditions such as precipitate formation. Table 5.1 shows some acid-base indicators and their colours in acidic and basic media.

Table 5.1: *Common acid-base indicators*

Indicator	Colour in acidic medium	Colour in basic medium
Litmus	Red	Blue
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Pink

The choice of the indicator to be used depends on the strength of the reactants. Table 5.2 shows the appropriate indicators for different acid-base titrations.

Table 5.2: *Appropriate indicators for different acid-base titrations*

Titration reaction	Example	Most suitable indicator
Weak base + Weak acid	Ammonia solution + Ethanoic acid	No suitable indicator
Strong base + Weak acid	Sodium hydroxide solution + Ethanoic acid	Phenolphthalein (POP)
Strong base + Strong acid	Sodium hydroxide solution + Hydrochloric acid	Any indicator
Strong acid + Weak base	Hydrochloric acid + Sodium carbonate solution	Methyl orange (MO)

Procedure for carrying out acid-base titrations

The following are the steps to be followed when performing acid-base titrations:

1. Clean the burette, conical flasks and the pipette before use.
2. Prepare the retort stand that will be used to hold the burette upright.
3. Use a funnel to fill the burette with the titrant, usually the acid.
4. Adjust the bubble at the tip of the burette.
5. Record the initial burette reading to two decimal places.
6. Pipette the alkali (base) into a conical flask. Record the volume of the base for the pipette used.
7. Add 2 to 3 drops of the indicator in the conical flask and record the colour of the solution.

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8. Slowly allow the acid to drain from the burette into the alkali in the flask while swirling carefully the contents in the flask until a permanent colour change of the indicator is observed. Record the final reading of the burette and the colour of the resulting solution. Note that when the end point approaches, the colour of the solution will start to change. At this point, add the acid in small drops until there is a permanent colour change in the mixture.
9. Calculate the *volume used,* sometimes known as the *titre value* by subtracting the initial burette reading from the final burette reading and record the volume of the acid used.
10. Repeat the Steps 6−9 three more times. The difference in volumes used should be one to three drops, of which one drop is approximately equal to 0.1 cm^3, thus the difference between one experiment to another should not exceed 0.3 cm^3.
11. Calculate the average titre value.
12. From the average titre value, the concentration of the substance is worked out.

Note: The first titration is usually referred to as the rough reading and recorded as a pilot; therefore, it is not included in the calculations.

To determine the concentration (molarity) of the acid, proceed as follows:

1. Calculate the average volume of the acid used (average titre volume).
2. Calculate the number of moles of the base used using the pipetted volume.
3. Write the balanced chemical equation. For example, if the reaction involved Na_2CO_3 and HCl solutions, the balanced chemical equation will be:

$$Na_2CO_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

4. Determine the stoichiometric coefficient of hydrochloric acid used, that is number of moles. From the equation, the stoichiometric coefficient of Na_2CO_3:HCl is 1:2. Therefore, the number of moles of hydrochloric acid are twice that of sodium carbonate.
5. Use the average volume for HCl to calculate the molarity.
   Alternatively, the mole ratio for the acid and base are related to their stoichiometric coefficients, that is number of reacting moles by:

$$\frac{M_A V_A}{M_B V_B} = \frac{n_A}{n_B}$$

where

M_A is the molar concentration of the acid;

V_A is the volume of the acid used;

n_A is the number of the moles of the acid used (found from a balanced equation);

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M_B is the molar concentration of the base;

V_B is the volume of the base used, and

n_B is the number of moles of the base used (found from a balanced equation).

A similar procedure is used to determine the concentration of hydrochloric acid by using a known concentration of sodium hydroxide. In this case, consider the mole ratio from a balanced chemical equation between NaOH and HCl. The balanced chemical equation is:

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

Activity 5.8



Aim: To standardise dilute hydrochloric acid by using a sodium carbonate solution.

Requirements: 50-mL burette, filter funnel, stand with a clamp, 25-mL pipette, three 250-mL conical flasks, white tile, standard sodium carbonate solution (0.0980 M), dilute hydrochloric acid, methyl orange indicator, and distilled water

Procedure

1. Fill the burette with dilute hydrochloric acid up to the zero mark. Record the initial burette reading.
2. Transfer 25.00 cm^3 of the 0.0980 M sodium carbonate solution into a clean conical flask.
3. Add two to three drops of methyl orange indicator into the conical flask.
4. Titrate the solution with hydrochloric acid until the colour of the indicator just changes and record the burette reading.
5. Repeat the Steps 2–4 three more times. Refill the burette when the level of the acid falls below half.
6. Record the results in a tabular form as shown in Table 5.3.

Table 5.3: *Titration results for the reaction between Na_2CO_3 and HCl.*

Burette readings	Pilot	1	2	3
Final burette reading (cm^3)				
Initial burette reading (cm^3)				
Volume of HCl (cm^3)				

The volume of the pipette used is_____.

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Questions

1. Explain the colour change of methyl orange before titration and at the end point.
2. Write a balanced chemical equation for the reaction.
3. Calculate the moles and concentration in mol/dm^3 and g/dm^3 of the hydrochloric acid used.
4. Calculate the number of moles of sodium carbonate used in the reaction.

Activity 5.9



Aim: To determine the concentration of hydrochloric acid by titration using standard sodium hydroxide solution

Requirements: Beakers, conical flasks, burette, pipette, stand and clamp, hydrochloric acid, 0.4000 M sodium hydroxide solution, and phenolphthalein indicator

Procedure

1. Measure 150.00 cm^3 of hydrochloric acid and 150.00 cm^3 of sodium hydroxide solution and place them in separate beakers.
2. Fill the burette with hydrochloric acid to the required mark and record the volume. The burette should be fixed on the retort stand.
3. Pipette 25.00 cm^3 of the base and put it into a clean conical flask. Then add three drops of phenolphthalein indicator.
4. Titrate the two components four times and record their results in a tabular form.

Questions

1. Calculate the molarity of the HCl.
2. Calculate the concentration of the HCl in g/dm^3.
3. Why is the phenolphthalein indicator suitable for this titration experiment?

Exercise 5.2



1. Describe how you can prepare a 0.2000 M standard solution of sodium carbonate by using sodium carbonate as a primary standard.
2. (a) Explain how you can prepare 2.0 dm^3 of 0.0500 M sodium hydroxide solution from a stock solution of 6.2500 M sodium hydroxide solution that is 90% pure.

   (b) What is the mass of the sodium hydroxide in the final solution?
3. Sodium hydroxide solution (25.00 cm^3) was titrated with 0.1000 M hydrochloric acid. The results obtained are presented in Table 5.4:

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Table 5.4: *Titration results for the reaction between HCl and NaOH*

Burette readings	Pilot	1	2	3
Final burette reading (cm^3)	50.00	40.00	30.00	42.60
Initial burette reading (cm^3)	36.50	27.45	17.50	30.00
Volume of HCl (cm^3)	13.50	12.55	12.50	12.60

Determine the molarity of the sodium hydroxide solution.

4. A 26.50 g sample of sodium carbonate was dissolved in distilled water to make 500.00 cm^3 of solution. A 25.00 cm^3 volume of the solution reacted completely with 20.00 cm^3 of dilute sulphuric acid. Calculate the molarity of the acid.

Determination of properties of chemical substances

Volumetric analysis is used to determine various properties of chemical substances such as the relative atomic mass of an unknown element in a particular substance, calculating the percentage purity of components in the sample, and calculating the water of crystallisation in the formula unit.

Determination of the relative atomic masses of unknown elements in an acid or base

The atomic mass of an unknown element in an acid or alkali can be determined by titrating its solution with a standard solution of a base or acid, respectively. The relative atomic mass of the metal, can be determined as follows:

1. Calculate the average volume of the acid used.
2. Calculate the number of moles of the acid used.
3. Write a balanced chemical equation for the reaction to obtain the mole ratio. For example;

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

$$Na_2CO_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

4. Determine the number of moles of the base used. Use the mole ratio of the base to the acid.
5. Calculate the concentration of the metal salt in 1 litre solution.
6. Work out the molarity of the base solution by dividing the number of moles with its respective volume. Remember to change the volume into litres.

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7. Use the following relation to calculate the molar mass of the base:

$$\text{Molarity} = \frac{\text{concentration (g/L)}}{\text{molar mass (g/mol)}}$$

In this relation, only the molar mass is unknown.

8. Calculate the relative atomic mass of the metal based on the formula of the base.

Example 5.5

A 2.12 g sample of an unknown monovalent metal carbonate (M_2CO_3) was dissolved in water to make 250 cm^3 of a solution. The following results were obtained in the experiment when 25 cm^3 of the monovalent metal carbonate was reacted with 0.2500 M hydrochloric acid solution by using a methyl orange indicator.

Burette readings (cm^3)	Pilot	1	2	3
Final burette reading (cm^3)	17.00	16.10	32.00	48.10
Initial burette reading (cm^3)	0.00	0.00	16.10	32.00
Volume of HCl (cm^3)				

(a) Calculate the average volume of the hydrochloric acid used in the reaction.

(b) Write a balanced chemical equation for the reaction.

(c) Calculate the number of moles of hydrochloric acid and those of the metal carbonate solution used in the reaction.

(d) Determine the relative molecular mass of metal carbonate.

(e) Determine the relative atomic mass of the metal.

(f) Identify the unknown element 'M' in the compound M_2CO_3.

Solution

(a) To calculate the average volume of the acid used, first calculate the titre values for each titration as follows:

Burette readings (cm^3)	Pilot	1	2	3
Final burette reading (cm^3)	17.00	16.10	32.00	48.10
Initial burette reading (cm^3)	0.00	0.00	16.10	32.00
Volume of HCl (cm^3)	17.00	16.10	15.90	16.10

The average volume is determined by averaging the three titre values.

$$\frac{(16.10 + 15.90 + 16.10)\ cm^3}{3} = 16.03\ cm^3$$

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Therefore, 16.03 cm^3 of hydrochloric acid completely reacted with 25.00 cm^3 of the metal carbonate solution.

(b) To write a balanced chemical equation for the reaction as follows,

$$M_2CO_3(aq) + 2HCl(aq) \longrightarrow 2MCl(aq) + H_2O(l) + CO_2(g)$$

(c) To calculate the numbers of moles of the hydrochloric acid and carbonate used;

Number of moles of HCl used will be obtained as follows:

The molarity of the acid is 0.2500 M. This means that 1000 cm^3 of HCl contains 0.25 moles of the acid.

$$\text{Therefore, 16.03 } cm^3 \text{ contain } \frac{0.25 \text{ mol} \times 16.03 \text{ } cm^3}{1000 \text{ } cm^3} = 0.004 \text{ moles}$$

The number of moles of the metal carbonate used will be calculated as follows:

From the balanced equation, the mole ratio of the metal carbonate to the acid is 1:2. Therefore, the number of moles of the metal carbonate used in the reaction will be:

$$\frac{0.004 \text{ mol}}{2} = 0.002 \text{ mol}$$

(d) To determine the relative molecular mass of metal carbonate:

Since 0.002 moles of the metal carbonate are contained in 25 cm^3, 1000 cm^3 contain;

$$\frac{0.002 \text{ mol} \times 1000 \text{ } cm^3/L}{25 \text{ } cm^3} = 0.08 \text{ mol/L}$$

The molarity of the carbonate solution is 0.0800 M.

Since 250 cm^3 of the solution contain 2.12 g of the metal carbonate; 1000 cm^3 of the solution contain

$$\frac{2.12 \text{ g} \times 1000 \text{ } cm^3/L}{250 \text{ } cm^3} = 8.48 \text{ g/L of the metal carbonate}$$

$$\text{Molarity} = \frac{\text{concentration } (g/dm^3)}{\text{molar mass } (g/mol)}$$

$$\text{Therefore, molar mass} = \frac{\text{concentration } (g/dm^3)}{\text{molarity } (mol/dm^3)}$$

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$$= \frac{8.48 \text{ g/dm}^3}{0.08 \text{ mol/dm}^3}$$

$$= 106 \text{ g/mol}$$

The relative molecular mass of the metal carbonate is 106.

(e) To determine the relative atomic mass of the metal: From relative molecular mass of $M_2CO_3 = 106$ g, and atomic masses of C = 12 and O = 16;

The relative atomic mass of the metal (M) in M_2CO_3 is calculated as;

$$2M + 12 + (16 \times 3) = 106$$
$$2M + 60 = 106$$
$$2M = 46$$
$$M = 23$$

The relative atomic mass of the metal M = 23

(f) From the Periodic Table, an element with atomic mass of 23 is sodium.

Determination of the percentage purity of a substance

Titration reactions are used to determine the percentage purity of common laboratory reagents.

The percentage purity of the sample of a chemical substance can be determined as follows:

1. Calculate the molarity of the sample.
2. Calculate the concentration of the sample in g/L.
3. Calculate the amount of the chemical substance in a total volume of solution. That is the mass of pure chemical substance.
4. Work out the percentage purity of the sample by using the previously described relation.

Activity 5.10



Aim: To determine the percentage purity of an impure sample of sodium carbonate

Requirements: 50-cm^3 burette, a stand with a clamp, 25-cm^3 pipette, three 250-cm^3 conical flasks, 250-cm^3 volumetric flask, white tile, standard solution of 0.5 M hydrochloric acid, an impure sample of sodium carbonate, methyl orange indicator, and distilled water

Procedure

1. Prepare a solution of the sodium carbonate by dissolving 6.63 g of the impure sample in distilled water to make a 250 cm^3 solution.
2. Fill the burette with a 0.5 M hydrochloric acid solution and record the initial reading of the burette.
3. Using a pipette, transfer 25 cm^3 of the sodium carbonate solution into a conical flask.
4. Add two to three drops of methyl orange indicator into the conical flask.

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5. Titrate the solution against hydrochloric acid until there is a permanent colour change.
6. Repeat the Steps 3–5 three times and record the results in a tabular form.

Questions

1. Identify the colour change observed during the reaction.
2. Calculate the average volume of hydrochloric acid used in the reaction.
3. Write a balanced chemical equation for the reaction.
4. Calculate the number of moles of hydrochloric acid and sodium carbonate used in the reaction.
5. Determine the percentage purity of the sample of sodium carbonate.

Example 5.6

A 5.125 g sample of impure sodium carbonate was dissolved to make 250.00 cm^3 of a solution. Calculate the percentage purity of the sodium carbonate sample if 25.00 cm^3 of the solution required 35.80 cm^3 of 0.0500 M sulphuric acid for complete neutralisation.

Solution

The number of moles of sulphuric acid used in the reaction is given by:

$$= \text{Molarity} \times \text{volume of solution } (dm^3)$$

$$= 0.0500 \text{ mol/dm}^3 \times \frac{35.80}{1000} \text{ dm}^3$$

$$= 0.00179 \text{ mol}$$

The equation for the reaction is;

$$Na_2CO_3(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$$

The mole ratio of Na_2CO_3 to H_2SO_4 is 1:1.

Therefore, the number of moles of sodium carbonate used in the reaction is 0.00179.

The molar mass of sodium carbonate is 106 g/mol. The mass of sodium carbonate in 0.00179 moles is obtained as follows:

$$\text{Mass} = \text{molar mass} \times \text{number of moles}$$

$$= 106 \text{ g/mol} \times 0.00179 \text{ mol}$$

$$= 0.18974 \text{ g}$$

Therefore, the mass of sodium carbonate in 25 cm^3 of the solution is 0.18974 g.

The mass in 250 cm^3 of the solution is calculated as:

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$$\text{Mass} = \frac{0.18974\ \text{g} \times 250\ \text{cm}^3}{25\ \text{cm}^3}$$

$$= 1.8974\ \text{g}$$

The percentage purity of the sodium carbonate sample is obtained using the following formula:

$$\%\text{purity} = \frac{\text{mass of the substance in a sample}}{\text{mass of the impure sample}} \times 100\%$$

$$= \frac{1.8974\ \text{g}}{5.125\ \text{g}} \times 100\%$$

$$= 37.02\%$$

Alternatively,

Solution

To calculate the concentration of impure Na_2CO_3;

$$\text{Concentration} = \frac{\text{Mass}}{\text{Volume}} = \frac{5.125\ \text{g}}{0.25\ \text{dm}^3} = 20.5\ \text{g/dm}^3$$

Chemical equation: $Na_2CO_3(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$

Mole ratio $n_A = 1$ and $n_B = 1$; n_A = Number of moles of acid
n_B = Number of moles of base

Molarity of acid (M_A) = 0.05 M

Volume of acid (V_A) = 35.8 cm^3

Molarity of base (M_B) = ?

Volume of base (V_B) = 25 cm^3

From,

$$\frac{M_A V_A}{M_B V_B} = \frac{n_A}{n_B}$$

$$M_B = \frac{M_A V_A}{V_B} = \frac{0.05\ \text{M} \times 35.8\ \text{cm}^3}{25\ \text{cm}^3}$$

$$M_B = 0.0716\ \text{M}$$

To calculate the concentration of pure Na_2CO_3

$$\text{Concentration} = \text{molarity} \times \text{molar mass}$$

$$= 0.0716\ \text{mol/dm}^3 \times 106\ \text{g/mol}$$

$$= 7.5896\ \text{g/dm}^3$$

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$$\text{Percentage purity} = \frac{\text{concentration of pure sample (g/dm}^3\text{)}}{\text{concentration of impure sample (g/dm}^3\text{)}} \times 100\%$$
$$= \frac{7.5896 \text{ g/dm}^3}{20.50 \text{ g/dm}^3} \times 100\%$$
$$= 37.02\%$$

Determination of water of crystallisation in a substance

Water of crystallisation is the water that is bound within the crystals of a substance. This water can be removed by heating. Titration reactions are used to determine the amount of water of crystallisation in a substance. For example, the number of molecules of water of crystallisation in a hydrated sodium carbonate sample can be determined as follows:

1. Calculate the concentration of the hydrated carbonate sample using 250 cm^3 solution and report it in g/L.
2. Calculate the molarity of sodium carbonate using the experimental titration results and the balanced chemical equation.
   Note: Water of crystallisation is not involved in the reaction.
3. Calculate the concentration of sodium carbonate by multiplying its molarity by the molar mass.
4. Calculate the R.M.M. of the hydrated carbonate.
   Use the following relation to calculate the molar mass of the hydrated sodium carbonate:

$$\frac{\text{molar mass of anhydrous } Na_2CO_3}{\text{molar mass of hydrated } Na_2CO_3} = \frac{\text{concentratio of anhydrous } Na_2CO_3}{\text{concentratio of hydrated } Na_2CO_3}$$

   where,

   The molar mass of anhydrous Na_2CO_3 is 106 g/mol;
   The molar mass of the hydrated Na_2CO_3 is unknown;
   The concentration of anhydrous Na_2CO_3 is obtained from Step 2; and
   The concentration of the hydrated sample is obtained from Step 1.

   Alternatively,

   Use the following relation to calculate the molar mass of the hydrated sample:

$$\text{Molarity} = \frac{\text{concentration of hydrated carbonate (g/L)}}{\text{molar mass of hydrated carbonate (g/mol)}}$$

   Since the molarity and concentration of the hydrated carbonate are known, then the molar mass of the hydrated carbonate can be calculated. Assume that, the molar mass of the hydrated carbonate is 286 g/mol; then proceed to the next step.

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5. Determine the number of molecules of water of crystallisation in the sample using the relative atomic masses of various atoms in the hydrated sodium carbonate ($Na_2CO_3 \cdot xH_2O$) as;

$$2Na + C + 3O + x(2H + O) = 286$$
$$(2 \times 23) + 12 + (3 \times 16) + x[(2 \times 1) + 16] = 286$$
$$106 + 18x = 286$$
$$18x = 286 - 106$$
$$18x = 180$$
$$x = 10$$

There are 10 molecules of water of crystallisation in the hydrated sodium carbonate and the formula of the sodium carbonate is therefore $Na_2CO_3 \cdot 10H_2O$.

Activity 5.11



Aim: To determine the number of molecules of water of crystallisation in a sample of hydrated sodium carbonate through titration

Requirements: Burette, retort stand, 25-cm^3 pipette, three 250-cm^3 conical flasks, white tile, 0.2400 M hydrochloric acid solution, methyl orange indicator, distilled water, and sodium carbonate solution made by dissolving 38.90 g of the hydrated solid ($Na_2CO_3 \cdot xH_2O$) to make one litre of solution

Procedure

1. Clean the burette, conical flasks and pipette thoroughly before use.
2. Fill the burette with the hydrochloric acid solution up to the zero mark.
3. Record the initial reading of the burette.
4. Measure 25.00 cm^3 of the sodium carbonate and transfer it into a conical flask.
5. Add 2 to 3 drops of methyl orange indicator to the solution in the conical flask.
6. Titrate the mixture against hydrochloric acid until a permanent colour change occurs and record the volume.
7. Repeat the Steps 4–6 three times and record the results in a tabular form.

Questions

1. Calculate the average volume of the acid used.
2. Calculate the number of moles of the hydrochloric acid used in the reaction.

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3. Write a balanced chemical equation for the reaction.
4. Calculate the number of moles of the sodium carbonate used in the reaction.
5. Calculate the number of molecules of water of crystallisation in the molecule of sodium carbonate.

Exercise 5.3

1. A 17.25 g sample of an unknown metal carbonate (M_2CO_3) was dissolved in water to make 250.00 cm^3 of a solution. A 20.00 cm^3 portion of the solution was then titrated with 0.8000 M hydrochloric acid. The following results were obtained.

Burette readings	Pilot	1	2	3
Final burette reading (cm^3)	25.50	45.95	45.00	48.95
Initial burette reading (cm^3)	0.00	20.90	20.00	24.00
Volume of HCl (cm^3)	25.50	25.05	25.00	24.95

(a) Determine the relative atomic mass of the metal in the metal carbonate.
(b) Identify an element represented by 'M' in the compound M_2CO_3.

2. A 25.00 cm^3 volume of 2.0000 M nitric acid reacted completely with 20.00 cm^3 of a solution containing 56.25 g of impure sodium carbonate in 250.00 cm^3 of solution. Determine the percentage purity of the sodium carbonate. (N = 14; C = 12; Na = 23; O = 16; H = 1).
3. Hydrated oxalic acid ($H_2C_2O_4 \cdot xH_2O$) was made by dissolving 20.16 g of the acid to make one litre of the solution. The solution was then titrated against 20.00 cm^3 of 0.4000 M sodium hydroxide solution. The following results were obtained during the titration.

Burette readings	Pilot	1	2	3
Final burette reading (cm^3)	26.50	25.00	45.05	25.00
Initial burette reading (cm^3)	0.00	0.00	20.00	0.00
Volume of acid used (cm^3)	26.50	25.00	25.05	25.00

The equation for the reaction is

$$H_2C_2O_4 \cdot xH_2O(aq) + 2NaOH(aq) \longrightarrow Na_2C_2O_4(aq) + (x + 2)H_2O(l)$$

Calculate the number of molecules of water of crystallisation in the acid.

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Challenges in volumetric analysis

The processes that are involved in volumetric analysis can be associated with several challenges. The common challenges include leaking burettes, broken pipette tips and tops, problems in colour identifications, failure to hold the calibration mark at the eye level, unclearly seen calibration marks, and breaking of glass apparatus. The leaking of burettes normally occurs at the tap or stopcock. This is caused by a loose connection or wear out of the tap. When the leakage is serious, replace the burette with a new one, but when the leakage is low, applying grease to the leaking part may help.

Pipettes with broken tips cannot deliver accurate volumes of solutions since their readings start from the bottom. Do not use pipettes with broken tips. If a pipette is broken at the top, the laboratory technician can fix the problem and it can be reused.

The correct reading of the volumes in measuring apparatus is necessary to have accurate results. It is advised to read the meniscus when reading volumes. Volumes should be read at an upright straight position. Enough bright light should be present to aid visualisation of the calibrations when reading the volumes.

Sometimes the glass volume measuring apparatus can lose their calibration marks on their sides. Such apparatus pose difficulties in reading accurate volumes, and therefore should not be used anymore. Great care should also be taken when handling glassware since they are fragile and their replacement is expensive.

Applications of volumetric analysis

Volumetric analysis is used in different fields including medicine, industries, and agriculture.

Medicine: Titration is used in the determination of the concentration of glucose in a blood sample and the amount of acid which can be neutralised by an antacid tablet.

Industries: There are various uses of volumetric analysis in industries. For example,

(a) to determine the amount of acetic acid in vinegar. Table vinegar should have a minimum amount of 5% acetic acid content. This is established by titrating a known volume of vinegar with a standard solution of a base.

(b) to determine the acidity of fruit juices. This is done by titrating a known volume of the fruit juice with a standard solution of a base.

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(c) to determine the concentration of household ammonia. This is obtained when a known volume of the ammonia is titrated with a standard solution of an acid.

Agriculture: Titrations are used to determine the compositions of different substances contained in the soil. These substances include nitrogen, potassium, and phosphorus.

Other applications

There are other applications of volumetric analysis that include commercial uses such as;

(a) Determination of the percentage of iron content in an iron ore.
(b) Determination of salt content in brine.
(c) Determination of the percentage by mass of copper in copper salt.

Chapter summary

1. Volumetric analysis is a chemical procedure for determining unknown concentration of a solution through volumetric titration.
2. Standard volumetric apparatus include the volumetric flask, conical flask, funnel, dropper, retort stand and clamp, white tile or paper, burette, pipette, analytical balance, and measuring cylinder.
3. Standardisation is a procedure used to determine the concentration of a solution which is known as analyte.
4. Primary standard solutions are prepared by dissolving a known amount of a primary standard to make a required volume of solution using a specific solvent.
5. Secondary standard solutions are prepared by standardisation using primary standard solutions.
6. Titration is a laboratory technique for determining the unknown concentration of a solution using a standard solution.
7. Types of titration reactions include acid-base titrations, redox titrations, and precipitation titrations.
8. Volumetric analysis is a common technique used in the laboratory for the determination of concentrations of analytes in solution. It is also used in the determination of the relative atomic masses of unknown elements in an acid, base or any other compounds, determination of the percentage purity of the substance, and determination of the amount of water of crystallisation in a substance.
9. Volumetric analysis is also used in determining the concentrations of analytes in manufacturing industries, agriculture, domestic activities, and other fields.

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Revision exercise 5

Choose the correct answer in items 1–4.

1. What mass of solid potassium chloride is needed to prepare 250.00 cm^3 of 0.2350 M solution? (K = 39, Cl = 35.5).
   (a) 9.23 g (b) 31.3 g
   (c) 15.6 g (d) 4.38 g

2. What volume of 6.0000 M hydrochloric acid is needed to prepare 1.50 dm^3 of 0.2000 M hydrochloric acid solution? (H = 1, Cl = 35.5).
   (a) 1800.00 cm^3 (b) 125.00 cm^3
   (c) 50.00 cm^3 (d) 150. 00 cm^3

3. A solution is prepared by diluting 25 cm^3 of 1.20 M aluminium chloride solution to 250.00 cm^3. What is the molarity of the solution (Al = 27, Cl = 35.5).
   (a) 0.260 M (b) 0.12 M
   (c) 6.50 M (d) 1.20 M

4. What is the molarity of a solution containing 1.75 moles of the solute in 3.00 litres of solution?
   (a) 5.25 M (b) 1.71 M
   (c) 0.58 M (d) 1.75 M

5. The following table contains some names of laboratory apparatus.

Tripods stand	Bunsen burner
Measuring cylinder	Petri dish
Pipette	Combustion tube
Conical flask	Burette
Wire gauze	Retort stand and clamp

   (a) Using the names given in the box, identify five apparatus that are used in volumetric analysis.
   (b) Explain the uses of each of the apparatus you have identified in 5(a).

6. Explain the role of a pilot experiment in titration.

7. Give reasons as to why a burette is calibrated from the top downward, while the pipette is calibrated in the opposite way of a burette.

8. To what volume should a 300.00 cm^3 of 0.6000 M sodium hydroxide solution be diluted to give a solution with a concentration of 0.0050 M?

9. The complete neutralisation of 25.00 cm^3 of 0.1200 M sodium hydroxide solution requires 28.40 cm^3 of a sulphuric acid solution. Calculate the:
   (a) Concentration of the acid in grams per litre.
   (b) Molarity of the acid.

10. A student prepared a standard solution of sodium hydroxide by dissolving 40.00 g of the alkali in distilled water to make one litre of a solution. The student used 25.00 cm^3 of the solution to standardise a solution of sulphuric acid. The following were the results obtained during the titration:

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Burette readings	Pilot	1	2	3
Final reading (cm^3)	12.90	12.90	25.50	37.90
Initial reading (cm^3)	0.00	0.40	12.90	25.50
Volume used (cm^3)				

(a) Copy and complete the table of results.
(b) Calculate the average volume of the acid which neutralised 25.00 cm^3 of the alkali.
(c) Write a balanced chemical equation for the reaction.
(d) Calculate the:
  (i) Number of moles of sodium hydroxide in the solution used in the reaction.
  (ii) Number of moles of sulphuric acid required to neutralise 25.00 cm^3 of the hydroxide solution.
  (iii) Molarity of the sulphuric acid solution.
  (iv) Concentration of the acid in grams per litre.
  (Na = 23, H = 1, O = 16, S = 32)

11. A 39.75 g sample of unknown metal carbonate (Y_2CO_3) was dissolved to make 200.00 cm^3 of a solution. The solution was titrated with 1.0000 M sulphuric acid. If 20.00 cm^3 of the solution reacted completely with 37.50 cm^3 of the sulphuric acid, calculate the:
  (a) Number of moles of sulphuric acid in the reaction.
  (b) Number of moles of the metal carbonate in 20.00 cm^3 of the metal carbonate solution.
  (c) Molar mass of the metal carbonate.
  (d) Relative atomic mass of the metal Y in the formula unit and identify the metal.
  (Note that Y is not the actual symbol of the metal; C = 12, O = 16, H = 1, S = 32).
12. A solution was made by dissolving 14.00 g of an impure sample of potassium hydroxide (KOH) to make 100.00 cm^3. The solution was then titrated against 1.2500 M nitric acid solution. If 40.00 cm^3 of the acid solution required 25.00 cm^3 of the alkali solution for complete neutralisation, work out the percentage purity of the potassium hydroxide sample.
13. A solution of hydrated sodium carbonate was titrated with 1.6800 M nitric acid solution. It was found that 30.00 cm^3 of the solution required 28.75 cm^3 of the nitric acid for a complete reaction. If the solution was prepared by dissolving 138.14 g of the carbonate to make 600.00 cm^3 of solution, determine the number of molecules of water of crystallisation in the hydrated sodium carbonate, and write its correct formula.

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14. A certain commercial sulphuric acid solution contains 33.0% acid by mass and is packed in 2.50 litres bottles. If the density of such acid is 1.28 g/cm^3, calculate the:
    (a) Mass of sulphuric acid present in the bottle.
    (b) Molarity of the commercial acid.
    (c) Volume of the acid required to be diluted into 2.00 litres of 0.1250 M of the acid solution.

15. Explain the applications of volumetric analysis in your daily life activities and chemical production industries.

16. A doctor prescribed antacid tablets made from magnesium hydroxide to a student who was suffering from excess stomach acid:
    (a) Why did the doctor prescribe the antacid tablets?
    (b) Write a balanced chemical equation that took place in the student's stomach after taking the tablets.
    (c) If the normal acid content in the stomach is 160.00 millimolar and the stomach secrets 0.50 litres of gastric juice in a day, calculate the mass of the tablets to be given to the student if the gastric juice had 210.00 millimolar hydrochloric acid.
    (d) If each tablet of an antacid weighs 145 mg of magnesium hydroxide, how many tablets will the student be given?
    (e) If a dose requires 1 tablet per every 12 hours, how many days should the student take to finish the full dose?

17. In a certain experiment, a student was provided with a container of commercial sulphuric acid with the following information on its label: Molar mass = 98 g mol^{-1}, density =1.804 g cm^{-3}, and purity = 98%. The student was assigned to prepare a 1 L of 0.2000 M dilute solution of H_2SO_4 from the commercial bottle. The student measured 10.00 cm^3 of concentrated acid and diluted it to 1000.00 cm^3.
    (a) Calculate the molarity of the acid solution prepared by the student.
    (b) What volume of concentrated H_2SO_4 could be added to the 10 cm^3 to obtain exactly 0.2000 M of H_2SO_4 solution?

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Chapter Six

Ionic theory and electrolysis

Introduction

In Form Two, you learnt about atoms, which are the most basic units of elements made up of protons, neutrons, and electrons. When electrons are added to, or removed from, neutral atoms, ions are formed. Ions are particles that are either positively or negatively charged. These ions allow an electric current to flow in chemical substances leading to some chemical changes. This is known as electrolysis which is well explained by the ionic theory. In this chapter, you will learn the concept of ionic theory, electrolysis, mechanism of electrolysis, laws of electrolysis, and the applications of electrolysis in daily life. The competencies developed will enable you to apply skills in various fields such as a purification of different metals, protection of metals from corrosion, and extraction of metals from their compounds or naturally occurring sources.

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Concept of ionic theory

The ionic theory was proposed by Svante Arrhenius to explain that during electrolysis, ionic substances (*electrolytes)* dissociate in solution to give randomly moving charged particles called *ions*. These ions are either *positively* or *negatively* charged, and are referred to as *cations* or *anions,* respectively. The substances that conduct electricity have mobile electrically charged particles. The movements of the charged particles enable the materials to conduct electricity. Some solid substances such as metals allow electric current to pass through them, while other solids such as non-metals and salts in solid state, do not conduct electricity. Most salts conduct electricity when are in solution or molten state. The substances that conduct electricity are referred to as *conductors,* and those which do not conduct electricity are known as *non-conductors.*

Activity 6.1



Aim: To investigate the materials that conduct electricity

Requirements: Strips of metals such as copper or lead, pieces of dry wood, graphite rods, molten wax, crucible, Bunsen burner, 1.5-V cell, 1.5-V bulb, bulb holder, connecting wires, beakers,

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solutions of salts such as sodium chloride, and copper sulphate, crystalline substances such as calcium nitrate and lead bromide, sugar, liquids such as distilled water, ethanol, and paraffin

Procedure

1. Use copper wires, a bulb, and a cell to form an open circuit as shown in Figure 6.1.

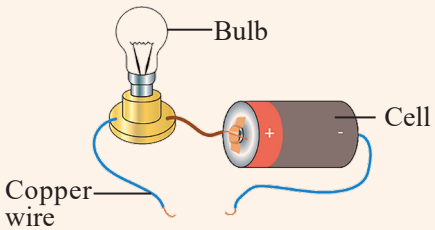


Figure 6.1: *Open circuit*

2. Close the gap in the circuit as shown in Figure 6.2 by using a solid material such as a dry wood and strips of metal, one at a time, and in each case observe what happens to the bulb.

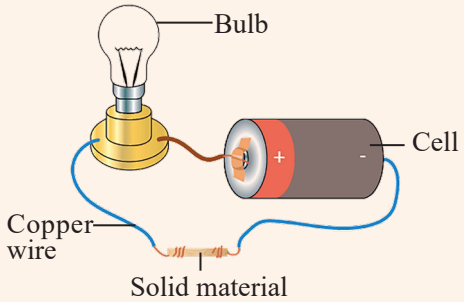


Figure 6.2: *Circuit closed by using a solid material*

3. Replace the strips of metal or dry wood with graphite rods and dip them in a beaker containing salt solutions one at a time as shown in Figure 6.3. Use a clean beaker for each solution. Observe what happens to the bulb.

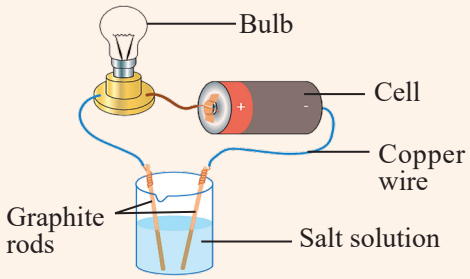


Figure 6.3: *Circuit closed by using a salt solution*

4. Remove the graphite rods and clean them thoroughly by using distilled water.
5. Repeat Step 3 by using the liquids one at a time. Use a clean beaker for each liquid. Observe what happens to the bulb.
6. Repeat Step 3 by using molten wax and crystalls of lead(II) bromide and calcium nitrate one at a time. Record the observations.
7. Heat the sugar in the crucible until it melts, and connect its molten form to the circuit using graphite rods. Record the observations.
8. Repeat Step 7 by using lead(II) bromide and calcium nitrate one at a time. Record the observations. **Note:** Lead(II) bromide will exist in molten form when it is still hot.
9. Record the results as shown in Table 6.1. Put a tick (√) if the bulb lights, or a cross (x) if the bulb does not light.

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Table 6.1: *Results for Activity 6.1*

Substances	Bulb lights
Dry wood	
Metal (copper, lead)	
Sodium chloride solution	
Copper sulphate solution	
Paraffin	
Ethanol	
Water	
Lead(II) bromide crystals	
Calcium nitrate crystals	
Molten wax	
Molten sugar	
Molten lead(II) bromide	
Molten calcium nitrate	

Questions

1. What is the function of the bulb in the electric circuit?
2. Which substances are good conductors of electricity?
3. What features are commonly present in good conductors? Explain.
4. Which substances do not conduct electricity in a solid state, but conduct electricity when in molten state? Give reasons to support your answer.

Electrolytes

An *electrolyte* is a substance which dissociates into free ions when in solution or molten state, thus allowing an electric current to pass through it. Electrolytes are either *strong* or *weak*. *Strong electrolytes* dissociate completely in aqueous solution, whereas *weak electrolytes* dissociate partially. The common electrolytes are solutions of acids, bases, and salts and their respective molten or fused states.

Dissociation of strong electrolytes

The dissociation of strong electrolytes is represented by a fully headed arrow (⟶) to indicate that the whole substance has dissociated into its respective ions.

Dissociation of salts

Salts such as sodium chloride, lead(II) nitrate, and ammonium sulphate dissociate completely in aqueous solutions. The following examples indicate how salts dissociate into their free ions:

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(a) Sodium chloride dissociates in an aqueous solution into free ions of sodium (positively charged) and chloride (negatively charged).

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

(b) Lead(II) nitrate dissociates in an aqueous solution into free positively charged lead and negatively charged nitrate ions.

$$Pb(NO_3)_2(s) \xrightarrow{H_2O} Pb^{2+}(aq) + 2NO_3^-(aq)$$

(c) Ammonium sulphate dissociates in an aqueous solution into free positively charged ammonium and negatively charged sulphate ions.

$$(NH_4)_2SO_4(s) \xrightarrow{H_2O} 2NH_4^+(aq) + SO_4^{2-}(aq)$$

Dissociation of strong acids

Strong acids such as sulphuric acid, nitric acid, hydrochloric acid, and phosphoric acid, dissociate in aqueous solutions into free ions. The following are examples on how these acids dissociate into their free ions:

(a) Sulphuric acid dissociates in an aqueous solution into free positively charged hydrogen and negatively charged sulphate ions:

$$H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$$

(b) Nitric acid dissociates in an aqueous solution into free positively charged hydrogen and negatively charged nitrate ions.

$$HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$$

(c) Hydrochloric acid dissociates in an aqueous solution into free positively charged hydrogen and negatively charged chloride ions.

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

(d) Phosphoric acid dissociates in an aqueous solution into free positively charged hydrogen and negatively charged phosphate ions.

$$H_3PO_4(aq) \longrightarrow 3H^+(aq) + PO_4^{3-}(aq)$$

Dissociation of strong alkalis

Strong alkalis dissociate into their respective charged ions when in aqueous solution. The following are examples of dissociations of strong alkalis:

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(a) Sodium hydroxide dissociates in an aqueous solution into free positively charged sodium and negatively charged hydroxide ions.

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

(b) Potassium hydroxide dissociates in an aqueous solution into free positively charged potassium and negatively charged hydroxide ions.

$$KOH(s) \xrightarrow{H_2O} K^+(aq) + OH^-(aq)$$

Dissociation of weak electrolytes

The dissociation of weak electrolytes is represented by two half headed arrows pointing in opposite directions ($\rightleftharpoons$) to indicate that part of the substance has dissociated into its respective ions, and the maximum dissociation has reached an equilibrium concentration.

Dissociation of weak acids

Weak acids are weak electrolytes that dissociate partially in aqueous solutions. The solution contains both ions and molecules of an electrolyte. The following are examples of the disociation of some weak acids:

(a) Ethanoic acid (acetic acid) dissociates partially in an aqueous solution to give positively charged hydrogen and negatively charged ethanoate (acetate) ions.

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

(b) Ethane-1,2-dioic acid (oxalic acid) dissociates partially in an aqueous solution to give free positively charged hydrogen and negatively charged oxalate ions.

$$H_2C_2O_4(s) \xrightleftharpoons{H_2O} 2H^+(aq) + C_2O_4^{2-}(aq)$$

(c) Carbonic acid dissociates partially in an aqueous solution into free positively charged hydrogen and negatively charged carbonate ions.

$$H_2CO_3(aq) \rightleftharpoons 2H^+(aq) + CO_3^{2-}(aq)$$

Dissociation of weak bases

Bases which form weak electrolytes dissociate partially in aqueous solutions into free positively charged ions and negatively charged ions. For example, ammonium hydroxide partially dissociates into free ammonium and hydroxide ions in an aqueous solution.

$$NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

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Non-electrolytes

A *non-electrolyte* is a substance that does not conduct an electric current in either aqueous solution or in the molten state. When non-electrolytes dissolve in water, they do not produce ions. Examples of non-electrolytes are cyclohexane, sugar, cooking oil, pure alcohols, kerosene, gasoline, molten wax, and diesel.

Activity 6.2



Aim: To identify weak electrolytes, strong electrolytes, and non-electrolytes

Requirements: Battery, connecting wires, two crocodile clips, 100-cm^3 beaker, 2.5-V bulb, bulb holder, spatula, switch, 25-cm^3 measuring cylinder, two graphite rods, common salt, sugar, potassium iodide solution, saturated lead(II) bromide solution, dilute ethanoic acid, dilute sulphuric acid, aqueous ammonia, distilled water, ethanol, and kerosene

Procedure

1. Prepare separate solutions of the solids by dissolving two spatulafuls of each solid to make about 25 cm^3 of solution.
2. Using the solutions one at a time, set the apparatus as shown in Figure 6.4.
3. Close and open the switch, observe and record the light intensity of the bulb.

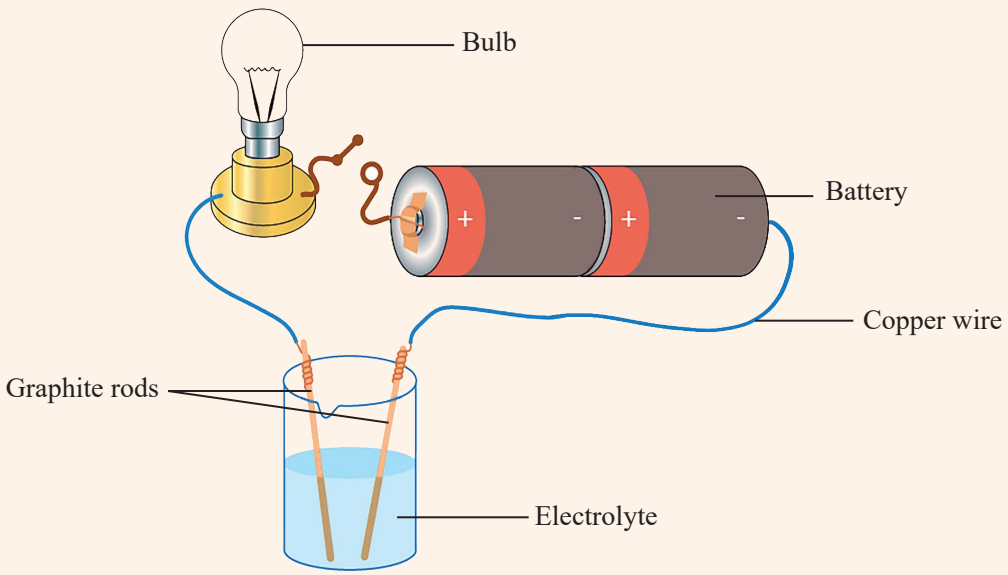


Figure 6.4: *Set-up to distinguish strong and weak electrolytes*

4. Repeat Steps 2−3 by using the other solutions and liquids, one at a time. Other solutions and liquids apart from those listed can also be used.

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5. Record the observations as indicated in Table 6.2. Put a tick (√) in the appropriate column to show whether the bulb lights brightly, dimly or does not light at all.

Table 6.2: *Results on effect of electrolytes and non-electrolytes on bulb lighting*

Solution/liquid	Bright light	Dim light	No light
Common salt solution			
Dilute sulphuric acid			
Ethanol			
Kerosene			
Sugar syrup			
Saturated lead bromide			
Potassium iodide			
Dilute ethanoic acid			
Dilute ammonia			
Distilled water			

Questions

1. What did the results indicate for the substances in which the bulb;
   (a) lighted brightly?
   (b) lighted dimly?
   (c) had no light?
2. Based on the light intensity, group the substances used in this experiment as strong electrolytes, weak electrolytes, or non-electrolytes.

Note: A substance that is an electrolyte in one solvent may become a non-electrolyte in another solvent. For example, when hydrogen chloride gas is dissolved in water, it becomes a strong electrolyte, while in an organic solvent such as methylbenzene, it is a non-electrolyte.

Electrolysis

Electrolysis is a process in which an electric current is passed through an electrolyte to produce a chemical reaction. In electrolysis, an external power source supplies the energy required to drive a chemical reaction. Therefore, *electrolysis* is referred to as a process in which electrical energy is used to cause a *non-spontaneous* chemical reaction to take place. A non-spontaneous reaction is a type of a reaction which requires an

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external source of energy for it to take place. The main components required to achieve electrolysis are an electrolyte, electrodes, and an external power source. When a direct electric current is passed through an electrolyte such as a molten salt or an aqueous salt, acidic or basic solutions, an is electron transfer takes place. The electron transfer may be between two different or same chemical species. Chemical reactions involving the transfer of electrons from one atom to another are called *Reduction-Oxidation* (*Redox*) reactions.

Electrolytic cells

The apparatus in which electrolysis is carried out is known as an *electrolytic cell* (*coulometer* or *voltameter*). A diagram to represent a simple electrolytic cell containing an electrolyte is shown in Figure 6.5.

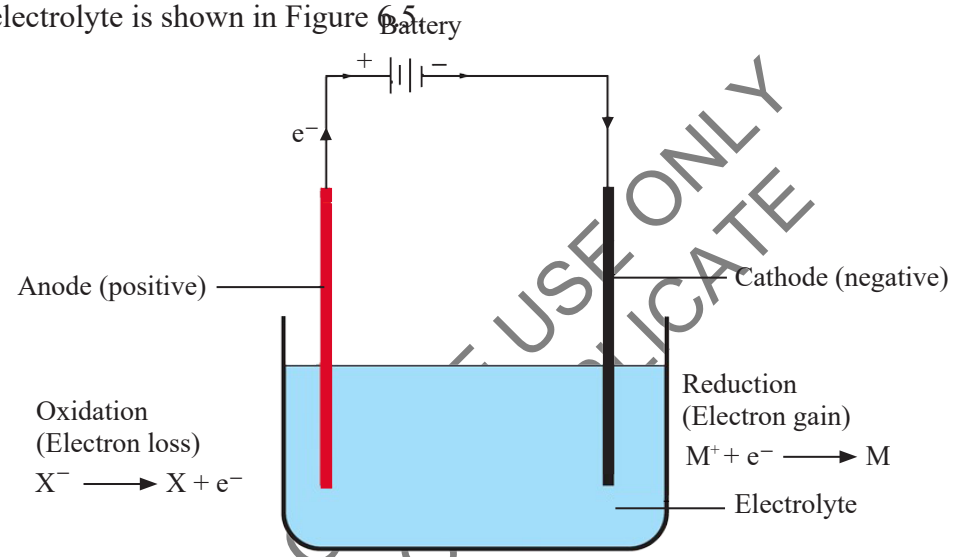


Figure 6.5: *Sketch of a simple electrolytic cell*

An electrolytic cell consists of three important components, which are; battery (source of direct electric current), electrolytes, and electrodes. *Electrodes* are rods through which electrons enter or leave the electrolyte. The electrode which is attached to a negatively charged terminal of the battery is called *cathode* (-). Therefore, cathode is a negatively charged electrode. Through the cathode, electrons flow towards the device or electrolyte and reduction occurs. *Reduction* in this context is referred to as the addition of electrons that leads to a decrease in oxidation state. Generally, a *cathode* is the negatively charged electrode at which the conventional electric current leaves the device or electrolyte.

The electrode which is attached to the positively charged terminal of the battery is called *anode* (+). Therefore, *anode* is a positively charged electrode. At the anode is where the electrons leave the device or electrolyte and oxidation takes place. *Oxidation* in this context is referred to as the removal of electrons that leads to an increase in oxidation

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state. Generally, an *anode* is a positive electrode at which the conventional electric current flows into the device or electrolyte. Most of the anode electrodes commonly used are inert in nature implying that, they can not react or undergo any changes during electrolysis. Examples of inert electrodes include carbon or graphite, gold, and platinum. Reactive electrodes such as copper, zinc, lead, and silver may dissolve or increase in size due to the deposition of the ionic species.

Concentration and strength of an electrolyte

The strength of an electrolyte is not the same as its concentration. The strength of an electrolyte depends on how the substance dissociates when dissolved in a solvent. If a substance dissociates completely, then the electrolyte is *strong,* but if it dissociates partially, then the electrolyte is *weak*. However, the *concentration of an electrolyte* is the amount of solute per unit volume of a solution. When one mole of a substance is dissolved to make one litre of a solution (electrolyte), the concentration of the electrolyte is said to be *one molar solution* written as 1 M. Thus, different electrolytes may have the same concentration but different strengths due to their differences in dissociation abilities. On the other hand, electrolytes can have similar strengths but different concentrations. Table 6.3 shows the comparison of strengths and concentrations for different electrolytes.

Table 6.3: *Concentrations and strengths of electrolytes*

Electrolytes	Concentrations	Strengths
6 M hydrochloric acid 6 M ethanoic acid	Both have the same concentration.	Hydrochloric acid is a strong electrolyte, while ethanoic acid is a weak electrolyte.
6 M ethanoic acid 1 M hydrochloric acid	The ethanoic acid is more concentrated than hydrochloric acid.	Ethanoic acid is a weak electrolyte, while hydrochloric acid is a strong electrolyte.
4 M hydrochloric acid 2 M sulphuric acid	The hydrochloric acid is more concentrated than sulphuric acid.	Both are strong electrolytes regardless of the differences in concentrations.
2 M ethanoic acid 2 M aqueous ammonia	The two electrolytes have the same concentration.	Both are weak electrolytes regardless of the same concentration.
2 M ethanoic acid 1 M aqueous ammonia	The ethanoic acid is more concentrated than aqueous ammonia	Both are weak electrolytes regardless of their differences in concentrations.

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Exercise 6.1

1. Explain the meaning of the following terms:
   (a) Electrolysis
   (b) Electrolytes
   (c) Cations
   (d) Anions
   (e) Concentration of an electrolyte
   (f) Strength of an electrolyte
2. Give difference between the following pairs of terms as applied in electrolysis:
   (a) Electrolytes and non-electrolytes
   (b) Cathode and anode
   (c) Strong electrolytes and weak electrolytes
   (d) Electrolytic cell and a battery
   (e) Reduction and oxidation
3. A dilute electrolyte is not necessarily weak and a concentrated electrolyte is not necessarily strong. Explain.

The mechanism of electrolysis

When an electric current is passed through an electrolyte in an electrolytic cell, the free ions move towards their oppositely charged electrodes. The movement of ions towards their oppositely charged electrodes, either cathode or anode is called *ionic migration.*

Ionic migration during electrolysis

The term *ion* is referred to as a charged particle which is either a *cation* or an *anion*. During electrolysis, positively charged ions migrate towards the negative electrode (cathode), while negatively charged ions migrate towards the positive electrode (anode) (Figure 6.6).

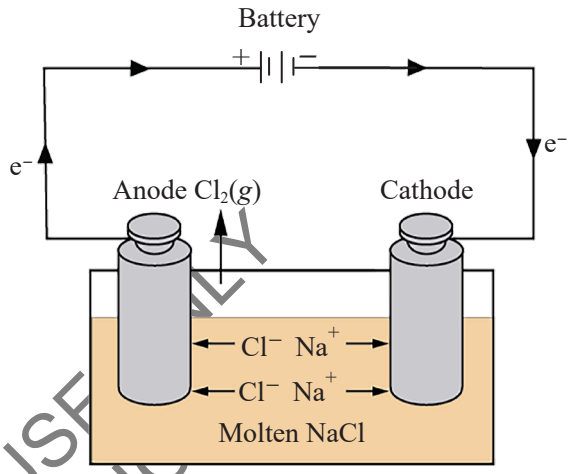


Figure 6.6: *Migration of ions during electrolysis of molten sodium chloride*

The migration of ions to the respective electrode is purely a physical change. On arrival at their respective electrodes, the ions undergo chemical changes, whereby ions are discharged into pure metals or gases.

Electrode reactions

The external battery maintains a potential difference between the two electrodes; cathode and anode. These two oppositely charged electrodes cause the migration of ions towards their opposite charges.

(a) Reactions at the anode

Anions carry the negative charge due to the extra electrons they have. On arrival at the surface of the anode, anions surrender

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the extra electrons to the electrode, thus become discharged. In other words, anions lose electrons leading to an increase in oxidation state, a process referred to as oxidation. The general equation for the reaction of the univalent or monovalent anion at the anode is given by:

$$X^- \longrightarrow X + e^-$$

For example, two monovalent anions such as chloride ions, each discharges one electron to form one molecule of chlorine gas as shown in the following equation:

$$2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$$

A general equation for a divalent anion is:

$$X^{2-} \longrightarrow X + 2e^-$$

For example, two divalent anions such as the oxide ions, each discharges two electrons to form one molecule of oxygen gas as shown in the following equation:

$$2O^{2-}(aq) \longrightarrow O_2(g) + 4e^-$$

(b) Reactions at the cathode

Cations are positively charged because of electron deficiency. On arrival at the surface of the cathode, cations receive electron(s) from the electrode, thus become discharged. In other words, cations gain electron(s) leading to a decrease in oxidation state, a process referred to as reduction. The reactions involved are:

Univalent or monovalent cations:

$$M^+ + e^- \longrightarrow M$$

Divalent cations:

$$M^{2+} + 2e^- \longrightarrow M$$

For example, one monovalent cation such as sodium ion gains one electron to form sodium atom.

$$Na^+(aq) + e^- \longrightarrow Na(s)$$

One divalent cation such as copper ion gains two electrons to form copper atom.

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

During electrolysis, reduction and oxidation reactions take place in a complete electrolytic cell. The two reactions are abbreviated as *redox reactions*. The oxidation

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half-reaction occurs at anode placed on the left hand side, whereas reduction occurs at the cathode on the right hand side of the electrolytic cell.

Example 6.1

The crystals of lead(II) nitrate and potassium iodide are used to show the migration of ions. The set-up for the experiment is shown in Figure 6.7. With the aid of the balanced ionic equations, explain the migration of ions when the switch is closed.

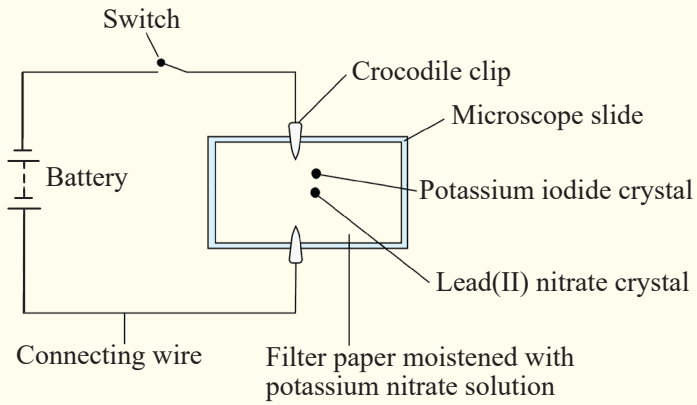


Figure 6.7: *Set-up to illustrate the movements of ions during electrolysis*

Solution

Potassium iodide contains potassium (K^+) and iodide (I^-) ions which are colourless. Lead(II) nitrate contains lead(II) (Pb^{2+}) and nitrate (NO_3^-) ions which are also colourless. When the switch is closed, K^+ and Pb^{2+} move towards the cathode, while I^- and NO_3^- move towards the anode. The formation of a yellow colour is an evidence of the migration of ions in solution. This occurs when $Pb^{2+}(aq)$ and $I^-(aq)$ ions meet.

$$Pb^{2+}(aq) + 2I^-(aq) \longrightarrow PbI_2(s)$$

Example 6.2

Crystals of copper(II) sulphate and potassium chromate(VI) are placed in an electrolytic cell that consists of inert electrodes and potassium nitrate solution. Explain the movements of the coloured ions in the two different terminals of the battery when the switch of the cell is closed.

Solution

Copper(II) sulphate is blue in colour which is caused by copper(II) ions (Cu^{2+}). Potassium chromate(VI) is yellow due to the colour of chromate(VI) ions (CrO_4^{2-}). When crystals of copper(II) sulphate and potassium chromate(VI) are placed in an

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electrolytic cell containing potassium nitrate solution, a small amount of each substance dissolves. A blue ring forms around the copper(II) sulphate, while a yellow ring forms around the potassium chromate(VI). Copper(II) ions move towards the cathode, while chromate(VI) ions move towards the anode. This leaves the copper(II) sulphate solution with an excess of negative ions (SO_4^{2-}), while the potassium chromate(VI) solution is left with excess positive ions (K^+), thus the solutions are no longer electrically neutral. A saturated solution of potassium nitrate helps to restore the neutrality by dissociating into free ions as shown in the following equation:

$$KNO_3(aq) \longrightarrow K^+(aq) + NO_3^-(aq)$$

The negative ions (NO_3^-) move into the potassium chromate(VI) solution to balance the excess potassium ions. Similarly, the positive ions (K^+) move into the copper(II) sulphate solution to balance the excess sulphate ions. Without a saturated potassium nitrate solution, the reactions that take place in the 'cell' would stop almost as soon as they start.

Activity 6.3

Aim: To investigate the migration of ions during electrolysis

Requirements: Distilled water, 2 M Na_2SO_4, U- tube, graphite electrodes, 12-V battery, universal indicator or litmus solution, connecting wires, and retort stand

Procedure

1. Pour a little amount of water into a U-tube to about a half.
2. Add about 2 cm^3 of litmus solution or universal indicator, shake the mixture.
3. Add about 2 cm^3 of sodium sulphate solution.
4. Set the apparatus as shown in Figure 6.8.

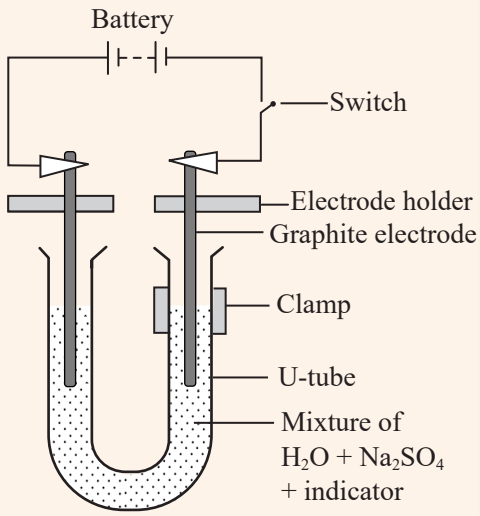


Figure 6.8: *Open circuit for investigating migration of ions during electrolysis of water*

5. Close the switch as shown in Figure 6.9 to allow electric current to pass through the mixture.

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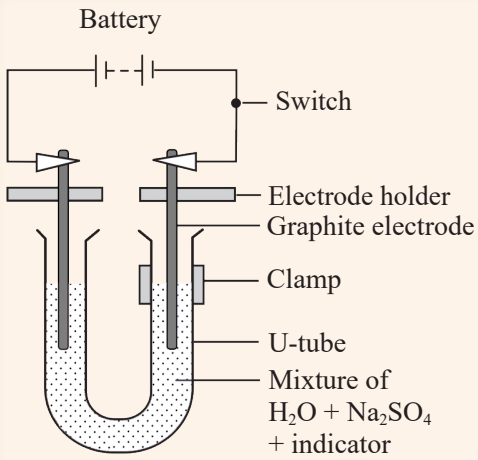


Figure 6.9: *Closed circuit for investigating migration of ions during electrolysis of water*

Questions

1. What colour is observed before the switch is closed?
2. What colours are observed after the switch is closed?
3. Explain the movements of ions towards the graphite electrodes.
4. Write the reactions taking place at each electrode.

Preferential discharge of ions during electrolysis

In addition to the ions of a solute, an aqueous solution of an electrolyte contains hydrogen (H^+) and hydroxide (OH^-) ions from water. The hydrogen ions together with the cations of the electrolyte migrate to the cathode, while the hydroxide ions and the anions of the electrolyte migrate to the anode. At the electrode, some of the ions gain or lose electrons. Only one type of the ions is discharged at each electrode.

When two or more ions of similar charge are present under similar conditions in a solution, for example H^+ and Na^+, or OH^- and SO_4^{2-}, one is preferentially selected for discharge. However, the selection of the ions for discharge depends on factors such as *the position of ion in the electrochemical series*, *concentration of ions in a solution*, and *the nature of the electrode*.

Position of ion in the electrochemical series

For metals and hydrogen, the ease of discharge of an ion is determined by the position of the ion in the electrochemical series shown in Figure 6.10. The *electrochemical series* is an arrangement of elements depending on their ease in gaining electrons.

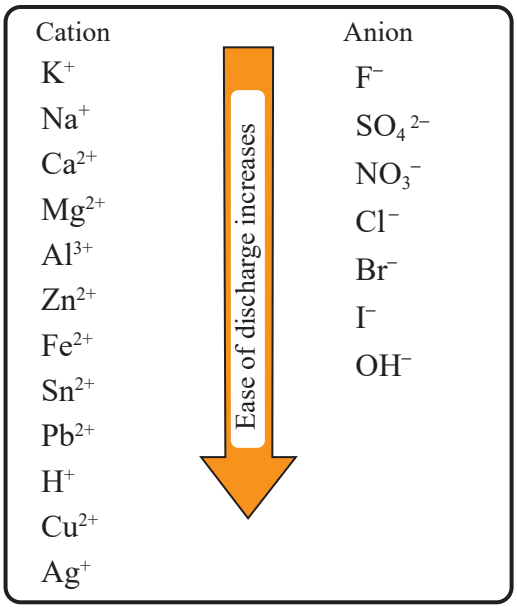


Figure 6.10: *The electrochemical series*

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The ions of the elements high in the electrochemical series are not readily discharged because their ions are very stable in solution. For instance, in the electrolysis of a solution of copper(II) sulphate, the ions present in the solution are Cu^{2+}, SO_4^{2-}, H^+, and OH^-. The Cu^{2+} and H^+ ions migrate to the cathode, while SO_4^{2-} and OH^- migrate to the anode. At the cathode, the copper ions are discharged in preference to the hydrogen ions because copper is lower than hydrogen in the electrochemical series, thus copper metal is deposited at the cathode.

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

At the anode, hydroxide ions are discharged in preference to sulphate ions because hydroxide is lower than sulphate in the electrochemical series, thus oxygen and water are formed at the anode.

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

The overall reaction equation in the entire cell is obtained by adding the two half-cell reactions and balancing their oxidation-reduction (redox) electrons as shown in the following equations:

$$+ \left[\begin{array}{l} 2Cu^{2+}(aq) + 4e^- \longrightarrow 2Cu(s) \\ 4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^- \end{array}\right.$$

Overall equation $2Cu^{2+}(aq) + 4OH^-(aq) \longrightarrow 2Cu(s) + 2H_2O(l) + O_2(g)$

In a redox reaction, the number of electrons lost must always be equal to the number of electrons gained. In this case, four electrons are produced at the anode for every mole of oxygen gas evolved. These electrons are used to discharge two moles of copper ions at the cathode.

Concentration of ions

A high concentration of ions in a solution favours their discharge. For example, in the electrolysis of a concentrated solution of hydrochloric acid, the anions present are hydroxide ions (OH^-) from water and chloride ions (Cl^-) from concentrated hydrochloric acid. These ions migrate to the anode.

Considering the preferential discharge of the two types of ions, OH^- should be discharged in preference to Cl^- since OH^- is below Cl^- in the electrochemical series. However, Cl^- is discharged in preference to OH^- because of the high concentration of Cl^- in the solution, thus chlorine gas is evolved instead of oxygen gas as shown in the following equation:

$$2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$$

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Nature of the electrodes

The nature of the electrodes, especially the cathode, determines which ions are discharged. For example, when sodium chloride solution is electrolysed by using a mercury cathode, a product formed is different from the one that is formed when a platinum cathode is used. Both Na^+ and H^+ ions present in sodium chloride solution migrate to the cathode. When a mercury cathode is used, sodium ions (Na^+) are discharged to give sodium metal which readily dissolves in mercury to form sodium amalgam, a solution of sodium in mercury.

$$Na^+(aq) + e^- \longrightarrow Na(s)$$

$$Na(s) + Hg(l) \longrightarrow \underset{\text{sodium amalgam}}{HgNa(l)}$$

If platinum is used, hydrogen ions (H^+) are discharged to give hydrogen gas.

$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

Note: Mercury electrodes are currently prohibited or restricted because of their toxicity to human beings and the environment.

Electrolysis of selected electrolytes

The products that are formed during electrolysis of aqueous solutions depend on the ions discharged during electrolysis. The ions discharged in turn depend on the three factors described in the previous section. Dilute sodium chloride, sodium hydroxide, copper(II) sulphate, potassium iodide, and sulphuric acid are the selected electrolytes to be discussed in this section.

Electrolysis of dilute sodium chloride solution

In electrolysis of a dilute sodium chloride solution, the cations present are H^+ from water and Na^+ from sodium chloride. These ions migrate to the cathode. Both OH^- and Cl^- are present in the solution. The two anions migrate to the anode.

The $H^+(aq)$ is discharged in preference to $Na^+(aq)$ because hydrogen is lower than sodium in the electrochemical series, and the solution is dilute, thus hydrogen gas is formed at the cathode. The OH^- is discharged in preference to Cl^-, thus oxygen gas is formed at the anode. The overall cell equation is presented as:

$$4H^+(aq) + 4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 2H_2(g)$$

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Two molecules of hydrogen gas are formed at the cathode for every one molecule of oxygen gas formed at the anode. The volume of the gas collected at the cathode is therefore twice that collected at the anode.

Note: Since OH^- and H^+ from water are the only ions that are discharged, electrolysis of a dilute sodium chloride solution is simply the electrolysis of water. The undischarged sodium and chloride ions combine to form sodium chloride. The solution therefore becomes more concentrated as the electrolysis proceeds. If the electrolysis continues for some time, the products that are formed at the anode change. As water molecules continue to be eliminated from the solution, the concentration of chloride and sodium ions increases. It reaches a point where chloride ions start to be discharged at the anode in preference to the hydroxide ions due to their high concentration, thus chlorine gas is produced.

Activity 6.4



Aim: To electrolyse a dilute solution of sodium chloride by using carbon (graphite) electrodes

Requirements: Two combustion tubes with corks, retort stand and clamp, switch, voltameter (a simple voltameter improvised from a plastic drinking water bottle), 6-V batteries, connecting wires, two crocodile clips, 2.5-V bulb, bulb holder, measuring cylinder, 1 M sodium chloride solution, and 2 carbon (graphite) rods

Procedure

1. Measure about 250 cm^3 of dilute sodium chloride and transfer it into a voltameter.
2. Set the apparatus as shown in Figure 6.11.

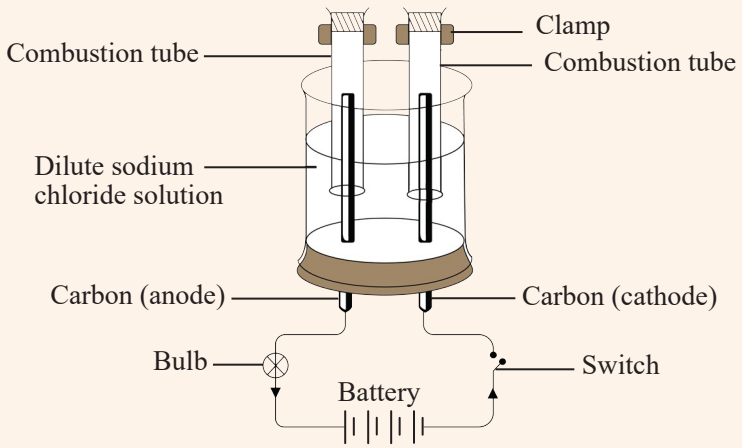


Figure 6.11: *Set-up for the electrolysis of sodium chloride*

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3. Close the switch and leave the reaction to proceed while observing and recording what is happening.

Questions

1. What was observed at each electrode immediately after the switch was closed?
2. Which ions migrated to the cathode and anode, respectively?
3. Write the balanced half equations for the reaction occurring at each electrode.
4. Write the overall cell reaction equation.
5. What will happen to the cell solution when all the H^+ and OH^- ions are completely discharged?

Electrolysis of sodium hydroxide solution

In electrolysis of sodium hydroxide solution, the ions present are Na^+, H^+, and OH^-.

Activity 6.5

Aim: To electrolyse an aqueous solution of sodium hydroxide

Requirements: Two combustion tubes with corks, retort stand and clamp, voltameter (improvised from a plastic drinking water bottle), 6-V batteries, connecting wires, 2 crocodile clips, 2.5-V bulb, bulb holder, measuring cylinder, and 0.1 M sodium hydroxide solution, and 2 carbon (graphite) rods

Procedure

1. Transfer 250 cm^3 of dilute sodium hydroxide solution into the voltameter.
2. Set the electrolytic cell as shown in Figure 6.12.

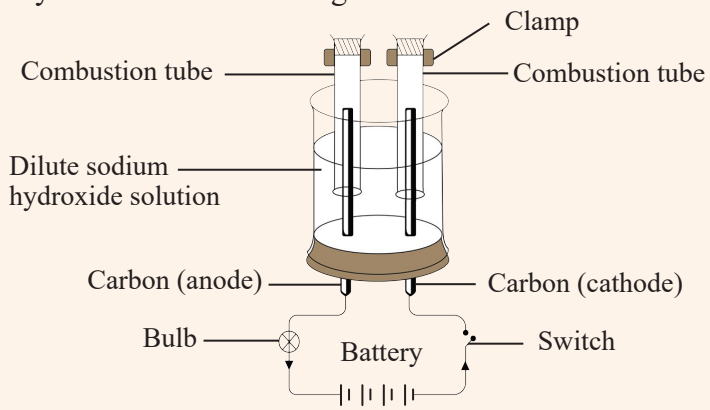


Figure 6.12: *Set-up for the electrolysis of sodium hydroxide solution*

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3. Close the switch, then observe and record what is happening.

Questions

1. What was observed at each electrode immediately after the switch was closed?
2. Which ions migrated to the cathode and anode, respectively?
3. Write the balanced half equations for the reaction occurring at each electrode.
4. Write an overall cell reaction equation.

Electrolysis of copper(II) sulphate solution

The ions present in a copper(II) sulphate solution are H^+ and OH^- from water, and Cu^{2+} and SO_4^{2-} from copper(II) sulphate. The ions that are discharged and the products formed during this electrolysis process depend on the type of electrodes used, that is, either carbon electrode or copper electrode.

Electrolysis by using carbon electrodes

When carbon electrodes are used, copper(II) ions (Cu^{2+}) are discharged at the cathode in preference to hydrogen ions (H^+). This is because copper is below hydrogen in the electrochemical series, thus copper is deposited as a brown coating on the cathode.

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

At the anode, OH^- ions are discharged in preference to SO_4^{2-} ions. This is because hydroxide ions are easier to discharge than sulphate ions, thus oxygen gas is produced at the anode.

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

The overall ionic equation for the reaction is:

$$2Cu^{2+}(aq) + 4OH^-(aq) \longrightarrow 2Cu(s) + 2H_2O(l) + O_2(g)$$

There is no change in the mass of the anode, but the cathode gains mass because of the deposition of copper metal. The concentration of copper(II) ions in the solution decreases. This causes the fading of the blue colour of the solution with time, and finally the solution becomes colourless as water and other colourless components predominate. Hydrogen and sulphate ions are not discharged. They combine to form sulphuric acid. The final colourless solution is therefore the dilute sulphuric acid.

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Electrolysis by using copper electrodes

When copper electrodes are used, the migration of ions is similar to that when carbon electrodes are used. However, when hydroxide and sulphate ions migrate to the anode, none of them is discharged. Instead, the copper anode dissolves to form copper(II) ions.

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

This is because it is easier for the copper anode to ionise than for the hydroxide ions to be discharged. The mass of the anode therefore decreases. At the cathode, copper ions are discharged in preference to hydrogen ions and are deposited as a brown coating on the cathode (Figure 6.13). The mass of the cathode therefore increases, while that of the anode decreases.

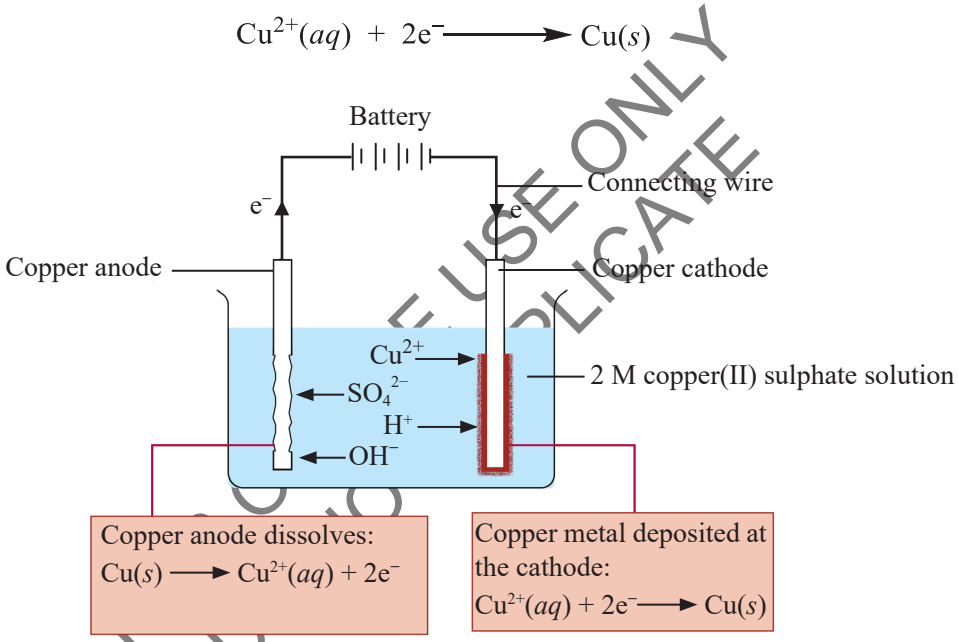


Figure 6.13: *Sketch of the electrolysis of copper sulphate by using copper electrodes*

For every atom of copper that ionises at the anode, one copper ion is discharged at the cathode. The loss in mass of the copper anode is therefore equal to the gain in mass of the copper cathode. The concentration of copper ions in the solution does not change. The intensity of the blue colour of the solution remains the same. This illustrates that the material that is used as the electrode determines the type of the ion that is discharged during electrolysis. If the anode is not an inert material, it may dissolve. Inert electrodes such as carbon and platinum are never affected chemically even when are used as anodes.

Note: The cathode never dissolves regardless of the materials used.

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Activity 6.6



Aim: To electrolyse copper(II) sulphate using different electrodes

Requirements: Beaker, 6-V batteries, analytical balance, copper strings, 2 carbon (graphite) rods, copper strings, connecting wires, 2 crocodile clips, 2.5-V bulb, bulb holder, steel wool or sandpaper, blue and red litmus papers, 2 M copper(II) sulphate solution, propanone, and ethanol

Procedure

1. Transfer about 200 cm^3 of copper(II) sulphate solution into a beaker.
2. Weigh the carbon electrodes and note their masses.
3. Set the electrolytic cell as in Activity 6.5.
4. Connect the cell with the power source.
5. Close the switch and observe how the electrolysis proceeds for about 15 minutes and record the changes observed.
6. Switch off the current and weigh the electrodes.
7. Wash the electrodes by dipping them in distilled water, then in ethanol and finally in propanone.
8. Re-weigh the electrodes again.
9. Test the final solution by using a blue litmus paper.
10. Repeat the experiment by using clean copper strings instead of carbon electrodes and a fresh solution of copper(II) sulphate.

Questions

1. Describe the ionic migration during the electrolysis of copper(II) sulphate solution by using carbon electrodes.
2. What is the pH of the remained electrolyte when both copper and carbon were used as electrodes?
3. Is there any mass change in the electrodes? Explain.

Electrolysis of dilute sulphuric acid

Dilute sulphuric acid contains water; thus, the ions present in the solution are H^+ and SO_4^{2-} from sulphuric acid, and H^+ and OH^- from water. Hydrogen ions migrate to the cathode, while the sulphate and hydroxide ions migrate to the anode.

At the anode, hydroxide ions are discharged in preference to sulphate ions as follows:

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

This is because the hydroxide ions have a greater tendency to lose electrons than

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sulphate ions. The gas collected at the anode is therefore oxygen.
At the cathode, hydrogen ions are the only cationic species present in the solution. These are discharged as hydrogen gas at the cathode.

$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

The volume of hydrogen gas collected is approximately twice that of oxygen. The ratio of the gases collected may not be exactly 1:2 because oxygen is more soluble in water than hydrogen.

The overall ionic equation is the electrolysis of water. Therefore, the overall equation shows that the electrolysis of dilute sulphuric acid is referred to as *electrolysis of water*. Since only water is removed from the solution, dilute sulphuric acid becomes more concentrated with time. The sketch of a Hofmann's voltameter shown in Figure 6.14 is a specifically designed electrolytic cell that is used for the electrolysis of acidified water by using dilute sulphuric acid.

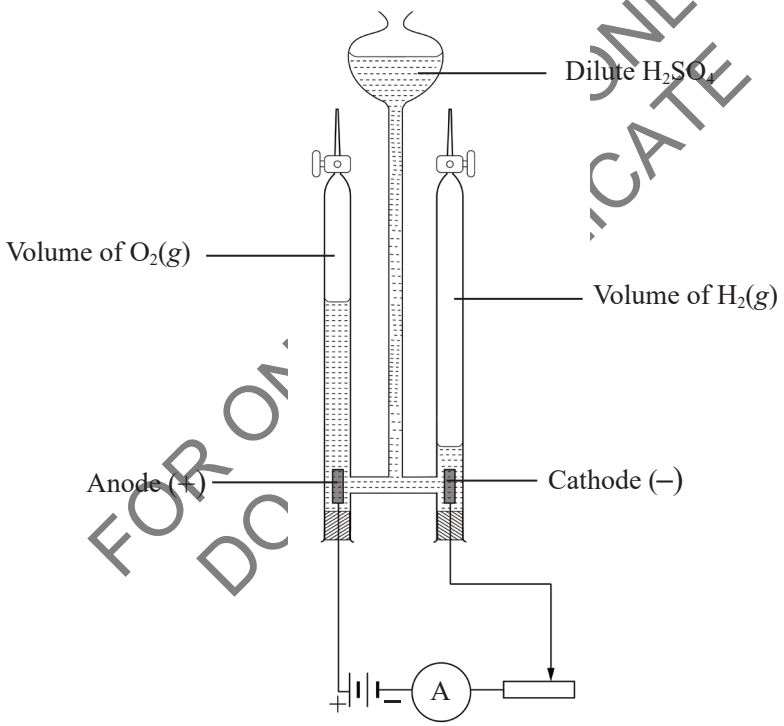


Figure 6.14: *Sketch of a Hofmann's voltameter for electrolysis of dilute sulphuric acid*

Activity 6.7

Aim: To investigate the electrolysis of dilute sulphuric acid by using carbon electrodes

Requirements: Two combustion tubes with corks, retort stand and clamp, dropper, wooden splint, stirring rod, voltameter (improvised from a plastic drinking water bottle), 6-V battery, connecting wires, 2

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crocodile clips, 2.5-V bulb, bulb holder, 2 carbon (graphite) rods, concentrated sulphuric acid, and distilled water

Procedure

1. Add distilled water into an electrolytic cell until it reaches half way.
2. Set the electrolytic cell as in Figure 6.15 and close the switch.

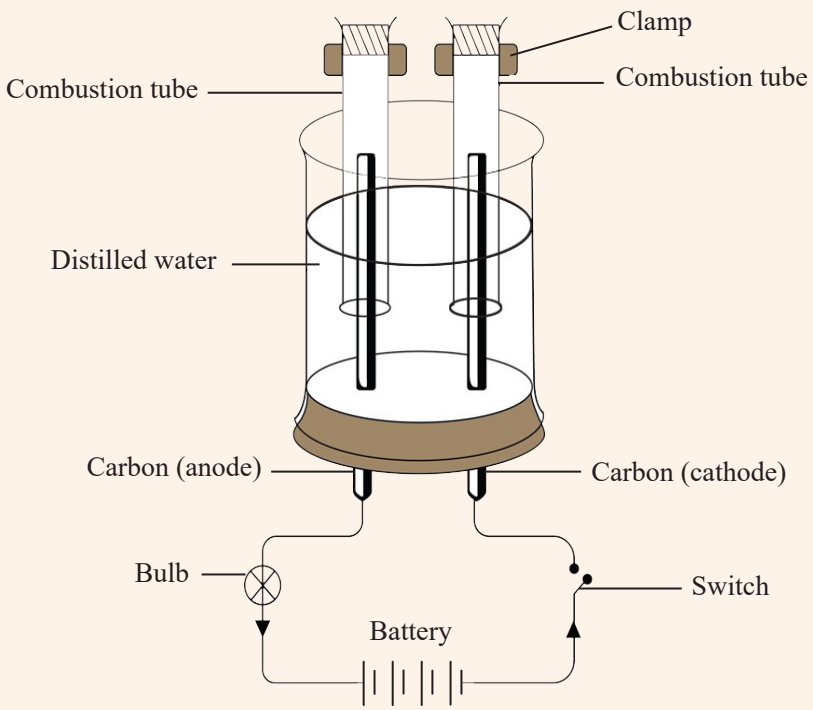


Figure 6.15: *Set-up for the electrolysis of dilute sulphuric acid using carbon electrodes*

3. Add concentrated sulphuric acid into the electrolytic cell dropwise while stirring until the bulb produces a constant amount of light.

   Caution! Be careful when adding concentrated acid into water because the reaction produces heat.

4. Remove the bulb while the switch is still on and continue collecting the gas in each combustion tube.
5. Test the gas collected from the anode tube by using a glowing wooden splint and the one collected from the cathode tube by using a burning wooden splint.
6. Record the results obtained.

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Questions

1. Identify the gas collected at each electrode.
2. Write the ionic equations for the reactions taking place at each electrode and the overall equation. Comment on the volume of the gases collected at each electrode.

Electrolysis of molten potassium iodide

The ions in molten potassium iodide are K^+ and I^-. Potassium ions migrate to the cathode where they are discharged because these are the only cations present. The iodide ions migrate to the anode and become discharged to form iodine since are the only anions present.

Activity 6.8



Aim: To demonstrate the electrolysis of molten lead bromide

Requirements: Crucible, Bunsen burner, tripod stand, wire gauze, connecting wire, crocodile clips, graphite rods, batteries (6 V), switch, and lead bromide crystals

Procedure

1. Place the lead bromide crystals in a crucible.
2. Dip the graphite electrodes in the crystals and connect them with the batteries.
3. Heat the crystals until they melt.
4. After they have melted, close the switch to allow the current to flow for about five minutes.
5. Record your observations.

Questions

1. Describe the ionic migration during the electrolysis of molten lead bromide.
2. Write the ionic equations to describe the reaction at each electrode.

Exercise 6.2



1. Explain why solid sodium chloride is a poor conductor of electricity, while when in molten state, it readily conducts electricity.
2. Show the products of electrolysis of molten sodium chloride by giving a balanced cell equation at each electrode.
3. Describe the products that are obtained during electrolysis of a concentrated sodium chloride solution.

Laws of electrolysis

There is a direct relationship between the amounts of products formed at the electrodes and the quantity of electricity that is passed through the electrolyte. This quantitative aspect of electrolysis was first investigated by Michael Faraday and is summarised into two laws known as *Faraday's Laws of Electrolysis.*

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The two laws of electrolysis help to determine the quantity of a substance liberated during electrolysis. The laws depend on three factors, which are; the time (t) of passing a steady current through an electrolyte, the magnitude of the current (I) supplied, and the charge on the ion of the element.

Faraday's First Law of electrolysis

Faraday's First Law of electrolysis relates the mass of a substance liberated and the quantity of electricity used to liberate it. The law states that, *the mass of a substance liberated at (or dissolved from) an electrode during electrolysis is directly proportional to the quantity of electricity passed through the electrolyte.*

Alternatively, this law is stated as *the mass of a substance produced or dissolved at an electrode during electrolysis is directly proportional to the moles of electrons transferred at that electrode.*

The law is used to calculate the masses of metals such as copper deposited at the cathode by working out the change in mass at the cathode. For example, if the initial mass of copper is 63.680 g and the final mass is 63.705 g, the mass of copper deposited is (63.705 – 63.680) g = 0.025 g

Note: The quantity of electricity is measured in coulumbs (C). One coulombs is the quantity of electric charge that passes through a given point in a circuit when a current of 1 ampere flows for 1 second. The current is measured in amperes (A), and time in seconds (s). Table 6.4 shows sample results for the verification of Faraday's First Law of electrolysis.

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Table 6.4: *Sample results for the verification of the Faraday's First Law of electrolysis*

Current (A)	Time (s)	Quantity of electricity (C)	Change in mass (g)
0.267	300	80	0.025
0.567	300	170	0.049
0.783	300	235	0.074
1.067	300	320	0.099
1.317	300	395	0.124
1.067	450	480	0.149
1.047	525	550	0.173
1.050	600	630	0.198

When a current of 0.267 A is passed for 300 s, the quantity of electric charge (Q) that passes is calculated from

$Q = It$, where I = current and t = time

$Q = 0.267\ \text{A} \times 300\ \text{s}$

$Q = 80.1\ \text{A s}$ or $80.1\ \text{C}$

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When the quantity of electricity passed is doubled, the amount of copper deposited almost doubles. A graph showing the mass of copper deposited (g) against the quantity of electricity passed (C) is a straight line passing through the origin as shown in Figure 6.16. This shows that the mass of copper deposited is directly proportional to the quantity of electricity passed. This is in accordance with the Faraday's First Law of electrolysis.

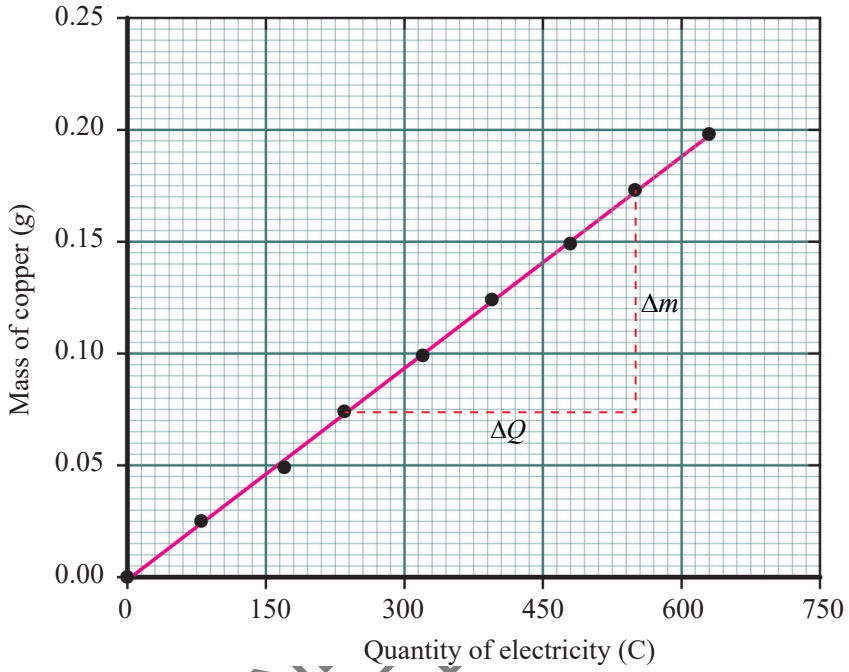


Figure 6.16: *Mass of copper deposited against the quantity of electricity used*

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If m is the mass of a substance deposited, I is the current of electricity passed, and t is the time in seconds, Faraday's First Law of electrolysis is expressed as:

$$m \propto It$$

$$m = ZIt$$

where;

m is the mass of a substance produced or consumed (in grams);

Z is a constant known as *electrochemical equivalent*;

I is the electric current (in amperes);

t is the time taken for deposition (in seconds).

$$\text{Thus; } Z = \frac{m}{It}$$

An *electrochemical equivalent* (ECE) is the mass of a substance produced or consumed when one coulomb of electricity is passed in an electrochemical reaction. It has the units of $g\ C^{-1}$.

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From the formula, ECE is the slope of a graph of the mass deposited (grams) against the quantity of electricity (in coulombs). Using the data for electrode deposition of copper, the slope of the graph of Figure 6.16 is determined as follows:

$$\text{The slope of the graph is} = \frac{\Delta m}{\Delta Q}$$

$$= \frac{(0.173 - 0.074)\ \text{g}}{(550 - 235)\ \text{C}}$$

$$= 3.14 \times 10^{-4}\ \text{g C}^{-1}$$

The actual electrochemical equivalent for copper is 3.29×10^{-4} g C^{-1}. The value obtained in the calculations is a bit different from the actual value because of experimental errors.

The electrochemical equivalents of some elements are given in Table 6.5.

Table 6.5: *Electrochemical equivalents of some elements*

Substance	Electrochemical equivalents (g C^{-1})
Hydrogen	1.045×10^{-4}
Silver	1.118×10^{-3}
Oxygen	8.290×10^{-5}
Copper	3.290×10^{-4}
Gold	6.810×10^{-4}
Chromium	1.800×10^{-4}
Iron	2.894×10^{-4}

Activity 6.9



Aim: To measure the amount of copper deposited during electrolysis of copper chloride solution.

Requirements: 250-mL beaker, 12-V battery, rheostat, ammeter, analytical balance, stopwatch or clock, 2 carbon (graphite) electrodes, connecting wires, 2 crocodile clips, switch, steel wool or sandpaper, 2 M copper(II) chloride solution, propanone, and ethanol

Procedure

1. Transfer 200 cm^3 of copper(II) chloride solution into a beaker.
2. Clean the carbon electrodes with steel wool or sandpaper and label them as cathode and anode.
3. Weigh the cathode before using it.

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4. Set the electrolytic cell as shown in Figure 6.17.

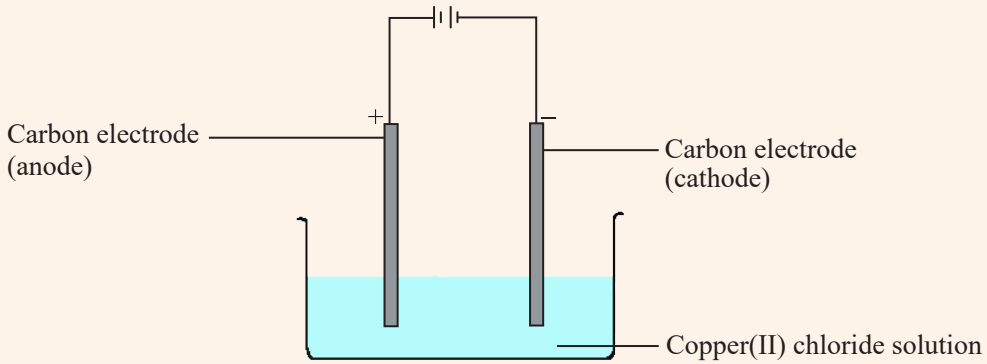


Figure 6.17: *Set-up for the electrolysis of copper(II) chloride*

5. Close the switch and immediately start the stopwatch and adjust the rheostat to a constant current of 0.25 A.
6. Remove the cathode after 5 minutes, dry and weigh it again.
7. Repeat the Steps 1–6 by using higher current values and time shown in Table 6.6.

Table 6.6: *Verification of the Faraday's First Law of electrolysis*

Current (A)	Time (s)	Current × time (A s)	Mass of copper deposited (*m*)
0.25	300		
0.50	300		
0.75	300		
1.00	300		
1.25	300		
1.00	450		
1.00	525		
1.00	600		

Questions

1. Fill the missing data in Table 6.6.
2. Plot a graph of the mass of copper deposited against quantity of electricity passed.
3. Determine the slope of the graph in Question (2) and explain the significance of this value.
4. Copy the electrolytic cell shown in Figure 6.17 and indicate what were observed at the two electrodes.

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Example 6.3

How long should a current of 5 A be passed through a solution of silver chloride in order to deposit 3.24 kg of silver at the cathode? (The electrochemical equivalent of silver = 1.118×10^{-3} g C^{-1}).

Solution

$m = 3.24$ kg, $I = 5$ A, $Z = 1.118 \times 10^{-3}$ g C^{-1}

$m = ZIt$

$t = \frac{m}{ZI}$

$t = \frac{(3.24 \text{ kg} \times 1000 \text{ g/kg})}{1.118 \times 10^{-3} \text{g C}^{-1} \times 5 \text{ A}}$

$= 5.80 \times 10^{5}$ s

Faraday's constant

The amount of electric charge carried by one mole of electrons is known as *Faraday's constant*. The charge of 1 electron is 1.602×10^{-19} C. One mole of electrons contains the Avogadro's number of particles which is equal to 6.022×10^{23} electrons.

Therefore, one mole of electrons $= 1.602 \times 10^{-19}$ C $\times 6.022 \times 10^{23}$ mol^{-1}

$= 96,472.44$ C mol^{-1}.

The charge of one mole of electrons is usually approximated to 96500 C. This is the Faraday's constant. It is possible to determine the Faraday's constant from the results of Activity 6.9. One mole of copper is equivalent to 63.5 g of copper and the electrochemical equivalent of copper is 3.290×10^{-4} g C^{-1}. Therefore, the quantity of electricity required to deposit 1 mole of copper (63.5 g) is:

$$\frac{63.5 \text{ g mol}^{-1}}{3.290 \times 10^{-4} \text{ g C}^{-1}} = 193009.12 \text{ C mol}^{-1}$$

The equation for the deposition of copper is:

$$Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)$$

This means, two moles of electrons are needed to deposit one mole of copper which is 63.5 g of copper atoms or 193009.12 C deposit one mole of copper atoms. The quantity of electricity carried by the two moles of electrons is 193009.12 C. Therefore, 1 mole of electrons carries;

$$\frac{193009.12 \text{ C}}{2} = 96,504.6 \text{ C}$$

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Faraday's constant as a unit of measurement

Faraday's constant (96500 C) is also used as a unit of measurement called faraday. It is abbreviated by the letter F. Therefore, one faraday (F) is the quantity of electric charge carried by one mole of electrons.

Generally, the amount in moles of atoms of an element liberated or deposited at an electrode is equal to:

$$\frac{\text{number of moles of electrons passed}}{\text{number of charges on each ion discharged (valence)}}$$

One faraday discharges one mole of silver ions or one mole of sodium ions.

$$Ag^{+}(aq) + e^{-} \longrightarrow Ag(s)$$

$$Na^{+}(aq) + e^{-} \longrightarrow Na(s)$$

Two faradays are required to discharge one mole of magnesium ions or calcium ions.

$$Mg^{2+}(aq) + 2e^{-} \longrightarrow Mg(s)$$

$$Ca^{2+}(aq) + 2e^{-} \longrightarrow Ca(s)$$

Three faradays are needed to discharge one mole of aluminium ions

$$Al^{3+}(aq) + 3e^{-} \longrightarrow Al(s)$$

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Example 6.4

A current of 3.2 A is passed through a solution of copper(II) sulphate for 30 minutes. What is the mass of copper deposited at the cathode? (Cu = 63.5, S = 32, O =16).

Solution

Data: $I = 3.2$ A, $t = 30$ minutes

$Q = It$

$= 3.2 \text{ A} \times (30 \times 60)$ seconds

$= 5760$ C

The number of moles of electrons;

$$= \frac{5760 \text{ C}}{96500 \text{ C mol}^{-1}}$$

$= 0.0597$ mol

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Since the copper ion (Cu^{2+}) has two charges, the number of moles of copper discharged is given by:

$$\frac{\text{number of moles of electrons passed}}{\text{number of charges on each ion discharged (valence)}}$$

$$= \frac{0.0597 \text{ mol}}{2}$$

$$= 0.0298 \text{ mol}$$

The mass of copper in 0.0298 moles of copper = number of moles × molar mass

But the molar mass of copper is 63.5 g.

Mass of copper deposited = 0.0298 mol × 63.5g mol^{-1}.
= 1.89 g

Alternatively, if the electrochemical equivalent of copper (Z = 3.290×10^{-4} g C^{-1}) is provided, then use the following approach:

$$m = ZIt$$

$$= 3.290 \times 10^{-4} \text{ g C}^{-1} \times (3.2 \text{ A} \times 30 \times 60) \text{ C} = 1.89 \text{ g}$$

Example 6.5

What mass of zinc will be liberated during electrolysis if a current of 15 A is supplied to an electrolytic cell for one and half hours?

Solution

Data: t = 1.5 hours = 90 min × 60 s = 5400 s, I = 15 A

$$m = \frac{A_r It}{VF} \qquad \text{Note: } Z = \frac{A_r}{VF}$$

where;

A_r is the relative atomic mass;
V is the number of moles of electrons required to deposit 1 mole of substance; and
F is the Faraday's constant.

$$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$$

2 moles of electrons are required for deposition of 1 mole of Zn;

$$m = \frac{65.38 \text{ g mol}^{-1} \times 15 \text{ A} \times 5400 \text{ s}}{2 \times 96500 \text{ C mol}^{-1}}$$

Mass of zinc = 27.44 g

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Example 6.6

What volume of oxygen gas measured at STP will be produced during electrolysis by a current of 5 A supplied for 30 minutes?

Solution

Data: $t = 30 \text{ minutes} \times 60 \text{ s} = 1800 \text{ s}$, $I = 5 \text{ A}$

$Q = It$

$= 5 \text{ A} \times 1800 \text{ s} = 9000 \text{ Coulombs}$

The number of moles of electrons is given by:

$$= \frac{9000 \text{ C}}{96500 \text{ C mol}^{-1}}$$

$$= 0.0932 \text{ F}$$

Discharge equation:

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

4 moles of electrons are required to liberate one mole of oxygen as stated in the discharge equation.

1 mole of electrons is equivalent to 1 faraday of charge.

$4 \text{ F} = \text{one mole of oxygen gas}$

$4 \text{ F} = 22.4 \text{ dm}^3 \text{ of } O_2 \text{ at STP}$

$0.09326 \text{ F} = y \text{ dm}^3 \text{ of } O_2$

$$y = \frac{22.4 \text{ dm}^3 \times 0.09326}{4 \text{ F}}$$

$$y = 0.522 \text{ dm}^3$$

The volume of oxygen gas that will be produced at STP is 0.522 dm^3.

Example 6.7

Calculate the current required to produce 20 dm^3 of chlorine gas by electrolysis at STP for one hour?

Solution

Data: Volume of chlorine gas = 20 dm^3, time 1 hr = 60 min × 60 s = 3600 s

The molar mass of chlorine gas is 71 g mol^{-1}.

71 g mol^{-1} of chlorine gas occupy 22.4 dm^3 at STP

x g of chlorine gas will occupy 20 dm^3 at STP

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$$x = \frac{20\ \text{dm}^3 \times 71\ \text{g mol}^{-1}}{22.4\ \text{dm}^3\ \text{mol}^{-1}}$$

$$= 63.39\ \text{g of Cl}_2$$

$$I = \frac{mV\text{F}}{tA_r}$$

From $2\text{Cl}^-(aq) \longrightarrow \text{Cl}_2(g) + 2\text{e}^-$, V is 2

$$I = \frac{63.39\ \text{g} \times 2 \times 96500\ \text{C mol}^{-1}}{3600\ \text{s} \times 71\ \text{g mol}^{-1}}$$

The current will be 47.86 A

Faraday's Second Law of electrolysis

This law relates the mass of deposited substances with their respective chemical equivalents. The consistence of this law is based on the fact that each element has its own equivalent weight. The law states that, *when the same quantity of electricity is passed through solutions of different electrolytes, the mass of the substances liberated or deposited at the electrodes is directly proportional to the chemical equivalents of the substances.*

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That is $\frac{m_1}{m_2} = \frac{\text{E}_1}{\text{E}_2}$

where m_1 and m_2 are the respective masses of substances liberated or deposited at the electrodes, and E_1 and E_2 are the chemical equivalents of the two substances, respectively.

The *chemical equivalent* (E) of a substance is the relative atomic mass (RAM) of the substance divided by the number of electrons required to oxidise or reduce each unit of a substance.

$$\text{Chemical equivalent} = \frac{\text{RAM}}{\text{number of charge(s) on an ion}}$$

Faraday's Second Law of electrolysis can also be stated as; *when the same quantity of electricity is passed through solutions of different electrolytes, the relative numbers of moles of the elements deposited are inversely proportional to the charge on the ions of the respective elements.*

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Activity 6.10

Aim: To verify Faraday's Second Law of electrolysis

Requirements: Two beakers (250-mL), 12-V batteries, rheostat, ammeter, analytical balance, stopwatch or clock, two copper electrodes, two silver electrodes or zinc electrodes, connecting wires, 2 crocodile clips, switch, steel wool or sandpaper, 0.05 M copper(II) sulphate, 0.05 M silver nitrate solution or 0.05 zinc nitrate solution, distilled water, propanone, and ethanol

Procedure

1. Transfer 200 cm^3 of copper(II) sulphate solution and silver nitrate solution or zinc nitrate solution into separate beakers.
2. Clean the copper and silver plates with steel wool or sand paper.
3. Prepare two electrolytic cells depending on the solution, and label the cathode for each of the copper and silver electrodes or zinc electrodes. Weigh the electrodes by using an analytical balance before starting the electrolysis process.
4. Connect the plates to an electrical circuit as indicated in Figure 6.18.

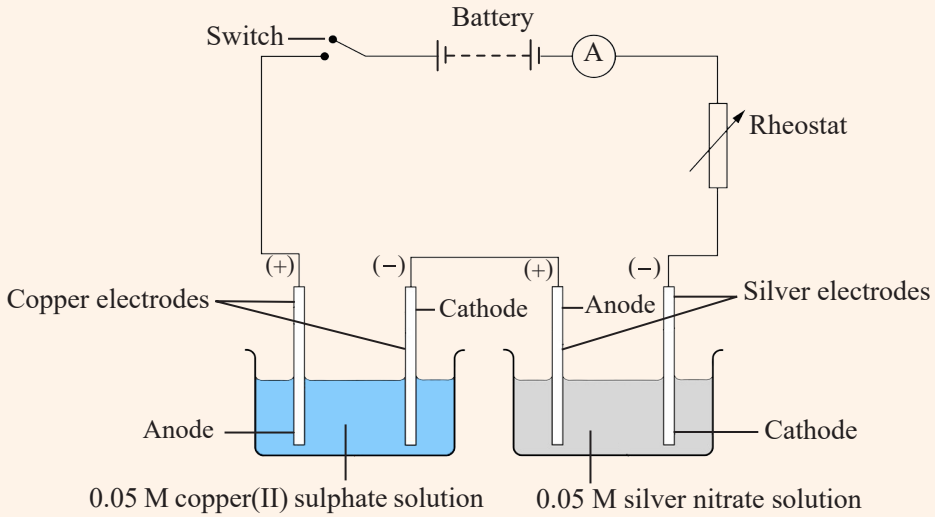


Figure 6.18: *Set-up for the verification of Faraday's Second Law of electrolysis*

5. Close the switch and immediately start the stopwatch. Adjust the rheostat to give a constant current of 0.65 A.
6. Maintain the steady current of 0.65 A for 35 minutes.
7. Switch off the current and carefully remove the electrodes, wash the electrodes by dipping into distilled water, ethanol, and finally in propanone. Leave the electrodes to dry.
8. Weigh the dried electrodes to determine the changes in mass.

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Questions

1. Determine the masses of copper and silver deposited at the electrodes.
2. Calculate the chemical equivalents of copper and silver.

Example 6.8

During purification of copper, a current of 10000 A is passed through an electrolytic cell for 12 hours.

(a) What mass of pure copper is deposited at the cathode?

(b) What mass of aluminium would be deposited if the cell was changed to refine aluminium? (Cu = 63.5, Al = 27, electrochemical equivalent of copper = 3.290×10^{-4} g C^{-1})

Solution

(a) Data

$I = 10000$ A, $t = 12$ hours, $Z = 3.290 \times 10^{-4}$ g C^{-1}

$m = ZIt$

$= 3.290 \times 10^{-4}$ g $C^{-1} \times 10{,}000$ A $\times$ 12 hours $\times$ 60 min $\times$ 60 s.

$= 1.421 \times 10^{5}$ g

$= 142.1$ kg

(b) Chemical equivalent of copper (E_1) $= \frac{63.5}{2} = 31.75$

Chemical equivalent of aluminium (E_2) $= \frac{27}{3} = 9$

Mass of copper (m_1) = 142.1 kg

The mass of aluminium (m_2):

$$\frac{m_1}{m_2} = \frac{31.75}{9} = 3.53$$

$$\frac{142.1 \text{ kg}}{m_2} = 3.53$$

$$m_2 = \frac{142.1 \text{ kg}}{3.53}$$

Mass of aluminium = 40.25 kg.

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Example 6.9

An electric current was passed in series through solutions of calcium chloride and copper(II) sulphate. Carbon electrodes were used in both electrolytes. If 2.5 litres of chlorine gas measured at STP were produced,

(a) what volume of oxygen would also be produced?
(b) what was the mass of copper produced?

Solution

(a) In this question, no data about current or time is given. The whole answer will depend on the analysis of the relevant electrode reactions. The electrode reaction for the discharge of chlorine gas is:

$$2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$$

The equation suggests that 2 faradays liberate one mole of Cl_2 gas.
The number of faradays that will liberate 2.5 dm^3 will be

$$= \frac{2.5\ dm^3}{22.4\ dm^3} \times 2\ F$$

$$= 0.2232\ F$$

Thus, 0.2232 F of electricity were passed through the two solutions in series. Having found the quantity of electricity supplied, find the corresponding volume of oxygen liberated. The discharge equation for the liberation of oxygen is:

$$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$$

The equation suggests that 4 faradays would liberate one mole of oxygen gas at STP which is equivalent to 22.4 dm^3. The volume produced by 0.2232 F will be:

$$= \frac{0.2232\ F}{4\ F} \times 22.4\ dm^3$$

$$= 1.25\ dm^3 \text{ of oxygen gas will be produced.}$$

(b) The discharge equation for the liberation of copper is:

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

The equation suggests that 2 faradays would liberate one mole of copper which is equivalent to 63.5 g of copper. The weight of copper liberated by 0.2232 F will be:

$$= \frac{0.2232\ F}{2\ F} \times 63.5\ g$$

$$= 7.09\ g \text{ of copper.}$$

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Exercise 6.3

1. How many faradays of electricity must be passed to deposit the following masses of metals?
   (a) 2.16 g of silver
   (b) 1.28 g of copper
   (Ag = 108, Cu = 63.5)

2. (a) If you re-plate a chromium spoon, would you make it the anode or the cathode? Explain your answer giving the half reaction chemical equations.
   (b) How many moles of electrons are needed to plate out 0.49 g of chromium(III)? (Cr = 51.99).

3. A current of 1.5 A was passed through 250 cm^3 of 0.5 M $AgNO_3$ aqueous solution for 1 hour and 10 minutes. The anode was platinum, and the cathode was silver.
   (a) Identify the species which carried the current through the solution and indicate their directions of movements.
   (b) Give the equations for the reactions which took place at the anode and cathode.
   (c) Write the changes that would occur if the platinum anode was replaced with a silver anode?

4. During the electrolysis of aqueous hydrochloric acid by using silver electrodes, the mass of the cathode increases by forming a silvery coat on its surface. Explain.

Applications of electrolysis

There are many industrial applications of electrolysis in our daily life. A few of them are discussed in this section.

Production of gases

The commercial production of chlorine gas is done during the manufacture of sodium hydroxide from brine or manufacture of sodium from molten sodium chloride. In both cases, chlorine gas is usually given out as a by-product. Different electrolytic cells such as *diaphragm* or *membrane cells* are used in the production of chlorine gas. In the *diaphragm cell* process, concentrated sodium chloride solution (brine) is electrolysed to produce chlorine gas at the anode, while sodium hydroxide (caustic soda) and hydrogen gas are produced at the cathode. The anode and cathode chambers are separated by a porous diaphragm to prevent hydrogen and sodium hydroxide to react with chlorine. In the *membrane cell*, brine is electrolysed to produce chlorine at the anode, while sodium hydroxide and hydrogen are produced at the cathode. An ion-selective membrane prevents the reaction between chlorine gas with either hydrogen gas or sodium hydroxide.

Purification of copper

Copper is refined through electrolysis of acidified copper(II) sulphate solution. Impure copper is made as the anode. A sheet of pure copper is made as the cathode as shown in Figure 6.19.

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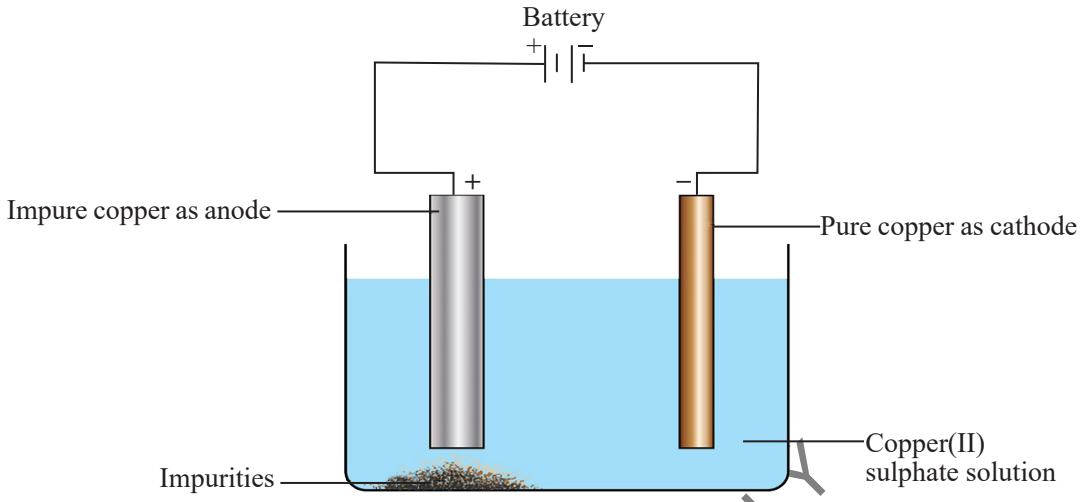


Figure 6.19: *Set-up for the refining of copper by electrolysis*

Impure copper ionises or dissolves to form copper ions which migrate to the cathode where they are reduced and deposited as pure copper on the pure copper plate. Eventually, the whole impure copper block dissolves, and the pure plate becomes larger in size.

The refining process is summarised as follows:
At the anode, impure copper is oxidised to copper(II) ions.

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

At the cathode, copper(II) ions are discharged (reduced) to copper.

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

Electroplating

Electroplating is the electrical deposition of one metal on the other. In electroplating, a cheaper metal is usually coated with a more expensive one. The purpose of electroplating is to protect the metal being electroplated from corrosion or rusting, decorate it or improve its appearance.

Electroplating is used to make ornaments such as rings, earrings, bangles and wrist watches to appear as if they are made of gold or silver. Steel parts of cars, bicycles and some tools appear shiny because they are electroplated with chromium.

During the electroplating process, the anode is usually made of the plating metal while the object to be plated is made the cathode. The electrolyte that is used should be a salt solution of the plating metal. During the electrolysis, the plating metal dissolves and it is transferred to the cathode where the object is being plated. For example, a steel spoon

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can be electroplated with copper, where dilute sulphuric acid is added to copper(II) sulphate to increase its conductivity. The reactions taking place during electrolysis of copper by using copper electrodes are as follows:

At the anode, the copper strip dissolves to form copper(II) ions by losing two moles of electrons.

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^{-}$$

At the cathode, copper is discharged by gaining two electrons and is deposited on the surface of the spoon.

$$Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)$$

In the course of electrolysis, a spoon is rotated through 180° to allow even plating.

Activity 6.11



Aim: To electroplate a steel spoon with copper

Requirements: Two beakers (250 mL), 6-V batteries, rheostat, ammeter, weighing balance, stopwatch or clock, copper strip (each 10 cm by 1 cm), connecting wires, steel spoon, switch, 1 M copper(II) sulphate solution, and dilute sulphuric acid

Procedure

1. Transfer 200 cm^3 of copper(II) sulphate solution into the beaker.
2. Add 2 to 3 drops of dilute sulphuric acid into the beaker.
3. Connect the copper strip with the steel spoon as indicated in Figure 6.20.

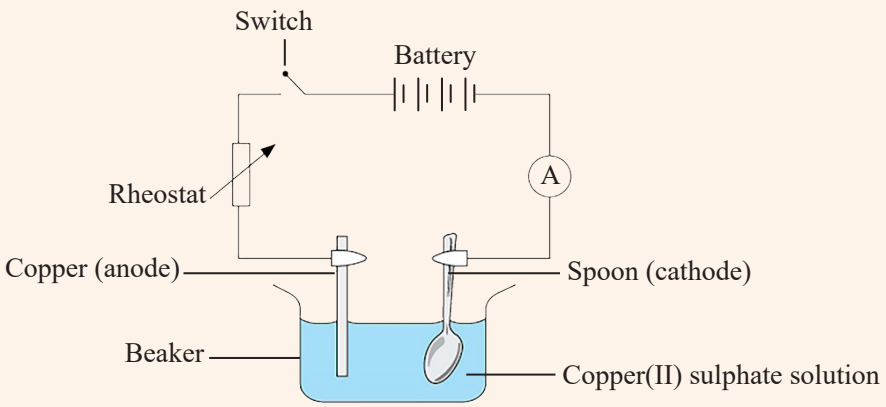


Figure 6.20: *Set-up for electroplating a silver spoon by using copper*

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4. Close the switch, immediately start the stopwatch and adjust the rheostat to give a constant current of 0.05 A.
5. Maintain the steady current of 0.05 A for 5 minutes and switch off the current.
6. Rotate the cathode at 180°, and then switch on the current once more for 15 minutes.

   Record the appearance of the spoon.

Questions

1. What happened to the anode and cathode?
2. What could happen if copper(II) sulphate was replaced with sodium chloride solution?

Chapter summary

1. Conductors of electricity are materials that allow an electric current to pass through.
2. Electrolytes are substances which dissociate into free ions when in molten state or solution form.
3. Electrolytes commonly exist as aqueous solutions of acids, bases, and salts or, molten salts.
4. Strong electrolytes dissociate completely into free ions when in solution or in molten form. These include mineral acids, bases, and ionic salts.
5. Weak electrolytes dissociate partially into free ions when in solution or molten state.
6. The classification of electrolytes is summarised in Figure 6.21.

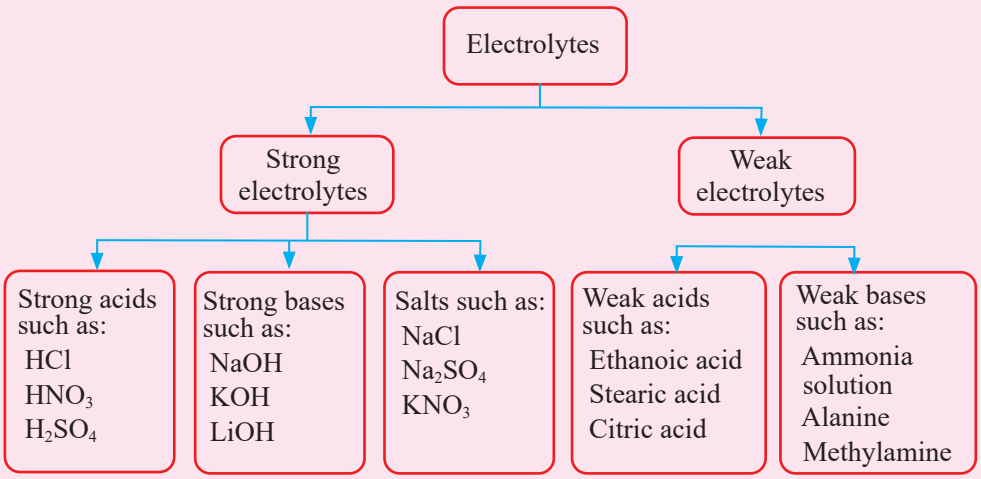


Figure 6.21: *Classification of electrolytes*

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7. Electrolysis is the process through which an electrolyte undergoes a chemical change when an electric current is passed through it.
8. Electrolysis is carried out in an electrolytic cell. The major components of an electrolytic cell are *battery, electrodes,* and *electrolyte*.
9. During electrolysis, cations; the positively charged ions migrate to the cathode, and anions; the negatively charged ions migrate to the anode.
10. The factors that affect the discharge of ions at the electrodes are:
    (a) Position in the electrochemical series (easiness of discharge),
    (b) Concentration of ions in solution, and
    (c) Nature of electrodes.
11. There are two Faradays Laws of electrolysis. The first law relates the quantity of electricity passed in a particular electrolyte to the substances deposited or dissolved. The second law relates the amount of the substances to their equivalent weights.
12. The Faraday's First Law of electrolysis states that, the mass of a substance liberated at (or dissolved from) an electrode during electrolysis is directly proportional to the quantity of electricity passed through the electrolyte.
13. The Faraday's Second Law of electrolysis states that, when the same quantity of electricity is passed through solutions of different electrolytes, the mass of the substances liberated or deposited at the electrodes is directly proportional to the chemical equivalents of the substances.
14. Electrochemical equivalent is the amount in grams of a substance liberated when one mole of electrons is passed through an electrolyte containing that substance.
15. The amount of a substance liberated depends on the number of moles of electrons required to liberate it in a balanced half reaction equation.
16. The Faraday constant is the quantity of electric charge released by one mole of electrons. This is equivalent to 96,500 coulombs.
17. Electrolysis is highly applied in industries. Its applications include the production of gases, purification of metals, and electroplating of objects.

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Revision exercise 6

Choose the correct answer in items 1–4.

1. Which process uses electricity to decompose molten sodium chloride into its components?
   (a) Electroplating
   (b) Electrolysis
   (c) Galvanisation
   (d) Hydrolysis

2. Which particles are produced when electrolytes are in molten state or in solution form?
   (a) Neutrons
   (b) Protons
   (c) Ions
   (d) Molecules

3. One of the following is not an electrolyte:
   (a) A very concentrated sulphuric acid
   (b) Molten aluminium oxide
   (c) Copper sulphate solution
   (d) Sodium chloride solution

4. Which of the following refers to the amount of charge carried by one mole of electrons?
   (a) Faraday
   (b) 96,500 C
   (c) Current
   (d) Ions

5. Write **TRUE** for a correct statement and **FALSE** for an incorrect statement for items (a)-(d).
   (a) An electric charge depends on the quantity of current passed for a particular period of time.
   (b) The charge carried by 1 mole of electrons is approximately equal to 96500 C.
   (c) An aluminium ion has a charge of +3. The passage of 1 mole of electrons through a solution of aluminium will produce 9 g of aluminium.
   (d) The amount of the material produced at an electrode during electrolysis is directly proportional to the quantity of electricity passed through that solution.

6. What would happen if the principles of electrolysis were not discovered up to now?

7. Using three examples in each, briefly explain any three industrial applications of electrolysis.

8. A current of 0.5 A is allowed to flow for six minutes and twenty seconds through two cells which are connected in series. If at the end of the reaction, 0.216 g of silver, and 0.659 g of nickel are deposited at their cathodes, respectively, calculate the following quantities:
   (a) The quantity of electricity passed through the two cells.

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(b) The quantity of electricity needed to deposit 1 g of silver atoms.

(c) The quantity of electricity needed to deposit 1 g of nickel atoms.

(The relative atomic masses of silver and nickel are 108 and 59, respectively).

9. An electric current was passed through the two electrolytic cells, A and B for an hour. Cell A contained a solution of copper(II) sulphate with copper electrodes, while Cell B had a solution of silver nitrate with a silver anode. The gains in mass of the cathodes were 0.189 g and 0.6428 g for copper and silver, respectively. Calculate;
   (a) the quantity of electricity which was used in such experiment.
   (b) the current required in the rheostat during the experiment.
   (c) the chemical equivalent of copper, given that the chemical equivalent of silver is 108.

10. Use the Faraday's First Law of electrolysis to verify the following results:
   (a) A current of 47.86 A is required to produce 20 dm^3 of chlorine gas by electrolysis at STP within one hour.
   (b) When a current of 10,000 A is passed through an electrolytic cell during purifying copper for 12 hours, 142.1 kg of copper will be deposited.
   (The electrochemical equivalent of Cu = 3.290×10^{-4} g C^{-1}).

11. Explain the chemical preference of decorating a copper necklace with silver or gold metals by using electrolysis method.

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Chapter Seven

Chemical kinetics, equilibrium, and energetics

Introduction

In Chapter One, you learnt about the concept of chemical reactions and equations. The chemical reactions proceed at different rates and involve energy. The study of reaction rate is called chemical kinetics. The relative stability of the reactants and products in the chemical reactions is referred to as chemical equilibrium. This state occurs when there is no net change in the amounts of reactants and products. In this chapter, you will learn about the rates of chemical reactions, factors affecting the rates of chemical reactions, reversible and irreversible reactions, chemical equilibrium, and chemical energetics. The competencies developed will enable you to apply the essential principles and theories associated with the rates of chemical reactions, chemical equilibrium, and energy changes in chemical reactions occurring in daily life.

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The rate of a chemical reaction

You might be aware that it takes a short time to cook potato chips than the whole potato of the same weight using the same source of heat. Both phenomena are chemical reactions taking place at different rates.

The rate of a chemical reaction is related with the time taken for the completion of the reaction. Thus, the rate at which a chemical reaction takes place is a measure of the progress of the reaction per unit time. The unit of time may be seconds, minutes, weeks, months or years. In a chemical reaction, the reactants are changed to products and the proportions of both the products and reactants change.

At the start of the reaction (zero time), there is no reaction between the reactants. At this point, the amount of the reactant(s) is in its maximum and there is no product formed. As the reaction progresses, the amounts of the reactants decrease while that of the products increase. At the end of the reaction (if the reaction goes to completion), there will be no reactants left in the reaction mixture. Therefore, the products at this point are in their maximum amounts.

The rate of a chemical reaction is determined by measuring the amount of reactant used up or the amount of product formed during a given time. This can be expressed as:

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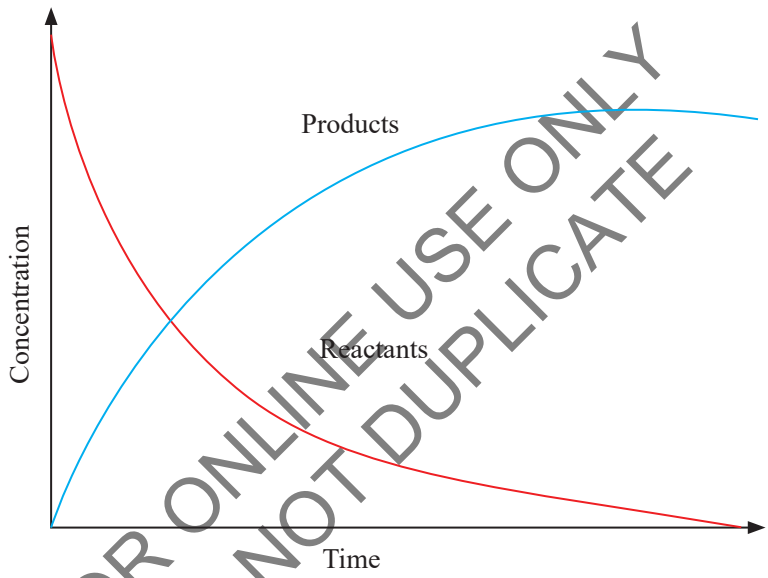
$$\text{Rate of reaction} = \frac{\text{amount of reactant used}}{\text{time taken}}$$

or

$$\text{Rate of reaction} = \frac{\text{amount of product formed}}{\text{time taken}}$$

For the reactants or products in solution form, the rate of reaction can also be expressed in terms of the change in reactant or product concentrations over change in time.

As a chemical reaction proceeds, the concentrations of the reactants (starting materials) decrease, but those of the products increase as shown in Figure 7.1.



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Figure 7.1: *Variation in the concentrations of reactants and products with time*

In general;

$$\text{Rate of reaction} = \frac{\text{measured change in reactants/products during a given time}}{\text{time taken for a change to occur}}$$

For the reactions that produce gaseous products, the rate can be measured in terms of volume of the gas produced or by calculating the reduction in the mass of reaction solution. For example, the reaction between zinc granules and dilute hydrochloric acid is a metal-acid reaction in which zinc chloride is formed and hydrogen gas is evolved as indicated in the following chemical equation:

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

The rate of this reaction can be determined by measuring the volume of the hydrogen gas produced per unit time. The change in mass of the reaction mixture with time can

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also be used to determine the rate of reaction. If the reaction produces a product in the form of a precipitate, the amount of the product formed can be used to determine the rate of reaction.

Some reactions involve coloured substances (either reactants or products). The progress of such reactions can be monitored by observing the colour changes. For example, potassium iodide and lead nitrate are colourless. The reaction between potassium iodide and lead nitrate is instantaneous, and the yellow colour of lead iodide appears immediately as the two colourless solutions are mixed. Thus, the colour change is an indication that the reaction has taken place.

$$Pb^{2+}(aq) + 2I^{-}(aq) \longrightarrow \underset{\text{Yellow}}{PbI_2(s)}$$

Comparing the rates of chemical reactions

On the basis of speed, chemical reactions are categorised into three types: instantaneous or rapid reactions, moderate reactions, and very slow reactions.

Instantaneous reactions are those reactions that occur in a very short period of time. Examples of such reactions include the neutralisation of an acid by a base when both reactants are in aqueous solution, and the precipitation of many substances such as silver chloride. For example, if two colourless solutions of sodium chloride and silver nitrate are mixed, they react immediately to give a white precipitate of silver chloride as per the following chemical equation:

$$Ag^{+}(aq) + Cl^{-}(aq) \longrightarrow \underset{\text{White}}{AgCl(s)}$$

The formation of silver chloride as a visible white precipitate requires a solution of silver nitrate of very low concentration.

Moderate reactions fall between instantaneous and very slow reactions. They take place at a measurable time at temperatures that are easily achieved in the laboratory. For example, the reaction between marble chips and dilute hydrochloric acid, and the reaction between sodium thiosulphate and a dilute mineral acid such as nitric acid are moderate reactions.

Very slow reactions are those reactions which occur in a long period of time. For example, the combination of hydrogen and oxygen in the absence of a catalyst at ordinary temperature is so slow that there is no observable change during the reaction. The rusting process is also relatively slow. Iron reacts with oxygen in the presence of

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water to form hydrated iron(III) oxide (rust). The chemical equation for this reaction is as follows:

$$4Fe(s) + 3O_2(g) \xrightarrow{xH_2O} 2Fe_2O_3{\cdot}xH_2O(s)$$

The number of molecules of the water of crystallisation (x) varies under different conditions. Another example of a reaction which is quite slow is the decomposition of hydrogen peroxide in the absence of sunlight and a catalyst such as manganese(IV) oxide. This explains why when hydrogen peroxide is in a dark sealed container can last for a long period of time.

Activity 7.1



Aim: To investigate the rate of the reaction between potassium iodide and lead(II) nitrate

Requirements: Two beakers, spatula, measuring cylinder, stirring rod, stopwatch, potassium iodide, lead(II) nitrate, and distilled water

Procedure

1. Put about 25 cm^3 of distilled water in each of the two beakers.
2. Add a spatulaful of potassium iodide in one beaker and a spatulaful of lead(II) nitrate in the other beaker. Stir each mixture until all the solid particles dissolve.
3. Mix the two solutions by either adding the potassium iodide solution to the lead(II) nitrate solution or vice versa. Start the stopwatch and observe the reaction. Note the time taken for any changes observed.

Questions

1. What did you observe when the two solutions were mixed?
2. How long did it take for the reaction in Step 3 to occur?
3. Comment on the rate of the reaction observed.

Activity 7.2



Aim: To investigate the rate of precipitation of silver chloride

Requirements: Two beakers, measuring cylinder, stopwatch, 0.05 M silver nitrate solution, and 1 M sodium chloride solution

Procedure

1. Measure about 25 cm^3 of sodium chloride solution and put it in a beaker.

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2. Measure about 25 cm^3 of silver nitrate solution and put it in another beaker.
3. Add the silver nitrate to the sodium chloride solution. Start the stopwatch and note the time it takes to observe any changes.

Questions

1. What was observed when the two solutions were mixed?
2. Comment on the time interval between the time the solutions were mixed and when the first observable changes were noted.
3. Explain the relationship between the observed time taken and the type of the reaction occurring.

Activity 7.3



Aim: To demonstrate the rusting of iron

Requirements: Test tube, two iron nails, measuring cylinder, sandpaper or steel wool, and water

Procedure

1. Clean the two iron nails by using sandpaper or steel wool.
2. Put the iron nails in a test tube.
3. Add about 10 cm^3 of tap water into the test tube.
4. Leave the set-up for three days.
5. Observe the iron nails daily.

Questions

1. Explain your observations after three days.
2. Describe the rate of rusting of the iron nails.

Measuring the rates of chemical reactions

The rate of a chemical reaction can be determined by either measuring the rate at which the reactants are used up, or measuring the rate at which the products are formed.

Consider the following hypothetical reaction:

$$A + B \longrightarrow C$$

In this reaction, one molecule of A reacts with one molecule of B to form one molecule of C. The rate of the appearance and disappearance of these chemical species are related by the following expression:

$$\text{Average rate} = \frac{-\Delta A}{\Delta t} = \frac{-\Delta B}{\Delta t} = \frac{+\Delta C}{\Delta t}$$

Measurement of the rate of reaction depends on the nature of the reaction. For reactions that produce gases, the volume of the gas produced can be used to determine the rate of the reaction. Table 7.1 shows sample results from an experiment conducted for the reaction between zinc granules and hydrochloric acid. This reaction produces hydrogen gas as one of the products. The graph derived from the results presented in Table 7.1 is shown in Figure 7.2.

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Table 7.1: *Volume of hydrogen gas recorded at 30 seconds intervals*

Time (s)	0	30	60	90	120	150	180	210	240	270	300
Volume (cm^3)	0	19	33	45	55	62	66	69	70	70	70

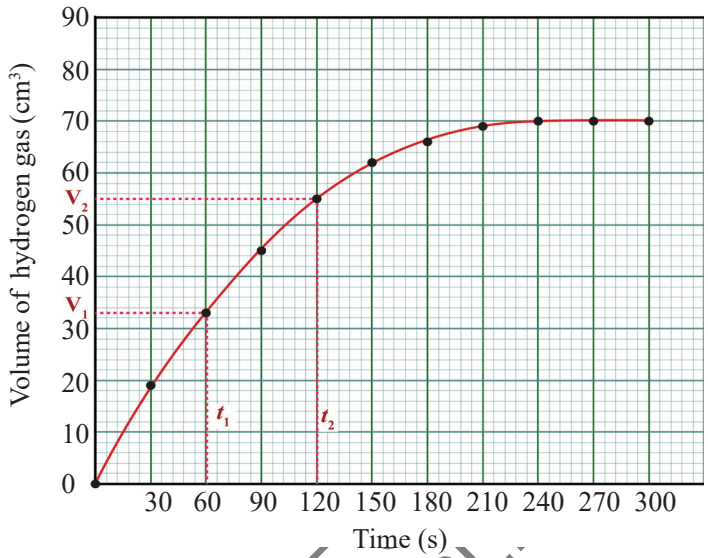


Figure 7.2: *Volume of hydrogen gas produced against time*

The graph indicates that the volume of hydrogen gas increases with time. Since the amounts of the reactants do decrease with time, the rate of evolution of hydrogen gas also decreases with time. The average rate of the reaction can be obtained by dividing the change in volume of hydrogen gas by the time interval.

$$\text{Rate of reaction} = \frac{\Delta\text{volume}}{\Delta\text{time}}$$

The rate of the reaction is, therefore, equal to the slope of the curve.

The average reaction rate at any time interval (t_1 and t_2) can be determined by getting the gradient of the line joining the two points.

Thus,

$$\text{Rate of reaction} = \frac{V_2 - V_1}{t_2 - t_1}$$

From the graph, the rate of reaction between the first minute and the second minute is:

$$\text{Rate of reaction} = \frac{(55-33)\ \text{cm}^3}{(120-60)\ \text{s}}$$

$$= \frac{22\ \text{cm}^3}{60\ \text{s}} = 0.37\ \text{cm}^3\ \text{s}^{-1}$$

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The gradient of the curve is the highest at the start and the lowest at the end of the reaction. From the graph, there is no change in the volume of the gas evolved between the fourth minute (240 s) and the fifth minute (300 s).

Therefore, $\frac{\Delta V}{\Delta t} = \frac{0}{60\ s} = 0$ implying that, at this time interval, the rate of reaction is zero. This shows that the reaction is complete since no reactants are being converted to products.

Activity 7.4

Aim: To investigate the rate of the reaction between dilute hydrochloric acid and zinc granules

Requirements: Conical flask, burette, measuring cylinder, two retort stands and clamps, trough, rubber bung to fit the conical flask, delivery tubes, stopwatch or clock, weighing balance, 2 M hydrochloric acid, and zinc granules

Procedure

1. Put about 100 cm^3 of 2 M hydrochloric acid in the conical flask.
2. Set the apparatus as shown in Figure 7.3.

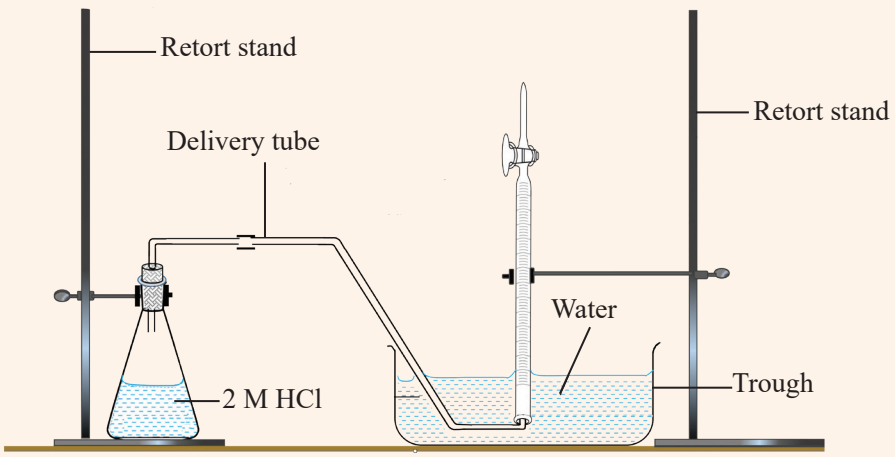


Figure 7.3: *Set-up for the reaction of zinc granules with dilute hydrochloric acid*

3. Remove the rubber bung and put some zinc granules (about 1.5 g) into the conical flask containing the hydrochloric acid. Quickly return the rubber bung and start the stopwatch.

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4. Read and record the volume of the gas collected in the burette at 20 seconds intervals for about seven minutes. Record your results in a tabular form as shown in Table 7.2.

Table 7.2: *Volume of hydrogen gas recorded at 20 seconds time intervals*

Time (s)	Volume of hydrogen gas (cm^3)
0	
20	
40	
60	

Questions

1. What is observed when the zinc granules were added into the 2 M hydrochloric acid?
2. Plot a graph of volume of the hydrogen gas liberated (cm^3) against time taken (s).
3. Interprete the plotted graph in Question 2.
4. From the graph, determine the rate of the reaction between zinc granules and hydrochloric acid.
5. Compare the nature of your graph with that shown in Figure 7.1.

Exercise 7.1

1. Explain the meaning of the following terms:
   (a) Chemical kinetics
   (b) Rate of reaction
2. Distinguish between instantaneous reactions and moderate reactions.
3. The reaction between carbon monoxide and nitrogen dioxide is represented by the following chemical equation:

$$CO(g) + NO_2(g) \longrightarrow CO_2(g) + NO(g)$$

If the concentration of NO is 0.0001 M at time $t_1 = 0.00$ s, and 0.0015 M at $t_2 = 2.005$ s after the reaction begins, calculate the rate of this reaction.

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4. In an experiment for the determination of the rate of reaction between butyl chloride (C_4H_9Cl) and water, the concentration of C_4H_9Cl was 0.200 M at the beginning of the reaction. At 4 seconds, the concentration of C_4H_9Cl was 0.100 M. Calculate the average reaction rate over the given time period.
5. A form three student wishes to measure the reaction rate of zinc with dilute hydrochloric acid. The chemical equation for the reaction is:

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

In a given experiment, 0.016 g of zinc granules was added into 1.00 L of 0.100 M HCI, and the following data were recorded:

Time (s)	0	4	8	12	16	20
Mass of zinc (g)	0.016	0.014	0.012	0.010	0.008	0.006

(a) Calculate the rate of the reaction in grams of the zinc granules consumed per second.
(b) Calculate the rate of the reaction in moles of the zinc granules consumed per second.
(c) Write the complete ionic equation for the reaction.
(d) What would happen to the acid as the reaction proceeds?

Factors affecting the rate of a chemical reaction

The rate or speed at which a reaction occurs depends on the collision of the reactant molecules. The more often reactant molecules collide, the more often they react with one another, and the faster the reaction rate. The theory that provides a qualitative explanation of the chemical reactions and the rate at which they occur is called the *collision theory*. According to the collision theory, reactions occur due to collisions of reactant particles. Not all collisions succeed in producing a chemical reaction; those with proper orientation and certain minimum amount of energy result in chemical reactions. These collisions are referred to as *effective collisions* and the minimum amount of energy is known as *activation energy*. Activation energy is the minimum amount of energy required by the reactant(s) in a chemical reaction to reach an *activated state* from which the product(s) can form. The activation energy is used to break the bonds in the reactants so that new bonds can be formed in the products.

In general, a factor that increases the number of collisions between reactant particles will increase the reaction rate and a factor that decreases the number of collisions between particles will decrease the reaction rate. The factors that can be changed to

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alter the rate of a reaction include: temperature, pressure, concentration, surface area, and a catalyst.

Temperature

The rate of a chemical reaction increases with an increase in temperature. An increase in temperature increases the kinetic energy of the reactant particles. This causes the particles to move faster, thus increasing the chances (frequencies) of effective collisions between the reacting particles. Table 7.3 shows the sample results from an experiment conducted to investigate the effect of temperature on the reaction between sodium thiosulphate and hydrochloric acid. Dilute hydrochloric acid reacts with sodium thiosulphate to give sodium chloride, water, sulphur, and sulphur dioxide gas. The chemical equation for this reaction is as follows:

$$Na_2S_2O_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + S(s) + SO_2(g)$$

The fine particles of sulphur formed from the reaction make the reaction mixture cloudy. Sulphur dioxide is detected by its irritating smell as the reaction progresses. During the experiment, the reaction vessel is usually placed on a white tile or piece of white paper marked with a cross (X). The cross disappears when a certain amount of sulphur is formed. The time taken for such amount to be formed can be monitored in order to determine the rate of this reaction.

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Table 7.3: *Effect of temperature on the rate of the reaction between sodium thiosulphate and dilute hydrochloric acid*

Temperature (°C)	20	30	40	50	60
Time taken for the cross to disappear (s)	420	240	165	125	90

Figure 7.4 shows a graph of temperature against time based on the results of the experiment (Table 7.3). The graph shows that the time taken for the cross to disappear decreases with an increase in temperature. Thus, the rate of the reaction increases with an increase in temperature.

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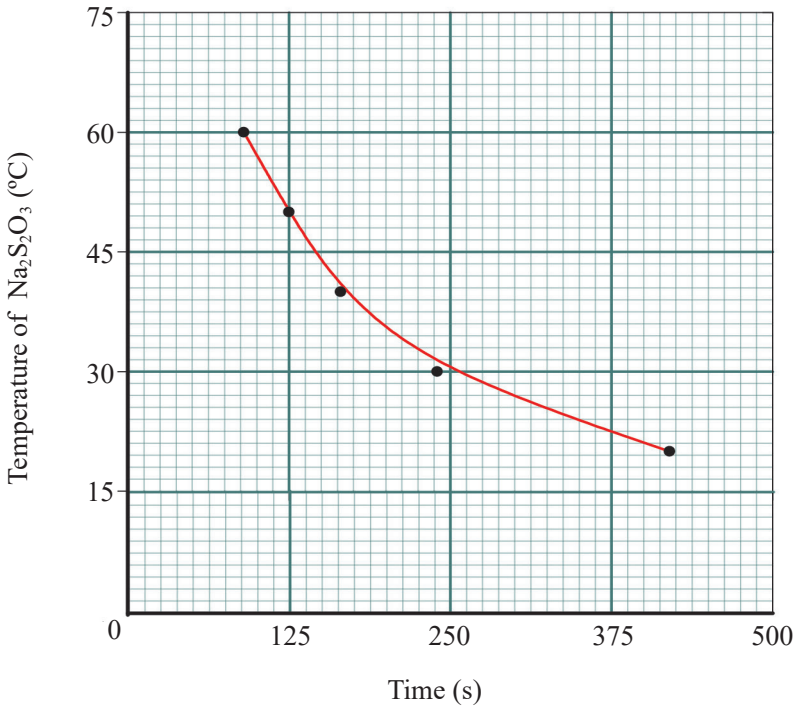


Figure 7.4: *Graph of temperature against time for the reaction between sodium thiosulphate and hydrochloric acid*

Activity 7.5

Aim: To investigate the effect of temperature on the rate of a chemical reaction

Requirements: Five 100-mL beakers, measuring cylinders, stopwatch or clock, thermometer (−10 ºC to 110 ºC), rubber bungs, source of heat, white paper, labels, black or blue felt-tip pen, 1 M hydrochloric acid, and 0.05 M sodium thiosulphate

Procedure

1. Label the five beakers as A, B, C, D, and E.
2. Measure five portions of about 50 cm^3 of 0.05 M sodium thiosulphate solution. Put each portion in a labelled beaker.
3. Record the temperature (T) of the solution in beaker A.
4. Draw a cross (X) on a white paper using a felt-tip pen.
5. Add about 10 cm^3 of 1 M hydrochloric acid into beaker A and at the same time start the stopwatch.

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6. Swirl the mixture in the beaker and place it on the cross (X) mark.
   Observe the mixture from above as shown in Figure 7.5. Then, note the time taken for the cross to disappear.

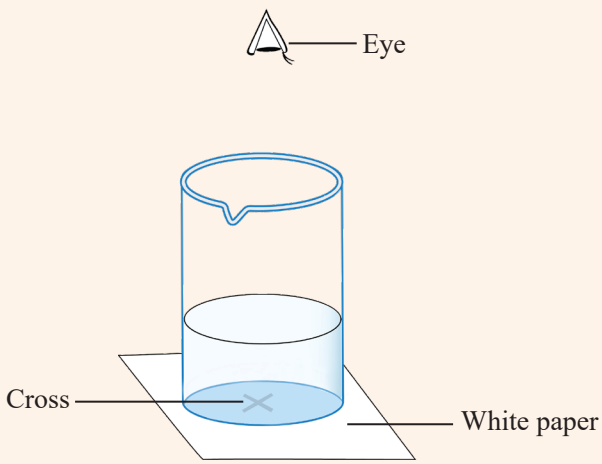


Figure 7.5: *Eye view for monitoring the disappearance of a cross* (X)

7. Note down any other observations such as smell from the reaction mixture.
8. Repeat Steps 1–7 using the solutions in the beakers B, C, D, and E at 30 ºC, 40 ºC, 50 ºC, and 60 ºC, respectively. To raise the temperature, gently heat the sodium thiosulphate solution using a Bunsen burner or a methanol lamp until the desired temperature is achieved. The hydrochloric acid should also be at the same temperature as the portion of sodium thiosulphate solution to be reacted with. This can be achieved by heating the acid separately in a water bath.
9. Record the results as shown in Table 7.4.

Table 7.4: *Time taken for the cross (X) to disappear at different reaction temperatures*

Temperature of $Na_2S_2O_3$ (*aq*) (ºC)	T	30	40	50	60
Time taken for the cross to disappear (s)					

Key: T is a temperature of solution recorded in beaker A.

Questions

1. Use the results obtained to plot a graph of temperature against time.
2. From the graph, comment on the effect of increasing the reaction temperature on the rate of the reaction.

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Pressure

Changing the pressure of the reactants in solid or liquid state has no effect in the rate of the reaction. Increasing the pressure for a reaction involving gases will increase the rate of the reaction. This is because the increase in pressure decreases the volume of the gaseous reaction mixture which causes the gas molecules to collide more frequently.

Concentration

Changing the concentration of one or more reactants alters the number of particles per unit volume of the reaction mixture. Thus, increasing the concentrations of reactants increases the number of collisions per unit time, and hence increases the rate of a chemical reaction. Table 7.5 shows sample results obtained from an experiment conducted to demonstrate the effect of changing concentration on the rate of the reaction between sodium thiosulphate and dilute hydrochloric acid.

Table 7.5: *Effect of concentration on the rate of reaction between sodium thiosulphate and dilute hydrochloric acid*

Volume, V, of $Na_2S_2O_3$ (*aq*) (cm^3)	Volume of water (cm^3)	Total volume (cm^3)	Time, t, taken for the cross to disappear (s)
30	0	30	0
25	5	30	12
20	10	30	36
15	15	30	72
10	20	30	152
5	25	30	280

The results in Table 7.5 are plotted in the graph of volume of $Na_2S_2O_3$ against the time taken for the cross to disappear (Figure 7.6). The graph in Figure 7.6 shows that the rate of the reaction is slow at low concentrations and faster at higher concentrations. Therefore, the rate of the reaction is directly proportional to the concentrations of the reacting substances.

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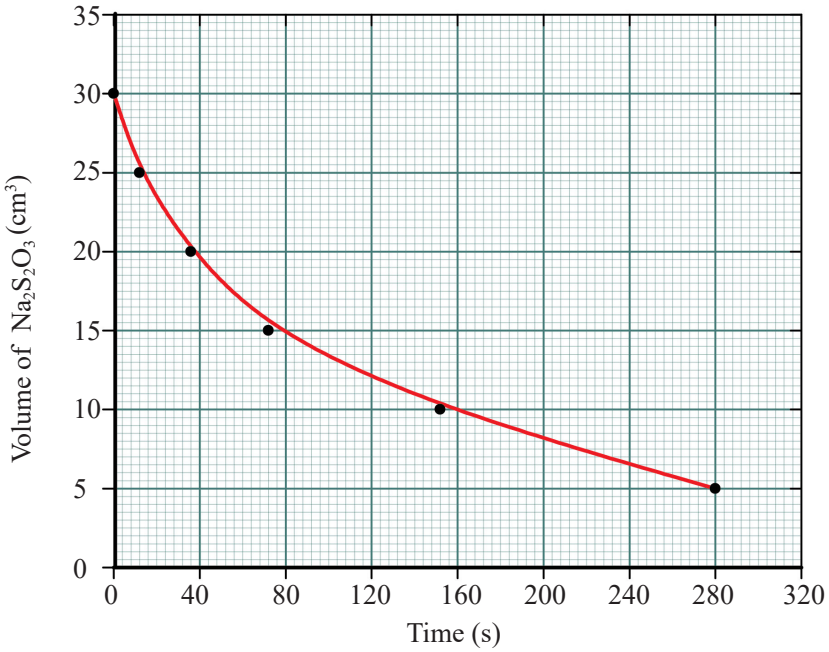


Figure 7.6: *Graph of volume against time for the reaction between sodium thiosulphate and dilute hydrochloric acid*

Activity 7.6



Aim: To investigate the effect of concentration on the rate of the reaction between hydrochloric acid and sodium thiosulphate

Requirements: Seven beakers, measuring cylinders, stopwatch or clock, white paper, black or blue felt-tip pen, labels, 0.5 M sodium thiosulphate, 1 M hydrochloric acid, and distilled water

Procedure

1. Label seven beakers as follows: 35 cm^3, 30 cm^3, 25 cm^3, 20 cm^3, 15 cm^3, 10 cm^3, and 5 cm^3.
2. Arrange the seven beakers in a line from left to right starting with the 35 cm^3 followed by 30 cm^3, 25 cm^3, 20 cm^3, 15 cm^3, 10 cm^3, and 5 cm^3.
3. Measure the required volumes of 0.5 M sodium thiosulphate solution and transfer into the seven beakers as per the labels. For example, in the beaker labelled 35 cm^3, put 35 cm^3 of sodium thiosulphate solution.
4. Add distilled water, starting from the left beaker to the right: 0 cm^3, 5 cm^3, 10 cm^3, 15 cm^3, 20 cm^3, 25 cm^3, and 30 cm^3.

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5. Draw a cross (X) on a white paper using the felt-tip pen. Place the first beaker (35 cm^3) on the X mark of the white paper.
6. Add 10 cm^3 of 1 M hydrochloric acid into the beaker on the white paper and at the same time start the stopwatch.
7. Swirl the mixture, then place the beaker on the white paper again.
8. Observe the solution mixture from above as shown in Figure 7.5. Switch off the stopwatch immediately when the cross mark disappears from the sight.
9. Record the time taken for the cross to disappear.
10. Repeat Steps 1–9 with the remaining six solutions. Record the results in a tabular form as indicated in Table 7.6.

Note: The concentration of the acid remains the same for all the solutions.

Table 7.6: *Time taken for the cross (X) to disappear for different concentrations of* $Na_2S_2O_3$

Volume, V, of $Na_2S_2O_3$ (*aq*) (cm^3)	Volume of water (cm^3)	Total volume (cm^3)	Time, t, taken for the cross to disappear (s)
35	0	35	
30	5	35	
25	10	35	
20	15	35	
15	20	35	
10	25	35	
5	30	35	

Questions

1. Plot a graph of the volume, V, of the sodium thiosulphate solution used against the inverse of time, $\frac{1}{t}$, it takes for the cross to disappear.
2. What does the shape of the graph indicate? Explain.
3. What happens to the concentration of sodium thiosulphate solution after adding distilled water?

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Surface area

The rate of chemical reaction can be raised by increasing the surface area of the solid reactant. This is done by cutting the substance into small pieces or by grinding it into a powder. Surface area is only significant in *heterogeneous reactions*. For example, in the reaction between a solid and a liquid, the surface area of the solid will have high impact on the rate of the reaction. Reactions involving solids take place on the surface of the solid. Thus, only the particles at the surface take part in the reaction. Increasing the surface area of the solid exposes more solid molecules to the liquid, which allows faster reaction. The solid molecules trapped within the body of the solid are not exposed, hence cannot react. Think of a cube measuring 1 cm × 1 cm × 1 cm cut into 1000 equal cubes measuring 1 mm × 1 mm × 1 mm as shown in Figure 7.7.

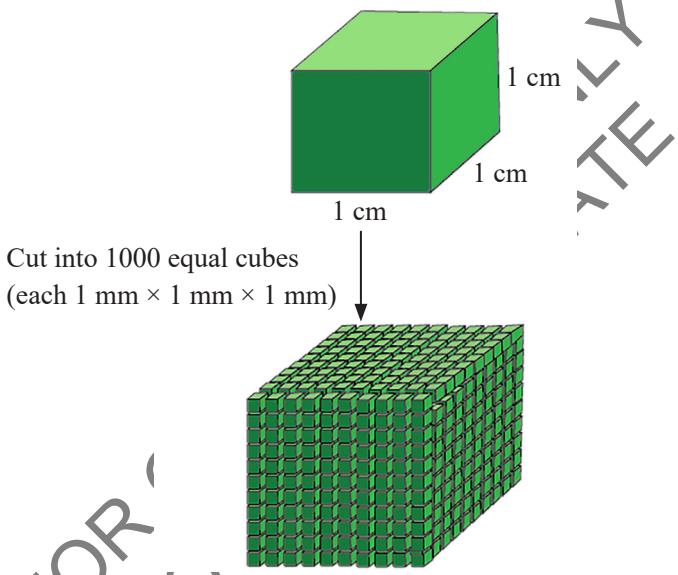


Figure 7.7: *Comparing the surface areas of solids*

The surface area of the big cube is 1 cm × 1 cm × 6 (faces) = 6 cm^2. The surface area of each small cube is 0.1 cm × 0.1 cm × 6 = 0.06 cm^2. The total surface area of the small cubes is 0.06 cm^2 × 1000 = 60 cm^2. Thus, the small cubes have a larger surface area than the big cube.

A simple experiment to demonstrate the effect of surface area on reaction rate is the reaction between calcium carbonate (marble) and hydrochloric acid in which the reaction rates of both powdered marble and marble chips can be compared. A sample of marble chips reacts with dilute hydrochloric acid to form calcium chloride, water and carbon dioxide as indicated in the following chemical equation:

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

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The reaction is very vigorous when the sample of powdered marble is used. This is because the powder has larger surface area compared to marble chips of equal mass. The surface area of the marble chips that is in contact with the acid is less, hence fewer particles react with the acid. On the other hand, the powdered marble chips have high surface area that is in contact with the acid, thus increases the number of particles that react with the acid. The greater the surface area that is in contact with the acid, the higher the rate of reaction. This is why the reaction between hydrochloric acid and the powdered marble chips is vigorous.

In the experiment involving the reaction between marble chips and hydrochloric acid, the volume of carbon dioxide produced per unit time may be used to determine the rate of the reaction. Table 7.7 shows the sample results from such an experiment.

Table 7.7: *Effect of surface area on the rate of the reaction*

Time (s)	Volume of CO_2 evolved with powdered marble (cm^3)	Volume of CO_2 evolved with marble chips (cm^3)
10	20	9
20	32	17
30	45	24
40	53	30
50	58	36
60	60	42
70	60	47
80	60	51
90	60	55
100	60	58
110	60	59
120	60	60
130	60	60

The plots drawn by using the results in Table 7.7 are shown in Figure 7.8. It can be observed that, the plot for the reaction of powdered marble and hydrochloric acid has a steeper gradient at the start which takes less time to become horizontal as compared to the graph for the reaction between marble chips and hydrochloric acid. This shows that the rate of the reaction between hydrochloric acid and powdered marble is higher than that of the reaction between hydrochloric acid and marble chips. This indicates that the rate of the reaction is greater when the surface area is increased.

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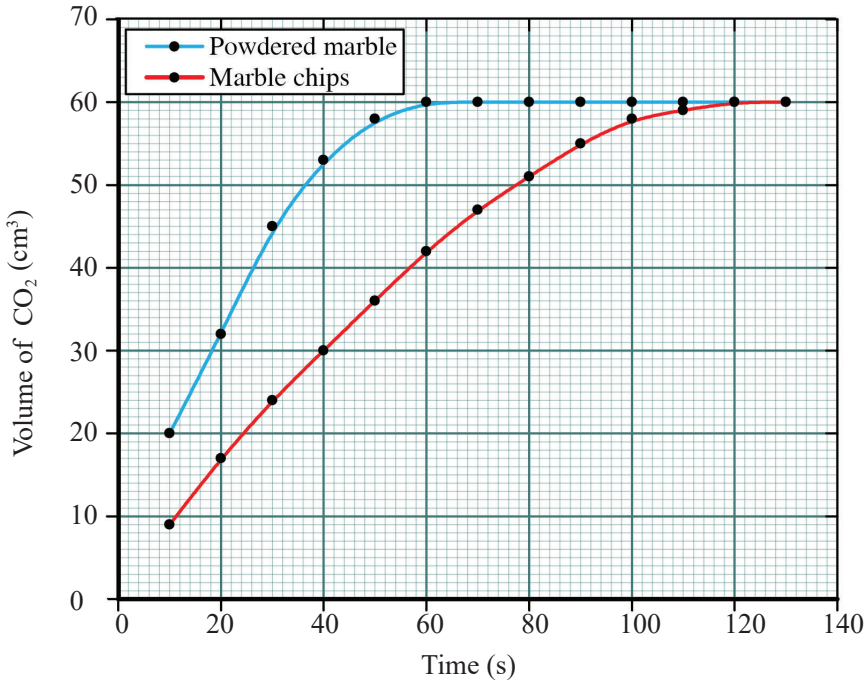


Figure 7.8: *Graph of volume of CO_2 produced against time for the reaction between marble and hydrochloric acid*

The rate of this reaction could also be determined by measuring the decrease of mass of the marble in the reaction flask. A graph similar to that in Figure 7.6 would be obtained from the results.

Activity 7.7



Aim: To investigate the effect of surface area on the rate of the reaction between hydrochloric acid and marble

Requirements: Mortar and pestle, weighing balance, 250-mL conical flask with a side-arm, 100-mL measuring cylinder, two watch glasses, 100-mL syringes, stopwatch or clock, rubber bung, retort stand and clamp, marble chips, and 1 M hydrochloric acid

Procedure

1. Measure 100 cm^3 of 1 M hydrochloric acid and put it in a conical flask.
2. Set the apparatus as shown in Figure 7.9.
3. Weigh two samples of marble chips, 0.5 g each, in two separate watch glasses. Grind one sample into powder and resize the other sample into small chips using a mortar and pestle.

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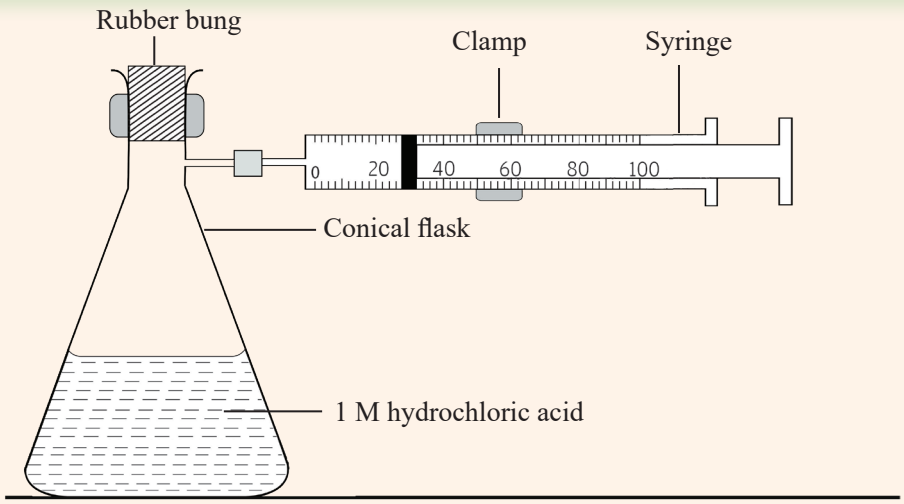


Figure 7.9: *Set-up for the reaction between hydrochloric acid and marble*

4. Remove the rubber bung and add the powdered sample of marble into the acid. Start the stopwatch immediately and at the same time, replace the rubber bung quickly.
5. Repeat Steps 1 and 4 by using the small crystals obtained by grinding marble chips. Record the results in a tabular form as shown in Table 7.8.

Table 7.8: *Volume of CO_2 evolved from the reaction of HCl with marble*

Time (s)	Volume of CO_2 evolved with powdered marble (cm^3)	Volume of CO_2 evolved with sample of marble chips (cm^3)
10		
20		
30		
40		
50		
60		
70		
80		
90		
100		
110		
120		
130		

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Questions

1. Which reaction took less time to complete?
2. Plot a graph of the volume of carbon dioxide gas evolved against time for the powdered marble and the marble chips on the same axes.
3. From the graph, explain the effect of grinding the marble chips into powder on the rate of reaction.

Catalyst

A *catalyst* is a substance that alters the rate of a chemical reaction but remains chemically unchanged at the end of the reaction. Although a catalyst is chemically unchanged in the chemical reaction, it can change its physical nature, for example, from coarse powder to fine powder in the course of a reaction. Most catalysts are used to increase the rates of chemical reactions by lowering the activation energy of the reaction. The phenomenon of increasing the rate of reaction by the use of a catalyst is called *catalysis*. For example, the decomposition reaction of hydrogen peroxide occurs very slowly at room temperature. When manganese(IV) oxide is added to a flask containing hydrogen peroxide, the reaction occurs very fast and a lot of oxygen is collected within a short time.

$$2H_2O_2(l) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

At the end of the experiment, the amount of manganese(IV) oxide that is left is equal to the original amount that was added to the reaction mixture. This shows that manganese(IV) oxide is not used up during the reaction. The dried manganese(IV) oxide can therefore be reused. Increasing the amount of the catalyst will also increase the rate of the reaction but only up to a certain limit. For example, if 3 g of manganese(IV) oxide are used, the volume of oxygen collected at each specific time interval is more than the corresponding volume collected when 1 g is used as shown in Table 7.9 and Figure 7.10. This is because increasing the amount of the catalyst increases the chances of each reactant molecule to come into contact with the catalyst for the initiation of the reaction. The volume of oxygen gas collected is however the same at the end of the experiment regardless of the amount of catalyst used. Manganese(IV) oxide only acts as a catalyst in this reaction by increasing the rate of the reaction but not the amounts of the products formed. It should be noted that, increasing the amount of the catalyst will not increase the rate of the reaction beyond a certain point. When more of the catalyst is added, it will come a point where all the reactant molecules will have a sufficient amount of the catalyst, and therefore, the reaction rate would not increase any further.

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Table 7.9: *Volume of oxygen gas evolved from the decomposition of hydrogen peroxide*

Time (s)	Volume of O_2 (cm^3)	
	1 g of manganese(IV) oxide catalyst	3 g of manganese(IV) oxide catalyst
0	0	0
15	9	15
30	18	29
45	26	42
60	35	53
75	44	64
90	50	70
105	57	75
120	62	78
135	67	80
150	71	80
165	74	80
180	76	80
195	78	80
210	80	80
225	80	80

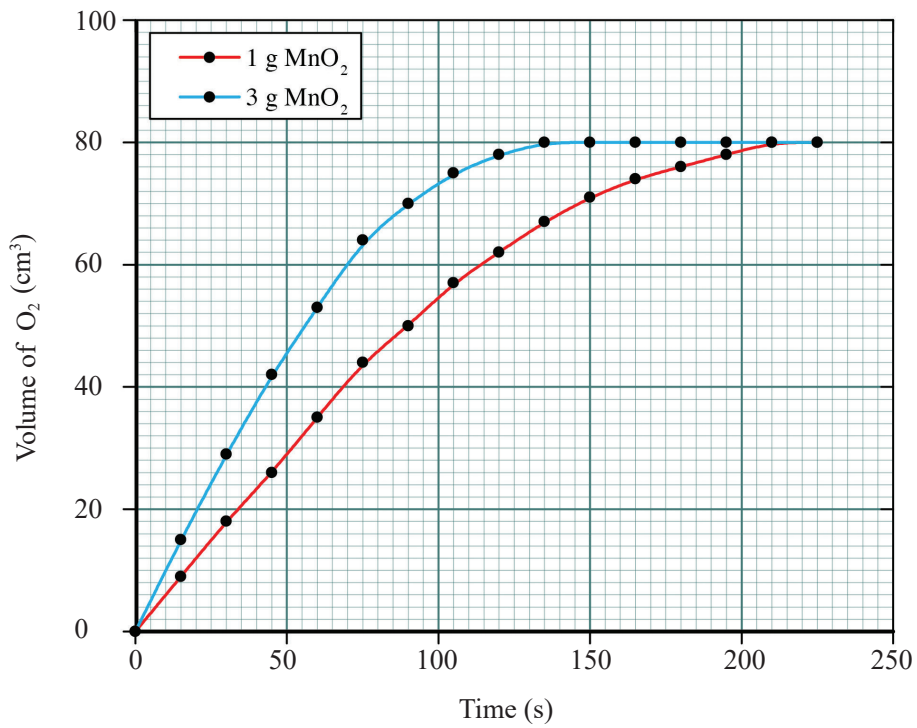


Figure 7.10: *Effect of the amount of a catalyst on the rate of decomposition of H_2O_2*

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Activity 7.8

Aim: To investigate the effect of the amount of a catalyst on the rate of decomposition of hydrogen peroxide

Requirements: Conical flask with a side-arm, two 100-mL syringes, two rubber bungs, sand bath, weighing balance, two stands and clamps, stopwatch or clock, wash bottle, measuring cylinder, delivery tubes, filter paper, 0.7-volume hydrogen peroxide, and manganese(IV) oxide

Procedure

1. Prepare 0.7-volume of hydrogen peroxide by diluting 7 mL of 100-volume hydrogen peroxide into one litre.
2. Measure about 100 mL of 0.7-volume hydrogen peroxide and put it in a conical flask.
3. Set the apparatus as shown in Figure 7.11.

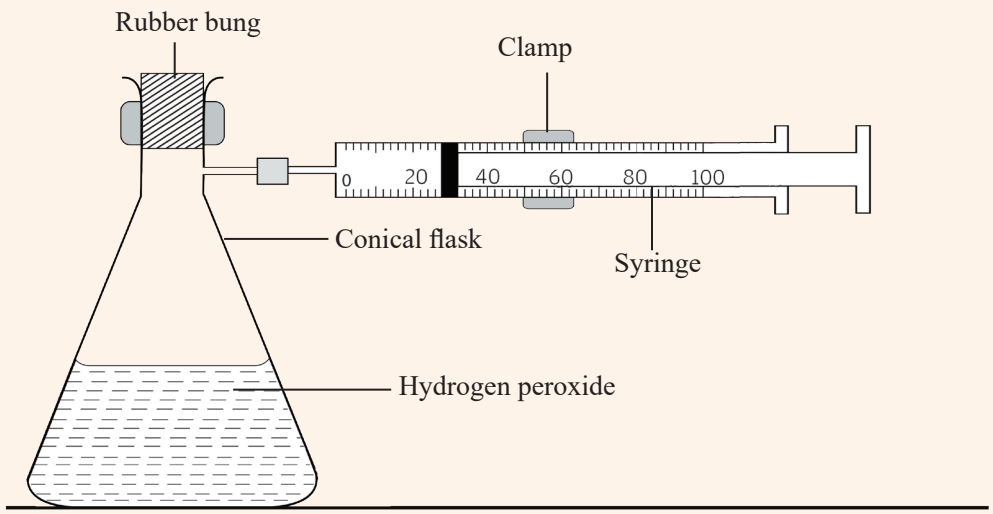


Figure 7.11: *Apparatus set-up for the decomposition of hydrogen peroxide*

4. Remove the rubber bung and add 1 g of manganese(IV) oxide. Start the stopwatch as you quickly replace the rubber bung.
5. Shake the flask gently. Note the volume of the gas collected in the syringe at 10 second intervals. Continue recording until there is no observable change in the volume of the gas collected in the syringe. Record your results as shown in Table 7.10.

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Table 7.10: *Volume of oxygen gas evolved from the decomposition of H_2O_2*

Time (s)	Volume of O_2 collected (cm^3)
0	
10	
20	
30	
40	

6. Weigh a piece of filter paper and use it to filter the suspension that is left in the flask. Remove all the particles from the conical flask by washing the flask with distilled water.
7. Dry the filter paper together with the residue over a sand bath. Weigh the residue together with the filter paper. Repeat the process until the mass measured does not change.
8. Repeat Steps 3–6 using 2 g of manganese(IV) oxide.

Questions

1. Explain your observation about the volume of the gas produced when;
   (a) 1 g of manganese(IV) oxide was used.
   (b) 2 g of manganese(IV) oxide was used.
2. On the same graph, plot volume of oxygen (y-axis) produced against time when 1 g and 2 g of manganese(IV) oxide were used as catalysts.
3. Comments on the observations made in Steps 5–7.

Exercise 7.2

1. Explain the following concepts:
   (a) Collision theory
   (b) Effective collision
   (c) Activation energy
2. According to the collision theory, not all collisions lead into a chemical reaction. Explain.
3. What is the relationship between the rate of a chemical reaction and each of the following?
   (a) Temperature of the reaction

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(b) Reactant concentration
(c) Physical properties of the reactants
(d) Surface area of the reactants
(e) Catalyst

4. Why is the reaction rate generally the fastest at the beginning of a reaction?
5. How does a catalyst increase the rate of reaction?
6. Carbon burns in air according to the following chemical equation:

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

Describe four ways that can be used to increase the rate of this reaction.

Reversible and irreversible reactions

A chemical reaction is said to have occurred when the concentrations of reactants decrease and the concentrations of products increase. Chemical reactions can be either irreversible or reversible depending on the extent to which they proceed.

Irreversible reactions

Irreversible reactions take place in a forward direction only and proceed to completion. Such reactions occur until at least one reactant is used up. Examples of irreversible reactions include the reactions between relatively reactive metals and dilute acids to form a salt and hydrogen gas. Consider the reaction between magnesium metal and dilute hydrochloric acid which forms magnesium chloride and hydrogen gas as indicated in the following chemical equation:

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

This reaction cannot have its direction reversed in such a way that the starting materials (magnesium and hydrochloric acid) are obtained from the products. Such a reaction is said to be irreversible. The fully headed arrow (⟶) is used to show that a reaction is irreversible, meaning that it proceeds in forward direction only. Another example of an irreversible reaction is the reaction between zinc and hydrochloric acid. When hydrochloric acid is added to zinc granules, effervescence is observed. The gas evolved burns with a pop sound, an indication of the production of hydrogen gas.

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

Reversible reactions

Reversible reactions take place in both the forward and backward (reverse) directions. In this case, none of the reactants is completely used up and so the reaction does not proceed to completion. This means that the conditions of the reaction can be maintained or changed so that the reactants are obtained from the products. A reaction that proceeds

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in both directions is said to be a reversible reaction and is indicated by double half headed arrows ($\rightleftharpoons$) pointing in opposite directions. Such double arrows indicate that the reactants can become products and the products can become reactants.

An example of a reversible reaction is the decomposition of ammonium chloride. On gentle heating, ammonium chloride dissociates into ammonia and hydrogen chloride gas as per the following chemical equation:

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

Ammonia is an alkaline gas, while hydrogen chloride gas is acidic. Ammonia is less dense than hydrogen chloride and diffuses at a fast rate. Therefore, if blue and red litmus papers are employed during the decomposition of ammonium chloride experiment, ammonia will turn the wet red litmus paper blue first. Hydrogen chloride gas diffuses more slowly and will turn the blue litmus paper red much later. The two gases recombine on cooling to form solid ammonium chloride. Thus, heating of ammonium chloride is a reversible chemical reaction.

Another example of reversible reaction is the heating of hydrated copper(II) sulphate crystals. When blue hydrated copper(II) sulphate crystals are heated, they change colour from blue to white.

$$\underset{\text{Blue}}{CuSO_4 \cdot 5H_2O(s)} \xrightarrow{\text{heat}} \underset{\text{White}}{CuSO_4(s)} + 5H_2O(l)$$

The blue colour is restored when distilled water is added to the white crystals.

$$\underset{\text{White}}{CuSO_4(s)} + 5H_2O(l) \longrightarrow \underset{\text{Blue}}{CuSO_4 \cdot 5H_2O(s)}$$

This chemical equation is a reverse of the decomposition of hydrated copper(II) sulphate. The two chemical reactions can therefore be presented as follows:

$$\underset{\text{Blue}}{CuSO_4 \cdot 5H_2O(s)} \overset{\text{heat}}{\rightleftharpoons} \underset{\text{White}}{CuSO_4(s)} + 5H_2O(l)$$

The reverse reaction which involves the addition of water molecules to an anhydrous copper(II) sulphate gives off heat. Thus, when touched, the reaction container (test tube) feels warm.

Activity 7.9



Aim: To investigate reversibility of the reactions of hydrated copper(II) sulphate

Requirements: Boiling tube, beaker, Bunsen burner, test tubes, delivery tube, rubber bung, spatula, stand and clamp, dropper, hydrated copper(II) sulphate, distilled water and ice-cold water.

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Procedure

1. Place two spatulafuls of hydrated copper(II) sulphate into a boiling tube as shown in Figure 7.12.
2. Heat the hydrated copper(II) sulphate gently until there is no further observable change in the colour. Record your observation.
3. Allow the apparatus to cool to room temperature and then divide the solid residue into two portions. Place each portion in a separate clean dry test tube.
4. Pour the liquid collected into the test tube containing one portion of the residue, drop by drop (about five drops).
5. After adding the liquid, touch the test tube and feel the temperature.
6. Repeat Steps 4 and 5 by using the residue in the second test tube and distilled water. Record all your observations.

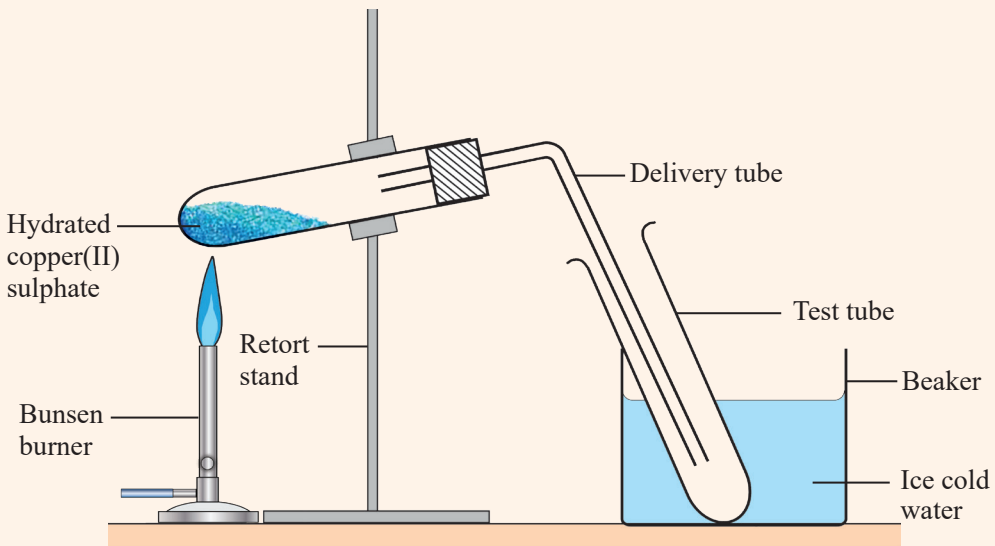


Figure 7.12: *Experimental set-up for the decomposition of hydrated copper(II) sulphate*

Questions

1. What happened when the liquid collected and distilled water were separately added into the solid left after heating?
2. Explain the observations based on reversible and irreversible reactions.
3. Why was the ice-cold water used in the experiment?

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Chemical equilibrium

In some reversible reactions, the forward and reverse reactions take place at the same time. The reactants react to form products while at the same time the products react to form the reactants. At the start of the reaction (zero time), there is maximum concentration of the reactant(s), and no products in the reaction. The rate of the forward reaction is maximum and the rate of the reverse reaction is therefore, zero. As the reaction proceeds, the concentrations of the products and the rate of the reverse reaction increase. Eventually, a point is reached when the rate of the forward reaction equalises with the rate of the reverse reaction. This balance is referred to as *equilibrium*. The equilibrium point can only be reached in a closed system, where no products or reactants escape to the outside environment. At equilibrium, the composition of the system no longer changes with time; the concentrations of reactants and products remain constant, but not necessarily equal. The mixture of reactants and products in the equilibrium state is called an *equilibrium mixture*. A chemical equilibrium is *dynamic* in nature, and is usually referred to as a *dynamic equilibrium*. This is because at equilibrium, the forward and reverse reactions are still occurring, but the two rates are equal and unchanging. Figure 7.13 shows the concept of dynamic equilibrium.

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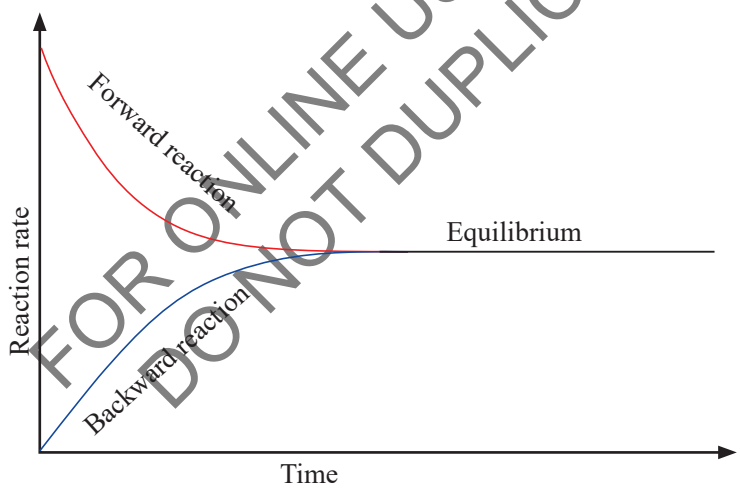


Figure 7.13: *Dynamic equilibrium for a reversible reaction*

An example of a chemical equilibrium reaction that can easily be performed is an equilibrium between liquid bromine (the dark red liquid) and bromine vapour (the orange gas). When a drop of liquid bromine is put into an empty conical flask, and the stopper is quickly fixed, some liquid evaporates (brown fumes are seen diffusing through the air from the drop); the amount of the liquid decreases, and the amount of the vapour increases. The brown fumes eventually fill the whole flask until there is no further change in colour inside the flask as shown in Figure 7.14.

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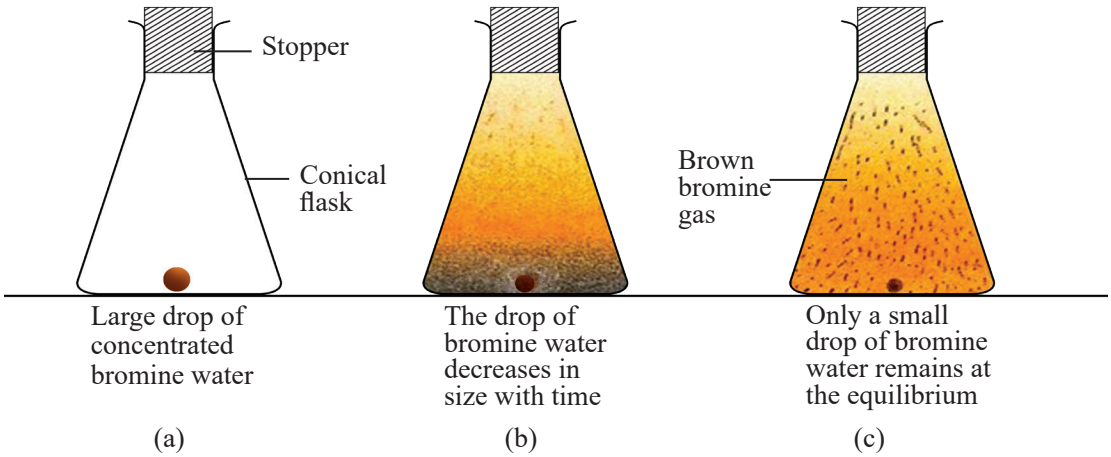


Figure 7.14: *Equilibrium reaction involving bromine water*

Since the conical flask is capped, no bromine vapour escapes and equilibrium between the liquid and the vapour is established as shown in Figure 7.15. A state of physical equilibrium exists when no further colour change is observed even if the flask is shaken. At equilibrium, the gaseous and liquid states exist together in a dynamic equilibrium.

$$Br_2(l) \rightleftharpoons Br_2(g)$$

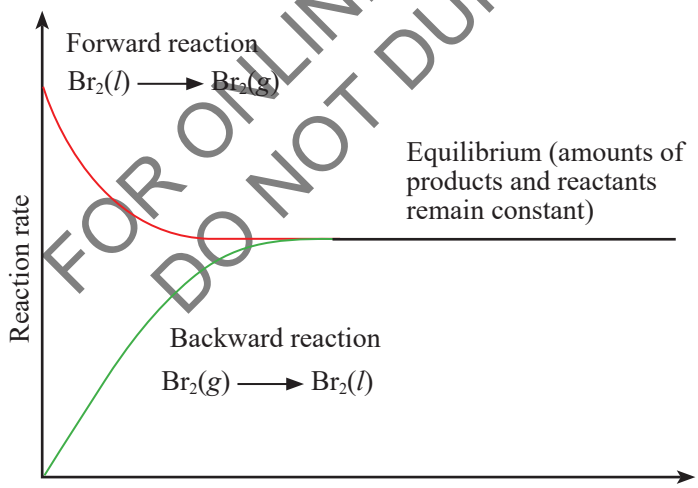


Figure 7.15: *Dynamic equilibrium between bromine water and bromine vapour*

As highlighted previously, the equilibrium point can only be reached in a closed system.

From Figure 7.14, if the stopper is removed, the molecules of bromine in the gaseous state would escape to the outside environment (fume chamber). The liquid remaining as a small drop would also diffuse into the air. This is because bromine water is no longer being produced from the condensation of bromine vapour.

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Another example of chemical equilibrium is observed in the chromate-dichromate equilibrium. The chromate ions (CrO_4^{2-}) are yellow in colour which turn orange in an acidic medium due to the formation of the dichromate ions ($Cr_2O_7^{2-}$).

$$2CrO_4^{2-}(aq) + 2H^+(aq) \longrightarrow Cr_2O_7^{2-}(aq) + H_2O(l)$$

Yellow — Orange

On the other hand, the dichromate ions, which are orange in colour, turn yellow in an alkaline solution due to the formation of the chromate ions.

$$Cr_2O_7^{2-}(aq) + 2OH^-(aq) \longrightarrow 2CrO_4^{2-}(aq) + H_2O(l)$$

Orange — Yellow

The chromate-dichromate reactions indicate that the two ions exist in a dynamic equilibrium in a neutral medium as illustrated in Figure 7.16. The chemical equilibrium of the two ions is presented as:

$$2CrO_4^{2-}(aq) + H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + OH^-(aq)$$

Yellow — Orange

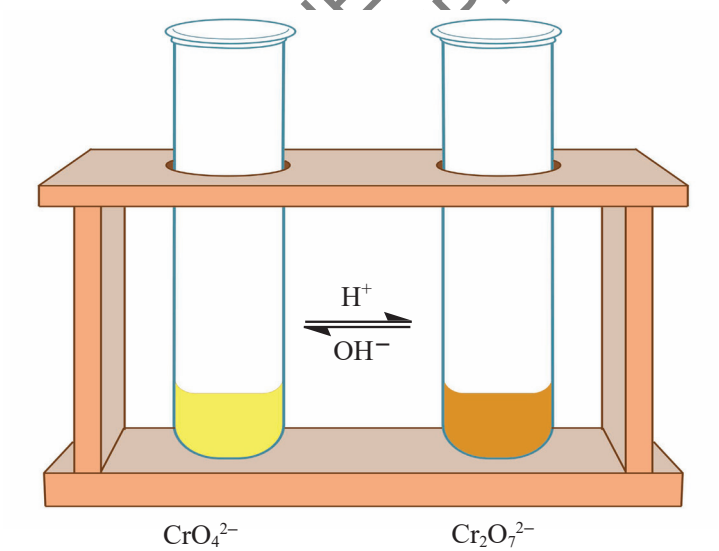


Figure 7.16: *Chromate-dichromate equilibrium*

When the concentration of hydrogen ions is increased, the equilibrium position shifts to the right; that is, more chromate ions are converted to dichromate ions. When the concentration of hydroxide ions is increased, the equilibrium shifts to the left. More dichromate ions are converted to chromate ions.

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Activity 7.10



Aim: To investigate the bromine water-bromine vapour equilibrium

Requirements: Conical flask, teat pipette, stopper, and concentrated bromine water

Caution! This experiment should be carried out in a fume chamber. Bromine is highly toxic. It should not be sucked into the pipette by mouth.

Procedure

1. Put one drop of concentrated bromine water in a clean dry conical flask by using a teat pipette. Fix the stopper immediately.
2. Record all the observations until no further change can be observed.
3. Shake the flask and note the observations.

Question

Based on the concept of chemical equilibrium, describe the changes observed in the conical flask with time.

Activity 7.11



Aim: To investigate the chromate-dichromate equilibrium

Requirements: 250-mL conical flasks, 2 droppers, measuring cylinder, 0.2 M potassium chromate solution, 0.2 M potassium dichromate solution, 1 M sulphuric acid, and 1 M sodium hydroxide

Procedure

1. Put 25 cm^3 of 0.2 M potassium chromate solution in a conical flask. Add 1 M sulphuric acid drop by drop until the colour of the solution changes.
2. To the resulting mixture in Step 1, add 1 M sodium hydroxide drop by drop until the colour of the solution changes.
3. Put 25 cm^3 of 0.2 M potassium dichromate solution in another conical flask. Add 1 M sodium hydroxide drop by drop until the colour of the solution changes.
4. To the resulting mixture in Step 3, add 1 M sulphuric acid drop by drop until the colour of the mixture changes.
5. Record all the observations.

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Question

Based on the concept of chemical equilibrium, describe all the observations recorded.

Equilibrium reactions versus reversible reactions

Equilibrium reactions and reversible reactions are similar in that they proceed in both directions, in which the conversion of reactants to products and the conversion of products to reactants occur simultaneously. The differences between equilibrium reactions and reversible reactions are summarised in Table 7.11.

Table 7.11: *Differences between equilibrium reactions and reversible reactions*

Equilibrium reactions	Reversible reactions
The rates of the forward and reverse reactions are equal.	The rates of the forward and reverse reactions are not necessarily equal.
They must be in a closed system for the equilibrium condition (state) to take place.	They do not have to take place in a closed system to be reversible.

Factors affecting the position of chemical equilibrium

The effects of different factors on chemical equilibrium positions were first investigated by a French chemist named Henri Louis Le Chatelier. He came up with a generalisation which is commonly known as Le Chatelier's Principle. *The Le Chatelier's Principle* states that, *if a system at equilibrium is subjected to a change, processes will occur which tend to counteract the change imposed*. The principle can also be stated as, *if a stress is applied to a system at equilibrium, the system re-adjusts, to reduce the effect of the stress imposed*.

The factors affecting the equilibrium position of a chemical reaction include; *concentration*, *pressure*, and *temperature*.

Concentration

On the basis of the collision theory, the more concentrated the reacting particles, the more often they collide with one another. This increases the rate of the reaction. Changing the concentrations of any reactants or products in a reaction at equilibrium affects the equilibrium position. When the concentration of a reactant is increased, the reaction proceeds in the forward direction to decrease the concentration of the reactant. When the concentration of a product is increased, the reaction proceeds in the reverse direction to decrease the concentration of the product.

Pressure

Pressure affects the equilibrium position of an equilibrium reaction that consists of gaseous components (reactants and products). An increase in pressure favours a

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reaction which forms fewer molecules or moles. A decrease in pressure favours a reaction which forms more molecules or moles. In a reaction where the number of molecules or moles of the reactants and the products are equal, change in pressure does not affect the state of the equilibrium. The equilibrium between nitrogen dioxide (NO_2) and dinitrogen tetraoxide (N_2O_4) gases can be used to show the effects of pressure on the state of equilibrium. The two gases are normally at equilibrium when they are in a closed system at room temperature.

$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$

Reddish brown — Colourless

Nitrogen dioxide is reddish brown when just prepared. When it is put in a syringe and allowed to attain the room temperature, it turns brown. At this point, the two gases (NO_2 and N_2O_4) exist together in a dynamic equilibrium. When the piston is pushed, the pressure is increased and the gas changes to colourless, the colour of dinitrogen tetraoxide. This shows that the formation of dinitrogen tetraoxide is favoured at high pressure. When the pressure is reduced, the colourless dinitrogen tetraoxide changes to reddish brown, the colour of nitrogen dioxide.

Temperature

The effects of temperature on the equilibrium position of a reaction depend on whether the reaction gives off heat (exothermic) or absorbs heat (endothermic) to and from the surroundings, respectively. An increase in temperature favours an endothermic reaction, while a decrease in temperature favours an exothermic reaction. Table 7.12 shows a summary of how the different factors affect the positions of equilibrium.

Table 7.12: *Factors that affect the position of equilibrium*

Factors	Changes	Effects
Temperature	Increase	Favours the reaction that absorbs heat (endothermic)
	Decrease	Favours the reaction that gives off heat (exothermic)
Pressure	Increase	Favours the reaction that forms fewer moles (volumes)
	Decrease	Favours the reaction that forms more moles (volumes)
Concentration	Increase	When concentration of the reactants is increased, the reaction proceeds in the forward direction to use up additional reactants.
	Decrease	When concentration of the reactants is decreased, the reaction proceeds in the reverse direction to use up the excess products.

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Chemical equilibria of industrial importance

The principles of chemical reaction rates and equilibrium play important roles in the design and working conditions of many industrial processes. Industrial plants are designed in such a way that they operate at the optimum production levels and at the least cost possible. This involves choosing the right conditions for obtaining reasonable yields. The knowledge of chemical equilibrium helps in the choice of the appropriate conditions.

The following are some industrial processes which use the principles of chemical equilibria.

The Haber process

The Haber process is used in the manufacture of ammonia from nitrogen and hydrogen gases. The process involves getting hydrogen from methane (natural gas) and reacting with atmospheric nitrogen. The equation for this equilibrium reaction is:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{heat}$$

Effects of temperature

At room temperature, the reaction is slow and the obvious solution would be to raise the temperature. Raising the temperature would favour the reverse reaction, since the reaction is exothermic. Decreasing the temperature would therefore cause the equilibrium to move to the right, resulting in a higher yield of ammonia. However, at low temperatures, the reaction is slow, which means it takes longer time to achieve an equilibrium mixture which contains a very high proportion of ammonia. Thus, a temperature of about 500 ºC is used in the production of ammonia. This temperature is high enough to enable the production of ammonia to proceed at an acceptable rate. To rectify the negative effects of temperature, a favourable pressure is applied to the system as discussed in the next section.

Effects of pressure

From the balanced chemical equation of the Haber process, there is a decrease in the number of moles in the forward reaction. The forward reaction is therefore favoured by high pressure. Increasing pressure causes the equilibrium to move to the right, resulting in a higher yield of ammonia. At 200 °C and pressures above 750 atmosphere (atm), there is an almost 100% conversion of the reactants to ammonia gas. However, it is difficult and expensive to maintain such a high pressure. Thus, a low pressure of about 200 atm is usually used. At a pressure of around 200 atmospheres and a temperature of about 500 ºC, the rate of production of ammonia is 10–20%. This gives an acceptable yield of ammonia within an acceptable period of time. This ensures that the costs and safety concerns during the operation are compromised. Usually, an iron catalyst is also used to speed up the rate of the reaction. However, the catalyst has no effect whatsoever on the position of the equilibrium. The Haber process for the manufacture of ammonia is illustrated in Figure 7.17.

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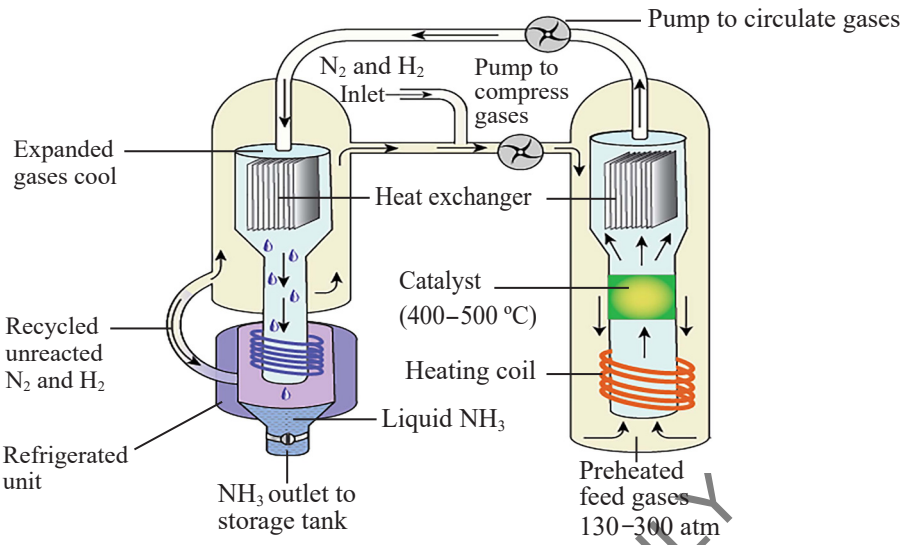


Figure 7.17: *Diagram of the Haber process for the manufacture of ammonia*

The Contact process

The Contact process is used in the industrial manufacture of sulphuric acid. The most important stage in the Contact process is the conversion of sulphur dioxide (SO_2) to sulphur trioxide (SO_3), which is an equilibrium reaction.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + \text{heat}$$

Effects of pressure

From the mole or volume ratio of reactants to products, there is a larger number of moles of the reactants than that of the products. The forward reaction, that is, the conversion of sulphur dioxide to sulphur trioxide is therefore favoured by high pressure. However, just like in the Haber process, high pressure is expensive and difficult to maintain. In the Contact process, the conversion of sulphur dioxide to sulphur trioxide is done at the atmospheric pressure. At this pressure, the conversion of sulphur dioxide to sulphur trioxide is quite high (over 90%). It would therefore not be economical to increase the pressure.

Effects of temperature

Since the reaction is exothermic, the conversion of sulphur dioxide to sulphur trioxide is favoured by low temperatures. However, at low temperatures, the rate of reaction is very slow. A compromise temperature of between 400 ºC and 500 ºC is used. This temperature is high enough to enable the reaction to take place at an acceptable rate. The most commonly used temperature in the Contact Process is 450 ºC at normal atmospheric pressure. The conversion rate of the reactants to products under these conditions is 97%. To attain a faster equilibrium, vanadium pentoxide or platinum is used as a catalyst. Figure 7.18 shows a diagram of the Contact process for the manufacture of sulphuric acid.

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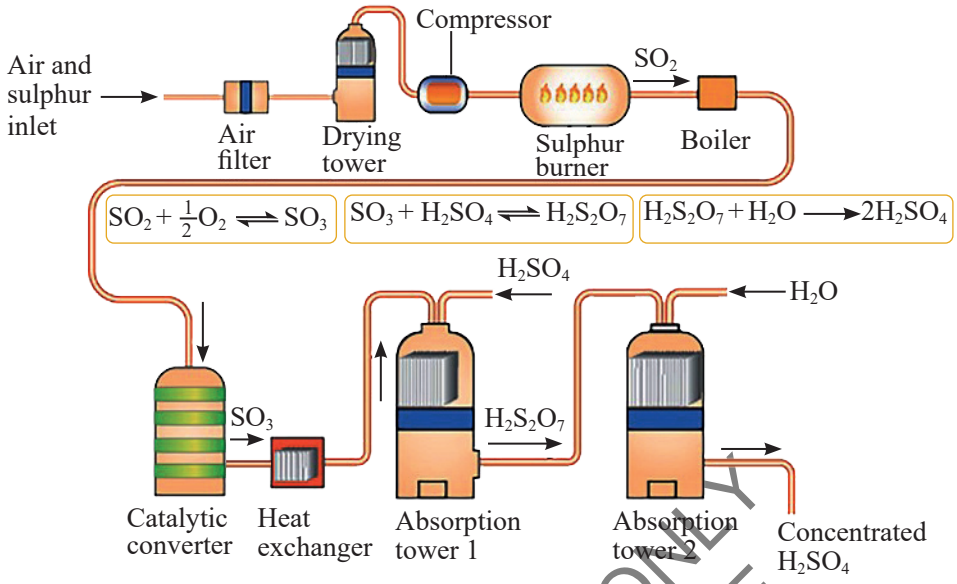


Figure 7.18: *Diagram of the Contact process for the manufacture of sulphuric acid*

Exercise 7.3

1. Explain the meaning of each of the following terms:
   (a) Reversible reaction
   (b) Dynamic equilibrium
   (c) Activation energy
2. How is a reversible reaction distinguished from an irreversible reaction?
3. The following equation shows a reaction between oxygen and hydrogen gases to form water:

   $2H_2(g) + O_2(g) \rightleftharpoons 2H_2O(g) + \text{heat}$

   What would happen if pressure is increased?
4. The following equation shows the reaction between hydrogen and iodine gases to form hydrogen iodide gas:

   $H_2(g) + I_2(g) \rightleftharpoons 2HI(g) + \text{heat}$

   Explain what would happen to the position of equilibrium if:
   (a) temperature is increased.
   (b) pressure is lowered.
   (c) hydrogen iodide gas is pumped into the system.

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Chemical energetics

In many chemical reactions, energy is converted to heat energy and sometimes to light energy. However, not all chemical changes are accompanied with the production of heat. Some chemical changes absorb heat. The study of energy changes that accompany chemical reactions is called *thermochemistry*. Molecules have *kinetic energy* due to the vibrational and rotational motions of the particles that make them. In addition to the kinetic energy, molecules also have *potential energy* due to their positions and composition. The potential energy is stored in molecular bonds that exist within molecules (*intramolecular*), and also between different molecules (*intermolecular*).

For example, in water there is energy stored in the bonds between oxygen and hydrogen as shown in Figure 7.19. This happens between oxygen and the two hydrogen atoms in each molecule (intramolecular energy), and between the oxygen atom of one molecule and one of the hydrogen atoms of another water molecule (intermolecular energy). The intermolecular bond in water molecules is also known as *hydrogen bond*.

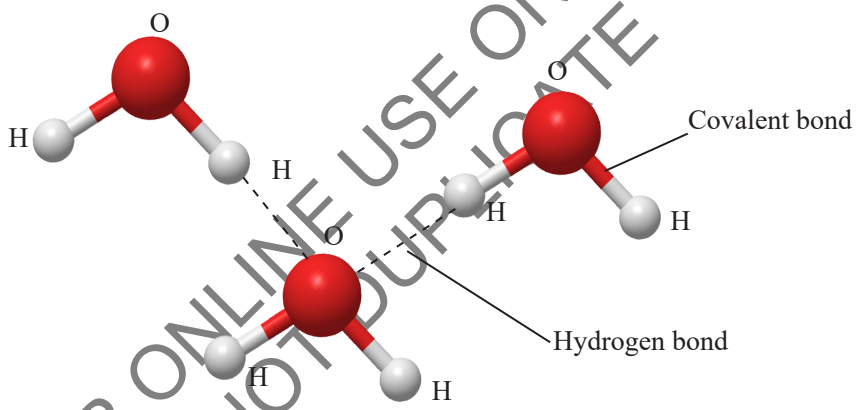


Figure 7.19: *Molecular bonds in water molecules*

The sum of the kinetic and potential energies of a substance is known as *heat content* or *enthalpy* (H). The unit of H is joule (J).

Chemical reactions involve the rearrangements of atoms and molecules. The rearrangements are achieved by breaking bonds in the reactant molecules and forming new bonds in the product molecules. Energy is required to break the bonds, whereas energy is given out when new bonds are formed. The difference between the energy released and the energy absorbed in a chemical reaction is the energy change for the reaction. This energy change is usually referred to as *enthalpy change*.

Enthalpy changes in chemical reactions

The enthalpy change in a chemical reaction is the energy gained or lost during a chemical reaction carried out at constant pressure. It is represented by the symbol ΔH (read as 'delta H').

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The units for ΔH are kilojoule per mole (kJ mol^{-1}). The enthalpy change in a chemical reaction is the difference between the heat contents of the reactants and that of the products, that is;

ΔH = heat content of the products (H_P) – heat content of the reactants (H_R).

$\Delta H = H_P - H_R$

The value of ΔH is usually shown at the end of a chemical equation as;

$$A + B \longrightarrow C + D \qquad \Delta H = x \text{ kJ mol}^{-1}$$

where x is the value for ΔH.

A chemical equation that includes the enthalpy change of the reaction is referred to as a *thermochemical equation.* Enthalpy change is directly proportional to the amounts of substances taking part in the reaction. This means that the enthalpy change will double if the amount (number of moles) of the reactants is doubled.

Exothermic and endothermic reactions

Energy changes occur in chemical reactions and in some physical processes, such as a solid (ice) melting into a liquid (water). In every chemical change, reactants are converted into products through a chemical reaction. As discussed earlier, chemical reactions happen when bonds between atoms in the substance are broken; the atoms rearrange, and new bonds between atoms are formed to make new substances. Breaking bonds between atoms requires energy, while creating new bonds releases energy. If a reaction absorbs or uses more energy than it releases, the reaction is *endothermic*. On the other hand, if a reaction releases more energy than it absorbs, the reaction is *exothermic.*

Consider a chemical reaction in which molecules A and B react to form molecules C and D.

$$A + B \longrightarrow C + D$$

Reactants Products

If the sum of the heat content of molecules A and B is more than the sum of the heat content of molecules C and D, the excess energy is released to the surroundings. Such a reaction is referred to as an exothermic reaction. *An exothermic reaction* is a reaction in which energy is lost from the reaction mixture. The energy is usually transferred as heat energy, causing the reaction mixture and its surroundings to become hotter.

For an exothermic reaction, H_R is greater than H_P, and therefore the value of ΔH is negative. This shows that energy has been lost from the reaction mixture. Examples of exothermic reactions are combustion reactions, mixing of water and strong acids such as concentrated sulphuric acid, the neutralisation reaction between acid and alkali, and respiration.

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If the heat content of molecules C and D (products) is greater than the heat content of molecules A and B (reactants), the extra energy required is absorbed from the surroundings. Such a reaction is an endothermic reaction. An *endothermic reaction* is a reaction in which the energy content of the products is more than that of the reactants. Since the energy is transferred in the form of heat, the reaction mixture and its surroundings become colder.

For an endothermic reaction, H_p is greater than H_R, and therefore the value of ΔH is positive. This shows that heat has been gained in the reaction mixture. Some examples of endothermic reactions are photosynthesis, dissolving ammonium chloride in water, mixing water and potassium chloride, and thermal decomposition of calcium carbonate.

Reversible reactions are both endothermic and exothermic. If the forward reaction is exothermic, then the reverse reaction is endothermic. The magnitude of ΔH for the two processes is the same. Only the signs differ (negative for the exothermic reaction and positive for the endothermic reaction).

Consider the reaction between sulphur dioxide and oxygen gases in the manufacture of sulphuric acid.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H = -196 \text{ kJ mol}^{-1}$$

From the thermochemical equation, the forward reaction is exothermic. Therefore, the reverse reaction is endothermic with the same value of ΔH, but different in sign. That is:

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g) \qquad \Delta H = +196 \text{ kJ mol}^{-1}$$

The characteristics of exothermic and endothermic reactions are summarised in Table 7.13.

Table 7.13: *Characteristics of exothermic and endothermic reactions*

Exothermic reaction	Endothermic reaction
Heat is given out and transferred from the chemical reaction to the surrounding environment.	Heat energy is absorbed and transferred from the surrounding environment to the reactants.
The temperature of the mixture rises, the container will therefore be hot.	The temperature of the mixture falls; the container will therefore be cold.

Energy level diagrams in chemical reactions

The energy change in a chemical reaction is presented graphically using a type of graph called an *energy level diagram*. The vertical axis (Y) of the energy level diagram represents the energy, while the horizontal axis (X) represents the reaction path as

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it progresses from the reactants to form the products. Energy level diagrams are used to show the total energy content of the reactants in relation to the total energy content of the products. They also show the energy difference between the reactants and the products, ΔH as the heat of reaction.

In an exothermic reaction, the reactants are at a higher energy level than the products. The difference between the two energy levels is the energy released to the surroundings during the reaction. This is shown as a vertical drop in the energy level diagram (Figure 7.20).

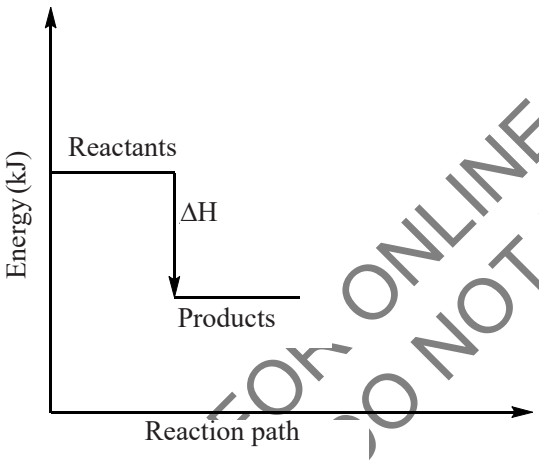


Figure 7.20: *Energy level diagram for an exothermic reaction*

In an endothermic reaction, the reactants are at a lower energy level than the products. The difference between the two energy levels is the energy gained from the surroundings during the reaction. This energy is represented by a vertical rise in an energy level diagram as shown in Figure 7.21.

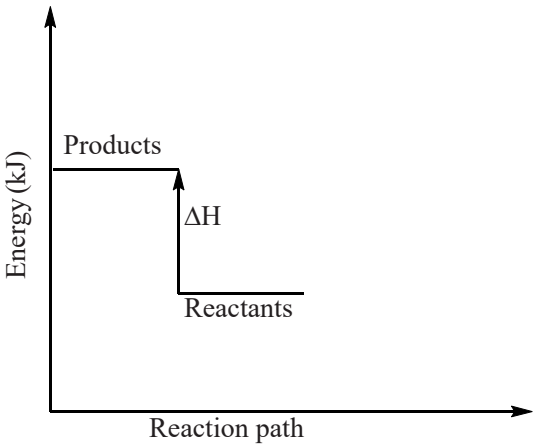


Figure 7.21: *Energy level diagram for an endothermic reaction*

For example, the dissolution of ammonium nitrate in water is an endothermic reaction. The temperature of the ammonium nitrate solution is less than that of the water before the dissolution of NH_4NO_3. The energy level diagram for the dissolution of ammonium nitrate can be presented as shown in Figure 7.22.

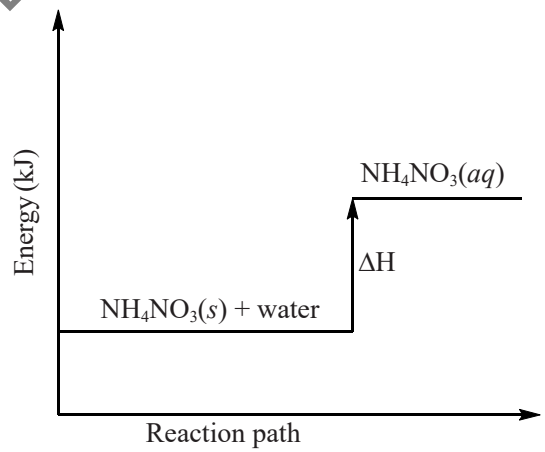


Figure 7.22: *Energy level diagram for the dissolution of ammonium nitrate*

Many cold packs that are used for relieving pain make use of the reaction of ammonium nitrate with water. Such

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packs come in a plastic bag that contains ammonium nitrate crystals and a smaller bag filled with water. The smaller bag is made of a thin weak plastic that is easy to break. To use the cold pack, one tears the inner bag. This releases water, which dissolves the ammonium nitrate in an endothermic reaction. The extra heat needed for the reaction is absorbed from the surrounding, quickly lowering the pack's temperature. Cold packs are a convenient replacement for crushed ice that is used as first aid for relieving pain and can be carried as first aid to remote areas where ice is unavailable.

The dissolution of concentrated sulphuric acid in water is a highly exothermic reaction in which a lot of heat is given out. When concentrated sulphuric acid is dissolved in water, the temperature of the mixture is higher than that of water before adding concentrated sulphuric acid. The diagram in Figure 7.23 shows an energy level diagram for the dissolution of concentrated sulphuric acid.

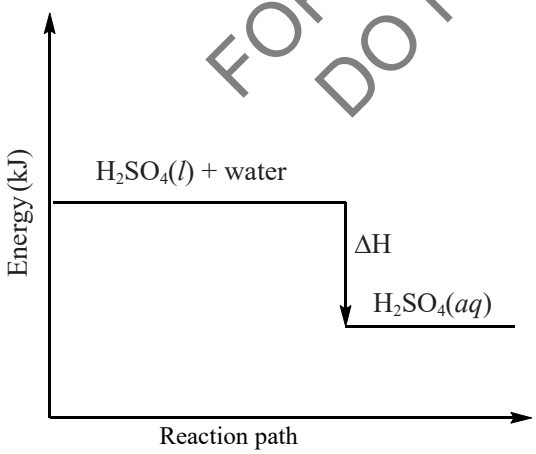


Figure 7.23: *Energy level diagram for the dissolution of concentrated sulphuric acid*

Since both exothermic and endothermic reactions cause changes in temperature of the reaction mixture and its surroundings, one can easily classify chemical reactions as exothermic or endothermic reactions.

Example 7.1

A solution of sodium hydroxide is added to a solution of hydrochloric acid.

(a) Given that both solutions are at room temperature, what type of reaction takes place if the final steady temperature of the reaction mixture is 36 ºC?

(b) Draw an energy level diagram for the reaction.

Solution

(a) The reaction is exothermic since there is a temperature rise in the reaction mixture. The heat is given out to the surroundings.

(b) The energy diagram for this reaction is shown in Figure 7.24.

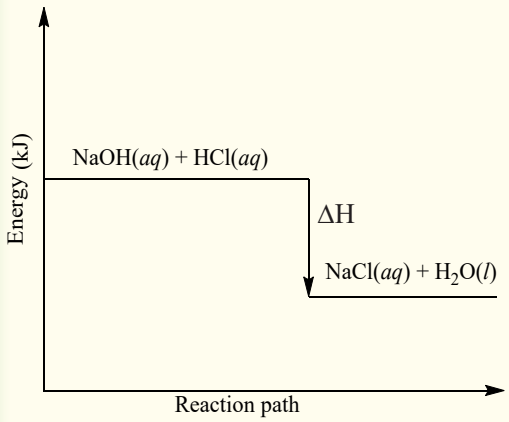


Figure 7.24: *Energy level diagram for the reaction between NaOH and HCl*

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Activity 7.12

Aim: To investigate heat changes when ammonium nitrate is dissolved in water

Requirements: Thermometer, 250-mL beaker, ammonium nitrate, and distilled water

Procedure

1. Put 100 cm^3 of distilled water in a beaker.
2. Insert the thermometer into the water as shown in Figure 7.25, and note the temperature of water.

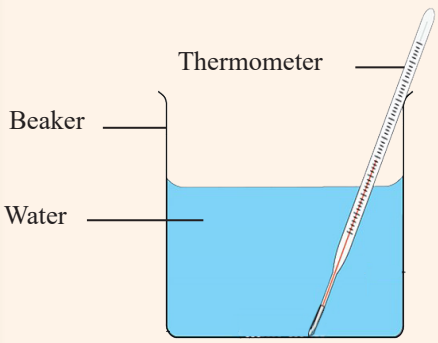


Figure 7.25: *Set-up for the investigation of heat changes when ammonium nitrate is dissolved in water*

3. Dissolve about 4 g of ammonium nitrate into the beaker containing the distilled water.
4. Stir the mixture gently by using the thermometer to dissolve the ammonium nitrate.
5. When all the ammonium nitrate has dissolved, note the temperature reading on thermometer.

Questions

1. Compare the initial temperature of the distilled water with that of the solution of ammonium nitrate.
2. What type of a chemical reaction takes place when ammonium nitrate is dissolved in water? Explain
3. Explain the reason for the changes in temperature observed in the experiment.

Activity 7.13

Aim: To investigate heat changes when concentrated sulphuric acid is dissolved in water

Requirements: 250-mL beaker, a test tube, a thermometer, concentrated sulphuric acid, and distilled water

Caution! The reaction between concentrated sulphuric acid and water is highly exothermic.

Procedure

1. Put about 100 cm^3 of distilled water into the beaker.
2. Dip the thermometer into the water and record the initial temperature of water.
3. Put 2 to 3 cm^3 of concentrated sulphuric acid into a test tube.

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4. Pour the acid carefully and slowly into the water from the side of the beaker.

5. Stir the mixture using the thermometer and note the final temperature of the mixture.

Questions

1. Compare the temperature of the distilled water with that of the acid solution.

2. Explain the reason for the change in temperature of the distilled water after the sulphuric acid was added.

Chapter summary

1. Chemical kinetics is the study of the rates of chemical reactions.

2. The rate of a chemical reaction is a measure of the progress of the reaction per unit time. The rate of a chemical reaction can be expressed as the rate of decrease of a reactant or the rate of increase of a product.

3. According to the collision theory, chemical reactions occur as a result of collisions between the reactants' particles. However, only particles which possess the minimum activation energy and have proper orientation result in effective collisions and hence chemical reactions.

4. The main factors that affect the rate of a chemical reaction are temperature, pressure, concentration, surface area, and presence of a catalyst.

5. Reversible reactions are those reactions that proceed in both the forward and reverse directions.

6. Irreversible reactions are those reactions that proceed to completion in the forward direction only.

7. In a chemical equilibrium, both reactions (forward and reverse) take place at equal rates.

8. Chemical equilibrium is described as a dynamic process because there are moments in which the forward and reverse reactions occur at the same rate to reach a point where the amount of the reactants and products remain unchanged with time.

9. The main factors that affect the position of equilibrium of a reaction are concentration, pressure, and temperature.

10. The Haber process and the Contact process are examples of chemical equilibria of industrial importance.

11. Chemical reactions are either exothermic or endothermic. Exothermic reactions give out heat, while endothermic reactions absorb heat.

12. Energy level diagrams are used to show the energy changes in chemical reactions.

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Revision exercise 7

Choose the correct answer in items 1–8.

1. Which of the following will lower the activation energy of a reaction?
   (a) Increasing pressure
   (b) Adding a catalyst for the reaction
   (c) Raising the temperature of the reaction
   (d) Increasing the pressure and the concentrations of reactants

2. Why a burning splint burns more vigorously in pure oxygen than in air?
   (a) Oxygen is a catalyst for combustion.
   (b) Oxygen is a product of combustion.
   (c) The concentration of oxygen is higher in pure oxygen than it is in air.
   (d) The low concentration in pure oxygen catalyses the combustion reaction.

3. Which of the following units can be used to express the rate of a reaction?
   (a) mL/s
   (b) mL/g
   (c) g/mL
   (d) mL/mol

4. Consider the following reaction:

   $Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

   Which of the following can be used to increase the rate of the production of zinc chloride?
   (a) Increasing the surface area of zinc
   (b) Increasing the volume of hydrogen
   (c) Increasing the concentration of zinc
   (d) Decreasing the concentration of hydrochloric acid

5. Which of the following factors affects the rate of heterogeneous reactions only?
   (a) Temperature
   (b) Nature of the reactants
   (c) Surface area of reactants
   (d) Concentrations of reactant

6. Magnesium reacts with hydrochloric acid according to the following equation:

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

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Which of the following properties could be used to measure the rate of this reaction carried out in an open container?

(a) Mass of magnesium

(b) Volume of hydrogen

(c) Colour of the solution

(d) Solubility of hydrochloric acid

7. Consider the conditions summarised in the table for the reaction of 1.0 g of powdered marble and hydrochloric acid solution:

$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

Experiment	Temperature (°C)	Concentration of HCl (M)
1	40	3.0
2	20	3.0
3	40	6.0

What will be the rate of the reaction in the order from the fastest to slowest?

(a) 1, 2, 3.

(b) 2, 1, 3.

(c) 3, 1, 2.

(d) 3, 2, 1.

8. Given the following thermochemical equation:

$4NH_3(g) + 7O_2(g) \rightleftharpoons 4NO_2(g) + 6H_2O(g) \qquad \Delta H = -1130 \text{ kJ}$

Which of the following sets of conditions will give the greatest yield of NO_2?

(a) Low pressure and low temperature

(b) Low pressure and high temperature

(c) High pressure and temperature

(d) High pressure and low temperature

9. What does the statement '*a reaction has reached equilibrium*' mean with regard to the following?

(a) Forward and reverse reaction rates

(b) The amounts or concentrations of the reactants and the products

10. A chemical equilibrium is described as a dynamic process. Explain.

11. A 0.124 mol sample of reactant A was placed in a flask and allowed to react to form product B according to the following reaction equation:

$A(g) \longrightarrow B(g)$

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The following data were obtained for the amount of reactant A during the reaction:

Time (s)	0	10	20	30	40
Moles of A	0.124	0.110	0.088	0.073	0.054

Calculate the average rate of disappearance of A between 10 s and 20 s.

12. Calcium reacts with water according to the following reaction equation:

$Ca(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$

At a certain temperature, 2.05 g of calcium reacted completely in 30.0 seconds. Calculate the rate of consumption of calcium.

13. A form three student wishes to monitor the rate of the following reaction:

$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

One of the properties that could be used to monitor the reaction rate is the mass of $CaCO_3$ which decreases as the reactants are converted to products. Identify other four properties that could be used to monitor the rate of this reaction, and describe the changes that would occur.

14. The manufacture of ammonia by Haber process is an example of a chemical equilibrium of industrial importance.

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92 \text{ kJ mol}^{-1}$

Describe the effects of temperature and pressure on this reaction.

15. The following are thermochemical equations for some reactions:

(a) $C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -393.5 \text{ kJ mol}^{-1}$

(b) $2C(s) + H_2(g) \longrightarrow C_2H_2(g) \quad \Delta H = +53.2 \text{ kJ mol}^{-1}$

(i) Explain whether the two reactions are endothermic or exothermic.

(ii) Draw an energy level diagram for each of the reactions.

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Chapter Eight

Extraction and properties of metals

Introduction

In your day to day life activities, you see and probably use substances such as spoons, kitchen utensils, knives, razor blades, chains, aluminium foil, and wires. These and many other substances are made up of different substances, mainly metals. Metals are found naturally in the Earth's crust, and some of them exist in combination with other elements like oxygen, sulphur, and carbon. To make these metals useful, they have to be extracted from their ores and purified. In this chapter, you will learn the concept of metals and their importance, occurrence and abundances of metals in the Earth's crust, occurrence and location of metals in Tanzania, properties of metals, extraction of metals, and environmental considerations regarding extraction of metals. The competencies developed will enable you to identify metal constituents in different materials in your environment in which metals can be extracted, and manipulate the uses of different metal-based substances for different purposes.

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Concept of metals and their importance

A metal is an element that readily loses electron(s) to form a positive ion. For monovalent metals the reaction is;

$$M(s) \longrightarrow M^{+}(aq) + e^{-}$$

For example,

$$Na(s) \longrightarrow Na^{+}(aq) + e^{-}$$

For divalent metals, the reaction is;

$$M(s) \longrightarrow M^{2+}(aq) + 2e^{-}$$

For example,

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^{-}$$

Metal compounds occur in the Earth's crust along with some impurities such as rock particles. These compounds are usually referred to as *minerals*. A mineral is a naturally occurring inorganic solid, with a definite chemical composition and an ordered atomic arrangement. A mineral that contains large proportions of metals or metallic compounds worth to be extracted and has an economic value is referred to as an *ore*.

Metals exist in nature either as free elements or as compounds. Some of the metals that exist as free elements include silver, gold, and platinum.

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These elements are the least reactive, hence do not form compounds as others. The compounds of metals can be oxides, carbonates, sulphates, or complex compounds made from two or more simple compounds.

Metals are widely available on the Earth and are commonly used in daily life activities. Many materials such as mobile phones, computers, electric wires, bridges, nails, planes, cooking pots, rails, door locks, and earings are made of metals. Metals are tremendously important. Their roles in economic, social, cultural, and technological development are exceptionally high. Metals are used in different areas such as; in agriculture, industries, automobiles, security systems, drilling and construction, buildings and bridges, electronics and communication, furniture, jewellery, clothing, electric lightening, medicine, and medical instruments.

Occurrence and abundances of metals in the Earth's crust

There are different elements in the Earth's crust. Aluminium makes about 8% of the Earth's crust and is the most abundant metal, followed by iron which constitutes about 6%. The abundances of calcium, sodium, potassium, and magnesium in the Earth's crust are 3.6%, 2.8%, 2.0%, and 2.3%, respectively. Other metals found in low abundances in the Earth's crust include the following: titanium (0.44%), manganese (0.095%), barium (0.043%), chromium (0.61%), nickel (0.01%), copper (0.006%), zinc (0.007%), cobalt (0.003%), lead (0.0013%), uranium (0.0003%), gold ($<5 \times 10^{-6}$%), mercury ($<8 \times 10^{-6}$%), and silver ($<7 \times 10^{-6}$%), The remaining metals make about 3% of the Earth's crust.

Occurrence and locations of metals in Tanzania

Tanzania is among the countries that have abundant mineral resources. The minerals found in Tanzania are categorised into metallic, gemstones, industrial, energy source, and construction minerals. The metallic minerals include gold, iron, silver, cobalt, aluminium, uranium, copper, and nickel. Gemstones include tanzanite, diamond, garnet, and ruby, among many others. Examples of industrial minerals found in Tanzania are gypsum, limestone, phosphates, soda ash, and salt. Energy source minerals include coal and uranium. Gravel, dimension stones, and sand are in the category of construction minerals.

Different metals are found in different places in Tanzania. For example, gold is found in Geita, Kahama, Buhemba (Butiama), Biharamulo, Nyamongo (Tarime), Majimoto (Ngara), Mananila, Magambazi and Lupa (Chunya), Nzega, Handeni, and Ikungi. Iron reserves are found in Liganga (Ludewa and Mbinga), Uluguru mountains, Mbabala near Lake Tanganyika, Karema, Manyara, and Itewe. The Liganga iron ore mine holds the biggest iron resources in Tanzania. Nickel is found in Kabanga (Ngara), aluminium is found in Mabughai-Mlomboza (Lushoto), and copper is found in Mpanda. Uranium is found in

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Mkuju (Namtumbo), Bahi, Galapo, Minjingu, Mbulu, Simanjiro, Manyoni, Songea, Tunduru, Madaba, and Nachingwea. Other minerals include common salt found in Uvinza, tanzanite in Merelani, diamond in Mwadui and ruby in Mahenge (Ulanga). The map in Figure 8.1 shows the distribution of major mineral resources in Tanzania.

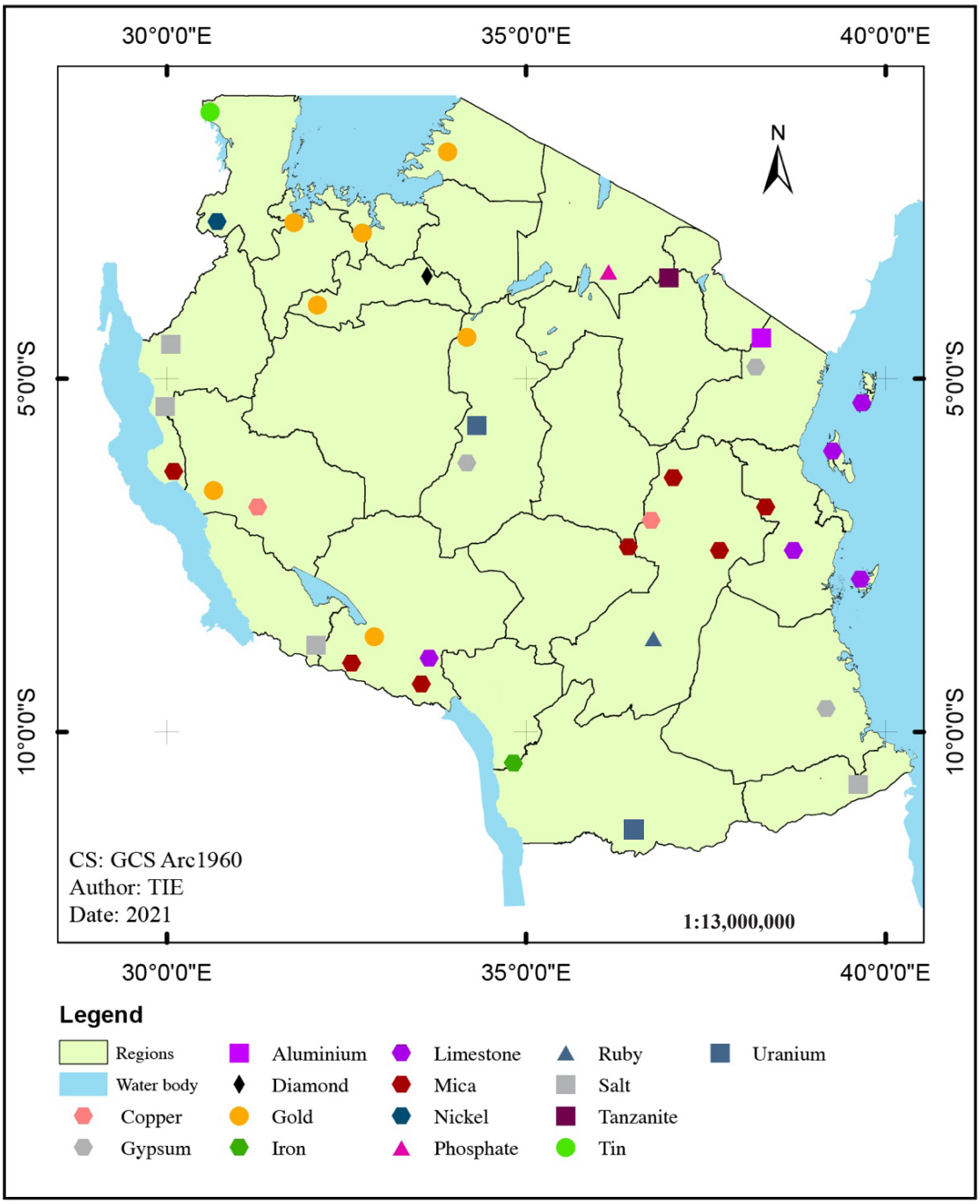


Figure 8.1: *Distribution of major mineral resources in Tanzania*

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Task 8.1

Tanzania is rich in minerals which are found in different regions. Use various sources to:

(a) Identify the districts and regions where the following minerals are found: copper, ruby, mica, gold, aluminium, limestone, tin, iron, tanzanite, and phosphate.

(b) Comment on the abundance of the minerals at the sites where they are found in Tanzania.

Properties of metals

Pure metals have specific physical and chemical properties. Their physical properties include physical strength, colour, melting point, boiling point, and solubility. The chemical properties include their ability to react with other substances.

Physical properties of metals

Generally, metals exhibit the following physical characteristics:

Appearance: They have a shiny lustre, especially when freshly cut. The lustre of a sharpened end of an axe and knife exemplify this property.

Ductility: They are ductile, which means they can be stretched into thin wires or rollers. Copper wires, tungsten wires and iron rollers are examples of such thin materials made due to the ductility of metals.

Malleability: They are malleable as they can be hammered into thin sheets without breaking. The corrugated iron sheets are examples of such sheets from a malleable metal.

Conductivity: They are good conductors of heat and electricity. They can be used in making cables for conducting electricity.

Melting points: They are solids with high melting points with the exception of mercury which is a liquid metal. The high melting points of metals are due to strong forces of attraction among the metal atoms.

Sonority: Some metals are sonorous. A sonorous substance is the one that is capable of producing a deep ringing sound when hit. For a metal to suit as a construction material for a bell, it must be sonorous. Examples of sonorous metals are aluminium, iron, and silver.

Strength: They are strong and tough with high tensile strength. Tensile strength is the resistance of a material to breaking under tension. It includes the resistance against bending. For a metal to qualify as a building material for bridges, it must have a high tensile strength.

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Chemical properties of metals

Metals are electropositive elements. In chemical reactions, metals donate electrons to other elements. The tendency of an element to lose electron and form a positive ion is called *electropositivity*. In solutions, metals form cations allowing them to become stable.

Chemical strengths versus physical strengths of metals

Metals have different chemical and physical strengths as discussed hereunder.

Physical strength

The physical strength of a metal is the ability to withstand the force applied on it without breaking. The physical strength of a metal is referred to as its *tensile strength*. Iron has a very high tensile strength, while sodium has a low tensile strength. Metals which are physically strong are not necessarily chemically strong, and metals which are chemically strong are not necessarily physically strong.

Activity 8.1



Aim: To investigate the physical strengths of sodium and iron metals

Requirements: Sharp knife, pair of tongs, watch glass, sodium metal, iron nail, and beaker

Procedure

1. Cut a very small piece of sodium using a dry sharp knife. This should be done by holding the sodium metal by using a pair of tongs, and the cutting should be done on a dry watch glass. Return the pieces of sodium to the oil and seal the container.
2. Try to cut an iron nail by using the same sharp knife, and observe if a knife can easily cut the iron nail.

Caution! Sodium is a very reactive metal and its reaction with oxygen and moisture may cause explosion. Wear protective gloves and goggles.

Question

Which metal was difficult to cut?

Chemical strength

The chemical strength of a metal is a measure of how readily the metal takes part in chemical reactions. Some metals are very reactive hence react rapidly to produce new substances. Potassium and sodium are some of the most reactive metals. Metals such as platinum, gold and silver are very unreactive. Such metals do not release their electrons easily. The reactivity of metals has potential roles towards the methods that are used in their extraction.

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Activity 8.2

Aim: To investigate the reactivity of different metals

Requirements: Bunsen burner, measuring cylinder, retort stand and clamp, beaker, test tubes, boiling tube, trough, delivery tube, cotton wool, calcium metal, iron filings, magnesium ribbon, aluminium foil, copper turnings, and distilled water

Procedure

1. Label four test tubes as A, B, C, and D.
2. Transfer about 10 cm^3 of cold distilled water in each of the test tubes.
3. (a) Put a lump of calcium metal in test tube A.
   (b) Put some iron filings in test tube B.
   (c) Put some copper turnings in test tube C.
   (d) Put some magnesium ribbon in test tube D.

   Make your observations in each of the test tubes.
4. Repeat Steps 2 and 3 by using hot water.
5. Put some cotton wool soaked in water in a boiling tube, and then insert a magnesium ribbon in the tube as shown in Figure 8.2.
6. Heat the boiling tube gently, then strongly. Record any observation. Observe if there is any water bubbling in the trough.
7. Repeat Steps 5 and 6 by using an aluminium foil instead of the magnesium ribbon, then with iron filings and finally with copper turnings. Record all your observations.

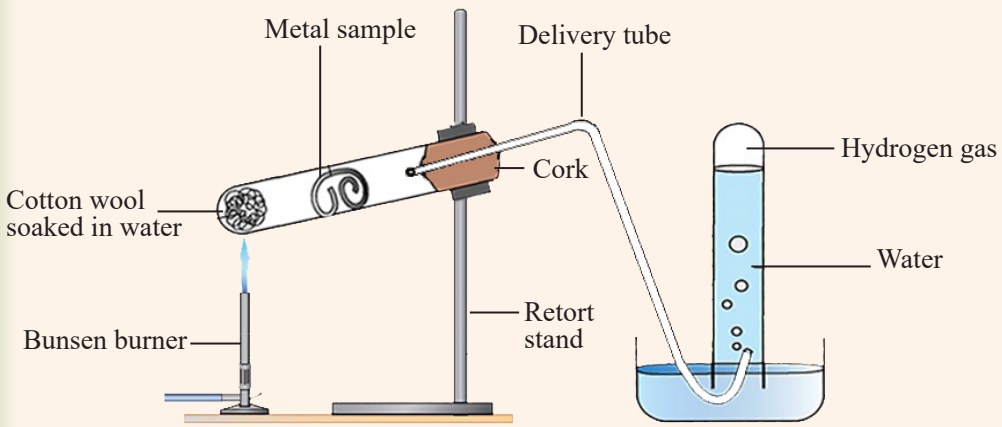


Figure 8.2: *Set-up for the action of steam on a metal*

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Questions

1. Based on your observations, compare the reactivities of calcium, iron, copper, and magnesium in cold and hot water.
2. What happened when the metals mentioned in Question 1 were treated with steam?
3. Comment on the reactivities of magnesium and aluminium.

Redox reactions involving metals

In redox reactions, oxidation and reduction take place simultaneously. *Oxidation* is a process which involves the loss of electron(s) from a substance. For example, sodium metal loses an electron to form sodium ion.

$$Na(s) \longrightarrow Na^+(g) + e^-$$

In this reaction, sodium has been oxidised to sodium ion. An equation which represents the loss of electron(s) is referred to as an *oxidation half reaction.* On the other hand, *reduction* involves the gaining of electron(s). For example, chlorine gas can be reduced to chloride ions by gaining electrons.

$$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(g)$$

An equation that represents a reduction reaction is referred to as a *reduction half reaction*. A *redox reaction* is therefore made up of the two half reactions, one representing oxidation and the other representing reduction.

Consider the reaction between sodium metal and chlorine gas to form sodium chloride.

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

This is a redox reaction and its half reactions are:

$$2Na(s) \longrightarrow Na^+(g) + 2e^- \quad \text{(oxidation)}$$

$$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(g) \quad \text{(reduction)}$$

The electrons lost by sodium are gained by chlorine. The substance that loses electrons is said to be a *reducing agent,* while the substance that gains electrons is said to be an *oxidising agent.* In the given reaction, sodium is a reducing agent, while chlorine is an oxidising agent.

The reducing powers of metals

In chemical reactions, metals tend to lose their outermost shell electron(s). The ease with which metals lose electrons increases down the group within the Periodic Table. Potassium with an electronic configuration of 2:8:8:1 is a stronger reducing agent than

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lithium which has an electronic configuration of 2:1. This is because the number of electron shells increases down the group causing the outermost electrons to be weakly attracted to the nucleus. The outermost electrons are therefore easily lost during chemical reactions. The reducing powers of metals thus increase down the group.

Across the period, that means from left to right, the number of electrons in the outermost shell increases. The number of protons also increases because the number of protons is equal to the number of electrons in a neutral atom. However, the number of shells does not change. This results in stronger attractions of the electrons in the outermost shell to the protons in the nucleus. This makes it difficult for the electrons to be lost during chemical reactions. The reducing powers of metals, therefore, decrease across the period from left to right. For example, sodium (2:8:1) is a stronger reducing agent than aluminium (2:8:3). Figure 8.3 shows the trends in the reducing powers of metals down the group and across the periods.

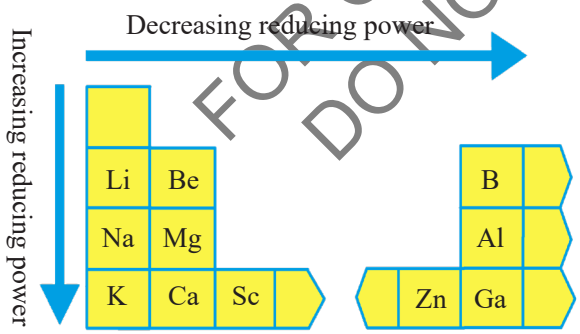


Figure 8.3: *Trends in the reducing powers of metals*

Reactivity series of metals

The reducing power of a metal determines its reactivity. The higher the reducing power, the more reactive the metal is. The arrangement of metals according to their reducing powers or reactivity is known as the *reactivity series of metals.* At the top of the reactivity series are metals with the highest reducing powers. The reactivity (activity) series of some metals is shown in Figure 8.4.

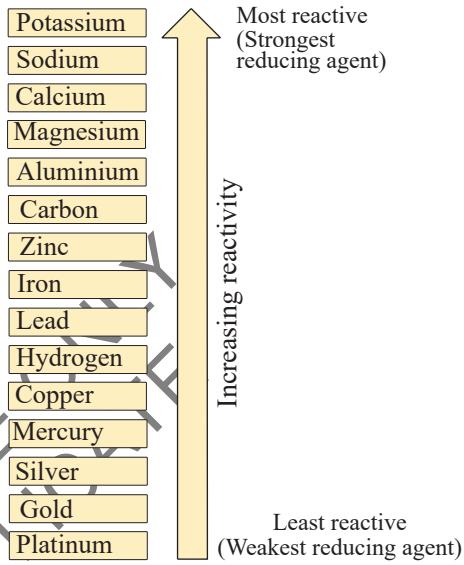


Figure 8.4: *Reactivity series of some metals*

Displacement reactions involving metals

A displacement reaction takes place when an atom or group of atoms takes the position of another in a compound. Consider the reaction between iron metal and hydrochloric acid to form iron(III) chloride.

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$$2Fe(s) + 6HCl(aq) \longrightarrow 2FeCl_3(aq) + 3H_2(g)$$

In this reaction, iron takes the place of hydrogen in the hydrochloric acid. Therefore, iron has displaced hydrogen from the hydrochloric acid. A more reactive metal displaces a less reactive metal from a solution of one of its compounds. It is therefore possible to predict whether a metal will displace another metal from its solution by comparing their positions in the reactivity series. For example, sodium will displace copper from copper(II) sulphate solution.

Another example is, when zinc is added to copper(II) sulphate solution and stirred, a brown solid is formed. The beaker becomes warm indicating that the reaction is exothermic. The brown solid is copper metal.

$$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$$

Zinc metal dissolves to form zinc ions (Zn^{2+}) which displace copper(II) ions from the copper(II) sulphate solution. In this reaction, the zinc metal acts as a reducing agent, reducing copper(II) ions to copper metal.

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

This reaction is possible because zinc is a stronger reducing agent than copper. On the other hand, Cu^{2+} in copper(II) sulphate acts as a stronger oxidising agent and oxidises zinc metal to zinc ions.

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$

If the reaction is left for a longer time, the blue colour caused by copper ions fades. If the zinc powder is in excess, the solution will be decolourised completely.

Some possible and impossible displacement reactions involving metals and metal ions are shown in Table 8.1. Some of the displacement reactions involve observable colour changes. The reaction between iron and copper(II) sulphate is a good example. The blue colour of the copper(II) sulphate changes to pale green which is the colour of the iron(II) sulphate solution.

$$\underset{\text{Blue}}{Fe(s) + CuSO_4(aq)} \longrightarrow \underset{\text{Pale green}}{FeSO_4(aq) + Cu(s)}$$

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Table 8.1: *Possible and impossible displacement reactions involving metals*

Ions present in solution	Metal				
	Magnesium (Mg)	Zinc (Zn)	Iron (Fe)	Lead (Pb)	Copper (Cu)
Potassium (K^+)	×	×	×	×	×
Sodium (Na^+)	×	×	×	×	×
Calcium (Ca^{2+})	×	×	×	×	×
Magnesium (Mg^{2+})	×	×	×	×	×
Aluminium (Al^{3+})	√	×	×	×	×
Zinc (Zn^{2+})	√	×	×	×	×
Iron (Fe^{2+})	√	√	×	×	×
Lead (Pb^{2+})	√	√	√	×	×
Copper (Cu^{2+})	√	√	√	√	×
Silver (Ag^+)	√	√	√	√	√

Key

√ Displacement reaction takes place.

× Displacement reaction does not take place.

Activity 8.3



Aim: To investigate displacement reactions involving metals

Requirements: Beakers, measuring cylinders, spatula, glass rods, 1 M magnesium nitrate solution, 1 M zinc sulphate solution, 1 M lead(II) nitrate solution, 1 M copper(II) sulphate solution, 1 M silver nitrate solution, 2 M iron(III) chloride solution, copper turnings, zinc granules, and iron filings

Procedure

1. Transfer about 20 cm^3 of 1 M copper(II) sulphate solution in a beaker.
2. Add few zinc granules and stir using a glass rod, then record your observation.
3. Leave the solution overnight and record your observation.
4. Repeat Steps 1–3 using the following combinations, and record your observation:
   (a) Zinc granules in:
   (i) magnesium nitrate solution; (ii) lead(II) nitrate solution;
   (iii) silver nitrate solution; and (iv) iron(III) chloride solution.

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(b) Iron filings in:
(i) magnesium nitrate solution; (ii) zinc sulphate solution;
(iii) lead(II) nitrate solution; (iv) copper(II) sulphate solution; and
(v) silver nitrate solution.

(c) Copper turnings in:
(i) magnesium nitrate solution; (ii) zinc sulphate solution;
(iii) lead(II) nitrate solution; (iv) silver nitrate solution; and
(v) iron(III) chloride solution.

Questions

1. Tabulate your observations indicating the possible and impossible displacement reactions.
2. Write balanced chemical equations for the reactions occurring in Steps 4(b) (iii)−(v).

Carbon and hydrogen in the reactivity series

Carbon and hydrogen are included in the reactivity series although they are non-metals because they are reducing agents just like metals.

Carbon

Carbon is not a metal, but its characteristics lie between metals and non-metals. Based on its reducing power, carbon is placed between aluminium and zinc in the reactivity series. Metals above carbon in the reactivity series are usually extracted by electrolysis. Metals below carbon can be extracted by reducing their oxides using carbon or carbon monoxide. For example, charcoal (carbon) is heated with copper(II) oxide to produce brown specks of copper metal. White specks of lead metal are also formed when lead(II) oxide is heated under carbon. These reactions are indicated in the following equations:

$$2CuO(s) + C(s) \longrightarrow 2Cu(s) + CO_2(g)$$

$$2PbO(s) + C(s) \longrightarrow 2Pb(s) + CO_2(g)$$

Extraction methods used for different metals are summarised in Table 8.2.

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Table 8.2: *Extraction methods used for different metals*

Metal	Extraction method	Reactivity
Potassium	Electrolysis of molten compounds	Most reactive ↑
Sodium		
Calcium		
Magnesium		
Aluminium		
(Carbon)		
Zinc	Heating with carbon	
Iron		
Copper		
Silver	Various chemical reactions	
Gold		Least reactive

Activity 8.4

Aim: To investigate the reactions between some metal oxides and carbon

Requirements: Combustion tube, spatula, retort stand and clamp, Bunsen burner, watch glass, limewater, copper(II) oxide, lead(II) oxide, and crushed charcoal

Procedure

1. Put 3 spatulafuls of copper(II) oxide powder and 3 spatulafuls of charcoal powder in a watch glass. Mix the two substances thoroughly until they form a homogeneous mixture.
2. Put the mixture in the combustion tube and set the apparatus as in Figure 8.5.

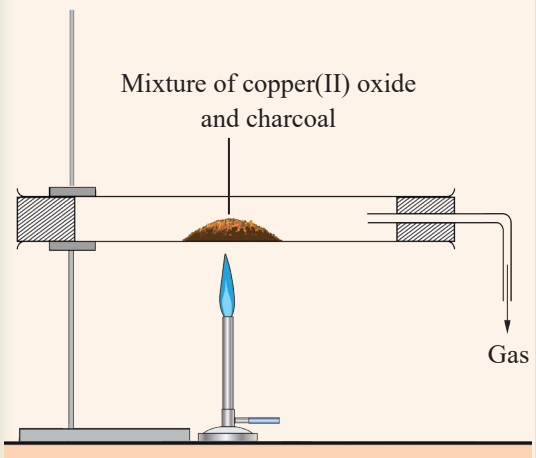


Figure 8.5: *Set-up for the reaction between copper(II) oxide and charcoal*

3. Heat the mixture strongly by using a non-luminous flame. Use limewater to test the gas evolved.
4. Note the colour of the residue when there is no more evolution of gas.

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5. Repeat Steps 1–4 by using lead(II) oxide instead of copper(II) oxide.

Questions

1. Name the gas evolved in each case and write a chemical equation for its identification.
2. State the colour of the residues found after the heating process.

Hydrogen

Hydrogen is placed between lead and copper in the reactivity series. Metals below hydrogen cannot react with acids to form hydrogen gas because they cannot displace hydrogen from a non-metal anion. Such metals are also not attacked by water or steam, and are therefore used in jewellery and in making coins. These metals include copper, silver, gold and platinum.

Zinc and magnesium react with hydrochloric acid to give salt solutions and hydrogen gas as indicated in the following equations:

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

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

The reaction between magnesium and hydrochloric acid is more vigorous than that of zinc and the acid. This is because magnesium is higher in the reactivity series than zinc. The reactions between hydrochloric acid and metals such as magnesium and zinc to give a salt solution and hydrogen gas are redox reactions. The hydrogen in the hydrochloric acid is an oxidising agent, and it is reduced to hydrogen gas. Table 8.3 shows the order of reactivity of metals with dilute acids.

Table 8.3: *Reactivity of metals with dilute acids*

Metal	Reaction
Potassium, sodium, and lithium	Very reactive, and hence too dangerous to add to dilute acids
Calcium, magnesium aluminium, zinc, iron, and lead	React moderately to produce a salt and hydrogen gas.
Copper, silver, and gold	Less reactive than hydrogen, and hence do not react with dilute acids

Note: The reactions between aluminium and dilute acids are slower than would be expected considering its position in the reactivity series. This is because of the formation of an oxide layer which prevents further reactions of the acids with the metal.

Activity 8.5



Aim: To investigate the reactions between metals and dilute hydrochloric acid

Requirements: Beaker, measuring cylinder, spatula, glass rod, test tube, 1 M hydrochloric acid, zinc granules, magnesium ribbon, and copper turnings

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Procedure

1. Put some zinc granules in a test tube.
2. Add about 10 cm^3 of 1 M hydrochloric acid into the test tube containing zinc granules. Record your observations.
3. Touch the test tube with the back of your hand and note its temperature.
4. Repeat Steps 1−3 using magnesium ribbon and then copper turnings in the place of zinc granules. Record your observations.

Questions

1. With the aid of balanced chemical equations, explain the observations in Step 2 for each metal.
2. State if the reactions are endothermic or exothermic.
3. Comment on the fact that, the reducing property of a metal is related to its reactivity.
4. Giving examples, differentiate between minerals and ores.

Exercise 8.1

1. A metal has both physical and chemical strengths. What does this statement imply?
2. With examples, explain any four chemical properties of metals.
3. When a piece of magnesium ribbon is added into a solution of dilute hydrochloric acid, some gas bubbles are observed, but when a piece of copper wire is used instead of magnesium, no bubbles are observed. Explain.

Extraction of metals

Metals are extracted from their ores. An ore is a naturally occurring solid material from which a metal or valuable mineral is extracted profitably. When an ore containing a metal is dug out from the source, the metal is separated from the rest of the materials in the ore by a process called *extraction*. The branch of science that deals with the properties, production and purification of metals is called *metallurgy*.

Processes for extraction of metals

The main processes involved in the extraction of metals from their ores are:

(a) crushing and grinding.

(b) concentration or dressing of the ore.

(c) calcination or roasting of the ore.

(d) reduction of a metal oxide to a free metal.

(e) purification and refining of the metal.

Crushing and grinding

Metal ores are generally obtained as large pieces of rocks. The metal ore has to be crushed and ground (pulverised) to convert the components into powder. This is done by using crushers and grinders.

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Concentration or dressing of the ore

Ores are usually found mixed with earthy impurities like sand, clay, and limestones. The impurities in the ore are called *gangue* or *matrix*. The process of removing impurities (gangue) from the metal ore is called *concentration*. This helps to obtain the materials richer in the metal or those worth working up with. There are various ways to carry out the concentration of the ores. The suitable method is chosen based on the physical and chemical properties of the metal. Some important methods used in concentrating the metal ores include *gravity separation* (hydraulic washing), *magnetic separation*, *froth flotation*, and *chemical methods*.

Gravity separation

This method is based on the differences in weight between the impurities and the metal. Impurities are lighter than the metal particles, the principle of gravity is used to separate them. In this method, the light earthy impurities are removed from the heavier metallic ore particles by washing with water.

Magnetic separation

In this method, the magnetic properties of certain metals are used to separate them from the non-magnetic impurities. The reverse may be the case where the impurities are magnetic, while metals are non-magnetic. In either case, magnetic separation is applicable.

Froth flotation

This is the process for selectively separating hydrophobic materials from hydrophilic materials. It uses the difference in the wetting characteristics between minerals and impurities. Froth flotation is primarily used for concentrating sulphide ores. Sulphide ores are hydrophobic in nature and are easily wetted by oil. Thus, sulphide ores can be selectively separated from hydrophilic impurities by froth flotation.

Leaching

Leaching is the chemical method of concentrating the ore. It involves dissolving the ore in a chemical solution in which the mineral dissolves leaving behind the impurities which are filtered out. The metal ore is then recovered from the chemical solution by a suitable chemical. For example, the concentration of the bauxite ore for extraction of aluminium. Bauxite is contaminated with iron, titanium, and silicon oxide. These impurities are removed by heating the powdered ore with an aqueous solution of sodium carbonate or sodium hydroxide at a certain temperature and pressure. Aluminium oxide dissolves in the base, whereas the impurities remain insoluble and are removed by filtration.

Calcination and roasting of the ore

The concentrated metal ore is converted into metal oxide by either calcination or roasting.

Calcination

Calcination is a process in which the ore is heated below its melting point in the absence of air. This method is commonly used for converting metal carbonates into

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their respective oxides. Calcination is generally carried out in a *reverberatory furnace* or in multiple heater furnace. During calcination, moisture, volatile gaseous substances, and water of hydration are driven out. Examples of calcination include the following:

(a) Expulsion of CO_2 from carbonate

$$ZnCO_3(s) \xrightarrow{\text{heat}} ZnO(s) + CO_2(g)$$

(b) Removal of water of hydration

$$Al_2O_3{\cdot}2H_2O(s) \xrightarrow{\text{heat}} Al_2O_3(s) + 2H_2O(g)$$

Roasting

Roasting involves converting the concentrated ore into a metal oxide by heating it below its melting point in the presence of excess air. While calcination is mostly used for converting metal carbonates to oxides, roasting is a method used for converting sulphide ores to their respective oxides as shown in the following examples:

$$2PbS(s) + 3O_2(g) \xrightarrow{\text{heat}} 2PbO(s) + 2SO_2(g)$$

$$2ZnS(s) + 3O_2(g) \xrightarrow{\text{heat}} 2ZnO(s) + 2SO_2(g)$$

As in calcination, during roasting, moisture and impurities in the form of gases such as sulphur dioxide are removed. A metal is obtained from its oxide by the processes of reduction or electrolysis. The oxides are easier to deal with in the reduction process than the sulphides or carbonates.

Reduction of the metal oxides to free metals

There are various ways by which the reduction of a metal oxide can be done. The methods used include the reduction by heating in the air (*auto-reduction*), reduction by carbon (*smelting*), reduction by *electrolysis* (electrolytic reduction), reduction by *precipitation,* and *amalgamation*.

The choice of the reduction method to be used for a given metal compound depends on the position of the metal in the reactivity series. The more reactive the metal, the more difficult it is for it to be extracted from its compounds. Metals high in the reactivity series require the strongest method of reduction from their compounds, which is *electrolysis*. The method is called the *electrolytic reduction*. Metals under this group include potassium, sodium, calcium, magnesium, and aluminium. Table 8.4 shows the methods used to reduce different metal compounds.

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Table 8.4: *Methods used to reduce (extract) different metal compounds (metal ores)*

Method	Metal compound
Reduction by using carbon (smelting)	Oxides of less electropositive metals such as lead, zinc, iron, tin, and copper
Reduction by heating in the air (auto-reduction)	Oxides and sulphides of less electropositive metals such as mercury, copper, and lead
Reduction by electrolysis	Oxides, hydroxides or chlorides of metals high in the reactivity series such as sodium, calcium, magnesium, and aluminium
Reduction by precipitation	Metal ores of less reactive metals. A solution containing metal ions is reacted with a more reactive metal. The more reactive metal displaces the less reactive metal from the solution.
Amalgamation	Less reactive metals like silver and gold. The finely crushed ore is brought into contact with mercury which forms an alloy (amalgam) with the metal. The metal is then recovered by distilling the amalgam.

Refining and purification

Metals obtained through reduction process are generally impure. They are known as *crude metals*. Crude metals contain impurities such as other metals, non-metals like silicon and phosphorus, and unreduced oxides and sulphides of the metals. Some impurities are also introduced in the process when preparing the metal ore for reduction. This is especially the case when the chemical method is used to concentrate the ore. Purification of the crude metals is done through three main methods which are *distillation*, *oxidation,* and *electro-refining*.

Distillation

Distillation involves heating the crude metal in a furnace until a pure metal evaporates, leaving behind the impurities. The vapour is then collected and condensed in a separate chamber.

Oxidation

The molten crude metal is exposed to air in the furnace so that the impurities are oxidised and escape as vapour or form a scum over the molten metal. The scum is then removed by skimming. This method is used only when the impurities have a greater affinity for oxygen than the metal.

Electro-refining

Most metals are purified by electrolysis (electro-refining). The crude metal is moulded into blocks and made the anode of an electrolytic cell. The cathode is usually made up of a thin plate of the

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pure metal. When a current is passed, the anode dissolves and the pure metal is deposited on the cathode. The soluble impurities go into the solution, while the insoluble impurities settle down below the anode as *mud* or *sludge*.

Extraction of metals by electrolytic reduction

The extraction of metals by electrolytic reduction involves the application of electric current as a source of electrons for the reduction process. The electrolytic reduction is a type of electrolysis. The oxides, hydroxides and chlorides of metals in a fused state are electrically reduced by using this method. The metals are extracted at the cathode. Some metals like sodium, potassium, magnesium, calcium, and aluminium are obtained by this process. This is a common extraction process for the more reactive metals.

Extraction of sodium

Occurrence

Sodium does not occur naturally as a free element because it is extremely reactive. However, it is abundant in different compounds such as sodium chloride (rock salt) and sodium sulphate.

Extraction

Sodium is extracted by the *electrolytic reduction* of purified molten rock salt. The electrolysis is mainly carried out in the *Down's cell* shown in Figure 8.6. The cell has a steel cathode and a graphite anode. Calcium chloride is added to sodium chloride to lower the melting point of sodium chloride from 774 °C to about 600 °C. The composition of the electrolyte is 40% sodium chloride and 60% calcium chloride. During the electrolysis, chloride ions move to the anode, while sodium ions move to the cathode.

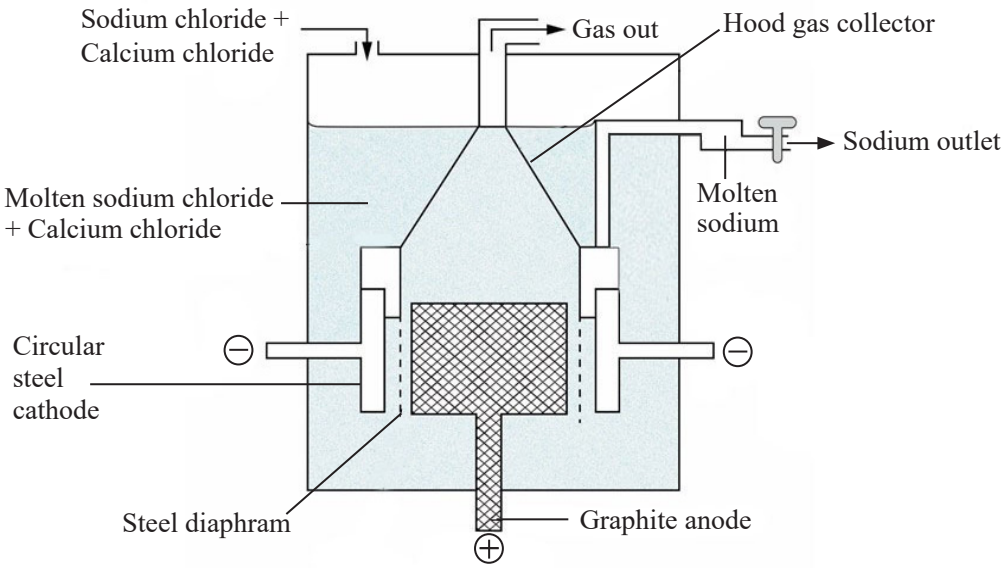


Figure 8.6: *Sketched Down's cell*

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The reactions taking place at the electrodes are:

(a) At the anode, chloride ions are discharged as chlorine gas as shown by the following equation:

$$2Cl^-(l) \longrightarrow Cl_2(g) + 2e^-$$

(b) At the cathode, sodium ions are discharged as a metal as shown by the following equation:

$$2Na^+(l) + 2e^- \longrightarrow 2Na(l)$$

If the sodium metal and chlorine gas produced are allowed to come into contact at 600 °C, a violent reaction occurs which oxidises the metal. To prevent this, the cell has a steel gauze as a diaphragm around the anode to keep the two products apart.

The large graphite anode is used to facilitate maximum oxidation of chloride ions to chlorine gas. This also maximises the formation of sodium metal at the cathode. Chlorine is collected in the inverted trough (hood) placed over the anode, and is tapped off through an iron vessel. The sodium metal is collected upwards in the *Down's cell* because of its low density which makes it float over the mixture. The sodium metal from the Down's cell contains some calcium, which is also formed through electrolysis. The calcium crystallises when the mixture cools and a relatively pure sodium metal is obtained.

Extraction of aluminium

Aluminium is the most abundant metal on the Earth's crust making about 8% of the crust. It does not occur naturally as a free element. It occurs in the forms of oxides, fluorides, silicates, and sulphates. The main ores of aluminium are bauxite ($Al_2O_3 \cdot 2H_2O$), cryolite (Na_3AlF_6), feldspar ($KAlSi_3O_8$), and kaolin ($Al_2Si_2O_7 \cdot 2H_2O$).

The chief ores from which aluminium is extracted are bauxite and cryolite. Aluminium is extracted from bauxite by electrolysis. This is done by using the *Hall-Héroult method* that involves the following stages:

(a) Purification of bauxite.

(b) Electrolytic reduction of alumina.

(c) Refining of the crude aluminium.

Purification of bauxite

Bauxite contains oxides of iron and silicon as impurities. These impurities are removed by using the *Hall's process* where the powdered ore is heated to bright red with sodium carbonate. Since aluminium oxide is amphoteric, it dissolves to form sodium aluminate ($NaAlO_2$). This reaction is indicated in the following equation:

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$$Al_2O_3{\cdot}2H_2O(s) + Na_2CO_3(s) \xrightarrow{\text{heat}} 2NaAlO_2(l) + 2H_2O(l) + CO_2(g)$$

The molten mass is then removed together with water, while the insoluble iron and silicon oxides are left behind as residue. When the filtrate is heated at a temperature between 50 °C and 60 °C and a stream of carbon dioxide is passed through the mixture, precipitation of aluminium hydroxide occurs.

$$2NaAlO_2(l) + 3H_2O(l) + CO_2(g) \xrightarrow{\text{heat}} 2Al(OH)_3(s) + Na_2CO_3(aq)$$

Aluminium hydroxide is filtered off, washed, dried, then calcinated at 1500 °C to give alumina.

$$2Al(OH)_3(s) \xrightarrow{\text{heat}} Al_2O_3(s) + 3H_2O(g)$$

Electrolysis of alumina

Alumina (Al_2O_3) cannot be reduced to aluminium metal by carbon because aluminium has a greater affinity to oxygen than carbon. Carbon can react with aluminium at high temperature to form aluminium carbide (Al_4C_3).

The electrolysis of alumina is carried out in the presence of cryolite (Na_3AlF_6) which improves the electrical conductivity of alumina, which is otherwise a bad conductor of electricity. Cryolite also lowers the melting point of aluminium oxide. This is very important because if electrolysis is carried out at very high temperatures, the aluminium formed will vapourise.

The sketch of an electrolytic cell for the electrolysis of alumina (aluminium oxide) is shown in Figure 8.7.

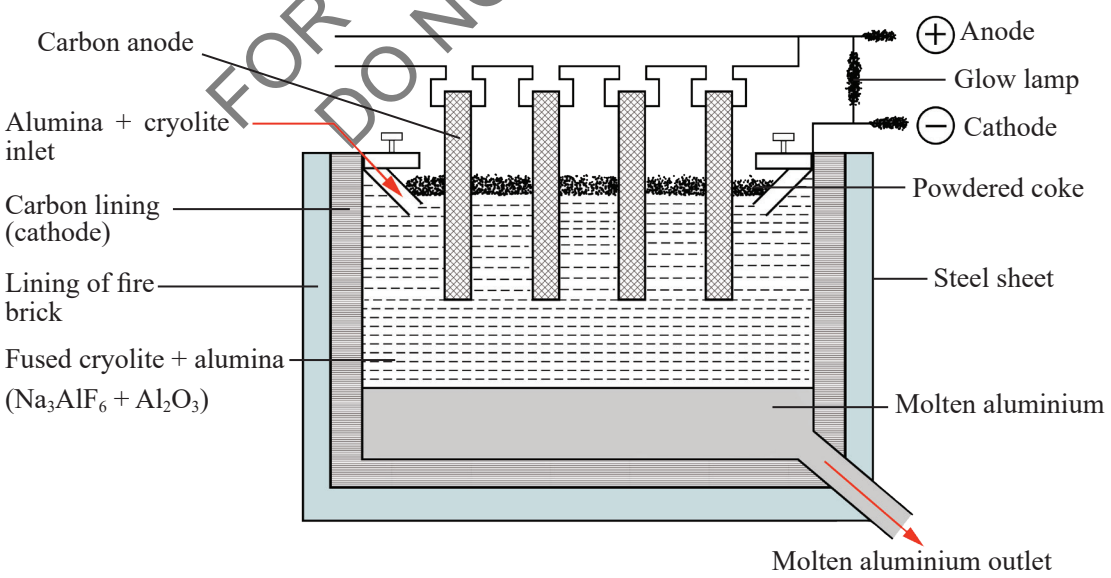


Figure 8.7: *Sketch of the electrolytic cell for electrolysis of alumina*

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The electrolytic cell for the electrolysis of alumina consists of an iron bath lined with graphite (carbon) which acts as the cathode. The anode consists of a number of carbon rods attached to copper clamps and suspended in the fused mass. The carbon rods are arranged in such a way that they can be raised, lowered or replaced when needed. The reactions taking place at the electrodes are as shown in the following equations:

Reaction at the cathode:

$$4Al^{3+}(l) + 12e^- \longrightarrow 4Al(l)$$

Reaction at the anode:

$$6O^{2-}(l) \longrightarrow 6O_2(g) + 12e^-$$

After it is discharged, the aluminium is collected at the bottom of the cell in molten state. The overall equation for the electrolysis of aluminium is:

$$2Al_2O_3(l) \longrightarrow 4Al(l) + 3O_2(g)$$

The production of aluminium by electrolysis is an expensive process. The temperature of the fused electrolyte has to be maintained between 800 °C and 900 °C. This is achieved by passing a steady current of about 100,000 ampere through the electrolyte. This makes aluminium expensive despite the fact that it is the most abundant metal on Earth. The process is only economical when electricity is cheap and readily available.

Refining of crude aluminium

The aluminium obtained from the Hall-Héroult process is 99% pure. It contains small quantities of iron, silicon, alumina and carbon. It is purified further by electrolysis using *Hoope's electrolytic cell* (Figure 8.8).

The Hoope's electrolytic cell is made up of an iron tank lined on the inside with carbon which serves as the anode. The tank has three different layers of molten substances:

(a) The bottom layer contains impure molten aluminium containing copper and silicon to increase the density.

(b) The middle layer contains a mixture of fluorides of sodium, barium, and aluminium in a fused form. This layer acts as an electrolyte.

(c) The top layer contains pure molten aluminium which together with the carbon rod serves as the cathode. The carbon cathodes are suspended from above inside the cell mixture.

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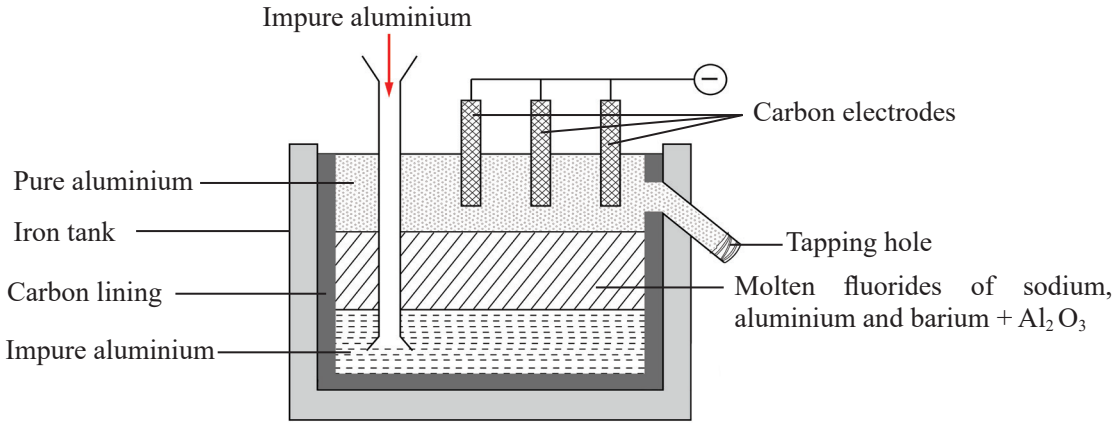


Figure 8.8: *Sketch of the Hoope's electrolytic cell*

When the electric current is passed through the cell, the aluminium formed in the middle layer passes to the top layer as pure aluminium. An equivalent amount of impure aluminium passes from the bottom layer to the middle layer and the process continues. Aluminium of about 100% purity is tapped from the tapping hole.

Exercise 8.2

1. Briefly explain how sodium metal is extracted from its ore.
2. Explain what would happen in the Down's cell if the molten sodium is mixed with chlorine at very high temperature during the electrolytic extraction of sodium.
3. Describe the processes involved in the extraction of aluminium.
4. Comment on the extraction of aluminium from bauxite in the absence of cryolite.

Extraction of metals by chemical reduction

Less electropositive metals such as iron and copper can best be extracted through chemical reduction which involves heating the ore with reducing agents. The common material used in reducing the ores is carbon.

Extraction of iron

Iron is the second most abundant metal occurring in the Earth's crust after aluminium. It does not occur as a free element because it is relatively reactive. The ores of iron include haematite (Fe_2O_3), magnetite (Fe_3O_4), limonite ($Fe_2O_3 \cdot H_2O$), siderite or spathic iron ore ($FeCO_3$), and iron pyrites (FeS_2). The most prominent ores of iron are haematite which contains about 70% iron, and magnetite which contains 72.4% iron. Although iron pyrites is abundant on the Earth's crust, it is not used as a source of iron, but it is mainly used in the production of sulphuric acid.

Iron is mainly extracted from the haematite ore by reduction using carbon (coke). Three main stages are involved in the extraction process, which are *concentration*, *roasting,* and *smelting* (*reduction*).

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Concentration

The haematite ore (Fe_2O_3) is first washed to remove earthy impurities such as clay, sand, and other non-magnetic impurities.

Roasting

The concentrated ore is strongly heated in excess air in shallow kilns and the following chemical changes occur during the process. In addition, calcination of hydrated iron(III) oxide and iron(II) carbonate occurs.

(a) Most of the moisture is removed.

$$Fe_2O_3 \cdot xH_2O(s) \xrightarrow{\text{heat}} Fe_2O_3(s) + xH_2O(l)$$

(b) Sulphur and arsenic impurities are oxidised to their oxides, which escape as gases.

$$S(s) + O_2(g) \xrightarrow{\text{heat}} SO_2(g)$$

$$4As(s) + 3O_2(g) \xrightarrow{\text{heat}} 2As_2O_3(g)$$

(c) Carbonates are decomposed into oxides giving off carbon dioxide gas. For example,

$$FeCO_3(s) \xrightarrow{\text{heat}} FeO(s) + CO_2(g)$$

(d) Iron(II) oxide is oxidised to iron(III) oxide

$$4FeO(s) + O_2(g) \xrightarrow{\text{heat}} 2Fe_2O_3(s)$$

This reduces the loss of iron, since iron(II) oxide has a tendency of forming a slag with silicon dioxide (sand).

$$FeO(s) + SiO_2(s) \xrightarrow{\text{heat}} FeSiO_3(l)$$

The entire mass becomes porous and suitable for reduction to metallic iron in the blast furnace.

Smelting

The roasted ore is reduced (smelted) by carbon in the *blast furnace*. The blast furnace is a metallurgical furnace used for smelting to produce metals such as iron, lead, and copper. The structure of the blast furnace for iron smelting is about 30 m high and 10 m in diameter at its widest part. Figure 8.9 shows the parts of the blast furnace for smelting iron. The inside lining of the furnace is made up of firebrick which is resistant to high temperature. The ore is mixed with carbon (coke) and limestone ($CaCO_3$) which are introduced into the furnace from the top. Blasts of hot air are blown into the furnace through small openings near the base known as *tuyéres*.

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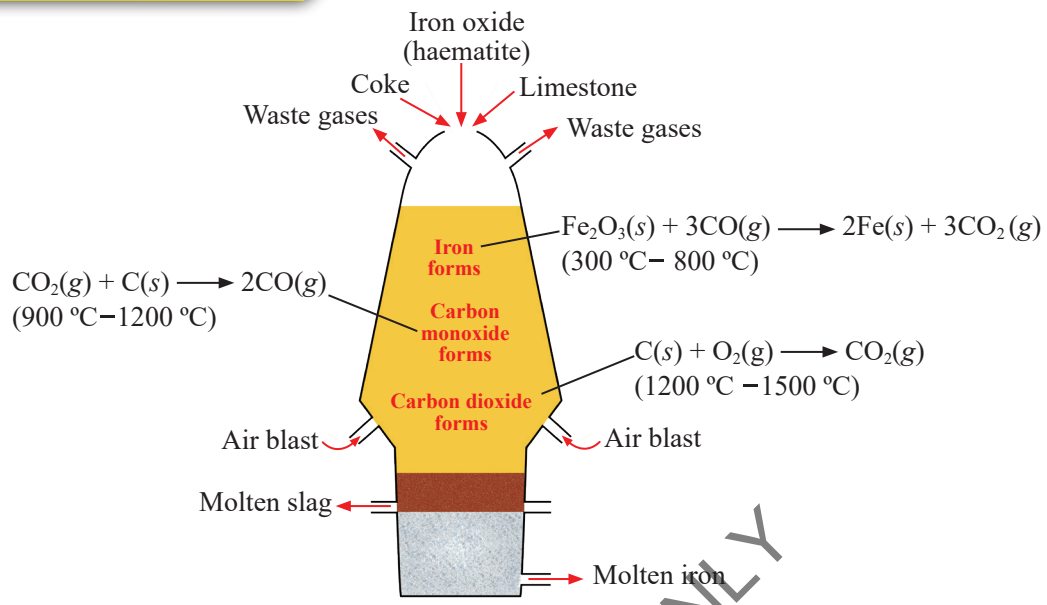


Figure 8.9: *Schematic diagram of the blast furnace for iron smelting*

The blast furnace is divided into three zones, according to the temperature profile in the furnace.

(a) The upper zone of reduction

This zone is near the top of the furnace with a temperature range of 300 °C–800 °C.

The reactions which take place in this zone depend directly on the reactions that occur in the lower and middle zones. In the lower zone, the carbon burns in oxygen of the hot air to form carbon dioxide.

$$C(s) + O_2(g) \xrightarrow{\text{heat}} CO_2(g)$$

The carbon dioxide formed reacts with more coke to form carbon monoxide at the middle zone.

$$C(s) + CO_2(g) \xrightarrow{\text{heat}} 2CO(g)$$

The carbon monoxide rises up to the upper zone of the furnace for reducing iron(III) oxide to spongy iron.

$$Fe_2O_3(s) + 3CO(g) \xrightarrow{\text{heat}} 2Fe(l) + 3CO_2(g)$$

It should be noted that, the main reducing agent in the blast furnace is carbon monoxide and not carbon. The metal obtained is called *spongy iron* because the temperature in this zone is too low to melt the iron. As the iron ore is being reduced, part of the limestone decomposes at 600 °C to quicklime (CaO).

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$$CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$$

The limestone also reacts with coke to form quicklime.

$$CaCO_3(s) + C(s) \xrightarrow{\text{heat}} CaO(s) + 2CO(g)$$

(b) The middle zone

This is the lower zone of reduction with the temperature range of 900 °C−1200 °C. The following are the reactions that take place in this zone:

(i) Carbon dioxide is reduced by coke to carbon monoxide.

$$C(s) + CO_2(g) \xrightarrow{\text{heat}} 2CO(g)$$

This reaction is accompanied by heat absorption.

(ii) At about 1000 °C, calcium carbonate is almost completely decomposed to quicklime (CaO)

$$CaCO_3(s) \xrightarrow{1000\ °C} CaO(s) + CO_2(g)$$

(iii) The calcium oxide acts as a flux and combines with the silica (silicon dioxide) to form calcium silicate (slag). A *flux* is a mixture of chemicals that react with impurities to form slag.

$$CaO(s) + SiO_2(s) \xrightarrow{1000\ °C} CaSiO_3(l)$$

(c) The lower zone of combustion

This is the lowest and the hottest part of the furnace. The temperature range at this zone is 1200 °C−1500 °C. The following reactions take place in this zone:

(i) Coke burns in the blast of hot air to produce carbon dioxide. This reaction is exothermic, therefore, it raises the temperature of the furnace.

$$C(s) + O_2(g) \xrightarrow{\text{heat}} CO_2(g)$$

(ii) The spongy iron melts in this zone at about 1300 °C and collects at the bottom of the furnace

(iii) Iron(II) oxide which might have escaped the reduction process in the middle zone is reduced by coke in this zone.

$$2FeO(s) + C(s) \xrightarrow{\text{heat}} 2Fe(l) + CO_2(g)$$

The molten iron sinks at the bottom of the furnace, while the less dense fusible slag floats over the molten iron forming a separate layer. This layer of slag prevents iron from being oxidised by hot air. The slag and iron are periodically removed through different outlets.

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The mixture of waste gases containing nitrogen, carbon dioxide, and carbon monoxide are known as *blast furnace gases*. This mixture of gases is burnt in air to produce heat which is used for pre-heating the air blast coming in through the tuyéres.

The iron obtained through the blast furnace is an impure type known as *pig iron*. Pig iron is further purified by re-smelting it using coke and lime in another furnace called *cupola*. The molten iron from the cupola is poured into moulds of desired shapes. The pure iron obtained is called *cast iron*.

Extraction of copper

Copper occurs in both free and combined states. Most copper occurs in combined forms, mainly as sulphides and oxides. The main ores of copper are copper pyrites ($CuFeS_2$), copper glance (Cu_2S), cuprite (Cu_2O), malachite [$Cu_2CO_3(OH)_2$] or [$CuCO_3 \cdot Cu(OH)_2$] and azurite [$Cu_3(CO_3))_2(OH)_2$] or [$2CuCO_3 \cdot Cu(OH)_2$]. Copper is mainly extracted from copper pyrites by using the dry process. There are five main stages that are involved in the extraction process, namely; *concentration, roasting, smelting, conversion,* and *refining.*

Concentration

The powdered ore is concentrated by froth flotation or other processes to remove non-ore impurities such as galena (PbS).

Roasting

The concentrated ore is heated strongly in the presence of excess air in a special furnace where it decomposes to form copper(I) and iron(II) sulphides.

$$2CuFeS_2(s) + O_2(g) \xrightarrow{\text{heat}} Cu_2S(s) + 2FeS(s) + SO_2(g)$$

Part of the sulphides may be oxidised to their respective oxides. The reactions are shown in the following equations:

$$2Cu_2S(s) + 3O_2(g) \xrightarrow{\text{heat}} 2Cu_2O(s) + 2SO_2(g)$$

$$2FeS(s) + 3O_2(g) \xrightarrow{\text{heat}} 2FeO(s) + 2SO_2(g)$$

As the roasting proceeds, any moisture present in the ore evaporates. Impurities such as sulphur, arsenic, and antimony are removed in the form of their volatile oxides. For example, sulphur is oxidised as shown in the following equation:

$$S(s) + O_2(g) \xrightarrow{\text{heat}} SO_2(g)$$

The roasting process results in a mixture of sulphides and oxides of copper and iron.

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Smelting

The roasted ore is mixed with coke and sand in the presence of excess air in a water jacketed blast furnace shown in Figure 8.10.

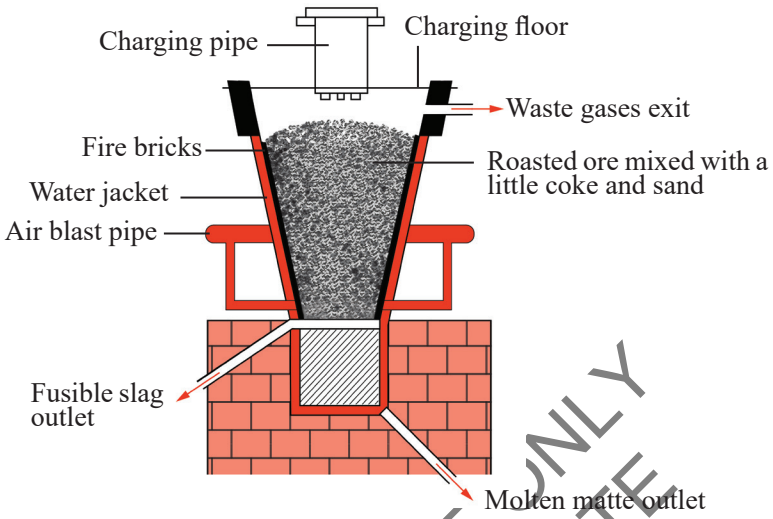


Figure 8.10: *Sketched blast furnance for the extraction of copper*

In the blast furnace, the oxidation of iron(II) sulphide which started during the roasting process continues as shown in the following equation:

$$2FeS(s) + 3O_2(g) \xrightarrow{\text{heat}} 2FeO(s) + 2SO_2(g)$$

The iron(II) oxide formed combines with sand to form a fusible slag.

$$FeO(s) + SiO_2(s) \xrightarrow{\text{heat}} \underset{\text{Slag}}{FeSiO_3(l)}$$

Copper(I) oxide also reacts with some of the iron(II) sulphide to form iron(II) oxide.

$$Cu_2O(s) + FeS(s) \xrightarrow{\text{heat}} Cu_2S(s) + FeO(s)$$

The reaction between copper(I) oxide and iron(II) sulphide takes place because iron has a greater affinity for oxygen than sulphur. The reaction results in the conversion of most of the iron sulphide into the oxide, whereby iron(II) oxide is removed as slag. The molten mass containing mainly copper(I) sulphide with little amount of iron(II) sulphide is known as *matte*. It is taken out through the outlet at the bottom.

Conversion

The molten matte is then transferred to a Bessemer converter shown in Figure 8.11. This is a pear-shaped furnace made of steel. The furnace is fitted with pipes known as *tuyéres* on the sides through which sand and hot air are blown.

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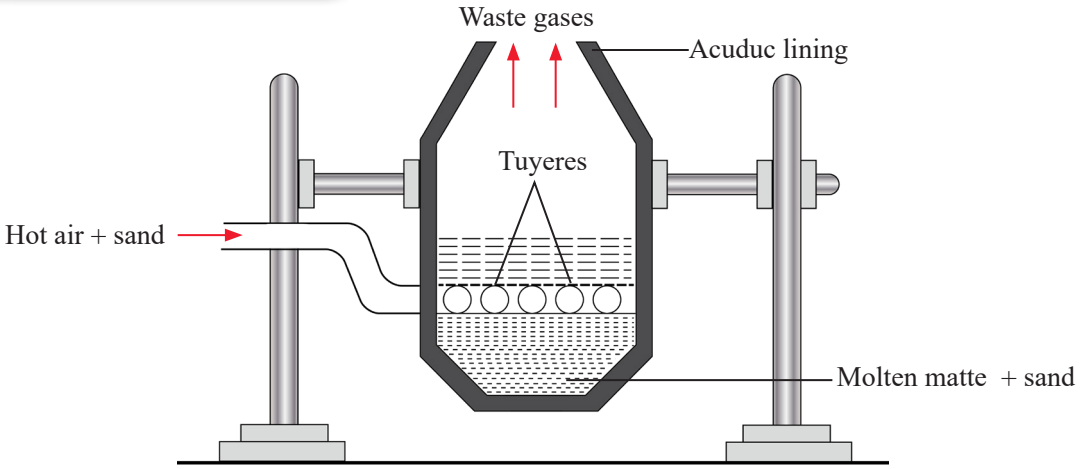


Figure 8.11: *Sketch of a Bessemer converter*

The following reactions take place in the converter:

(a) Iron(II) sulphide reacts with oxygen to form iron(II) oxide.

$$2FeS(l) + 3O_2(g) \xrightarrow{\text{heat}} 2FeO(l) + 2SO_2(g)$$

(b) The iron(II) oxide formed reacts with sand (silicon dioxide) to form slag.

$$FeO(l) + SiO_2(s) \xrightarrow{\text{heat}} FeSiO_3(l)$$

Slag

(c) Some of the copper(I) sulphide is oxidised to copper(I) oxide.

$$2Cu_2S(l) + 3O_2(g) \xrightarrow{\text{heat}} 2Cu_2O(l) + 2SO_2(g)$$

(d) Copper(I) oxide is reduced by copper(I) sulphide to copper.

$$2Cu_2O(l) + Cu_2S(l) \xrightarrow{\text{heat}} 6Cu(l) + SO_2(g)$$

The reaction taking place in (d) is known as *self-reduction of copper* because the substances involved are both copper compounds.

The molten metal is poured into sand moulds and allowed to stand. On cooling, any dissolved sulphur dioxide escapes leaving blisters on the surface of copper. The copper obtained from the Bessemer converter is therefore known as *blister copper*.

Refining of copper

Blister copper is 97% to 99% pure. This means that it contains a small amount of impurities which affect its quality. Trace amounts of sulphur in copper significantly reduce its electrical conductivity. Blister copper is refined further by electrolysis in which copper(II) sulphate solution is an electrolyte. The blister copper is cast into blocks and made the anode. The cathode is made up of a sheet of pure copper. Figure 8.12 shows a sketch of an electrolytic cell that is used in purifying the blister copper.

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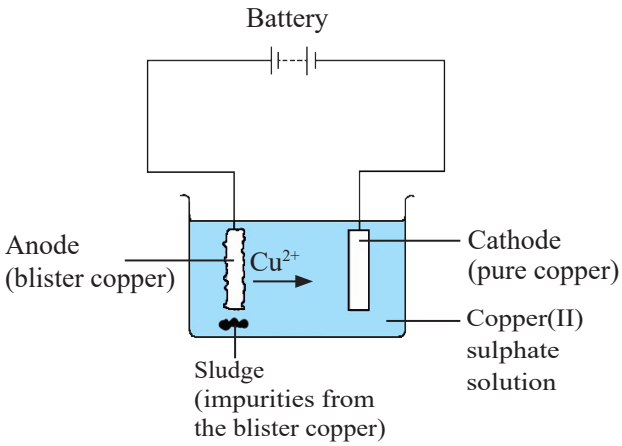


Figure 8.12: *Sketch of an electrolytic cell for refining blister copper*

At the anode, blister copper is oxidised to copper(II) ions.

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^{-}$$

The ions move to the cathode where they are reduced to copper metal.

$$Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)$$

The copper is deposited on the cathode sheet. As the electrolysis continues, the whole blister copper block dissolves. This eventually increases the size of the pure copper plate.

Exercise 8.3

1. Describe any four chief ores of each of the following metals:
   (a) Iron
   (b) Copper
2. Differentiate the extraction of copper from that of iron.
3. What is the role of limestone in the extraction of iron?

Environmental considerations

The extraction of metals is a violent process which may have great impacts on the environment. Such impacts may cause serious environmental problems as discussed in the next subsection.

Environmental degradation caused by extraction of metals

The extraction of metals causes environmental degradation due to production of solid waste, chemical waste, energy, gaseous pollutants, such as greenhouse gases, and dust.

Solid wastes

A small amount of ore obtained from extraction process produces a huge mountain of solid wastes. Mining wastes are not inert as they often contain toxic chemicals such as cyanide that are added to the ore during extraction of metals.

Chemical wastes

Many of the chemicals that are used in the separation of metals are toxic. Froth flotation uses a range of chemicals, and leaching utilises acids as well as toxic metals such as mercury. If such chemicals contaminate the environment (water, air, soil, and organisms), they may cause health problems such as headache, nausea, genetic defects, and respiratory disorders in human beings when they get exposed to them.

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Energy production

Some processes such as smelting produce a large amount of mechanical, heat, and sound energy. These interfere with the environment in some ways. The mechanical energy interferes with the earth's structure, while the noise pollution causes health problems such as hearing loss and cardiovascular diseases.

Gaseous pollutants

Smelting releases gases such as carbon dioxide, nitrogen dioxide and sulphur dioxide. These gases contribute to the formation of acid rain. The increase of carbon dioxide in the atmosphere contributes to global warming.

Dust

The crushing and separation processes of ores produce dust. The dust pollutes the environment and may cause health problems such as respiratory problems, and irritation of eyes, nose, throat, and skin to human beings.

Intervention measures to environmental degradation

Extraction of metals is inevitable for economic reasons. However, it is important to consider environmental health for better life. The following are some intervention measures to environmental degradation due to metal extraction:

Laws and regulations: To minimise the environmental effects due to metal extraction, the laws and regulations that require the extraction companies to restore the land to its previous state should be implemented. This involves flattening the heaps of wastes, rehabilitation of landfill sites and landscapes, eliminating ponds, and afforestation after extraction.

Recycling: Minerals do not exist forever, thus once exploited they are gone. Recycling of metals such as iron reduces the need for further extraction. Recycling saves the valuable finite natural mineral resources and reduces the problems of waste disposal.

Controlling chemicals before release: Toxic gases such as carbon monoxide which can be used as fuel must be controlled before being released into the atmosphere. Harmful acidic gases such as sulphur dioxide and nitrogen dioxide can also be treated with alkali solutions. For example, limewater can be used to neutralise the sulphur dioxide gas to harmless calcium sulphate.

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Exercise 8.4

1. Briefly explain the impacts of extraction of metals to the environment.
2. Describe the major ways which can be used to control the environmental impacts that are caused by extraction of metals through mining.
3. Suggest the best solution for solid waste management in the mining industry.

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Chapter summary

1. Metals occur in nature either as free elements or compounds such as oxides, sulphides, or chlorides.
2. An ore is a naturally occurring solid material from which metals or valuable minerals can be extracted profitably.
3. Metallurgy is the science that is concerned with the properties of metals, their production and purification.
4. The most common methods that are used to extract metals from their ores are electrolytic reduction and reduction by using carbon (smelting).
5. The method that is used to extract a metal from its ore depends on the position of the metal in the reactivity series.
6. Very reactive metals such as sodium, potassium, and calcium are extracted by electrolysis. Less reactive metals such as zinc, iron, lead and copper are extracted by reduction using carbon.
7. The arrangement of metals according to their reducing powers is known as the reactivity series.
8. A substance that loses electrons is a reducing agent, while the one that gains electrons is an oxidising agent. Metals are reducing agents because they readily lose electrons in a chemical reaction.
9. Carbon and hydrogen are not metals, but they are also reducing agents.
10. In the Periodic Table, the reducing strengths of metals increase down the group, but decrease across the period.

Revision exercise 8

Chose the correct answer in each of the items 1–4.

1. Which of the following methods is most suitable for extracting iron from its ore?
   (a) Reduction using carbon.
   (b) Electrolysis of its solution.
   (c) Roasting of the iron ore in air.
   (d) Electrolysis of its molten forms.
2. Which of the following statements is not true about the blast furnace during the extraction of iron?
   (a) The reactions in the blast furnace keep the iron molten.
   (b) Limestone is reacted with the impurities to form slag.
   (c) Coke is added to reduce iron oxide to iron metal.
   (d) Coke combines with some impurities in the ore.

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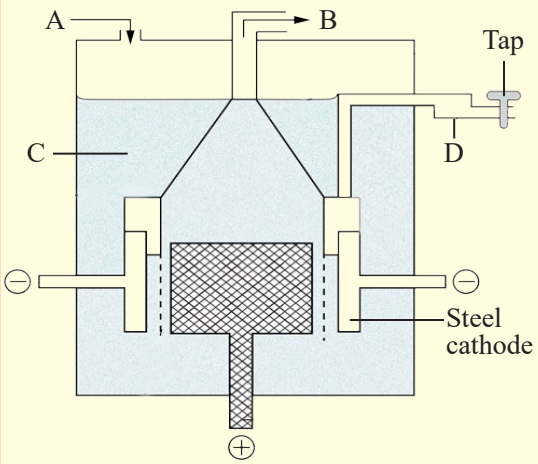
3. A less electropositive metal may be extracted by;
   (a) ore concentration and electrolytic reduction.
   (b) ore concentration and chemical reduction.
   (c) ore concentration and purification.
   (d) ore concentration, reduction and purification.

4. In the Down's cell, the substance considered as a by-product is;
   (a) oxygen.
   (b) chlorine.
   (c) sodium.
   (d) hydrogen.

5. Describe three methods that are used in extracting metals from metal ores. Give one example of a metal that is extracted by using each of the methods.

6. Metal ores contain impurities that must be removed during the extraction process.
   (a) Give three common impurities found in metal ores.
   (b) Why should the impurities be removed before the reduction process?

7. The following figure is a set-up used in the extraction of sodium by the Down's cell.



   (a) What is the major ore from which sodium is extracted in this cell?
   (b) What form of the ore is placed in the cell for the extraction?
   (c) Why is sodium not extracted by reduction using carbon?
   (d) What role does calcium chloride play in the extraction of sodium from sodium chloride?
   (e) Explain the roles of the parts marked A, B, C, and D.

8. Explain the following processes which are involved in the extraction of aluminium:
   (a) Bauxite ore should be heated to bright red with sodium carbonate.
   (b) An electrolyte must be added in the electrolysis of bauxite.

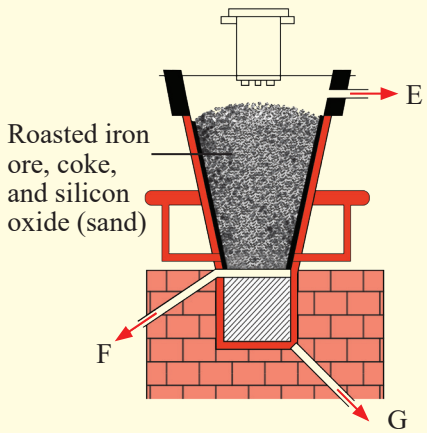
9. Iron is extracted from various ores by reduction in the blast furnace.
   (a) Why is the furnace called a blast furnace?

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(b) Identify the chief ore from which the iron metal is extracted.
(c) Describe the role of the blast furnace in the extraction process.
(d) What roles are played by tuyéres in the furnace?
(e) Write the equation for the reduction of the ore into iron.
(f) Explain the purpose of adding limestone and coke in the blast furnace.

10. The following figure is a sketch of the furnace for the extraction of copper:



(a) Explain the roles of parts E, F, and G.
(b) Give four ores from which copper can be extracted.
(c) What is the chief ore from which copper is extracted?
(d) Write an equation to show the reaction between copper(I) oxide and iron(II) sulphide.
(e) Why is the reaction in (d) possible?

11. Some metals were mixed with salt solutions as outlined below:
(a) Copper and sodium chloride
(b) Iron and copper(II) sulphate
(c) Potassium and sodium carbonate
(d) Calcium and magnesium chloride
(e) Lead and silver nitrate

In which mixture(s) did a reaction take place? Write the ionic equation for each reaction.

12. The following equation represents a reaction between potassium and iodine:

$$2K(s) + I_2(s) \longrightarrow 2KI(s)$$

(a) Identify the oxidising agent, and the reducing agent.
(b) Write the balanced half reaction equations for the oxidation and reduction reactions.

13. Suggest the ways which can be used to minimise environmental pollution due to liquid and solid wastes from mining.

14. Aluminium is mostly abundant on the Earth's crust but not cheaply available. Explain.

15. All ores are minerals, but not all minerals are ores. Explain.

16. Although iron pyrites is abundant in the Earth's crust, it is not used as a source of iron. Explain.

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Chapter Nine

Compounds of metals-Part I: oxides, hydroxides, carbonates, and hydrogencarbonates

Introduction

Metals react with other chemical substances, especially non-metals, to form different compounds. The compounds formed include metal oxides, hydroxides, carbonates, hydrogencarbonates, nitrates, chlorides, and sulphates. In this chapter, you will learn about the preparation, chemical properties, and uses of metal oxides, hydroxides, carbonates, and hydrogencarbonates. The competencies developed will enable you to solve various issues related to metal oxides, hydroxides, carbonates, and hydrogencarbonates in laboratories, industries, medicine, agriculture, the environment, and other areas.

Metal oxides

An *oxide* is a compound formed by the chemical combination of oxygen and another element. Metal oxides are compounds of metals and oxygen. These compounds require at least two elements and always contain a minimum of one oxygen atom. They tend to be solid, basic, and denser than their non-metallic oxide counterparts. When they react with water or acids, they produce bases or salts, respectively.

Preparation of metal oxides

Metal oxides can be prepared by direct and indirect methods.

Preparation of metal oxides by direct method

In the direct method, metal oxides are prepared by heating the respective metal in air or oxygen. Group I metals burn in air or oxygen to produce various ionic oxides such as simple oxides, peroxides, or superoxides. For example, when sodium is exposed to air, it reacts with the oxygen to form a whitish grey solid of sodium oxide. This is the reason as to why sodium is stored under oil.

$$4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$$

In excess air, sodium burns with a yellow flame to form sodium peroxide which is a pale yellow solid. Depending on the degree of heating, a mixture of sodium peroxide and sodium oxide can be formed at the same time.

$$6Na(s) + 2O_2(g) \longrightarrow Na_2O_2(s) + 2Na_2O(s)$$

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When sodium peroxide dissolves in water at room temperature, it forms an alkaline solution of NaOH, which changes a phenolphthalein indicator from colourless to pink. Oxygen gas is also evolved.

$$2Na_2O_2(s) + 2H_2O(l) \longrightarrow 4NaOH(aq) + O_2(g)$$

With ice-cold water, sodium peroxide forms sodium hydroxide and hydrogen peroxide.

$$Na_2O_2(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2O_2(l)$$

Sodium oxide reacts with water to give sodium hydroxide solution.

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

Magnesium burns in air with a very bright light to give a white solid of magnesium oxide.

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

Magnesium oxide reacts with water to form basic magnesium hydroxide solution.

$$MgO(s) + H_2O(l) \longrightarrow Mg(OH)_2(aq)$$

Similarly, calcium burns in air with a red flame to form calcium oxide which is white.

$$2Ca(s) + O_2(g) \longrightarrow 2CaO(s)$$

Calcium oxide reacts with water to form basic calcium hydroxide, which is sparingly soluble in water.

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$$

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Activity 9.1



Aim: To prepare metal oxides by direct method

Requirements: Bunsen burner, a pair of tongs, deflagrating spoon, cutting knives, watch glass, measuring cylinder, test tubes, filter paper, sandpaper or steel wool, calcium metal, magnesium ribbon, distilled water, and phenolphthalein indicator

Procedure

1. Clean two test tubes and label them A and B.
2. Transfer about 2 cm^3 of distilled water into each of the test tubes.
3. Add three drops of phenolphthalein indicator into each test tube.

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4. Cut a small piece of calcium metal and place it on a deflagrating spoon as shown in Figure 9.1, and then heat the metal until it starts to burn. Note the colour of its flame.

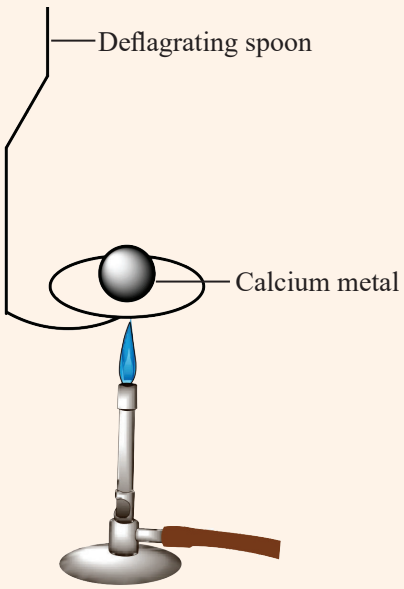


Figure 9.1: *Heating calcium metal on a deflagrating spoon*

5. Turn off the burner and allow the metal to continue burning until the flame goes off. Note the colour of the residues.

6. Leave the residues to cool, and then transfer them into test tube A. Shake the mixture and record your observations.

7. Repeat Steps 2–6 using magnesium ribbon (polished with sandpaper/ steel wool) in the place of calcium metal.
Caution! Do not look directly at a burning magnesium ribbon because the bright flame can harm your eyes. Wear safety goggles.

8. Transfer the magnesium residues into test tube B. Shake the mixture and record your observations.

Questions

1. What were the observed colours of the flames when magnesium and calcium metals burnt in air?
2. Write balanced chemical equations for the reactions in Question 1.
3. What were the colours of the residues obtained after magnesium and calcium metals burnt in air?
4. Use chemical equations to explain the reactions occurring when the metal residues were separately added into the test tubes containing distilled water.
5. Explain the significance of using phenolphthalein indicator in the experiments performed.

Preparation of metal oxides by indirect methods

The indirect methods of preparing metal oxides involve thermal decomposition of metal carbonates and hydroxides. They are carried out by:

(a) preparing a compound and then decomposing it by heating to give the respective metal oxide, or

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(b) decomposing an already prepared compound by heating it to give the oxide.

Preparation of metal oxides from metal carbonates

Metal carbonates decompose upon heating to give respective metal oxides and carbon dioxide gas. The degree of heating required to decompose the metal carbonates varies from one metal to another.

The carbonates of sodium and potassium are resistant to heat. These carbonates can decompose only when heated to temperatures greater than 800 °C to give metal oxides, $M_2O(s)$ (M is Na or K), and carbon dioxide. This is the reason why the indirect method is not preferred for the laboratory preparation of sodium oxide and potassium oxide.

The common carbonates of Group II are magnesium carbonate and calcium carbonate. These decompose when strongly heated to form metal oxide and carbon dioxide gas.

$$CaCO_3(s) \xrightarrow{\text{strong heat}} CaO(s) + CO_2(g)$$

Magnesium carbonate similarly decomposes on strong heating to give magnesium oxide and carbon dioxide gas. The carbonates of zinc, lead, copper, and iron decompose readily when heated to give their metal oxides and carbon dioxide. Zinc oxide is white when cold and yellow when hot, whereas lead(II) oxide is reddish-brown when hot and yellow when cold. Iron(II) carbonate decomposes to iron(II) oxide which oxidises immediately to iron(III) oxide.

$$FeCO_3(s) \xrightarrow{\text{heat}} FeO(s) + CO_2(g)$$

$$4FeO(s) + O_2(g) \longrightarrow 2Fe_2O_3(s)$$

The gas produced (CO_2) when metal carbonates decompose forms a white precipitate of calcium carbonate with calcium hydroxide solution (limewater).

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$$

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Activity 9.2



Aim: To prepare metal oxides by heating metal carbonates

Requirements: Test tubes, test tube rack, Bunsen burner, bent delivery tube, spatula, retort stand and clamp, small rubber bung, limewater, and carbonates of copper, zinc, lead, potassium, sodium, and magnesium

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Procedure

1. Clean six test tubes and label them J, K, L, M, N, and O.
2. Transfer a spatulaful of each of the six carbonates into the test tubes J, K, L, M, N, and O, respectively.
3. Set up the experiment as shown in Figure 9.2.

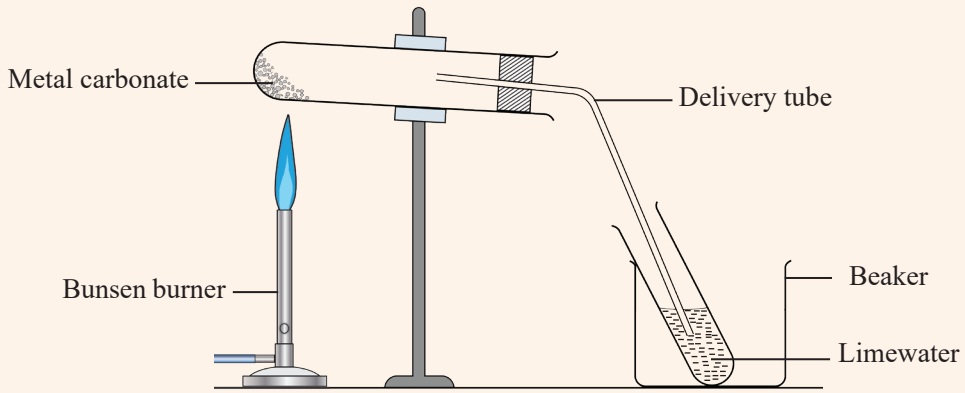


Figure 9.2: *Set up for the preparation of metal oxides from metal carbonates*

4. Heat each carbonate at a time.
5. Blow the gas evolved (if any) from each carbonate into limewater. Record any changes occurring in the limewater.
6. Record the colour of the residue left in the test tube (if any) both when hot and after cooling.

Questions

1. Give chemical equations for the reactions involved in the experiments performed.
2. With the aid of a balanced chemical equation, explain the significance of Step 5.
3. Comment on the reactivity of the used metal carbonates towards the action of heat.

Preparation of metal oxides from metal hydroxides

Metal hydroxides decompose upon heating to produce metal oxides and water. For example, when bluish green copper(II) hydroxide is heated, it decomposes to a black copper(II) oxide.

$$Cu(OH)_2(s) \xrightarrow{\text{heat}} CuO(s) + H_2O(g)$$

Lead(II) hydroxide, a white coloured compound decomposes on heating to form lead(II) oxide and water.

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$$Pb(OH)_2(s) \xrightarrow{\text{heat}} PbO(s) + H_2O(g)$$

Reddish-brown iron(III) hydroxide decomposes on heating to form iron(III) oxide which is reddish-brown in colour.

$$2Fe(OH)_3(s) \xrightarrow{\text{heat}} Fe_2O_3(s) + 3H_2O(g)$$

The water vapour produced after heating metal hydroxides turns cobalt chloride paper pink.

The hydroxides of zinc, magnesium, and aluminium similarly decompose to produce the corresponding metal oxides and water vapour. The hydroxides of sodium and potassium are stable and do not decompose on heating. Hydroxides of silver, gold, and mercury do not exist.

Activity 9.3

Aim: To prepare metal oxides by heating metal hydroxides

Requirements: Test tubes, Bunsen burner, spatula, retort stand and clamp, cobalt chloride paper, and hydroxides of zinc, magnesium, and aluminium

Procedure

1. Transfer a spatulaful of each hydroxide into a separate test tube.
2. Mount the test tubes as shown in Figure 9.3.

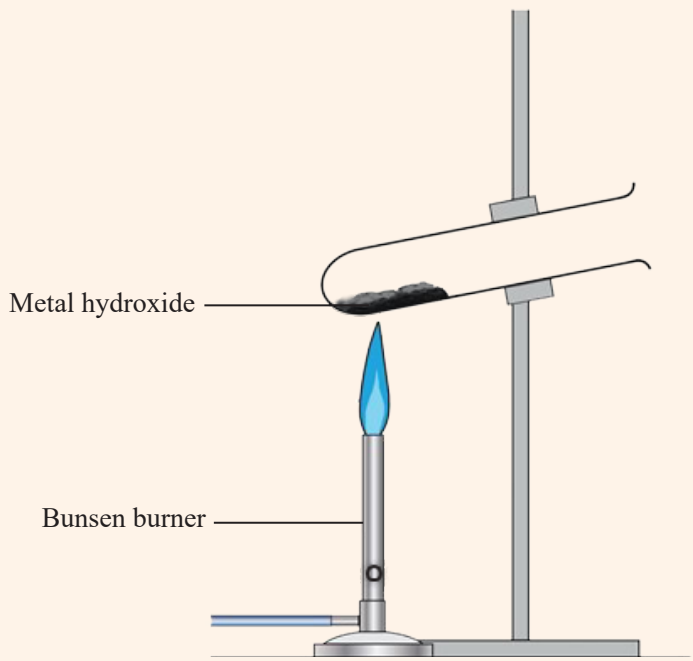


Figure 9.3: *Set-up for the preparation of metal oxides from metal hydroxides*

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3. Heat each metal hydroxide over a Bunsen burner flame until there is no further change in the colour of the residue left inside the test tube. Note the colour of the hot residue. Test any vapour produced with a blue cobalt chloride paper.

4. Allow the residue to cool and note its colour. Record all the observations as shown in the following table:

Hydroxide	Colour of the hydroxide	Colour of the residue when hot	Colour of the residue when cold

Questions

1. Write balanced chemical equations the reactions in all the experiments performed.
2. Based on your observations, explain the use of cobalt chloride paper in the experiments.
3. Explain what will happen if the hydroxides of sodium and potassium are used in the places of zinc and magnesium hydroxides.

Preparation of metal oxides from metal nitrates

Most metal nitrates decompose when heated to produce the corresponding metal oxide, brown fumes of nitrogen dioxide, and oxygen gas. For example, magnesium nitrate decomposes to form magnesium oxide, nitrogen dioxide, and oxygen.

$$2Mg(NO_3)_2(s) \xrightarrow{\text{heat}} 2MgO(s) + 4NO_2(g) + O_2(g)$$

Nitrates of zinc, iron, lead and copper also decompose to metal oxides on heating. Nitrates of sodium and potassium do not form oxides, instead they decompose slowly upon strong heating to give nitrites and oxygen gas.

$$2NaNO_3(s) \xrightarrow{\text{heat}} 2NaNO_2(s) + O_2(g)$$

In Activity 9.4, copper(II) oxide will be prepared by a decomposition of copper(II) nitrate synthesised by reacting concentrated nitric acid with copper. Concentrated nitric acid is a strong oxidising agent, so it oxidises copper to copper(II) nitrate. Nitric acid is itself reduced to nitrogen dioxide which is observed as brown fumes.

$$Cu(s) + 4HNO_3(l) \longrightarrow Cu(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$$

When the blue solution of copper(II) nitrate is heated to dryness, blue crystals of copper(II) nitrate are observed on the sides of the boiling tube. The prepared copper(II) nitrate is then decomposed by heating to produce copper(II) oxide.

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Activity 9.4



Aim: To prepare copper(II) oxide from copper granules and concentrated nitric acid

Requirements: Beakers, measuring cylinders, glass rod, boiling tube, test tube holder, Bunsen burner, wooden splint, concentrated nitric acid, and copper granules

Note: This experiment should be performed in a fume chamber

Caution! Concentrated nitric acid is corrosive. Handle it with care.

Procedure

1. Transfer 20 cm^3 of concentrated nitric acid into a beaker.
2. Add copper granules into the beaker containing concentrated nitric acid and stir the mixture using a glass rod.
3. Continue adding the granules until there is no more dissolution.
4. Decant the mixture into a separate beaker.
5. Transfer 3 cm^3 of the decanted mixture into a boiling tube and heat to completely evaporate the solvent.
6. Continue heating the solids in the tube until you observe no colour change on the solid product.
7. Test any evolving gas with a glowing wooden splint. Record your observations.

Questions

1. Write a balanced chemical equation to show the reaction in Step 6.
2. What is the colour of the residue left in the reaction occurring in Step 6?
3. What is the colour of the fumes produced in Step 6 (if any)?
4. What does the colour of the fumes suggest about the identity of the gas produced?
5. Explain the importance of Step 7.
6. Why did you add more granules in Step 3?

Classification of metal oxides

Metal oxides can be grouped into soluble or insoluble oxides based on their solubility in water and as basic or amphoteric oxides based on their reactions with acids and bases.

Soluble metal oxides readily dissolve in water to form metal hydroxide solutions. They include potassium oxide, sodium oxide, and magnesium oxide. Soluble metal oxides are known as *alkalis*. Other oxides are *insoluble* in water. Examples of such oxides include aluminium oxide, copper(II) oxide, zinc oxide, lead(II) oxide, iron(II) oxide, and iron(III) oxide.

Basic metal oxides react with acids to give salt and water as the only products. Such reactions are referred to as *neutralisation reactions*. Basic metal oxides include all

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the oxides of Group I and Group II metals such as sodium, potassium, lithium, magnesium, and calcium. The oxides of aluminium, zinc, and lead react with both acids and bases. They are known as *amphoteric oxides* as they have alkaline and acidic properties.

Chemical properties of metal oxides

Metal oxides react differently with water, acids, and alkalis. The following are some of the chemical properties of metal oxides:

Reactions of metal oxides with water

The oxides of sodium and potassium react vigorously with water to form basic solutions as indicated in the following equations:

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

$$K_2O(s) + H_2O(l) \longrightarrow 2KOH(aq)$$

The oxides of calcium and magnesium similarly react with water to form basic metal hydroxides. For example, when calcium oxide is added to water, it reacts with water to form basic calcium hydroxide.

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$$

Aluminium oxide, zinc oxide, lead(II) oxide, copper(II) oxide, iron(II) oxide, and iron(III) oxides do not react with water. They are all insoluble in water.

Activity 9.5



Aim: To investigate the reactions of metal oxides with water

Requirements: Beaker, glass rod, spatula, test tube, glass tube, red litmus paper, oxides of calcium, magnesium, aluminium, lead, zinc, copper, and iron, and distilled water

Procedure

1. Transfer a spatulaful of calcium oxide into a clean beaker.
2. Add about 10 cm^3 of distilled water in the beaker containing the oxide and stir using a glass rod.
3. Dip a red litmus paper into the mixture and record any colour change.
4. Filter the mixture into a test tube. Blow air from the mouth into the solution using a glass tube. Record the results.
5. Repeat Steps 1–4 using the oxides of magnesium, aluminium, lead, copper, iron, and zinc, and record your observations.

Questions

1. Write balanced chemical equations for the formation of the reaction products (if any).
2. What are the colours of the formed reaction products (if any)?
3. Explain the significance of using a red litmus paper.
4. From your observations, comment on the reactivity of the metal oxides with water.

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Reactions of metal oxides with acids and alkalis

Most metal oxides react with acids such as nitric acid, hydrochloric acid, and sulphuric acid to form salt and water. Basic oxides react with acids, whereas amphoteric oxides react with both acids and alkalis to give salt and water.

Dilute nitric acid reacts with metal oxides to form metal nitrates and water. All metal nitrates are soluble in water. Dilute hydrochloric acid reacts with metal oxides to give metal chlorides and water. Dilute sulphuric acid reacts with metal oxides to give metal sulphates and water. When dilute sulphuric acid and hydrochloric acid react with lead(II) oxide, they form insoluble salts of lead(II) sulphate, and lead(II) chloride, respectively. The insoluble salts form a protective coating on the oxides, thus preventing further reaction.

$$PbO(s) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + H_2O(l)$$

$$PbO(s) + 2HCl(aq) \longrightarrow PbCl_2(s) + H_2O(l)$$

The reaction of calcium oxide with sulphuric acid also produces an insoluble calcium sulphate. The reactions of metal oxides with dilute mineral acids are shown in Table 9.1.

Table 9.1: *Reactions of metal oxides with dilute mineral acids*

Metal oxides	Reaction
Dilute nitric acid	
Calcium oxide	$CaO(s) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + H_2O(l)$
Magnesium oxide	$MgO(s) + 2HNO_3(aq) \longrightarrow Mg(NO_3)_2(aq) + H_2O(l)$
Aluminium oxide	$Al_2O_3(s) + 6HNO_3(aq) \longrightarrow 2Al(NO_3)_3(aq) + 3H_2O(l)$
Iron(II) oxide	$FeO(s) + 2HNO_3(aq) \longrightarrow Fe(NO_3)_2(aq) + H_2O(l)$
Lead(II) oxide	$PbO(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + H_2O(l)$
Copper(II) oxide	$CuO(s) + 2HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + H_2O(l)$
Dilute hydrochloric acid	
Calcium oxide	$CaO(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l)$
Magnesium oxide	$MgO(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2O(l)$
Aluminium oxide	$Al_2O_3(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2O(l)$
Iron(II) oxide	$FeO(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2O(l)$
Lead(II) oxide	$PbO(s) + 2HCl(aq) \longrightarrow PbCl_2(s) + H_2O(l)$
Copper(II) oxide	$CuO(s) + 2HCl(aq) \longrightarrow CuCl_2(aq) + H_2O(l)$
Dilute sulphuric acid	
Calcium oxide	$CaO(s) + H_2SO_4(aq) \longrightarrow CaSO_4(s) + H_2O(l)$
Magnesium oxide	$MgO(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2O(l)$
Aluminium oxide	$Al_2O_3(s) + 3H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3H_2O(l)$
Iron(II) oxide	$FeO(s) + H_2SO_4(aq) \longrightarrow FeSO_4(aq) + H_2O(l)$
Lead(II) oxide	$PbO(s) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + H_2O(l)$
Copper(II) oxide	$CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$

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Activity 9.6



Aim: To investigate the amphoteric properties of some metal oxides

Requirements: Spatula, droppers, test tubes, Bunsen burner, aluminium(III) oxide, zinc oxide, lead(II) oxide, dilute nitric acid, and dilute sodium hydroxide solution

Procedure

1. Label three test tubes with the names of the given oxides.
2. Transfer a spatulaful of each oxide into its respective test tube.
3. Using a dropper, add dilute nitric acid into the test tubes dropwise. Observe any changes. If no observable changes occur in the mixture, warm it.
4. Repeat Steps 1–3 using dilute sodium hydroxide in the place of dilute nitric acid. Record your observations.

Question

Discuss the observations made from the experiments.

Activity 9.7



Aim: To investigate the reactions of metal oxides with dilute acids

Requirements: Measuring cylinder, glass rod, spatula, tripod stand, wire gauze, Bunsen burner, test tubes, oxides of calcium, magnesium, lead, iron(II), and copper(II), 2 M nitric acid, 2 M hydrochloric acid, and 2 M sulphuric acid

Procedure

1. Label five test tubes with the names of the given oxides.
2. Transfer a spatulaful of each metal oxide into its respective test tube.
3. Add about 10 cm^3 of 2 M nitric acid into each test tube. If no observable changes are made at room temperature, warm the mixture. Record the observations.
4. Repeat Steps 1–3 using 2 M sulphuric acid and 2 M hydrochloric acid in the place of 2 M nitric acid.

Questions

1. Discuss the observations made from the experiments.
2. Write the balanced chemical equations for the reaction of each metal oxide with the three dilute acids.

Uses of metal oxides

Metal oxides have different uses which are related to the properties of the specific metal oxide. The uses of selected metal oxides are described in this section.

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Calcium oxide

The following are the uses of calcium oxide:

(a) Preparation of calcium carbide
Calcium carbide is prepared by strongly heating calcium oxide and coke in an electric furnace at 250 °C.

$$CaO(s) + 3C(s) \longrightarrow CaC_2(s) + CO(g)$$

Calcium carbide is used in synthesising calcium cyanamide fertiliser, steel making, manufacture of acetylene gas, and in carbide lamps.

(b) Lining of furnaces
Calcium oxide is used as a refractory in furnaces. A *refractory* is a solid material that resists decomposition by heat, pressure, or chemical attack, and retains its strength and form at high temperatures. Calcium oxide has a high melting point of 2580 °C, making it suitable for use as a refractory.

(c) Formation of slag
During the extraction of iron, calcium carbonate is added into the blast furnace. The carbonate decomposes to calcium oxide which acts as a flux and also combines with silica (SiO_2) present as an impurity in iron to form calcium silicate (fusible slag).

$$CaO(s) + SiO_2(s) \xrightarrow{\text{heat}} CaSiO_3(l)$$

The fusible slag floats over molten iron and prevents iron from being oxidised.

(d) Drying agent
Calcium oxide is used as a drying agent in the preparation of ammonia gas and ethanol.

(e) Preparing mortar
Calcium oxide reacts with water to make slaked lime (calcium hydroxide). Slaked lime, sand and water are mixed to make mortar which is used to stick bricks together. The mixture of slaked lime, sand, and water is also used to plaster buildings.

(f) Manufacture of cement
In the manufacture of cement, clay which contains hydrated aluminium silicate is heated together with lime (calcium oxide) or limestone ($CaCO_3$) to form a mixture of calcium silicates and aluminates of calcium, commonly known as *cement.*

(g) Soil treatment
Lime is added to soils with low pH (acidic soils) to raise their pH. Lime is also used to make clay soil more porous, thus improving the texture of the soil. A mixture of magnesium oxide and calcium oxide, commonly known as *magmax*, is also added to acidic soils to improve their pH. The mixture also acts as a supplier of calcium and magnesium nutrients which are needed for the growth of plants.

Magnesium oxide

Magnesium oxide is used for making the lining for open-hearth and steel furnaces because of its high melting point of about 2800 °C.

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Zinc oxide

Zinc oxide is used as a white pigment in paints, a filler in rubber and as a component of glasses, enamels and antiseptic ointments. In drug products, zinc oxide is used as a skin protectant and a sunscreen. Zinc oxide works as a sunscreen by reflecting and scattering ultraviolet radiation. Sunscreen reduces or prevents sunburn and premature aging of the skin.

Copper(II) oxide

Copper(II) oxide is used as a starting material for the production of other copper salts. For example, many wood presevatives are produced from copper(II) oxide. In addition, copper(II) oxide is used as a pigment in ceramics to produce blue, red, and green, and sometimes gray, pink, or black glazes. It is also used when welding with copper alloys.

Aluminium oxide

The major uses of aluminium oxide are in refractories, ceramics, polishing, and abrasive applications. For example, corundum which is the most common naturally occurring crystalline form of aluminium oxide is used as an abrasive component in sandpaper.

Lead oxide

Lead oxide is an important industrial material. It is widely used in batteries, gas sensors, pigments and paints, ceramics, glass industry, and as a catalyst in synthesis of organic compounds.

Iron oxides

The oxides are used as iron ores for extraction of iron, manufacture of pigments, and as catalysts. In addition, iron oxide is used to make thermite. From iron oxides, cheap and durable pigments, coatings, and coloured concretes are manufactured.

Exercise 9.1

1. (a) Give two methods of preparing metal oxides.
   (b) Using balanced chemical equations, give one example of an oxide that can be prepared by using each of the two methods you have given.
2. Write a balanced chemical equation to show:
   (a) The reaction of sodium with atmospheric air.
   (b) The burning of sodium in air.
   (c) The reaction of sodium peroxide with water at room temperature.
   (d) The reaction of sodium peroxide with ice-cold water.
3. (a) Name two metal oxides which are soluble in water.
   (b) Give five uses of metal oxides.

Metal hydroxides

A *hydroxide* is any inorganic compound that consists of hydroxide ions (OH^-). Most metal hydroxides are bases and have pH values greater than 7. Alkali metal hydroxides such as sodium hydroxide are soluble in water and form very strong bases. Alkaline earth metal hydroxides such as calcium hydroxide are only sparingly soluble in water.

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Preparation of metal hydroxides

Metal hydroxides can be prepared either by *direct* or *indirect* methods.

Preparation of metal hydroxides by the direct method

In the direct method, the metals high in the reactivity series are reacted with water to form hydroxides. Majority of the Group I and Group II metals react with water to give metal hydroxides and hydrogen gas. Group I metals are known to react vigorously with water. For instance, potassium reacts vigorously with water to form an alkaline solution of potassium hydoxide which turns red litmus paper blue, with the evolution of hydrogen gas.

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

Sodium also reacts vigorously with water to form an alkaline solution of sodium hydoxide and hydrogen gas.

When calcium metal is added into water, effervescence are observed. The gas produced burns with a pop sound indicating that it is hydrogen gas. A white suspension of calcium hydroxide is left when the reaction is complete. Since calcium hydroxide is sparingly soluble in water, a small amount of it dissolves in water to give an alkaline solution. Magnesium forms magnesium hydroxide and hydrogen gas with water, but the reaction is quite slow.

$$Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$$

Hydroxides of metals below hydrogen in the reactivity series cannot be prepared by the direct method. Details about the reactivity series of metals are given in Chapter 6.

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Activity 9.8



Aim: To prepare metal hydroxides by direct method

Requirements: Beaker, measuring cylinder, test tube, red litmus paper, wooden splint, calcium metal, sodium metal, and distilled water

Procedure

1. Transfer about 10 cm^3 of distilled water into a beaker.
2. Put a clean piece of calcium metal into the water. Record your observations.
3. Test the evolving gas using a burning wooden splint and record your observations.
4. Dip a red litmus paper into the reaction mixture and record the colour change (if any).
5. Repeat Steps 1–4 using sodium metal instead of calcium. Ensure you use a small piece of sodium (size of a rice grain).

Questions

1. Which gas evolved when the two metals reacted with water? How do you confirm the identity of the gas?

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2. Write balanced chemical equations for the reactions that occurred in this experiment.
3. What was the significance of using a litmus paper?
4. What will happen if a large piece of sodium metal is used in Step 5?

Preparation of metal hydroxides by indirect method

In the preparation of metal hydroxides using the indirect method, the aqueous salts of metals are reacted with the aqueous solutions of alkalis such as NaOH(*aq*) and KOH(*aq*). The reactions are generally referred to as precipitation reactions as the hydroxides are obtained as precipitates. For example copper(II) sulphate solution reacts with dilute sodium hydroxide to form a blue precipitate of copper(II) hydroxide. Its ionic equation is presented as follows:

$$Cu^{2+}(aq) + 2OH^{-}(aq) \longrightarrow Cu(OH)_2(s)$$

Iron(III) chloride solution also reacts with dilute sodium hydroxide to form a precipitate of iron(III) hydroxide. Zinc nitrate solution reacts with just enough dilute sodium hydroxide to form a white precipitate of zinc hydroxide. The reaction is represented by the following ionic equation:

$$Zn^{2+}(aq) + 2OH^{-}(aq) \longrightarrow Zn(OH)_2(s)$$

If excess sodium hydroxide is used, the zinc hydroxide will dissolve to form a zincate complex ion.

$$Zn(OH)_2(s) + 2OH^{-}(aq) \longrightarrow \left[Zn(OH)_4\right]^{2-}(aq)$$

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Activity 9.9



Aim: To prepare metal hydroxides by indirect method

Requirements: Beakers, dropper, measuring cylinder, wash bottle, filter funnel, filter papers, 1 M sodium hydroxide solution, 1 M copper(II) sulphate solution, 1 M zinc nitrate solution, 1 M iron(III) chloride solution, and distilled water

Procedure

1. Label three beakers as X, Y, and Z, respectively, and transfer about 10 cm^3 of 1 M copper(II) sulphate solution into beaker X.
2. Add a few drops of sodium hydroxide solution into beaker X until a significant amount of precipitates forms. Record the colour of the precipitates formed.

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3. Filter the precipitates.
4. Wash the precipitates with distilled water and leave them on the filter paper or watch glass to dry.
5. Repeat Steps 2–4 using 1 M zinc nitrate solution for beaker Y, and 1 M iron(III) chloride solution for beaker Z.

Note: Just enough sodium hydroxide solution should be added into beaker Y to allow the precipitation.

Questions

1. Write balanced chemical equations to show the products formed in each beaker.
2. What is the colour of the precipitates formed in each beaker?
3. Explain what will happen if potassium hydroxide is used instead of sodium hydroxide.

Classification of metal hydroxides

Metal hydroxides are classified based on their solubility in water as soluble and insoluble hydroxides. They are also classified into *basic* and *amphoteric* hydroxides based on their reactions with acids and bases. Basic hydroxides react with acids to give salt and water only.

When excess sodium hydroxide is added to the insoluble metal hydroxides of zinc, lead, and aluminium, they dissolve to form complex ions. These hydroxides are said to be amphoteric as they react with both bases and acids.

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Reactions with bases

Zinc hydroxide dissolves in excess sodium hydroxide to form sodium zincate complex.

$$Zn(OH)_2(s) + 2NaOH(aq) \longrightarrow Na_2[Zn(OH)_4](aq)$$

Lead(IV) hydroxide dissolves in excess sodium hydroxide to form sodium plumbate solution.

$$Pb(OH)_4(s) + 2NaOH(aq) \longrightarrow Na_2PbO_3(aq) + 3H_2O(l)$$

Aluminium hydroxide dissolves in excess sodium hydroxide to form sodium aluminate solution.

$$Al(OH)_3(s) + NaOH(aq) \longrightarrow NaAlO_2(aq) + 2H_2O(l)$$

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Reactions with acids

Amphoteric hydroxides react with acids to form salts and water. For example,

$$Pb(OH)_2(s) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + 2H_2O(l)$$

$$Al(OH)_3(s) + 3HCl(aq) \longrightarrow AlCl_3(aq) + 3H_2O(l)$$

Activity 9.10



Aim: To investigate the solubility of metal hydroxides in sodium hydroxide

Requirements: Test tubes, droppers, measuring cylinder, test tube rack, 1 M of each of the following solutions: sodium hydroxide, copper(II) chloride, iron(III) chloride, magnesium chloride, zinc chloride, lead(II) nitrate, and sodium chloride

Procedure

1. Place seven clean test tubes in a test tube rack and label them with the names of metal salts given.
2. Transfer about 3 cm^3 of each of the salt solutions into respective test tubes.
3. Add dilute sodium hydroxide dropwise in each of the test tubes until excess.
4. Record all the observations made.

Questions

1. With the aid of chemical equations, explain what will happen if the sodium hydroxide in Step 3 is not added in excess. Give balanced chemical equations.
2. Comment on the solubility of the metal hydroxides in sodium hydroxide solution.

Chemical properties of metal hydroxides

Metal hydroxides take part in different chemical reactions. These include thermal decomposition and neutralisation reactions.

Action of heat on metal hydroxides

Hydroxides of the metals high in the reactivity series such as potassium and sodium do not decompose on heating. Other metal hydroxides decompose to give metal oxides and water. For example, lead(II) hydroxide decomposes to give lead oxide and water.

$$Pb(OH)_2(s) \xrightarrow{\text{heat}} PbO(s) + H_2O(l)$$

Action of acids on metal hydroxides

Metal hydroxides react with mineral acids to give salt and water.

For example, copper(II) hydroxide is a blue solid and reacts with both hydrochloric

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acid and sulphuric acid to form copper(II) chloride and copper(II) sulphate solutions, respectively.

$$Cu(OH)_2(s) + 2HCl(aq) \longrightarrow CuCl_2(aq) + 2H_2O(l)$$

$$Cu(OH)_2(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + 2H_2O(l)$$

Similarly, iron(II) hydroxide which is a green solid, reacts with both hydrochloric acid and sulphuric acid to form green solutions of iron(II) chloride and iron(II) sulphate, respectively.

$$Fe(OH)_2(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + 2H_2O(l)$$

$$Fe(OH)_2(s) + H_2SO_4(aq) \longrightarrow FeSO_4(aq) + 2H_2O(l)$$

When left to stand for a few minutes, the light green solutions of iron(II) chloride and iron(II) sulphate change to reddish-brown solutions due to oxidation of iron(II) ions (Fe^{2+}) into iron(III) ions (Fe^{3+}) by air.

$$Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^-$$

The reactions of lead(II) hydroxide with hydrochloric acid and sulphuric acid stop immediately after they start because of the formation of insolube lead(II) chloride and lead(II) sulphate, respectively.

$$Pb(OH)_2(s) + 2HCl(aq) \longrightarrow PbCl_2(s) + 2H_2O(l)$$

$$Pb(OH)_2(s) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + 2H_2O(l)$$

These salts form a coating on the lead hydroxide and prevent further action of the acid on the lead(II) hydroxide.

Sodium hydroxide and potassium hydroxide react with both hydrochloric acid and sulphuric acid to form soluble metal chlorides and sulphates, respectively.

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

$$KOH(aq) + HCl(aq) \longrightarrow KCl(aq) + H_2O(l)$$

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

Reactions between metal hydroxides and dilute nitric acid give metal nitrates and water. All metal nitrates are soluble in water. The reactions proceed as shown in the following general equation:

$$M(OH)_x(aq) + zHNO_3(aq) \longrightarrow M(NO_3)_x(aq) + yH_2O(l)$$

where M is the metal forming the hydroxide, x is the metal valence, z and y are stoichiometric coefficients.

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Activity 9.11



Aim: To investigate the action of mineral acids on metal hydroxides

Requirements: Measuring cylinder, beakers, spatula, blue and red litmus papers, copper(II) hydroxide, aluminium hydroxide, zinc hydroxide, lead(II) hydroxide, 1 M sulphuric acid, and 1 M hydrochloric acid

Procedure

1. Put a spatulaful of copper(II) hydroxide in a beaker.
2. Add 5 cm^3 of 1 M hydrochloric acid into the beaker containing copper(II) hydroxide. Record your observations.
3. Dip a blue litmus paper, then a red one into the mixture. Note any change.
4. Leave the reaction to proceed for about 5 minutes and record any changes.
5. Repeat Steps 1–4 for aluminium hydroxide, lead(II) hydroxide, and finally with zinc hydroxide, one after another.
6. Repeat Steps 1–4 for all the metal hydroxides using 1 M sulphuric acid in the place of the hydrochloric acid. Record your observations.

Questions

1. Write balanced chemical equations for the reaction of each metal hydroxide with hydrochloric acid and sulphuric acid.
2. What are the colours of the formed products?
3. Explain the importance of the two litmus papers used.

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Uses of metal hydroxides

The uses of metal hydroxides are very similar to those of the corresponding metal oxides. The following are some areas where metal hydroxides are used:

Agriculture: Calcium hydroxide (slaked lime) is used in soil treatment to reduce acidity.

Medicine: Aluminium hydroxide and magnesium hydroxide are used as antacids to neutralise stomach acidity. Magnesium hydroxide (milk of magnesia) is a common *laxative*.

Building and construction: Whitewash is a thick suspension of slaked lime in water. It is smeared on the walls of buildings to give a smooth protective finish. Slaked lime mixed with sand and water is used to stick bricks together.

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Bleaching of pulp: Calcium hydroxide is used in the preparation of calcium hydrogensulphite, $Ca(HSO_3)_2$ which is required in preparing sulphite pulp from wood. The pulp is used in making paper and artificial silk.

Softening of water: Calcium hydroxide is used to remove temporary hardness of water caused by metal hydrogencarbonates. The metal hydrogencarbonates in water are converted into their carbonates and precipitated out. For example,

$$Ca(OH)_2(aq) + Ca(HCO_3)_2(aq) \longrightarrow 2CaCO_3(s) + 2H_2O(l)$$

Manufacture of paints: Calcium hydroxide is used in the manufacture of undercoat paints which are applied as the first coat on plaster walls or on wood before applying the final gloss paint.

Qualitative analysis: Calcium hydroxide is sparingly soluble in water forming limewater that is used as a qualitative test for carbon dioxide. When carbon dioxide is passed through limewater, a white precipitate of calcium carbonate is formed.

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$$

Sodium hydroxide is also an important reagent in the qualitative analysis of metal cations. Zinc, lead, and aluminium ions form precipitates with sodium hydroxide solution which dissolve in excess alkali.

Extraction of metals: Sodium hydroxide is used in the extraction of aluminium from bauxite ore.

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Exercise 9.2

1. Amphoteric hydroxides react with both acids and alkalis.
   (a) Give three examples of amphoteric hydroxides.
   (b) Write equations to show the reactions between each of the three metal hydroxides with dilute nitric acid and excess sodium hydroxide solution.
2. Complete the following reactions:
   (a) $NaOH(s) \xrightarrow{\text{heat}}$
   (b) $Pb(OH)_2(s) \xrightarrow{\text{heat}}$
3. Give five uses of metal hydroxides.

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Carbonates and hydrogencarbonates of metals

Metal carbonates and hydrogencarbonates are derivatives of carbonic acid. Metal carbonates are formed when both hydrogen atoms in carbonic acid are replaced by a metal, while metal hydrogencarbonates are formed when one of the hydrogen atoms in carbonic acid is replaced by a metal ion.

Metal carbonates

When atmospheric carbon dioxide dissolves in water, a dibasic carbonic acid is formed. This means that it has two replaceable hydrogen atoms in its molecule. Metal carbonates are formed when both hydrogen atoms in carbonic acid are replaced by a metal.

Preparation of metal carbonates

The methods used to prepare carbonates depend on whether the carbonate is soluble or insoluble in water.

Preparation of soluble metal carbonates

Carbonates of alkaline metals, except lithium carbonate are soluble in water and give colourless solutions. Examples of soluble carbonates are sodium carbonate and potassium carbonate. They are prepared in the laboratory by reacting their corresponding alkalis with carbon dioxide as shown in the following general equation:

$$2MOH(aq) + CO_2(g) \longrightarrow M_2CO_3(aq) + H_2O(l)$$

where M is Na or K.

When the carbonate solutions are evaporated to saturation, crystals of the corresponding metal carbonates are formed. The crystalline sodium carbonate is known as *washing soda*. The reacting carbon dioxide gas to produce metal carbonate should be free from any acid spray, otherwise the acid will react with the carbonate formed to produce carbon dioxide and water.

$$2H^+(aq) + CO_3^{2-}(aq) \longrightarrow CO_2(g) + H_2O(l)$$

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Activity 9.12



Aim: To prepare sodium carbonate (a soluble metal carbonate)

Requirements: Beaker, Bunsen burner, evaporating dish, measuring cylinder, carbon dioxide generator, bent glass tube, and dilute sodium hydroxide

Procedure

1. Transfer about 25 cm^3 of dilute sodium hydroxide into a beaker.
2. Pass some carbon dioxide through the solution from a carbon dioxide gas generator at a moderate rate for about 5 minutes. The gas should be free from any acid spray.

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3. Transfer 5 cm^3 of the solution from the beaker and put in an evaporating dish.
4. Heat the solution to complete evaporation, and then leave it to cool. Note the colour of the crystals formed.

Question

Write a balanced chemical equation for the reaction between sodium hydroxide and carbon dioxide.

Preparation of insoluble metal carbonates

All insoluble metal carbonates can be prepared by precipitation method. This involves the reaction between a soluble carbonate and the solution of the metal salt. For example, a solution of sodium carbonate reacts with a solution of calcium chloride to give calcium carbonate which is insoluble. Table 9.2 shows some chemical equations for the preparation of some insoluble metal carbonates.

Table 9.2: *Chemical equations for the preparation of some insoluble metal carbonates*

Metal carbonate	Reaction
Magnesium carbonate	$MgCl_2(aq) + Na_2CO_3(aq) \longrightarrow MgCO_3(s) + 2NaCl(aq)$
Zinc carbonate	$ZnCl_2(aq) + K_2CO_3(aq) \longrightarrow ZnCO_3(s) + 2KCl(aq)$
Iron(II) carbonate	$FeCl_2(aq) + Na_2CO_3(aq) \longrightarrow FeCO_3(s) + 2NaCl(aq)$

Note that sodium carbonate is not used to prepare carbonates of zinc, copper, and lead because it precipitates basic carbonates of the metals which are $ZnCO_3 \cdot Zn(OH)_2$, $CuCO_3 \cdot Cu(OH)_2$, and $PbCO_3 \cdot Pb(OH)_2$, respectively instead of precipitating their simple carbonates. Aluminium carbonate and iron(III) carbonate do not exist because they are unstable and therefore they decompose if exposed to heat or moisture.

Activity 9.13



Aim: To prepare calcium carbonate (an insoluble metal carbonate)

Requirements: Beakers, filter funnel, filter paper, stand and clamp, measuring cylinder, 2 M sodium carbonate, 2 M calcium chloride, and distilled water

Procedure

1. Transfer about 10 cm^3 of sodium carbonate solution into a beaker.
2. Add about 10 cm^3 of calcium chloride into the sodium carbonate solution in the beaker. Shake the mixture.

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3. Filter the precipitates and wash them with distilled water.
4. Leave the precipitate to dry under the sun.

Questions

1. Explain what will happen if potassium carbonate is used instead of sodium carbonate. Give a chemical equation to support your answer.
2. Write a balanced chemical equation for the reaction occurring in Step 2 of the experimental procedure.

Classification of metal carbonates

Carbonates can be classified as *soluble* or *insoluble* based on their solubility in water. Soluble carbonates dissolve in water to form solutions. As highlighted previously, carbonates of alkali metals, and ammonium which is the only non metal carbonate, are soluble in water except lithium carbonate. All other metal carbonates are insoluble in water. Table 9.3 shows the classification of some metal carbonates based on their solubility in water.

Table 9.3: *List of some soluble and insoluble metal carbonates*

Group of carbonates	Examples
Soluble carbonates	Potassium carbonate Sodium carbonate
Insoluble carbonates	Calcium carbonate Magnesium carbonate Zinc carbonate Iron(II) carbonate Lead(II) carbonate Copper(II) carbonate

Chemical properties of metal carbonates

The chemical properties of a metal carbonate is related to the position of the metal in the reactivity series. Metals which are higher in the reactivity series are more reactive, thus they form more stable metal carbonates than those lower in the series. The following are some of the chemical properties of metal carbonates:

Action of heat on metal carbonates

The carbonates of potassium and sodium are very stable towards heating. The other metal carbonates decompose to give the corresponding metal oxide and carbon dioxide. Magnesium carbonate and calcium carbonate only decompose on strong heating.

$$CaCO_3(s) \xrightarrow{\text{strong heat}} CaO(s) + CO_2(g)$$

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Some carbonates decompose to give more than two products as shown in the following examples:

(a) Mercury(II) carbonate decomposes on heating to give mercury, carbon dioxide, and oxygen.

$$2HgCO_3(s) \xrightarrow{\text{heat}} 2Hg(l) + 2CO_2(g) + O_2(g)$$

Note: Mercury containing compounds are prohibited due to their high toxicity, including carcinogenic nature.

(b) Silver carbonate decomposes upon heating to give silver, carbon dioxide, and oxygen.

$$2Ag_2CO_3(s) \xrightarrow{\text{heat}} 4Ag(l) + 2CO_2(g) + O_2(g)$$

(c) Ammonium carbonate decomposes upon heating to give ammonia gas, carbon dioxide, and water vapour.

$$(NH_4)_2CO_3(s) \longrightarrow 2NH_3(g) + CO_2(g) + H_2O(g)$$

Activity 9.14

Aim: To investigate the effects of heat on different metal carbonates

Requirements: Test tubes, test tube rack, Bunsen burner, delivery tube, spatula, stand and clamp, limewater, and carbonates of magnesium, zinc, lead, and copper

Procedure

1. Put a spatulaful of magnesium carbonate in a test tube.
2. Set the apparatus as shown in Figure 9.4.

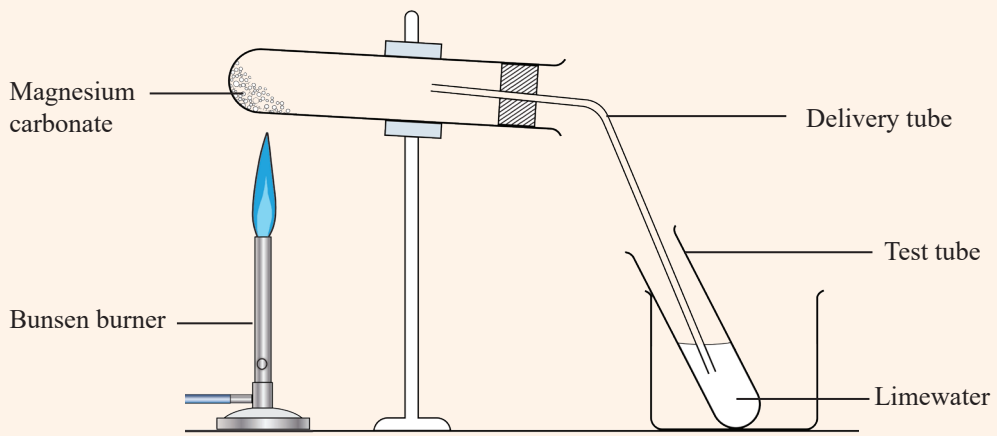


Figure 9.4: *Set-up for investigating the effects of heat on metal carbonates*

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3. Heat the contents strongly for 3 to 5 minutes. Blow the gas evolved into limewater and note any colour change or precipitation. Note also the colour of the residue after heating.
4. Leave the set-up to cool. Note the final colour of the residue after cooling.
5. Repeat Steps 1–4 for the other carbonates.
6. Record your observations as shown in the following table:

Metal Carbonate	Colour of the residue when hot	Colour of the residue on cooling

Questions

1. Based on your observations, comment on the relationships between the decomposition temperature and the stability of the metal carbonates.
2. With the aid of a chemical equation, explain what happens when the gas evolved in Step 3 is blown into limewater.
3. Write balanced chemical equations for the decomposition of the metal carbonates used in the experiment.

Action of dilute acids on metal carbonates

All metal carbonates react with dilute mineral acids to give metal salts, carbon dioxide gas and water. For example, sodium carbonate reacts with dilute hydrochloric acid to give a colourless sodium chloride solution, water, and carbon dioxide gas.

$$Na_2CO_3(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$$

Copper(II) carbonate and zinc carbonate react in a similar way. When lead(II) carbonate is reacted with dilute hydrochloric acid, a slight reaction occurs with evolution of carbon dioxide. The reaction stops immediately due to the formation of insoluble lead(II) chloride which forms a coating on the lead(II) carbonate that inhibits further action of the acid on the carbonate.

$$PbCO_3(s) + 2HCl(aq) \longrightarrow PbCl_2(s) + CO_2(g) + H_2O(l)$$

Dilute sulphuric acid reacts with metal carbonates to give metal sulphates, carbon dioxide, and water. Lead(II) carbonate and calcium carbonate react with dilute sulphuric acid to form insoluble lead(II) sulphate and calcium sulphate, respectively.

$$PbCO_3(s) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + CO_2(g) + H_2O(l)$$

$$CaCO_3(s) + H_2SO_4(aq) \longrightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$$

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The insoluble sulphates form a coating on the carbonates, thus preventing further reaction between the carbonates and the acid. Dilute nitric acid reacts with metal carbonates to liberate carbon dioxide, water and soluble metal nitrates.

$$PbCO_3(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

$$CaCO_3(s) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

Activity 9.15



Aim: To investigate the action of dilute mineral acids on metal carbonates

Requirements: Test tube with a side arm, test tube, teat pipette, beaker, measuring cylinder, spatula, stand and clamp, delivery tube, rubber bung to fit the test tube with a side arm, 2 M hydrochloric acid, carbonates of copper, zinc, and lead

Procedure

1. Put one spatulaful of copper(II) carbonate into the test tube with a side arm.
2. Draw some hydrochloric acid into the teat pipette and insert in the mouth of the test tube as shown in Figure 9.5.

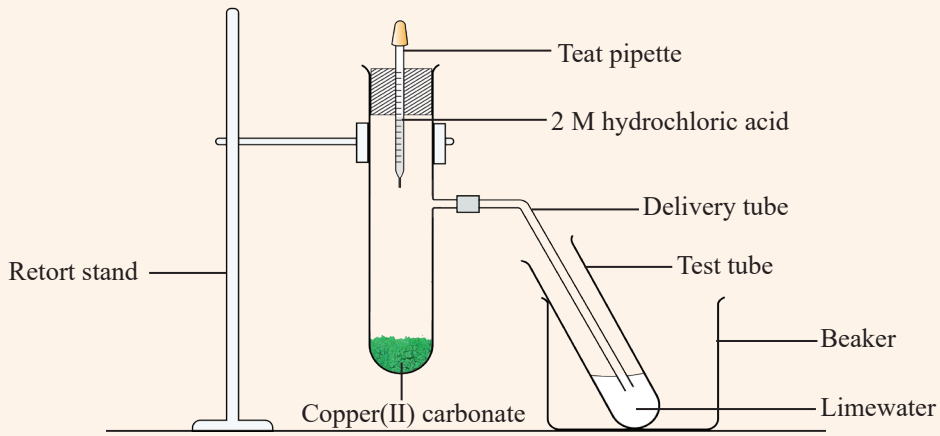


Figure 9.5: *Set-up for investigating the action of dilute hydrochloric acid on copper(II) carbonate*

3. Transfer 2 cm^3 of 2 M hydrochloric acid from the teat pipette into the copper(II) carbonate. Record your observations.
4. Repeat Steps 1–3 using other metal carbonates, one at a time.

Questions

1. Which of the reaction(s) in the experiment gave (or was expected to give) the smallest amount of the gas? Explain with a chemical equation.

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2. Write the balanced chemical equations for the reactions between the carbonates of copper(II) and zinc with dilute hydrochloric acid.
3. Explain how you can confirm the identity of the gas evolved in the performed experiments.
4. Use balanced chemical equations to show the reaction products that will be formed when dilute hydrochloric acid is replaced with dilute nitric acid in the experiment.

Test for metal carbonates

All metal carbonates react with dilute nitric acid to give metal nitrates, water and carbon dioxide gas. Carbon dioxide gas forms a white precipitate with limewater. This is used as a confirmatory test for metal carbonates.

Metal hydrogencarbonates

Metal hydrogencarbonates are formed when only one of the two hydrogen atoms in the carbonic acid is replaced with a metal ion. Most metal hydrogencarbonates are soluble in water.

Preparation of metal hydrogencarbonates

Metal hydrogencarbonates can be prepared by a two-step reaction of carbon dioxide and metal hydroxides. For example, when excess carbon dioxide gas is passed through concentrated sodium hydroxide solution, sodium hydrogencarbonate is formed. The reaction takes place in two stages. In the first stage, sodium carbonate is formed.

$$2NaOH(aq) + CO_2(g) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$

When excess carbon dioxide is passed through the solution, sodium hydrogencarbonate is formed.

$$Na_2CO_3(aq) + CO_2(g) + H_2O(l) \longrightarrow 2NaHCO_3(aq)$$

Activity 9.16



Aim: To prepare sodium hydrogencarbonate

Requirements: Beaker, filter funnel, filter paper, measuring cylinder, carbon dioxide generator, concentrated sodium hydroxide solution, sodium carbonate solution, and distilled water

Procedure

1. Transfer about 100 cm^3 of concentrated sodium hydroxide solution into a beaker.
2. Pass some carbon dioxide gas which is free from any acid through the concentrated sodium hydroxide, until excess.

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3. Repeat Steps 1–2 using sodium carbonate solution in the place of concentrated sodium hydroxide solution.

Questions

1. Why is carbon dioxide passed in excess during the preparation of hydrogencarbonates?
2. Based on your observations, comment on the products formed in the experiment.

Activity 9.17



Aim: To prepare calcium hydrogencarbonate

Requirements: Beaker, carbon dioxide gas generator, bent glass tube, test tubes, test tube holder, and aqueous calcium hydroxide (limewater)

Procedure

1. Transfer about 100 cm^3 of calcium hydroxide solution into a beaker.
2. Pass some carbon dioxide gas which is free from any acid through the calcium hydroxide solution until excess and record your observations.

Questions

1. With the aid of chemical equations, explain the reaction(s) occurring in this experiment.
2. Explain what will happen if a small amount of carbon dioxide is passed through limewater.
3. What will happen if calcium carbonate is used instead of aqueous calcium hydroxide?
4. Magnesium hydrogencarbonate can also be prepared using the same procedure. Write the chemical equation for its preparation.

Chemical properties of hydrogencarbonates

The following are some of the properties of hydrogencarbonates:

Action of heat on hydrogencarbonates

Hydrogencarbonates, whether in solution or in solid form, decompose to give carbonates, water and carbon dioxide when heated. For example calcium hydrogencarbonate decomposes upon heating to produce calcium carbonate, water, and carbon dioxide. Sodium hydrogencarbonate decomposes in the same way.

$$2NaHCO_3(s) \xrightarrow{\text{heat}} Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

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Action of acids on hydrogencarbonates

Hydrogencarbonates react with dilute acids to give a salt, water, and carbon dioxide. For example, when dilute hydrochloric acid is added to sodium hydrogencarbonate, some effervescence are observed. This is due to the production of carbon dioxide which forms a white precipitate with limewater. A colourless solution of sodium chloride is also formed.

$$NaHCO_3(s) + HCl(aq) \longrightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

The hydrogencarbonates of potassium, calcium, and magnesium will also react with hydrochloric acid in a similar way to give a metal chloride, carbon dioxide, and water.

Activity 9.18



Aim: To investigate the action of dilute acids on hydrogencarbonates

Requirements: Test tube with a side arm, a test tube, teat pipette, beaker, measuring cylinder, spatula, retort stand and clamp, delivery tube, small rubber bung with a hole, 2 M hydrochloric acid, solid potassium hydrogencarbonate, and calcium hydrogencarbonate solution

Procedure

1. Set the experiment as shown in Figure 9.5.
2. Put a spatulaful of potassium hydrogencarbonate into the test tube with a side arm.
3. Fill the teat pipette with 2 M hydrochloric acid.
4. Add about 2 cm^3 of the 2 M hydrochloric acid from the teat pipette into the potassium hydrogencarbonate. Record your observations.
5. Repeat Steps 2–4 using about 5 cm^3 of calcium hydrogencarbonate solution in the place of patassium hydrogencarbonate.

Questions

1. (a) Which gas is evolved when the metal hydrogencarbonates react with dilute hydrochloric acid?
   (b) Explain how you will experimentally confirm the identity of the gas.
2. Write the balanced chemical equations for the reactions occurring in the performed experiments.

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3. With the aid of a balanced chemical equation, explain what will happen if a magnesium hydrogencarbonate is used in the place of the calcium hydrogencarbonate.

Test for soluble carbonates and hydrogencarbonates

Metal carbonates and hydrogencarbonates can be distinguished by the use of magnesium sulphate. When a metal carbonate is reacted with magnesium sulphate, a white precipitate of magnesium carbonate is formed. For example, potassium carbonate reacts with magnesium sulphate to give magnesium carbonate and potassium sulphate as shown in the following equation:

$$K_2CO_3(aq) + MgSO_4(aq) \longrightarrow MgCO_3(s) + K_2SO_4(aq)$$

On the other hand, the reaction between a metal hydrogencarbonate and magnesium sulphate does not produce any precipitate. This is because metal hydrogencarbonates and magnesium sulphate react to form magnesium hydrogencarbonate which is soluble in water.

$$2NaHCO_3(s) + MgSO_4(aq) \longrightarrow Mg(HCO_3)_2(aq) + Na_2SO_4(aq)$$

Generally, soluble carbonates form white precipitates with magnesium sulphate solution, while hydrogencarbonates do not. This test is used to distinguish soluble carbonates from hydrogencarbonates.

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Activity 9.19



Aim: To test for soluble carbonates and hydrogencarbonates

Requirements: Beakers, measuring cylinder, 2 M magnesium sulphate, 2 M sodium carbonate solution, and 2 M calcium hydrogencarbonate solution

Procedure

1. Put about 20 cm^3 of the sodium carbonate solution in a beaker.
2. Add about 20 cm^3 of magnesium sulphate solution into the beaker containing sodium carbonate solution. Record your observations.
3. Repeat Steps 1–2 using about 20 cm^3 of calcium hydrogencarbonate solution in the place of the sodium carbonate solution.

Questions

1. With the aid of chemical equations, explain the results observed.
2. What will happen if the products in Step 3 are heated? Give a balanced chemical equation.

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Uses of carbonates and hydrogencarbonates

Sodium carbonate and sodium hydrogencarbonate are used in different areas including the following:

Softening hard water: Sodium carbonate is used to remove both temporary and permanent hardness of water. It reacts with the hydrogencarbonate and sulphate ions to form insoluble carbonates. The insoluble carbonates are then filtered out.

$$Ca(HCO_3)_2(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + 2NaHCO_3(aq)$$

$$MgSO_4(aq) + Na_2CO_3(aq) \longrightarrow MgCO_3(s) + Na_2SO_4(aq)$$

Manufacture of glass: Glass is a mixture of sodium and calcium silicates. It is made by heating together sand, calcium carbonate, and sodium carbonate at a temperature between 1300 °C and 1400 °C. During the process, carbon dioxide is given off, leaving a mixture of sodium silicate and calcium silicate.

$$CaCO_3(s) + SiO_2(s) \xrightarrow{\text{heat}} CaSiO_3(s) + CO_2(g)$$

$$Na_2CO_3(s) + SiO_2(s) \xrightarrow{\text{heat}} Na_2SiO_3(s) + CO_2(g)$$

Qualitative and quantitative analyses: Sodium carbonate is used in the qualitative analysis involving carbonates because it is soluble in water. Sodium carbonate is used in quantitative analysis because the moisture it absorbs from the air can readily be removed by heating without decomposing the salt. Pure and dry sodium carbonate can be obtained in this manner for use as a primary standard and in chemical analysis.

Manufacture of sodium hydroxide: Slaked lime (calcium hydroxide) is added to sodium carbonate in the preparation of sodium hydroxide. The precipitates of calcium carbonate formed are filtered off leaving sodium hydroxide in solution.

$$Na_2CO_3(aq) + Ca(OH)_2(s) \longrightarrow CaCO_3(s) + 2NaOH(aq)$$

Sodium hydroxide crystals will be obtained from the sodium hydroxide filtrate after evaporation of the water to dryness.

Manufacture of water glass: When sodium carbonate is heated together with silicon dioxide (SiO_2), sodium silicate (Na_2SiO_3) and carbon dioxide are produced.

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$$Na_2CO_3(s) + SiO_2(s) \xrightarrow{heat} Na_2SiO_3(s) + CO_2(g)$$

A concentrated solution of sodium silicate in water, known as 'water glass' is used as a preservative for eggs. It is also used as an adhesive in paper making and in television tubes.

Baking: Sodium hydrogencarbonate (baking soda) is used in baking industry. When the baking soda is mixed with a dough, the sodium hydrogencarbonate decomposes to give bubbles of carbon dioxide which causes the dough to rise.

$$2NaHCO_3(aq) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

Commercial baking powder is a mixture of baking soda and organic acids such as citric or tartaric acids.

Health salts: Health salts are a mixture of sodium hydrogencarbonate, tartaric acid, and sodium potassium tartrate. Sodium hydrogencarbonate acts as an antacid which neutralises the hydrochloric acid produced in the stomach. Its high solubility enables it to be absorbed very fast in the body, thus providing a quick relief for the symptoms caused by excess acid in the stomach.

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Exercise 9.3

1. Write chemical equations to show the reactions between the following compounds:
   (a) Potassium hydrogencarbonate and zinc chloride.
   (b) Sodium hydrogencarbonate and lead(II) nitrate.
   (c) Sodium hydrogencarbonate and copper(II) sulphate.

2. Write chemical equations to show the effects of heat on the following carbonates:
   (a) Potassium carbonate
   (b) Magnesium carbonate
   (c) Silver carbonate
   (d) Ammonium carbonate

3. What happens when dilute hydrochloric acid is added to powdered lead carbonate in a beaker? Use an equation to explain your answer.

4. Describe how you can distinguish a hydrogencarbonate from a soluble carbonate.

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5. Give five uses of metal carbonates.
6. Describe how you can prepare sodium carbonate in the laboratory.

Chapter summary

1. A compound is a substance formed by the chemical combination of two or more elements.
2. Metals react with other substances, especially non-metals, to form different compounds. The compounds formed include metal oxides, hydroxides, carbonates, and hydrogencarbonates.
3. Metal oxides are compounds of metals and oxygen. The oxides can be prepared by direct and indirect methods. In the direct method, metal oxides are prepared by heating the respective metal in air. The indirect methods involve thermal decomposition of metal carbonates, hydroxides, and nitrates to prepare the respective metal oxides.
4. Metal oxides are classified into two groups, which are: basic metal oxides; those which react with acids to give salt and water as the only products, and amphoteric oxides; those which react with both acids and bases.
5. Metal oxides are used in different fields depending on the type of the oxide. Some of the uses include preparation of calcium carbide, lining of furnaces, formation of slag, preparation of mortar, and manufacture of cement.
6. Metal hydroxides consist of a metal and hydroxide ions, (OH^-). The hydroxides can be prepared either by the direct method in which metals, especially those high in the reactivity series, are reacted with water to form their hydroxides, or by the indirect method in which the aqueous salts of metals are reacted with the aqueous solutions of sodium hydroxide or potassium hydroxide.
7. Metal hydroxides are classified into basic hydroxides and amphoteric hydroxides depending on their reactions towards acids and bases.
8. Metal hydroxides are used in various areas including agriculture, medicine, building and construction, bleaching of pulp, softening of water, manufacture of paints, in qualitative analysis, and in extraction of metals.
9. Both metal carbonates and metal hydrogencarbonates are derived from carbonic acid (H_2CO_3). Metal carbonates are formed when both hydrogen atoms in carbonic acid are replaced by a metal, whereas metal hydrogencarbonates are formed when only one of the two hydrogen atoms in the carbonic acid is replaced with a metal.

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10. Metal carbonates can be classified as soluble or insoluble carbonates based on their solubility in water.

11. Soluble carbonates are prepared in the laboratory by reacting their corresponding alkalis with carbon dioxide. All insoluble metal carbonates can be prepared by precipitation method which involves the reaction between a soluble carbonate and the solution of the metal salt.

12. Metal hydrogencarbonates can be prepared by a two-step reaction of carbon dioxide and metal hydroxides.

13. Metal carbonates and hydrogencarbonates have different uses depending on the type of the carbonate and hydrogencarbonate. Some of the uses include softening hard water, manufacture of glass, manufacture of sodium hydroxide, manufacture of 'water glass', manufacture of medicine, for qualitative and quantitative analyses, and in baking industry.

Revision exercise 9

Choose the correct answer in items 1 and 2.

1. Sodium hydroxide was added to three different test tubes, X, Y, and Z. Blue, green and reddish-brown precipitates were formed, respectively. The metal ions present in the test tubes X, Y, and Z, respectively were likely to be:
   (a) Ca^{2+}, Pb^{2+}, and Al^{3+}
   (b) Cu^{2+}, Fe^{2+}, and Fe^{3+}
   (c) Ag^{+}, Pb^{2+}, and Mg^{2+}
   (d) Pb^{2+}, Mg^{2+}, and Ca^{2+}

2. A sample of a solid compound X was strongly heated in a test tube. No observable changes were noted during the heating. When dilute nitric acid was added to a sample of solid X, some effervescence were observed. The gas produced formed a white precipitate with limewater. Compound X is likely to be:
   (a) zinc nitrate.
   (b) copper(II) sulphate.
   (c) potassium carbonate.
   (d) magnesium chloride.

3. Give the colour of the following oxides when hot, and cold:
   (a) Copper oxide
   (b) Zinc oxide
   (c) Magnesium oxide

4. Write balanced chemical equations for each of the following reactions:
   (a) Sodium peroxide and water
   (b) Sodium peroxide and ice-cold water
   (c) Sodium oxide and water

5. Consider the following metal oxides: calcium oxide, aluminium oxide, potassium oxide, and copper(II) oxide. Write balanced chemical equations to show the reaction of

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each of the metal oxides with the following:
(a) Water
(b) Dilute nitric acid
(c) Sodium hydroxide solution

6. Using balanced chemical equations, explain what happens when dilute hydrochloric acid is added to the following substances:
(a) Calcium oxide
(b) Sodium hydroxide solution

7. With examples, classify each of the following metal hydroxides as either basic or amphoteric:
(a) Sodium hydroxide
(b) Zinc hydroxide
(c) Aluminium hydroxide
(d) Potassium hydroxide
(e) Lead(II) hydroxide
(f) Magnesium hydroxide

8. With the aid of a chemical equation, explain what will happen if excess sodium hydroxide is added to zinc nitrate solution.

9. Name one metal compound that can be used to remove hardness in water. Give the chemical equation for the reaction involved.

10. Explain why the indirect method is not preferred for the laboratory preparation of sodium oxide and potassium oxide.

11. (a) Give chemical equations to show the reaction occurring when each of the following metals are burnt in air:
(i) Sodium
(ii) Magnesium
(iii) Calcium

(b) Explain what happens when the three metals in 11(a) are exposed to air.

12. A chemist observed that, when a small amount of carbon dioxide was blown into limewater, a white precipitate was formed. However, when more carbon dioxide was added, the precipitate dissolved to form a colourless solution. Explain these observations with the aid of chemical equations.

13. A form three student was investigating the reaction of calcium metal with water. When the metal was placed in a beaker with water, effervescence were observed. A white suspension was left after the reaction was complete.

(a) Name the white suspension left after completion of the reaction. Comment on the solubility of the reaction product in water.
(b) Name the gas that evolved.
(c) How can you experimentally confirm the identity of the gas in (b)? Explain.
(d) How can you test the pH of the solution left in the beaker above? Explain.
(e) Write balanced chemical equations for the reactions involved in the investigation.

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14. (a) You are provided with silver carbonate, magnesium carbonate, copper carbonate, and sodium carbonate. Identify the compound which;
    (i) only decomposes on strong heating.
    (ii) decomposes readily when heated to give the corresponding metal oxide and carbon dioxide gas.
    (iii) does not decompose under normal laboratory temperatures.
    (iv) decomposes to produce the corresponding metal, carbon dioxide gas and oxygen gas.
    (b) Write balanced chemical equations to show the decomposition products of the metal carbonate in 14 (a) (i).
    (c) Indicate the conditions under which the metal carbonate in 14 (a) (iii) can be decomposed.

15. During an investigation on the properties of different compounds of metals, a chemist carried out a reaction between the aqueous solutions of potassium carbonate and magnesium sulphate in which a white precipitate was formed. However, when the chemist reacted sodium hydrogencarbonate solution and magnesium sulphate, no precipitate was formed. Explain these observations with the aid of chemical equations.

16. When dilute nitric acid is added to a green solid P, a blue solution Q is formed and a gas R that forms a white precipitate with limewater is also produced. Evaporation of the blue solution to dryness, and then strongly heating it in a crucible forms a black solid S, brown fumes of gas T, and a gas that relights a glowing splint.
    (a) Identify the;
        (i) solids P and S.
        (ii) gases R and T.
    (b) Write a chemical equation for the;
        (i) reaction between solid P and dilute nitric acid.
        (ii) formation of solid S.

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Chapter Ten

Compounds of metals-Part II: nitrates, chlorides, and sulphates

Introduction

In Chapter Nine, you learnt about the preparation, chemical properties, and uses of metal oxides, hydroxides, carbonates, and hydrogencarbonates. In this chapter, you will learn about the preparation, chemical properties, and uses of metal nitrates, chlorides, and sulphates. The competencies developed will enable you to solve various problems related to metal nitrates, chlorides, and sulphates in industries, medicine, agriculture, households, schools, and the environment.

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Metal nitrates

Metal nitrates are derivatives of nitric acid after replacing the hydrogen with a metal or an ammonium radical.

Preparation of metal nitrates

Metal nitrates are soluble in water, thus cannot be prepared by precipitation. Nitric acid is the most common reagent in the preparation of metal nitrates from its reactions with either metals, metal hydroxides, metal carbonates or metal oxides.

Reactions of metals with nitric acid

The reactions of metals with nitric acid depend on the nature of the metals and the concentration of the acid. Nitric acid, unlike other acids, does not give hydrogen gas with metals, except with magnesium and manganese when the acid is very dilute. This is because nitric acid is a strong oxidising agent, so it oxidises hydrogen gas into water and oxide of nitrogen. The following equations represent the reactions of magnesium and manganese with dilute nitric acid.

$$Mg(s) + 2HNO_3(aq) \longrightarrow Mg(NO_3)_2(aq) + H_2(g)$$

$$Mn(s) + 2HNO_{3(aq)} \longrightarrow Mn(NO_3)_2(aq) + H_2(g)$$

When concentrated nitric acid reacts with metals, brown fumes of nitrogen dioxide gas are normally given off. When dilute nitric acid is used, nitrogen monoxide gas is normally produced which immediately reacts with oxygen in the air to produce brown fumes of nitrogen dioxide gas.

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Magnesium reacts with concentrated nitric acid to form a solution of magnesium nitrate nitrogen dioxide gas, and water.

$$Mg(s) + 4HNO_3(l) \longrightarrow Mg(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

When a solution is evaporated and then cooled, white crystals of magnesium nitrate are formed. The solution should not be evaporated to dryness because the nitrates will undergo thermal decomposition.

Copper reacts with concentrated nitric acid in a similar way as magnesium to give a blue solution of copper nitrate, brown fumes of nitrogen dioxide gas, and water.

$$Cu(s) + 4HNO_3(l) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

When moderate concentrated nitric acid is used, nitrogen monoxide gas (NO) is formed instead of nitrogen dioxide gas.

$$3Cu(s) + 8HNO_3(aq) \longrightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(l)$$

Zinc reacts with hot concentrated nitric acid to give zinc nitrate, nitrogen dioxide, and water.

$$Zn(s) + 4HNO_3(l) \longrightarrow Zn(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

The reaction of nitric acid with aluminium stops immediately after starting due to the formation of an oxide layer (Al_2O_3) on the metal's surface. The oxide layer prevents further reaction between the acid and the metal.

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Activity 10.1



Aim: To prepare metal nitrates by the reactions of metals with concentrated nitric acid

Requirements: Beakers, measuring cylinder, evaporating dish, Bunsen burner, tripod stand, wire gauze, glass rod, concentrated nitric acid, copper granules, iron, and aluminium

Caution!

1. This experiment should be carried out in a fume cupboard.
2. Concetrated nitric acid is corrosive; handle it with care.

Procedure

1. Transfer about 25 cm^3 of concentrated nitric acid into a beaker.
2. Add some copper granules little by little, and then stir the mixture using glass rod until the reaction stops.

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3. Decant the mixture and evaporate about 10 cm^3 of the solution in an evaporating dish to saturation.
4. Leave the solution to cool and crystallise.
5. Repeat Steps 1–4 by using iron and aluminium, one at a time in the place of copper granules.
6. Record your observations as shown in the following table:

Metal	Colour of the solution	Colour of the residue

Questions

1. (a) What was the colour of the fumes in Step 2 (if any)?
   (b) What does this colour suggest on the identity of the gas?
2. Give the balanced chemical equations for the reactions involved in the experiments.
3. What did you observe when aluminium was reacted with concentrated nitric acid? Explain.

Reactions of metal hydroxides with dilute nitric acid

Dilute nitric acid reacts with metal hydroxides through a neutralisation reaction to give a salt and water. For example, the acid reacts with potassium hydroxide to form potassium nitrate and water.

$$KOH(aq) + HNO_3(aq) \longrightarrow KNO_3(aq) + H_2O(l)$$

White crystals of potassium nitrate are formed when the solution of potassium nitrate is heated to saturation, then cooled. Similary, white crystals of sodium nitrate are formed when the solution of sodium nitrate is heated to saturation, then cooled.

Activity 10.2



Aim: To prepare a sodium nitrate by reacting sodium hydroxide with dilute nitric acid

Requirements: Beaker, measuring cylinder, evaporating dish, Bunsen burner, tripod stand, wire gauze, glass rod, 2 M sodium hydroxide, and 2 M nitric acid

Procedure

1. Transfer about 25 cm^3 of 2 M sodium hydroxide solution into a beaker.

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2. Add about 25 cm^3 of 2 M nitric acid and stir using a glass rod.
3. Put some of the mixture into an evaporating dish and heat it to saturation.
   Note: Do not overheat.
4. Leave the solution to cool and crystallise.

Questions

1. Write a balanced chemical equation for the reaction in the performed experiment.
2. What is the colour of the solid product formed in Step 4?
3. Name the type of the reaction occurring in the experiment.

Reactions of metal carbonates with dilute nitric acid

Dilute nitric acid reacts with metal carbonates to form a metal nitrate, carbon dioxide, and water. For example, sodium carbonate reacts with dilute nitric acid to form a colourless solution of sodium nitrate, carbon dioxide gas, and water.

$$Na_2CO_3(s) + 2HNO_3(aq) \longrightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$$

White crystals of sodium nitrate are formed when the solution is heated to saturation, then cooled.

Copper(II) carbonate reacts in a similar way to produce a blue solution of copper(II) nitrate.

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Activity 10.3



Aim: To prepare nitrates of metals by reacting dilute nitric acid with metal carbonates

Requirements: Beakers, spatula, glass rod, measuring cylinder, evaporating dish, tripod stand, wire gauze, Bunsen burner, copper(II) carbonate, calcium carbonate, limewater, and 2 M nitric acid

Procedure

1. Transfer about 10 cm^3 of 2 M nitric acid into a beaker.
2. Add some copper(II) carbonate little by little and stir.
3. Test the gas given off by using a drop of limewater on a glass rod.
4. Add more carbonate continuously until no more effervescence is observed.
5. Allow the mixture to settle, then decant the mixture.
6. Evaporate the solution into an evaporating dish to saturation.

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7. Leave the saturated solution to cool.
8. Repeat Steps 1–7 by using calcium carbonate in the place of copper(II) carbonate and record your observations.

Questions

1. Use a balanced chemical equation to explain the importance of Step 3 in the procedure.
2. Write the balanced chemical equations for the reactions occurring in each of the two experiments.
3. Identify the colour of the solutions and crystals formed after saturation in each reaction.
4. What happens if the solution in Step 6 is heated to dryness?

Reactions of metal oxides with dilute nitric acid

Neutralisation reactions occur between dilute nitric acid and metal oxides in which salt and water are formed. For example, zinc oxide forms a colourless solution of zinc nitrate and water.

$$ZnO(s) + 2HNO_3(aq) \longrightarrow Zn(NO_3)_2(aq) + H_2O(l)$$

White zinc nitrate crystals are formed when the colourless solution is heated to saturation. Calcium oxide, magnesium oxide, and copper(II) oxide react in a similar way to form their corresponding metal nitrates. The nitrate solutions of calcium and magnesium are colourless, which upon crystallisation form white crystals. The blue solution of copper(II) nitrate forms blue crystalline solids upon crystallisation.

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Activity 10.4



Aim: To prepare metal nitrates by reacting dilute nitric acid with metal oxides

Requirements: Beakers, spatula, measuring cylinder, Bunsen burner, tripod stand, wire gauze, glass rod, evaporating dish, copper(II) oxide, magnesium oxide, calcium oxide, and 2 M nitric acid

Procedure

1. Transfer about 25 cm^3 of 2 M nitric acid into a beaker.
2. Add copper(II) oxide little by little while stirring using glass rod until a small amount of the oxide remains undissolved in the reaction mixture.

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3. Decant the mixture and heat the solution gently to saturation.
4. Repeat Steps 1–3 by using magnesium oxide, then calcium oxide in the place of copper(II) oxide. Record all the observations as shown in the following table:

Metal nitrate	Colour of the solution	Colour of the crystals

Questions

1. Write the balanced chemical equations for the reactions between dilute nitric acid and the three metal oxides that are involved in the experiment.
2. Explain the reasons behind gentle heating versus strong heating of metal nitrates solutions towards crystallisation.

Chemical properties of metal nitrates

Nitrates behave differently towards heating and in different chemical reactions. Some of their chemical properties are discussed in the following sections:

Effect of heat on nitrates

Metal nitrates decompose when heated to give different products depending on the extent of heating and the reactivity of each metal.

When potassium nitrate is gently heated, it first melts to form a colourless liquid. On strong heating, it slowly decomposes to give potassium nitrite (KNO_2) and oxygen gas.

$$2KNO_3(s) \xrightarrow{\text{heat}} 2KNO_2(s) + O_2(g)$$

Sodium nitrate decomposes in a similar way to give sodium nitrite ($NaNO_2$) and oxygen gas.

$$2NaNO_3(s) \xrightarrow{\text{heat}} 2NaNO_2(s) + O_2(g)$$

When hydrated copper nitrate is heated, it first releases its water of crystallisation, and then dissolves in water to form a blue solution. On further heating, the water evaporates. When most of the water has evaporated, decomposition starts. The nitrate decomposes to form copper(II) oxide, a reddish brown gas (nitrogen dioxide) and oxygen gas.

$$2Cu(NO_3)_2(s) \xrightarrow{\text{heat}} 2CuO(s) + 4NO_2(g) + O_2(g)$$

Lead(II) nitrate decomposes with a characteristic cracking sound to form a yellow lead(II) oxide (when hot), brown nitrogen dioxide, and colourless oxygen gases. The cracking sound is produced as a result of the air trapped inside the crystals expanding and breaking the crystals.

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$$2Pb(NO_3)_2(s) \xrightarrow{\text{heat}} 2PbO(s) + 4NO_2(g) + O_2(g)$$

Zinc nitrate decomposes to form zinc oxide, nitrogen dioxide, and oxygen.

$$2Zn(NO_3)_2(s) \xrightarrow{\text{heat}} 2ZnO(s) + 4NO_2(g) + O_2(g)$$

Nitrates of metals that are low in the reactivity series such as silver nitrate and mercury nitrate decompose to give the metal, nitrogen dioxide gas, and oxygen gas. For example, silver nitrate decomposes to give silver, nitrogen dioxide gas, and oxygen gas as in the following equation:

$$2AgNO_3(aq) \xrightarrow{\text{heat}} 2Ag(s) + 2NO_2(g) + O_2(g)$$

Table 10.1 shows a summary of the effects of heat on different metal nitrates.

Table 10.1: *Effects of heat on metal nitrates*

Nitrates	Products
Potassium nitrate Sodium nitrate	Metal nitrite + oxygen gas
Calcium nitrate Magnesium nitrate Aluminium nitrate Zinc nitrate Iron nitrate Lead nitrate Copper nitrate	Metal oxide + nitrogen dioxide gas + oxygen gas
Silver nitrate Mercury nitrate	Metal + nitrogen dioxide gas + oxygen gas

Activity 10.5



Aim: To investigate the effects of heat on different metal nitrates

Requirements: Test tubes, test tube holder, test tube rack, Bunsen burner, wooden splint, nitrates of sodium, calcium, and silver

Procedure

1. Clean three test tubes and label each with the name of the metal nitrate.
2. Put a spatulaful of each nitrate into its respective test tube.
3. Heat the metal nitrates one at a time while observing the changes in the colour of the nitrate and that of the gases produced.
4. Test any gas produced by using a glowing wooden splint.
5. Note the colour of the residue both when hot and cold.
6. Record all the observations.

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Questions

1. Based on the observations, which metal nitrate(s) evolve(s) only oxygen gas when heated? Explain.
2. Based on the observations, which metal nitrate(s) evolve(s) both nitrogen dioxide gas and oxygen gas when heated? Explain.
3. Write the balanced chemical equations for the decomposition of the three metal nitrates that are involved in the experiment.

Tests for nitrates

Various tests are used to identify metal nitrates in a solution or in solid form. These include the action of heat on nitrates, brown ring test, and copper turnings test.

With exceptions of sodium and potassium nitrates, all metal nitrates produce reddish brown fumes of nitrogen dioxide gas when heated.

The brown ring test is performed by adding freshly prepared iron(II) sulphate to the nitrate solution, followed by dropwise addition of concentrated sulphuric acid. When concentrated sulphuric acid is poured down along the side of the test tube, it moves down and settles at the bottom of the test tube. This is because the acid is denser than the mixture of the test solution and iron(II) sulphate. If the test solution is a nitrate, a brown ring forms between the layer of concentrated sulphuric acid, and the mixture of the test solution and iron(II) sulphate. All nitrates form a brown ring. This test is therefore known as the brown ring test for nitrates.

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The brown ring complex is a combination of iron(II) sulphate and nitrogen monoxide gas.

$$FeSO_4(aq) + NO(g) \rightleftharpoons FeSO_4{\cdot}NO(aq)$$

In the copper turnings test, the nitrate solution is heated along with copper turnings and concentrated sulphuric acid. The production of brown fumes of nitrogen dioxide gas which turns moist blue litmus paper red indicates the presence of a metal nitrate. Basically, sulphuric acid reacts with the nitrate ion to give nitric acid. The nitric acid then reacts with the copper turnings to form copper(II) nitrate, brown fumes of nitrogen dioxide gas, and water.

$$H^+(aq) + NO_3^-(aq) \longrightarrow HNO_3(aq)$$

$$4HNO_3(aq) + Cu(s) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

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Activity 10.6



Aim: To identify a metal salt in an unknown solution

Requirements: Test tube, teat pipette or dropper, measuring cylinder, iron(II) sulphate solution, concentrated sulphuric acid, blue litmus paper, distilled water, and test solution (unknown salt solution).

Note: Unknown salt solution should be prepared by the laboratory technician or teacher.

Caution! Concentrated sulphuric acid is very corrosive and produces a lot of heat when it comes into contact with water. Handle it with care.

Procedure

1. Transfer about 2 cm^3 of a test solution into a test tube.
2. Add about 2 cm^3 of freshly prepared iron(II) sulphate into the test tube and shake well to mix the two solutions.
3. Tilt the test tube containing the mixture and add about 2 cm^3 of concentrated sulphuric acid. Pour the acid carefully by using a teat pipette or dropper down the side of the test tube without shaking the tube. Record your observations.
4. Shake the mixture carefully, and if there is any gas evolved, test with a moist blue litmus paper. Record your observations.

Question

Based on the observations, identify the type of the metal salt provided. Explain the answer with the aid of chemical equations.

Activity 10.7



Aim: To identify a metal salt in an unknown solid sample

Requirements: Test tubes, Bunsen burner, dropper, tripod stand, wire gauze, concentrated sulphuric acid, copper turnings, blue litmus paper, distilled water, and unknown solid sample.

Procedure

1. Put 2 to 3 spatulafuls of the unknown solid into a test tube.
2. Add about 5 cm^3 of distilled water and shake the mixture to dissolve the solid.
3. Add some copper turnings into the resulting solution followed by 3 drops of concentrated sulphuric acid.

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4. Warm the mixture and test any gas evolved with a moist blue litmus paper. Record all the observations.

Question

Based on the observations, identify the type of the metal salt provided. Explain the answer with the aid of chemical equations.

Uses of metal nitrates

Nitrates have a number of uses in different areas, including the following:

Agriculture: Nitrogenous fertilisers manufactured from nitric acid are mainly nitrates. These include ammonium nitrate, potassium nitrate, and calcium nitrate.

Photography: Silver nitrate is used in photography. A solution of silver nitrate reacts with potassium bromide to form silver bromide.

$$AgNO_3(aq) + KBr(aq) \longrightarrow KNO_3(aq) + AgBr(s)$$

Photographic films and plates contain an emulsion in the form of silver bromide which decomposes to silver metal on exposure to light.

$$2AgBr(s) \xrightarrow{\text{light}} 2Ag(s) + Br_2(g)$$

The photochemical change involving silver bromide is the key reaction in black and white photographs.

Weapons: Gunpowder is a mixture of potassium nitrate, charcoal, and sulphur. Ammonium nitrate is also used in making explosives and blasting agents which are used in mines and quarries.

Food preservation: Nitrates and nitrites are used in curing meat and fish. They kill bacteria and also produce a characteristic flavour as well as give meat a pink or red colour. Sodium nitrate or potassium nitrate is used as a source of nitrite (NO_2^-). The nitrite breaks down in the meat into nitric oxide (NO) which helps to prevent oxidation.

Antiseptics: Antiseptics are chemical agents that are used on the skin and mucous membranes to kill germs. Silver compounds such as silver nitrate and silver sulphadiazine are used to prevent the infections of burns, some eye infections, and to destroy warts.

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Exercise 10.1

1. Give the methods that are used to prepare metal nitrates in the laboratory.
2. Use balanced chemical equations to show the decomposition of the following metal nitrates by heat:
   (a) Potassium nitrate
   (b) Aluminium nitrate
   (c) Silver nitrate
   (d) Copper nitrate
3. Describe one chemical test that can be used to identify a metal nitrate in an unknown solution.
4. Give three uses of metal nitrates.

Metal chlorides

Chlorides are salts that are derived from hydrochloric acid.

Preparation of metal chlorides

The preparation methods of metal chlorides depend on whether a particular metal chloride is soluble or insoluble in water. The methods available include:

(a) Precipitation
(b) Direct method, where a metal reacts with dry chlorine gas
(c) Reaction of hydrochloric acid with a metal hydroxide
(d) Reaction of hydrochloric acid with a metal oxide
(e) Reaction of hydrochloric acid with a metal carbonate
(f) Reaction of hydrochloric acid with a metal

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Precipitation method

The precipitation method is used to prepare insoluble chlorides. Silver chloride and lead chloride are insoluble in water. The two chlorides are prepared by using the precipitation method. Silver nitrate solution precipitates silver chloride when it reacts with dilute hydrochloric acid.

$$AgNO_3(aq) + HCl(aq) \longrightarrow HNO_3(aq) + AgCl(s)$$

Silver chloride is a white solid. When exposed to light, it changes from white to grey colour. This is due to the conversion of silver chloride into silver and chlorine. This reaction only takes place in the presence of light.

$$2AgCl(s) \xrightarrow{\text{light}} 2Ag(s) + Cl_2(g)$$

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Lead(II) chloride on the other hand is precipitated out as a white crystalline solid when dilute hydrochloric acid reacts with lead(II) nitrate.

$$Pb(NO_3)_2(aq) + 2HCl(aq) \longrightarrow 2HNO_3(aq) + PbCl_2(s)$$

Activity 10.8



Aim: To prepare insoluble metal chlorides by precipitation

Requirements: Beakers, filter funnel, measuring cylinder, stirring rod, filter paper, watch glass, 0.05 M silver nitrate, 0.05 M lead(II) nitrate, dilute hydrochloric acid, and distilled water

Procedure

1. Transfer about 10 cm^3 of 0.05 M silver nitrate into a beaker.
2. Add about 10 cm^3 of the dilute hydrochloric acid in the beaker.
3. Stir the mixture using a stirring rod and filter.
4. Wash the filtrate with distilled water.
5. Press the residue between two filter papers, or place on the watch glass to dry.
6. Repeat Steps 1–5 with 0.05 M lead(II) nitrate in the place of silver nitrate.

Questions

1. What is the colour of the product formed when lead(II) nitrate reacts with dilute hydrochloric acid?
2. Write the balanced chemical equations for the reactions between dilute hydrochloric acid and the metal nitrates in the experiment.

Preparation of metal chlorides by direct method

In this method, metal chlorides are prepared by the reactions of metals and chlorine gas. For example, iron(III) chloride, a reddish brown solid is prepared by passing dry chlorine gas over heated iron. The iron(III) chloride sublimes on heating and is collected in a conical flask as shown in Figure 10.1. During the preparation, the reaction initially forms iron(II) chloride. The equation for this reaction is as follows:

$$Fe(s) + Cl_2(g) \xrightarrow{\text{heat}} FeCl_2(s)$$

Because chlorine gas is a strong oxidising agent, it oxidises iron(II) chloride to iron(III) chloride.

$$2FeCl_2(s) + Cl_2(g) \xrightarrow{\text{heat}} 2FeCl_3(s)$$

Iron(III) chloride is a deliquescent substance as when it is exposed to air, it absorbs water vapour to form a solution. Aluminium chloride and zinc chloride are prepared by using the same method.

$$2Al(s) + 3Cl_2(g) \xrightarrow{\text{heat}} 2AlCl_3(s)$$

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Activity 10.9

Aim: To prepare iron(III) chloride by direct method

Requirements: Chlorine gas generator, U-tube fitted with rubber bungs, combustion tube, two test tubes, conical flask, Bunsen burner, three stands and clamps, guard tube, two rubber bungs to fit the combustion tube, anhydrous calcium chloride, a rust free coil wire of iron, and zinc granules

Caution! Chlorine gas is highly poisonous. This experiment should be carried out in a fume chamber.

Procedure

1. Set the apparatus as shown in Figure 10.1.

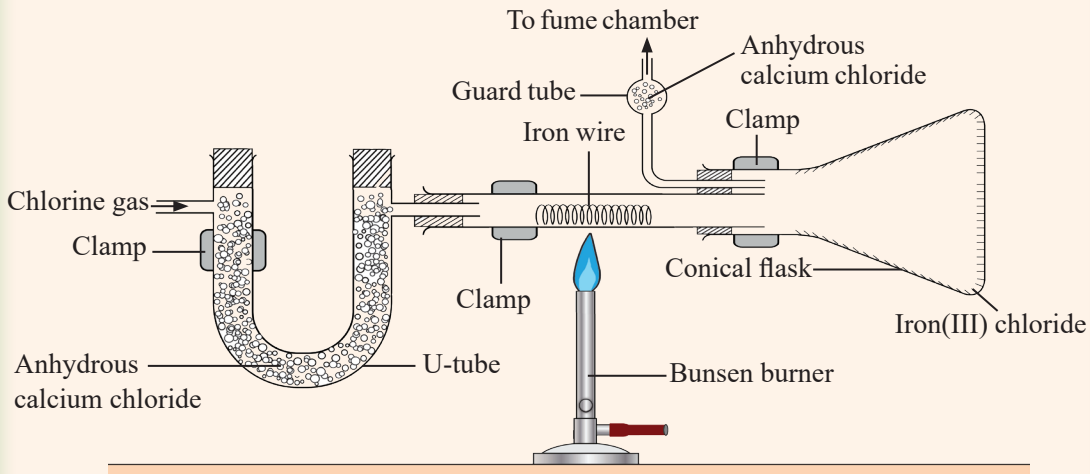


Figure 10.1*: Set-up for the preparation of iron(III) chloride by direct method*

2. Pass some chlorine gas from the chlorine generator through the U-tube containing anhydrous calcium chloride, and then over the coil of a hot rust-free iron wire as shown in the figure.
3. Stop heating the wire as soon as it becomes red-hot. Heat the combustion tube gently once the reaction stops. Record all the observations.
4. Repeat Steps 1–3 by using zinc granules in the place of the iron wire.

Note: The outlet guard tube helps to reduce the pressure in the apparatus, otherwise there would be an explosion. It also allows excess chlorine gas to be passed out. The anhydrous calcium chloride in the guard tube prevents moisture from getting into the conical flask.

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Question

With the aid of a balanced chemical equation, identify the products that are formed when zinc reacts with chlorine gas.

Reactions of hydrochloric acid with a metal hydroxide

Metal hydroxides and dilute hydrochloric acid undergo a neutralisation reaction to produce metal chlorides and water. For example, sodium hydroxide reacts with dilute hydrochloric acid to form sodium chloride and water.

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

Potassium hydroxide and calcium hydroxide are also neutralised by dilute hydrochloric acid to form their corresponding metal chlorides which are soluble in water. Sodium chloride, potassium chloride, and calcium chloride appear as white crystalline solids.

Activity 10.10



Aim: To prepare metal chlorides by reacting metal hydroxide with dilute hydrochloric acid

Requirements: Beakers, measuring cylinders, spatula, evaporating dish, Bunsen burner, tripod stand, wire gauze, glass rod, 2 M potassium hydroxide, calcium hydroxide solids, and 2 M hydrochloric acid

Procedure

1. Transfer about 10 cm^3 of 2 M potassium hydroxide into a beaker.
2. Add about 10 cm^3 of 2 M hydrochloric acid and stir the mixture by using a glass rod.
3. Put the mixture in an evaporating dish and boil to saturation. Care needs to be taken so as not to evaporate to dryness.
4. Leave the saturated mixture to cool and crystallise.
5. Repeat Steps 1–4 with a spatulaful of calcium hydroxide in the place of 2 M potassium hydroxide. Record the observations.

Questions

1. Identify the colours of the solutions and crystals that are formed in the two experiments.
2. Write the balanced chemical equations for the two reactions.
3. Name the type of the reaction that occurred in the experiments.

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Reaction of hydrochloric acid with a metal oxide

The reaction between a metal oxide and dilute hydrochloric acid is also a neutralisation reaction. Consider the reaction between copper(II) oxide and dilute hydrochloric acid, the oxide reacts with the acid to form a green solution of copper(II) chloride.

$$CuO(s) + 2HCl(aq) \longrightarrow CuCl_2(aq) + H_2O(l)$$

The solution of copper(II) chloride crystallises to form green crystals of copper(II) chloride ($CuCl_2$). Iron(III) oxide, which is reddish-brown in colour, reacts with dilute hydrochloric acid to give a brown solution of iron(III) chloride which crystallises to dark green crystals.

$$Fe_2O_3(s) + 6HCl(aq) \longrightarrow 2FeCl_3(aq) + 3H_2O(l)$$

Zinc oxide and calcium oxide show similar reactions with dilute hydrochloric acid.

When dilute hydrochloric acid is added to lead(II) oxide, the reaction starts then stops after a short time due to the formation of insoluble lead(II) chloride which forms a coating on the oxide, hence preventing any further acid action. The reaction between hydrochloric acid and lead oxide cannot therefore be used to prepare lead chloride.

$$PbO(s) + 2HCl(aq) \longrightarrow PbCl_2(s) + H_2O(l)$$

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Activity 10.11



Aim: To prepare metal chlorides by reacting metal oxides with hydrochloric acid

Requirements: Beakers, glass rod, spatula, measuring cylinders, evaporating dish, Bunsen burner, tripod stand, wire gauze, 2 M hydrochloric acid, oxides of magnesium, zinc, calcium, and lead

Procedure

1. Transfer about 20 cm^3 of 2 M hydrochloric acid into a beaker.
2. Add magnesium oxide to the acid little by little and stir by using a glass rod.
3. Continue adding magnesium oxide and stir until some of the magnesium oxide is left undissolved in the reaction mixture.
4. Decant the mixture into an evaporating dish.
5. Heat the solution to saturation, then leave it to cool.
6. Repeat Steps 1–5 by using the oxides of lead, zinc, and calcium. Record all the observations.

Questions

1. Identify the colours of the solutions and crystals that are formed in the experiments.

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2. What did you observe when lead oxide reacted with dilute hydrochloric acid? Explain.
3. Write the balanced chemical equations for the reactions that occurred in the experiments.
4. Name the type of the reaction that is involved when the metal oxides in the experiment react with dilute hydrochloric acid.

Reactions of hydrochloric acid with metal carbonates

Dilute hydrochloric acid reacts with different metal carbonates to give metal chloride, water, and carbon dioxide gas. For example, sodium carbonate forms a colourless solution of sodium chloride when treated with dilute hydrochloric acid.

$$Na_2CO_3(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

Copper(II) carbonate and calcium carbonate exhibit a similar reaction with dilute hydrochloric acid.

Lead(II) carbonate cannot be used to prepare lead(II) chloride. Lead(II) chloride is insoluble and would form a coating on the carbonate, preventing any further contact of the oxide with the acid.

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Activity 10.12



Aim: To prepare metal chlorides by reacting metal carbonates with dilute hydrochloric acid

Requirements: Beakers, evaporating dish, measuring cylinders, spatula, glass rod, Bunsen burner, tripod stand, wire gauze, calcium carbonate, copper(II) carbonate, 2 M hydrochloric acid and limewater

Procedure

1. Transfer about 20 cm^3 of 2 M hydrochloric acid into a beaker.
2. Add some calcium carbonate little by little until a small amount of it remains unreacted in the reaction mixture.
3. Test the gas produced by using a drop of limewater on a glass rod.
4. Decant the mixture into an evaporating dish and heat the solution to saturation. Cool the mixture to crystallise.
5. Repeat Steps 1– 4 by using copper(II) carbonate. Record the observations as shown in the following table:

Metal chloride	Colour of the solution	Colour of the crystals

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Questions

1. Which gas evolved in the reaction for the experiments perfomed? Explain.
2. Write the balanced chemical equations for the reactions between the two metal carbonates and dilute hydrochloric acid used in the experiment.
3. Give the chemical equation for the reaction between the gas evolved and limewater.

Reactions of hydrochloric acid with metals

Metal chlorides can also be prepared by the reactions of metals and dilute hydrochloric acid. For example, magnesium granules react with dilute hydrochloric acid to form a solution of magnesium chloride which crystallises to form white crystals of magnesium chloride.

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

Zinc granules react in a similar way to give a colourless crystalline solid which readily dissolves in water.

There is no reaction when copper is reacted with hydrochloric acid. Copper is lower than hydrogen in the reactivity series and cannot displace hydrogen from hydrochloric acid to produce hydrogen gas. Sodium cannot be used because the reaction would be too vigorous and difficult to control. There is also a possibility of an explosion if hydrogen gas catches fire.

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Activity 10.13



Aim: To prepare metal chlorides by reacting metals with dilute hydrochloric acid

Requirements: Beakers, measuring cylinders, evaporating dish, glass rod, filter funnel, Bunsen burner, tripod stand, wire gauze, filter paper, 2 M hydrochloric acid, zinc granules, calcium metal, copper powder or granules, and silver metal

Procedure

1. Transfer about 20 cm^3 of 2 M hydrochloric acid into a beaker.
2. Add zinc granules, little by little to the acid and stir with a glass rod.
3. Continue adding the granules until the effervescence stops and some granules remain unreacted in the reaction mixture. Test the gas evolved with a lighted splint.
4. Filter the mixture and put the filtrate in an evaporating dish.
5. Heat the filtrate to saturation, then leave to cool and crystallise.

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6. Repeat Steps 1–5 by using copper, calcium, and silver metals one at a time. Record the observations.

Questions

1. Based on the observations, comment on the reactivities of the metals with dilute hydrochloric acid.
2. What are the colours of the crystals formed (if any) upon reacting the metals with dilute hydrochloric acid? Write the balanced chemical equations for the reactions.
3. How can you confirm the identity of the evolved gas experimentally? Explain.

Chemical properties of metal chlorides

Metal chlorides respond differently to heat and reactions with other chemical compounds. Some of the chemical properties of metal chlorides are discussed in the following sections:

Effects of heat on metal chlorides

With the exception of ammonium chloride, other chlorides do not decompose on heating. Ammonium chloride decomposes on heating to form ammonia gas and hydrogen chloride gas.

$$NH_4Cl(s) \overset{\text{heat}}{\rightleftharpoons} NH_3(g) + HCl(g)$$

However, hydrated chlorides undergo dehydration when heated. For example, hydrated copper chloride gives up its water of crystallisation when gently heated.

$$CuCl_2 \cdot 2H_2O(s) \xrightarrow{\text{heat}} CuCl_2(s) + 2H_2O(g)$$

The water of crystallisation is observed as a colourless liquid at the cooler parts of the reaction vessel. The liquid turns cobalt chloride paper pink and anhydrous copper(II) sulphate from white to blue confirming that the liquid is water.

When hydrated magnesium chloride is gently heated, it is partially hydrolysed to form a basic salt, hydrogen chloride gas, and water vapour.

$$MgCl_2 \cdot 6H_2O(s) \xrightarrow{\text{heat}} Mg(OH)Cl(s) + HCl(g) + 5H_2O(g)$$

When strongly heated, the basic salt formed decomposes to form magnesium oxide and hydrogen chloride gas.

$$Mg(OH)Cl(s) \xrightarrow{\text{heat}} MgO(s) + HCl(g)$$

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Activity 10.14

Aim: To investigate the effects of heat on hydrated copper(II) chloride

Requirements: Test tubes, Bunsen burner, retort stand and clamp, spatula, cobalt chloride paper, hydrated copper(II) chloride crystals, distilled water, calcium chloride, anhydrous copper(II) sulphate, and phenolphthalein indicator

Procedure

1. Transfer two spatulafuls of hydrated copper(II) chloride crystals into a test tube.
2. Set the apparatus as shown in Figure 10.2.
3. Heat the crystals gently, and stop heating when some liquid collects in the cooler parts of the test tube.
4. Test any liquid produced by using cobalt chloride paper, then with anhydrous copper(II) sulphate. Record the observations.
5. Allow the residue to cool.

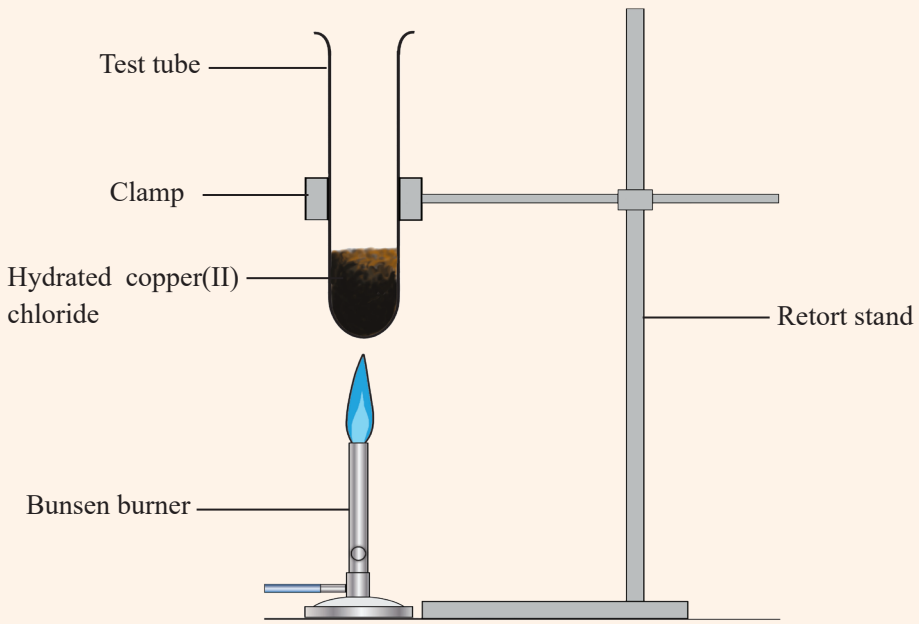


Figure 10.2: *Set-up for the effect of heat on hydrated copper(II) chloride*

Questions

1. Explain the importance of using cobalt chloride paper, and anhydrous copper(II) sulphate in the experiment.
2. Use a balanced chemical equation to explain the reaction taking place in Step 3.

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Test for chlorides

It is possible to identify chlorides both in solid form and in solution by simple chemical reactions.

Test for solid metal chlorides

Solid metal chlorides are identified by a reaction with concentrated sulphuric acid. When concentrated sulphuric acid is added to a solid chloride, hydrogen chloride gas is given off. Consider the reaction between potassium chloride and concentrated sulphuric acid shown in the following equation:

$$KCl(s) + H_2SO_4(l) \longrightarrow KHSO_4(aq) + HCl(g)$$

The evolution of hydrogen chloride gas is a confirmatory test that the unknown solid is a chloride. The presence of the gas can be confirmed by using concentrated ammonia gas, acidified silver nitrate solution or a moist blue litmus paper.
Hydrogen chloride gas forms dense white fumes with ammonia gas due to the formation of ammonium chloride.

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

Hydrogen chloride gas also forms white precipitates with acidified silver nitrate due to the formation of silver chloride.

$$HCl(g) + AgNO_3(aq) \longrightarrow AgCl(s) + HNO_3(aq)$$

The hydrogen chloride gas is acidic, thus it turns wet blue litmus paper red.

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Activity 10.15



Aim: To identify a metal salt from an unknown solid

Requirements: Test tube, test tube holder, Bunsen burner, glass rod, measuring cylinder, spatula, blue litmus paper, unknown solid, concentrated sulphuric acid, 0.05 M acidified (using nitric acid), silver nitrate solution, and concentrated aqueous ammonia

Procedure

1. Put a spatulaful of the unknown solid into a test tube.
2. Add 3 cm^3 of concentrated sulphuric acid.
3. Dip a glass rod into concentrated aqueous ammonia and place at the mouth of the test tube.
4. Repeat Step 3 by using acidified silver nitrate solution in the place of aqueous ammonia.
5. Place a wet blue litmus paper at the mouth of the test tube.

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Questions

1. Based on the observations, identify the type of the unkown metal compound provided.
2. Identify the reaction products that are formed when the unkown solid reacts with concentrated sulphuric acid.
3. With the aid of chemical equations, explain the significance of a silver nitrate solution and concentrated aqueous ammonia in the experiment.
4. What was the aim of using the blue litmus paper?

Test for metal chloride solutions

When a few drops of dilute nitric acid are added to an unknown solution of a metal chloride followed by silver nitrate solution, a white precipitate of silver chloride is formed.

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

Silver chloride does not dissolve even when dilute nitric acid is added to it.

Dilute nitric acid is added to prevent precipitation of other silver salts such as a carbonates as shown in the following equation, otherwise they will be mistaken for a chloride:

$$2Ag^+(aq) + CO_3^{2-}(aq) \longrightarrow Ag_2CO_3(s)$$

If a carbonate is present, it will react with the nitric acid to give carbon dioxide and water.

$$2H^+(aq) + CO_3^{2-}(aq) \longrightarrow CO_2(g) + H_2O(l)$$

Chlorides are not affected by dilute nitric acid.

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Activity 10.16



Aim: To identify a metal salt in an unknown solution

Requirements: Beaker, measuring cylinder, dropper, 2 M nitric acid, unknown test solution, and 0.05 M silver nitrate

Procedure

1. Transfer about 5 cm^3 of the test solution into a beaker.
2. Add 3 to 4 drops of 2 M nitric acid and shake.
3. Add 3 drops of 0.05 M silver nitrate in the mixture and then about three more drops of 2 M nitric acid. Record all the observations.

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Questions

1. Identify the anion in the salt solution provided. Explain the answer with the aid of chemical equations (use 'Q' as a representative symbol for the metal).
2. Explain the use of dilute nitric acid in the experiment. Support your answer with a balanced chemical equation.
3. What happens if dilute hydrochloric acid is used instead of dilute nitric acid?

Uses of metal chlorides

The following are some of the areas where metal chlorides are used:

Manufacture of dry batteries: Zinc chloride mixed with ammonium chloride is used in manufacturing dry batteries.

Household uses: Common salt (sodium chloride) is added in food to give it taste.

Food preservation: Canned foods contain sodium chloride, which is in this case used as a preservative to prevent the growth of bacteria.

Petroleum industry: Aluminium chloride acts as a catalyst in many organic reactions whose reactants are derived from petroleum. Aluminium chloride is also used in the manufacture of synthetic lubricating oils.

Spinning industry: Magnesium chloride is used to lubricate cotton threads in the spinning industry.

Dentistry: Magnesium chloride is a constituent of the cement that is used to fill teeth cavities.

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Exercise 10.2

1. Describe any five methods that are used to prepare soluble metal chlorides in the laboratory.
2. Using equations, give one example of a metal chloride that is prepared by using each of the methods described in Question 1.
3. Describe a chemical test that is used to confirm the presence of a metal chloride in an unknown solid. Use balanced chemical equations to support your answer.
4. Explain the uses of metal chlorides.

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Metal sulphates

Sulphates are salts that are derived from sulphuric acid. There are various methods that are used to prepare metal sulphates. These methods are applied depending on whether the sulphate is soluble or insoluble in water. All metal sulphates are soluble in water except the sulphates of barium and lead. Calcium sulphate, on the other hand, is sparingly soluble in water.

Preparation of soluble sulphates

Soluble sulphates can be prepared by reacting a metal, metal carbonate, metal hydroxide or metal oxide with dilute sulphuric acid.

Reactions of dilute sulphuric acid with metals

Metal sulphates are prepared by the reactions of metals and dilute sulphuric acid. For instance, dilute sulphuric acid reacts with magnesium to give a colourless solution of magnesium sulphate and hydrogen gas.

$$Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$$

White crystals of magnesium sulphate are formed when the solution is heated to saturation and then cooled. Other metal sulphates such as zinc sulphate and iron(III) sulphate are prepared by using this method.

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Activity 10.17



Aim: To prepare zinc sulphate by reacting zinc granules with dilute sulphuric acid

Requirements: Beaker, glass rod, evaporating dish, tripod stand, wire gauze, Bunsen burner measuring cylinder, filter papers, filter funnel, 2 M sulphuric acid, and zinc granules

Procedure

1. Transfer about 25 cm^3 of 2 M sulphuric acid into a beaker.
2. Add some zinc granules and stir using a glass rod. Add more granules continuously until no effervescence observed and some zinc granules remain unreacted.
3. Note the smell of the mixture as the reaction proceeds.
4. Filter the mixture and transfer the filtrate into an evaporating dish.
5. Heat the filtrate to saturation, then leave it to cool and crystallise.
6. Dry the crystals between filter papers taking care not to break them.

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Note: Granulated zinc contains impurities which act as catalysts. The impurities cause a choking smell of the gas produced. When pure zinc is used, the reaction is very slow.

Questions

1. What is the colour of the solution that is formed by the reaction of zinc granules and dilute sulphuric acid?
2. What is the colour of the crystals that are formed after crystallisation?
3. Write a balanced chemical equation for the reaction between zinc granules and and dilute sulphuric acid.
4. Explain how you can experimentally confirm the identity of the gas evolved.

Reactions of metal carbonates with dilute sulphuric acid

Metal carbonates react with dilute sulphuric acid to form metal sulphates, carbon dioxide, and water. For example, zinc carbonate reacts with dilute sulphuric acid to give a colourless solution of zinc sulphate, carbon dioxide gas, and water.

$$ZnCO_3(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + CO_2(g) + H_2O(l)$$

Other metal sulphates that are prepared in a similar way are copper(II) sulphate and magnesium sulphate which both appear as white solids.

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Activity 10.18



Aim: To prepare copper(II) sulphate by reacting copper(II) carbonate with dilute sulphuric acid

Requirements: Beaker, tripod stand, wire gauze, glass rod, evaporating dish, spatula, measuring cylinder, Bunsen burner, filter funnel, filter papers, 2 M sulphuric acid, and copper(II) carbonate powder

Procedure

1. Transfer about 25 cm^3 of 2 M sulphuric acid into a beaker.
2. Add some copper(II) carbonate, a little at a time, until the reaction stops with excess copper(II) carbonate in the reaction mixture.
3. Filter the mixture into an evaporating dish.
4. Heat the solution to saturation, then leave it to cool and crystallise.
5. Dry the crystals between filter papers.

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Questions

1. What is the colour of the crystals formed in Step 4?
2. Write a balanced chemical equation for the reaction between copper(II) carbonate and dilute sulphuric acid.
3. Explain how you can confirm the identity of the gas evolved in the reaction in Question 2.

Reactions of metal hydroxides with dilute sulphuric acid

The reactions of metal hydroxides and dilute sulphuric acid produce metal sulphates and water. For example, zinc hydroxide reacts with dilute sulphuric acid in a neutralisation reaction to form zinc sulphate and water.

$$Zn(OH)_2(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + 2H_2O(l)$$

Zinc sulphate solution is colourless, but forms white crystals of zinc sulphate when heated to saturation, then cooled. Other sulphates that can be prepared by using the same method are copper(II) sulphate, magnesium sulphate and aluminium sulphate.

Activity 10.19



Aim: To prepare magnesium sulphate by the reaction of magnesium hydroxide with dilute sulphuric acid

Requirements: Beaker, Bunsen burner, evaporating dish, glass rod, spatula, tripod stand, wire gauze, measuring cylinder, filter papers, filter funnel, magnesium hydroxide, and 2 M sulphuric acid

Procedure

1. Transfer about 25 cm^3 of 2 M sulphuric acid into a beaker.
2. Add magnesium hydroxide, a little at a time, and stir using a glass rod. Continue adding the hydroxide until a small amount of it remains in the reaction mixture.
3. Filter the mixture and put the filtrate in an evaporating dish.
4. Heat the filtrate to saturation, then leave it to cool and crystallise.
5. Dry the crystals between filter papers.

Questions

1. Write a balanced chemical equation for the reaction between magnesium hydroxide and dilute sulphuric acid.
2. What is the colour of the crystals formed in Step 4?
3. What would be the products of the reaction if magnesium hydroxide is replaced by aluminium hydroxide? Give a balanced chemical equation for the reaction.

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Reactions of metal oxides with dilute sulphuric acid

Metal oxides react with dilute sulphuric acid to form metal sulphates and water. For example, magnesium oxide reacts with dilute sulphuric acid to form a colourless solution of magnesium sulphate.

$$MgO(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2O(l)$$

The colourless solution crystallises to form white crystals when heated to saturation then cooled. The reaction between magnesium oxide and dilute sulphuric acid is a neutralisation reaction. Other metal sulphates that can be prepared in a similar way include copper(II) sulphate, calcium sulphate, and aluminium sulphate.

Activity 10.20



Aim: To prepare copper(II) sulphate by reacting copper(II) oxide with dilute sulphuric acid

Requirements: Beaker, measuring cylinder, evaporating dish, tripod stand, wire gauze, Bunsen burner, glass rod, spatula, filter papers, filter funnel, copper(II) oxide, and 2 M sulphuric acid

Procedure

1. Transfer about 25 cm^3 of 2 M sulphuric acid into a beaker.
2. Heat the acid and add copper(II) oxide. Stir the mixture using a glass rod until all the copper(II) oxide dissolves.
3. Continue heating the solution while adding more copper(II) oxide until no further dissolution is observed.
4. Filter the mixture into an evaporating dish.
5. Heat the solution to saturation, then leave it to cool and crystallise.
6. Dry the crystals between filter papers.

Questions

1. What is the colour of the solution that is formed by the reaction of copper(II) oxide and dilute sulphuric acid?
2. What is the colour of the crystals formed in Step 5?
3. Compare the colour of the solution to that of the crystals. Comment on your comparison.
4. Write a balanced chemical equation for the reaction between copper(II) oxide and dilute sulphuric acid.

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Preparation of insoluble sulphates

Barium sulphate and lead sulphate are the two common metal sulphates which are insoluble in water. Insoluble metal sulphates are prepared by precipitation reactions. Barium sulphate and lead sulphate are prepared by reacting soluble salts of barium and lead with dilute sulphuric acid, respectively. For example, a white precipitate of barium sulphate is formed when barium nitrate reacts with dilute sulphuric acid as shown in the following chemical equation:

$$Ba(NO_3)_2(aq) + H_2SO_4(aq) \longrightarrow BaSO_4(s) + 2HNO_3(aq)$$

A similar reaction takes place if barium chloride is used instead of barium nitrate.

Activity 10.21



Aim: To prepare lead sulphate by the reaction of lead nitrate and dilute sulphuric acid

Requirements: Two beakers, measuring cylinder, glass rod, filter funnel, stand and clamp squeeze bottle, filter papers, 2 M lead(II) nitrate solution, 2 M sulphuric acid, and distilled water

Procedure

1. Transfer about 25 cm^3 of 2 M lead(II) nitrate solution into a beaker.
2. Add about 25 cm^3 of 2 M sulphuric acid to the lead(II) nitrate and stir the mixture by using a glass rod.
3. Filter the mixture and wash the residue with distilled water.
4. Dry the residue between filter papers and record all the observations.

Questions

1. Give a balanced chemical equation for the reaction between lead(II) nitrate and dilute sulphuric acid.
2. What is the colour of the precipitates formed in Question 1?
3. Identify the type of the reaction that occurred in the experiment.
4. What would be the reaction products if lead(II) nitrate is replaced by lead(II) chloride in the experiment? Support your answer with a balanced chemical equation.

Activity 10.22



Aim: To prepare barium sulphate by the reaction of barium chloride with dilute sulphuric acid

Requirements: Two beakers, glass rod, filter funnel, stand and clamp, measuring cylinder, squeeze bottle, filter papers, 2 M barium chloride solution, 2 M sulphuric acid, and distilled water

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Procedure

1. Transfer 25 cm^3 of 2 M barium chloride into a beaker.
2. Add about 25 cm^3 of 2 M sulphuric acid into the beaker containing the barium chloride and stir the mixture using a glass rod.
3. Filter the mixture and wash the residues with distilled water.
4. Dry the residue under the sun or by rubbing the residue between filter papers and record the observations.

Questions

1. Write a balanced ionic equation for the reaction between barium chloride solution and dilute sulphuric acid.
2. What is the colour of the precipitate formed in Step 2?
3. Name the type of the reaction that occurred in the experiment.

Chemical properties of metal sulphates

The chemical properties of metal sulphates include the action of heat on their solids and reactions with other chemical compounds.

Effects of heat on metal sulphates

In general, sulphates are more stable to heat than their corresponding nitrates. The sulphates of metals lower in the reactivity series must be strongly heated to decompose.

When crystals of hydrated copper(II) sulphate are heated, the colour changes from blue to white due to the formation of anhydrous copper(II) sulphate.

$$\underset{\text{Blue}}{CuSO_4 \cdot 5H_2O(s)} \xrightarrow{\text{heat}} \underset{\text{White}}{CuSO_4(s) + 5H_2O(g)}$$

On strong heating, the anhydrous copper(II) sulphate decomposes to copper(II) oxide and sulphur trioxide gas which turns a wet blue litmus paper red.

$$CuSO_4(s) \xrightarrow{\text{strong heat}} CuO(s) + SO_3(g)$$

When hydrated iron(II) sulphate is gently heated, it gives up its water of crystallisation to give anhydrous iron(II) sulphate which is accompanied with colour change from green to white.

$$\underset{\text{Green}}{FeSO_4 \cdot 7H_2O(s)} \xrightarrow{\text{heat}} \underset{\text{White}}{FeSO_4(s) + 7H_2O(g)}$$

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On strong heating, the anhydrous iron(II) sulphate decomposes to form a reddish-brown iron(III) oxide, sulphur trioxide, and sulphur dioxide.

$$2FeSO_4(s) \xrightarrow{\text{strong heat}} Fe_2O_3(s) + SO_3(g) + SO_2(g)$$

Sulphur dioxide turns the filter paper dipped in acidified potassium dichromate solution green.

When hydrated iron(III) sulphate is gently heated, it releases its water of crystallisation to form a reddish-brown solid of anhydrous iron(III) sulphate.

$$Fe_2(SO_4)_3 \cdot 9H_2O(s) \xrightarrow{\text{heat}} Fe_2(SO_4)_3(s) + 9H_2O(g)$$

On strong heating, the anhydrous iron(III) sulphate decomposes to a reddish-brown iron(III) oxide and sulphur trioxide. Sulphur trioxide turns a moist blue litmus paper red.

$$Fe_2(SO_4)_3(s) \xrightarrow{\text{strong heat}} Fe_2O_3(s) + 3SO_3(g)$$

Zinc sulphate decomposes in a similar way on strong heating.

Activity 10.23



Aim: To investigate the effects of heat on metal sulphates

Requirements: Test tubes, Bunsen burner, test tube holder, spatula, filter paper, blue litmus paper, blue cobalt chloride paper, hydrated copper(II) sulphate ($CuSO_4 \cdot 5H_2O$), hydrated iron(II) sulphate ($FeSO_4 \cdot 7H_2O$), hydrated iron(III) sulphate ($Fe_2(SO_4)_3 \cdot 9H_2O$), and acidifed potassium dichromate(VI) solution

Procedure

1. Transfer one spatulaful of hydrated copper(II) sulphate into a test tube.
2. Heat the sample gently and test any liquid formed by using a cobalt(II) chloride paper.
3. Place a piece of wet blue litmus paper at the mouth of the test tube and a filter paper dipped in acidified potassium dichromate solution, then heat strongly record the observations.
4. Repeat Steps 1–3 by using the hydrated iron(II) sulphate and iron(III) sulphate in the place of hydrated copper(II) sulphate.

Questions

1. What is the significance of using cobalt chloride paper?
2. Write a balanced chemical equation for the effect of heat on zinc sulphate.

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Test for sulphates

Barium ions (Ba^{2+}) and lead ions (Pb^{2+}) are used to test the presence of sulphate ions because they form insoluble sulphates. Dilute hydrochloric acid is used to acidify the test solution and then a few drops of a barium chloride or lead chloride solution is added. If the test solution contains sulphate ions, then a white precipitate of barium sulphate or lead(II) sulphate will be formed.

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$$

$$Pb^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s)$$

Any carbonate ions present in the test solution are removed by the addition of hydrochloric acid. The test can also be performed by adding barium nitrate or lead nitrate solution to the test solution acidified by nitric acid.

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Activity 10.24



Aim: To test for the presence of a sulphate in an aqueous solution

Requirements: Test tubes, measuring cylinder, dropper, test solution, barium chloride solution, barium nitrate solution, dilute hydrochloric acid, and dilute nitric acid

Procedure

1. Transfer about 2 cm^3 of the test solution into a clean test tube.
2. Add about 3 to 4 drops of dilute hydrochloric acid followed by two drops of barium chloride. Record the observations.
3. Repeat the Steps 1 and 2 by using dilute nitric acid in the place of hydrochloric acid and barium nitrate in the place of barium chloride.

Questions

1. From the observations, identify the type of the metal salt present in the test solution. Give balanced chemical equations to support your answer.
2. Explain what will happen if the test solution is not acidified in both experiments. Give balanced chemical equations to support your answer.
3. Write a balanced ionic equation for the acidification process in the experiment.
4. Can sulphuric acid be used to acidify the test solution? Explain.

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Uses of metal sulphates

The following are some of the uses of some metal sulphates:

Calcium sulphate

Calcium sulphate, which is commonly known as Plaster of Paris (POP), is used in the following ways:

(a) in making moulds, statues, and chalks,
(b) as a plaster,
(c) in hospitals to immobilise broken bones,
(d) as a fireproofing materials, and
(e) in making decorative materials and as cement in ornamental casting.

Iron(II) sulphate

Iron(II) sulphate is used:

(a) to make tablets prescribed to patients who have iron deficiency.
(b) to prepare a reddish-brown iron(III) oxide. The oxide commonly referred to as '*red oxide*' is used as a pigment.
(c) to make potassium hexacyanoferrate(II) which reacts with iron(III) ions to form a dark blue insoluble crystalline solid called Prussian blue which is used extensively as a dye for blue print paper ink.
(d) as a weed killer.
(e) for the treatment of wastewater.
(f) to coagulate (bind together) blood in slaughterhouses.
(g) in tanning leather (iron-tanning).
(h) as a fungicide.

Barium sulphate

Barium sulphate is used in the following ways:

(a) It is used as a white pigment in white paints.
(b) When taking X-ray pictures of the digestive tract, barium meal is taken. Barium ions (Ba^{2+}) have a high ability to scatter X-rays which helps to detect the illnesses and disorders in the digestive tract.

Aluminium sulphate

(a) As a mordant in dyeing. A *mordant* is a substance that is used to fix a pigment into a fabric during the dyeing process.
(b) As a size in paper making. A *size* is a substance that is added to the paper pulp to reduce the tendency of the paper to absorb liquids.

Potassium aluminium sulphate

(a) Potassium aluminium sulphate (alum) is used in the dyeing industry as a source of aluminium ions (Al^{3+}) which are uncontaminated by iron(III) ions.
(b) It is used in water treatment plants as a coagulant. As a coagulant, it binds together very fine suspended particles (colloids) into large particles that can be removed by settling and filtration.
(c) Alum is used in the tanning of leather.

Uses of other metal sulphates

(a) *Glauber's* salt which is hydrated sodium sulphate ($Na_2SO_4 \cdot 10H_2O$), is used as a mild laxative.
(b) Copper(II) sulphate is used as a catalyst in the preparation of ethanol.
(c) Hydrated zinc sulphate ($ZnSO_4.7H_2O$) is used in the dyeing industry as an antiseptic and in preserving wood. It is also used in zinc-plating by electrolysis.

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Exercise 10.3

1. (a) Explain four methods that can be used to prepare soluble metal sulphates in the laboratory.

   (b) Give one example of a metal sulphate that can be prepared by using each of the methods in 1(a).

2. Describe how you can prepare barium sulphate in the laboratory using barium nitrate and sulphuric acid as the starting materials.

3. Describe a chemical test that can be used to test the presence of a metal sulphate in a solution.

4. Explain five uses of metal sulphates.

Chapter summary

1. Nitrates are salts derived from nitric acid.
2. Metal nitrates can be prepared by the reaction of dilute nitric acid with metals, metal oxides, metal hydroxides, or metal carbonates.
3. The presence of nitrate ions in an unknown sample can be tested by performing the brown-ring test, copper turnings test, or by decomposition of the nitrates (except for the nitrates of sodium and potassium).
4. Metal nitrates are used in agriculture, photography, weapons, food preservation, and antiseptics.
5. Metal chlorides are salts that are derived from hydrochloric acid.
6. Metal chlorides are either prepared by the direct action of dry chlorine gas on metals or by the reactions of dilute hydrochloric acid with metals, metal oxides, metal hydroxides, and metal carbonates. Insoluble metal chlorides are prepared by the precipitation method.
7. Metal chlorides find applications in the manufacture of dry batteries, food preservation, petroleum chemistry, spinning industry, and dentistry.
8. Sulphates are salts that are derived from sulphuric acid.
9. The methods of preparing metal sulphates depend on whether the sulphate is soluble or insoluble in water. All metal sulphates are soluble in water except the sulphates of barium and lead. Soluble sulphates are prepared by reacting a metal, metal carbonate, metal hydroxide or metal oxide with dilute sulphuric acid. Insoluble metal sulphates are prepared by precipitation reactions.
10. Metal sulphates find applications in medicine, agriculture, sewage treatment, domestic water treatment, manufacturing, and building industries.

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Revision exercise 10

Choose the correct answer for questions 1 to 5.

1. A sample of a solid compound X was strongly heated in a test tube. No observable changes were noted during the heating. When dilute nitric acid was added to a sample of solid compound X, some effervescence was observed. The gas produced formed a white precipitate with limewater. Compound X is likely to be:
   (a) zinc nitrate.
   (b) copper(II) sulphate.
   (c) potassium carbonate.
   (d) magnesium chloride.

2. Which of the following pairs of compounds can be used to test the presence of chloride ions in a solution?
   (a) Nitric acid and silver nitrate solution.
   (b) Barium chloride and sodium hydroxide solution.
   (c) Aqueous ammonia and silver nitrate solution.
   (d) Nitric acid and sodium chloride.

3. Which equation shows the direct method of preparing metal chlorides?
   (a) $NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$
   (b) $Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$
   (c) $2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$
   (d) $AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(aq) + NaNO_3(aq)$

4. When iron(II) sulphate is heated strongly, it decomposes to form;
   (a) iron(III) oxide and sulphur trioxide.
   (b) iron(III) oxide, sulphur dioxide, and oxygen.
   (c) iron(III) oxide, sulphur trioxide, and sulphur dioxide.
   (d) iron(III) oxide and sulphur dioxide.

5. When an unknown salt solution is heated with concentrated sulphuric acid and copper turnings, brown fumes of nitrogen dioxide gas are observed. The unknown metal salt solution is likely to be;
   (a) a metal sulphate.
   (b) a metal chloride.
   (c) a metal nitrate.
   (d) a metal carbonate.

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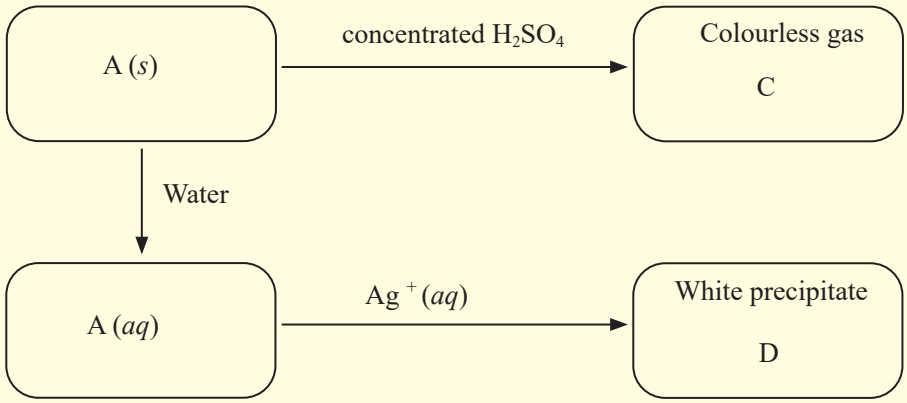
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6. Give the balanced chemical equations for all the reactions occurring in Question 5.
7. Using the balanced chemical equations, explain what happens when dilute hydrochloric acid is added to the following:
   (a) Lead(II) nitrate solution
   (b) Iron(II) oxide
   (c) Sodium carbonate
   (d) Magnesium granules
8. In each of the following statements, write TRUE if the statement is correct, and FALSE if the statement is not correct.
   (a) Barium chloride is used to test the presence of sulphates in a compound.
   (b) Sodium nitrate is used to make fertilisers.
   (c) Silver nitrate is highly soluble in water.
   (d) Barium chloride and lead chloride are not soluble in water.
   (e) All metal nitrates produce reddish brown fumes of nitrogen dioxide when heated.
   (f) Silver and lead chlorides are insoluble in water.
   (g) All metal sulphates are soluble in water except the sulphates of barium and lead.
   (h) Metal nitrates are generally more stable to heat than the corresponding sulphates.
9. Write the balanced chemical equations for the reaction between;
   (a) copper and concentrated nitric acid.
   (b) copper and dilute nitric acid.
   (c) zinc and concentrated nitric acid.
   (d) magnesium and a very dilute nitric acid.
   (e) magnesium and concentrated nitric acid.
10. Under certain conditions, iron reacts with chlorine to give anhydrous iron(III) chloride.
   (a) Why an excess of chlorine gas has to be used? Explain.
   (b) Why should moisture not be allowed into the apparatus during the reaction?
   (c) Why iron(II) chloride is not formed? Show using a balanced chemical equation.

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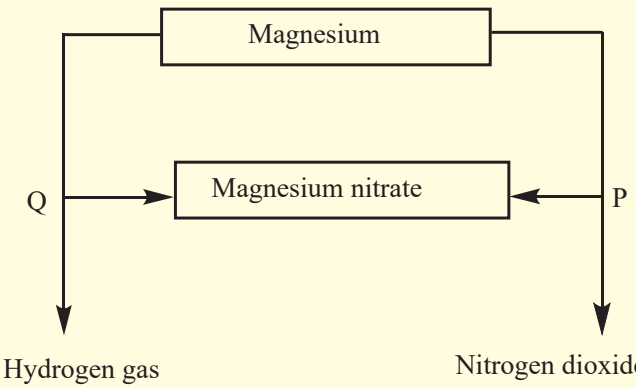
11. Study the following scheme and answer the questions that follow.



(a) Identify the metal salt labelled A.
(b) Identify the colourless gas labelled C.
(c) What is the white precipitate labelled D?
(d) Write an ionic equation for the formation of the white precipitate.

12. When dilute nitric acid is added to a green solid P, a blue solution X is formed and a gas R that forms a white precipitate with limewater. Heating the blue solution of the hydrated salt X, it initially loses its water of crystallisation, and then decomposes to form a black solid S, brown fumes of gas Y, and a gas Z that relights a glowing splint.
(a) Identify the solid X and gases Y and Z.
(b) Write an equation for the reaction between the solid P and dilute nitric acid.
(c) Write an equation for the formation of the solid S.

13. The following scheme shows the preparation of magnesium nitrate. Study the scheme and then answer the questions that follow.



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(a) Identify the substances P and Q.
(b) Write the equation for the reaction between magnesium and substance Q.

14. Write short notes on the identification of the following:
(a) Nitrates.
(b) Chlorides.
(c) Sulphates.

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Answers to numerical questions

Chapter Four

Exercise 4.1

1. (a) 1.75×10^{23} particles
   (b) 4.03×10^{22} particles
   (c) 1.14×10^{25} particles
   (d) 3.31×10^{25} particles
2. (a) 3.55×10^{-22} mol
   (b) 1.25×10^{27} mol
   (c) 3.32×10^{-24} mol
   (d) 6.14×10^{-27} mol
3. (b) 3.01×10^{23} lead ions and 6.02×10^{23} nitrate ions
4. (a) 0.4 mol (b) 0.53 mol
   (c) 16.87 mol
5. (a) 224 g (b) 7.3 g
   (c) 5.58 g (d) 0.27 g

Exercise 4.2

1. (a) 0.54 mol (b) 2.68×10^{-4} mol
   (c) 334.82 mol (d) 1.21×10^{-4} mol
2. (a) 3.58×10^{-3} dm^3 (b) 392 dm^3
   (c) 6.27 dm^3 (d) 470.4 dm^3
3. 0.086 M
4. 0.8 g
5. (a) 80 g dm^{-3}
   (b) 1 mol
   (c) 50 cm^3
6. 10.2 cm^3
7. 829.48 g
8. (a) 0.252 mol (b) 0.168 mol
   (c) 3.76 dm^3 (d) 15.97 g
9. (a) 3.31 g (b) 2.23 g
   (c) 0.448 dm^3

Revision exercise 4

7. 6.97×10^{-12} mol
8. 0.5 mol
9. (a) 0.087 mol
   (b) 5.24×10^{22} molecules
10. 44
11. 0.7143 M
12. 3.36 dm^3
13. 21.4 g
14. 0.567 mol
16. (b) 156.53 g (c) 1.127 dm^3
17. 0.104 dm^3 or 104 cm^3
18. 0.112 dm^3 or 112 cm^3
19. 90 g mol^{-1}
20. 3.011×10^{23} ions
21. (a) 0.286 M (b) 7
22. (b) 1.635 g
23. 34.8%
24. 1.344 g

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25. (a) (i) 18 tonnes (ii) 37 tonnes
(b) 5.6×10^6 dm^3
(c) 1.506×10^{29} molecules

Chapter 5

Exercise 5.2

2. 4.44 g
3. 0.0502 M
4. 0.625 M

Exercise 5.3

1. (a) 39 (b) Potassium
2. 58.89%
3. 2 molecules

Revision exercise 5

8. 36000 cm^3
9. (a) 5.176 g/dm^3 (b) 0.0528 M
10. (a)

Pilot	1	2	3
12.90	12.50	12.60	12.40

(b) 12.50 cm^3
(d) (i) 0.025 mol
(ii) 0.0125 mol
(iii) 1 M
(iv) 98 g/L

11. (a) 0.0375 mol (b) 0.0375 mol
(c) 106 (d) 23, Y is sodium

12. 80%

13. 10 molecules, $Na_2CO_3 \cdot 10H_2O$

14. (a) 1056 g in 2.5 dm^3
(b) 4.3080 M (c) 58 cm^3

16. (c) 0.725 g (d) 5 tablets
(e) $2\frac{1}{2}$ days

17. (a) 0.1804 M (b) 1.09 cm^3

Chapter 6

Exercise 6.3

1. (a) 0.02 F (b) 0.04 F
2. (b) 0.028 mol

Revision exercise 6

8. (a) 190 C (b) 879.63 C
(c) 288.32 C
9. (a) 574 C (b) 0.16 A (c) 31.75
10. (a) 47.86 A (b) 1.421×10^5 g or 142.1 kg

Chapter 7

Exercise 7.1

3. 0.0007 M s^{-1}
4. 0.025 M s^{-1}
5. (a) 0.0005 g s^{-1}
(b) 0.0000076 mol s^{-1}

Revision exercise 7

11. 0.0022 mol s^{-1}
12. 0.0683 g/s

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Glossary

Acid	a substance which produces hydrogen ions (H^+) in water as the positively charged ions
Activated state	a state in which chemicals species are at their highest energy level
Activation energy	the minimum amount of energy required by colliding reactant particles in order to react
Amphoteric oxides	oxides which can react with both acids and bases
Analyte	a chemical species of interest in a particular analytical procedure
Anions	are negatively charged particles
Anode	a positively charged electrode in an electrolytic cell through which electrons leave the electrolyte
Auto-reduction	the reduction by heating in the air
Avogadro's Law	states that equal volumes of all gases measured at the same temperature and pressure, contain the same number of particles
Base	a substance that neutralises an acid by reacting with hydrogen ions
Basic oxides	oxides which react with acids to give salt and water as the only products
Basicity of an acid	the number of ionisable hydrogen ions per molecule of an acid
Blast furnace	a metallurgical furnace used for smelting to produce metals such as iron
Blister copper	is an impure form of copper having a blister like surface due to the release of gas during cooling
Cast iron	the pure iron obtained from the furnace which contains 2% to 4% of carbon
Catalytic decomposition reaction	a decomposition reaction where a catalyst is introduced to alter the rate of a chemical reaction but itself remains unchanged at the end of the reaction
Cathode	a negatively charged electrode in an electrolytic cell through which electrons enter the electrolyte

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Cations	are positively charged particles
Collision theory	a theory that predicts the occurrence of a chemical reaction
Concentration	the quantity of the solute in the solution
Conductors	substances that conduct electricity/heat
Crystallisation	a process of forming solid crystals from a homogeneous solution
Crystallisation point	the temperature at which crystals start to be formed as the solution evaporates
Deliquescent salt	a substance which when exposed to air absorbs water vapour and forms a solution
Diaphragm cell	an electrolytic cell used in the production of chlorine gas from sodium chloride brine
Dilution factor	the number of times the volume of the more concentrated solution is diluted to obtain the less concentrated solution
Double displacement reaction	a reaction in which two compounds react by an exchange of ions to form two new compounds
Dynamic equilibrium	an equilibrium state that exists when the forward and backward reactions occur at the same rate
Effective collision	collision which results in a chemical reaction
Efflorescent substance	a substance when left in air loses all its water of crystallisation
Electrochemical equivalent	the amount of a substance produced or consumed in grams when one coulomb of electricity is passed in an electrochemical reaction
Electrochemical series	an arrangement of elements depending on their ability to gain or lose electrons
Electrodes	terminal rods through which electrons enter or leave the electrolyte
Electrolysis	a process in which an electric current is passed through an electrolyte to drive a non-spontaneous chemical reaction
Electrolyte	a substance in solution or molten state that gives randomly moving ions
Electrolytic cell	the apparatus in which electrolysis is carried out
Electrolytic decomposition	a reaction achieved by exposing an electrolyte to electricity
Electropositivity	the tendency of a metal to lose electrons and form positive ions

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End point	the point in titration process whereby reactants have reacted completely and is indicated by a physical change such as colour change of an indicator used or precipitates formation
Endothermic reaction	a reaction that absorbs or uses more energy than it releases
Energy level diagram	the energy change in a chemical reaction presented in a graphical way
Enthalpy change	the difference between the energy released and absorbed in a chemical reaction
Equilibrium mixture	a mixture in which components of both reactants and products are present in equal ratio with each other (a mixture with constant concentrations of products and reactants).
Exothermic reaction	a reaction that releases more energy than it absorbs
Extraction of metal	a process of separating a metal from the rest of the materials in its ore
Faraday's constant	the amount of electric charge carried by one mole of electrons which is approximately 96,500 coulombs per mole
Faraday's First Law of electrolysis	relates the mass of a substance liberated or deposited and the quantity of electricity used.
Faraday's Second Law of electrolysis	relates the mass of a substance liberated or deposited and the equivalent weights.
Hard water	the water that does not readily form lather with soap
Hard water purification	a process by which undesirable chemical compounds, organic, and inorganic materials are removed from water
Heat content (Enthalpy)	the amount of heat transferred during chemical reaction at a constant pressure
Heterogeneous reaction	chemical reaction in which the reactants are components of two or more phases (solid and gas, or solid and liquid)
Hydroxide	is a base or alkali containing the hydroxide ion
Hygroscopic salt	a substance which absorbs water from air without necessarily forming a solution
Indicator	a substance that changes its colour or smell to indicate the presence or absence of some chemical species
Inorganic acid	an acid formed from inorganic species

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Intermolecular energy	the potential energy stored in the bonds of two different molecules
Intramolecular energy	the potential energy stored in the bonds of the same molecules
Ionic migration	the movement of ions towards their oppositely charged electrodes
Irreversible reaction	a reaction that proceeds in only one direction
Le Chatelier's Principle	describes the effect of change in condition on the equilibrium of a system
Liming materials	are basic chemicals such as calcium oxide and calcium hydroxide which are added in soil to adjust the pH by neutralising the soil acidity
Metallurgy	the branch of science that deals with the properties, production and purification of metals
Micropipette	a pipette which delivers very small volumes of liquids
Mineral	a metallic compound occurring in the Earth's crust along with some impurities like rock particles
Molar concentration (Molarity)	the amount of substance in moles that has dissolved in a solvent to form a litre of solution
Molar mass	the mass of one mole of a substance
Molar volume	the volume of one mole of a substance
Mole	the amount of the substance which contains as many particulate entities as there are carbon atoms in 12 g of carbon–12 isotope
Mole ratio	numbers of chemical species appearing before each reactant and product in a balanced chemical equation to show their quantity relationships
Mordant	a substance used to fix a pigment into a fabric during the dyeing process
Neutralisation	a process where an oxide or hydroxide ions of a base combine with an acid in the correct ratio to form salt and water
Non-conductors	a substance that does not conduct electricity

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Non-spontaneous reaction	a reaction which requires an external source of energy for it to occur
Organic acid	an acidic organic substance synthesised in industry or obtained directly from natural organic materials
Oxide	a binary compound formed by the chemical combination of oxygen and another element
Oxidising agent	a substance that gains electron(s) or donates oxygen to another substance
Permanent hard water	the water that contains sulphates or chlorides of calcium, magnesium or sometimes copper, iron, and manganese
pH	the negative logarithm of hydrogen ion concentration
Pig iron	the impure iron obtained by smelting iron ore in a blast furnace
Potential energy	the energy due position and composition
Precipitate	are insoluble chemical species formed after the chemical reaction
Primary standard	a known substance of high purity (about 99.9% pure) and stability which may be dissolved in a known volume of solvent to give a primary standard solution
Product	a substance that is formed after the process of a chemical reaction
Reactant	a chemical species that begins the process of a chemical reaction
Reactivity series	the arrangement of metals according to their ability of losing electrons or reducing power
Redox reaction	a reaction in which both oxidation and reduction occur together
Reducing agent	a substance that donates electron(s)
Reduction half reaction	a half equation which represents the gaining of electron(s)
Refractory	a solid material that resists decomposition when subjected to high heat
Relative atomic mass	the atomic masses of all the atoms of a particular element compared to the atomic mass of carbon–12 isotope

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Relative molecular mass	is the average mass of a molecule relative to $\frac{1}{12}$ of the mass of a carbon–12 atom
Reversible reaction	a reaction that proceeds in both the forward and backward directions
Salt	an ionic compound which is made up of cation(s) and anion(s)
Saturated solution	the solution that cannot hold any more solute at room temperature and standard atmospheric pressure
Secondary standard	a reagent determined by direct weighing and standardised by using a primary standard
Smelting	the extraction of metal from its ore by a process which involves heating and melting
Soft water	the water that readily forms lather with soap
Soluble basic oxide	an oxide which dissolves in water to form an alkali
Spectator ions	ions that do not change their oxidation states in the reaction
Standard solution	a solution with a known concentration
Standardisation	the process of establishing the unknown concentration of the solution by using a standard solution
Stoichiometric coefficients	the numbers which represent the actual ratios of the chemical species quantities in a balanced chemical equation
Stoichiometry	the relationship between the relative quantities of reactants and products in a balanced chemical reaction
Strong electrolyte	a substance which dissociates completely in aqueous solution to form ions
Supernatant	a clear liquid that lies above the solid residue after centrifugation, precipitation, crystallisation, or settling
Temporary hard water	the water containing hydrogencarbonates of calcium or magnesium
Tensile strength	the ability of a metal to withstand forces applied on it without breaking

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Thermal decomposition	a reaction that occurs when a compound is exposed to direct heat
Thermite	a mixture of finely powdered aluminium and iron oxide that produces a very high temperature on combustion, used in welding and for incendiary bombs
Thermochemical equation	a chemical equation that includes the enthalpy change of the reaction
Thermochemistry	the study of energy changes that accompany chemical reactions
Titrand	a solution of unknown concentration usually added to the conical flask during titration
Titrant	a standard solution usually added to the burette during titration
Titration	The process of determining the quantity of analyte by adding measured volumes of a titrant
Transfer pipette	a pipette used for transfer of small volumes of liquids
Universal indicator	a chemical substance that displays several colour changes across a wide range of pH values
Volumetric analysis	an analytical method based on the measurements of volumes of one solution required to react completely with a definite volume of the other solution
Water softening	a process in which the sulphates or hydrogencarbonates of calcium, magnesium and sometimes iron, copper, and manganese are removed from hard water
Weak electrolytes	a chemical species which dissociates partially in aqueous solution to give the respective ions
Word equation	an equation which represents a chemical reaction by using the names of the chemical species that are involved

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Appendices

Appendix 1: *Table of atomic numbers and atomic masses of elements*

Name of the element	Symbol	Atomic number	Atomic mass
Actinium	Ac	89	227
Aluminium	Al	13	26.985
Americium	Am	95	243
Antimony	Sb	51	121.75
Argon	Ar	18	39.948
Arsenic	As	33	74.926
Astatine	At	85	210
Barium	Ba	56	137.34
Berkelium	Bk	97	247
Beryllium	Be	4	9.0122
Bismuth	Bi	83	208.90
Bohrium	Bh	107	264
Boron	B	5	10.811
Bromine	Br	35	79.909
Cadmium	Cd	48	112.40
Caesium	Cs	55	132.95
Calcium	Ca	20	40.08
Californium	Cf	98	251
Carbon	C	6	12.0115
Cerium	Ce	58	140.40
Chlorine	Cl	17	35.453
Chromium	Cr	24	51.996
Cobalt	Co	27	58.9332
Copernicium	Cn	112	285
Copper	Cu	29	63.54
Curium	Cm	96	247
Darmstadtium	Ds	110	281
Dubnium	Db	105	262
Dysprosium	Dy	66	162.50
Einsteinium	Es	99	252
Erbium	Er	68	167.26
Europium	Eu	63	151.96
Fermium	Fm	100	257
Flerovium	Fl	114	289
Fluorine	F	9	18.9984
Francium	Fr	87	223

Name of the element	Symbol	Atomic number	Atomic mass
Gadolinium	Gd	64	157.25
Gallium	Ga	31	69.72
Germanium	Ge	32	72.64
Gold	Au	79	196.967
Hafnium	Hf	72	178.49
Hassium	Hs	108	277
Helium	He	2	4.0026
Holmium	Ho	67	164.930
Hydrogen	H	1	1.0079
Indium	In	49	114.818
Iodine	I	53	126.904
Iridium	Ir	77	192.2
Iron	Fe	26	55.847
Krypton	Kr	36	83.80
Lanthanum	La	57	138.91
Lawrencium	Lr	103	262
Lead	Pb	82	207.19
Lithium	Li	3	6.939
Livermorium	Lv	116	293
Lutetium	Lu	71	174.97
Magnesium	Mg	12	24.312
Manganese	Mn	25	54.9380
Meitnerium	Mt	109	268
Mendelevium	Md	101	258
Mercury	Hg	80	200.59
Molybdenum	Mo	42	95.94
Moscovium	Mc	115	289
Neodymium	Nd	60	144
Neon	Ne	10	20.183
Neptunium	Np	93	237
Nickel	Ni	28	58.71
Nihonium	Nh	113	286
Niobium	Nb	41	93
Nitrogen	N	7	14.0467
Nobelium	No	102	259
Oganesson	Og	118	294

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Appendix 1: (*Continued*)

Name of the element	Symbol	Atomic number	Atomic mass
Osmium	Os	76	190.23
Oxygen	O	8	15.9994
Palladium	Pd	46	106.4
Phosphorus	P	15	30.9734
Platinum	Pt	78	195.09
Plutonium	Pu	94	244
Polonium	Po	84	209
Potassium	K	19	39.102
Praseodymium	Pr	59	141
Promethium	Pm	61	145
Protactinium	Pa	91	231
Radium	Ra	88	226
Radon	Rn	86	222
Rhenium	Re	75	186
Rhodium	Rh	45	102.905
Roentgenium	Rg	111	280
Rubidium	Rb	37	85.47
Ruthenium	Ru	44	101
Rutherfordium	Rf	104	261
Samarium	Sm	62	150
Scandium	Sc	21	44.946
Seaborgium	Sg	106	269
Selenium	Se	34	78.96

Name of the element	Symbol	Atomic number	Atomic mass
Silicon	Si	14	28.086
Silver	Ag	47	107.870
Sodium	Na	11	22.9898
Strontium	Sr	38	87.62
Sulphur	S	16	32.064
Tantalum	Ta	73	181
Technetium	Tc	43	98
Tellurium	Te	52	127.604
Tennessine	Ts	117	294
Terbium	Tb	65	159
Thallium	Tl	81	204.37
Thorium	Th	90	232.038
Thulium	Tm	69	169
Tin	Sn	50	118.71
Titanium	Ti	22	47.867
Tungsten	W	74	183.85
Uranium	U	92	238.03
Vanadium	V	23	50.942
Xenon	Xe	54	131.30
Ytterbium	Yb	70	173
Yttrium	Y	39	89
Zinc	Zn	30	65.38
Zirconium	Zr	40	91.22

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Appendix 2: *Preparation of common gases*

Gases	Reagents	Methods	Chemical equations	Methods of collection
Oxygen (O_2)	6% hydrogen peroxide + manganese(IV) oxide (catalyst)	Add hydrogen peroxide to about 5 g of manganese(IV) oxide	$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$	Over water
	Potassium permanganate	Heat solid potassium permanganate	$2KMnO_4(s) \xrightarrow{heat} K_2MnO_4(s) + MnO_2(s) + O_2(g)$	Over water
Hydrogen (H_2)	Zinc (granulated) + 6 M hydrochloric acid	Add 6 M hydrochloric acid to 5–10 g of granulated zinc	$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$	Over water
Nitrogen (N_2)	Ammonium chloride + sodium nitrite	Heat an aqueous solution of ammonium chloride and sodium nitrite	$NH_4Cl(aq) + NaNO_2(aq) \xrightarrow{heat} NaCl(aq) + 2H_2O(l) + N_2(g)$	Over water. Purified by passing through an aqueous sulphuric acid solution containing potassium dichromate
	Ammonium dichromate	Heat ammonium dichromate	$(NH_4)_2Cr_2O_7(s) \xrightarrow{heat} N_2(g) + 4H_2O(l) + Cr_2O_3(s)$	Over water
	Oxidation of ammonia	Pass ammonia gas over a red hot copper oxide	$2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3H_2O(l) + 3Cu(s)$	Over water
Chlorine (Cl_2)	Potassium permanganate + concentrated hydrochloric acid	Add concentrated hydrochloric acid, drop by drop, onto a small amount of potassium permanganate crystals in a flask	$2KMnO_4(s) + 16HCl(aq) \longrightarrow 2KCl(aq) + 2MnCl_2(aq) + 8H_2O(l) + 5Cl_2(g)$	Upward displacement of air
Carbon dioxide (CO_2)	Calcium carbonate + 6 M hydrochloric acid	Add 6 M hydrochloric acid to 5–10 g calcium carbonate	$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$	Upward displacement of air

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Ammonia (NH_3)	Ammonium chloride + calcium hydroxide	Gently heat a mixture of ammonium chloride and calcium hydroxide in water.	$2NH_4Cl(s) + Ca(OH)_2(aq) \xrightarrow{heat} CaCl_2(aq) + 2H_2O(l) + 2NH_3(g)$	Upward displacement of air
Hydrogen chloride (HCl)	Sodium chloride + concentrated sulphuric acid	Slowly add concentrated sulphuric acid to solid sodium chloride	$2NaCl(s) + H_2SO_4(l) \longrightarrow Na_2SO_4(aq) + 2HCl(g)$	Upward Displacement of air
Nitrogen monoxide (NO)	Copper (turnings) + 6 M nitric acid	Add 6 M nitric acid to 5–10 g copper turnings	$3Cu(s) + 8HNO_3(aq) \xrightarrow{heat} 3Cu(NO_3)_2(aq) + 4H_2O(l) + 2NO(g)$	Over water
Nitrogen dioxide (NO_2)	Copper (turnings) + 10 M nitric acid	Add concentrated nitric acid to 5–10 g copper	$Cu(s) + 4HNO_3(aq) \xrightarrow{heat} Cu(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$	Upward displacement of air
Sulphur dioxide (SO_2)	Sodium sulphite (or hydrogensulphite) + 2 M hydrochloric acid	Add dilute hydrochloric acid to 5–10 g of sodium sulphite (or hydrogensulphite)	$Na_2SO_3(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + SO_2(g)$	Upward displacement of air

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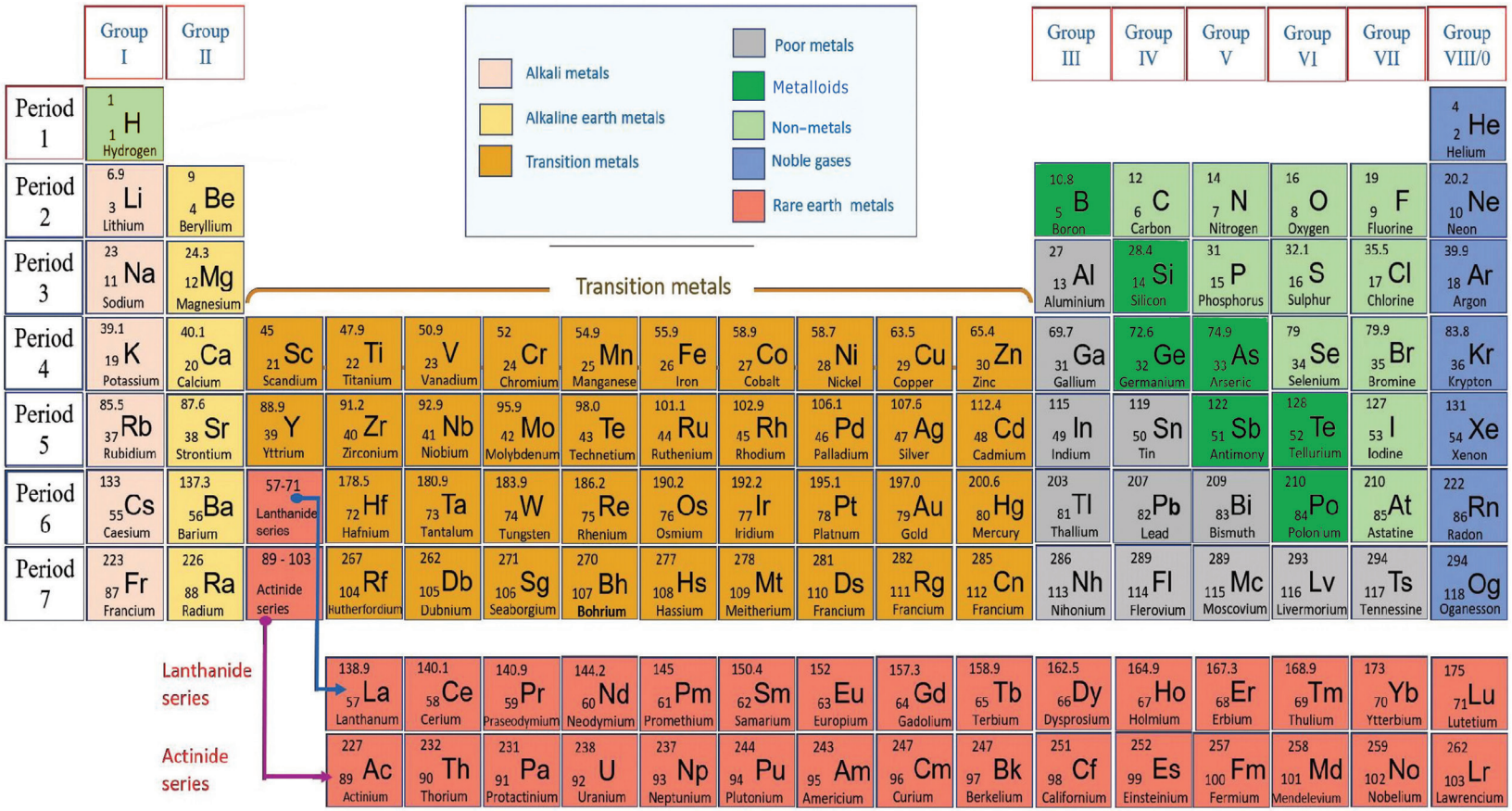
Appendix 3: *Common names and IUPAC names*

Formula	Common name	IUPAC name
H_2SO_4	Sulphuric acid	Tetraoxosulphate(VI) acid
HNO_3	Nitric acid	Trioxonitrate(V) acid
HCl	Hydrochloric acid	Hydrocholoric acid
H_2CO_3	Carbonic acid	Trioxocarbonate(IV) acid
HNO_2	Nitrous acid	Dioxonitrate(III) acid
H_2SO_3	Sulphurous acid	Trioxosulphate(IV) acid
Na_2CO_3	Sodium carbonate	Sodium trioxocarbonate(IV)
$NaNO_3$	Sodium nitrate	Sodium trioxonitrate(V)
CO_2	Carbon dioxide	Carbon(IV) oxide
CO	Carbon monoxide	Carbon(II) oxide
SO_2	Sulphur dioxide	Sulphur(IV) oxide
SO_3	Sulphur trioxide	Sulphur(VI) oxide
N_2O	Nitrous oxide	Nitrogen(I) oxide
NO	Nitric oxide	Nitrogen(II) oxide
NO_2	Nitrogen dioxide	Nitrogen(IV) oxide
N_2O_4	Dinitrogen tetraoxide	Dinitrogen tetraoxide
Na_2SO_4	Sodium sulphate	Sodium tetraoxosulphate(VI)
$CuSO_4$	Copper(II) sulphate	Copper(II) tetraoxosulphate(VI)
$Al(NO_3)_3$	Aluminium nitrate	Aluminium trioxonitrate(V)
$CaCO_3$	Calcium carbonate	Calcium trioxocarbonate(IV)
CO_3^{2-}	Carbonate ion	Trioxocarbonate(IV) ion
$FeCO_3$	Iron(II) carbonate (Ferrous carbonate)	Iron(II) trioxocarbonate(IV)
PbO	Lead(II) oxide	Lead(II) oxide
FeO	Ferrous oxide	Iron(II) oxide
Fe_2O_3	Ferric oxide	Iron(III) oxide
SO_4^{2-}	Sulphate ion	Tetraoxosulphate(VI) ion
SO_3^{2-}	Sulphite ion	Trioxosulphate(IV) ion
NO_3^-	Nitrate ion	Trioxonitrate(V) ion
HCO_3^-	Hydrogencarbonate ion	Hydrogen trioxocarbonate(IV) ion
MnO_4^-	Permanganate ion	Tetraoxomanganate(VII) ion
$Cr_2O_7^{2-}$	Dichromate ion	Heptaoxodichromate (VI) ion

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Appendix 4: Periodic Table



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