MINISTRY OF EDUCATION

GHANA ASSOCIATION OF SCIENCE TEACHERS

(GAST)

CHEMISTRY

for Senior High Schools

Year 3



• Alhassan Kuvidana Mustapha • Christian Dzikunu
• Yirwelle Santus

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3

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Ghana Education Service (GES)

(GAST)



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FOREWORD

Ghana's new Senior High School Curriculum aims to ensure that all learners achieve their potential by equipping them with 21st Century skills, knowledge, character qualities and shared Ghanaian values. This will prepare learners to live a responsible adult life, progress to further studies and enter the world of work. This is the first time that Ghana has developed a Senior High School Curriculum which focuses on national values, attempting to educate a generation of Ghanaian youth who are proud of our country and can contribute effectively to its development.

The Ministry of Education is proud to have overseen the production of these Learner Materials which can be used in class and for self-study and revision. These materials have been developed through a partnership between the Ghana Education Service, teacher unions (Ghana National Association of Teachers-GNAT, National Association of Graduate Teacher -NAGRAT and the Coalition of Concerned Teachers- CCT) and National Subject Associations. These materials are informative and of high quality because they have been written by teachers for teachers with the expert backing of each subject association.

I believe that, if used appropriately, these materials will go a long way to transforming our Senior High Schools and developing Ghana so that we become a proud, prosperous and values-driven nation where our people are our greatest national asset.

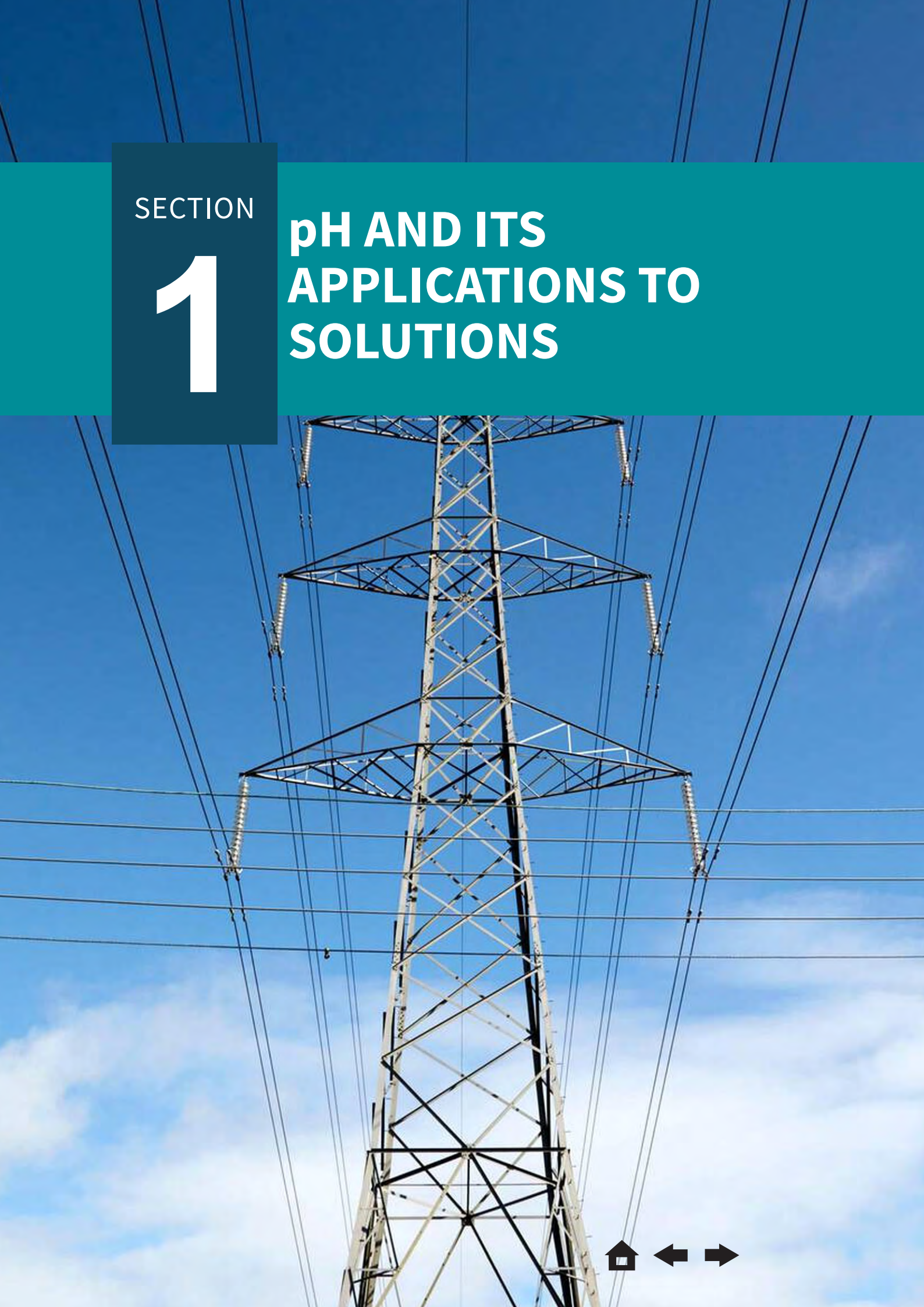
Haruna Iddrisu MP

Minister for Education

SECTION

1

pH AND ITS APPLICATIONS TO SOLUTIONS



PHYSICAL CHEMISTRY

Equilibria – 1

INTRODUCTION

In this section, you will learn about pH and pOH — important ideas that help us understand acids and bases. You will discover how the concentration of ions in solutions affects how acidic or basic it is. You will also be able to:

- use the pH scale to tell if a solution is an acid or a base
- do simple calculations of pH or pOH for strong and weak acids and bases
- analyse how buffers work to keep pH steady in solutions
- explain why pH is important in our bodies, the environment, and in industries.

KEY IDEAS

- **Buffer action** is how a buffer solution works to prevent changes in pH when small amount of acid or base is added.
- **Buffer solution** is a solution that keeps the pH constant, even when small amounts of acid or base are added.
- **Ionic product of water** is the product of the concentrations of hydrogen ions and hydroxide ions in pure water at a given temperature.
- **pH** tells us how acidic or basic a solution is.
- **The pH scale** is a numerical scale used to measure how acidic or basic (alkaline) a solution is.
- **pOH** measures the amount of hydroxide ions (OH^-) in a solution. It helps show how basic or acidic a solution is.

EXPLORATION OF PH AND POH

pH AND pOH

Have you ever tasted something sour like a lemon? Or touched something slippery like soap? That is because some of these substances around us are **acids**, and others are **bases**. Scientists use a special number called **pH** to tell how **acidic** or **basic** or **neutral** a solution is. There is also another number called **pOH** that helps us understand how **basic** a liquid is. pH shows how active the tiny **hydrogen ions** (**H^+**) are in a solution.

If there are **more H^+ ions**, the solution is **acidic** (pH less than 7).

If there are **more OH^- ions** (called hydroxide ions), the solution is **basic** (pH more than 7).

If there are **equal H^+ and OH^- ions**, the solution is **neutral** (pH is 7).

These numbers help us know:

1. if something is safe to touch
2. if it is good for plants
3. if it is right to taste, or even if it is right to drink

The word *pH* stands for "potency of hydrogen". This tells us how many hydrogen ions are in a solution. To check how acidic or basic a liquid is, we use a special scale called the pH scale. So, the pH scale helps scientists know how acidic or basic something is by checking the balance between H^+ and OH^- ions.

pH *is defined as the negative logarithm to the base ten of the hydrogen ion concentration of a substance.*

$$pH = -\log[H^+]$$

$$pH = \log\left(\frac{1}{[H^+]}\right)$$

$$[H^+] = 10^{-pH}$$

This means that:

1. when a liquid has more hydrogen ions (H^+), its pH number is low (more acidic).
2. when a liquid has fewer hydrogen ions, its pH number is high (less acidic or more basic).

This helps scientists show even tiny amounts of hydrogen ions using simple small numbers on the pH scale.

A change of 1 pH unit represents a 10-fold change in the hydrogen ion concentration, $[H^+]$

pOH is the potential of hydroxyl ions in solution. It is defined as the negative logarithm to the base ten of the hydroxyl ion concentration of a substance.

$$pOH = -\log[OH^-]$$

$$pOH = \log\left(\frac{1}{[OH^-]}\right)$$

$$[OH^-] = 10^{-pOH}$$

Table 1.1: A table showing the pH and its corresponding $[H^+]$ and $[OH^-]$

pH	$[H^+]$ (mol/dm^{-3})	$[OH^-]$ (mol/ dm-3)
0	$1.0\times 10^{0} = 1$	$1.0\times 10^{-14} = 0.00000000000001$
1	$1.0\times 10^{-1} = 0.1$	$1.0\times 10^{-13} = 0.0000000000001$
2	$1.0\times 10^{-2} = 0.01$	$1.0\times 10^{-12} = 0.000000000001$
3	$1.0\times 10^{-3} = 0.001$	$1.0\times 10^{-11} = 0.00000000001$
4	$1.0\times 10^{-4} = 0.0001$	$1.0\times 10^{-10} = 0.0000000001$
5	$1.0\times 10^{-5} = 0.00001$	$1.0\times 10^{-9} = 0.000000001$
6	$1.0\times 10^{-6} = 0.000001$	$1.0\times 10^{-8} = 0.00000001$
7	$1.0\times 10^{-7} = 0.0000001$	$1.0\times 10^{-7} = 0.0000001$
8	$1.0\times 10^{-8} = 0.00000001$	$1.0\times 10^{-6} = 0.000001$
9	$1.0\times 10^{-9} = 0.000000001$	$1.0\times 10^{-5} = 0.00001$
10	$1.0\times 10^{-10} = 0.0000000001$	$1.0\times 10^{-4} = 0.0001$
11	$1.0\times 10^{-11} = 0.00000000001$	$1.0\times 10^{-3} = 0.001$
12	$1.0\times 10^{-12} = 0.000000000001$	$1.0\times 10^{-2} = 0.01$
13	$1.0\times 10^{-13} = 0.0000000000001$	$1.0\times 10^{-1} = 0.1$
14	$1.0\times 10^{-14} = 0.00000000000001$	$1.0\times 10^{0} = 1$

pH Scale

The pH scale is a scale that is used to distinguish between acidic, neutral and basic (alkaline) a solution. The pH scale ranges from zero (0) to fourteen (14).

pH Scale Chart (0–14)

pH Value	Type of Solution	H^+ and OH^- Balance	Examples
0 – 6	Acidic	More **H^+ ions** than OH^- ions	Lemon juice, vinegar
7	Neutral	Equal **H^+** and **OH^-** ions	Pure water
8 – 14	Basic (Alkaline)	More **OH^- ions** than H^+ ions	Soap, baking soda water

The pH scale

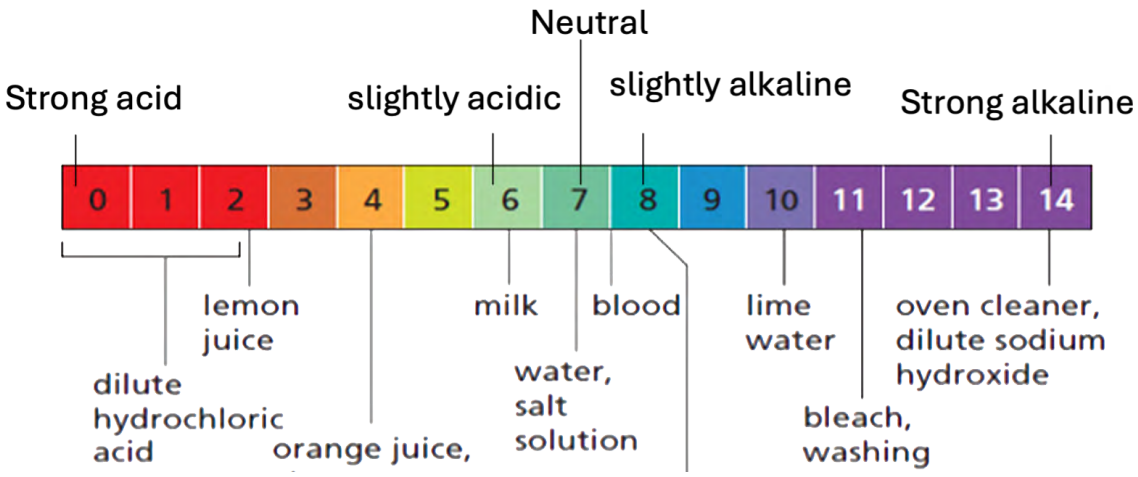


Figure 1.1: *The pH scale showing the pH of household items*

Significance of the Values of pH in Everyday Life

pH tells us how **sour** (acidic) or **soapy** (alkaline/basic) things are. This helps us make safe choices in farming, health, food, water, and cleaning!

1. **Farming (Agriculture)**
   a. Soil pH affects how plants grow.
   b. Most plants thrive in soil that is *slightly* acidic or neutral (around 6–7).
   c. Farmers check soil pH and add:

   i. Lime (to make it less acidic)
   ii. Sulphur (to make it more acidic)

2. **Health & Medicine**
   a. Human blood is slightly basic, about pH 7.4—Slight changes in this pH can make you very sick.
   b. Your stomach has strong acid (pH 1–2) to help digest food.
   c. Doctors check pH in body fluids (like urine) to diagnose health issues.
3. **Food & Cooking**
   a. pH affects taste, freshness, and safety of food.
   b. Vinegar is very acidic and keeps food from spoiling.
   c. Baking needs just the right pH (7 – 8) so cakes rise and taste good. This is done by usually adding baking soda
4. **Water Safety**
   a. Drinking water should be about **pH 7** (neutral).
   b. If water is too acidic or too basic, it can:
      - Damage pipes
      - Harm fish and plants
   c. Water plants adjust pH during purification process to make water safe for consumption.
5. **Cleaning Products**
   a. Some cleaners (like vinegar) need to be **acidic** to remove hard water stains.
   b. Others are **alkaline** (like soap and bleach) to remove grease.
   c. Choosing the right pH cleaner makes cleaning better and safer.

Activity 1.1 Comparing pH of Household Substances

Option A

Materials needed

pH strips or indicator, transparent plastic cup or test tube, spoon, spatula, water samples, lemon juice, vinegar, milk, toothpaste solution, baking soda solution, soap solution.

Procedure

1. Use the link to watch the Video: https://youtu.be/V5Mq_cL9Bck
2. Carry out the following activities and record the colour and the pH of the various solutions in the Table below.

S/n	Test	Colour change observed	pH
a.	Dip a pH strip in 10 cm^3 distilled water in a test tube		
b.	Dip a new pH strip in 10 cm^3 tap water in a test tube		
c.	Dip a new pH strip in small lemon juice in a test tube		
d.	Dip a new pH strip in small orange juice in a test tube		
e.	Dip a new pH strip in small vinegar in a test tube		
f.	Dip a new pH strip in small baking soda solution in a test tube		
g.	Dip a new pH strip in small soap solution in a test tube		
h.	Dip a new pH strip in small milk solution in a test tube		
i.	Dip a new pH strip in small toothpaste solution in a test tube		

3. Compare results, rank the solutions from most acidic to most basic.

Option B

Materials needed

Beetroot juice (natural pH indicator) or universal indicator, pH scale chart (0–14), transparent plastic cups or test tubes, dropper or spoon, marker, lemon juice, vinegar, orange juice, baking soda solution, milk, soap solution, toothpaste solution, tap water

Procedure

1. Create an indicator from red beetroot juice (Beetroot juice is a natural indicator which changes colour with pH. It is red in acid and purple in base).
2. Boil chopped beetroot in water, cool, and use the juice.
3. Label each cup or test tube with the name of the household substance.
4. Pour a small amount of each substance into its corresponding cup.
5. Add a few drops of red beetroot juice to each sample.
6. Observe and record the colour change in each cup.
7. Match the colour to the pH scale chart to estimate the pH value.
8. Record the pH value in a table.
9. Classify each substance as acidic (pH < 7), neutral (pH $= 7$), or basic (pH > 7).

Observation Table

Substance	Colour Change	Estimated pH	Acidic, Neutral or Basic?
Lemon juice			
Vinegar			
Orange juice			
Baking soda			
Milk			
Soap solution			
Toothpaste solution			
Tap water			

Discussion

1. Which substance was the most acidic?
2. Which one was the most basic?
3. Were there any surprises in your results?
4. Why is it important to know the pH of substances we use or consume daily?

IONIC PRODUCT OF WATER

Meaning of Ionic Product of Water

Ionic product of water (Kw) is the product of the concentrations of hydrogen ions [H+] and hydroxide ions [OH−] in pure water at a particular temperature. It shows how much water dissociates (breaks) into hydrogen and hydroxide ions. Even though water mostly stays as a molecule (H_2O), a tiny amount splits like this: $H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$

It expresses the product of the molar concentrations of hydrogen and hydroxide ions at equilibrium:

$$K_c = \frac{[H^+][OH^-]}{[H_2O]}$$

$$K_c [H_2O] = [H^+][OH^-]$$

$$K_c [H_2O] = K_w$$

$$K_w = [H^+][OH^-]$$

K_w is temperature dependent.

At a temperature of 25°C (298K);

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}\ mol^2\,dm^{-6}$$

In pure water, $[H^+] = [OH^-]$.

Thus:

$$[H^+][H^+] = 1 \times 10^{-14}$$

$$[H^+]^2 = 1 \times 10^{-14}$$

$$[H^+] = \sqrt{1 \times 10^{-14}}$$

$$[H^+] = 1 \times 10^{-7}\ mol/dm3$$

$$\therefore [OH^-] = 1 \times 10^{-7}\ mol/dm3$$

As temperature increases, water ionises more, so the value of Kw increases with temperature. See **Table 1.2** for ionic product of water (Kw) at various temperatures.

Table 1.2: Ionic Product of Water (Kw) at Various Temperatures

Temperature (°C)	K_w ($mol^2 \cdot dm^{-6}$)	$pK_w = -\log_{10}(K_w)$
0 °C	1.14×10^{-15}	14.94
10 °C	2.93×10^{-15}	14.53
25 °C	1.00×10^{-14}	14.00
40 °C	2.92×10^{-14}	13.53
50 °C	5.47×10^{-14}	13.26
60 °C	9.61×10^{-14}	13.02
100 °C	5.13×10^{-13}	12.29

Note

- As **temperature increases**, **K_w increases**.
- This means that **water ionises more** at higher temperatures.
- However, the solution can still be **neutral** (i.e., $[H^+] = [OH^-]$) even though pH changes slightly.
- Kw links $[H^+]$ and $[OH^-]$: if one rises, the other falls so that their product remains constant at given temperature.
- Kw is central when calculating pH in acidic or basic solutions.
- You are encouraged to use the link to watch the video on ionic product of water (Kw): https://youtu.be/Xfdj8JCCwI0

Relationship Between pH, pOH and pKw

The relationship between pH, pOH, and pKw is fundamental to understanding acid-base chemistry. It is expressed mathematically as: pH + pOH = pKw

At 25°C, the ionic product of water (Kw) is 1.0×10^{-14}, so:

$pKw = -\log_{10}(Kw)$

$= -\log_{10} 1.0 \times 10^{-14}$

$= 14$

$\therefore pH + pOH = 14$

As temperature increases, Kw increases and pKw decreases, hence the sum of pH and pOH becomes less than 14.

Summary of pH, pOH, and pKw at Different Temperatures:

Condition	Kw ($mol^2 \cdot dm^{-6}$)	pK_w	pH + pOH
At 25°C (neutral)	1.0×10^{-14}	14.00	14.00
At 50°C (neutral)	5.47×10^{-14}	13.26	13.26

Example 1

At 298 K, the pOH of a solution is 12.5. Calculate the pH of the solution. Give your answer to 1 decimal place.

Answer

Step 1: States the correct relation: pH + pOH = 14

pH = 14 − pOH

Given: pOH = 12.5

Step 2: Substitutes correct values into the equation: pH = 14.00 – 12.5 = 1.5

Example 2

At 298 K, the pH of a solution is 10.5. Determine the concentration of hydroxide ions, $[OH^-]$, in the solution. Give your answer in mol/dm^3 to 2 significant figures.

Answer

Step 1: Use the relation pH + pOH = 14 (at 298 K)

Step 2: Calculate the pOH

pOH = 14 - 10.5 = 3.5

Step 3: Use the formula to find $[OH^-]$:

$[OH^-] = 10^{-pOH}$

$= 10^{-3.5}$

Step 4: Calculate $[OH^-]$:

$[OH^-] = 3.16 \times 10^{-4}$ mol/dm^3

$= 3.2 \times 10^{-4}$ mol/dm^3

Example 3

A solution has a hydroxide ion concentration of $[OH^-] = 1.0 \times 10^{-4}$ mol/dm³. Calculate the pOH of the solution.

a. Determine the pH of the solution.

b. State whether the solution is acidic, basic, or neutral.

Answer

a. Calculating pOH

Step 1: Use the formula $pOH = -\log_{10}[OH^-]$

Step 2: Substitute the value: $pOH = -\log_{10}(1.0 \times 10^{-4}) = 4.0$

$pOH = 4.0$

Determining pH

Step 1: Use the relation $pH + pOH = 14$

Step 2: Substitute the value of pOH, $pH = 14 - 4.0 = 10.0$

$pH = 10.0$

b. Since pH = 10.0, which is greater than 7, the solution is basic

Calculation of pH and pOH of Strong Acids and Bases

Strong acids dissociate completely in water to release hydrogen ions (H^+). This means that for monoprotic acids like HCl and HNO_3, the amount of hydrogen ions in the solution is the same as the amount of acid added, because each acid molecule gives one hydrogen ion.

General Steps

1. Write the dissociation equation—it goes to completion for strong species.

$HCl_{(aq)} \rightarrow H^+ + Cl^-$

$H_2SO_4(aq) \rightarrow 2H^+ + SO_4^{2-}$

$NaOH_{(aq)} \rightarrow Na^+ + OH^-$

2. Use stoichiometry to determine $[H^+]$ or $[OH^-]$.

e.g., HCl, HBr, HNO_3

$[H^+] = [\text{acid}]_{\text{initial}}$

$e.g., H_2SO_4$

$2[H^+] = [\text{acid}]_{\text{initial}}$

3. Calculate pH or pOH using
   pH = –log $[H^+]$ or
   pOH = –log $[OH^-]$
4. If you have pOH, convert it to pH using: pH = 14 – pOH

Example 4

What is the pH of 0.0025 mol/dm³ HCl?

Answer

Step 1: HCl is a strong acid, so it dissociates completely in water:
$HCl \rightarrow H^+ + Cl^-$

This means the concentration of hydrogen ions $[H^+]$ is equal to the concentration of HCl. So, $[H^+] = [HCl] = 0.0025$ mol/dm³

Step 2: Use the pH formula, $pH = -\log_{10}[H^+]$

Step 3: Substitute the value, $pH = -\log_{10}(0.0025)$

Step 4: Calculate the value
$pH = -\log_{10}(2.5 \times 10^{-3})$

$pH = -(\log_{10}(2.5) + \log_{10}(10^{-3}))$

$pH = -(0.398 - 3)$

$pH = 2.60$

pH of Very Dilute Strong Acids

When the concentration of a strong acid is less than 1.0×10^{-6} mol/dm³, we must consider the contribution of water to the hydrogen ion concentration.

Why include water's contribution?
Water naturally ionizes slightly: $H_2O \rightleftharpoons H^+ + OH^-$
At 25°C (298 K), this results in: $[H^+] = [OH^-] = 1.0 \times 10^{-7}$ mol/dm³

So, when the acid concentration is below 1.0×10^{-6} mol/dm³, the hydrogen ions from water are significant and must be added to those from the acid.

Example 5

Given: $[H^+]$ from acid $= 1.0 \times 10^{-7}$ mol/dm³

Contribution from water $= 1.0 \times 10^{-7}$ mol/dm³

Total $[H^+] = 1.0 \times 10^{-7} + 1.0 \times 10^{-7} = 2.0 \times 10^{-7}$ mol/dm³

Now calculate the pH:

$pH = -\log_{10}(2.0 \times 10^{-7}) \approx 6.70$

For a diprotic strong acid like H_2SO_4, the concentration of H+ ions is approximately equal to twice the initial concentration of the acid. The same pattern will also apply if the acid is triprotic (Phosphoric acid); it will be triple to initial concentration.

Calculating pOH for Strong Bases

Steps to Calculate pOH for Strong Bases

1. Write the dissociation equation. Example: $NaOH \rightarrow Na^+ + OH^-$
2. Determine $[OH^-]$ from concentration of the strong base. For monoprotic bases like NaOH or KOH, $[OH^-] = [NaOH]$
3. Calculate pOH using: $pOH = -\log_{10}[OH^-]$
4. Convert pOH to pH using the equation: $pH + pOH = 14$

Example 6

What is the pOH of a 0.05 mol/dm³ potassium hydroxide (KOH) solution?

Answer

Step 1: Identify the base and its dissociation.

Potassium hydroxide (KOH) is a strong base, so it dissociates completely in water:

$KOH \rightarrow K^+ + OH^-$

$[OH^-] = [KOH]$

$[OH^-] = 0.05$ mol/dm³

Step 2: Use the pOH formula

$pOH = -\log_{10}[OH^-]$

Step 3: Substitute the value

$pOH = -\log_{10}(0.05)$

Step 4: Calculate the value

$0.05 = 5 \times 10^{-2}$

$\log_{10}(0.05) = \log_{10}(5 \times 10^{-2})$

$= \log_{10}(5) + \log_{10}(10^{-2})$

$= 0.699 - 2 = -1.301$

$pOH = -(-1.301) = 1.30$

"This calculation is valid only when the base concentration is above the hydroxide level of pure water at 298 K (i.e., above 1.0×10^{-7} M). In that case, water's own OH^- contribution is insignificant and can be ignored when calculating pOH."

Example 7

What is the pH of a 1×10^{-8} mol/dm³ sodium hydroxide (NaOH) solution?

Answer

Step 1: Contributors to $[OH^-]$

a. From NaOH dissociation: 1.0×10^{-8} mol/dm³

b. From water autoionisation: 1.0×10^{-7} mol/dm³ (in pure water at 298 K)

Step 2: Add $[OH^-]$ together

$$[OH^-]_{total} = 1.0 \times 10^{-8} + 1.0 \times 10^{-7} = 1.1 \times 10^{-7}\,\text{mol/dm}^3$$

Step 3: Calculate pOH

$pOH = -\log^{10}[OH^-]$

$= -\log^{10}(1.1 \times 10^{-7})$

$= 6.96$

Step 3: Find pH using pH + pOH=14.00

$pH = 14\text{-}pOH$

$= 14\text{-}6.96$

$= 7.04$

Activity 1.2 Exploring the pH Scale

Materials needed

Blank pH scale diagrams (0–14) – one per group, coloured markers or crayons, substance cards with names and pH values (e.g., lemon juice – pH 2, blood – pH 7.4, soap – pH 10), calculators for H^+ concentration conversion

Procedure

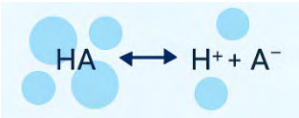
1. Explain how pH and hydrogen ion concentration are related.
2. Explain the formula in your own words.
3. On a large blank pH scale paper, label the regions:
   a. Red for Acidic (pH 0–6)
   b. Green for Neutral (pH 7)
   c. Blue for Basic (pH 8–14)
4. Place each card at the correct position on the scale.
5. Write the approximate $[H^+]$ concentration below the pH points using: $[H+] = 10-pH$
6. Pick one region of the pH scale to present: Acidic, Neutral, or Basic
7. Include
   a. examples of substances in that region
   b. their pH values
   c. their $[H^+]$ concentration (pH 0 → $[H^+] = 1$ M, pH 7 → $[H^+] = 1\times10^{-7}$ M, pH 14 → $[H^+] = 1\times10^{-14}$ M)
   d. why pH matters (e.g., in food, body fluids, cleaning)

Discuss

1. How can knowing pH help us in cooking?
2. Why is it important for farmers or doctors?
3. Why do we use a pH scale instead of saying the exact H^+ concentration?

pK_A and pK_B Ionisation Constant K_a for Weak Acids

Weak acids only partially break apart (dissociate) in water.



This means not all acid molecules produce hydrogen ions (H^+). The process is reversible, so both the acid and ions stay in the solution. Because fewer H^+ ions are formed, weak acids are less acidic than strong acids. The strength of a weak acid is measured using a value called the ionisation constant (Ka).

Ka is the dissociation constant for a weak acid. It measures how much a weak acid breaks apart (dissociates) in water.

Ka is not affected by concentration changes but is affected by temperature. A higher Ka means the acid dissociates more and is therefore a stronger weak acid. Ka is used only for weak acids, not strong acids (which fully dissociate).

Table 1.3: Ka and pKa values of monoprotic weak acids at a temperature of 298K

Weak Monoprotic Acid	K_a	pK_a
Acetic acid (CH_3COOH)	1.80×10^{-5}	4.74
Formic acid ($HCOOH$)	1.80×10^{-4}	3.75
Hydrofluoric acid (HF)	6.60×10^{-4}	3.17
Benzoic acid (C_6H_5COOH)	6.30×10^{-5}	4.2
Propanoic acid (CH_3CH_2COOH)	1.30×10^{-5}	4.87
Chloroacetic acid ($ClCH_2COOH$)	1.40×10^{-3}	2.86
Nitrous acid (HNO_2)	4.50×10^{-4}	3.35
Hydrazoic acid (HN_3)	1.90×10^{-5}	4.72
Hypochlorous acid ($HOCl$)	3.00×10^{-8}	7.4

Table 1.4: Ka and pKa values of polyprotic weak acids at a temperature of 298K

Weak Polyprotic Acid	K_a	pK_a
Carbonic acid (H_2CO_3) – 1st dissociation	4.30×10^{-7}	6.37
Carbonic acid (H_2CO_3) – 2nd dissociation	5.60×10^{-11}	10.25
Phosphoric acid (H_3PO_4) – 1st dissociation	7.10×10^{-3}	2.15
Phosphoric acid (H_3PO_4) – 2nd dissociation	6.30×10^{-8}	7.2
Phosphoric acid (H_3PO_4) – 3rd dissociation	4.50×10^{-13}	12.35
Sulfurous acid (H_2SO_3) – 1st dissociation	1.40×10^{-2}	1.85
Sulfurous acid (H_2SO_3) – 2nd dissociation	6.30×10^{-8}	7.2

Weak Polyprotic Acid	K_a	pK_a
Oxalic acid ($H_2C_2O_4$) – 1st dissociation	5.40×10^{-2}	1.27
Oxalic acid ($H_2C_2O_4$) – 2nd dissociation	5.40×10^{-5}	4.27

Deriving K_a Expression for a Weak Monoprotic Acid

Let's say the weak monoprotic acid is HA. It reacts with water like this:

$HA\ (aq) \leftrightarrows H^+(aq) + A^-(aq)$

HA = weak acid

H^+ = hydrogen ion

A^- = conjugate base

From the reaction, we apply the equilibrium law to get:

$$K_a = \frac{[H^+]_{eq}[A^-]_{eq}}{[HA]_{eq}}$$

This is the acid dissociation constant (Ka).

Use concentration assumptions

Let the initial concentration of HA be 'c' mol/dm³.

Let '*x*' mol/dm³ of it dissociate at equilibrium.

Then at equilibrium:

$[HA] = c - x$;

$[H^+] = x$;

$[A^-] = x$

Substitute into the Ka Expression

$$K_a = \frac{x.x}{c - x}$$

$$= \frac{x^2}{c - x}$$

When dealing with **weak acids**, only a **small fraction** of the acid molecules dissociates into ions. This means:

$x \ll c \text{ or } c \gg x$

$\Longrightarrow c - x = c$

So, in the expression:

$$K_a = \frac{x^2}{c}$$

$$x^2 = K_a \times c$$

$$x = \sqrt{K_a \times c}$$

$$[H^+] = x$$

$$\therefore [H^+] = \sqrt{K_a \times c}$$

To determine the pH of a weak monoprotic acid, take the negative logarithm to the base ten of both sides of the equation

$$\log[H^+] = -\log\left(K_a \times c\right)^{\frac{1}{2}}$$

$$-\log[H^+] = -\tfrac{1}{2}\log K_a - \tfrac{1}{2}\log c$$

But $-\log[H^+] = pH$ and $-\log K_a = pK_a$

$$pH = \tfrac{1}{2}pK_a - \tfrac{1}{2}\log c$$

Example 8

Calculation: $[H^+]$, pH, and pKa of a weak acid

Given:

Weak acid: Acetic acid (CH_3COOH)

Initial concentration (c) = 0.10 mol/dm³

$Ka = 1.8 \times 10^{-5}$

Step 1: Write the dissociation equation $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$

Step 2: Set up the Ka expression

Let '*x*' be the amount that dissociates:

At equilibrium:

$[H^+] = x$

$[CH_3COO^-] = x$

$[CH_3COOH] = 0.10 - x$

$$K_a = \frac{[H^+]_{eq}[CH_3COO^-]_{eq}}{[CH_3COOH]_{eq}}$$

$$= \frac{x^2}{0.10 - x}$$

Since x is small, assume $0.10 - x \approx 0.10$

Step 3: Substitute and solve for x

$$K_a = \frac{x^2}{0.10}$$

$$x^2 = 1.8 \times 10^{-5} \times 0.10$$

$$= 1.8 \times 10^{-6}$$

$$x = \sqrt{1.8 \times 10^{-6}} = 1.34 \times 10^{-3}$$

So, $[H^+] = 1.34 \times 10^{-3}$ mol/dm³

Step 4: Calculate pH

$$pH = -\log[H^+]$$

$$= -\log(1.34 \times 10^{-3})$$

$$= 2.87$$

Step 5: Calculate pKa

$$pK_a = -\log(K_a)$$

$$= -\log(1.8 \times 10^{-5})$$

$$= 4.74$$

Ionisation Constant K_b for Weak Bases

Weak bases only partially break apart (dissociate) in water. This means not all the base molecules produce hydroxide ions (OH-). The process is reversible, so both the base molecules and ions stay in the solution. Because fewer OH- ions are formed, weak bases are less basic than strong bases. The strength of a weak base is measured using a value called the ionisation constant (Kb).

Kb is the dissociation constant for a weak base. It measures how much a weak base break apart (dissociates) in water. Kb is not affected by concentration changes but is affected by temperature. A higher Kb means the acid dissociates more and is therefore a stronger weak base. Kb is used only for weak bases, not strong bases (which fully dissociate).

Table 1.5: Kb and pKb values of weak bases at a temperature of 298K

Weak Base	K_b (mol/dm³)	pK_b
Ammonia (NH_3)	1.80×10^{-5}	4.74
Methylamine (CH_3NH_2)	4.40×10^{-4}	3.36
Ethylamine ($C_2H_5NH_2$)	5.60×10^{-4}	3.25

Weak Base	K_b (mol/dm³)	pK_b
Aniline ($C_6H_5NH_2$)	4.30×10^{-10}	9.37
Hydroxylamine (NH_2OH)	1.10×10^{-8}	7.96
Pyridine (C_5H_5N)	1.70×10^{-9}	8.77
Triethylamine ($(C_2H_5)_3N$)	6.30×10^{-5}	4.2
Dimethylamine ($(CH_3)_2NH$)	5.40×10^{-4}	3.27
Morpholine ($O(CH_2CH_2)_2NH$)	6.30×10^{-4}	3.2

Deriving K_b Expression for Weak Bases

Let's say the weak monoprotic or monoacidic base BOH, it reacts with water like this:

$BOH\ (aq) \leftrightarrows B^+(aq) + OH^-(aq)$

BOH = weak base

OH^- = hydroxide ion

B^+ = conjugate acid

From the reaction, we apply the equilibrium law to get:

$$K_b = \frac{[B^+]_{eq}[OH^-]_{eq}}{[BOH]_{eq}}$$

This is the base dissociation constant (Kb).

Use concentration assumptions

Let the initial concentration of BOH be 'c' mol/dm³.

Let '*x*' mol/dm³ of it dissociate at equilibrium.

Then at equilibrium:

$[BOH] = c - x$;

$[OH^-] = x$;

$[B^+] = x$

Substitute into the Kb Expression

$$K_b = \frac{x.x}{c - x}$$

$$= \frac{x^2}{c - x}$$

When dealing with **weak base**, only a **small fraction** of the base molecules dissociates into ions. This means:

$x \ll c \; or \; c \gg x$

$\Longrightarrow c - x = c$

So in the expression:

$K_b = \frac{x^2}{c}$

$x^2 = K_b \times c$

$x = \sqrt{K_b \times c}$

$[OH^-] = x$

$\therefore [OH^-] = \sqrt{K_b \times c}$

To determine the pOH of a weak base, take the negative logarithm to the base ten of both sides of the equation

$\log[OH^-] = -\log\left(K_b \times c\right)^{\frac{1}{2}}$

$-\log[OH^-] = -\frac{1}{2}\log K_b - \frac{1}{2}\log c$

But $-\log[OH^-] = pOH$ and $-\log K_b = pK_b$

$pOH = \frac{1}{2}pK_b - \frac{1}{2}\log c$

Example 9

Calculation: $[OH^-]$, pOH, pH, and pKb of a weak base

Example: Ammonia (NH_3)

Given:

Initial concentration of $NH_3 = 0.20$ mol/dm³

$Kb = 1.8 \times 10^{-5}$

Step 1: Write the base dissociation equation

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

Step 2: Set up the Kb expression

Let x be the amount that dissociates:

At equilibrium:

$[OH^-] = x$

$[NH_4^+] = x$

$[NH_3] = 0.10 - x$

$$K_b = \frac{[NH^{4+}]_{eq}[OH^-]_{eq}}{[NH^3]_{eq}}$$

$$= \frac{x^2}{0.20 - x}$$

Assume x is small:

$$K_b = \frac{x^2}{0.20}$$

Step 3: Substitute Kb and solve for x

$$1.8 \times 10^{-5} = \frac{x^2}{0.20}$$

$$x^2 = 1.8 \times 10^{-5} \times 0.20$$

$$= 3.6 \times 10^{-6}$$

$$x = \sqrt{3.6 \times 10^{-6}} = 1.90 \times 10^{-3}$$

So, $[OH^-] = 1.90 \times 10^{-3}$ mol/dm³

Step 4: Calculate pOH

$$pOH = -\log[OH^-]$$

$$= -\log(1.90 \times 10^{-3})$$

$$= 2.72$$

Step 5: Calculate pH

$$pH = 14 - pOH$$

$$= 14 - 2.72$$

$$= 11.28$$

Step 6: Calculate pKb

$$pKb = -\log(Kb)$$

$$= -\log(3.6 \times 10^{-5})$$

$$= 4.44$$

Activity 1.3 Understanding Partial Dissociation of Weak Acids and Bases

Materials needed

Two labelled beakers (A and B), Universal indicator solution or paper, Vinegar (weak acid), Hydrochloric acid (strong acid), distilled water

Procedure

1. **a.** Discuss acids and bases in terms of Arrhenius and Brønsted-Lowy concepts.

b. Name everyday acids and bases that you encounter.

c. Discuss the difference between strong and weak acids or bases.

2. **a.** Label Beaker A as 'Weak Acid (Vinegar)' and Beaker B as 'Strong Acid (HCl)'.

b. Pour equal volumes (e.g. 50 cm3) of vinegar into Beaker A and HCl into Beaker B.

c. Add a few drops of universal indicator to both solutions.

d. Observe and compare the colour change in each beaker.

e. Record the pH of each solution

Use the table below to record your observations during the experiment comparing vinegar (a weak acid) and HCl (a strong acid).

Beaker Label	Type of Acid	Colour with Indicator	Measured pH
A	Weak Acid (Vinegar)		
B	Strong Acid (HCl)		

i. What do the colours and pH values tell us about ion concentration?

ii. Repeat the experiment using weak and strong bases (e.g., Ammonia versus Sodium hydroxide).

3. **a.** What do you notice about the pH of vinegar compared to hydrochloric acid?

b. Why is the pH of vinegar higher than HCl even though both are acids?

c. What does this tell us about the number of hydrogen ions in each solution?

d. How does this relate to the concept of partial break apart (dissociation)?

e. Would you expect a similar result with weak and strong bases? Why or why not?

4. Consider a generic weak acid HA that dissociates as follows: $HA \rightleftharpoons H^+ + A^-$

a. Write the ionisation constant expression

b. Derive an expression for pH of weak acids in terms of C and *x*.

Assume the initial concentration of HA is 'C' mol/dm³.

Let '*x*' be the concentration of H^+ at equilibrium

Activity 1.4 K_a, K_b, pK_a, pK_b & pH Calculations

Materials needed: Calculator

Procedure

1. Use the formula pKa = -log (Ka) and pKb = -log (Kb) to calculate the missing values in the table below.

2. Use a calculator and write your steps.

Substance	K_a or K_b (given)	pK_a or pK_b (your answer)
Benzoic acid	$K_a = 6.3 \times 10^{-5}$	
Dimethylamine	$K_b = 5.4 \times 10^{-4}$	

3. Calculate the pH of a 0.100 moldm-3 acetic acid solution. Use the steps below:

a. Write the equation: $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

b. Set up Ka expression:

$$K_a = \frac{[H^+]^2}{[CH_3COOH]}$$

c. Assume $[H^+] = x$, solve $= \frac{x^2}{0.10}$

d. Find x (which is $[H^+]$), then pH = -log(x)

4. Calculate the pH of 0.1 moldm-3 ammonia. Use $Kb = \frac{x^2}{0.10}$

a. Find $x = [OH^-]$

b. Calculate pOH = -log $[OH^-]$

c. Use pH = 14 - pOH

BUFFER SOLUTIONS AND BUFFER ACTIONS

Buffer Solutions and their Composition

What is a Buffer Solution?

A buffer solution is a special mixture that resists changes in pH when small amounts of an acid or a base is added. It helps to keep the solution's pH stable, which is important in many chemical and biological systems.

What is it made of?

A buffer solution usually contains

1. A weak acid and its salt (conjugate base) (acidic buffer), e.g., ethanoic acid (CH_3COOH) and sodium ethanoate (CH_3COONa). *Or*
2. A weak base and its salt (conjugate acid (basic buffer), e.g., ammonia (NH_3) and ammonium chloride (NH_4Cl)

Why is it important?

Buffer solutions are used to

1. Keep the pH of our blood stable
2. Control pH in chemical reactions
3. Maintain pH in soaps, shampoos, and medicines

Buffer Action

Buffer action is how a buffer solution works to keep the pH steady, even when we add a little amount of acid or base.

Acidic Buffer (e.g., Vinegar and its Salt – CH_3COOH/CH_3COONa)

An acidic buffer is a mix of a weak acid and its salt (which has a base part).

1. If you add acid (H^+), the salt (CH_3COONa) soaks up the extra acid and turns it into the weak acid CH_3COOH. This stops the pH from dropping too much.

   Reaction: $H^+ + CH_3COO^- \rightarrow CH_3COOH$
2. If you add base (OH^-), the weak acid reacts with the base to make water and salt. This stops the pH from rising too much.

   Reaction: $OH^- + CH_3COOH \rightarrow CH_3COO^- + H_2O$

Basic Buffer (e.g., Ammonia and its salt – NH_3/NH_4Cl)

A basic buffer is a mixture of a weak base and its salt (which has an acid part).

1. If you add acid (H^+), the weak base reacts with the acid to form the salt (NH_4^+). This stops the pH from dropping. Reaction: $H^+ + NH_3 \rightarrow NH_4^+$
2. If you add base (OH^-), the acid part (NH_4^+) reacts with the base to make NH_3 and water. This stops the pH from rising too much. Reaction: $OH^- + NH_4^+ \rightarrow NH_3 + H_2O$

Real-life Analogies to Understand Buffer Action

1. **Body Temperature Analogy**

   When you are too hot, your body sweats to cool you down. When you are too cold, your body shivers to warm you up. It keeps your body temperature stable — just like a buffer keeps pH stable!

   a. Sweating = the buffer removing excess acid

   b. Shivering = the buffer adding acid when base is added

2. **Cooking Analogy** (Jollof Rice Example)

   You accidentally add too much salt → You add a potato to soak it up or more rice to dilute it. It keeps the food flavour balanced — just **like** a buffer keeps the pH balanced!

   a. Potato = absorbs extra salt → like the conjugate base absorbing acid

   b. More rice or seasoning = balances taste → **like** the acid part of the buffer reacting with base

Calculation of pH of a Buffer Solution

Acidic Buffer

For an acidic buffer (weak acid + salt), we use the Henderson-Hasselbalch equation:

$$pH = pKa + \log\frac{[\text{Salt } (A^-)]}{[\text{Acid } (HA)]}$$

Where

pKa is the negative log of Ka (acid dissociation constant)

[Salt] is the concentration of the conjugate base (A^-)

[Acid] is the concentration of the weak acid (HA)

Example

Calculate the pH of a buffer made from:

0.10 mol dm-3 acetic acid (CH_3COOH)

0.10 mol dm-3 Sodium acetate (CH_3COONa)

Ka of acetic acid = 1.8×10^{-5}

Step 1: Find pKa

pKa = -log (1.8×10^{-5}) ≈ 4.74

Step 2: Use the formula

$$pH = pKa + \log \frac{[Salt\ (A^-)]}{[Acid\ (HA)]}$$

$$pH = 4.74 + \log \frac{0.10}{0.10} = 4.74$$

Basic Buffer

For a basic buffer (weak base + salt), use the formula:

$$pOH = pKb + \log \frac{[Salt\ B\,H^+]}{[Base\ B]}$$

Then: pH = 14 - pOH

Example 9

Calculate the pH of a buffer solution that contains:

0.20 mol dm-3 acetic acid (CH_3COOH)

0.10 mol dm-3 sodium acetate (CH_3COONa)

Ka of acetic acid = 1.8×10^{-5}

Answer

Step 1: Calculate pKa

pKa = -log (1.8×10^{-5}) ≈ 4.74

Step 2: Use the Henderson-Hasselbalch equation

$$pH = pKa + \log \frac{[Salt\ (A^-)]}{[Acid\ (HA)]}$$

$$pH = 4.74 + \log \frac{0.10}{0.20}$$

$$pH = 4.74 + \log(0.50)$$

$pH = 4.74 + (-0.30)$

$pH = 4.44$

Example 10

Calculate the pH of a buffer solution that contains:

0.30 mol dm-3 ammonia (NH_3)

0.20 mol dm-3 ammonium chloride (NH_4Cl)

Kb of ammonia $= 1.8 \times 10^{-5}$

Answer

Step 1: Calculate pKb

$pKb = -\log (1.8 \times 10^{-5}) \approx 4.74$

Step 2: Calculate pOH

$pOH = pKb + \log \frac{[Salt\ ]}{[Base]}$

$= 4.74 + \log \frac{0.20}{0.30}$

$= 4.74 + \log(0.667)$

$= 4.74 + (-0.18)$

$= 4.56$

Step 3: Calculate pH

$pH = 14 - pOH = 14 - 4.56$

$pH \approx 9.44$

Activity 1.5 Collaborative Learning

Topics

1. Buffer Solutions and Their Behaviour
2. pH Calculations

Overview of the Activity

1. Organise yourself into groups of 3 or 4.
2. Each group member assumes a specific role- Researcher: Gathers key facts and relevant examples

   Note-taker: Records important information and group ideas

   Timekeeper: Keeps the group on task and on time

   Presenter: Shares the group's findings with the class
3. Topic Assignment and Research
   a. Group A: Investigates buffer solutions, including their components, behaviour, and real-life applications.
   b. Group B: Studies pH calculations using the Henderson-Hasselbalch equation, including how to calculate pH and component concentrations.

   Groups use chemistry textbooks, printed guides, and teacher-approved resources for your research.
4. Collaboration Guidelines
   a. Listen actively and respect one another's ideas
   b. Share tasks and support one another
   c. Provide constructive feedback
5. Compile your findings and create a short, creative presentation. This can be a: poster, diagram, slide show or oral report.
6. Present your work to the class
7. Reflection
   a. What is a buffer solution?
   b. How does a buffer help maintain pH?
   c. How do you calculate the pH of a buffer?
   d. What did you enjoy or learn from your group experience?

PREPARATION OF BUFFER SOLUTIONS

Buffer Solution is made by combining

1. A weak acid and its conjugate base (salt) → Acidic Buffer
2. A weak base and its conjugate acid (salt) → Basic Buffer

Acidic Buffer Preparation (e.g., CH_3COOH / CH_3COONa)

Method A

Mixing a Weak Acid and its Salt

1. Mix a measured amount of acetic acid (CH_3COOH) with sodium acetate (CH_3COONa).
2. Dissolve both in distilled water.
3. Adjust their concentrations based on the required pH.
4. Use the Henderson-Hasselbalch equation:

$$pH = pKa + \log\frac{[Salt\ (A^-]}{[Acid\ (HA)]}$$

Method B

Reacting a Weak Acid with a Strong Base

1. Add NaOH (strong base) gradually to CH_3COOH (excess).
2. This neutralizes some of the CH_3COOH to form CH_3COONa.
3. The final solution contains both CH_3COOH and CH_3COONa — forming a buffer.

Basic Buffer Preparation (e.g., NH_3 / NH_4Cl)

Method A: Mixing a Weak Base and Its Salt

1. Mix a known concentration of ammonia (NH_3) with ammonium chloride (NH_4Cl).
2. Dissolve both in water to make the basic buffer.
3. Use the Henderson-Hasselbalch equation:

$$pOH = pK_b + \log\left(\frac{[Salt]}{[Base]}\right)$$

Method B: Reacting a Weak Base with a Strong Acid

1. Add HCl (strong acid) slowly to NH_3 (excess)
2. The reaction forms NH_4Cl, resulting in a buffer solution containing both NH_3 and NH_4^+.

Applications of Buffer Solutions in Everyday Life

Buffer solutions are like helpful 'pH bodyguards'. They stop things from becoming too acidic or too basic, keeping things balanced. Here are some fun and simple ways buffers help us every day:

1. **In Our Body (Especially in Our Blood)**

   Our blood must stay at a steady pH (around 7.4) to keep us healthy. Buffers in the blood prevent it from becoming too acidic or too basic when we eat, run, or get sick. Imagine buffers as tiny superheroes protecting our blood!
2. **In Toothpaste**

   When we eat sweet food, our mouth becomes more acidic, which can hurt our teeth. Toothpaste has buffers that help balance the pH in our mouth and protect our teeth from decay. Buffers help us to keep our smiles strong and healthy! (no mouth odour)
3. **In Science Experiments**

   In school labs, we mix chemicals. The pH must stay steady so the results are correct. Buffers keep the pH from changing too much, so we get the right results. They make science safer and more accurate!
4. **In Food and Drinks**

   Some drinks and foods need the right pH to taste good and stay fresh. Buffers help keep the pH steady so the food doesn't spoil quickly or taste bad. For example, soft drinks have buffers to balance their fizzy taste.
5. **In Plants and Soil**

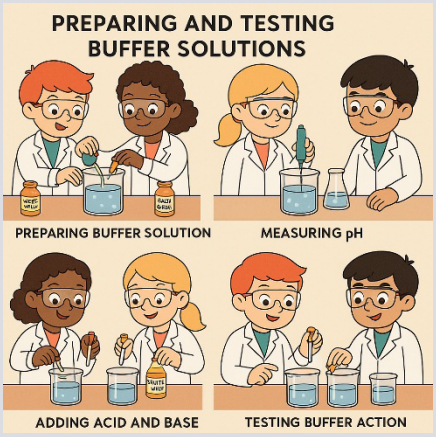
   Plants need healthy soil to grow. Buffers in the soil help keep the pH just right so the plants can absorb nutrients. Healthy soil = happy plants!
6. **In Fish Tanks (Aquariums)**
7. Fish can get sick if the water is too acidic or too basic. Buffers keep the water just right for the fish to live happily.

Activity 1.6 Preparing and Testing Buffer Solutions

Materials needed

1. Ethanoic acid solution (CH_3COOH) (weak acid)
2. Sodium ethanoate, CH_3COONa (salt of the weak acid)
3. Ammonium hydroxide, NH4OH (weak base)
4. Ammonium chloride, NH4Cl (salt of the weak base)

5. pH paper or universal indicator
6. Beakers, measuring cylinders, glass rods
7. Distilled water
8. Safety goggles and gloves



Step 1

a. Put on your goggles and gloves.

b. Mix a weak acid (0.100 mol dm-3) (like vinegar or ethanoic acid) with its salt (0.10 mol dm-3) (like sodium ethanoate) in a beaker.

c. Stir the solution gently using a glass rod.

This mixture creates a buffer solution that can resist pH changes.

Step 2: Measuring pH

a. Use pH paper or a universal indicator to measure the solution's pH.

b. Record the value.

You are checking how acidic or basic the buffer is before testing it.

Step 3: Adding Acid and Base

a. Divide the buffer into 3 small beakers

b. Add a few drops of dilute HCl (acid) to the first one.

c. Add a few drops of dilute NaOH (base) to the second one.

d. Leave the third as a control (no change).

You are simulating what happens when an acid or base is added to a buffer.

Step 4: Testing Buffer Action

a. Measure the pH of all solutions in the three beakers again.

b. Compare the pH changes.

Buffers resist pH changes. You should see only small changes in pH when acid or base is added!

Repeat Steps 1-4, using a mixture of 0.100 mol dm-3 NH3(aq) and 0.1 mol dm-3 NH4Cl. This is a basic buffer.

Discussion Questions

1. Which sample had the biggest change in pH?

2. Did your buffer solution work well?

3. Why do you think some buffers work better than others?

Activity 1.7 Applications of Buffer Systems

This inquiry-based activity helps you explore real-world applications of buffer systems in biological and industrial contexts. Learners will work in groups, research a buffer system, and present their findings using guiding questions and a video resource.

Steps

1. Use the following video link to begin your investigation:

https://www.youtube.com/watch?app=desktop&v=10Al4Z7W_zI

The video explores how buffer systems work and their importance in biological and industrial environments.

2. Form small groups and choose a buffer system to investigate from one of these categories:

a. Biological buffers (e.g., blood, intracellular fluid, saliva)

b. Industrial buffers (e.g., food production, pharmaceuticals, agriculture)

3. Find answers to the following questions:

a. What are the components of your buffer system?

b. What is the optimal pH range maintained by the buffer?

c. What would happen if the buffer system failed?

d. How does this buffer system relate to the Henderson-Hasselbalch equation?

4. Group Presentations and Class Discussion

After all groups have presented, engage in a comparative discussion. Focus on:

a. Common features of all buffer systems

b. How buffer composition matches the required pH range

c. The effects of buffer failure in different systems

d. Innovations in medicine and industry made possible through buffer chemistry

5. Design a Buffer System

Choose one of the following scenarios and propose a buffer system for it:

a. A sports drink that maintains pH for optimal electrolyte absorption

b. A soil additive that stabilises pH for specific crops

c. A pharmaceutical formulation that keeps a safe pH inside the body

Activity 1.8 Exploring Acids, Bases, pH, and Buffer Systems

Materials needed

pH paper, beakers, measuring cylinders, household substances (e.g., vinegar, baking soda solution, lemon juice, soap solution), weak acids and bases (e.g., acetic acid, ammonia), indicators, buffer salts (e.g., CH_3COONa, NH_4Cl), and distilled water.

Steps

1. Comparing Household Substances by pH

a. Collect common substances (e.g., lemon juice, baking soda, vinegar, soap solution).

b. Use pH paper to determine the pH of each and record.

c. Plot the results on a - pH scale.

d. What makes a substance acidic or basic, and identify neutral substances.

2. Understanding Acid/Base Strength and Dissociation

a. Classify the tested substances as strong acids, weak acids, neutral, weak bases, or strong bases based on pH values.

b. Write dissociation equations for:

i. Strong acid (e.g., HCl)

ii. Weak acid (e.g., CH_3COOH)

iii. Weak base (e.g., NH_3)

c. Compare extent of dissociation and relate it to pH values from Step 1.

3. Performing Ka, Kb, pKa, pKb and pH Calculations

Given values of Ka(CH_3COOH) = 1.8×10^{-5} and Kb (NH3) = 1.8×10^{-5}, calculate:

a. pKa and pH of CH_3COOH

b. pKb and pH of NH3

4. Preparing and Testing Buffer Solutions.

a. Prepare buffer solutions using:

i. Weak acid and its conjugate base (e.g., CH_3COOH + CH_3COONa)

ii. Weak base and its conjugate acid (e.g., NH_3 + NH_4Cl)

b. Test the buffer's ability to resist pH change by adding small amounts of HCl or NaOH.

5. Record and analyse results.

6. Exploring Real-Life Applications of Buffer Systems. Name buffer systems used in:

a. biological systems

b. industry

c. agriculture (soil pH control)

EXTENDED READING

1. Brown, LeMay, Bursten & Murphy (2017). Chemistry: The Central Science (14th Edition), Chapter 16: Acid–Base Equilibria.
2. Khan Academy: Interactive lessons on pH and pOH – https://www.khanacademy.org
3. Silberberg, M. S. & Amateis, P. G. (2017). Chemistry: The Molecular Nature of Matter and Change (7th Edition), Chapter on Acids and Bases.
4. Zumdahl, S. S. & Zumdahl, S. A. (2014). Chemistry (9th Edition), Sections 14.3–14.6.

REVIEW QUESTIONS

1. **a.** Define the term pH.

b. What is the mathematical relationship between pH and hydrogen ion concentration?

c. State the range of values on the pH scale and identify the type of solutions they represent.

i. $pH < 7$

ii. $pH = 7$

iii. $pH > 7$

2. A solution has a hydrogen ion concentration of $[H^+] = 3.2 \times 10^{-4}$ mol/dm³.

a. Calculate the pH and pOH of the solution.

b. State whether the solution is acidic, basic, or neutral.

c. Explain briefly the importance of this acidity level in everyday use.

d. How does temperature affect the ionic product of water (Kw)?

3. The pH of a solution is 5.2.

a. Calculate the concentration of hydrogen ions, $[H^+]$, in mol/dm³.

b. Comment on how the hydrogen ion concentration would change if the pH decreased to 4.2.

c. Give **one** example of a substance that could have this pH value.

4. A sample of a basic solution has a hydroxide ion concentration of $[OH^-] = 1.0 \times 10^{-3}$ mol/dm³.

a. Calculate the pOH of the solution.

b. Determine the pH of the solution at 25°C.

c. Explain the significance of the equation $pH + pOH = 14$. When is it valid?

d. Describe one real-life use of a basic solution and explain why its pH is suitable for that use.

5. Lemon juice has a pH of 2.5, while blood has a pH of 7.4.

a. Calculate how many times greater the hydrogen ion concentration is in lemon juice compared to blood.

b. Describe how this large difference affects their properties and uses.

c. Explain why even a small change in blood pH can be harmful to the body.

6. You are designing a poster to educate your community on the importance of pH in everyday life.

a. Identify and describe three different areas of life where pH plays a key role.

b. For each area, state:

i. a common substance involved

ii. its pH value

iii. why its pH matters in that context

c. Discuss how incorrect pH levels can lead to problems in any one of the areas identified.

7. The table below shows the pH values of five household substances:

Substance	pH Value
Vinegar	3.0
Baking soda	9.0
Lemon juice	2.0
Milk	6.5
Ammonia solution	11.0

a. Arrange the substances in increasing order of acidity.

b. Which substance is closest to neutral pH?

c. Suggest a suitable universal indicator colour that might be observed for:

i. Lemon juice

ii. Baking soda

d. Why might ammonia solution be unsafe for use on skin?

8. A solution is prepared by mixing 3.60×10^{-3} mol NaOH with 5.95×10^{-4} mol HCl, then diluted to 1.00 L.

a. Determine the pH of the resulting solution.

b. Explain why direct use of $pH = -\log[H^+]$ would be incorrect before evaluating the reaction.

9. The pH of soil is measured at 8.42.

a. Calculate the $[H^+]$ and $[OH^-]$.

b. Suggest two practical ways farmers could adjust this soil pH if it were too high (basic).

10. **a.** Explain why using the pH scale is more practical than giving hydrogen ion concentrations in everyday life.

b. Describe how pH knowledge is useful in two of these areas: cooking, gardening, medicine, environmental water monitoring, or cleaning. Explain why specific pH ranges matter in each.

11. Chemical runoff from nearby farmland caused the pH of a pond to drop from 7.8 to 6.4.

a. Calculate the change in hydrogen-ion concentration between the two pH values.

b. Explain the potential effects of this pH change on fish and aquatic plant life.

c. Propose two sustainable farming practices that could prevent such acidification of water bodies.

Multiple Choice Questions

1. What does Ka represent in acid-base chemistry?

A. Acid boiling point

B. Base dissociation constant

C. Acid dissociation constant

D. Activation energy

2. Which of the following correctly describes pKa?

A. The negative logarithm of the acid concentration

B. The product of Ka and Kb

C. The negative logarithm of Ka

D. The inverse of pH

3. Which statement is TRUE about strong acids?

A. They have high pKa values

B. They have low Ka values

C. They ionize partially in solution

D. They have low pKa values

4. The Kb of ammonia is 1.8×10^{-5}. What is its pKb?

A. 5.0

B. 4.74

C. 6.22

D. 9.2

5. What is the relationship between Ka and the strength of a weak acid?

A. Higher Ka = weaker acid

B. Lower Ka = stronger acid

C. Higher Ka = stronger acid

D. Ka has no effect on acid strength

6. Which pair of values would most likely represent a strong base?

A. Low Kb, high pKb

B. High Kb, low pKb

C. Low Kb, low pKb

D. High Ka, high pKb

7. A solution has $[H^+] = 1 \times 10^{-6}$ mol/dm³. What is its pH?

A. 6

B. 7

C. 8

D. 10

8. The Ka of acetic acid is 1.8×10^{-5}. Calculate its pKa.

A. 5.00

B. 4.74

C. 2.34

D. 3.76

9. If the pKa of a weak acid is 4.75, what can you infer about the pH of a buffer solution when $[A^-] = [HA]$?

A. pH < 4.75

B. pH = 4.75

C. pH > 4.75

D. pH = 7

10. A weak base has a Kb of 3.2×10^{-6}. What is its pKb?

A. 5.5

B. 6.5

C. 4.0

D. 3.0

Essay-type Questions

1. What is a buffer solution?

2. What are the two main components of an acidic buffer?

3. What is the function of a buffer solution?

4. How does an acidic buffer neutralise added acid?

5. How does an acidic buffer neutralise added base?

6. What is the Henderson-Hasselbalch equation?

7. Calculate the pH of a buffer made by mixing 0.2 M CH_3COOH and 0.1 M CH_3COONa. Ka of $CH_3COOH = 1.8 \times 10^{-5}$

8. Explain why buffer solutions are important in biological systems.

9. Calculate the pH of a buffer prepared by mixing 0.3 M NH_3 and 0.2 M NH_4Cl. Kb of $NH_3 = 1.8 \times 10^{-5}$

10. Design a buffer solution with a target pH of 5.0. Choose an appropriate weak acid and justify your choice.

11. What is the formula for calculating the pH of an acidic buffer solution?

12. Calculate the pH of a buffer solution containing 0.1 M acetic acid (Ka = 1.8×10^{-5}) and 0.1 M sodium acetate.

13. Calculate the Ka of a weak acid if a buffer made from it and its salt has a pH of 5.0 and [Salt]/[Acid] = 2.

14. You want to prepare a buffer solution with pH = 9.25 using ammonia (Kb = 1.8×10^{-5}) and ammonium chloride. What should be the ratio [Base]/[Salt]?

15. Describe how you would prepare 250 mL of a buffer solution with pH = 4.76 using acetic acid (Ka = 1.8×10^{-5}) and sodium acetate.

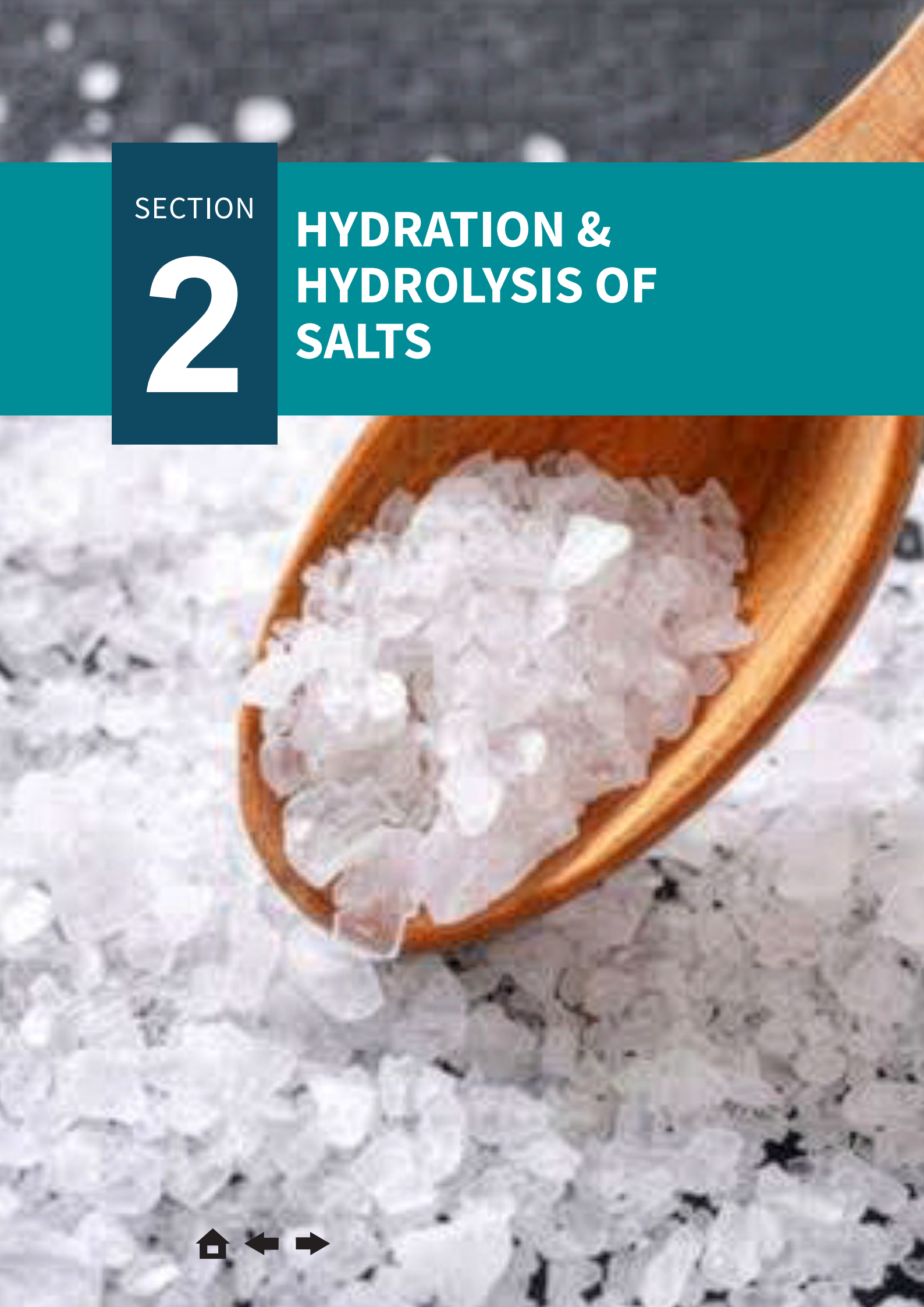
16. Design a buffer system to maintain pH 7.4 for a biochemical experiment. Choose appropriate components and justify your choice.

17. Explain the impact of dilution on the pH and buffer capacity of a solution made from a weak acid and its conjugate base.

SECTION

2

HYDRATION & HYDROLYSIS OF SALTS



PHYSICAL CHEMISTRY

Equilibria – 2

INTRODUCTION

In this section, you will learn what happens when salts dissolve in water. The two main ideas are hydration and hydrolysis. By the end of the section, you will be able to:

- describe what hydration is and what affects it.
- know how to tell if a salt solution will be acidic, basic, or neutral.

KEY IDEAS

- Hydration (of salts) is when water molecules surround and stick to the ions of a salt when it dissolves in water.
- Hydrolysis (of salts) is when a salt reacts with water to make the solution either acidic or basic.

HYDRATION AND HYDROLYSIS OF SALTS

Hydration of Salts

What is Hydration of Salts?

When you put a salt in water, it breaks apart into tiny, charged pieces called ions. These ions do not float around by themselves — water molecules quickly surround them. This process is called hydration.

When you stir table salt into water, you cannot see it anymore because the salt ions are now hydrated — water has surrounded them.

Example using Sodium chloride (NaCl)

$NaCl\ (s) \rightarrow Na^+\ (aq) + Cl^-\ (aq)$

This equation means that solid salt (NaCl) breaks apart into sodium ions (Na^+) and chloride ions (Cl^-) in water. These ions are now hydrated, meaning they are surrounded by water molecules.

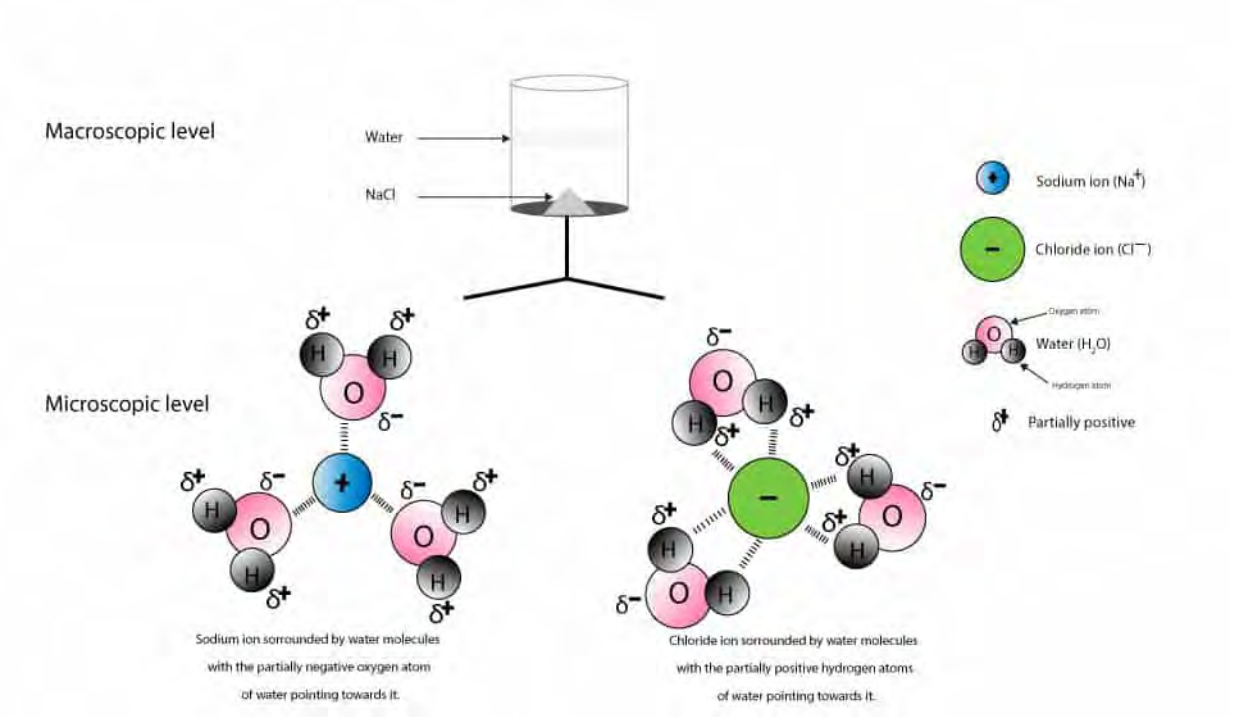


Figure 2.1: Dissolution of NaCl in H2O

Hydrolysis of Salt

Hydrolysis of salts happens when a salt dissolves in water and the ions it forms react with water, changing the pH of the solution — making it either acidic or basic.

What Causes Hydrolysis?

Salts are made from acids and bases. If the salt comes from a strong acid and strong base (like NaCl), it stays neutral. If the salt comes from a weak acid or weak base, it can react with water to make the solution acidic or basic.

The hydrolysis of salt ions in water is an example of a Brønsted-Lowry acid-base reaction, where the ions act as either proton donors or proton acceptors when they interact with water molecules. The extent of hydrolysis depends on the strength of the parent acid or base. The weaker the parent acid or base, the greater the extent of hydrolysis of its conjugate ion.

Types of Salt Hydrolysis

Table 2.1: Resulting solution pH based on salt type

Salt Type	Parent Acid/Base	Resulting Solution
Strong acid + Strong base	e.g., NaCl	Neutral
Weak acid + Strong base	e.g., CH_3COONa	Basic
Strong acid + Weak base	e.g., NH_4Cl	Acidic
Weak acid + Weak base	e.g., NH_4CH_3COO $CH_3COO\ NH_4$	Depends on relative strengths (K_a and K_b values)

Cation Hydrolysis

Cation hydrolysis happens when a positive ion (cation) from a salt reacts with water and releases H^+ ions, making the solution acidic.

When Does Cation Hydrolysis Occur?

It usually happens when the salt comes from a weak base and a strong acid. The cation (positive ion) from the weak base reacts with water, forming an acidic solution.

Example 1

Step-by-Step Cation Hydrolysis of Ammonium Chloride (NH_4Cl)

1. Dissolution of Salt
   When NH_4Cl is added to water, it dissociates completely into its ions:
   $NH_4Cl\ (s) \rightarrow NH_4^+\ (aq) + Cl^-\ (aq)$
   a. NH_4^+ is the cation from the weak base ammonia (NH_3).
   b. Cl^- is from the strong acid HCl and does not hydrolyse.
2. Identify the ion that hydrolyses
   Only the NH_4^+ ion reacts with water. This is a cation hydrolysis process.

3. Reaction of NH_4^+ with Water

   $NH_4^+ (aq) + H_2O (l) \rightleftharpoons NH_3 (aq) + H_3O^+ (aq)$

   a. NH_4^+ donates a proton (H^+) to water.

   b. This forms NH_3 and H_3O^+ (hydronium ion).

   c. The H_3O^+ ion makes the solution acidic.

Summary

1. NH_4Cl dissolves into NH_4^+ and Cl^-.
2. NH_4^+ undergoes cation hydrolysis by reacting with water.
3. It produces H_3O^+, which lowers the pH.
4. The resulting solution is acidic.

Example 2

Step-by-Step Cation Hydrolysis of Aluminium Chloride ($AlCl_3$)

1. Dissolution of salt

   When aluminium chloride ($AlCl_3$) dissolves in water, it dissociates into its ions:

   $AlCl_3 (s) \rightarrow Al^{3+} (aq) + 3Cl^- (aq)$

   a. Al^{3+} is the cation that undergoes hydrolysis.

   b. Cl^- does not hydrolyse because it comes from a strong acid (HCl).

2. Identify the hydrolysing ion

   Al^{3+} is a small, highly charged cation with a high charge density. It strongly attracts water molecules.

3. Reaction of Al^{3+} with water (hydrolysis)

   a. $Al^{3+} (aq) + 6H_2O (l) \rightleftharpoons [Al(H_2O)_6]^{3+}$

   b. The hydrated ion then loses a proton from one of its water molecules:

   $[Al(H_2O)_6]^{3+} \rightleftharpoons [Al(H_2O)_5(OH)]^{2+} + H_3O^+$

   c. The release of H_3O^+ (hydronium ion) makes the solution acidic.

Summary

1. $AlCl_3$ dissolves in water and releases Al^{3+} ions.
2. Al^{3+} attracts water molecules to form hydrated aluminium ions.
3. These hydrated ions donate protons to water, increasing H_3O^+ concentration.

4. The solution becomes acidic due to this cation hydrolysis.

Anion Hydrolysis

Anion hydrolysis occurs when a negative ion (anion) from a salt reacts with water to form hydroxide ions (OH^-), making the solution basic (alkaline).

When Does Anion Hydrolysis Happen?

Anion hydrolysis usually occurs when the salt comes from a strong base (like NaOH) and a weak acid (like acetic acid, CH_3COOH). In such cases, the anion (from the weak acid) reacts with water.

Example 3

Sodium Acetate (CH_3COONa)

When sodium acetate dissolves in water: $CH_3COONa\ (s) \rightarrow CH_3COO^-\ (aq) + Na^+\ (aq)$

1. Na^+ does not hydrolyse (it comes from a strong base).
2. CH_3COO^- (acetate ion) does hydrolyse: $CH_3COO^-\ (aq) + H_2O\ (l) \rightleftharpoons CH_3COOH\ (aq) + OH^-\ (aq)$

   This reaction:

   a. produces OH^-, increasing the pH.

   b. makes the solution basic.

Summary

1. Anion hydrolysis makes solutions basic.
2. It happens when salts come from a strong base and weak acid.
3. The anion which is from a weak acid (like CH_3COO^-, CN^-, or CO_3^{2-}) reacts with water, forming OH^-.

Example 4

Step-by-Step Anion Hydrolysis of Sodium trioxocarbonate(IV) (Na_2CO_3)

1. Dissolution of the salt

   When sodium carbonate (Na_2CO_3) dissolves in water, it dissociates completely:

   $Na_2CO_3\ (s) \rightarrow 2Na^+\ (aq) + CO_3^{2-}\ (aq)$

   a. Na^+ does not hydrolyse (it comes from a strong base).

b. CO_3^{2-} is the conjugate base of a weak acid (H_2CO_3) and does hydrolyse.

2. Hydrolysis of CO_3^{2-} ion

   $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$

   a. The carbonate ion reacts with water to form bicarbonate (HCO_3^-) and hydroxide (OH^-).

   b. OH^- increases the pH, making the solution basic.

3. Further hydrolysis (optional)

   $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

   Bicarbonate can also undergo further hydrolysis, but this reaction is weaker.

Summary

1. Na_2CO_3 dissociates into Na^+ and CO_3^{2-}.
2. CO_3^{2-} reacts with water to form OH^-, increasing the pH. The solution becomes basic due to anion hydrolysis.

Whilst these examples can seem complex, the general principle of **Table 2.1** holds for all of them.

Real-World Applications of Acidic, Basic and Neutral Salt Solutions

Acidic Salt Solutions (pH < 7)

1. Fertilisers (e.g., Ammonium salts like NH_4Cl or NH_4NO_3): These release H^+ ions in soil, lowering pH and aiding in the absorption of certain nutrients.
2. Industrial Metal Cleaning: Iron (III) chloride ($FeCl_3$) solutions are used to etch metals or clean machinery due to their mild acidity.
3. Electroplating and Dye Fixing: Acidic salts help in fixing dyes onto fabrics or plating metals in manufacturing.

Basic Salt Solutions (pH > 7)

1. Water Softening (e.g., Sodium carbonate Na_2CO_3): Removes hardness from water by precipitating calcium and magnesium ions.
2. Baking & Food Preparation (e.g., Sodium bicarbonate $NaHCO_3$): Baking soda is mildly basic and helps in leavening and neutralising acids in recipes.
3. Soap and Detergent Production: Basic salts are involved in the saponification process to make soap.

4. Antacid Formulations: Some basic salts (e.g., Magnesium hydroxide) are used in over-the-counter medicines to relieve heartburn.

Neutral Salt Solutions (pH ≈ 7)

1. Intravenous (IV) Fluids (e.g., NaCl - normal saline): Used in hospitals for hydration; being neutral prevents damage to tissues and blood cells.
2. Food Seasoning and Preservation: Table salt (NaCl) in solution is neutral and used in cooking, pickling, and curing meats.
3. Laboratory Reference Solutions: Neutral salts are used as controls or standards in chemical experiments due to their stable pH.

Factors Affecting the Extent of Hydrolysis

Several factors influence how much a salt undergoes hydrolysis in aqueous solution. These include:

1. **Strength of the Parent Acid or Base**

   The weaker the parent acid or base, the greater the extent of hydrolysis of its conjugate ion.

   Example: Salts of weak acids like CH_3COO^- or weak bases like NH_4^+ hydrolyse more than those from strong acids/bases.

2. **Charge on the ion**

   Ions with higher charges, especially metal cations like Al^{3+} or Fe^{3+}, tend to hydrolyse more strongly because they attract water molecules more intensely.

3. **Size of the ion**

   For ions with the same charge, smaller ions exhibit higher charge density, leading to stronger hydrolysis.

   Example

   a. Al^{3+} hydrolyses more than Na^+ due to its smaller size and higher charge.

   b. Among the ions Mg^{2+} and Ca^{2+}, Mg^{2+} hydrolyses more than Ca^{2+} due to its smaller size and higher charge density.

4. **Concentration of the Salt Solution**

   Dilute solutions often show a greater proportion of hydrolysis because water is in excess and equilibrium favours ion-water interaction.

5. Temperature

An increase in temperature usually increases the extent of hydrolysis, especially if the hydrolysis reaction is endothermic.

Activity 2.1 Exploring Salt Hydrolysis

Materials needed

Use the Reference Table on salt hydrolysis below

Type of Salt (Source)	pH	Cation Type	Anion Type
Strong acid + strong base (neutral salt) e.g. $MgCl_2$, Na	No hydrolysis (pH = 7)	Strong base cations (e.g. Na^+, K^+, Ba^{2+})	Strong acid anions (Cl^-, Br^-, I^-, ClO_4^-)
Strong acid + weak base (acidic salt) e.g. NH_4Br	Cationic hydrolysis (pH < 7)	Weak base cations (e.g. NH_4^+, Cu^{2+}, Zn^{2+})	Strong acid anions (Cl^-, Br^-, I^-, ClO_4^-)
Weak acid + strong base (basic salt) e.g. CH_3COOLi	Anionic hydrolysis (pH > 7)	Strong base cations (Na^+, K^+, Ca^{2+})	Weak acid anions (CH_3COO^-, S^{2-}, CO_3^{2-}, PO_4^{3-})

Steps

1. Identify salts that produce acidic aqueous solutions from the list below:

   NH_4Cl, CH_3COONa, $FeCl_3$, NaCl, Na_2CO_3, KNO_3

2. Write balanced ionic dissociation equations (with state symbols) for $NH_4Cl(aq)$ and $FeCl_3(aq)$.
   - **a.** For each Step 2 cation, write the equation showing how it reacts with water (hydrolysis) to produce hydronium, H_3O^+.
   - **b.** Using your equations, justify the acidity of the corresponding salt solutions.

3. From the salts listed in step 1, identify salts that produce basic aqueous solutions.
   - **a.** Write the balanced dissociation equations for the basic salt(s) identified in Step 3
   - **b.** Explain why these salts solutions are basic.

4. Identify salts that produce neutral aqueous solutions. Explain why aqueous solution(s) of this /these salts is/are neutral.

Extension Task

1. Verify your findings by testing each of these salts in turn
2. Dissolve a spatula of each salt into 50 ml of water.
3. Using Universal indicator, pH paper or another suitable indicator test the resulting solutions for pH.
4. Compare this to your conclusions above.

EXTENDED READING

1. https://www.chemguide.co.uk/physical/acids/salthydrolysis.html
2. https://chem.libretexts.org/Bookshelves/General_Chemistry/Book%3A_General_Chemistry_(Petrucci_et_al.)/17%3A_Acid-Base_Equilibria/17.06%3A_Salt_Hydrolysis

REVIEW QUESTIONS

Multiple choice Questions

1. Which of the following salts forms an acidic solution in water?

A. NaCl

B. NH_4Cl

C. Na_2CO_3

D. KNO_3

2. What is the product of the hydrolysis of CH_3COO^- in water?

A. CH_4

B. OH^-

C. H_3O^+

D. Na^+

3. The hydration of an ion refers to:

A. Precipitation of salts

B. Reaction of a salt with water

C. Surrounding of ions by water molecules

D. Evaporation of water

4. Which of the following salts is derived from a strong acid and strong base?

A. NH_4Cl

B. Na_2CO_3

C. NaCl

D. $FeCl_3$

5. Which ion contributes to the basicity of an aqueous solution?

A. NH_4^+

B. CH_3COO^-

C. Fe^{3+}

D. H_3O^+

6. What type of solution does Na_2CO_3 form in water?

A. Acidic

B. Basic

C. Neutral

D. Amphoteric

7. The hydrolysis of NH_4^+ in water results in:

A. Release of OH^-

B. Formation of NH_3 and H_3O^+

C. No reaction

D. Formation of Na^+

8. Which factor increases the extent of hydrolysis?

A. High salt concentration

B. Use of strong acids

C. Use of strong bases

D. Dilute solution

9. Which salt will produce a neutral solution upon dissolution in water?

A. NH_4Cl

B. NaCl

C. CH_3COONa

D. $FeCl_3$

10. Which of the following is correct about hydration energy?

A. It is the energy needed to evaporate water

B. It is the heat absorbed during hydrolysis

C. It is the energy released when ions are surrounded by water molecules

D. It is the heat required to ionise a salt

11. Which combination of ions will most likely produce a neutral salt?

A. Strong acid + strong base

B. Weak acid + strong base

C. Weak base + strong acid

D. Weak acid + weak base

12. Hydration involves the interaction of:

A. Water molecules with covalent bonds

B. Water molecules and ionic solutes

C. Salt and oxygen

D. Heat and acids

13. A salt solution has a pH of 5.0. This indicates the salt:

A. is neutral

B. undergoes cation hydrolysis

C. undergoes anion hydrolysis

D. cannot be hydrolysed

14. Which of the following explains why $FeCl_3$ produces an acidic solution?

A. Fe^{3+} is a weak acid

B. Fe^{3+} reacts with water to release H^+

C. $FeCl_3$ is a basic salt

D. Cl^- reacts with water

Structured Questions

1. Define hydration as it applies to ionic compounds.

2. What is meant by hydrolysis of salts?

3. State the difference between salts that produce acidic and basic solutions.

4. Explain why NaCl does not hydrolyse in water.

5. Write a balanced equation for the hydrolysis of NH_4^+ in water.

6. Write the hydrolysis equation for CH_3COO^- in water. State the nature (acidic/basic/neutral) of the solution.

7. Predict the pH of an aqueous solution of $FeCl_3$ and justify your answer with an equation.

8. Explain why aqueous Na_2CO_3 is basic using relevant ionic equations.

9. A student dissolves KNO_3 in water. Predict the pH of the solution and explain your reasoning.

10. Using chemical equations, differentiate between hydrolysis of NH_4Cl and CH_3COONa.

11. Classify NH_4Cl, NaCl, CH_3COONa, and $AlCl_3$ as forming acidic, basic, or neutral solutions. Explain.

12. Explain how the extent of hydrolysis is influenced by the strength of the parent acid or base and give examples.

13. Design an experiment to compare the acidity of NH_4Cl, NaCl, and Na_2CO_3. Include expected results.

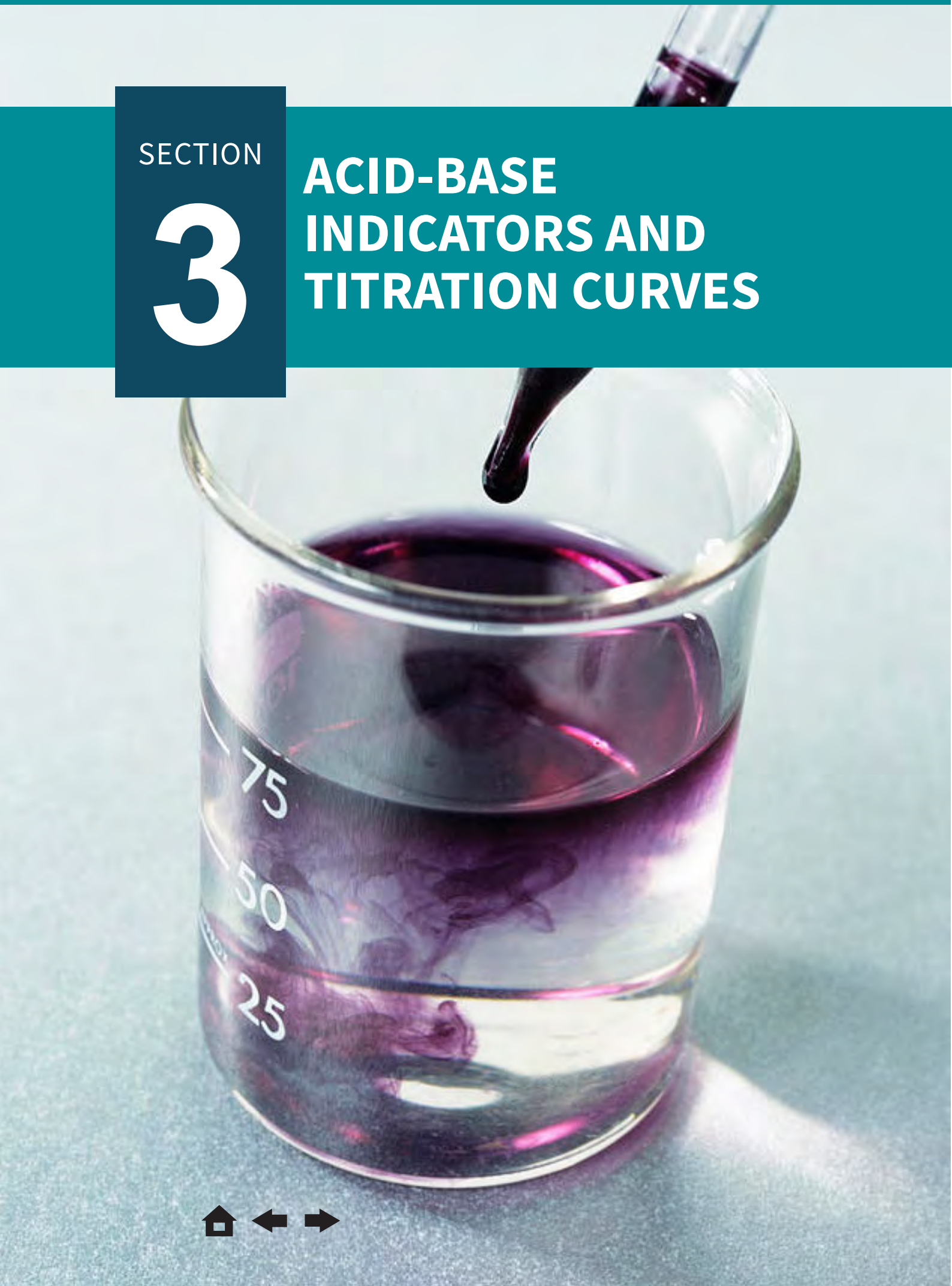
14. Discuss the environmental significance of salt hydrolysis.

15. Clarify the pH change of CH_3COONa in water with diagrams and equations.

SECTION

3

ACID-BASE INDICATORS AND TITRATION CURVES



PHYSICAL CHEMISTRY

Equilibria – 3

INTRODUCTION

In this section, you will learn about acid-base indicators, which are special chemicals that change colour when added to acids or bases. These colour changes help scientists know when a reaction is complete, especially during a process called titration. Indicators are usually weak acids or bases, and they show different colours depending on the pH of the solution. Common indicators include: phenolphthalein – turns pink in bases and colourless in acids and methyl orange – turns red in acids and yellow in bases. You will learn:

1. why indicators change colour at different pH values
2. which indicator to choose for different types of acid and base reactions
3. how to draw and read a titration curve – a graph that shows how the pH changes as one solution is added to another
4. important parts of the curve include the equivalence point (where the acid and base neutralise) and buffer region (where pH changes slowly)

Through fun activities and practice problems, you will also see how indicators are used in real-life jobs, like: making medicine, checking clean water and testing products in factories. To understand Acid-Base Indicators and Titration Curves, refer to this section in the previous years

KEY IDEAS

- Analyte is the solution in the beaker or flask that is being tested.
- Endpoint is the point where the indicator changes colour, showing the titration is almost done.
- Equivalence point is the point in titration when the acid and base have completely reacted.
- Indicator is a chemical that changes colour in an acid or a base.
- Titration is a method where an acid is slowly added to a base (or vice versa) to find the exact point they react completely.

- Titration curve is a graph that shows how pH changes as more titrant is added.

ACID-BASE INDICATOR

Acid-Base Indicator

An acid-base indicator is a weak acid or base that changes colour when the pH of a solution changes (when it becomes more acidic or more basic). An acid-base indicator is a special chemical that changes colour when added to an acid or a base. It is made from a weak acid or base. It shows one colour in acid and a different colour in base.

How Does It Work?

The indicator exists in two forms: (unionised molecules in equilibrium with their ions)

1. One form (HIn), which represents the molecular form gives Colour 1
2. The other (In^-) representing the ionic form gives Colour 2

These two forms are in equilibrium: HIn (Colour 1) $\rightleftharpoons$ $H^+ + In^-$ (Colour 2)

When the amount of HIn and In^- is the same, the colour looks like a mix of the two colours. The colour changes over a small pH range, depending on the indicator.

When Can We See the Colour Change?

Our eyes notice the colour change when the amount of HIn and In^- changes a lot. This usually happens around a pH range that is about 1 unit above or below the indicator's pKa value.

Common Acid-Base Indicators and Their Colour Changes

Table 2.2 shows some common indicators and how their colours change in acids and bases.

Table 2.2: Common Acid-Base Indicators

Indicator	Colour in Acid	Colour in Base
Litmus	Red	Blue
Methyl Orange	Red	Yellow
Phenolphthalein	Colourless	Pink
Bromothymol Blue	Yellow	Blue
Universal Indicator	Red to Yellow (pH 1–6)	Green to Purple (pH 7–14)

Choice of Indicator in Acid-Base Titration

CHOOSING THE RIGHT INDICATOR IN ACID-BASE TITRATION

When performing a titration (a method to find out how much acid reacts with a base), we use an indicator to tell us when the reaction is finished. This point is called the endpoint or equivalence point. The equivalence point is the point at which stoichiometric amount added of titrant reacts completely with the analyte, leaving no excess. It represents the theoretical end point of the reaction.

End point in the titration is the experimental physical point where the indicator changes colour signalling the end of the titration

But not all indicators work for all types of titrations. We must choose the one that changes colour at the right pH.

Simple Guide for Choosing Indicators

Table 2.3: Guide for choosing indicators

Type of Titration	pH at Equivalence Point	Best Indicator	Colour Change
Strong Acid + Strong Base	Around 7 (neutral)	Any (e.g. Litmus, Phenolphthalein, Methyl orange)	Varies depending on choice
Strong Acid + Weak Base	Less than 7 (acidic)	Methyl orange	Red → Yellow

Type of Titration	pH at Equivalence Point	Best Indicator	Colour Change
Weak Acid + Strong Base	Greater than 7 (basic)	Phenolphthalein	Colourless → Pink
Weak Acid + Weak Base	No sharp pH change	No suitable indicator	Use a pH meter instead

Note

- Choose an indicator that changes colour close to the equivalence point pH.
- For sharp colour changes, the indicator must have its working range within the steep part of the titration curve.

Titration Curves

A titration curve is a graph that shows how pH changes as one solution (like an acid or base) is slowly added to another during a titration. The x-axis shows the volume of the titrant added, and the y-axis shows the pH of the solution in the flask.

What Does the Curve Tell Us?

1. At first, the pH changes slowly.
2. Then there is a sudden jump in pH—this is the equivalence point.
3. After the jump, the pH levels off again.

Types of Titration Curves

Table 2.4: Types of titration curves and their equivalence point pH

Titration Type	Curve Shape	Equivalence Point pH
Strong Acid + Strong Base	Steep jump around pH 7	About pH 7
Weak Acid + Strong Base	Gentle rise, steep jump above pH 7	Around pH 8–9
Strong Acid + Weak Base	Steep jump below pH 7	Around pH 4–6
Weak Acid + Weak Base	Very gentle curve, no sharp jump	No clear equivalence point

Why Is It Useful?

1. Helps choose the best indicator (look for the steepest part).
2. Shows where the reaction is complete.
3. Used in chemistry labs, medicine, and water testing.

Titration Curve Diagrams

Refer to ***Figures 3.1, 3.2 & 3.3*** for different titration curves.

1. **Strong Acid vs Strong Base**

   pH curve for the titration of 25.0 cm^3 of 0.10 mol dm^{-3} HCl with 0.10 mol dm^{-3} NaOH

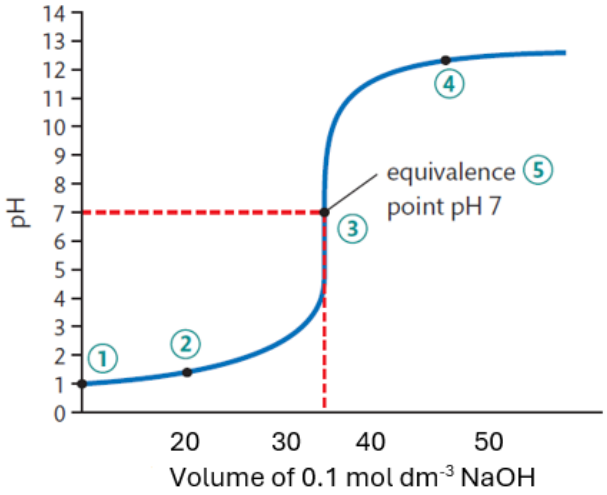


Figure 3.1: Strong Acid vs Strong Base

   The following points can be deduced from the graph in ***Figure 3.1***

   1. initial pH = 1 (pH of 0.10 mol dm^{-3} HCl)
   2. pH changes only gradually until equivalence
   3. very sharp jump in pH at equivalence: from pH 3 to pH 11
   4. after the equivalence point the curve flattens out at a high value (pH of strong base).
   5. *pH at equivalence = 7*

2. **Weak Acid vs Strong Base**

   pH curve for the titration of 25.0 cm^3 of 0.10 mol dm^{-3} CH_3COOH with 0.10 mol dm^{-3} NaOH

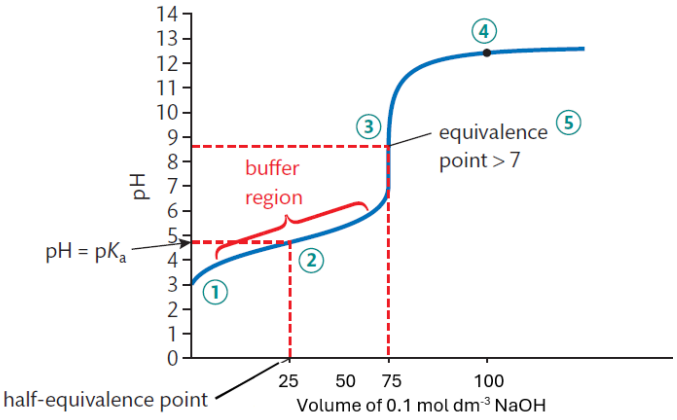


Figure 3.2: Weak Acid vs Strong Base

The following points can be deduced from the graph in ***Figure 3.2***

1. pH > 1 (pH of 0.10 mol dm–3 weak acid)
2. Addition of the strong base converts some of the weak acid to its conjugate base and creates a buffer. pH stays almost constant through the buffer region where both HA and A– are present. At half-equivalence point (12.5.0 cm3); pH = p*K*a
3. jumps in pH at equivalence from about pH 7.0 to 11.0, which is not as much of a jump as for a strong acid–strong base titration
4. after the equivalence point the curve flattens out at a high value (pH of strong base).
5. *pH at equivalence is > 7*

3. Strong Acid vs Weak Base

pH curve for the titration of 25.0 cm^3 of 0.10 mol dm^{-3} HCl with 0.10 mol dm^{-3} NH_3

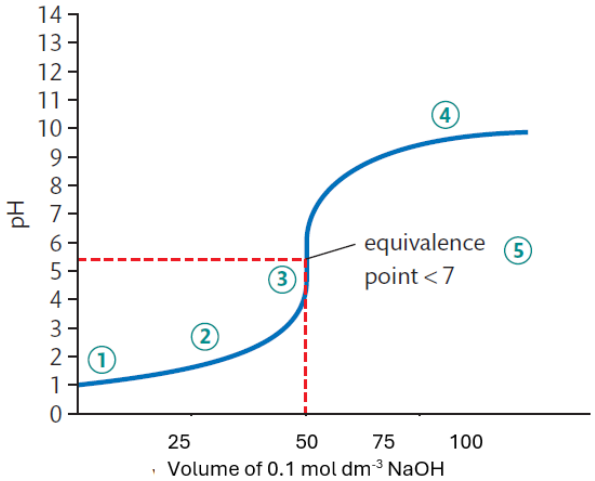


Figure 3.3: Strong Acid vs Weak Base

The following points can be deduced from the graph in ***Figure 3.3***

1. initial pH = 1 (pH of 0.10 mol dm^{-3} HCl)
2. pH stays relatively constant through the buffer region to equivalence
3. jumps in pH at equivalence from about pH 3.0 to 7.0
4. after the equivalence point the curve flattens out at a low pH (pH of weak base).
5. *pH at equivalence is < 7*

4. Titration of a weak acid and a weak base

pH curve for the titration of 25.0 cm^3 of 0.10 mol dm^{-3} CH_3COOH with 0.10 mol dm^{-3} NH_3

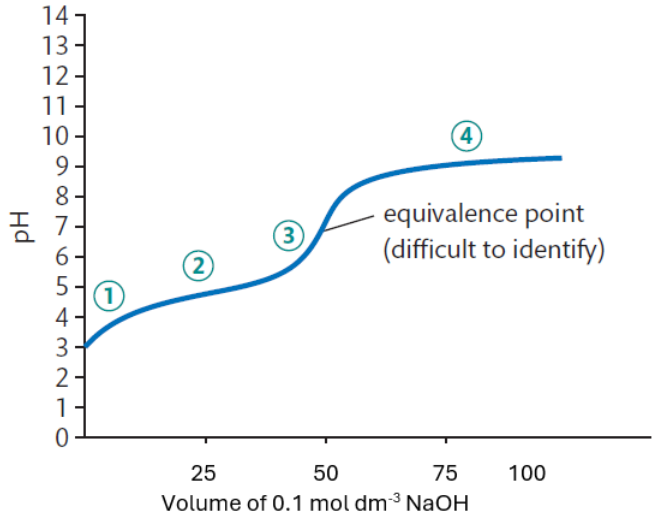


Figure 3.4: Weak Acid vs Weak Base

The following points can be deduced from the graph in ***Figure 3.4***

1. pH > 1 (pH of 0.10 mol dm^{-3} weak acid)
2. addition of base causes the pH to rise steadily
3. changes in pH at the equivalence point is much less sharp than in the other titrations
4. after the equivalence point the curve flattens out at a low pH (pH of weak base).
5. *pH at equivalence point is difficult to determine*

Drawing Titration Curves

1. Start with the pH of the first solution: Find the pH before adding anything (usually the acid or base being tested).
2. Add a little bit of the other solution (titrant): Each time you add, calculate the new pH.
3. Find the point where acid and base are equal (equivalence point): This is where the reaction is complete.
4. Add more titrant beyond that point: See how the pH keeps changing after the reaction ends.
5. Draw a graph: Put pH on the y-axis and volume added on the x-axis.
6. Join your points with a smooth line: The graph will have a curve, not sharp corners.
7. Mark the equivalence point: Show where the sharp rise or drop in pH happens.
8. If you see a slow change area, mark it as the buffer region: This is where the solution resists pH change.
9. Add the indicator's pH range on the graph: This helps you see if it changes colour at the right time.

Using the Curve to Pick the Right Indicator

1. Find the equivalence point on your graph (big jump in pH).
2. Look at the steep part of the curve near that point.
3. Choose an indicator that changes colour in that steep range.
4. Check if the indicator's colour change fits
5. It should happen where the big pH change is.

Activity 3.1 Investigating Acid-Base Indicators

Materials needed

1. Common indicators: phenolphthalein, methyl orange, litmus
2. Labelled solutions: distilled water (A), orange juice (B), baking soda (C), vinegar (D), toothpaste solution (E), water (F), soap solution (G)
3. Droppers, test tubes, and white tiles or paper for clearer colour observation

Steps

1. Put 10 cm3 of the solutions labelled A, B, C, D, E, F, and G.
2. Add two to three drops of phenolphthalein to solutions in the test tube.
3. Carefully observe the colour changes. Record the observations in the table below.

Test	Observation	Deduction
A + phenolphthalein		
B + phenolphthalein		
C + phenolphthalein		
D + phenolphthalein		
E + phenolphthalein		
F + phenolphthalein		
G + phenolphthalein		

4. Repeat steps 1,2 and 3 using methyl orange or litmus

Activity 3.2 Drawing a Strong Acid–Strong Base Titration Curve

Instructions

In this activity, you will draw the pH titration curve for a titration between 25.0 cm^3 of 0.10 mol dm^{-3} of a strong acid and 0.10 mol/dm^3 strong base. Use the titration data table below to plot the graph of the reaction

Steps

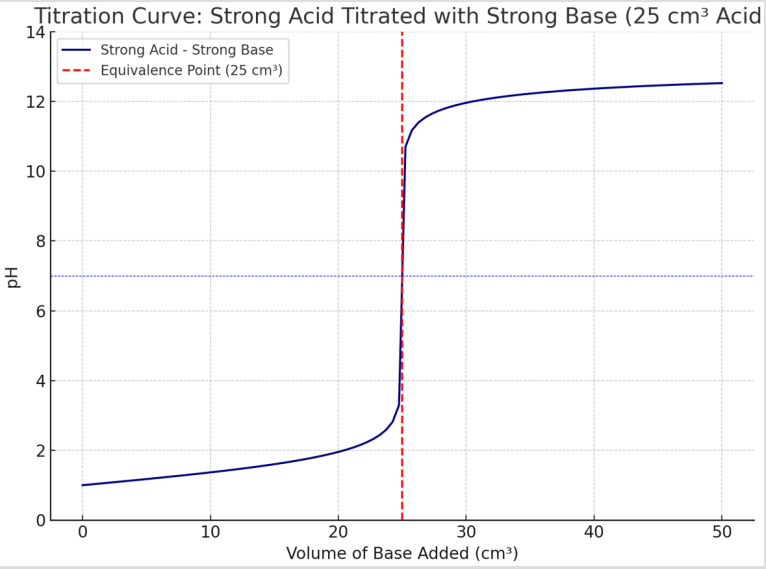
1. Draw a horizontal axis labelled "Volume of Base Added (cm^3)" from 0 to 50.
2. Draw a vertical axis labelled "pH" from 0 to 14.
3. Use the data table below to plot the pH at different volumes of base added.
4. Connect the points with a smooth curve.
5. Mark the equivalence point (25 cm^3, pH = 7) with a dashed vertical line.

Selected Titration Data (Strong Acid with Strong Base)

Volume of Base Added (cm^3)	pH
0.0	1.00
5.0	1.30
10.0	1.60
15.0	2.30
20.0	3.50
24.0	6.10
24.9	6.90
25.0	7.00
25.1	7.10
26.0	10.00
30.0	12.00
40.0	13.00
50.0	13.50

Note

This is how the curve will look when accurately plotted:



6. Repeat steps 1 to 5. Then, using the titration data table for Titration A below, plot the graph for the titration of a weak acid with a strong base.

Titration A

Volume of Base Added (cm^3)	pH
0.00	2.87
5.05	4.15
10.10	4.58
15.15	4.93
20.20	5.37
25.25	10.70
30.30	11.98
35.35	12.23
40.40	12.37
45.45	12.46

7. At what volume does the steepest change in pH occur?

8. Why is the pH at the equivalence point exactly 7.0?

9. Which indicator is most suitable for this titration?

10. What would happen if you used methyl orange for this titration?

Activity 3.3 Titration Technique

Materials needed

1. Glassware: Burette, Conical Flask (Erlenmeyer flask), Pipette (usually 25.0 cm^3), Pipette Filler, Beaker, Funnel, retort stand and Clamp, White Tile

2. Solutions: a solution of known concentration (e.g., NaOH), a solution of unknown concentration (e.g., HCl), Distilled Water

3. Indicator: Phenolphthalein or Methyl Orange

4. Safety Equipment: Gloves, Lab Coat, Safety Goggles

Steps

Step 1: Preparing the Burette (Image 1)

A gloved hand carefully fills a clean burette with a titrant (usually a strong acid or base) using a small glass beaker.

Safety Tips

1. Always wear gloves and protective goggles when handling acids and bases.
2. Ensure the burette is clamped securely to a stand at eye level.
3. Fill below eye level to avoid splashes into the face.
4. Check for air bubbles in the nozzle and remove them before starting.

Step 2: Initial Setup (Image 2)

The full titration setup includes:

1. A vertical burette clamped to a stand.
2. An Erlenmeyer flask (conical flask) beneath the burette, containing a measured volume of analyte (unknown solution) and an indicator (e.g., phenolphthalein, which appears pink in basic solution).

Precision Tips

1. Record the initial burette reading (to the nearest 0.10cm3).
2. Ensure the tip of the burette is just above the flask but not submerged.
3. Use a white tile under the flask to make colour changes more visible.

Step 3: Performing the Titration (Image 3)

1. The titrant is added slowly and steadily from the burette into the flask while swirling the flask gently with the other hand.
2. As the titrant nears the endpoint, slow down the flow to dropwise.

Observation Tips

1. Watch for the first permanent colour change (e.g., pink to colourless if acid is added to base).
2. Avoid overshooting the endpoint by adding titrant too quickly.
3. Record the final burette reading immediately after the endpoint is reached.

4. Repeat the titration at least 2–3 times until two concordant results (within ±0.2cm³) are obtained.

Titration results are recorded in a table looking like this:

Burette Reading/cm³	1	2	3
Final	22.80	46.80	22.50
Initial	0.00	24.40	0.00
Volume of acid used	22.80	22.40	22.50

Step 4: After the Endpoint (Image 4)

1. The flask now contains a neutralised solution, shown as clear and colourless.
2. Dispose of chemicals properly and rinse all glassware.

Step 5: Calculate the average titre using two concordant titre values to two decimal places.

$$\text{Average titre} = \frac{22.40 + 22.50}{2} = 22.45 \text{ cm}^3$$

Step 6: Calculate the concentration of their unknown solution using their titration results and the formula

$$\frac{C_A V_A}{C_B V_B} = \frac{n_A}{n_B} = \text{mole ratio}$$

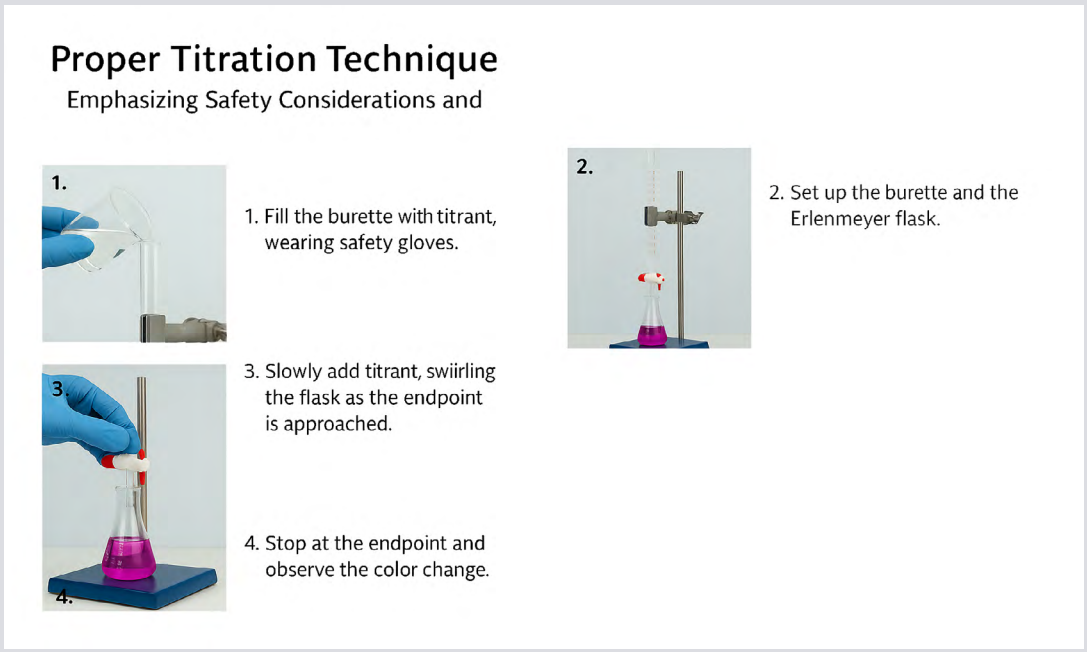


Figure 3.5: Proper Titration Technique

Activity 3.4 Exploring Indicators, Titration Curves, and Titration Technique

Materials needed

Phenolphthalein, methyl orange, bromothymol blue, hydrochloric acid (HCl), sodium hydroxide (NaOH), burette, pipette, conical flask, white tile, beakers, pH meter or pH paper, graph paper or digital plotting tool.

Steps

Step 1: Investigating Acid-Base Indicators

1. Test a range of acid-base indicators (e.g., litmus, phenolphthalein, methyl orange) using known acidic and basic substances (water, distilled water, vinegar, baking soda, hydrochloric acid solution, sodium hydroxide solution)
2. Record the colour change at different pH levels and match to a pH scale.
3. Discuss how each indicator works and the concept of indicator transition ranges.

Step 2: Performing a Strong Acid–Strong Base Titration

Perform a titration of 25.0 cm^3 of 0.10 mol/dm^3 HCl with 0.10 mol/dm^3 NaOH using phenolphthalein as the indicator. Accurately record volume readings and note the endpoint.

Step 3: Drawing and Interpreting the Titration Curve

1. Use the titration data to plot a titration curve (pH vs volume of base added). Label the initial pH, steep section, and equivalence point on the curve.
2. Interpret the graph and relate it to the theoretical understanding of acid-base neutralisation. Highlight where the indicator's transition range fits on the curve.

EXTENDED READING

1. IUPAC Compendium of Chemical Terminology
2. Khan Academy: Acid-Base Indicators and Titration Curves
3. RSC (Royal Society of Chemistry): Online indicator resources
4. Silberberg & Amateis: Chemistry: The Molecular Nature of Matter and Change

REVIEW QUESTIONS

1. **a.** Explain how acid-base indicators work in terms of pH and colour change.

   b. Explain how an acid-base indicator works using the example of phenolphthalein.

2. **a.** Describe the colour changes of the following indicators across pH ranges:

   i. Methyl orange

   ii. Phenolphthalein

   iii. Bromothymol blue

   b. Identify the observable colour changes of methyl red in acidic and basic solutions.

3. **a.** Compare the suitability of phenolphthalein and methyl orange in a titration between:

   i. Strong acid and strong base

   ii. Weak acid and strong base

   iii. Strong acid and weak base

   b. Describe the difference between a strong acid–strong base titration and a weak acid – strong base titration.

4. Label the key parts (initial pH, steep rise, equivalence point) on a typical strong acid strong base titration curve.

5. **a.** A student titrates 25.0 cm^3 of 0.10 mol/dm^3 HCl with 0.10 mol dm^{-3} NaOH and observes the endpoint using phenolphthalein. Explain why this is an appropriate indicator choice for this titration.

   b. In a titration between ethanoic acid (CH_3COOH) and NaOH: Justify the choice of phenolphthalein over methyl orange.

6. **a.** Analyse the differences in titration curves for the following combinations:

   i. Strong acid vs. strong base

   ii. Weak acid vs. strong base

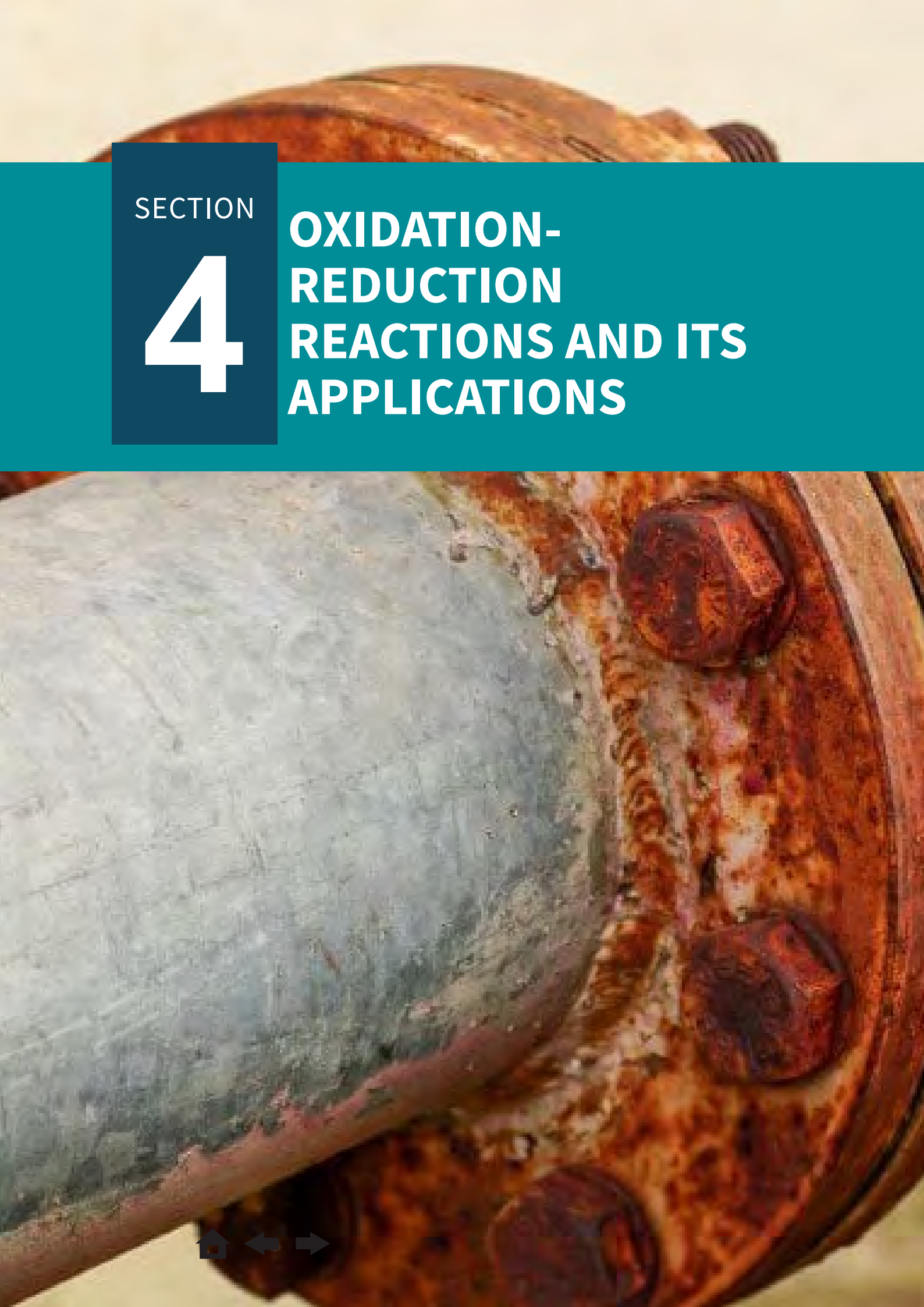
iii. Strong acid vs. weak base

iv. Weak acid vs. weak base

b. Explain how the equivalence point of a titration helps in the selection of a suitable indicator.

7. a. Draw and annotate the titration curve for a weak acid–strong base titration. Indicate the buffer region and explain its significance.

b. Explain why a weak acid–weak base titration typically does not have a steep vertical portion on the curve, and why indicators are not ideal for it.



SECTION

4

OXIDATION-REDUCTION REACTIONS AND ITS APPLICATIONS

PHYSICAL CHEMISTRY

Equilibria – 4

INTRODUCTION

In this section, you will discover the fascinating world of redox reactions, where substances gain or lose electrons during chemical changes. You will learn what oxidation and reduction mean, how to identify substances that are gaining or losing electrons, and explore the real-life applications of redox reactions, including their roles in batteries, rusting, and metal extraction. Through simple explanations and hands-on experiments, you will understand how metals react, investigate oxidising and reducing agents, and practice writing half-equations in redox reactions. Additionally, you will discover how redox reactions generate electricity in batteries and fuel cells, learn about electrolytic cells and their applications like electroplating and purifying metals, and understand why metals corrode/rust(for iron) and how to prevent corrosion in everyday objects.

KEY IDEAS

- **Corrosion** a redox process in which metals react with substances in the environment, often oxygen and moisture, leading to rusting (in the case of iron) or deterioration.
- **Displacement reaction**: a reaction in which a more reactive metal displaces a less reactive metal from its compound.
- **Electron transfer**; the movement of electrons from one substance to another during a redox reaction.
- **Oxidation** is the loss of electrons by a substance during a chemical reaction; it may also involve the gain of oxygen or loss of hydrogen as well as increase in oxidation number.
- **Oxidising agent**: a substance that gains electrons and causes another substance to be oxidised.
- **Reactivity series**: a ranking of metals from most reactive to least reactive based on their ability to lose electrons and form positive ions.

- **Redox reaction**: a chemical reaction in which oxidation and reduction occur simultaneously.
- **Reducing agent**: a substance that loses electrons and causes another substance to be reduced.
- **Reduction**: the gain of electrons by a substance. It may also involve the loss of oxygen or gain of hydrogen as well as decrease in oxidation number.

OXIDATION AND REDUCTION (REDOX) REACTIONS

Redox Reactions

Redox reactions cause changes in chemicals where some substances lose electrons (oxidation) and others gain electrons (reduction). These reactions happen all around us in Ghana — in our homes, farms, schools, and towns.

Redox reaction can be defined as a reaction in which electrons are transferred from one reactant to another, resulting in changes in oxidation states of the relevant species.

Real-Life Examples in Ghana

1. Gold Mining (Obuasi & Tarkwa): In gold mining areas, redox reactions help to remove gold from rocks.
2. Traditional Iron Work (Northern Ghana): Blacksmiths heat iron ore with charcoal. This causes a redox reaction that turns iron ore into real iron for tools and decorations.
3. Cocoa Fermentation: When cocoa beans are processed, redox reactions change their smell and colour, giving Ghanaian cocoa its world-famous taste.
4. Rusting Near the Coast (Elmina & Cape Coast): Metal gates, pipes, and buildings rust quickly in salty air. This is an oxidation reaction that damages metals and costs a lot of money to fix.

Definitions of Oxidation and Reduction

Oxidation and reduction can be defined in 4 broad ways. All of which could be applicable to any given situation. Choosing the best one to describe the processes will often depend upon the context.

In Terms of Oxygen

Oxidation: The gain of oxygen by a substance. For example, $4Fe + 3O_2 \rightarrow 2Fe_2O_3$

Iron (Fe) gained oxygen, so Fe is oxidised.

Reduction: The loss of oxygen from a substance.

For example, $CuO + H_2 \rightarrow Cu + H_2O$

Oxygen is removed from CuO (Copper oxide), hence, CuO is reduced

In Terms of Hydrogen

Oxidation: The loss of hydrogen from a substance. Example, in $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Hydrogen is removed from methane (CH_4), CH_4 is oxidised.

Reduction: The gain of hydrogen by a substance. In $C_2H_{4\ (g)} + H_{2\ (g)} \rightarrow C_2H_{6\ (g)}$, hydrogen is added to ethene (C_2H_4), C_2H_4 is reduced.

In Terms of Electron Transfer

Oxidation: The loss of electrons. Example: $Na \rightarrow Na^+ + e^-$ (oxidation),

Mnemonic: OIL — Oxidation Is Loss of electrons

Reduction: The gain of electrons. Example, $Cl_2 + 2e^- \rightarrow 2Cl^-$ (reduction)

Mnemonic: RIG — Reduction Is Gain of electrons

In Terms of Oxidation Numbers

Oxidation: An increase in oxidation number of relevant species.

Example: $Fe^{2+} \rightarrow Fe^{3+}$ (oxidation)

The oxidation number of Fe^{2+} increased from +2 to +3 in Fe^{3+}. Fe^{2+} is oxidised.

Reduction: A decrease in oxidation number of relevant species.

Example: $MnO_4^- \rightarrow Mn^{2+}$ (reduction)

The oxidation number of Mn in MnO_4^- decreased from +7 to +2 in Mn^{2+}. MnO_4^- is reduced.

Real-Life Examples of Redox Reactions

1. **Combustion of Fuel (in Cars or Stoves)**

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

   Methane (CH_4) is oxidised (loses hydrogen). Oxygen (O_2) is reduced (gains hydrogen). This reaction produces energy (heat and light).

2. **Displacement Reaction (Zinc and Copper Sulphate)**

   Equation: $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

   Zinc loses electrons → Oxidised.

   $Zn \rightarrow Zn^{2+} + 2e^-$

   Copper ions (Cu^{2+}) from $CuSO_4$ gain electrons → Reduced.

   $Cu^{2+} + 2e^- \rightarrow Cu$

   This is used in metal recovery and batteries.

3. **Rusting of Iron (in Moist Air)**

   Equation: $4Fe + 3O_2 + 6H_2O \rightarrow 4Fe(OH)_3$

   Iron (Fe) is oxidised to iron ions while oxygen (O_2) is reduced. This is a slow redox reaction that causes rust.

4. **Respiration (in Living Cells)**

   Equation: $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + \text{energy}$

   Glucose ($C_6H_{12}O_6$) is oxidised; Oxygen is reduced. This reaction gives energy to living things.

5. **Electrolysis of Water (in Labs or Hydrogen Production)**

   Equation: $2H_2O \rightarrow 2H_2 + O_2$

   Water is split: $H_2O \rightarrow H_2$ (reduction at cathode) and $H_2O \rightarrow O_2$ (oxidation at anode).

   H_2O is both oxidised and reduced. Oxidation number of hydrogen in H_2O is decreased from +1 to 0 (reduction) in H_2 while that of oxygen is increased from -2 in H_2O to 0 in O_2 (oxidation)

6. **Corrosion of Copper (Green Patina)**

   Equation: $2Cu + O_2 + CO_2 + H_2O \rightarrow Cu_2(OH)_2CO_3$

   Copper reacts with moist air and CO_2 to form green copper carbonate. It is a slow oxidation process.

Activity 4.1 Oxidation and Reduction in Everyday Observations

Step 1

1. Observe these items or materials in ***Figures 4.1 and 4.2***
2. What changes do you observe in these materials?
3. Why do you think the colour or texture has changed?
4. Have you seen anything like this in your community?

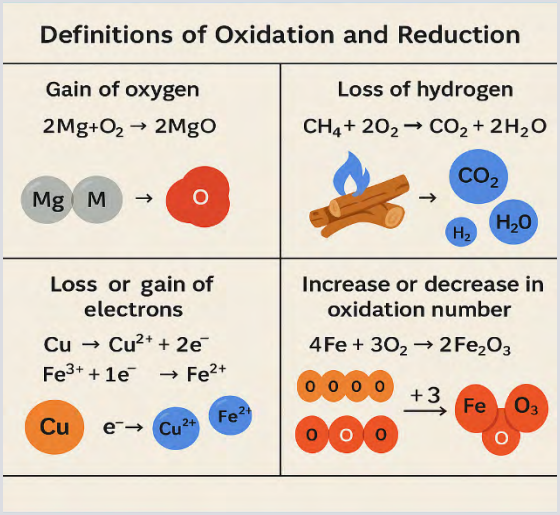


Figure 4.1: *Images of fresh items*



Figure 4.2: *Images showing redox processes*

Step 2



Using the poster above define oxidation and reduction in terms of

1. Gain of oxygen / Loss of oxygen
2. Loss of hydrogen / Gain of hydrogen
3. Loss of electrons / Gain of electrons
4. Increase or decrease in oxidation number

Step 3

Using the reactions below:

1. Assign oxidation numbers
2. Identify the substances oxidised and reduced
3. Explain your reasoning in (1) and (2)

$2Mg + O_2 \rightarrow 2MgO$

$2Na + Cl_2 \rightarrow 2NaCl$

$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

$H_2 + Cl_2 \rightarrow 2HCl$

$CH_3CH_2OH + [O] \rightarrow CH_3COOH + H_2O$

$2H_2O_2 \rightarrow 2H_2O + O_2$

$4Fe + 3O_2 \rightarrow 2Fe_2O_3$

$PbO_2 + 4H^+ + 2Cl^- \rightarrow PbCl_2 + Cl_2 + 2H_2O$

$2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$

REACTIVITY OF METALS

The reactivity of metals refers to their tendency to lose electrons and form positive ions (cations). Some metals react quickly, while others react slowly or not at all. A very reactive metal reacts easily with water, air, or acid. A less reactive metal reacts slowly or needs more heat to react.

The Reactivity Series

The Reactivity Series is a list that shows metals from the most reactive to the least reactive.

Here is a part of the list showing the most common metals:

Potassium (K) – **most reactive**

Sodium (Na)

Calcium (Ca)

Magnesium (Mg)

Aluminium (Al)

Zinc (Zn)

Iron (Fe)

Tin (Sn)

Lead (Pb)

Copper (Cu) – not very reactive

Gold (Au) – does not react easily

Platinum (Pt) – **Least Reactive**

Experimental Determination of Reactivity of Metals

We can find out how reactive a metal is by doing simple experiments. The faster or more strongly a metal reacts, the more reactive it is.

Materials required

Metals: Potassium, Sodium, Calcium, Magnesium, Zinc, Iron, Copper

Cold water, dilute hydrochloric acid (HCl), test tubes and test tube rack, measuring cylinder, droppers or pipettes, Bunsen burner or source of steam, copper (II) sulphate solution, sandpaper (to clean metal surfaces), tongs and gloves (for safety)

Procedure

Experiment 1: Reaction with Water

1. Your teacher will demonstrate the reactions of potassium, sodium and calcium, watch these demonstrations and take notes.
2. Clean small pieces of each of the remaining metals using sandpaper.
3. Add about 10 cm^3 cold water into a test tube.
4. Add a clean piece of metal and observe what happens.
5. Record the speed of reaction (fizzing, movement, heat).

Expected Experimental Results

Experiment 1: Reaction of Metals with Cold Water

Metal	Observation in Cold Water	Expected Result
Potassium (K)	Reacts violently, sparks, lilac flame	Very reactive
Sodium (Na)	Reacts vigorously, forms NaOH, hydrogen released	Very reactive
Calcium (Ca)	Bubbles (effervescence), forms $Ca(OH)_2$	Reactive
Magnesium (Mg)	Very slow/no reaction	Slightly reactive
Zinc (Zn)	No reaction	Unreactive in cold water
Iron (Fe)	No reaction	Unreactive in cold water
Copper (Cu)	No reaction	Unreactive in cold water
Gold (Au)	No reaction	Unreactive in cold water

Experiment 2: Reaction with Dilute Acid

1. Pour dilute hydrochloric acid into a test tube.
2. Add a clean piece of metal to the acid. DO NOT REACT POTASSIUM, SODIUM OR CALCIUM WITH ACID.
3. Observe bubbling or temperature change.
4. Record how fast hydrogen gas is produced.

Expected Experimental Results

Metal	Observation	Expected Result
Magnesium (Mg)	Rapid effervescence, produces H_2 gas	Highly reactive
Zinc (Zn)	Moderate effervescence	Reactive
Iron (Fe)	Slow bubbling	Less reactive
Copper (Cu)	No reaction	Unreactive
Gold (Au)	No reaction	Unreactive

Experiment 3: Metal Displacement Reaction

1. Pour copper (II) sulphate solution into a test tube.
2. Add a piece of another metal (e.g., zinc or iron).
3. Leave it for a few minutes and observe any colour change or metal deposit.
4. Record whether displacement occurred.

Expected Experimental Results

Metal (Solid)	Salt Solution	Observation	Expected Result
Magnesium (Mg)	Copper (II) sulphate ($CuSO_4$)	Brown deposit of copper, blue solution fades	Magnesium is more reactive than copper
Zinc (Zn)	Copper (II) sulphate ($CuSO_4$)	Brown copper forms on zinc, blue fades	Zinc displaces copper
Iron (Fe)	Copper (II) sulphate ($CuSO_4$)	Reddish-brown copper forms on iron	Iron is more reactive than copper
Copper (Cu)	Zinc sulphate ($ZnSO_4$)	No visible reaction	Copper is less reactive than zinc
Copper (Cu)	Iron (II) sulphate ($FeSO_4$)	No visible reaction	Copper cannot displace iron
Zinc (Zn)	Iron (II) sulphate ($FeSO_4$)	Grey iron forms on zinc, solution fades	Zinc displaces iron

Experiment 4: Reaction with Oxygen

1. Clean the surface of each metal using sandpaper to remove any oxide layer. Your teacher will demonstrate the reactions of potassium, sodium and calcium. DO NOT ATTEMPT THESE YOURSELF.
2. Hold the metal with tongs and place it in a crucible or on a metal spoon.
3. Heat the metal strongly using a Bunsen burner or spirit lamp.
4. Observe the reaction: Look for signs like flame, colour change, or formation of oxide.

5. Allow the crucible to cool and note the colour and texture of the product formed.
6. Repeat for each metal and compare the intensity of the reaction.

Expected Experimental Results

Metal	Observation (upon heating in air/ oxygen)	Expected Result
Potassium (K)	Burns with a lilac flame	Very reactive
Sodium (Na)	Burns with a bright yellow flame	Very reactive
Calcium (Ca)	Burns with a red/orange flame	Very reactive
Magnesium (Mg)	Burns with a bright white flame	Highly reactive
Zinc (Zn)	Burns with bluish-green flame	Moderately reactive
Iron (Fe)	Glows red-hot and forms black oxide	Less reactive
Copper (Cu)	Forms black coating slowly	Low reactivity
Gold (Au)	No reaction	Unreactive

Activity 4.2 Investigating Metal Reactivity

Materials needed

Test tubes and test-tube rack, beakers (100 mL and 250 mL), measuring cylinders (50 mL), metal samples: magnesium, zinc, iron, copper, salt solutions: copper (II) sulphate, zinc sulphate, iron (II) sulphate, dilute hydrochloric acid (HCl, ~1 M), distilled water, sandpaper (for cleaning metal surfaces), dropper or pipette, thermometer

Stopwatch or clock, stirring rod, gloves, goggles, and lab coats (for safety), tongs or forceps, waste container for used solutions.

This activity should be carried out in group.

Steps

1. Select a set of metals (e.g., magnesium, zinc, iron, copper), choose suitable reactions to investigate (e.g., with acids, water, or salt solutions), decide

on measurements to record (e.g., fizzing rate, temperature change), and prepare a data table to organise the results.

2. Use the table below to record your observations and measurements during the metal reactivity experiment. Test each metal with a chosen solution (acid, water, or salt solution) and observe the results.

Metal Tested	Reaction Type	Rate of Fizzing	Temp. Before (°C)	Temp. After (°C)	Colour Change	Reaction Observations

3. Discuss
   a. what you observed and how it matched or differed from what you expected.
   b. what may have affected the results (purity of metals, amount of acid, etc.).
4. Identify common patterns and create a class reactivity series. Rank the metals from most to least reactive based on your results.
5. Compare the group reactivity series with the standard reactivity series.

OXIDISING AND REDUCING AGENTS

Oxidising Agent

An oxidising agent is a substance that gains electrons during a redox reaction and causes another substance to be oxidised (i.e., to lose electrons). The oxidising agent itself is reduced.

Example: $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

Here, Cu^{2+} is the oxidising agent because, it gains 2 electrons from Fe. It is reduced to Cu(s).

Characteristics of Oxidising Agents

Refer to **Table 4.1** for the Characteristics of oxidising agents.

Table 4.1: Characteristics of oxidising agents

Characteristic	Explanation
Electron acceptor	Gains electrons from another substance during a redox reaction.
Causes oxidation	Oxidises another substance by removing electrons from it.
Itself undergoes reduction	While oxidising others, the oxidising agent is reduced (oxidation number decreases).
Often contains high oxidation states	Begins in a high oxidation state and becomes reduced during the reaction.
Commonly non-metals or positive ions	Examples include O_2, Cl_2, $KMnO_4$, $K_2Cr_2O_7$, H_2O_2, NO_3^-, Cu^{2+}, Fe^{3+}.
Strong oxidisers react readily	They have a high tendency to gain electrons (e.g., halogens, $KMnO_4$).
Used in bleaching and disinfection	Destroy organic materials and microbes (e.g., chlorine, ozone).

Reducing Agents

A reducing agent is a substance that loses electrons during a redox reaction and causes another substance to be reduced. It is itself oxidised in the process. Reducing agents are often metals or species in low oxidation states.

Characteristics of Reducing Agents

Refer to **Table 4.2** for the Characteristics of Reducing agents.

Table 4.2: Characteristics of Reducing agents

Characteristic	Explanation
Electron donor	Loses electrons to another substance during a redox reaction.
Causes reduction	Reduces another substance by donating electrons to it.

Characteristic	Explanation
Itself undergoes oxidation	While reducing others, the reducing agent is oxidised (oxidation number increases).
Often contains low oxidation states	Begins in a low oxidation state and becomes more positive after oxidation.
Commonly metals or negative ions	Examples include Fe^{2+}, Zn, H_2, CO, and SO_3^{2-}.
Strong reducers react easily	Alkali and alkaline earth metals (e.g., Na, K) are strong reducing agents.
Used in metal extraction and antioxidants	Important in both industrial processes and biological systems.
Colour change may indicate reaction	Can be observed when reducing agents react with coloured oxidising agents (e.g., $KMnO_4$ becomes colourless).

Example of reaction: $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

Fe is the reducing agent because it loses electrons and is oxidised to Fe^{2+}, reducing Cu^{2+} to Cu.

Tests for Oxidising and Reducing Agents

In **Table 4.3** are common tests used to identify oxidising and reducing agents, along with observable changes and chemical explanations.

Table 4.3: Tests for Oxidising and Reducing Agents

Test Substance	Used To Test For	Observation	Explanation
Potassium iodide (KI)	Oxidising agent	Solution turns brown or forms purple vapor	Oxidising agent oxidises I^- to I_2. Example: $Cl_2 + 2KI \rightarrow 2KCl + I_2$
Fe^{2+} ions	Oxidising agent	Pale green turns yellow-brown	Fe^{2+} is oxidised to Fe^{3+}.

Test Substance	Used To Test For	Observation	Explanation
Starch + Iodide	Oxidising agent	Solution turns blue-black	Oxidising agent releases I_2, which reacts with starch to form blue-black complex.
Potassium permanganate ($KMnO_4$, acidified)	Reducing agent	Purple solution becomes colourless	Reducing agent reduces MnO_4^- to Mn^{2+}. $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$
Potassium dichromate ($K_2Cr_2O_7$, acidified)	Reducing agent	Orange solution turns green	Reducing agent reduces $Cr_2O_7^{2-}$ to Cr^{3+}. $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
Hydrogen peroxide (H_2O_2)	Both	Effervescence (O_2 released)	Decomposes; acts as reducing agent by giving off O_2.

Applications of Oxidising and Reducing Agents

Table 4.4 shows summary of Real-Life Applications of Oxidising and Reducing Agents Across Fields Such as Industry, Healthcare, and Water Treatment, and metallurgy.

Table 4.4: Applications of Oxidising and Reducing Agents

Application Area	Agent Type	Examples	Explanation
Bleaching & Disinfection	Oxidising agent	Chlorine (Cl_2), Hydrogen peroxide (H_2O_2), Ozone (O_3)	Oxidises stains, kills bacteria by breaking down cell walls.
Water Treatment		Chlorine, Ozone	Disinfects water by oxidising microbes and organic waste.
Rocket Propulsion		Potassium nitrate (KNO_3), Ammonium perchlorate	Provides oxygen for combustion in fuels.

Application Area	Agent Type	Examples	Explanation
Electrochemical Cells		Oxygen, Permanganate (MnO_4^-)	Accepts electrons at the cathode to drive the cell reaction.
Metallurgy		Air (O_2), Oxygen-enriched blast	Oxidises impurities in ores such as FeS to Fe_2O_3.
Metal Extraction	Reducing agent	Carbon (C), Carbon monoxide (CO), Aluminum (Al)	Reduces metal oxides to obtain pure metals.
Photography (Developing)		Hydroquinone, Metol	Reduces silver ions (Ag^+) to metallic silver (Ag) on photographic paper or film.
Antioxidants in Biology		Ascorbic acid (Vitamin C), Glutathione	Donate electrons to neutralise free radicals and protect cells.
Redox Titration		Iron (II) sulfate ($FeSO_4$)	Used to determine the concentration of oxidising agents.
Industrial Processes		Hydrogen gas (H_2), Carbon monoxide (CO)	Used in refining metals and removing oxygen from compounds.

Activity 4.3 Understanding Oxidising and Reducing Agents

Using the following redox reaction write the definitions of oxidising and reducing agents using three different approaches: $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$

1. a. Oxygen/Hydrogen Transfer

Copper (II) oxide loses oxygen to form copper.

Hydrogen gains oxygen to form water.

Write your definitions:

Oxidising Agent:

Reducing Agent:

b. Electron Transfer Approach

Cu^{2+} in CuO gains electrons (is reduced) to form Cu.

H_2 loses electrons (is oxidised) to form H^+ in H_2O.

Write your definitions:

Oxidising Agent:

Reducing Agent:

c. Oxidation Number Change Approach

Cu changes from +2 to 0 (reduction).

H changes from 0 to +1 (oxidation).

Write your definitions:

Oxidising Agent:

Reducing Agent:

2. Using the following equations calculate the oxidation numbers and identify oxidising and reducing agents:

a. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

b. $MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$

3. Share your definitions and reaction analyses with your colleagues.

4. Using the equation: $H_2O_2 + 2FeSO_4 + H_2SO_4 \rightarrow Fe_2(SO_4)_3 + 2H_2O$, apply all three definitions to identify the oxidising and reducing agents.

5. Using the iron-oxygen reaction: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$, identify:

a. the oxidising agent and justify your choice

b. the reducing agent and justify your choice

BALANCING OF REDOX REACTIONS

Oxidation Half-Reaction

An oxidation half-reaction shows the part of a redox (oxidation-reduction) reaction where a substance loses electrons. It focuses only on the oxidation process, not the entire redox equation.

Features of Oxidation Half-Reactions

Feature	Description
Electron loss	Electrons appear on the product side of the equation.
Oxidation number	Increases during oxidation.
Charge balance	Must be balanced for both mass and charge.
Occurs at	The anode in electrochemical cells.

Examples of Oxidation Half-Reactions

Species Oxidised	Oxidation Half-Reaction	Explanation
Zinc (Zn)	$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$	Solid zinc loses 2 electrons to form Zn^{2+}.
Iron (Fe)	$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$	Iron (II) ion loses 1 electron to form Fe^{3+}.
Hydrogen gas (H_2)	$H_2(g) \rightarrow 2H^+(aq) + 2e^-$	Hydrogen gas is oxidised to hydrogen ions.
Carbon monoxide (CO)	$CO + H_2O \rightarrow CO_2 + 2H^+ + 2e^-$	CO is oxidised in fuel cells.

Reduction Half-Reactions

A reduction half-reaction shows the part of a redox (oxidation-reduction) reaction where a substance gains electrons. It focuses only on the reduction process, not the full redox equation.

Key Features of Reduction Half-Reactions

Feature	Description
Electron gain	Electrons appear on the reactant side of the equation.
Oxidation number	Decreases during reduction.
Charge balance	Must be balanced for both mass and charge.
Occurs at	The cathode in electrochemical cells.

Examples of Reduction Half-Reactions

Species Reduced	Reduction Half-Reaction	Explanation
Copper (II) ion (Cu^{2+})	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	Cu^{2+} gains 2 electrons to form solid copper.
Iron (III) ion (Fe^{3+})	$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	Fe^{3+} gains 1 electron to form Fe^{2+}.
Hydrogen ion (H^+)	$2H^+(aq) + 2e^- \rightarrow H_2(g)$	Hydrogen ions gain electrons to form hydrogen gas.
Manganate (VII) ion (MnO_4^-)	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	MnO_4^- is reduced to Mn^{2+} in acidic conditions.

Balancing Redox Reactions in Acidic Medium

Balancing redox (oxidation-reduction) reactions involves ensuring both mass and charge are conserved. In an acidic medium, this means using H^+ ions and H_2O molecules to balance atoms and electrons to balance charge.

Systematic Steps

1. **Split the reaction into two half-reactions**

   Identify the species undergoing oxidation and reduction and write each half-reaction separately.

2. **Balance all elements except H and O**

   Start by balancing atoms such as metals or nonmetals that are not hydrogen or oxygen.

3. **Balance oxygen atoms by adding H_2O**

   Add one H_2O molecule for each missing oxygen atom.

4. **Balance hydrogen atoms by adding H^+**

   Add one H^+ ion for each missing hydrogen atom.

5. **Balance charge by adding electrons (e^-)**

   Add electrons to the more positive side to balance the charges.

6. **Multiply half-reactions to equalise electrons**

   Multiply one or both half-reactions so that the number of electrons lost equals those gained.

7. **Add the half-reactions**

   Combine both balanced half-reactions and cancel out electrons and any identical species on both sides (these are known as spectator ions).

8. **Final check**

   Ensure atoms and charges are balanced. Confirm the reaction is appropriate for an acidic medium (no OH^- used).

Example 4.1

Unbalanced equation: $Cr_2O_7^{2-} + HNO_2 \rightarrow Cr^{3+} + NO_3^-$ (Acidic Medium)

Step 1: Split into half-reactions

Oxidation: $HNO_2 \rightarrow NO_3^-$

Reduction: $Cr_2O_7^{2-} \rightarrow Cr^{3+}$

Step 2: Balance atoms except H, O

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+}$

$HNO_2 \rightarrow NO_3^-$

Step 3: Balance O and H

Cr half: $Cr_2O_7^{2-} + 14\ H^+ \rightarrow 2\ Cr^{3+} + 7\ H_2O$

N half: $HNO_2 + H_2O \rightarrow NO_3^- + 3\ H^+$

Step 4: Balance charge with electrons

Cr: $6\ e^- + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

N: $HNO_2 + H_2O \rightarrow NO_3^- + 3\ H^+ + 2\ e^-$

Step 5: Equalise electrons (multiply N half by 3)

Nhalf × 3: $3\ HNO_2 + 3\ H_2O \rightarrow 3\ NO_3^- + 9\ H^+ + 6\ e^-$

Step 6: Combine and cancel duplicates

Final: $Cr_2O_7^{2-} + 3\ HNO_2 + H^+ \rightarrow 2\ Cr^{3+} + 3\ NO_3^- + H_2O$ $Cr_2O_7^{2-} + 3\ HNO_2 + 5H^+ \rightarrow 2\ Cr^{3+} + 3\ NO_3^- + 4H_2O$

Balancing Redox Reactions in Basic Medium

Step-by-Step Method in Basic Solution

1. Divide the reaction into two halfreactions: oxidation and reduction.
2. Balance atoms other than H and O in each half-reaction.
3. Balance oxygen by adding H_2O.
4. Balance hydrogen by adding H^+ as needed.
5. Neutralise H^+ by adding OH^- ions to both sides—this is the critical extra step in basic medium.
6. Combine H^+ and OH^- to form water, then simplify by removing excess H_2O.
7. Balance the charges by adding electrons (e^-) to the more positive side of each halfreaction.
8. Multiply halfreactions so that electrons cancel when added together.
9. Add them, cancel identical species (H_2O, OH^-, e^-), and check that atoms and charges balance.

Example 4.2

$Br^- + MnO_4^- \rightarrow MnO_2 + BrO_3^-$ in Basic Medium

1. Write halfreactions and balance atoms except H and O

   Oxidation: $Br^- \rightarrow BrO_3^-$

   Reduction: $MnO_4^- \rightarrow MnO_2$

   Balancing O and H (in acid medium first)

   $Br^- + 3H_2O \rightarrow BrO_3^- + 6H^+$

   $MnO_4^- + 4H^+ \rightarrow MnO_2 + 2H_2O$

   Add OH^- to both sides of the equations to neutralise H+

   $Br^- + 3H_2O + 6OH^- \rightarrow BrO_3^- + 6H^+ + 6OH^-$

   $MnO_4^- + 4H^+ + 4OH^- \rightarrow MnO_2 + 2H_2O + 4OH^-$

   Combine OH- and H+ to form water and simplify

   $Br^- + 3H_2O + 6OH^- \rightarrow BrO_3^- + 6H_2O$

$MnO_4^- + 4H_2O \rightarrow MnO_2 + 2H_2O + 4OH^-$

Simplify to net amounts

$Br^- + 6OH^- \rightarrow BrO_3^- + 3H_2O$

$MnO_4^- + 2H_2O \rightarrow MnO_2 + 4OH^-$

Balancing charges using electrons

$Br^- + 6OH^- \rightarrow BrO_3^- + 3H_2O + 6e^-$

$MnO_4^- + 2H_2O + 3e^- MnO_2 + 4OH^-$

Equalise electrons by scaling halfreactions accordingly.

$Br^- + 6OH^- \rightarrow BrO_3^- + 3H_2O + 6e^-$

$2MnO_4^- + 4H_2O + 6e^- \rightarrow 2MnO_2 + 4OH^-$

Cancel water and OH^- duplicates, simplify.

Final balanced net ionic: $H_2O + 2MnO_4^- + Br^- \rightarrow 2MnO_2 + BrO_3^- + 2OH^-$

Example 4.3

Balance the reaction in basic medium: $ClO^- + Mn(OH)_2 \rightarrow MnO_2 + Cl^-$

Step 1: Assign oxidation numbers

ClO^-: Cl is +1

Cl^-: Cl is –1 → reduction

$Mn(OH)_2$: Mn is +2

MnO_2: Mn is +4 → oxidation

Step 2: Write half-reactions

Oxidation (Mn^{2+} to Mn^{4+} in MnO_2): $Mn(OH)_2 \rightarrow MnO_2$

Reduction (ClO^- to Cl^-): $ClO^- \rightarrow Cl^-$

Step 3: Balance each half-reaction

Oxidation Half-Reaction (Mn)

1. Balance Mn: Already balanced
2. Balance O: $Mn(OH)_2$ has 2 OH (2 O, 2 H), MnO_2 has 2 O

   Add 2 H^+ to RHS to match 2 H in LHS

   $Mn(OH)_2 \rightarrow MnO_2 + 2H^+$
3. Neutralise H^+: Add 2 OH^- to both sides

   $Mn(OH)_2 + 2OH^- \rightarrow MnO_2 + 2H^+ + 2OH^-$

4. $Mn(OH)_2 + 2OH^- \rightarrow MnO_2 + 2H_2O$Balance charge:

   LHS: $Mn(OH)_2$ is neutral 0+ -2 from (OH-) = -2

   RHS: MnO_2 (neutral) 0 + 2 OH^- (–2) + 2 from ($2H^+$) = 0

   Add 2 e^- to RHS

   Oxidation half-reaction: $Mn(OH)_2 + 2OH^- \rightarrow MnO_2 + 2H_2O + 2e^-$

Reduction Half-Reaction (Cl)

1. Balance Cl: already balanced
2. Balance O: ClO^- has 1 O → add 1 H_2O to RHS
3. Balance H:2 H^+ to RHS then $2OH^-$ to both sides of the equation

   $ClO^- + 2H^+ + 2OH^- \rightarrow Cl^- + H_2O + 2OH^-$
4. Balance charge:

   LHS: $ClO^- + 2\ OH^- + 2H^+ = -1$

   RHS: $Cl^- = -1$, $H_2O = 0$, 2OH- = -2→ net = -3

   Add 2 e^- to LHS

Reduction half-reaction: $ClO^- + 2H^+ + 2OH^- + 2e^- \rightarrow Cl^- + H_2O + 2OH^-$

$ClO^- + H_2O^- + 2e^- \rightarrow Cl^- + 2OH^-$

Step 4: Add both half-reactions

$Mn(OH)_2 + 2OH^- \rightarrow MnO_2 + 2H_2O + 2e^-$

$ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$

Add and cancel electrons:

$Mn(OH)_2 + ClO^- + 2OH^- \rightarrow MnO_2 + 2OH^- + Cl^- + H_2O$

Cancel **$2OH^-$** from both sides (spectator ions):

Balanced Redox Reaction in Basic Medium: $Mn(OH)_2 + ClO^- \rightarrow MnO_2 + Cl^- + H_2O$

Activity 4.4 Understanding Redox Reactions



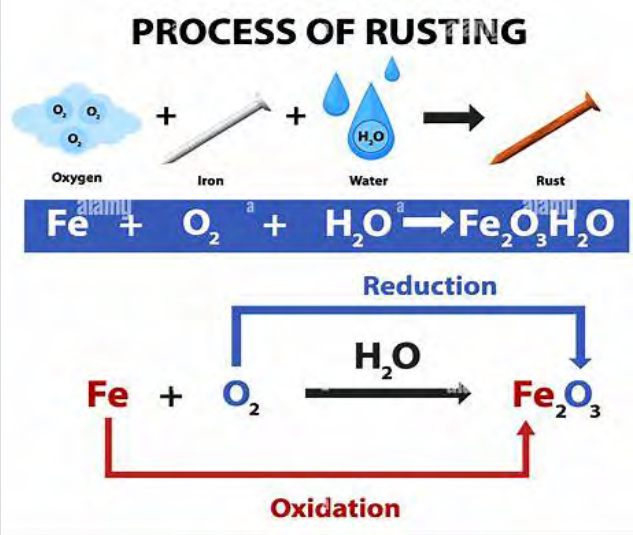
Figure 4.3: Images of rusting iron roofs (typical in Ghanaian buildings).



Figure 4.4: Image of rusting



Figure 4.5: Imagery related to gold mining (galamsey/artisanal sites)



1. Observe the images in Figures 4.3-4.5
   - **a.** Describe what you observed in these images.
   - **b.** What chemical changes might be happening in all situations?

2. Consider the reaction: $4Fe + 3O_2 + H_2O \rightarrow 2Fe_2O_3.H_2O$
   - **a.** Break the reaction down into oxidation and reduction halves.
   - **b.** Which species is oxidised? Which is reduced? How can you tell?

3. **a.** Model galvanic action like water pipes or roofs in coastal towns
   $Cu^{2+} + Fe^0 \rightarrow Cu^0 + Fe^{2+}$
   - i. Assign oxidation states
   - ii. Identify which is oxidised/reduced
   - iii. Write corresponding halfequations

   b. Represents gold recovery via the displacement method used in artisanal mining
   $\mathbf{Au^{3+} + Zn^0 \rightarrow Au^0 + Zn^{2+}}$
   - i. Write and balance oxidation and reduction half-reactions
   - ii. Ensure electrons cancel and show the balanced overall equation

   c. This redox pair can relate to water contamination issues near mining sites
   $\mathbf{Fe^{3+} + SO_3^{2-} \rightarrow Fe^{2+} + SO_4^{2-}}$
   - i. Balance halfreactions in an acidic medium
   - ii. Combine into a full balanced reaction

d. Balance halfreactions and write the overall balanced reactions

i. $MnO_4^- + C_2O_4^{2-} \rightarrow Mn^{2+} + CO_2$ (acidic medium)

ii. $Pb(OH)_4^{2-} + ClO^- \rightarrow PbO_2 + Cl^-$ (basic medium) Show handling of OH^- and H_2O to balance O, H, electrons.

iii. $Fe(OH)_2 \rightarrow Fe_3O_4 + H_2$ (basic/anaerobic corrosion context)

4. Identify one redox process in your home or community.

REDOX TITRATIONS

Redox Titrations

Redox titrations are a way to measure how much of a substance is in a solution using a redox reaction. This is different from acid–base titrations, where the reaction involves hydrogen ions (H^+) instead of electrons.

When Does the Reaction End?

1. The reaction ends when the right amount of each substance has been added and they have completely reacted.
2. This point is called the equivalence point – it is when the electrons lost and gained are equal.

How Do We Know It's Done?

1. A colour change tells us the titration is complete.
2. Some reactions change colour on their own (self-indicating).
3. Others need a redox indicator to help us see the change.

Types of Redox Titrations

1. Permanganate titrations: utilising the strong oxidising power of MnO_4^-
2. Dichromate titrations: using $Cr_2O_7^{2-}$ as the oxidising agent
3. Iodometric/iodimetric titrations: based on the redox chemistry of iodine

Summary Table

Type	Titrant	Colour Change	Indicator	Common Use
Permanganate	$KMnO_4$	Purple → Colourless	Self-indicating	Fe^{2+}, H_2O_2

Type	Titrant	Colour Change	Indicator	Common Use
Dichromate	$K_2Cr_2O_7$	Orange → Green/Blue	Diphenylamine	Fe^{2+}, Alcohols
Iodometric	$Na_2S_2O_3$	Brown → Blue → colourless	Starch	Bleach, Cu^{2+}, H_2O_2
Iodimetric	I_2	Yellow-brown → colourless	None/Starch	Reducing agents (e.g. SO_3^{2-})

Standardisation of Potassium Tetraoxomanganate (VII) ($KMnO_4$) Solutions

Potassium permanganate ($KMnO_4$) is a strong oxidising agent, but it is not a primary standard because:

1. it is not 100% pure (may contain MnO_2)
2. it decomposes over time, especially in light or in neutral/alkaline solution

Therefore, it must be standardised (accurately determined) before use in titrations.

Standardisation Procedure

To determine the exact concentration of $KMnO_4$ by titration with a primary standard reducing agent, such as:

1. Ethanedioic acid (oxalic acid), $H_2C_2O_4$
2. Iron (II) sulfate, $FeSO_4$
3. Sodium oxalate, $Na_2C_2O_4$

Example: Standardising $KMnO_4$ with $FeSO_4$

Apparatus & Chemicals

1. Burette, pipette, conical flask
2. $KMnO_4$ solution (in burette)
3. Standard $FeSO_4$ solution (in flask)
4. Dilute H_2SO_4 (acidic medium)
5. White tile

Reaction

$MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

Procedure

1. Rinse burette and fill with $KMnO_4$ solution.
2. Pipette 25 cm^3 of $FeSO_4$ into conical flask.
3. Add about 20 cm^3 of dilute H_2SO_4.
4. Titrate with $KMnO_4$ until the first permanent pale pink colour appears.
5. Repeat for concordant titres.

Calculation

Use:

1. Volume and known molarity of Fe^{2+}
2. Balanced redox equation
3. Moles of Fe^{2+} → moles of $KMnO_4$
4. Calculate molarity of $KMnO_4$ using:

$$\text{Molarity} = \frac{\text{Moles of } KMnO_4}{\text{Volume used (in } dm^3)}$$

Precautions

1. Use acidic medium only (dil. H_2SO_4) – never HCl (Cl^- gets oxidised)
2. Protect $KMnO_4$ solution from light and organic contamination
3. Use freshly prepared Fe^{2+} solution to avoid oxidation by air

Example 4.4: Standardising $KMnO_4$ with $FeSO_4$

To find the exact concentration of a $KMnO_4$ solution by titrating it against a 0.0200 mol/dm^3 $FeSO_4$ solution.

Titration Data

Volume of $FeSO_4$ solution = 25.00 cm^3

Molarity of $FeSO_4$ = 0.0200 mol/dm^3

Volume of $KMnO_4$ used = 20.00 cm^3

Balanced Redox Equation

$MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

From the equation, 1 mole of $KMnO_4$ reacts with 5 moles of Fe^{2+}

Step-by-Step Calculation

Step 1: Calculate moles of Fe^{2+}

Moles of $Fe^{2+} = \frac{25.00}{1000} \times 0.0200 = 5.00 \times 10^{-4}$ mol

Step 2: Use mole ratio to find moles of $KMnO_4$

From the equation:

Moles of $KMnO_4 = 1 \Rightarrow \times 5.00 \times 10^{-4} = 1.00 \times 10^{-4}$ mol

Step 3: Calculate molarity of $KMnO_4$

Molarity of $KMnO_4 = \frac{\text{Moles}}{\text{Volume in dm}^3} = \frac{1.00 \times 10^{-4}}{20.00} \times 1000 = 0.00500$ mol dm^{-3}

The exact concentration of the $KMnO_4$ solution is 0.00500 mol/dm^3

Standardisation of $KMnO_4$ with Ethanedioic Acid

Determining the exact concentration of potassium permanganate ($KMnO_4$) solution using a primary standard solution of ethanedioic acid (oxalic acid, $H_2C_2O_4$).

Balanced Redox Equation

$2MnO_4^- + 5H_2C_2O_4 + 6H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

Given Data

Volume of ethanedioic acid solution = 25.00 cm^3

Molarity of ethanedioic acid = 0.0200 mol/dm^3

Volume of $KMnO_4$ used = 20.00 cm^3

Step-by-Step Calculation

1. Calculate the moles of $H_2C_2O_4$ used:
   Moles = $(\frac{25.00}{1000}) \times 0.0200 = 5.00 \times 10^{-4}$ mol
2. Use the mole ratio from the balanced equation:
   Moles of $KMnO_4 = (\frac{2}{5}) \times 5.00 \times 10^{-4} = 2.00 \times 10^{-4}$ mol
3. Calculate the concentration of $KMnO_4$
   Molarity = moles / volume (in dm^3) = $\frac{2.00 \times 10^{-4}}{(\frac{20.00}{1000})} = 0.0100$ mol/dm^3

The exact concentration of the $KMnO_4$ solution is 0.0100 mol/dm^3.

Standardisation Using Sodium Thiosulphate (Indirect Method)

Standardisation of Sodium Thiosulphate – Indirect Iodometric Method

Determining the exact concentration of sodium thiosulphate ($Na_2S_2O_3$) solution using the indirect iodometric method. This involves generating iodine from a primary standard (e.g., potassium dichromate) and titrating it with sodium thiosulphate.

Chemical Reactions Involved

Step 1: Iodine generation from potassium dichromate:

$$Cr_2O_7^{2-} + 14H^+ + 6I^- \rightarrow 2Cr^{3+} + 3I_2 + 7H_2O$$

Step 2: Titration of iodine with sodium thiosulphate:

$$I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$$

Apparatus and Reagents

1. Standard potassium dichromate solution ($K_2Cr_2O_7$)
2. Excess potassium iodide (KI)
3. Dilute sulphuric acid (H_2SO_4) or hydrochloric acid (HCl)
4. Sodium thiosulphate solution ($Na_2S_2O_3$)
5. Starch indicator (freshly prepared)
6. Burette, pipette, conical flask, white tile

Procedure

1. Pipette 25 cm^3 of standard $K_2Cr_2O_7$ solution into a conical flask.
2. Add excess KI and acidify with dilute H_2SO_4.
3. Iodine (I_2) is liberated and turns the solution brown.
4. Titrate immediately with $Na_2S_2O_3$ solution from the burette.
5. When the solution turns pale yellow, add a few drops of starch indicator.
6. Continue titration until the blue colour disappears (end point).
7. Repeat to obtain concordant titres.

Calculations

1. Calculate the moles of $K_2Cr_2O_7$ used (primary standard).
2. Determine the moles of I_2 produced using the balanced redox equation.

3. Use the stoichiometry to find moles of $Na_2S_2O_3$ that reacted.
4. Calculate the concentration of $Na_2S_2O_3$: Molarity = moles / volume (in dm^3)

Example 4.5

Calculation – Standardisation of Sodium Thiosulphate

This example demonstrates how to calculate the exact concentration of sodium thiosulphate ($Na_2S_2O_3$) solution by titrating it with iodine generated from potassium dichromate ($K_2Cr_2O_7$).

Relevant Equations

Iodine generation: $Cr_2O_7^{2-} + 14H^+ + 6I^- \rightarrow 2Cr^{3+} + 3I_2 + 7H_2O$

Titration with thiosulphate: $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$

Given Data

1. Volume of $K_2Cr_2O_7$ used = 25.00 cm^3
2. Molarity of $K_2Cr_2O_7$ = 0.0200 mol/dm^3
3. Volume of $Na_2S_2O_3$ used = 30.00 cm^3

Step-by-Step Calculation

1. Calculate moles of $K_2Cr_2O_7$: Moles = $(25.00 / 1000) \times 0.0200 = 5.00 \times 10^{-4}$ mol
2. Use mole ratio from the equation: 1 mol $Cr_2O_7^{2-}$ produces 3 mol I_2
   So, moles of $I_2 = 5.00 \times 10^{-4} \times 3 = 1.50 \times 10^{-3}$ mol
3. Use $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$: 1 mol I_2 reacts with 2 mol $Na_2S_2O_3$
   Moles of $Na_2S_2O_3 = 1.50 \times 10^{-3} \times 2 = 3.00 \times 10^{-3}$ mol
4. Calculate concentration of $Na_2S_2O_3$: Molarity = moles / volume in dm^3 = $3.00 \times 10^{-3} / (30.00 / 1000) = 0.100$ mol/dm^3

The exact concentration of the $Na_2S_2O_3$ solution is 0.100 mol/dm^3.

Standardisation of Sodium Thiosulphate Using Potassium Iodate (KIO_3)

This guides learners through the process of determining the exact concentration of sodium thiosulphate ($Na_2S_2O_3$) solution using potassium iodate (KIO_3), a stable primary standard. Iodine is liberated by reaction with iodide ions in acidic medium and then titrated with sodium thiosulphate.

Key Chemical Reactions

1. Iodine generation: $IO_3^- + 5I^- + 6H^+ \rightarrow 3I_2 + 3H_2O$
2. Iodine titration with sodium thiosulphate: $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$

Example Data

Volume of KIO_3 solution = 25.00 cm^3

Molarity of KIO_3 = 0.0200 mol/dm^3

Volume of $Na_2S_2O_3$ used = 36.00 cm^3

Step-by-Step Calculation

1. Calculate moles of KIO_3: Moles = (25.00 / 1000) × 0.0200 = 5.00×10^{-4} mol
2. Determine moles of I_2 produced (from 1 mol KIO_3 → 3 mol I_2):
   Moles of $I_2 = 5.00 \times 10^{-4} \times 3 = 1.50 \times 10^{-3}$ mol
3. Calculate moles of $Na_2S_2O_3$ needed (1 mol I_2: 2 mol $Na_2S_2O_3$):
   Moles of $Na_2S_2O_3 = 2 \times 1.50 \times 10^{-3} = 3.00 \times 10^{-3}$ mol
4. Determine concentration of $Na_2S_2O_3$:
   Molarity = moles / volume (in dm^3) = 3.00×10^{-3} / (36.00 / 1000) = 0.0833 mol/dm^3

The concentration of the sodium thiosulphate solution is 0.0833 mol/dm^3.

Example 4.6: *Redox Titration Example – Potassium Permanganate and Iron (II)*

Chemicals and Apparatus

1. Solution A: 0.05 mol/dm^3 potassium tetraoxomanganate (VII) ($KMnO_4$)
2. Solution B: Iron (II) solution of unknown concentration
3. Dilute sulphuric acid (1M)
4. Burette (50 cm^3), Pipette (25 cm^3), Conical flask (250 cm^3)
5. White tile, Retort stand and clamp

Procedure

1. Set up the titration apparatus and fill the burette with solution A ($KMnO_4$).
2. Using a pipette, transfer exactly 25.0 cm^3 of solution B into a clean conical flask.
3. Add approximately 10 cm^3 of dilute sulphuric acid and swirl gently to mix.
4. Place the conical flask on a white tile under the burette.

5. Titrate against solution A until a permanent pale pink colour appears (endpoint).
6. Record your readings in a titration table.
7. Repeat the titration to obtain at least two concordant titres and calculate the average.

Balanced Redox Equation

$MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

Analysis and Questions

1. Using your results and the equation above, calculate the concentration of Fe^{2+} in mol/dm³.
2. Solution B was prepared by dissolving 71.5 g of ammonium iron (II) sulphate hydrate, $(NH_4)_2Fe(SO_4)_2 \cdot xH_2O$, per dm³ of solution. Calculate the value of x, the number of water molecules of crystallisation in the salt.
3. A groundwater sample from a community near a mining site contains iron. The water appears clear when collected but turns reddish-brown after standing. Explain the chemical transformations occurring in the water and how this titration method accounts for these changes.

Answer

1. Concentration of Fe^{2+}

   Assume average titre of $KMnO_4$ used = 23.50 cm³ (0.02350 dm³)

   Molarity of $KMnO_4$ = 0.050 mol/dm³

   Moles of $KMnO_4 = 0.050 \times 0.02350 = 1.175 \times 10^{-3}$ mol

   From the balanced equation:

   1 mol MnO_4^- reacts with 5 mol Fe^{2+}

   Moles of $Fe^{2+} = 5 \times 1.175 \times 10^{-3} = 5.875 \times 10^{-3}$ mol

   Volume of Fe^{2+} solution used = 25.00 cm³ = 0.02500 dm³

   Concentration of $Fe^{2+} = 5.875 \times 10^{-3} / 0.02500 = 0.235$ mol/dm³

2. Value of x in $(NH_4)_2Fe(SO_4)_2 \cdot xH_2O$

   Mass of salt used per dm³ = 71.5 g

   Moles of Fe^{2+} = 0.235 mol (from above)

   Molar mass of anhydrous part = (2×18.04) + 55.85 + (2×96.06) = 284.06 g/mol

   Let x = number of water molecules (18.02 g/mol each)

   Molar mass = 284.06 + 18.02x

$71.5 / (284.06 + 18.02x) = 0.235$

$\Rightarrow 284.06 + 18.02x = 71.5 / 0.235 \approx 304.26$

$\Rightarrow 18.02x = 304.26 - 284.06 = 20.2$

$\Rightarrow x \approx 1.12 \rightarrow$ round to 1 (indicating monohydrate, but commonly $x \approx 6$ for hexahydrate)

Note that actual known hydrate is $x = 6$

3. Iron Transformation in Groundwater

Fresh groundwater contains soluble Fe^{2+} ions, which are colourless.

Upon exposure to air (oxygen), Fe^{2+} is oxidised to Fe^{3+}:

$4Fe^{2+} + O_2 + 4H^+ \rightarrow 4Fe^{3+} + 2H_2O$

Fe^{3+} reacts with water to form reddish-brown precipitate of iron (III) hydroxide:

$Fe^{3+} + 3H_2O \rightarrow Fe(OH)_3 + 3H^+$

The titration method detects Fe^{2+} before oxidation or after converting Fe^{3+} back to Fe^{2+}.

Activity 4.5 Redox Titration

1. Laboratory preparation

Set up workstations with:

a. Equipment: burettes, pipettes, flasks, beakers, measuring cylinders, stands, clamps, thermometers, hot water baths.

b. Labelled Solutions:

i. Standard: $KMnO_4$ (permanganate), $Na_2S_2O_3$ (thiosulphate)

ii. Unknowns: Fe^{2+}, $C_2O_4^{2-}$ (oxalate), I_2/KI

2. Form groups with roles:

a. Equipment Manager

b. Titrator (Operator)

c. Recorder

d. Analyst/Reporter

3. Procedure Distribution

a. Redox Titration Procedure Sheets

b. Conduct the titrations on rotational bases

Titration 1: $KMnO_4$ vs Fe^{2+}

Purpose: To determine the concentration of Fe^{2+} using standard $KMnO_4$.

You may use the link below to watch a video on titration of $KMnO_4$ vs Fe^{2+}

https://www.youtube.com/watch?v=ci4cHGLVZQY

Materials

a. 0.020 mol/dm³ potassium permanganate ($KMnO_4$)
b. Iron (II) sulphate or Mohr's salt (unknown concentration)
c. Dilute sulphuric acid (H_2SO_4)
d. Burette, pipette, conical flask, white tile, funnel, clamp stand

Procedure

a. Rinse and fill the burette with $KMnO_4$ solution. Record initial volume.
b. Pipette 25.0 cm³ of Fe^{2+} solution into a conical flask.
c. Add about 15 cm³ of dilute H_2SO_4.
d. Titrate with $KMnO_4$ while swirling.
e. Stop when a faint pink colour persists for 30 seconds.
f. Record final volume and repeat for consistent titres.

Observations

a. $KMnO_4$ is deep purple.
b. Fe^{2+} solution becomes faint pink at endpoint.

Safety

a. $KMnO_4$ stains—avoid contact.
b. H_2SO_4 is corrosive—use gloves and goggles.

Titration 2: $KMnO_4$ vs Ethanedioate ($C_2O_4^{2-}$)

Purpose: To determine the concentration of sodium ethanedioate using standard $KMnO_4$.

You may use the link below to watch a video on titration of $KMnO_4$ vs $C_2O_4^{2-}$ https://www.youtube.com/watch?v=GkroZP3oNEw

Materials

a. 0.020 mol/dm³ $KMnO_4$
b. Sodium ethanedioate solution (unknown)

c. Dilute sulphuric acid
d. Water bath or hotplate, thermometer
e. Burette, pipette, conical flask

Procedure

a. Pipette 25.0 cm³ of ethanedioate solution into a conical flask.
b. Add 20 cm³ of dilute H_2SO_4.
c. Warm the flask in a 60–70 °C water bath.
d. Fill the burette with $KMnO_4$ and record initial volume.
e. Titrate while swirling; endpoint is faint pink persisting.
f. Repeat to get concordant titres.

Observations

a. Purple $KMnO_4$ is decolourised.
b. Faint pink appears at endpoint in warm solution.

Safety

a. Use thermometer to monitor temperature.
b. Avoid skin contact with $KMnO_4$ and H_2SO_4.

Titration 3: Iodine/Potassium Iodide vs Sodium Thiosulphate

Purpose: To determine the iodine concentration using sodium thiosulphate. You may use the link below to watch a video on titration of Iodine/Potassium Iodide vs Sodium Thiosulphate

https://www.youtube.com/watch?v=lkJyP6OYmAY&t=255s

Materials

a. Iodine/potassium iodide solution (I_2/KI)
b. 0.1 mol/dm³ sodium thiosulphate ($Na_2S_2O_3$)
c. Fresh starch solution
d. Burette, pipette, conical flask

Procedure

a. Pipette 25.0 cm³ of iodine solution into a conical flask.
b. Add 10 cm³ of potassium iodide (KI).
c. Titrate with sodium thiosulphate.
d. When brown fades to pale yellow, add 1–2 cm³ starch solution.
e. Continue titration until solution turns colourless.

f. Repeat for accuracy.

Observations

a. Brown iodine fades.

b. Blue-black colour appears with starch.

c. Solution becomes colourless at endpoint.

Safety

a. Iodine is an irritant—handle with care.

b. Use starch only near the endpoint. (when the solution turns pale yellow)

4. Problem-solving and analysis

a. Complete data tables: Record titration volumes and average results for all three reactions as shown in the titration table.

Titration table format

Burette readings/ cm^3	1	2	3
Final reading	22.50	22.20	22.30
Initial reading	0.00	0.00	0.00
Titre	22.50	22.20	22.30

Average titre $= \frac{22.20 + 22.30}{2} = 22.25cm3$

b. Balanced redox equations: Review balanced equations and identify mole ratios

$MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$ (mole ratio 1:5)

$2MnO_4^- + 5C_2O_4^- + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$ (mole ratio 2:5)

$I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$ (mole ratio 1:2)

c. Calculations (with units and ratios): Calculate Unknown Concentrations

$$\frac{C_{MnO_4^-} V_{MnO_4^-}}{C_{Fe^{2+}} V_{Fe^{2+}}} = 1 \_ \_$$

$$\frac{C_{MnO_4^-} V_{MnO_4^-}}{C_{C_2O_4^-} V_{C_2O_4^-}} = 2 \_ \_$$

$$\frac{C_{I_2} V_{I_2}}{C_{S_2O_3^{2-}} V_{S_2O_3^{2-}}} = 1 \_ \_$$

5. Identify potential experimental errors
   - **a.** Endpoint detection difficulties
   - **b.** Air bubbles in burette
   - **c.** Misreading meniscus
   - **d.** Temperature variations
6. Relate calculated concentrations to:
   - **a.** water quality standards
   - **b.** iron content in supplements
   - **c.** vitamin C and iodine analysis

ELECTROCHEMICAL CELLS

Electrochemistry

What is Electrochemistry?

Electrochemistry is the study of how electricity and chemical reactions are connected. It helps us understand:

1. how batteries work
2. why some metals rust /corrode
3. how we can use electricity to break or decompose substances apart (like water into hydrogen and oxygen)

In electrochemistry, we look at how electrons move from one substance to another. This movement can either:

1. Produce electricity (like in batteries), or
2. Use electricity to cause a chemical change (like in electrolysis).

Electrochemical Cell

An electrochemical cell is a device that uses chemical reactions to produce electricity or uses electricity to cause chemical changes. There are two main types of electrochemical cells:

1. Voltaic (or Galvanic) Cells – Generate electricity from chemical reactions (like in batteries).
2. Electrolytic Cells – Use electricity to make chemical reactions happen (like splitting water into gases).

Inside an electrochemical cell, electrons move from one part of the cell to another. This movement of electrons creates electric current. The cell has two metal parts called electrodes, dipped in a liquid called an electrolyte.

Example: A Simple Voltaic Cell

If you connect a zinc rod (Zn) and a copper rod (Cu) in two solutions and join them with a wire and a salt bridge, electricity flows. This is a simple battery — a voltaic cell in action.

When a metal plate is dipped into a solution containing its own ions, an electrochemical process begins.

Example: Zinc metal in zinc ion solution (Zn in Zn^{2+})

What happens?

1. Equilibrium is set up between the metal atoms and metal ions: $Zn\ (s) \rightleftharpoons Zn^{2+}(aq) + 2e^-$
2. Two things can occur depending on the metal and the ion concentration:

a. Oxidation: Some metal atoms lose electrons and go into the solution as positive ions.

b. Reduction: Some metal ions in solution gain electrons and stick to the metal plate as solid metal.

3. This causes a build-up of electrical charge on the metal plate, and a potential difference (voltage) is created between the metal and the solution.

Result

1. The metal becomes positively or negatively charged relative to the solution.
2. If connected in a complete circuit with another metal and its ions, an electrochemical cell is formed, and electricity can flow.

This process is the basis of all electrochemical cells – where metals "push" or "pull" electrons based on how easily they form ions.

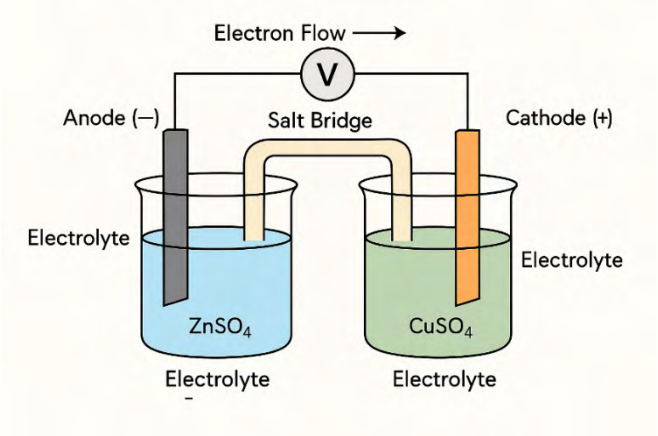


Figure 4.6: Diagram of an electrochemical cell (Voltaic cell)

Composition of an Electrochemical Cell

Electrochemical Cell consists of the following main components

1. **Electrodes:** Electrodes are solid conductors, usually metals, where redox reactions take place.

   Anode: Site of oxidation (loss of electrons). Electrons are released here. In voltaic cells, the anode is negative.

   Cathode: Site of reduction (gain of electrons). Electrons are accepted here. In voltaic cells, the cathode is positive.
2. **Electrolytes:** Electrolytes are ionic solutions that allow the movement of ions between electrodes and complete the internal circuit. They may be molten (melted) salts or aqueous solutions of salts
3. **Salt or ion Bridge (or Porous Partition):** A salt bridge allows ion flow between the two half-cells. It maintains electrical neutrality by letting ions balance charges.
4. **External Circuit:** The external circuit is usually a wire that connects the electrodes. Electrons flow from the anode to the cathode.
5. **Voltmeter:** The voltmeter is used to measure the potential difference (voltage) between the electrodes. It shows the cell's electrical output.

Electrode Potential

This is the measure of the tendency of an electrode (metal or non-metal) in contact with its ions in solution to gain or lose electrons, resulting in the creation of an electric potential (voltage) between the electrode and the solution. In an *electrochemical cell*, each half-cell has its own electrode potential. When two

half-cells are connected, electrons flow from the electrode with the lower (more negative) potential to the one with the higher (more positive) potential. This flow of electrons generates electricity.

Standard Hydrogen Electrode (SHE)

The Standard Hydrogen Electrode (SHE) is the reference electrode used to measure the standard electrode potentials (E°) of other half-cells.

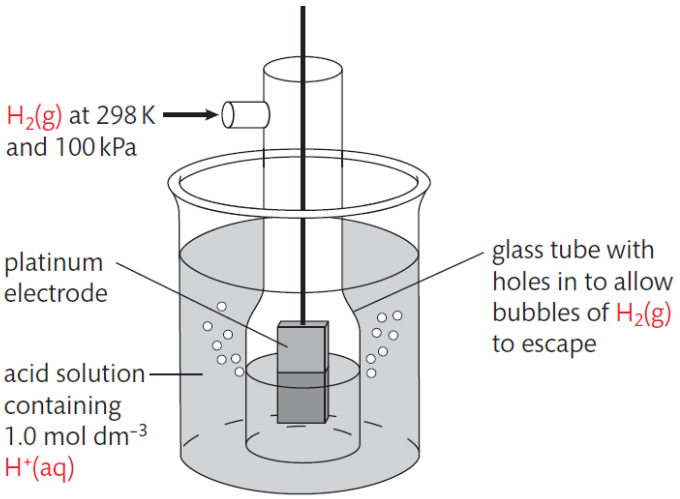


Figure 4.7: A standard hydrogen electrode

$Pt(s) \mid H_2(g, 1\ atm) \mid H^+(aq, 1\ M)$

Features of the SHE

Electrode reaction: $2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$

Electrode potential (E°): 0.00 V (by definition)

Electrode material: Platinum (Pt) – inert and conducts electrons

Gas used: Hydrogen gas (H_2)

Pressure of gas: 1 atm or 100 kPa

Concentration of H^+: 1.0 mol/dm^3 (1 M solution, e.g., HCl)

Temperature: 298 K (25°C) – standard temperature

Function of SHE

1. Acts as a reference half-cell in electrochemical measurements.
2. Used to measure standard electrode potentials of other half-cells.

Uses of the SHE

1. Measuring Standard Electrode Potentials: Connect the SHE to another half-cell to form a complete electrochemical cell and measure the voltage.
2. Comparing Oxidising/Reducing Powers
   a. Positive E°: species is a better oxidising agent than H^+.
   b. Negative E°: species is a better reducing agent than H_2.

Limitations

1. Difficult to set up and maintain under laboratory conditions.
2. Platinum surface can be poisoned by impurities.
3. Not practical for routine lab use – replaced by secondary reference electrodes (e.g., calomel or Ag/AgCl).

Standard Electrode Potential (E°)

To compare different electrodes, we use standard electrode potentials measured under standard conditions. All values are compared against the Standard Hydrogen Electrode (SHE), which is assigned a potential of 0.00 V.

The *Standard Electrode Potential*, symbol E°, is the voltage of a half-cell under standard conditions, measured relative to the Standard Hydrogen Electrode (SHE), which is defined as 0.00 V.

Standard Conditions

1. Concentration: 1.0 mol/dm³ of ions
2. Temperature: 298 K (25°C)
3. Pressure: 1 atm (for gases)
4. Electrode: A clean metal or inert electrode (like platinum) if the substance is not a metal

E° values are measured under standard conditions. A positive E° means the half-cell has a greater tendency to gain electrons (be reduced) than the hydrogen electrode. A negative E° means the half-cell is more likely to lose electrons (be oxidised). This is used to predict the direction of redox reactions and to calculate cell potentials.

Example

1. $E°(Zn^{2+}/Zn) = -0.76$ V → zinc prefers to lose electrons (oxidation) when compared to the SHE.

2. $E°(Cu^{2+}/Cu) = +0.34\ V \rightarrow$ copper prefers to gain electrons (reduction) when compared to the SHE

Applications

1. Predicting redox reaction feasibility or spontaneity
2. Designing electrochemical cells
3. Determining which species acts as oxidising or reducing agent
4. Used to predict the emf (electromotive force) of an electrochemical cell

Factors That Affect Electrode Potential Values

The main factors that affect electrode potential values are:

1. **Concentration of ions**

   Changes in the concentration of the metal ions in solution can shift the electrode potential. According to the Nernst equation, increasing the concentration of the ions involved in the half-cell increases the electrode potential for reduction reactions and decreases it for oxidation reactions.

   The Nernst Equation

   The Nernst Equation is used to calculate the electrode potential of a half-cell or full electrochemical cell under non-standard conditions. It accounts for the effect of ion concentration or gas pressure on the cell potential. The two forms are:

   a. Base-10 Logarithmic Form (Common in Textbooks)

      Standard Form (at 25°C or 298 K)

      $$E = E° - \frac{0.0591}{n}\log\left(\frac{[Red]}{[Ox]}\right)$$

   b. Full Nernst Equation

      $$E = E° - \frac{RT}{nF}\ln\left(\frac{[Red]}{[Ox]}\right)$$

   Where

   E = electrode potential (V)

   E° = standard electrode potential (V)

   R = gas constant = 8.314 J mol^{-1} K^{-1}

   T = temperature in Kelvin (K)

   n = number of electrons transferred

   F = Faraday constant = 96500 C mol^{-1}

[Red], [Ox] = concentrations or partial pressures of reduced and oxidised species

Example 4.7

Cell reaction: $Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$

Standard electrode potential, $E° = +0.77$ V

Temperature = 25°C = 298 K

$[Fe^{2+}] = 0.10$ mol/dm³, $[Fe^{3+}] = 1.00$ mol/dm³

n = 1 (number of electrons transferred)

a. Using the Base-10 Logarithmic Form

$$E = E° - \frac{0.0591}{n}\log\left(\frac{[Red]}{[Ox]}\right)$$

$$E = 0.77 - \frac{0.0591}{1}\log\left(\frac{0.10}{1.00}\right)$$

$$= 0.77 - 0.0591(-1)$$

$$= 0.77 + 0.0591 = 0.8291\ V$$

b. Using the Natural Logarithmic Form

$$E = E° - \frac{RT}{nF}\ln\left(\frac{[Red]}{[Ox]}\right)$$

$$E = 0.77 - \frac{8.314 \times 298}{1 \times 96500}\ln\left(\frac{0.10}{1}\right)$$

$$= 0.77 - 0.0257 \times (-2.3026)$$

$$= 0.77 + 0.0591 = 0.8291\ V$$

Conclusion

Both forms of the Nernst equation give the same result (0.8291 V) at 25°C.

- Use the simplified log version at 25°C for quick calculations.
- Use the natural log version (ln) for accuracy and variable temperature calculations.

2. **Temperature:** Electrode potentials vary with temperature.

   The standard electrode potential is measured at 25°C (298 K). Deviating from this temperature can affect reaction kinetics and equilibrium, thereby altering the potential.

3. **Pressure (for Gaseous Electrodes):** Affects electrode potential when gases are involved (e.g. in hydrogen or oxygen electrodes). Changes in the

partial pressure of a gas can shift the position of equilibrium and thus affect the potential.

4. **Nature of the Electrode Metal or Material:** Different materials have different tendencies to lose or gain electrons. The intrinsic properties of the metal (e.g., ionisation energy, electronegativity, lattice energy) influence its electrode potential.
5. **Presence of Complexing Agents:** Complexation can lower or raise the free ion concentration, thus changing the electrode potential. Complexes can stabilise certain oxidation states, shifting the equilibrium and affecting the potential (e.g., $Ag^+ + 2NH_3 \rightarrow [Ag(NH_3)_2]^+$).
6. **Type of Solvent:** The solvent can influence ion solubility and reaction kinetics. Solvent properties like dielectric constant and solvation energy affect the electrode reaction and therefore the potential.
7. **Surface Condition of the Electrode:** Oxidation, corrosion, or impurities on the electrode surface can alter potential. Surface films may inhibit electron transfer or change the effective surface area.

Activity 4.6 Understanding Electrode Potential

Step 1: Team up with your peers to explore and carry out this task. In your group appoint a recorder, researcher, presenter, and coordinator.

You may use the video link below to carry out *Activity 4.6*.

https://chatgpt.com/c/68931de5-8d9c-832a-9226-f87bcc491972

Step 2: In your groups discuss

a. What happens when a metal is placed in a solution of its own ions?

b. Why does a potential difference arise?

c. Use drawings to explain how a potential is developed at the metal-solution interface.

d. What determines the direction and magnitude of electron movement?

Step 3: Research and discuss how the following factors affect electrode potential:

Factor	Research Questions
(a) Nature of electrode	How does the metal's reactivity influence potential?
(b) Temperature of electrolyte	What happens to the potential when temperature increases or decreases?
(c) Concentration of ions	How does ion concentration shift the equilibrium and potential?
(d) Pressure of gas (for gaseous electrodes)	How does gas pressure (e.g., $H_2(g)$ in SHE) affect potential values?

Step 4: Complete the chart below as a group. Indicate whether each factor increases, decreases, or has no effect on electrode potential, and explain why.

Factor	Effect on E° (↑ / ↓ / ↔)	Explanation	Real-World Example
Nature of the electrode			
Temperature of electrolyte			
Concentration of electrolyte			
Pressure of gas species			

Step 5: Group presentation of charts

Step 6: Whole-class discussion, clarifying misconceptions

Activity4.7 Understanding Electrode Potentials and Voltaic Cells

1. Observe *Figure 4.8.*

- **a.** Can you measure the voltage of just one end of it? Why or why not?
- **b.** What do you need to measure a voltage or potential difference?
- **c.** Does voltage exist at a single point, or is it always between two points?
- **d.** How do you think this relates to measuring the potential of a single electrode in electrochemistry?

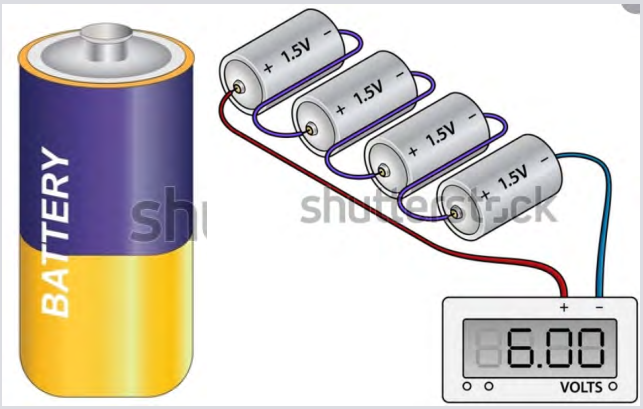


Figure 4.8: Dry cell batteries wired to a voltmeter

2. Observe *Figure 4.9*. Describe SHE in their own words.

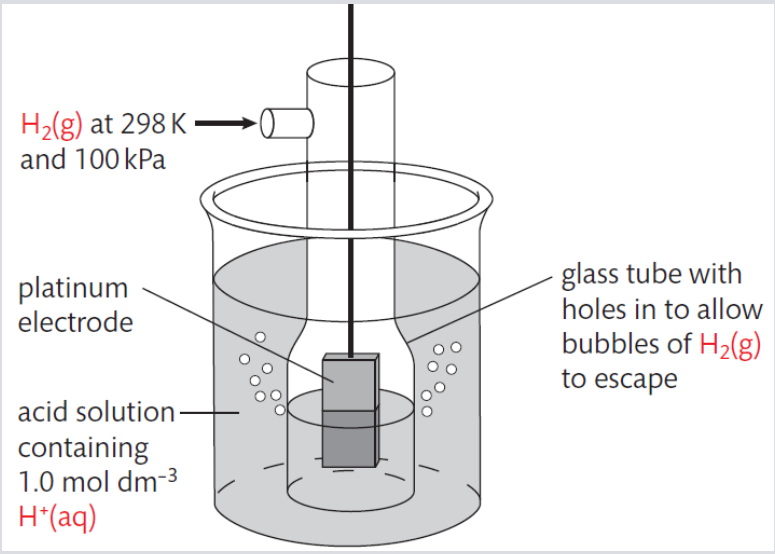


Figure 4.9: The standard hydrogen electrode (SHE)

3. Observe *Figure 4.10*.

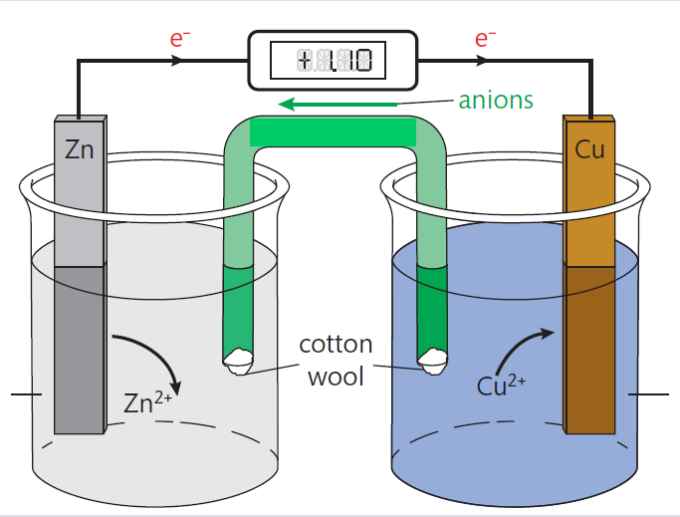


Figure 4.10: A typical voltaic cell

a. Identify the anode, cathode, salt bridge, voltmeter and electrolytes in the diagram.

b. State the standard conditions.

c. Write the redox equation for this cell

4. **a.** Explain how SHE is used as a reference.

b. Discuss examples like SHE || Cu^{2+}/Cu and interpret the meaning of positive or negative E° values.

5. Representing Voltaic Cells from Half-Cells: Consider the half cells

a. Metals in contact with ions: Mg/Mg^{2+},

b. Gas - ions: Cl_2/Cl^-

Draw, label and describe each cell, noting oxidation and reduction.

6. Writing Cell Notation

Example: Zn | Zn^{2+} || Cu^{2+} | Cu

a. Write cell notation for Voltaic Cells drawn in step 5 above.

b. Identify anode and cathode

Note

Symbols: single line (phase), double line (salt bridge), left = anode and right = cathode.

7. Write the following cell notations

a. Zn^{2+}/Zn and Ag^{+}/Ag

b. Fe^{3+}/Fe^{2+} and SHE

c. Cu^{2+}/Cu and SHE

Activity 4.8 Video Explanation of a Voltaic Cell

Materials needed: Laptop or tablet

1. You may use the link below to watch a video on how a Voltaic cell works.

https://chatgpt.com/c/6894d49e-ea04-832b-a653-62b7d2eb6607

As you watch the video, identify:

a. What materials make up the anode and cathode?

b. How do electrons move through the external circuit?

c. Which species is oxidised, and which is reduced?

d. How does the salt bridge help maintain neutrality?

e. Which electrode becomes negatively charged and why?

2. Pause the video at key moments:

a. What process occurred at each electrode?

b. In which direction are the electrons moving? How do we know?"

c. Ask: "What is the salt bridge doing here?"

3. **a.** What might happen if one electrode material changes?

b. How would the voltage change with different ion concentrations?

c. What do you predict will happen if the salt bridge is removed?

d. How would the reaction cease if one reactant runs out?

CALCULATIONS OF STANDARD ELECTRODE POTENTIALS

Calculations of standard electrode potentials

The standard electrode potential (E°) of a half-cell is the potential difference between it and the Standard Hydrogen Electrode (SHE), when both are under standard conditions:

For a voltaic cell:

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

Cathode → where reduction occurs (gain of electrons)

Anode → where oxidation occurs (loss of electrons)

Steps in Calculating E°cell

1. Write the two half-equations for the redox reaction.
2. Identify which is the oxidation and which is the reduction.
3. Assign the correct E° values from (reduction potentials only).
4. Apply the formula: $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
5. Interpret the sign:
   a. If E°cell > 0, reaction is spontaneous under standard conditions.
   b. If E°cell < 0, reaction is non-spontaneous under standard conditions.

Example 4.8

A cell is made from a magnesium electrode in $Mg^{2+}_{(aq)}$ and a copper electrode in $Cu^{2+}_{(aq),}$ both at 1 mol dm^{-3}.

Given:
$E°(Mg^{2+}/Mg) = -2.37$ V

$E°(Cu^{2+}/Cu) = +0.34$ V

Calculate E°cell and state whether the reaction is spontaneous.

Answer

Step 1 – Write the half-equations:

Reduction (Cathode): $Cu^{2+} + 2e^- \rightarrow Cu$ $E° = +0.34$ V

Oxidation (Anode): $Mg \rightarrow Mg^{2+} + 2e^-$ $E° = -2.37$ V

Step 2 – Apply formula:

$E°cell = (+0.34) - (-2.37)$

$E°cell = 0.34 + 2.37 = +2.71$ V

Step 3 – Interpretation:

$E°cell = +2.71$ V → Positive, so the reaction is spontaneous.

Overall cell reaction: $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$

Common Variations in Questions

1. Given E°cell and one half-cell value, find the other half-cell potential.
2. Predict spontaneity of a reaction.
3. Rank metals or ions by oxidising/reducing strength using E° values.
4. Use E° values to predict displacement reactions.

Predicting the Feasibility of a Reaction Using Standard Cell Potential and Gibbs Free Energy

1. **Relationship Between $E°_{cell}$ and $\Delta G°$**

   The link between standard cell potential and Gibbs free energy is given by the thermodynamic equation: $\Delta G° = -nF\,E°_{cell}$

   Where:

   $\Delta G°$ = Standard Gibbs free energy change (J or kJ mol^{-1})

   n = Number of moles of electrons transferred in the balanced redox equation

F = Faraday constant = 96,500 C mol^{-1}

E°cell = Standard cell potential (V)

2. **Prediction Rules**
   a. If $E°_{cell} > 0 \rightarrow \Delta G° < 0 \rightarrow$ Reaction is feasible/spontaneous under standard conditions.
   b. If $E°cell = 0 \rightarrow \Delta G° = 0 \rightarrow$ Reaction is at equilibrium under standard conditions.
   c. If $E°cell < 0 \rightarrow \Delta G° > 0 \rightarrow$ Reaction is non-feasible/non-spontaneous under standard conditions.

3. **Step-by-Step Procedure**
   a. Write the balanced redox reaction and determine n (electrons transferred).
   b. Look up or calculate $E°_{cell}$: $E°_{cell} = E°_{cathode} - E°_{anode}$
   c. Calculate $\Delta G°$ using: $\Delta G° = -nF\,E°_{cell}$
   d. Interpret the sign of $\Delta G°$ to predict feasibility.

Example 4.9

Given

$E°(Zn^{2+}/Zn) = -0.76$ V

$E°(Cu^{2+}/Cu) = +0.34$ V

Predict whether the reaction: $Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$ is feasible under standard conditions.

Answer

Step 1 – Identify electrodes

Cathode (reduction): $Cu^{2+} + 2e^- \rightarrow Cu$ $E° = +0.34$ V

Anode (oxidation): $Zn \rightarrow Zn^{2+} + 2e^-$ $E° = -0.76$ V

Step 2 – Calculate $E°_{cell}$

$E°_{cell} = (+0.34) - (-0.76) = 1.10$ V

Step 3 – Calculate $\Delta G°$

n = 2

F = 96,500 C mol^{-1}

$\Delta G° = -(2)\,(96{,}500)\,(1.10) = -212{,}300$ J $mol^{-1} = -212.3$ kJ mol^{-1}

Step 4 – Interpretation

ΔG° is negative → reaction is spontaneous under standard conditions.

Experiment to Determine the Emf of a Voltaic Cell

Aim: To construct a Daniell cell and determine its electromotive force (EMF) under standard conditions.

Materials needed

1. Zinc strip (clean, polished)
2. Copper strip (clean, polished)
3. 1.0 mol dm^{-3} zinc sulphate solution ($ZnSO_4$)
4. 1.0 mol dm^{-3} copper sulphate solution ($CuSO_4$)
5. Salt bridge (e.g., KNO_3 or NH_4NO_3 in agar gel)
6. Digital voltmeter (high resistance, ≥10 MΩ)
7. Connecting wires and crocodile clips
8. Two 250 cm^3 beakers

Procedure

1. Pour 1.0 mol dm^{-3} $ZnSO_4$ into one beaker and 1.0 mol dm^{-3} $CuSO_4$ into another.
2. Place the zinc strip in the $ZnSO_4$ solution and the copper strip in the $CuSO_4$ solution.
3. Connect the zinc electrode to the negative terminal of the voltmeter and the copper electrode to the positive terminal using wires and clips.
4. Insert the ends of the salt bridge into each beaker to allow ion flow and maintain electrical neutrality.
5. Observe and record the voltage displayed on the voltmeter as soon as the circuit is complete.

Expected Reaction in the Daniell Cell

Anode (oxidation): $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Cathode (reduction): $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Overall reaction: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Expected Result

Standard EMF of Daniell cell:

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$E^\circ_{cell} = (+0.34) - (-0.76) = +1.10\ V$

Measured EMF under ideal conditions ≈ 1.10 V.

Precautions

1. Ensure electrodes are clean and free from oxides before use.
2. Avoid drawing current for long periods to prevent polarisation.
3. Keep the salt bridge saturated and ensure no air bubbles block ion flow.
4. Use solutions at 1.0 mol dm^{-3} and maintain temperature at 25 °C for standard conditions.

Applications of Principles of Electrode Potentials

Application	Principle	Example
Predicting Feasibility of Redox Reactions	A redox reaction is feasible under standard conditions if $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} > 0$.	Zn displaces Cu^{2+} because $E^\circ(Cu^{2+}/Cu) = +0.34$ V > $E^\circ(Zn^{2+}/Zn) = -0.76$ V.
Predicting Direction of Electron Flow	Electrons flow from electrode with lower E° (more negative, stronger reducing agent) to one with higher E°.	In a Daniell cell, electrons flow from Zn (−0.76 V) to Cu (+0.34 V).
Electrochemical Series and Reactivity	Elements arranged in order of increasing E° help predict metal reactivity and displacement reactions.	Mg (−2.37 V) is more reactive than Zn (−0.76 V).
Calculating EMF of Cells	$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ gives the voltage of a cell.	For Zn–Cu cell: E°_{cell} = 0.34 − (−0.76) = +1.10 V.
Determining Gibbs Free Energy Change	$\Delta G^\circ = -nF\,E^\circ_{cell}$ links electrical work to thermodynamic feasibility.	For E°_{cell} = +1.10 V, n = 2, $\Delta G^\circ = -212.3$ kJ mol^{-1}.
Industrial Electrolysis	E° values predict which electrode reactions occur and the potential required.	Electroplating: E° values help choose conditions for desired metal deposition.

Application	Principle	Example
Corrosion Prevention	More reactive metals (negative E°) act as sacrificial anodes to protect less reactive metals.	Zn or Mg blocks protect steel hulls from rusting.
Electroplating and Metal Refining	E° values predict deposition order during electrolysis.	Silver plating: Ag^+ reduced at cathode, silver anode dissolves.
Designing Batteries and Fuel Cells	Half-cells with large positive $E°_{cell}$ produce high-voltage cells.	Li-ion batteries use Li/Li^+ (very negative E°) with strong oxidising cathode.
Analytical Chemistry	Redox titrations rely on predictable E° values to detect equivalence points.	Permanganate titration of Fe^{2+} uses known E° values for endpoint detection.

Activity 4.9 Understanding Standard Electrode Potentials and Connection with Gibbs Free Energy (ΔG°)

Materials

Mini E° table: Mini Electrochemical Series (Standard Reduction Potentials)

Half-reaction (reduction)	E° / V
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	+1.51
$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.36
$Ag^+ + e^- \rightarrow Ag$	+0.80
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.77
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34
$2H^+ + 2e^- \rightarrow H_2(g)$	0.00
$Pb^{2+} + 2e^- \rightarrow Pb$	−0.13
$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44
$Zn^{2+} + 2e^- \rightarrow Zn$	−0.76
$Al^{3+} + 3e^- \rightarrow Al$	−1.66

All tabulated E° values are for reduction

1. Understanding Standard Electrode Potentials

E° values refer to *reduction* under standard conditions (1 M, 25 °C, 1 atm) e.g.

$Zn^{2+} + 2e^- \rightarrow Zn$ $\quad E° = -0.76$ V

$Cu^{2+} + 2e^- \rightarrow Cu$ $\quad E° = +0.34$ V

More positive E° → stronger oxidising agent (more likely to be reduced).

More negative E° → stronger reducing agent (more likely to be oxidised).

2. Calculating Standard Cell Potentials ($E°_{cell}$)

Using the Mini Electrochemical Series above:

a. Identify the two half-reactions and their E° values

b. Label both as *reduction*

c. Determine which is to be oxidised (more negative E°) vs. reduced (more positive E°)

d. Reverse the oxidation reaction's E° sign

e. Apply: $E°cell = E°cathode - E°anode$

f. $\Delta G° = -nF\ E°cell$

where n= moles of electrons transferred.

$F = 96500\ C\,mol^{-1}$

Interpretation

a. $E°cell > 0 \Rightarrow$ **spontaneous** under standard conditions

b. $E°cell < 0 \Rightarrow$ **non-spontaneous** under standard conditions.

c. A negative $\Delta G°$ confirms spontaneity

3. Calculate $E°$ cell and $\Delta G°$ for a Zn/Cu cell.

a. Consider the simple Cell

$Cu^{2+} + 2e^- \rightarrow Cu$ $\quad E° = +0.34$ V

$Zn^{2+} + 2e^- \rightarrow Zn$ $\quad E° = -0.76$ V

i. Identify anode and cathode.

Cu^{2+}/Cu is more positive ⇒ cathode (reduction);

Zn^{2+}/Zn becomes oxidation.

Reverse oxidation: $Zn \rightarrow Zn^{2+} + 2e^-$

($E° = +0.76$ V after sign change).

ii. Calculate E°cell.

E°cell = E°cathode − E°anode

E°cell = 0.34 − (−0.76) = +1.10 V

iii. Determine ΔG° for this reaction.

$\Delta G° = -nF\,E°cell$

n = 2 (two electrons)

$\Delta G° = -2(96500)\,(1.10)$

$= -2.123\times10^5$ J ≈ −212 kJ

Negative ΔG°, positive E°cell ⇒ feasible/spontaneous.

b. Consider the simple Cell

$Fe^{3+} + e^- \rightarrow Fe^{2+}$ E° = +0.77V

$Al^{3+} + 3e^- \rightarrow Al$ E° = –1.66V

i. Identify which is oxidised and reduced.

ii. Calculate E°cell and ΔG°.

Activity 4.10 Build and Explore Voltaic Cells

Goal: Assemble simple cells and discover which pairings give higher emf.

Materials needed

1. Metals (strips/nails): Zn, Cu, Fe, Pb, graphite (pencil lead)
2. Electrolytes: 0.5 – 1.0 M solutions (e.g., $CuSO_4$, $ZnSO_4$, $FeSO_4$, NaCl)
3. Salt bridge: filter paper soaked in saturated KNO_3/$NaNO_3$ (or salt-soaked paper towel)
4. Beakers/label tape; multimeter (DC, 0–2 V range); crocodile clips; stopwatch

Mini E° (reduction) reference (25 °C, 1 M, 1 atm)

$Ag^+ + e^- \rightarrow Ag$ +0.80 V

$Fe^{3+} + e^- \rightarrow Fe^{2+}$ +0.77 V

$Cu^{2+} + 2e^- \rightarrow Cu$ +0.34 V

$2H^+ + 2e^- \rightarrow H_2$ 0.00 V

$Pb^{2+} + 2e^- \rightarrow Pb$ −0.13 V

$Fe^{2+} + 2e^- \rightarrow Fe \quad -0.44\ V$

$Zn^{2+} + 2e^- \rightarrow Zn \quad -0.76\ V$

Procedure

1. In small groups, set up two half-cells e.g.,

$Zn/ZnSO_4$ and $Cu/CuSO_4$

$Fe/FeSO_4$ and $Cu/CuSO_4$

$Pb/Pb(NO_3)_2$ and $Cu/CuSO_4$

$Cu/CuSO_4$ and $Ag/AgNO_3$, connected by a salt bridge.

Which electrode do you expect to be the anode/cathode? Why?

2. Connect to the multimeter or voltmeter (red to predicted **cathode**).

a. Record **emf** and **polarity**.

b. Draw cell diagrams

3. Using standard reduction potentials from the mini electrochemical series, calculate the theoretical E°cell.

a. Compare measured versus calculated values

b. Discuss possible sources of error: non-standard concentrations, temperature changes, contact resistance, electrode polarisation

4. Electrode Reactions Exploration using Worksheet. Research and complete the following for each battery type:

a. Lead-acid battery

Anode half-reaction (oxidation):_____

Cathode half-reaction (reduction): _____

Overall balanced equation: _____

b. Alkaline cell

Anode half-reaction (oxidation):_____

Cathode half-reaction (reduction): _____

Overall balanced equation: _____

c. Lithium-ion battery

Anode half-reaction (oxidation):_____

Cathode half-reaction (reduction): _____

Overall balanced equation:_____

d. Nickel-cadmium battery (NiCd)

Anode half-reaction (oxidation):_____

Cathode half-reaction (reduction): _____

Overall balanced equation: _____

e. Fuel cells (H_2–O_2)

Anode half-reaction (oxidation):_____

Cathode half-reaction (reduction): _____

Overall balanced equation:_____

5. Discuss why certain battery types are used for specific applications.

Activity 4.11 Storage & Fuel Cell Technologies

1. Watch the following videos

How batteries work - Adam Jacobson | TED-Ed

How a Fuel Cell Electric Vehicle Works on Vimeo

https://www.youtube.com/watch?v=Rwy6WozMbt8

https://www.youtube.com/watch?v=SRt1KOP9a2g

2. For each video, pause at key moments to discuss and

a. List the components *you can recognise*. What does each do in one phrase?

b. Write the anode and cathode.

c. State features that make fuel cells better suited than batteries for long-haul vehicles.

d. the component of fuel cell systems that improves their lifespan to match that of diesel engines.

Fill the table as you watch the video

System	Key components spotted	Evidence of anode vs cathode action	Energy transformations	Advantages over diesel-powered	Challenges/ risks
H_2 fuel-cell EV					
Hydrogen-rich fuel system (on-board reformer, etc.)					
Battery electric (reference)					

3. Concept Mapping for Hydrogen Fuel Cells: Use the diagram in *Figure 4.11* as visual reference for Hydrogen Fuel Cells. Develop four main branches from the central topic and add key subcomponents to each.

a. Fuel Cell Components

b. Electrochemical Reactions

i. *Anode (oxidation)*

ii. *Cathode (reduction)*

Overall

c. Energy Transformations

d. Advantages over Batteries & Diesel

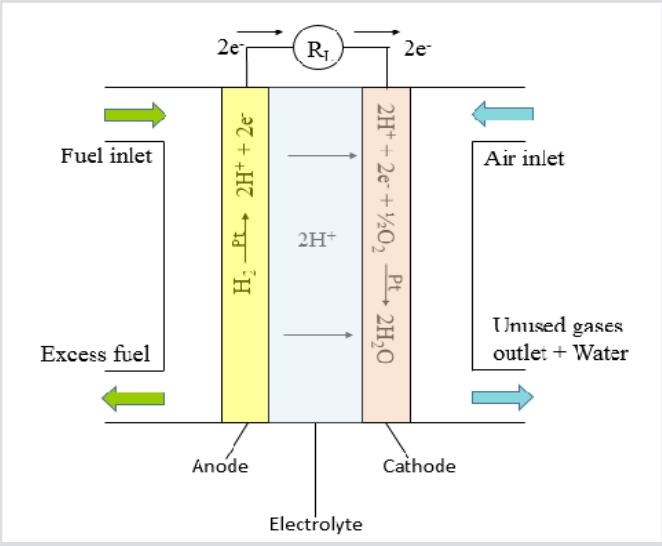


Figure 4.11: Hydrogen fuel cell

ELECTROLYSIS

Electrolysis

What is Electrolysis?

Electrolysis is a process that uses electricity to break down chemical compounds.

How Does it Work?

Electricity passes through a solution or liquid of the substance, called an electrolyte. The electricity causes chemical changes, breaking down the substance. This process stores energy in the form of new chemical bonds.

Key Points

1. Electrolysis needs electricity to work.
2. It is the opposite of a battery, which makes electricity from chemicals.
3. Electrolysis helps create new substances and store energy.

The word "electrolysis" comes from Greek words "electron" (electricity) and "lysis" (breaking down).

Electrolytic Cell

What is an Electrolytic Cell?

An electrolytic cell is a setup that uses electrical energy from an external source (like a battery or power supply) to force a chemical reaction that would not happen on its own. It is the opposite of a galvanic or voltaic cell (which produces electricity from a spontaneous reaction). In an electrolytic cell, electricity is supplied to drive a non-spontaneous redox reaction.

Basic Components

1. Power Source – a battery or DC supply to push electrons into the system.
2. Electrodes – two conductors dipped in the electrolyte:
   a. Cathode (–): where reduction happens (gain of electrons).
   b. Anode (+): where oxidation happens (loss of electrons).

Note

In electrolytic cells, the cathode is negative, and the anode is positive because of the external power supply.

3. Electrolyte – a molten ionic compound or ionic solution that conducts electricity (provides free-moving ions).

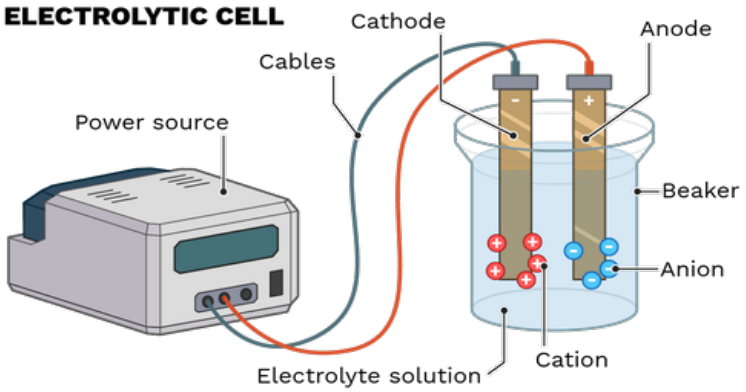


Figure 4.12: An Electrolytic Cell

How it Works

1. The external power supply pushes electrons into the cathode (negative electrode). Positive ions (cations) in the electrolyte move to the cathode to gain electrons → reduction occurs.
2. The power supply pulls electrons from the anode (positive electrode). Negative ions (anions) in the electrolyte move to the anode to give up electrons → oxidation occurs.
3. The overall reaction is a decomposition (breaking down) of the electrolyte or compound.

Factors That Affect Selective Discharge of Species During Electrolysis

1. **Position in the Electrochemical Series (discharge/reduction potential)**

   Cathode (−): Ions with a more positive E° (easier to reduce) discharge first.

   Example: In aqueous NaCl, H^+ (from water) is reduced to H_2 rather than Na^+ to Na metal.

   Anode (+): Species that are easier to oxidise discharge first.

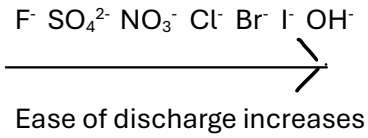
   Example: With enough Cl^-, Cl_2 forms; otherwise, OH^- from water is oxidised to O_2.

   Electrochemical Series for Cations

$$K^+\ Na^+\ Ca^{2+}\ Mg^{2+}\ Al^{3+}\ Zn^{2+}\ Fe^{2+}\ Sn^{2+}\ Pb^{2+}\ H^+\ Cu^{2+}\ Hg^{2+}\ Ag^+\ Au^+ \longrightarrow$$

Ease of discharge increases

Electrochemical Series for Anions



The electrochemical series is an arrangement of ions in order of their relative ability to accept or release electrons. Cations lower in the series readily accepts electrons at the cathode and get reduced than those higher in the series

Anions lower in the series give out electrons more easily and are therefore discharged in preference to those higher in the series

2. Concentration of ions

High concentration favours that ion (shifts the Nernst potential).

This applies mainly to anions in aqueous solutions in which the halide ion concentration is higher than OH^- concentration. In the case of SO_4^{2-} and NO_3^- ions, OH^- ions are still preferentially discharged even if their concentrations are higher. They are extremely difficult to discharge in aqueous solutions because they have high oxidation potential and will not easily give up electrons.

Example: Concentrated brine (NaCl): Cl^- is discharged to form Cl_2 at the anode instead of OH^- ions from H_2O.

In Dilute brine: $O_{2(g)}$ is formed at the anode instead (from OH^-).

3. Nature of the electrodes (inert vs active/corrodible)

Inert anodes (Pt/graphite): They are electrodes that conduct electrons but do not undergo oxidation or reduction. The electrolyte supplies the anion for oxidation (e.g., Cl_2 or O_2)

Active anodes: These are electrodes that actively participate in the electrochemical cell reaction (oxidation or reduction). They can dissolve and supply metal ions.

Example: Cu anode in $CuSO_4$: the anode ($Cu_{(s)}$) dissolves to Cu^{2+}; at the cathode, Cu^{2+} plates out (forms $Cu_{(s)}$) This is the basis of copper refining.

No $O_{2(g)}$ forms.

Electrolysis of Molten Compounds in Inert Electrodes

When an **ionic compound is melted** (molten), its ions are **free to move** and carry a charge.

1. **Molten lead (II) bromide, $PbBr_{2(l)}$**
   a. Cathode: $Pb^{2+} + 2e^- \rightarrow Pb_{(l)}$
   b. Anode: $2Br^- \rightarrow Br_{2(g)} + 2e^-$
   c. Overall: $PbBr_{2(l)} \rightarrow Pb_{(l)} + Br_{2(g)}$
2. **Molten sodium chloride, $NaCl_{(l)}$**
   a. Cathode: $Na_{(aq)}^{+} + e^- \rightarrow Na_{(l)}$
   b. Anode: $2Cl_{(aq)}^{-} \rightarrow Cl_{2(g)} + 2e^-$
   c. Overall: $2NaCl_{(l)} \rightarrow 2Na_{(l)} + Cl_{2(g)}$
3. **Molten alumina (Al_2O_3)**
   a. Cathode: $Al_{(aq)}^{3+} + 3e^- \rightarrow Al_{(l)}$
   b. Anode: $2O_{(aq)}^{2-} \rightarrow O_{2(g)} + 4e^-$
   c. Overall: $2Al_2O_{3(s)} \rightarrow 4Al_{(l)} + 3O_{2(g)}$

Electrolysis of Aqueous Solutions using Inert Electrodes

Assume inert electrodes (Pt/graphite) and room temperature. For each electrolyte: (i) half-reactions, (ii) observations, (iii) products, (iv) change to electrolyte.

Dilute NaCl (aq)

Item	Details
Half-reactions	Cathode (–): $2H_2O + 2e– \rightarrow H_{2(g)} + 2OH^-$ Anode (+): $4OH^- \rightarrow O_{2(g)} + 2H_2O + 4e^-$ ($water/OH^-$ oxidised)
Observations	Cathode: Colourless bubbles (H_2). Anode: Colourless bubbles (O_2), relights glowing splint.
Products	Cathode: Hydrogen gas (H_2). Anode: Oxygen gas (O_2).

Item	Details
Electrolyte change	Solution becomes alkaline (OH^- builds up) → $NaOH_{(aq)}$ forms; $[Cl^-]$ largely unchanged.

Concentrated NaCl (aq) (brine)

Item	Details
Half-reactions	Cathode (–): $2H_2O + 2e^- \rightarrow H_{2(g)} + 2OH^-$ Anode (+): $2Cl^- \rightarrow Cl_{2(g)} + 2e^-$ (high $[Cl^-]$ favours formation of Cl_2)
Observations	Cathode: Effervescence (H_2). Anode: Greenish-yellow gas with pungent smell (Cl_2).
Products	Cathode: Hydrogen gas (H_2). Anode: Chlorine gas (Cl_2).
Electrolyte change	Solution becomes alkaline (OH^-) → NaOH(aq) forms; $[Cl^-]$ decreases as Cl_2 leaves; some dissolved Cl_2 may form $HOCl/ClO^-$.

$CuSO_4$ (aq)

Item	Details
Half-reactions	Cathode (–): $Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$ Anode (+): $2H_2O \rightarrow O_{2(g)} + 4H+ + 4e^-$ (SO_4^{2-} not discharged)
Observations	Cathode: Reddish-brown copper deposit. Anode: Bubbles of O_2; blue colour of solution fades over time.
Products	Cathode: Copper metal (Cu). Anode: Oxygen gas (O_2).
Electrolyte change	$[Cu^{2+}]$ falls (blue fades); $[H^+]$ rises → solution becomes more acidic (effectively forming H_2SO_4 with SO_4^{2-}).

Dilute H_2SO_4 (aq)

Item	Details
Half-reactions	Cathode (–): $2H^+ + 2e^- \rightarrow H_{2(g)}$ Anode (+): $2H_2O \rightarrow O_{2(g)} + 4H^+ + 4e^-$
Observations	Cathode: Bubbles of H_2. Anode: Bubbles of O_2. Gas volume ratio ~2:1 (H_2:O_2).
Products	Cathode: Hydrogen gas (H_2). Anode: Oxygen gas (O_2).
Electrolyte change	Overall acid concentration ~unchanged (H^+ consumed at cathode ≈ H^+ produced at anode).

$AgNO_3$ (aq)

Item	Details
Half-reactions	Cathode (–): $Ag^+ + e^- \rightarrow Ag_{(s)}$ Anode (+): $2H_2O \rightarrow O_{2(g)} + 4H^+ + 4e^-$ (NO_3^- not discharged)
Observations	Cathode: Grey/silver crystals deposit. Anode: Bubbles of O_2.
Products	Cathode: Silver metal (Ag). Anode: Oxygen gas (O_2).
Electrolyte change	[Ag^+] decreases; [H^+] increases → solution becomes acidic (effectively some HNO_3 formed).

Concentrated KI (aq)

Item	Details
Half-reactions	Cathode (–): $2H_2O + 2e^- \rightarrow H_2(g) + 2OH^-$ (K^+ not reduced) Anode (+): $2I^- \rightarrow I_2 + 2e^-$

Item	Details
Observations	Cathode: Bubbles of H_2. Anode: Brown colour near anode (I_2); with starch, blue-black colour.
Products	Cathode: Hydrogen gas (H_2). Anode: Iodine (I_2).
Electrolyte change	Solution becomes alkaline (OH^-) near cathode; [I^-] falls; I_2 may form I_3^-/hypoiodite in alkaline medium.

Electrolysis with Active Electrodes

1. **Copper (II) sulphate solution with copper electrodes (active anode)**

Half-reactions

Cathode (−): $Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$ *(copper plates out)*

Anode (+): $Cu_{(s)} \rightarrow Cu^{2+} + 2e^-$ *(copper anode dissolves)*

Observations

Cathode: Reddish-brown Cu coating forms, gets thicker.

Anode: Copper electrode thins; flakes.

Solution: Blue colour ~constant (if anode and cathode areas/currents are balanced).

Products at each electrode

Cathode: Copper metal.

Anode: Copper (II) ions (from dissolving anode).

Change to the electrolyte

[Cu^{2+}] is maintained ~constant (Cu^{2+} used at cathode is replenished at anode); SO_4^{2-} unchanged; pH ~unchanged.

2. **Concentrated sodium chloride solution with a mercury cathode**

Half-reactions

Cathode (Hg, −): $Na^+ + e^- + Hg \rightarrow Na\,(Hg)$ *(sodium amalgam)*

Anode (+): $2Cl^- \rightarrow Cl_2(g) + 2e^-$

Observations

Anode: Greenish-yellow Cl_2 gas (pungent) bubbles off.

Cathode: Silvery amalgam film forms on Hg surface.

Smell: Sharp, irritating odour of Chlorine near the anode.

Products at each electrode

Cathode: Na (Hg) (sodium amalgam, intermediate).

Anode: Chlorine gas (Cl_2).

Change to the electrolyte

[Cl^-] Cl^- decreases (Cl_2 leaves).

[Na^+] Na^+ effectively becomes NaOH(aq) (solution turns more alkaline).

Safety note

Mercury is toxic—industrial process, not for school labs.

3. **Electrolysis of water using nickel-based electrodes**

(Ni works in alkaline media, e.g., 20–30% KOH/NaOH; Ni surfaces cycle between Ni $(OH)_2$ at the anode.)

Half-reactions (alkaline)

Cathode (−): $2H_2O + 2e^- \rightarrow H_2(g) + 2OH^-$

Anode (+): $4OH^- \rightarrow O_2(g) + 2H_2O + 4e^-$

Observations

Cathode: H_2 gas bubbles.

Anode: O_2 gas bubbles (about half the cathode volume for equal time).

Electrodes: Anode may darken (Ni $(OH)_2$ film). Slight warming of electrolyte.

Products at each electrode

Cathode: Hydrogen gas (H_2).

Anode: Oxygen gas (O_2).

Gas volume ratio: H_2:$O_2 \approx 2:1$

Change to the electrolyte

Net water is consumed → alkaline solution becomes slightly more concentrated over time; pH stays high (OH^- made at cathode ≈ used at anode), composition (KOH/NaOH) essentially unchanged aside from dilution effects.

4. **Electrolysis of copper (II) sulphate solution ($CuSO_4$(aq)) using graphite (inert) electrodes**

Half-reactions

Cathode (–): $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Anode (+): $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

Reason: SO_4^{2-} is very hard to oxidise; water is oxidised to oxygen.

Overall (net) ionic equation

Multiply the cathode reaction by 2 and add:

$2Cu^{2+}(aq) + 2H_2O(l) \rightarrow 2Cu(s) + O_2(g) + 4H^+(aq)$

In the presence of SO_4^{2-}, the rise in H^+ effectively means H_2SO_4 is formed in solution.

Observations

Cathode (–): Reddish-brown copper metal plates onto the electrode; cathode mass increases.

Anode (+): Colourless gas (O_2) bubbles; a glowing splint relights. Graphite remains largely intact (may slowly erode at high current).

Solution: The blue colour fades with time (as $[Cu^{2+}]$ falls).

Products at each electrode

Cathode: Copper metal (Cu) (solid coating).

Anode: Oxygen gas (O_2).

Change to the electrolyte

$[Cu^{2+}]$ decreases → blue colour fades.

[H+] increases → solution becomes more acidic (effectively $H_2SO_4(aq)$ forms with the sulphate present).

SO_4^{2-} essentially unchanged (spectator).

Activity 4.12 Components of an Electrolytic Cell

Materials needed

Blank cell template or access to one digital simulation or short video, and a printed reference diagram

1. Use the link below to watch the video on electrolysis or use the diagram in ***Figure 4.13***.

   https://chatgpt.com/c/68ad0880-8cdc-8325-a4f3-a96f75ea370e

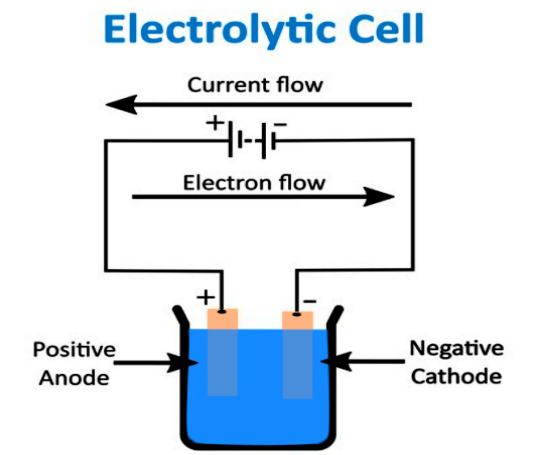


Figure 4.13: Electrolytic cell parts

2. Draw a labelled electrolytic cell diagram. The details of the diagram should include:

   Anode, cathode, electrolyte, power supply/source, direction of electron flow and direction of current flow

3. Explain the function(s) of the various parts of the diagram.
   a. the roles of electrodes,
   b. describes electron/ion flow, another discusses the power source, etc.

Activity 4.13 Modelling an Electrolytic Cell

Materials

1. Desk signs: Anode Bank (+), Cathode Bank (−), Power Supply (Manager)
2. Electron tokens (paper chips)
3. Role cards: Cations (+), Anions (−), Tellers, Manager
4. Tape arrows on floor: Wire path (external circuit) + Electrolyte zone (between banks)
5. A4 recording sheets (diagram + notes)

Roles

1. Bank Manager (Power Supply): pushes tokens (electrons) from the Cathode → Anode through the wire.
2. Tellers: receive or give out tokens at each bank.

3. Cations (+): move to Cathode Bank.
4. Anions (−): move to Anode Bank.

Procedure

Step 1

Create a visual setup: Draw or position "Anode Bank" (left, positive) and "Cathode Bank" (right, negative), include the Power Supply "Manager" in the middle, and arrowed paths for electron flow (Anode → Power Supply → Cathode).

Step 2: Introduce Aqueous Ions & Discharge Rules

1. Add water into the mix: Introduce H^+ and OH^- ions from water.
2. Discharge rules for ions:
   a. Cathode (–): If metal ions are less reactive than H^+, they get reduced; otherwise, $H^+ \rightarrow H_2$ forms.
   b. Anode (+): Halides (Cl^-, Br^-, I^-) are discharged first. If none, $OH^- \rightarrow O_2$ is released.
3. Use priority slips: Give each ion a "priority card" (e.g., high to low) and role-play their turn to get discharged.

Step 3: Trace movements: As a class, recap the paths each ion and electron took. Discuss the Manager's role (Power Supply):

Step 4: Diagram & Record

In small groups, sketch a simple diagram with:

1. Anode / Cathode "banks"
2. Electron flow arrows
3. Ion paths (H^+, OH^-, etc.) and discharge outcomes

Activity 4.14 Selective Discharge and Products

Step 1: Write down factors that affect ions that are discharged during electrolysis.

Step 2: Discus the following factors

a. Position in reactivity series

b. Ion concentration

c. Electrode type

d. Applied voltage

Step 3: Model by example comparisons (demonstration)

a. Molten salt (e.g., NaCl): Na^+ discharged at cathode, Cl^- at anode.

b. Dilute aqueous NaCl: H_2 forms at cathode if H^+ discharge is more favourable; Cl_2 at anode unless OH^- (from water) is favoured.

Step 4: Predict the outcomes of Electrolysis with inert versus active electrodes

a. Molten NaCl

b. Inert electrodes in dilute NaCl solution

c. Inert electrodes in concentrated NaCl solution

d. Active copper electrode in $CuSO_4$ solution

Write half-equations and state the key factor accounting for the discharge.

Step 5: Share your findings with your class.

Activity 4.15 Electrolysis Circus Inquiry

Materials needed

1. Electrodes: Inert (graphite rods, platinum wires, or carbon rods)

2. Power source: 6–12 V DC supply or battery pack

3. Electrolysis cell/container: Beaker (100 – 250 mL)

4. Connecting wires with crocodile clips

5. Electrolytes (examples)

- **a.** Copper (II) sulphate solution ($CuSO_4$, 0.1 – 0.5 M)
- **b.** Sodium chloride solution (NaCl, dilute and concentrated)
- **c.** Potassium iodide solution (KI, 0.1 – 0.5 M)
- **d.** Dilute sulphuric acid (H_2SO_4, 0.5 M)

6. Measuring cylinders and droppers
7. Safety equipment: Goggles, gloves, lab coats
8. Indicators/test reagents
   a. Blue and red litmus paper
   b. Glowing splint (for O_2 test)
   c. Burning splint (for H_2 test)

Safety Precautions

1. Always wear goggles and gloves.
2. Do not touch electrodes or solutions while current flows.
3. Handle acids and copper salts with care; wash spills immediately.
4. Work with low voltage DC only (not mains electricity).

Steps

1. In small groups investigate and state the ions present, possible products, half- equations of the following:
   a. Electrolysis of $CuSO_4$ solution using Cu electrodes
   b. Electrolysis of NaCl solution (dilute vs. concentrated) using C electrodes.
   c. Electrolysis of KI solution using C electrodes.
   d. Electrolysis of dilute H_2SO_4 using C electrodes.
2. Circus Experiment Stations

   Group Work Instructions
   a. Roles within each group (e.g., setup technician, recorder, product tester, safety officer).
   b. Rotate roles between trials so everyone participates.
   c. Record all observations in a results table (electrolyte, cathode product, anode product, evidence/test used).
   d. Compare group results and discuss any differences.
3. Experimental Setup
   a. Place 100 mL of electrolyte in the beaker.
   b. Insert two inert electrodes into the solution, keeping them apart but not touching.

   c. Connect the electrodes to the DC power supply using wires and crocodile clips. Mark the anode (+) and cathode (−) clearly.
   d. Prepare test papers/reagents for identifying gases and products.
   e. Switch on the power supply (6–9 V recommended).
4. Observe bubbling at the electrodes (gas evolution).
5. Collect gases (if possible) using an inverted test tube or test near the electrode.
6. Test products
   a. At cathode (−): Test gas with a burning splint (H_2 "pop" test).
   b. At anode (+): Test with glowing splint (O_2 relights), starch (iodine turns blue-black), or chlorine test (bleaches litmus).
7. Record observations (colour changes, gas evolution, electrode deposits).
8. Repeat the experiment for different electrolytes:
   a. $CuSO_4$ → copper deposit at cathode, O_2 at anode.
   b. NaCl (dilute) → H_2 at cathode, O_2 at anode.
   c. NaCl (concentrated) → H_2 at cathode, Cl_2 at anode.
   d. KI → H_2 at cathode, I_2 at anode (starch test).
   e. Dilute H_2SO_4 → H_2 at cathode, O_2 at anode.

Discussions

1. What factors determine which ion is discharged at the electrodes?
2. How do results differ between dilute and concentrated NaCl solutions?
3. Why is copper deposited in the electrolysis of $CuSO_4$?
4. How does this link to industrial applications (e.g., electroplating, chlor-alkali process)?

APPLICATIONS OF ELECTROLYSIS

Electroplating means covering one metal with a thin layer of another metal using electricity. The object to be coated is made the cathode (–), and the coating metal is the anode (+). Both are placed in a solution that has ions of the coating metal. When electricity passes, metal ions move from the solution and stick to the object (cathode). The coating metal (anode) dissolves to replace the ions in solution.

Practical Applications of Electrolysis

1. Extraction of Metals: Used to extract reactive metals (e.g., aluminium, sodium, magnesium) from their ores.
2. Electroplating: *Example:* Silver plating spoons; chrome plating car parts.
3. Manufacture of Chemicals
   a. Production of important substances: Chlorine, hydrogen, sodium hydroxide (by electrolysis of brine).
   b. Oxygen and hydrogen (by electrolysis of water).
4. Purification of Metals
5. Medical and Everyday Uses
   a. Electrolysis used for hair removal (cosmetic).
   b. Rechargeable batteries involve reversible electrolysis processes.

Experiment

Copper Electroplating

Materials

1. Copper sulphate solution
2. Copper strip (anode, +)
3. Metal object to coat (key, coin, etc.)
4. DC power supply (3 – 6 V battery or adapter)
5. Wires with clips
6. Beaker or plastic cup
7. Sandpaper or steel wool
8. Soap water (for grease)
9. Mass balance
10. Vinegar (for cleaning)

Procedure

Step 1: Clean the object

1. Rub with sandpaper to make it shiny.
2. Wash in soap water, rinse.
3. Dip quickly in vinegar, rinse again. *(Clean objects plate better!)*

Step 2: Set up the cell

1. Pour copper sulphate solution into the beaker.
2. Determine the mass of the object to be coated using a mass balance.
3. Connect the object to the negative (–) terminal → cathode.
4. Determine the mass of the copper strip using a mass balance.
5. Connect the copper strip to the positive (+) terminal → anode.
6. Put both into the solution, but don't let them touch.

Step 3: Electroplate

1. Switch on power (3 – 4 V).
2. Let it run for 15 – 30 minutes.
3. Watch: copper slowly coats the object.

Step 4: Finish

1. Turn off power.
2. Take out the object, rinse in water, and dry.
3. Determine the mass of the object.
4. Look at the new copper coating!
5. Determine the mass of the copper strip using a mass balance.
6. Compare the mass gained by the coated object to the mass lost by the copper strip.

What you should see

A thin, reddish-brown copper layer form on the object. The copper strip gets thinner as copper moves from it. If plating is uneven, it means the object was not cleaned well or current was too high/low.

Extraction of Aluminium

Aluminium is the most common metal in the Earth's crust. It is never found unreacted because it reacts easily with oxygen. We cannot use carbon to extract it (like iron) because aluminium holds oxygen more strongly (aluminium is higher in the electrochemical series than hydrogen). So, we use electrolysis.

Step 1: From Bauxite to Alumina (Bayer Process)

1. Crush bauxite ore.
2. Mix with hot sodium hydroxide → aluminium dissolves.
3. Filter out red mud (iron oxides, silica).

4. Cool the solution → aluminium hydroxide forms.
5. Heat aluminium hydroxide strongly → pure alumina (Al_2O_3).

Step 2: Setting up the Electrolytic Cell (Hall – Héroult Process)

1. A steel tank lined with carbon acts as the cathode (–).
2. Carbon anodes (+) hang from above.
3. Alumina is dissolved in molten cryolite (Na_3AlF_6) to lower melting point (to ~960 °C).
4. Direct current (DC) is applied:
   a. Carbon lining = negative terminal.
   b. Carbon anodes = positive terminal.

Step 3: What Happens During Electrolysis

1. Strong current passes through the cell.
2. At cathode (–): $Al^{3+} + 3e^- \rightarrow Al$ (molten aluminium collects at the bottom).
3. At anode (+): $O^{2-} + C \rightarrow CO_2$ (carbon is used up).
4. Aluminium is tapped out regularly.
5. Anodes must be replaced often.
6. More alumina is added from time to time.

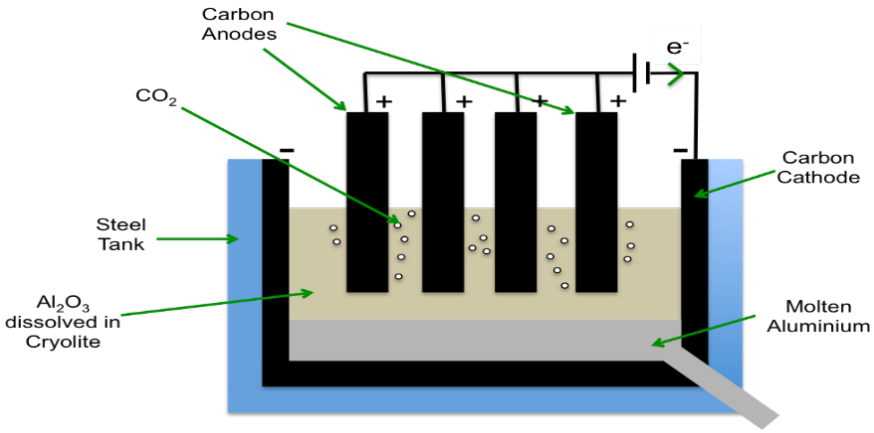


Figure 4.14: Electrolysis Process of Aluminium from Alumina

Extraction (Refining) of Gold by Electrolysis

Gold is often found mixed with other metals (like silver and copper). For jewellery and electronics, we need very pure gold. This is done using the Wohlwill process.

Step 1: Preparation

1. Impure gold (95–98% pure) is melted and shaped into anodes (+).

2. Thin pure gold sheets are prepared as cathodes (–).
3. The electrolyte is a solution of gold chloride in hydrochloric acid.
4. The tank is kept warm (about 60–70 °C).

Step 2: Setup

1. Impure gold anodes are hung in the solution.
2. Pure gold cathodes are placed in the same solution.
3. Anodes are connected to the positive terminal, cathodes to the negative terminal of a DC supply.

Step 3: Electrolysis Process

1. At anode (+): $Au \rightarrow Au^{3+} + 3e^-$ (impure gold dissolves).
2. At cathode (–): $Au^{3+} + 3e^- \rightarrow Au$ (pure gold is deposited).

Result

1. Pure gold coats the cathode.
2. Impurities fall off as anode slime or stay in solution.
3. Valuable metals (like silver and platinum) can later be recovered from the slime.

Faraday's Laws of Electrolysis

Introduction to Faraday's Laws of Electrolysis

What Are Faraday's Laws?

These laws, formulated by Michael Faraday in 1833, establish the quantitative relationship between electric charge and chemical changes at electrodes during electrolysis. They are foundational in electrochemistry.

Faraday's First Law

Statement: The mass of a substance produced (deposited or liberated) at an electrode during electrolysis is directly proportional to the total electric charge passed through the electrolyte.

Mathematical Expression: $m \propto Q$

$$\Rightarrow m = Z \times Q$$

m: mass of substance (g)

Q: electric charge passed (Coulombs, $Q = I \times t$)

Z: electrochemical equivalent (g/C)

Interpretation

More current or longer duration → more substance deposited, assuming the efficiency of the process remains the same.

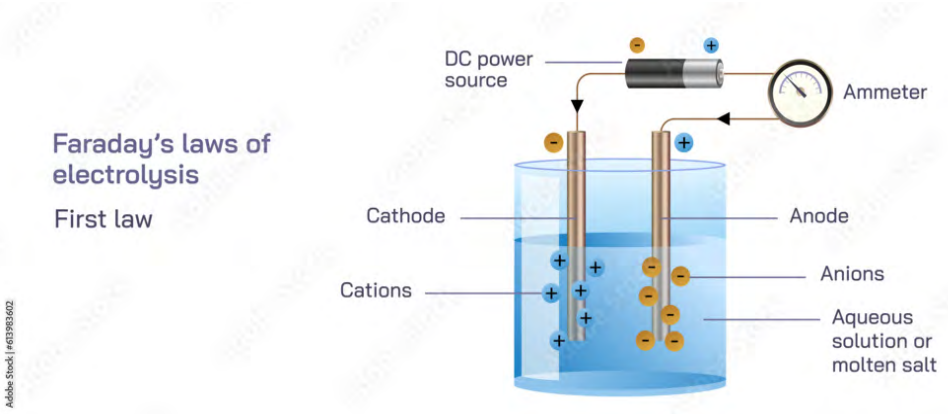


Figure 4.15: Visual overview of Faraday's First Law of Electrolysis

Faraday's Second Law

Statement: When the same amount of charge passes through different electrolytes, the masses of substances liberated at the electrodes are directly proportional to their chemical equivalent weights (equivalent weights = molar mass ÷ valence).

Mathematical Expression, $m_1 : m_2 = E_1 : E_2$

m_1, m_2: Masses of produced substances

E_1, E_2: Equivalent weights of the substances

Implication

Given equal charge, substances with higher equivalent weights yield more mass in proportion.

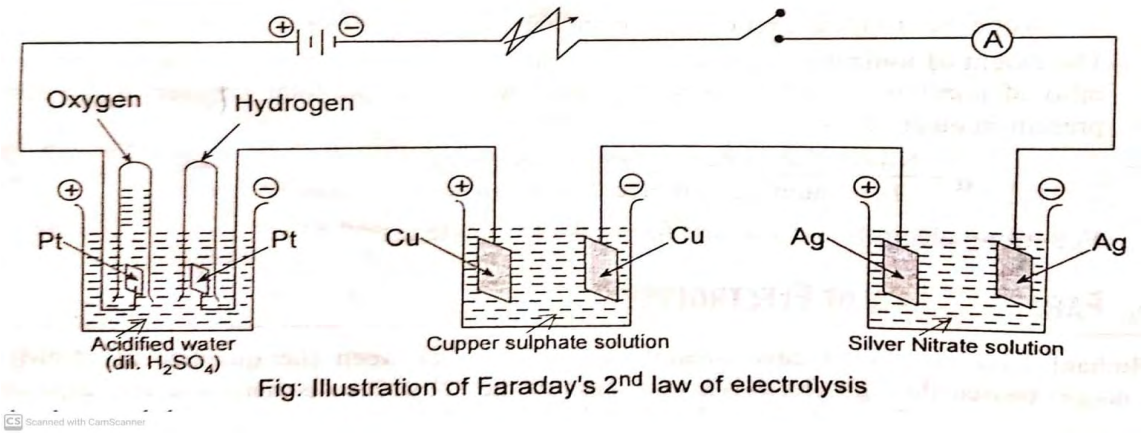


Fig: Illustration of Faraday's 2nd law of electrolysis

Figure 4.16*:* Visual overview of Faraday's second Law of Electrolysis

Combined Mathematical Form

Both laws can be combined into a single equation:

$$m = \frac{(M \times I \times t)}{(n \times F)}$$

Where:

m = mass of substance produced (g)

M = molar mass of the substance (g/mol)

I = current (amperes)

t = time (seconds)

n = number of electrons transferred per ion

F = Faraday constant (96,500 C/mol)

Given

I = 1.0 A, t = 1.0 h = 3600 s; M(Cu) = 63.55 g mol^{-1}; F = 96,500 C mol^{-1}

Reaction at the cathode: $Cu^{2+} + 2e^- \rightarrow Cu(s)$

from the balanced equation,

2 mol of e^- deposited 1mol of Cu

$$n = \frac{It}{zF} = \frac{1.0 \times (1 \times 60 \times 60)}{2 \times 96500} = 0.01865\ mol$$

$$m(Cu) = n \times M$$

$$= 0.01865 \text{ mol} \times 63.55 \text{ gmol}^{-1}$$

$$= 1.19 \text{ g}$$

Activity 4.16 Metal Extraction



Images of Bauxite ore → aluminium (left) and Alluvial gold ore → refined gold (right)

1. State the differences you can see between the raw ore and the final shiny metal in the image above.

2. **a.** How aluminium is obtained

i. Bauxite → purified to alumina
ii. Alumina is dissolved in molten cryolite
iii. **At the cathode:** aluminium metal forms
iv. **At the anode:** oxygen gas is released

b. Why is electricity needed to separate aluminium from its ore?

3. **a.** How gold is refined

i. Impure gold is made into an anode (+).
ii. Pure gold sheet is the cathode (–).
iii. Electrolyte = chloroauric acid.
iv. Pure gold slowly collects on the cathode

b. How does Faraday's First Law explain why more electricity produces more gold?

4. **a.** List both negative and positive impacts of mining in Ghana.

i. Problems (Negatives)

ii. Benefits (Positives)

b. How can we balance the benefits of mining with its problems?

5. Discuss how metal extraction supports Ghana

a. How does mining help Ghana grow?

b. What responsibilities come with this growth?

Activity 4.17 Exploring Faraday's Laws of Electrolysis

1. **a.** Calculate the mass of copper deposited when a current of 2.0 A passes through copper (II) sulphate solution for 30 minutes."

Discuss the steps and answer as a class.

b. How long must a current of 3 A be passed through molten aluminium oxide (Al_2O_3) to produce 2.7 g of aluminium? *(M(Al) = 27 g/mol, n = 3)*

2. **a.** Think about what Faraday's Laws say

i. The more electricity you use, the more substance is deposited.

ii. Different metals deposit in amounts based on their charge and mass.

b. Where are these laws used in Ghana. Use these prompts

i. Electroplating

ii. Purification

iii. Extraction

iv. Batteries

3. Write down two benefits of using electrolysis in Ghana and two challenges

4. Share your findings with colleagues.

CORROSION OF METALS

Corrosion in Ghana

Ghana's warm climate, long coastline, and growing industries make metals wear out quickly. In coastal towns, the salty sea breeze causes corrosion. In the north, metals face wet and dry seasons that speed up damage. Everyday items

like iron roofing sheets rust fast if they are not protected. Even historic places such as Cape Coast and Elmina Castles need constant repair because of the sea air. Understanding corrosion science is important for Ghana's future. It helps us protect metals by learning how corrosion works and how to prevent it.

What is Corrosion?

Corrosion is when a metal changes into a compound (like an oxide, hydroxide, or salt) because of chemical reactions with its surroundings.

It is an electrochemical process:

1. The metal loses electrons (oxidation).
2. Other substances in the environment gain electrons (reduction).

The Reactions

General Process

1. Anodic (Oxidation): $M \rightarrow M^{n+} + ne^-$ (The metal atom becomes a metal ion).
2. Cathodic (Reduction):
   a. In acidic solutions: $2H^+ + 2e^- \rightarrow H_2$
   b. In neutral/alkaline solutions: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

Example: Zinc in Acid

Anode: $Zn \rightarrow Zn^{2+} + 2e^-$

Cathode: $2H^+ + 2e^- \rightarrow H_2$

Overall: $Zn + 2H^+ \rightarrow Zn^{2+} + H_2$

In short: Corrosion is when metals 'rust away' because of electrochemical reactions with the environment. Learning how it works helps Ghana protect its buildings, industries, and heritage.

The Electrochemical Nature of Corrosion

Why Electrochemical?

Corrosion happens like a tiny battery on the surface of the metal

1. Some parts of the metal act like the anode (they lose electrons).
2. Other parts act like the cathode (they gain electrons).
3. Water and oxygen around the metal help carry the electrons and ions.

The Process Step by Step

1. **At the Anode (Oxidation)**
   The metal loses electrons and turns into ions.
   Example: $Fe \rightarrow Fe^{2+} + 2e^-$
2. **At the Cathode (Reduction)**
   Electrons released at the anode travel to another part of the metal.
   In acidic water: $2H^+ + 2e^- \rightarrow H_2$ (hydrogen gas forms)
   In neutral/alkaline water: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$
3. **Overall Effect**
   The metal surface changes into compounds like iron oxides (rust), zinc salts, or hydroxides. The bond between the metal and the oxide layer becomes weaker and flaky, causing it to come off with much less force than if it were pure metal.

Everyday Example

An iron nail left in salty water rusts quickly. One part of the nail becomes the anode (metal dissolves). Another part becomes the cathode (oxygen reacts with electrons). Together, the nail behaves like a little battery until it is eaten away.

Why It Matters in Ghana

1. Roofing sheets rust in the salty sea breeze.
2. Old bridges, cars, and even historic castles need constant maintenance.

Understanding corrosion helps us protect and save money by painting, coating, or galvanising metals.

Factors Affecting the Rate of Corrosion

The rate of corrosion depends on both the environment and the properties of the metal. The **Table 4.3** below summarises key factors, their effects, and everyday examples learners can relate to.

Table 4.3: Factors that affect the rate of corrosion

Factor	Effect on Corrosion Rate	Example
Moisture (Water Availability)	More moisture speeds up corrosion.	Iron rusts faster in rainy seasons.

Factor	Effect on Corrosion Rate	Example
Oxygen Supply	Higher oxygen speeds up cathodic reactions.	Steel near the sea corrodes quickly.
Electrolytes (Salts/Acids)	Salts and acids increase conductivity, making corrosion faster.	Metals corrode faster in seawater.
Temperature	Higher temperature increases reaction speed.	Faster rusting in hot, humid coastal Ghana.
Metal Purity & Composition	Impurities form anode–cathode sites; alloys resist corrosion.	Stainless steel resists rust better than iron.
pH of the Environment	Acidic conditions increase corrosion.	Acid rain corrodes buildings and cars.
Contact with Dissimilar Metals	The more reactive metal corrodes faster (galvanic corrosion).	Zinc corrodes when in contact with copper.
Surface Condition	Scratches or rough surfaces corrode faster.	Painted iron resists rust longer than bare iron.

Rusting

Rusting is a specific form of corrosion that affects iron and its alloys. It produces iron (III) oxide ($Fe_2O_3 . nH_2O$), commonly known as rust.

Rusting is defined as an electrochemical process in which iron reacts with oxygen and water to form hydrated iron (III) oxide ($Fe_2O_3.nH_2O$). Refer to ***Figure 4.17*** for the rusting of iron.

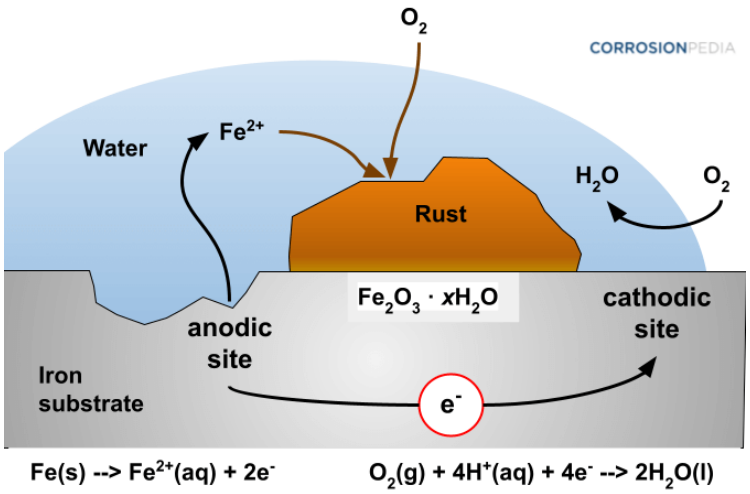


Figure 4.17: Rusting of iron

Experiment: Conditions Needed for Rusting

Aim: Find out which conditions (water, oxygen/air, salt) are needed for iron to rust.

Materials needed

1. 5 clean iron nails (same size)
2. 5 test tubes + rack (or 5 small clear jars with lids)
3. Tap water
4. Boiled (hot) water (recently boiled to remove dissolved air)
5. Cooking oil (a few mL)
6. Calcium chloride or silica gel (drying agent) in a little gauze or paper sachet
7. Salt (NaCl) to make salt water (about 1 teaspoon in 50 mL)
8. Labels/marker, measuring cylinder/spoon, tongs
9. A few drops of vinegar/lemon juice (acid rain test)
10. Paper towels, gloves, eye protection

Safety

1. Wear goggles and gloves.
2. Hot water can burn—use tongs and let it cool a bit before handling.
3. Do not touch silica gel/calcium chloride with bare hands.
4. Dispose of liquids down the sink with plenty of water; dry agents go to trash (or as per school rules).

Set-ups (label the tubes)

Tube A – Water + Air (control for rusting)

1. Add enough tap water to cover the nail halfway.
2. Insert clean nail so part is in water and part in air.
3. Leave open or loosely capped.

Tube B – No Oxygen (water present)

1. Fill with freshly boiled, hot water (nearly to the top).
2. Add a clean nail, so that it is fully submerged
3. Add a thin layer of oil to float on top (keeps air out).

Tube C – No Water (air present)

1. Put a small sachet of calcium chloride/silica gel in the tube (keeps air dry).
2. Add a dry nail (do not add water). Cap tightly.

Tube D – Salt Water (water + air + salt)

1. Make salt solution (1 tsp salt in 50 mL water).
2. Add enough to cover a clean nail halfway. Leave open/loosely capped.

Tube E – Water + Air and acid (acidic environment)

1. Add enough tap water to cover the nail halfway.
2. Add a few drops of vinegar or lemon juice to acidify the conditions.
3. Insert clean nail so part is in water and part in air.
4. Leave open or loosely capped.

Procedure

1. Clean each nail with a paper towel (remove grease/rust).
2. Set up Tubes A–E as above and label them.
3. Leave the rack undisturbed for 2–5 days (check daily).
4. Observe and record: colour changes, brown rust, cloudiness, bubbles.
5. Take photos or draw what you see each day.

Results table

Tube	Conditions	Day 1	Day 2	Day 3	Day 4/5	Rust? (Yes/ No/Rate)
A	Water + Air					
B	Water, no oxygen (oil layer)					
C	Dry air, no water					
D	Salt water + air					
E	Water and air and acid					

Conclusion prompts

1. Which tubes rusted? Which did not?
2. What does this show about the conditions needed for rusting?
3. Which factor made rusting faster? (salt, acid)

4. How can we prevent rusting in real life? (paint, oil, galvanising, stainless steel)

Explanations

1. Tube A (water + air): Rust forms at the waterline.

   Both oxygen and water are present → Fe → $Fe^{2+} + 2e^-$ (anode); $O_2 + H_2O + e^- \rightarrow OH^-$ (cathode) → rust ($Fe_2O_3 \cdot xH_2O$).

   Tube B (no oxygen): No rust.

   Water but no oxygen (oil stops air).
2. Tube C (no water): No rust.

   Oxygen present but no water → reaction cannot proceed.
3. Tube D (salt water): Fastest rusting.

   Salt increases conductivity→ speeds the electrochemical reactions.
4. Optional E (acidic): Faster rusting than A.

   Acid provides H^+, which speeds oxidation of iron.

Experiment to Investigate the Conditions That Affect Rate of Rusting

Materials

1. Iron nails of uniform size
2. Test tubes with stoppers
3. Test tube rack
4. Sodium chloride solution (various concentrations)
5. Distilled water
6. Dilute acids (HCl)
7. Tap water
8. Thermometer
9. Water bath
10. Iron filings and iron nails (different surface areas)
11. Labels

Procedure

1. Effect of Electrolytes

   Place identical iron nails in separate test tubes containing

   a. Distilled water

   b. 1% NaCl solution

   c. 5% NaCl solution

   d. 10% NaCl solution

   e. Observe for 3-5 days and record the extent of rusting.

2. Effect of pH

   Place identical iron nails in separate test tubes containing

   a. Neutral water (pH 7)

   b. Slightly acidic solution (pH 5)

   c. Acidic solution (pH 3)

   d. Observe for 3-5 days and record the extent of rusting.

3. Effect of Temperature

   Prepare three identical setups with iron nails in salt water.

   a. Keep one at room temperature,

   b. One in a refrigerator and

   c. One in a warm water bath.

   d. Observe for 3-5 days and record the extent of rusting.

4. Effect of Surface Area

   a. Place an iron nail in one test tube

   b. Place an equivalent mass of iron filings in another test tube.

   c. Add equal volumes of salt water to both.

   d. Observe for 3-5 days and record the extent of rusting.

Expected Results

1. Electrolytes: Higher salt concentration leads to faster rusting
2. pH: Lower pH (more acidic) accelerates rusting
3. Temperature: Higher temperature increases the rate of rusting
4. Surface Area: Larger surface area (iron filings) results in faster rusting

Conclusion

The rate of rusting increases with

1. higher concentration of electrolytes
2. lower pH (more acidic conditions)

3. higher temperature
4. greater surface area of the metal

Experiment to Investigate Ways of Preventing Rusting

Materials

1. Clean iron nails
2. Test tubes with stoppers
3. Test tube rack
4. Paint
5. Oil
6. Zinc strips
7. Magnesium strips
8. Copper strips
9. Grease
10. Galvanised nails
11. Saltwater solution (electrolyte)
12. Labels

Procedure

1. Physical barriers
   a. Coat different iron nails with paint, oil and grease
   b. Leave one nail uncoated as a control.
   c. Place each in separate test tubes with salt water.
2. Sacrificial protection
   a. Connect clean iron nails to zinc strip, magnesium strip and copper strip
   b. Place each pair in separate test tubes with salt water.
   c. Include an unconnected iron nail as a control.
3. Galvanisation comparison

   Place a galvanised nail and a regular iron nail in separate test tubes with salt water.
4. Observe all setups for 1 week and record observations.

Expected results

1. Physical Barriers: Nails coated with paint, oil, or grease show little to no rusting
2. Sacrificial Protection: Nails connected to zinc or magnesium show little to no rusting (these metals are more reactive and sacrifice themselves). Nails connected to copper show accelerated rusting (galvanic corrosion)
3. Galvanisation: Galvanised nail shows resistance to rusting

Conclusion

Various methods can effectively prevent rusting by either creating a barrier to oxygen and water or by providing sacrificial protection through more reactive metals.

Preventing Rusting

Rusting happens when iron + oxygen + water react. We can stop or slow down rusting in two main ways:

1. **Non-Redox Methods (Blocking air or water)**

   These methods do not involve chemical reactions; they just cover the iron to stop oxygen and water touching it.

   Painting → e.g., painted gates, cars.

   Oiling or Greasing → e.g., bicycle chains.

   Plastic Coating → e.g., wire coverings.

   Tin Plating → food cans are coated with tin to stop rust.

2. **Redox Methods (Sacrificial protection using another metal)**

   These methods use another metal that reacts more easily than iron. That metal is "sacrificed" (it corrodes first) to protect the iron.

   a. Galvanising → coating iron with zinc. Zinc rusts first, protecting the iron underneath.

   b. Sacrificial Anode → attaching blocks of magnesium or zinc to ships, underground pipes, or oil rigs. These metals corrode instead of the iron.

   c. Alloying → making stainless steel (iron + chromium + nickel). Chromium forms a thin, protective oxide layer that stops rusting.

Everyday Examples (Ghana)

1. Painted roofing sheets → non-redox protection.
2. Galvanised roofing sheets → redox protection (zinc protects the iron).

3. Stainless steel cutlery → alloy protection.

Impact of Rusting and Corrosion in Everyday Life

1. On Buildings and Homes
   a. Roofing sheets: Rust causes holes → leakages during rain.
   b. Pipelines: Corrode and burst → water wastage and high repair costs.
   c. Bridges and structures: Rust weakens iron beams → can cause accidents.
2. On Transport
   a. Cars, buses, and bicycles rust if not painted or maintained → dangerous and costly repairs.
   b. Ships and boats corrode quickly in salty sea water → need regular maintenance.
   c. Airplanes must be protected from corrosion to avoid mechanical failures.
3. On Environment and Heritage
   a. Historic buildings like Cape Coast and Elmina Castles require constant maintenance because of salty coastal air.
   b. Corroded tanks and containers can leak harmful chemicals into soil and water.
4. On the Economy
   a. Billions are spent worldwide each year repairing or replacing corroded equipment.
   b. In Ghana, money that could improve schools and hospitals is also used to fix rust damage (e.g., bridges, vehicles, smelters).
5. On Daily Life
   a. Rust makes tools blunt and unsafe.
   b. Household items like nails, cutlery, and fences lose strength.
   c. Corroded cooking utensils can contaminate food.

Activity 4.18 Understanding Rusting and Corrosion

1. Use the video link below to watch corrosion of metals
   https://chatgpt.com/c/68b0bef7-ad60-8324-96c7-43d8c372569b

2. Observe the two metal objects below.



a. What happened to this one (the brown metal)?

b. Why do you think the other one is not brown?

3. **a.** Write your own definition of "rusting" and "corrosion."

b. Compare your definitions with a partner and improve it.

4. What differences do you notice in how these metals have changed?



5. **a.** What differences do you notice in colour and texture?

b. What might those colours tell us about the **products** formed

c. Write the reaction steps of rusting process (anodic and cathodic processes).

d. Write the overall reaction.

e. Compare your answers.

6. Compare galvanic and non-galvanic methods for protecting iron from rust. What are the advantages and disadvantages of each method?

7. What do the words 'rust' and 'decay' have in common?

a. Use the link below to watch the video on rusting

https://www.youtube.com/watch?v=h1A0yAcavu0

i. What changes do you see on the metal surface?

ii. What role do water and air play in this transformation?

b. Use the link below to watch the video on Fruit Decay
https://www.youtube.com/watch?v=S12zZhdOckc&list=PL0FB8D619E5C752AE

i. How does the fruit change?

ii. What do you think causes those changes?

Activity 4.19 Rusting of Metals

1. Mystery Rust Stations

Aim: Work out which conditions made some nails rust and others not.

Materials

Pre-made station cups with nails (teacher prepares dry air; water only; salty water; oil-covered water; vinegar; painted nail; galvanized nail if available), labels **A–F**.

Procedure

a. Move around Stations A–F. Look only (no touching).

b. Observe nail colour, flakes, liquid, bubbles, smell.

c. State the condition you think is inside (air / water / salt / oil / acid / paint).

d. Write one why sentence for each station: "We think A is salt water because ___."

2. Rust Race (3–5 days)

Goal: Make a setup that rusts an identical nail the fastest.

Rules (one-variable only): 1 nail • 1 cup • 50 ml liquid • add only ONE extra (1 tsp salt or 1 tsp sugar or 1 tsp vinegar).

Do this

a. Choose the one change you'll test (salt? acid?). Label cup with team + variable.

b. Set up the nail in the liquid.

c. Each day, score rust 0–3. Note colour/flakes/cloudiness.

d. Final day: Share results, which set up generated the most rust? How did you ensure that this race was fair?

Safety

Do not touch rusty nails with bare hands; use paper towel; wash hands.

Rust Score Key

a. **0** = shiny, no change

b. **1** = slight colour change

c. **2** = obvious brown spots or flakes

d. **3** = heavy rust / lots of flakes

3. Crime Scene Investigation: The Rusty Disaster

Task: Your class has been called to investigate a *crime scene* where a metal structure failed because of corrosion.



a. Look at the photos and objects from the "crime scene" (e.g., rusted bolt, broken gate hinge, roof sheet).

b. Write down the clues you see rust colour, flakes, cracks, peeled paint, or green coating.

c. Read the short witness cards

d. In groups, decide the main cause of the failure

e. Fill in a simple Claim-Evidence-Reasoning sheet:

- **i.** Claim: What caused the failure.
- **ii.** Evidence: What clues you saw.
- **iii.** Reasoning: Why that evidence shows corrosion.

f. Share your group's conclusion with the class.

4. Anti-Rust Challenge checks (2–3 days)

Goal: Create the best rust-prevention method with limited materials.

Do this

a. Coat half of a nail (e.g., oil, wax, paint, tape, plastic wrap). Leave the other half bare.

b. Put all nails in the same liquid (class agrees). Label and predict which works best and why (barrier keeps out water/oxygen).

c. Each day, score rust 0–3 on each half; take photos/sketches.

d. Make a poster: method, before/after, scores, and why it worked (or not).

Rust Score Key

a. **0** = shiny, no change

b. **1** = slight colour change

c. **2** = obvious brown spots or flakes

d. **3** = heavy rust / lots of flakes

5. One-Variable Rust Test + Graph

Task: Work in groups to test how *one factor* (salt, temperature, covering, or oxygen) affects rusting.

Steps

1. Choose one variable to test

- **a.** Salt level (0, ½, 1, 2 tsp in water)
- **b.** Temperature (sunlight, shade, fridge)
- **c.** Oxygen (cup half-full vs filled to the top and sealed with oil)
- **d.** Covering (painted vs unpainted nail)

2. Set up your experiment

- **a.** Place identical nails in cups of water with your chosen change.
- **b.** Label each cup clearly.

3. Daily check (3–5 days)

- **a.** Look for colour change, flakes, or cloudiness.
- **b.** Score rust on a scale: 0 (no rust) → 3 (lots of rust).
- **c.** Record results in a table.

4. Make a graph

- **a.** Plot rust level (0–3) on the Y-axis.
- **b.** Plot your variable change (salt level, temp, etc.) on the X-axis.

5. Share findings
   a. Show your graph to the class.
   b. Complete the sentence: *"Rust was fastest when ___ because ___."*

Safety

Do not touch rusty nails with bare hands; use paper towels and wash hands after.

Rust Score Key

a. **0** = shiny, no change
b. **1** = slight colour change
c. **2** = obvious brown spots or flakes
d. **3** = heavy rust / lots of flakes

Activity 4.20 Rust Investigations & Anti-Rust Methods

1. Gallery Walk
   a. Post your experiment results and rust prevention method.
   b. Rotate around the room.
   c. Add a sticky note with a *question* or *comment* to each group's poster.
2. Rust Prevention Advertisement
   a. Create a 60-second commercial for your best prevention method.
   b. Show why it works and convince others to use it.
3. Real-Life Impact Stations
   a. Visit stations with photos of rust on bridges, cars, and pipes.
   b. Discuss with your group:
      i. How does rust affect safety?
      ii. How does it cost money?
4. Concept Mapping Challenge: Work in teams to draw a big map linking
   a. Rusting process
   b. Prevention methods
   c. Real-life impacts

REVIEW QUESTIONS

Part A

Multiple Choice Questions

1. Which of the following best defines oxidation?

A. Loss of protons

B. Gain of electrons

C. Loss of electrons

D. Gain of neutrons

2. Which of the following metals is the most reactive?

A. Copper

B. Zinc

C. Iron

D. Potassium

3. In the reaction: $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$, which substance is reduced?

A. Zinc

B. Copper

C. Sulphate

D. Oxygen

4. Which observation best indicates a metal is reacting with acid?

A. Effervescence

B. Formation of a precipitate

C. Colour change

D. Condensation

5. What is the reducing agent in the reaction: $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$?

A. Fe_2O_3

B. CO

C. Fe

D. CO_2

6. Which of the following metals will not displace hydrogen from dilute acid?

A. Zinc

B. Magnesium

C. Copper

D. Iron

7. A student places iron in a solution of copper (II) sulphate. What is observed?

A. Gas evolved

B. Copper is deposited

C. A white precipitate form

D. Nothing happens

8. Which metal will displace iron from iron (II) sulphate?

A. Copper

B. Gold

C. Magnesium

D. Silver

9. Which of the following is not an oxidation reaction?

A. Rusting of iron

B. Combustion of methane

C. Electrolysis of water

D. Freezing of water

10. What is the role of electrons in redox reactions?

A. Gained in oxidation only

B. Lost in both oxidation and reduction

C. Transferred from reducing to oxidising agents

D. Used up in neutralisation

11. Which of the following reactions is NOT a redox reaction involving changes in oxidation states of elements?

A. $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

B. $HCl + NaOH \rightarrow NaCl + H_2O$

C. $2Mg + O_2 \rightarrow 2MgO$

D. $4Fe + 3O_2 \rightarrow 2Fe_2O_3$

12. Which of the following reactions does NOT involve oxidation or reduction?

A. $Ag^+ + Br^- \rightarrow AgBr$

B. $2H_2S + SO_2 \rightarrow 3S + 2H_2O$

C. $2[Ag\ (NH_3)_2]^+ + Cu \rightarrow Cu[(NH_3)_4]^{2+} + 2Ag$

D. $2Al + 2OH^- + 6H_2O \rightarrow 2Al\ (OH)_4^- + 3H_2$

13. In the reaction below, which substances act as oxidising agents?
$2MnO_2 + 4KOH + O_2 + Cl_2 \rightarrow 2KMnO_4 + 2KCl + 2H_2O$

A. $KMnO_4$ only

B. MnO_2 only

C. MnO_2 and O_2

D. O_2 and Cl_2

Essay Questions

1. **a.** List the metals above in order of decreasing reactivity.

b. Define oxidation, in electronic terms, using one example from above.

c. Define reduction, in terms of oxidation number, using one example from above.

d. State and explain which is the strongest reducing agent in the examples above.

e. State and explain which is the strongest oxidising agent in the examples above.

f. Deduce whether a gold coin will react with aqueous magnesium nitrate.

2. How can you use displacement reactions to test which metals are more reactive than others?

3. You have magnesium, zinc, and copper. How would you carry out an experiment to find out which one reacts the fastest with dilute hydrochloric acid?

4. Why do some metals like potassium and sodium react with water, but others like copper do not?
5. Redox reactions happen in many places around us. Give three real-life examples of redox reactions and explain their importance.
6. What is the chemical equation for rusting? Explain how both oxidation and reduction happen during the rusting of iron.
7. Plan a simple classroom experiment that shows both oxidation and reduction happening at the same time.
8. Some metals react with oxygen, some with water, and others with acids. Compare their reactions and use your observations to arrange the metals in order of reactivity.
9. Imagine putting a piece of iron into a solution of silver nitrate. Explain what will happen.
10. A science laboratory in Kumasi received five unlabelled bottles, each containing a clear solution. Preliminary tests were conducted, and the following observations were recorded:

Solution	Test with KI	Test with $KMnO_4$	Test with $K_2Cr_2O_7$	pH
A	Brown colour formed	Purple colour remains	Orange colour remains	3.5
B	No change	Decolourised	Green colour formed	6.5
C	Brown colour formed	Decolourised	Green colour formed	2.0
D	No change	Purple colour remains	Orange colour remains	7.0
E	Brown colour formed	Purple colour remains	Green colour formed	4.0

The laboratory supervisor believes these solutions might be: hydrogen peroxide (H_2O_2), sodium chloride (NaCl), iron (II) sulphate ($FeSO_4$), sodium sulphite (Na_2SO_3), and copper (II) sulphate ($CuSO_4$).

a. Based on the test results, identify each solution (A through E), justifying your answers with specific reference to the chemical behaviour of oxidising and reducing agents.

b. Design a flowchart showing how you could systematically identify these five solutions using only two chemical tests. Explain your reasoning.

11. a. Write the oxidation and reduction half-equations for the following reaction

$2Mg + O_2 \rightarrow 2MgO$

b. Assign oxidation numbers to all elements in this reaction and identify the species oxidised and reduced

$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$

12. a. Balance the redox equation in an acidic medium

$MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$

b. Balance the equation in basic medium

$Cr(OH)_3 + ClO^- \rightarrow CrO_4^{2-} + Cl^-$

Redox titrations

13. a. Design an experiment to compare the reactivity of three metals: Mg, Zn, and Cu, using dilute HCl. Describe how to determine which is most reactive.

b. A student mixed iron nails with copper (II) sulphate solution and observed a brown deposit. Explain the redox changes using half-equations.

14. Evaluate the environmental implications of redox reactions in mining (e.g., gold or bauxite extraction). Suggest one way to reduce negative impacts.

15. a. Define the term standardisation in titration.

b. Write the balanced redox equation for the reaction between $KMnO_4$ and Fe^{2+} in acidic solution.

c. What colour change is observed at the endpoint when $KMnO_4$ is used as the titrant?

d. Identify the oxidising and reducing agents in this redox titration.

16. a. If 25.0 cm^3 of Fe^{2+} solution requires 22.60 cm^3 of 0.0200 mol/dm^3 $KMnO_4$, calculate the number of moles of $KMnO_4$ used.

b. Using your answer above, calculate the number of moles of Fe^{2+} that reacted.

c. From the moles and volume of Fe^{2+}, calculate its concentration in mol/dm³.

d. Describe two safety precautions specific to $KMnO_4$ during titration.

17. a. A student obtained titre values of 21.30 cm³, 22.60 cm³, and 22.50 cm³. Identify the concordant titres and justify the selection.

b. Compare the advantages of using $KMnO_4$ as a self-indicator versus using an external indicator like starch.

c. A sample of Fe^{2+} was exposed to air for some time. Explain how this affects the standardisation titration results.

d. Explain the role of dilute sulphuric acid in the titration and why hydrochloric acid is not preferred.

18. a. A solution was made by dissolving 55.6 g of $(NH_4)_2Fe(SO_4)_2 \cdot 6H_2O$ per dm³. Calculate the theoretical concentration of Fe^{2+} in mol/dm³ and compare it with an experimentally determined value. Discuss possible sources of error.

b. A student standardised $KMnO_4$ and obtained a concentration higher than expected. Propose and evaluate at least two experimental errors that may have caused this deviation.

c. Design an inquiry-based experiment to compare the rate of reaction between $KMnO_4$ and Fe^{2+} under different acid concentrations. Describe how you would collect and interpret the data.

19. a. State the chemical formula for ethanedioic acid and its role in redox titration.

b. What is the chemical formula of the ethanedioate ion?

c. Write the balanced redox equation for the reaction between $KMnO_4$ and $C_2O_4^{2-}$ in an acidic medium.

d. What colour change is observed at the endpoint of this titration?

e. Identify the oxidising agent and the reducing agent in the reaction.

20. a. If 25.0 cm³ of ethanedioate solution requires 20.00 cm³ of 0.0200 mol/dm³ $KMnO_4$ to reach the endpoint, calculate the moles of $KMnO_4$ used.

b. Use the balanced equation to calculate the moles of ethanedioate ions that reacted.

c. Calculate the concentration of the ethanedioate solution in mol/dm³.

d. Explain why the titration mixture must be heated to 60–70°C before starting the titration.

21. a. Describe the effect of not heating the titration mixture on the accuracy of the results.

b. A student consistently observed a delayed colour change at the endpoint. Suggest possible reasons for this.

c. Compare the endpoint visibility of $KMnO_4$ vs ethanedioate titrations and $KMnO_4$ vs Fe^{2+} titrations.

d. Discuss the importance of an acidic medium in the $KMnO_4$ and ethanedioate titration.

22. a. A student dissolved 6.30 g of ethanedioic acid dihydrate ($H_2C_2O_4 \cdot 2H_2O$, molar mass = 126 g/mol) in 250 cm³ of solution.

Calculate the theoretical concentration in mol/dm³ and compare with a value obtained from titration.

b. Suppose the $KMnO_4$ solution was not properly standardised.

Describe how this would affect the accuracy of the ethanedioate concentration obtained.

c. Design an investigation to determine the effect of temperature on the rate of the redox reaction between $KMnO_4$ and ethanedioate. Include how you would ensure reliability and safety.

23. a. What is the role of sodium thiosulphate in redox titration?

b. Write the ionic equation for the reaction between iodine and sodium thiosulphate.

c. What is the function of potassium iodide (KI) in the titration process?

d. What indicator is commonly used in this titration, and what colour change does it show at the endpoint?

e. Identify the oxidising and reducing agents in the titration.

24. a. A student titrated 25.0 cm³ of iodine solution with 0.100 mol/dm³ sodium thiosulphate. The average titre was 23.50 cm³. Calculate the moles of $Na_2S_2O_3$ used.

b. From the balanced equation, determine the moles of iodine that reacted.

c. Calculate the concentration of iodine in the 25.0 cm^3 sample.

d. Explain why potassium iodide is added in excess during the titration.

25. a. Compare the visibility and sharpness of the endpoint in this titration versus that of a $KMnO_4$ titration.

b. A student added starch at the beginning of the titration instead of near the endpoint. What could go wrong and why?

c. During the titration, the solution in the conical flask was left to stand for too long before the titration. What could happen, and how would it affect the results?

d. Why is it important to carry out the titration away from direct sunlight?

26. a. A solution containing iodine was standardised using sodium thiosulphate. After analysis, the concentration of iodine was found to be lower than expected. Suggest possible experimental errors that may explain this observation.

b. Design an experiment to investigate the effect of light exposure on the stability of iodine solution. Include a control and a method of measuring changes in iodine concentration.

c. Suppose you want to analyse iodine content in a seaweed extract. Outline the titration procedure and modifications needed for sample preparation and reliability.

Part B

Multiple Choice Questions

1. In a voltaic cell, which type of reaction occurs at the anode?

A. Reduction

B. Oxidation

C. Neutralisation

D. Precipitation

2. The function of the salt bridge in a voltaic cell is to:

A. Provide a pathway for electrons to flow

B. Maintain electrical neutrality by allowing ion movement

C. Increase the voltage of the cell

D. Reduce resistance in the external circuit

3. Which of the following is a correct example of a redox reaction producing electricity?

A. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

B. $HCl + NaOH \rightarrow NaCl + H_2O$

C. $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

D. $CaCO_3 \rightarrow CaO + CO_2$

4. In a Daniell cell, the zinc electrode is in $ZnSO_4$ solution and the copper electrode is in $CuSO_4$ solution. Which statement is correct?

A. Electrons flow from copper to zinc

B. Zinc is reduced

C. Copper is oxidised

D. Zinc is the anode

5. In a voltaic cell, electrons flow:

A. From the cathode to the anode through the salt bridge

B. From the anode to the cathode through the external circuit

C. From the cathode to the anode through the external circuit

D. From the anode to the cathode through the salt bridge

6. Which of the following best explains why a redox reaction can produce electricity?

A. The transfer of electrons between reactants is directed through an external circuit

B. The reaction occurs very slowly

C. The reaction involves gases that move freely

D. The reaction increases the temperature of the solution

7. A voltaic cell consists of a magnesium electrode in Mg^{2+} solution and a silver electrode in Ag^+ solution. Which electrode is the anode and why?

A. Silver, because Ag is more reactive than Mg

B. Silver, because Ag^+ ions are more easily reduced

C. Magnesium, because Mg is more reactive and more easily oxidised

D. Magnesium, because Mg^{2+} ions are more easily reduced

Essay Type Questions

1. Describe the basic components of a voltaic cell and explain the role of each in generating electricity.

2. Explain, with the aid of a diagram, how the Daniell cell works, including: redox reactions, electron flow, and purpose of salt bridge.

3. Compare and contrast the operation of a voltaic cell and an electrolytic cell.

4. A magnesium–copper voltaic cell is set up under standard conditions. Write the half-reactions, calculate E°cell, state spontaneity.

5. Given the following half-cell reduction potentials:

$E°(Ag^+/Ag) = +0.80$ V1

$E°(Zn^{2+}/Zn) = -0.76$ V1

Calculate the standard cell potential (E°cell) for the cell:

$Zn(s) \mid Zn^{2+}(aq) \parallel Ag^+(aq) \mid Ag(s)$

6. Given the following reduction potentials:

$E°(Cu^{2+}/Cu) = +0.34$ V

$E°(Pb^{2+}/Pb) = -0.13$ V

a. Write the balanced overall cell reaction for the voltaic cell.

b. Calculate E°cell.

c. Predict whether the reaction is spontaneous under standard conditions.

7. A galvanic cell is constructed from a magnesium half-cell and a tin half-cell.

$E°(Mg^{2+}/Mg) = -2.37$ V

$E°(Sn^{2+}/Sn) = -0.14$ V

a) Identify the anode and cathode

b) Calculate the standard cell potential

c) Determine the value of $\Delta G°$ for the cell reaction at 298 K.

($F = 96{,}500\ C\ mol^{-1}$; $n = 2$)

8. Consider the following half-reactions and their E° values:

$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ $E° = +1.36\ V$

$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$ $E° = +1.09\ V$

a) Which species is the stronger oxidising agent?

b) Write the overall cell reaction when these two half-cells are connected.

c) Calculate E°cell.

9. You are given the following data

$E°(Fe^{3+}/Fe^{2+}) = +0.77\ V$

$E°(I_2/I^-) = +0.54\ V$

a) Calculate E°cell for the reaction1

$2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(s)$

b) Calculate $\Delta G°$ for the reaction ($n = 2$; $F = 96{,}500\ C\ mol^{-1}$).

c) Comment on the spontaneity of the reaction.

d) Explain how changing the concentration of Fe^{3+} ions might affect the cell potential.

10. (a) Define standard electrode potential.

(b) State the condition under which a redox reaction is feasible based on E°cell.

(c) List two industrial applications of electrode potential principles.

(d) State the electrode potential sign convention for a reduction reaction.

11. (a) Given $E°(Zn^{2+}/Zn) = -0.76\ V$ and $E°(Cu^{2+}/Cu) = +0.34\ V$, calculate the standard cell potential for a Daniell cell and predict if the reaction is spontaneous.

(b) Arrange the following metals in order of increasing reactivity based on their E° values: Fe^{2+}/Fe ($-0.44\ V$), Al^{3+}/Al ($-1.66\ V$), Cu^{2+}/Cu ($+0.34\ V$).

(c) State two ways in which electrode potentials are used in corrosion prevention.

12. A fuel cell operates with the following half-reactions

$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$ $E^\circ = +1.23$ V

$H_2(g) \rightarrow 2H^+(aq) + 2e^-$ $E^\circ = 0.00$ V

(a) Identify the anode and cathode.

(b) Calculate E°cell.

(c) State whether the reaction is feasible under standard conditions.

(d) Explain how the concept of sacrificial protection works using electrode potentials.

13. In an experiment, a student measures the EMF of a Zn–Ag cell under non-standard conditions and obtains a value lower than the standard EMF.

(a) Suggest two factors that could cause this decrease.

(b) Using the Nernst equation, explain how ion concentrations affect the EMF.

(c) Discuss the role of electrode potentials in designing rechargeable batteries, giving examples of suitable electrode materials and their E° values.

14. **(a)** Define **electrode potential**.

(b) State two standard conditions for measuring standard electrode potential.

(c) State **two applications** of standard electrode potentials in chemistry.

15. **(a)** Arrange the following metals in order of increasing reactivity using their standard electrode potentials: Mg, Cu, Zn.

(b) Using standard electrode potentials, predict whether a solution of silver nitrate will oxidise copper metal.

(c) Explain why potassium cannot be obtained by the electrolysis of aqueous potassium chloride solution, using electrode potentials.

16. **(a)** A galvanic cell is constructed from Zn/Zn^{2+} and Cu/Cu^{2+} half-cells.

(i) Draw a labelled diagram of the cell.

(ii) Write the half-cell reactions and the overall cell reaction.

(iii) Calculate E°cell given: $E^\circ(Zn^{2+}/Zn) = -0.76$ V, $E^\circ(Cu^{2+}/Cu) = +0.34$ V.

(iv) State whether the reaction is spontaneous.

(b) Explain, with the aid of standard electrode potentials, why hydrogen is liberated when zinc reacts with dilute hydrochloric acid but copper is not.

17. Given the standard electrode potentials

$E°(Fe^{3+}/Fe^{2+}) = +0.77$ V,

$E°(Zn^{2+}/Zn) = -0.76$ V,

$E°(Cu^{2+}/Cu) = +0.34$ V,

(a) Predict whether zinc metal will reduce Fe^{3+} to Fe^{2+} in solution.

(b) Justify your answer with a suitable calculation of E°cell.

(c) Suggest one industrial process that uses this principle.

18. An unknown metal M has a standard electrode potential of –0.44 V for M^{2+}/M.

(a) Predict whether M will displace copper from a $CuSO_4$ solution.

(b) Predict whether M will liberate hydrogen gas from dilute H_2SO_4.

(c) Support your predictions with E°cell calculations.

Part C

Electrolytic Cells and Applications

1. **(a)** Define an electrolytic cell.

(b) State the role of the anode and cathode in an electrolytic cell.

(c) Identify one everyday application of electrolysis.

(d) List two examples of electrolytes commonly used in electrolytic cells.

2. **(a)** Explain why a direct current (DC) power supply is used in electrolytic cells instead of alternating current (AC).

(b) Describe what happens to cations and anions during electrolysis.

(c) Give two examples of how electrolysis is applied in industry.

(d) Illustrate with a simple diagram the movement of ions in an electrolytic cell.

3. **(a)** Compare the processes at the electrodes of an electrolytic cell with those in a voltaic cell.

(b) Predict the products at the electrodes when aqueous sodium chloride solution is electrolysed. Explain your reasoning.

(c) Analyse the importance of electroplating in protecting metals used in daily life.

(d) Justify the use of graphite electrodes in the electrolysis of molten sodium chloride instead of metal electrodes.

4. Evaluate the environmental impact of using electrolysis in industries such as aluminium extraction or the chlor-alkali process.

5. Design a simple classroom demonstration to show the electrolysis of water and explain how this connects to sustainable fuel production.

6. Synthesise information from industrial and household uses of electrolysis to propose new innovative applications in modern technology.

7. Critically assess the economic advantages and disadvantages of using electrolysis for large-scale metal extraction.

Part D

Faraday's Laws of Electrolysis

1. a. State Faraday's First Law of Electrolysis.

b. Write the formula that combines current, time, molar mass, and charge to calculate the mass deposited at an electrode.

c. What is the value of the Faraday constant (F)?

d. Which scientist formulated the laws of electrolysis?

e. At which electrode is copper deposited when a solution of copper (II) sulphate is electrolysed with copper electrodes?

2. a. Explain why more copper is deposited when electrolysis runs for a longer time.

b. A current of 2.0 A passes through silver nitrate solution for 30 minutes. Calculate the mass of silver deposited.

(M_{Ag} = 108 g mol^{-1}, n=1, F=96,500 C mol^{-1})

3. a. Using Faraday's Second Law, explain why passing the same charge through solutions of silver nitrate and copper sulfate produces more silver than copper.

b. Why does aluminium extraction require a large amount of electricity?

c. Which electrode process in the Hall–Héroult cell releases oxygen gas?

4. A current of 3.0 A is passed through molten sodium chloride for 10 minutes. Calculate the mass of sodium formed. ($M_{Na} = 23$ g mol^{-1}, $n=1$)

5. Compare the processes of electroplating copper onto jewellery and electrorefining impure gold using Faraday's Laws.

6. In electrolysis of copper (II) sulfate using inert electrodes, explain why the solution becomes less blue with time.

7. Describe how Faraday's First Law can be used to determine the value of Avogadro's number experimentally.

8. A plating industry requires 5.0 g of silver to coat a set of ornaments. Calculate the quantity of electricity needed.

9. Ghana's VALCO smelter consumes massive amounts of electricity for aluminium production. Use Faraday's First Law to explain why electricity cost is the largest factor in aluminium price.

10. Mining in Ghana provides both gold and bauxite (aluminium ore). Discuss how Faraday's Second Law helps explain differences in refining costs between gold and aluminium.

11. Design an experiment using Faraday's First Law to calculate the charge on an electron. What measurements would be needed?

12. Evaluate the environmental impacts of large-scale aluminium and gold electrolysis plants in Ghana, connecting them to Faraday's Laws.

13. A gold refinery passes a current of 500 A for 24 hours through an impure gold anode. Estimate the mass of gold refined and discuss how this contributes to Ghana's export revenue.

Part E

1. a. Define corrosion and distinguish it from rusting.

b. State two conditions necessary for the rusting of iron.

c. Write a balanced chemical equation for the rusting process in the presence of water and oxygen.

2. a. List three methods of preventing rusting and state how each works.

b. Why does a zinc coating (galvanisation) protect iron better than a paint coating?

c. Explain briefly why aluminium appears resistant to corrosion even though it is very reactive.

3. A bridge built near the sea rusted severely within five years, while a similar bridge inland remained rust-free

a. Suggest reasons for the difference

b. Design a simple experiment to compare the effect of salt water and fresh water on the rusting of nails.

4. A class tested rusting using identical iron nails:

Condition	Rust score after 72 h (0–3)
Dry air	0
Tap water	1
Salt water	3
Boiled water + oil layer	0

a. Identify the independent variable and state two controlled variables.

b. Rank the conditions from slowest to fastest rusting and give a brief reason.

c. Explain why boiled water plus oil prevented rusting.

d. Predict what happens if a copper strip touches the iron nail in salt water. State which metal corrodes; explain using anode/cathode ideas.

5. **a.** Explain how corrosion affects transportation, infrastructure, and the economy.

b. Suggest and evaluate two large-scale strategies to minimise corrosion damage.

c. If you were a materials engineer, which corrosion-prevention method would you recommend for pipelines in Ghana and why?

SECTION

5

PHYSICAL AND CHEMICAL PROPERTIES OF THE TRANSITION ELEMENTS



SYSTEMATIC CHEMISTRY OF THE ELEMENTS

Periodicity

INTRODUCTION

In this section, you will learn about a fascinating group of metals known as transition elements. These metals are distinct from the main group elements because of their unique properties. For instance, they can display different oxidation states, meaning they have more than one charge. You will also discover that transition elements often form colourful compounds, can act as catalysts in chemical reactions, and readily form complex compounds with other molecules.

By the end of this section, you will have the skills to find transition elements on the Periodic Table, explain their special properties through hands-on experiments, and understand how complex compounds are formed and named. Additionally, you will compare transition elements with main group metals and appreciate their numerous applications in daily life, including their roles in construction, electronics, medicine, and environmental protection.

KEY IDEAS

- **Transition elements**: a special group of metals in the middle of the periodic table (d-block) with unique properties.
- **Variable oxidations state**: the ability to show more than one charge
- **Complex compound**: a compound where a transition metal ion bonds with other molecules/ions called ligands (e.g., $[Cu(NH_3)_4]^{2+}$).

PROPERTIES OF TRANSITION ELEMENTS

Transition Elements

What are Transition Elements?

Transition elements are elements with a partially filled d sub-shell or the capacity to produce cations with an incomplete d sub-shell. Transition elements are special metals found in the middle of the periodic table (*Groups 3–12*). They have a partly filled d-orbital. Because of this, they can form different ions (charges) and show many unique properties.

Note

Zinc, cadmium, and mercury usually have their d-orbitals filled. They do not always behave like typical transition metals, but we still study them as part of this group.

The general way we write electron configurations is: **[noble gas] $d^{1–10}$ $ns^{1–2}$**. Some elements (like chromium and copper) are exceptions because it is more stable for them to have half-filled or filled d-orbitals.

Properties of Transition Elements

Physical Properties

1. Good conductors of heat and electricity.
2. High melting and boiling points (except mercury).
3. Hard and strong metals (used in construction).
4. High density compared to main group metals.
5. Malleable and ductile (can be shaped and stretched).

Characteristic Properties of Transition Elements

1. **Variable Oxidation States**

   Transition metals can lose different numbers of electrons, forming ions with different charges. For example, Iron can exist as Fe^{2+} (iron II) or Fe^{3+} (iron III).

   Application: This allows them to take part in many redox reactions. For example, manganese dioxide (MnO_2) with Mn^{4+} is used in dry cells (batteries).

2. Coloured Compounds

Transition metals often form salts and solutions that have bright colours due to d–d electron transitions. For example, Copper (II) sulphate is blue, potassium dichromate is orange, nickel compounds are green.

Application: Used as colouring agents in paints, inks, ceramics, stained glass, and fireworks.

3. Catalytic Properties

Transition metals and their compounds speed up reactions by providing an alternative pathway with lower activation energy. They are not used up in the process. For instance, Iron is a catalyst in the Haber process (making ammonia).

Application: Used as industrial catalysts e.g.,

a. Iron (Fe) for ammonia production.

b. Vanadium(V) oxide (V_2O_5) in manufacturing of sulfuric acid (Contact process).

c. Platinum (Pt) as catalytic converters in car exhaust systems.

4. Complex Formation

Transition metals form complex ions by bonding with other ions or molecules (ligands). Example: $[Cu(NH_3)_4]^{2+}$ is where copper forms a complex with ammonia.

Application: Complexes are used in medicine (cisplatin, a platinum complex, is used to treat cancer) and analytical chemistry (detecting and identifying ions).

5. Magnetism

Some transition elements (iron, cobalt, nickel) have unpaired d-electrons that make them magnetic.

Application: Used to make magnets for electric motors, transformers, loudspeakers, and data storage devices.

6. High Melting and Boiling Points

Transition metals have strong metallic bonds due to delocalised d-electrons, giving them high melting and boiling points.

Application: Used in construction of buildings, bridges, and tools where strong and heat-resistant materials are needed (e.g., tungsten filaments in bulbs).

7. **Good Conductors of Heat and Electricity**

   The presence of delocalized electrons allows easy flow of charge and heat.

   Application: Copper wires for electrical wiring, silver in electronics, platinum in electrodes.

Experiments to Show Catalytic Behaviour

General Safety (read before you start)

Wear goggles and gloves. Handle 3% H_2O_2 with care; avoid eyes/skin.

Do not inhale powders (e.g., MnO_2). Clean spills with plenty of water.

Dispose of solutions per school policy; do not pour heavy-metal salts down the sink.

Decomposition of Hydrogen Peroxide with Manganese (IV) Oxide (MnO_2)

Aim: Show that a transition element catalyst speeds decomposition of H_2O_2 and is not consumed.

Materials

1. 3% H_2O_2 (30–50 ml)
2. A pinch of MnO_2
3. 100 ml beaker
4. Gas collection (syringe or inverted cylinder)
5. Stopwatch
6. Filter paper (to recover MnO_2)

Procedure

1. Set up to collect oxygen gas from the beaker.
2. Add H_2O_2 to the beaker and start the timer the moment you add a pinch of MnO_2.
3. Record the O_2 volume every 10–20 s for 2–3 minutes.
4. Repeat a control run with the same H_2O_2 but without MnO_2.

Table 1A: Catalysed run (MnO_2 present)

Time / s	O_2 volume / mL

Table 1B: Control run (no catalyst)

Time / s	O_2 volume / mL

Record/Observe

Compare bubble rate catalysed vs control.

Filter, wash, dry MnO_2 and note that mass/appearance remains (not consumed). Confirm through a suitable gas test that the gas collected is in fact oxygen.

Calculations (show working)

1. Plot O_2 volume (mL) vs time (s) for both runs on the same axes.
2. From initial slope, compare relative rates.
3. Explain in terms of activation energy and surface sites.
4. Write the balanced chemical reaction for this experiment (remember to include the state symbols).

Iodide-Catalysed Decomposition of Hydrogen Peroxide (KI)

Aim: Show that a dissolved iodide catalyst speeds up O_2 release and is regenerated.

Materials

1. 3% H_2O_2 (20–30 ml)
2. KI solution (0.1 M) or a few crystals
3. Dish soap (1 drop) + food colouring (optional)
4. Small beaker/cylinder
5. Ruler (to measure foam height)
6. Stopwatch

Procedure

1. Place H_2O_2 and one drop of soap in the beaker. Start the timer when KI is added.
2. Measure foam height every 10–15 s until it stops rising.
3. Repeat a control without KI.

Table 2A: With KI (catalyst present)

Time / s	Foam height / cm

Table 2B: Control (no KI)

Time / s	Foam height / cm

Record/Observe

Note immediate vs delayed foam formation.

Record maximum foam height and time to reach it.

Calculations (show working)

1. Plot foam height vs time for both runs.
2. Relate observations to the iodide catalytic cycle ($I^- \rightarrow I_2 \rightarrow I^-$).

Check-Your-Understanding

1. Why does adding more KI (small increase) usually shorten the time to peak height?
2. State how the catalyst is regenerated in this reaction.

Transition-Metal Catalysis: Fe^{3+}/Fe^{2+} in the Persulphate–Iodide Reaction ('Iodine Clock' variant)

Aim: Show a transition metal ion catalyst speeds a reaction via an electron-shuttle and is regenerated.

Materials

1. 0.02 M $K_2S_2O_8$ (persulfate)

2. 0.02 M KI (iodide)
3. Starch indicator
4. M $FeCl_3$ (few drops as catalyst)
5. Beakers, measuring cylinders, stopwatch

Procedure

1. Prepare two identical mixtures of persulphate + iodide + starch.
2. Add a few drops of $FeCl_3$ to only one mixture (catalysed). Start timing both.
3. Record time to blue colour (starch–iodine complex) for each.
4. Optional: vary Fe^{3+} volume and repeat.

Table 3A: Time to blue (with and without Fe^{3+})

Trial	Fe^{3+}? (Y/N)	Fe^{3+} volume / drops	Time to blue / s

Table 3B: Extension — Effect of Fe^{3+} amount

Fe^{3+} drops	Time to blue / s

Record/Observe

Compare times catalysed vs uncatalysed.

Describe the colour change clearly.

Calculations (show working)

Explain Fe cycling between Fe^{2+} and Fe^{3+} and how that speeds the net reaction.

Check-Your-Understanding

1. What evidence shows Fe^{3+} is a catalyst, not a reactant?
2. Predict what happens if Fe^{3+} is doubled.

Activity 5.1 Understanding Transition Elements

1. Form a group with 4-5 classmates.
2. Use the periodic table to identify the first-row transition elements (Sc to Zn). Point them out together.
3. Write your own simple definition of transition elements.
4. Identify the characteristic properties of transition elements and describe each property with the help of textbooks.
5. Discuss and note key ideas as a whole class
   a. Where transition elements are on the periodic table.
   b. Their electronic configuration and d-orbitals.
   c. How properties are linked to electron structure.
   d. Real-life uses in Ghana (e.g., mining, manufacturing, catalysts).

COMPLEXES OF TRANSITION ELEMENTS

Copper Complex Formation Experiment

Aim: To observe how copper (II) ions form a deep-blue complex with ammonia and understand applications of complexes.

Materials

1. Copper (II) sulphate solution (dilute)
2. Ammonia solution (dilute)
3. Sodium hydroxide solution (dilute)
4. 3 test tubes + rack
5. Droppers/pipettes
6. Wooden splint + matches
7. White paper (to see colours clearly)
8. Beaker for waste, water for rinsing

Safety

1. Wear goggles.
2. Do not sniff ammonia; work in a ventilated space.
3. Wash hands after the experiment.

4. Teacher handles any spills.

Procedure

1. Put 2–3 ml of $CuSO_4$ in Test Tube 1. Note the pale blue colour.
2. Add ammonia drop by drop. At first, a light blue solid may appear.
3. Add more ammonia (excess) – the solid dissolves to give a deep royal blue solution.
4. In Test Tube 2, put 2–3 ml $CuSO_4$ and add NaOH solution drop by drop. A light blue solid forms.
5. Add ammonia to this tube – the solid dissolves and turns deep blue.
6. For the mystery test, add ammonia to two clear solutions (A and B). The one that turns deep blue contains copper ions.

Results Table

Tube	What we added	What we saw	What it means
1	$CuSO_4$ + **excess NH_3**	**Deep royal blue** solution	Copper–ammonia **complex** formed
2	$CuSO_4$ + **NaOH**	**Light blue solid** ($Cu(OH)_2$)	Hydroxide **precipitate**
2 → then NH_3	Added ammonia to the solid	Solid **dissolves** → **deep blue**	Complex **re-dissolves** the solid
Mystery	+ NH_3	Deep blue or no change	**Cu^{2+} present** or **absent**

Note

- Copper ions bond with ammonia to form a deep-blue complex.
- Sodium hydroxide only gives a pale blue solid ($Cu(OH)_2$).
- Complexes absorb light differently, giving colours.
- Complexes are important in industry, mining, medicine, and testing for ions.

Real-life Applications

1. Identification test: The deep-blue colour helps confirm copper in a sample (school labs, water testing).
2. Cleaning & extraction: Ammonia and other ligands help dissolve or separate metals in mining and recycling.

3. Medicine & analysis: Metal complexes are used for drug delivery and detecting ions (analytical chemistry).
4. Textiles & materials: Some copper complexes colour fabrics, glass, and ceramics.

Shapes of Common Complex Compounds

a. Linear Complexes (2 ligands opposite each other)

Shape: Straight line.

Hyhridisation of central atom: sp

Examples: $[Ag(NH_3)_2]^+$ (Tollen's reagent), $[CuCl_2]^-$

b. Square Planar Complexes (4 ligands in one flat plane)

Shape: Flat square arrangement.

Hyhridisation of central atom: dsp^2

Examples: $[Ni(CN)_4]^{2-}$, $[Pt(NH_3)_2Cl_2]$ (cisplatin – used in cancer treatment)

c. Tetrahedral Complexes (4 ligands around the metal, not flat)

Shape: Like a 3D pyramid with a triangular base.

Hyhridisation of central atom: sp^3

Examples: $[CuCl_4]^{2-}$, $[ZnCl_4]^{2-}$

d. Octahedral Complexes (6 ligands around the metal)

Shape: Looks like two square pyramids joined base to base.

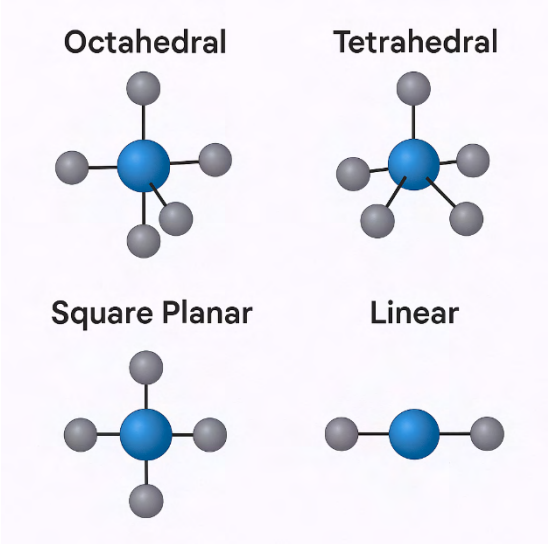
Hyhridisation of central atom: sp^3d^2

Examples: $[Cu(H_2O)_6]^{2+}$, $[Fe(CN)_6]^{3-}$, $[Co(NH_3)_6]^{3+}$

Note

The **number of ligands** (coordination number) determines the shape. Transition metals can form **different shapes** depending on the ligands.

Diagrams of octahedral, tetrahedral, square planar, linear shapes



Naming Complex Compounds

Scientists use special naming rules **(IUPAC rules)** to describe them clearly.

The name tells us:

1. The ligands (what is attached to the metal).
2. The number of ligands.
3. The metal in the centre.
4. The oxidation state (charge) of the metal.

This system makes sure every chemist around the world understands the same compound in the same way.

Example

$[Cu(NH_3)_4]SO_4$ is called **tetraamminecopper (II) tetraoxosulphate (VI)**.

"Tetraammine" = 4 ammonia ligands.

"Copper (II)" = copper with a +2 charge.

"Sulphate" = the outside ion.

Step-by-Step Guide

1. **Name the Ligands (in alphabetical order)**
   Negative ligands: change ending to –o (Cl^- → chloro, CN^- → cyano)
   Neutral ligands: keep normal name (NH_3 → ammine, H_2O → aqua)
2. **Use Prefixes for Number of Ligands**
   1 = mono-, 2 = di-, 3 = tri-, 4 = tetra-, 5 = penta-, 6 = hexa-

3. **Name the Central Metal**
   Neutral/positive complex → normal metal name (e.g., copper, iron)
   Negative complex → add –ate (Fe → ferrate, Cu → cuprate)
4. **Add the Metal's Oxidation State**
   Shown in Roman numerals in brackets (e.g., Copper (II) = Cu^{2+})
5. **Add Any Outside Ions**
   If part of a salt, add the outside ion last (e.g., SO_4^{2-} = sulfate, Cl^- = chloride)

Examples

1. $[Cu(NH_3)_4]SO_4$
   4 ammonia ligands = tetraammine
   Metal = copper
   Oxidation state = (II)
   Outside ion = sulphate
   Name: tetraamminecopper(II) tetraoxsulphate
2. $[Fe(CN)_6]^{3-}$
   6 CN^- ligands = hexacyano
   Negative complex → ferrate
   Oxidation state = (III)
   Name: hexacyanoferrate (III) ion
3. $[CoCl_4]^{2-}$
   4 Cl^- ligands = tetrachloro
   Negative complex → cobaltate
   Oxidation state = (II)
   Name: tetrachlorocobaltate (II) ion

Practice Questions

Name the following complexes

1. $[Cr(H_2O)_6]^{3+}$
2. $[Zn(NH_3)_4]^{2+}$
3. $[Pt(NH_3)_2Cl_2]$
4. $[Fe(H_2O)_6]^{2+}$
5. $[Ni(CN)_4]^{2-}$

Tip: Always follow the order: [Ligands → Metal + Oxidation State → Outside Ion]

Activity 5.2 Solubilising Insoluble Compounds

Aim: To explore how some insoluble compounds dissolve when ligands form complexes.

Work in small groups of learners. Each group will carry out the steps and record observations.

Steps

1. Prepare test solutions: Copper (II) sulphate ($CuSO_4$), Silver nitrate ($AgNO_3$), Zinc sulphate ($ZnSO_4$).
2. Form insoluble precipitates by adding sodium hydroxide (NaOH) or sodium chloride (NaCl).
3. Add ligand solutions like ammonia (NH_3) or excess NaOH drop by drop to the precipitate
4. Observe colour changes and note if precipitates dissolve.
5. Record results in a table.
6. Write balanced equations for
   a. Precipitate formation.
   b. Complex ion formation.
7. Discuss applications of complex formation in real life.
8. Explain why complex formation is important in
   a. Water treatment (removing unwanted ions).
   b. Medicine (metal complexes in drugs).
   c. Mining and extraction (dissolving ores).
   d. Analytical chemistry (detecting ions).

Observation Table

Compound + Reagent	Precipitate Formed	Effect of Excess Ligand	Final Colour/ Observation
$CuSO_4$ + NaOH	Blue $Cu(OH)_2$ solid	Dissolves in excess NH_3	Deep blue solution
$AgNO_3$ + NaCl	White AgCl solid	Dissolves in excess NH_3	Colourless solution
$ZnSO_4$ + NaOH	White $Zn(OH)_2$ solid	Dissolves in excess NaOH	Colourless solution

Results Table

Compound + Reagent	Precipitate Formed	Effect of Excess Ligand	Final Colour/ Observation
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

Balanced Equations

1. _____

2. _____

3. _____

Activity 5.3 Modelling Complex Geometries

Build a 3D model → Draw a 2D picture → Name the complex

Aim: Work in groups to:

1. build a 3D model of a metal complex
2. draw the 2D representation
3. correctly name the complex (IUPAC)

Group Roles

Reader – reads the task card

Builder – leads the model construction

Angle Checker – checks shape & angles

Artist/Recorder – draws 2D + writes name

Presenter – explains to class

Materials

Modelling kits or: beads/buttons (metal), small beads (ligands)

Pipe cleaners/straws/toothpicks (bonds), plasticine (joints)

Rulers, markers, A4 paper

Shape cards:

1. Tetrahedral
2. Square planar
3. Octahedral

Step 1: Get your shape card

Tetrahedral: 4 bonds, angles ≈ 109.5°

Square planar: 4 bonds in a flat square, angles 90°/180°

Octahedral: 6 bonds, angles 90°/180°

Pick one example to build:

1. Tetrahedral: $[ZnCl_4]^{2-}$
2. Square planar: $[Pt(NH_3)_2Cl_2]$
3. Octahedral: $[Fe(H_2O)_6]^{2+}$, $[Co(NH_3)_6]^{3+}$

Step 2 : Build your 3D model

Put the metal bead in the middle

Add the correct number of ligand beads around it

Use straws/pipe cleaners as bonds

Angle Checker: confirm flat (square planar), 109.5° (tetrahedral), 6 at right angles (octahedral)

Step 3: Draw the 2D picture

Square planar: draw a *square* with M in the centre; ligands at corners

Tetrahedral/Octahedral (use cues):

Solid line = in the plane

Wedge ▲= coming out of the page

Dash ▬ ▬ = going behind the page

Step 4: Name your complex (IUPAC)

Order ligands alphabetically (ignore di-/tri- when alphabetising)

Add number prefixes: di-, tri-, tetra-, penta-, hexa-

Write the metal name with oxidation state in Roman numerals if known

Examples

1. $[Fe(H_2O)_6]^{2+}$ → hexaaquairon (II) ion
2. $[Pt(NH_3)_2Cl_2]$ → diamine dichloro platinum (II) ion
3. $[ZnCl_4]^{2-}$ → tetrachloro zincate (II) ion (anionic complexes often end in "-ate")

Step 5: Present to the class

Show your 3D model and 2D drawing

Say the name and why the shape fits (number of ligands + angles)

Step 6: Connect to properties

CN = 4 can be tetrahedral or square planar (affects colour & magnetism)

Octahedral is common for d-block ions in water (e.g., $[Fe(H_2O)_6]^{2+}$)

Real life: catalysts (Pt), pigments/colours (Cu/Ni), medicine (cis-platin)

Quick Naming Helper

Common ligands	Prefixes	Metals
ammine (NH_3)	di- (2)	iron (II/III)
aqua (H_2O)	tri- (3)	copper(I/II)
chloro (Cl^-)	tetra- (4)	zinc (II)
cyano (CN^-)	penta- (5)	nickel (II)
	hexa- (6)	platinum (II)

Success Check (tick √)

Correct geometry (tetrahedral / square planar / octahedral).

Clear 2D drawing (use wedge/dash where needed).

Ligands in alphabetical order + correct prefixes.

Correct metal name with oxidation state.

Presenter explains shape: # ligands + angles.

SIMILARITIES AND DIFFERENCES BETWEEN TRANSITION METALS AND MAIN GROUP METALS

Similarities Between Transition Metals and Main Group Metals

1. Metallic Nature – Both are metals, so they are shiny (lustrous).
2. Conductivity – Both can conduct heat and electricity.
3. Malleability and Ductility – Both can be hammered into sheets (malleable) and stretched into wires (ductile).
4. High Melting and Boiling Points – Most of them are strong and have high melting/boiling points.
5. Form Positive Ions – Both lose electrons easily to form cations.
6. Form Metallic Bonds – In both, atoms are held together by metallic bonding (delocalised electrons).
7. Useful in Everyday Life – Both are used in construction, tools, coins, and machines.

Differences Between Transition Metals and Main Group Metals

Table 5.1: Differences between transition metals and group metals

Feature	Transition Metals	Main Group Metals
Position on Periodic Table	Found in the middle (Groups 3–12, d-block)	Found on the left and right (s- and p-block)
Oxidation States	Show variable charges (e.g., Fe^{2+}, Fe^{3+})	Usually have fixed charges (e.g., Na^{+}, Mg^{2+})
Compound Colour	Form coloured compounds/ solutions (e.g., blue Cu^{2+})	Mostly form colourless compounds

Feature	Transition Metals	Main Group Metals
Catalytic Activity	Many acts as catalysts (e.g., Fe in Haber process)	Rarely act as catalysts
Complex Formation	Readily form complex ions with ligands	Rarely form complexes
Magnetic Properties	Some (Fe, Co, Ni) are magnetic	Usually not magnetic
Strength & Hardness	Generally hard and strong	Many are softer and lighter (e.g., Na, K)

Uses of Transition Metals in Everyday Life

1. **Building & Tools**
   a. Iron/steel (Fe): beams for buildings, bridges, nails, cutlery.
   b. Chromium (Cr) + nickel (Ni) in stainless steel: sinks, spoons, roof fasteners (resist rust).
   c. Manganese (Mn): makes steel tougher.
2. **Electricity & Electronics**
   a. Copper (Cu): wires, phone chargers, motors (great conductor).
   b. Gold (Au) / Silver (Ag): tiny connectors that don't tarnish easily.
3. **Magnets & Motors**
   Iron, cobalt, nickel (Fe, Co, Ni): fridge magnets, speakers, electric motors, generators.
4. **Cleans Air in Cars**
   Platinum, palladium, rhodium (Pt, Pd, Rh) in catalytic converters: change harmful gases into less harmful ones.
5. **Batteries & Energy**
   Nickel, cobalt, manganese (Ni, Co, Mn): parts of many rechargeable batteries.
6. **Health & Medicine**
   a. Iron (Fe) in haemoglobin helps blood carry oxygen.
   b. Cobalt (Co) is part of vitamin B_{12}.

c. Platinum (Pt) medicines (like cis-platin) help treat some cancers (teacher note).

7. **Colours in Glass & Ceramics**
   a. Cobalt → deep blue glass/tiles.
   b. Chromium → green glass/glazes.
   c. Iron → brown/amber bottles.
8. **Coins & Jewellery**
   a. Copper & nickel in many coins.
   b. Gold & silver for jewellery.
9. **Stopping Rust**
   Zinc (Zn) coating (galvanised steel) protects fences, roofs, and water tanks.
10. **Heat & Cutting**
    Tungsten (W), molybdenum (Mo): very hard, used in drill bits, high-temperature parts.

Activity 5.4 Exploring Transition & Main Group Metals

Step 1

1. Open an interactive periodic table app/website/periodic table chart.
2. Highlight: Transition metals → middle block (Groups 3–12) and Main group metals → left and right side.
3. Underline the metals you see in daily life (coins, wires, pots, jewellery) on the periodic table.
4. Sorts the underlined metals into
   a. Transition metals (iron, copper, gold, silver, nickel).
   b. Main group metals (sodium, magnesium, aluminium).
5. Work in groups and fill in your comparison chart.
   a. Hardness – Compare which are harder: transition metals or main group metals.
   b. Reactivity – Watch a video/demo of metals with water or dilute acids. Record which type reacts faster.

c. Charges (Oxidation States) – Research examples of metals with variable charges (e.g., Fe^{2+}, Fe^{3+}) and metals with only one charge (e.g., Na^{+}).

d. Complex Ion Formation – Watch a demo/video of copper with ammonia (deep blue solution). Note which type forms complexes more easily.

e. Coloured Compounds – Observe salts dissolved in water ($CuSO_4$ = blue, Ni^{2+} = green, Fe^{3+} = yellow brown). Record colours.

f. Catalysts – Research or watch a video of MnO_2 breaking down H_2O_2. Record which metals are used as catalysts.

Comparison Chart: Transition vs Main Group Metals

Property	Transition Metals (Examples & Observations)	Main Group Metals (Examples & Observations)	Conclusion (Which type shows this property more?)
Hardness			
Reactivity (with water/acid)			
Charges (Oxidation States)			
Complex Ion Formation			
Coloured Compounds			
Catalysts			

6. With your group, create a **poster** showing uses in:

a. Construction (iron, steel).

b. Technology (copper, silver, gold in electronics).

c. Medicine (cobalt in vitamin B_{12}, platinum drugs).

d. Transportation (steel for cars, trains, ships).

e. Household items (coins, jewellery, cookware).

7. Share your work with the class.

REVIEW QUESTIONS

1. **a.** State two physical properties of transition elements.

b. State two chemical properties of transition elements.

c. Give one reason why transition metals are often used in construction.

2. **a.** Describe briefly why compounds of transition metals are often coloured.

b. Explain why many transition elements show variable oxidation states while main group metals do not.

c. Give two examples of transition elements used as catalysts and the processes in which they are applied.

3. You are provided with $CuSO_4$ solution, ammonia solution, MnO_2, and H_2O_2.

a. Design an experiment to show that:

i. Transition metal ions can form coloured solutions.

ii. Transition metals or their compounds act as catalysts.

b. Write down the expected observations and chemical equations for the reactions in (a.) above

4. Transition metals play important roles in science and society.

a. Discuss how electronic configuration is linked to:

i. Variable oxidation states.

ii. Formation of coloured compounds.

b. Compare the properties of transition metals with main group metals under:

- Oxidation states
- Compound colour
- Catalytic properties

c. Explain with examples three ways transition metals are applied in everyday life in Ghana.

5. **a.** State two similarities between transition metals and main group metals.

b. State two differences between transition metals and main group metals.

c. Give two everyday uses of transition metals.

6. **a.** Using examples, explain why transition metals often form coloured compounds while main group metals do not.

b. Why are transition metals generally less reactive with water and acids than main group metals?

c. Outline three specific everyday uses of transition metals in construction, technology, and medicine.

7. **a.** Complete the table with suitable entries.

b. Explain how these properties make transition metals more useful in industry than main group metals.

c. Suggest why aluminium, a main group metal, is still widely used despite these differences.

8. **a.** Compare transition metals and main group metals under oxidation states, compound colours, catalytic activity, and everyday applications.

b. Discuss the economic importance of transition metals in Ghana.

c. If you were to design a new industrial catalyst, explain why you would choose a transition metal and not a main group metal.

SECTION

6

ALKANOIC ACID DERIVATIVES AND LIPIDS



CHEMISTRY OF CARBON COMPOUNDS

Organic Functional Groups

INTRODUCTION

This section introduces you to a family of organic compounds derived from carboxylic acids and their natural forms found in living organisms. You will explore alkanoic acid derivatives, including esters, which are sweet-smelling compounds used in perfumes and flavours; amides, which are present in proteins and medicines; and acyl halides, which are very reactive compounds used in synthesis. Additionally, you will study fats and oils (lipids), which are natural compounds found in food and living organisms. By the end of this section, you will understand the structure, preparation, and properties of alkanoic acid derivatives and their everyday uses, such as esters in food flavours and amides in fabrics and drugs. You will also learn to identify the functional groups of these derivatives, explain their synthesis, compare their reactivity, describe the structure and properties of fats and oils, and appreciate their vital roles in everyday life, including food, nutrition, medicine, cosmetics, and industry.

KEY IDEAS

- **Acyl halides** – Very reactive derivatives of carboxylic acids where the –OH group is replaced by a halogen (Cl, Br).



- **Alkanoic acids** – Organic acids containing the –COOH (carboxyl) group.



- **Amides** – Organic compounds formed when the –OH group of a carboxylic acid is replaced by $–NH_2$ or substituted amines.



- **Derivatives** – Compounds obtained from another compound by replacing part of its structure.

- **Esters** – Sweet-smelling organic compounds formed from a carboxylic acid and an alcohol.



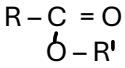
ALKANOIC ACID DERIVATIVES

Alkyl Alkanoates (Esters)

Esters are organic compounds formed when an alcohol (alkanol) reacts with a carboxylic acid (alkanoic acid).

In this reaction, the –H from the acid is replaced by an alkyl group (from the alcohol).

General form: RCOOR′



R = hydrogen or an alkyl group

R′ = an alkyl group (from the alkanol)

Structure and Functional Group

Esters contain the **–COO– group** (called the ester group). It has: **C=O** (carbonyl group) and **C–O** single bond attached to the same carbon. The alkyl groups (R and R′) may be the same or different.

Formula

General formula: $\mathbf{C_nH_{2n}O_2}$ (where n = 2, 3, 4, 5, …). The first ester (n = 2) has formula $\mathbf{C_2H_4O_2}$.

ESTERIFICATION

Alkanoic acid reacts directly with the alkanol in an acid-catalysed reversible reaction process.

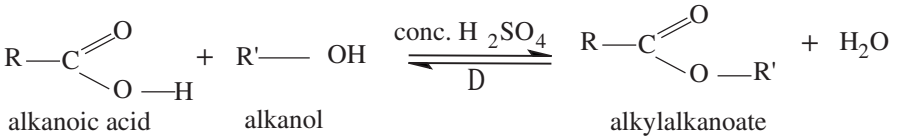
Reagents: Alkanoic acid and alkanol

Catalyst: H_2SO_4 or dry HCl

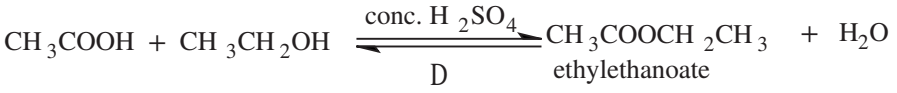
Condition: heat

Formation Reaction

Alcohol + Carboxylic acid → Ester + Water



Examples



Methyl butanoate (apple): $\underset{\text{Methanol}}{CH_3OH} + \underset{\text{butanoic acid}}{C_3H_7COOH} \rightarrow \underset{\text{methyl butanoate}}{C_3H_7COOCH_3} + \underset{\text{water}}{H_2O}$

Ethyl methanoate (rum essence): $\underset{\text{ethanol}}{C_2H_5OH} + \underset{\text{methanoic acid}}{HCOOH} \rightarrow \underset{\text{ethyl methanoate}}{HCOOC_2H_5} + \underset{\text{water}}{H_2O}$

Ethyl butanoate (pineapple): $\underset{\text{ethanol}}{C_2H_5OH} + \underset{\text{butanoic acid}}{C_3H_7COOH} \rightarrow \underset{\text{ethyl butanoate}}{C_3H_7COOC_2H_5} + \underset{\text{water}}{H_2O}$

Everyday Relevance

Esters often smell sweet and fruity, so they are used in perfumes and flavourings.

The esterification reaction is analogous to the neutralisation reaction between a mineral acid and a base, which produces a salt and water, but with important differences.

Table 6.1: Everyday relevance and use of esters

Esterification reaction	Neutralisation reaction
An alkanoic acid reacts with an alkanol	An acid reacts with a base
Products are alkyl alkanoate and water	Products are salt and water
Does not require an indicator	Requires an indicator to show end-point
Reaction is reversible	Reaction is irreversible
It is acid-catalysed	It requires no catalyst

Esterification reaction	Neutralisation reaction
Large molecules are reacting, hence very slow	Ions are reacting in solution, hence very fast

Sources of Alkyl Alkanoates

Natural sources

1. Fruits & flowers (aroma compounds): Tiny amounts of esters give many smells/flavours, e.g.,
   a. Isoamyl acetate → banana
   b. Ethyl butanoate / ethyl hexanoate → apple, pineapple
   c. Methyl salicylate → wintergreen/mint
2. Waxes in nature (long-chain alkyl alkanoates): Plant leaf waxes, beeswax, jojoba oil, and some animal waxes (e.g., cetyl palmitate in spermaceti) are mostly long-chain fatty acid + long-chain alcohol esters (R–COO–R′).
3. Fermented foods & drinks: Yeast makes esters such as ethyl ethanoate (ethyl acetate) and isoamyl acetate, contributing fruity notes in bread, beer, wine.
4. Natural signals: Some insect pheromones are acetate esters (help insects communicate).

Human-made (industrial & lab) sources

1. Solvent manufacture: Paints, inks, and glues use ethyl acetate and butyl acetate, produced by reacting the acid with the alcohol (see methods below).
2. Flavour & fragrance industry: "Fruit" esters (banana, pear, pineapple notes) are often synthesised to standard purity for food flavourings and perfumes.
3. Biodiesel (FAME): Fatty Acid Methyl Esters are made by transesterifying vegetable oils or animal fats with methanol (e.g., from palm, soybean, or used cooking oil).
4. Plastics & materials: Many coatings and polymers are esters (e.g., polyacrylates). (Note: big plastics like PET are esters too, though not simple "alkyl alkanoates".)

IUPAC Naming of Alkyl Alkanoates (Esters)

General Rule

Esters have the form RCOOR′: R comes from the carboxylic acid and R′ comes from the alcohol.

Name = Alkyl (from alcohol) + Alkanoate (from acid)

Steps to Name an Ester

1. Identify the alcohol part (R′)
   a. This gives the first word in the name.
   b. Write it as an alkyl group (methyl, ethyl, propyl, butyl…).
2. Identify the acid part (RCOO)
   a. This gives the second word.
   b. Change the "-oic acid" ending to "-oate" (methanoic acid → methanoate, ethanoic acid → ethanoate).
3. Combine them
   Name = [Alkyl group from alcohol] + [Alkanoate from acid].

Examples

1. CH_3COOCH_3 → Methyl ethanoate
   Alcohol part: CH_3OH → methyl
   Acid part: CH_3COOH → ethanoate
2. $CH_3COOCH_2CH_3$ → Ethyl ethanoate
   Alcohol part: ethanol → ethyl
   Acid part: ethanoic acid → ethanoate
3. $C_2H_5COOCH_3$ → Methyl propanoate
   Alcohol part: methanol → methyl
   Acid part: propanoic acid → propanoate
4. $C_3H_7COOCH_2CH_3$ → Ethyl butanoate
   Alcohol part: ethanol → ethyl
   Acid part: butanoic acid → butanoate

Preparation of Alkyl Alkanoates (Esters)

1. **Esterification (Fischer Method)**
   An acid (alkanoic acid) reacts with an alcohol (alkanol) in the presence of a strong acid catalyst (conc. H_2SO_4 or HCl). The reaction is slow at room temperature, so it needs heating. It is a reversible (equilibrium) reaction.
   Equation: $RCOOH + R'OH \rightleftharpoons RCOOR' + H_2O$
   Example: $CH_3COOH + CH_3CH_2OH \rightleftharpoons CH_3COOCH_2CH_3 + H_2O$

(Ethanoic acid + Ethanol ⇌ Ethyl ethanoate + Water)

2. **Reaction of Acyl Halides with Alcohols**

   Acyl halides (like acid chlorides, RCOCl) are **very reactive**. They react quickly with alcohols at **room temperature** to form esters. Hydrogen halide (HCl or HBr) is also produced.

   Equation: $RCOCl + R'OH \rightarrow RCOOR' + HCl$

   Example: $CH_3COCl + CH_3CH_2OH \rightarrow CH_3COOCH_2CH_3 + HCl$

   (Acetyl chloride + Ethanol → Ethyl ethanoate + Hydrogen chloride)

3. **Reaction of Acid Anhydrides with Alcohols**

   Acid anhydrides (RCO–O–COR) are **less reactive than acyl halides**, but more reactive than acids. They react with alcohols to form an **ester and a carboxylic acid**. No catalyst is needed.

   Equation: $(RCO)_2O + R'OH \rightarrow RCOOR' + RCOOH$

   Example: $(CH_3CO)_2O + CH_3CH_2OH \rightarrow CH_3COOCH_2CH_3 + CH_3COOH$

   (Ethanoic anhydride + Ethanol → Ethyl ethanoate + Ethanoic acid)

4. **Transesterification**

   One ester reacts with a **different alcohol**. The alkoxy group (–OR) is **exchanged**. Needs an **acid or base catalyst**.

   Equation: $RCOOR' + R''OH \rightleftharpoons RCOOR'' + R'OH$

   Example: $CH_3COOCH_3 + CH_3CH_2OH \rightleftharpoons CH_3COOCH_2CH_3 + CH_3OH$

   (Methyl ethanoate + Ethanol ⇌ Ethyl ethanoate + Methanol)

Physical Properties of Alkyl Alkanoates (Esters)

1. State and Appearance: Most low-molecular mass esters are colourless liquids with a sweet, fruity smell. Higher esters (larger molecules) can be waxy solids.
2. Odour: Many esters have pleasant fruity odours (e.g., banana, pineapple, apple). Used in flavourings and perfumes.
3. Boiling Points: Esters have moderate boiling points. Usually lower than carboxylic acids of similar molecular mass (because esters do not form hydrogen bonds as strongly as acids).
4. Solubility in Water: Small esters are slightly soluble in water (due to some hydrogen bonding with water). Larger esters are insoluble in water (non-polar parts dominate). Esters are soluble in organic solvents (e.g., ethanol, ether, chloroform).

5. Density: Esters are usually less dense than water (they float on water).
6. Volatility: Small esters are volatile (evaporate easily), which is why their smell is easily noticed.

Chemical Reactions of Alkyl Alkanoates (Esters)

1. Hydrolysis (Breaking Down with Water): Esters can be broken back into acid + alcohol.
   a. Acidic hydrolysis (using dilute HCl or H_2SO_4, with heat):

      $RCOOR' + H_2O \rightleftharpoons RCOOH + R'OH$ (Reversible reaction)

   b. Alkaline hydrolysis (saponification) (using NaOH):

      $RCOOR' + NaOH \rightarrow RCOONa + R'OH$ (Forms salt of the acid + alcohol, irreversible)

2. Reaction with Ammonia: Esters react with ammonia (NH_3) to give amides and alcohols.

   Equation: $RCOOR' + NH_3 \rightarrow RCONH_2 + R'OH$

3. Reaction with Alcohols (Transesterification): One ester reacts with a different alcohol to form a new ester and an alcohol. Needs an acid or base catalyst.

   Equation: $RCOOR' + R''OH \rightleftharpoons RCOOR'' + R'OH$

4. Reduction of Esters: Esters can be reduced (with strong reducing agents like $LiAlH_4$) to form primary alcohols.

   Equation: $RCOOR' \rightarrow RCH_2OH + R'OH$

5. Combustion: Like other organic compounds, esters burn in oxygen to form carbon dioxide and water.

   Equation: $RCOOR' + O_2 \rightarrow CO_2 + H_2O$

Amides

Structure and Bonding of Amides

Amides have the functional group $–CONH_2$. They come from acids called alkanoic acids.

Amides can be classified as

1. Primary: $R–CONH_2$ (one group attached to nitrogen)
2. Secondary: $R–CONHR'$ (two groups attached to nitrogen)

3. Tertiary: R–CONR′R″ (three groups attached to nitrogen).

The atoms in an amide lie in the same flat plane. This happens because the nitrogen shares its lone pair with the C=O group (this sharing is called resonance). Resonance makes amides more stable than expected. In amides, the nitrogen uses sp^2 orbitals (flat shape), while in amines (NH_3 family) nitrogen uses sp^3 orbitals (pyramid shape).

Nomenclature of Amides (How to Name Them)

1. Start with the parent acid (alkanoic acid).
   Example: ethanoic acid (CH_3COOH).
2. Replace the "–oic acid" ending with "–amide."
   Ethanoic acid → ethanamide (CH_3CONH_2).
   Methanoic acid → methanamide ($HCONH_2$).
3. For secondary and tertiary amides (when there are substituted groups on the nitrogen):
   Show these groups using N- to indicate their position in the name.
   Example: $CH_3CONHCH_3$ → N-methyl ethanamide.
   Example: $CH_3CON(CH_3)_2$ → N,N-dimethyl ethanamide.

Preparation Methods of Amides

1. From Alkanoic Acids (via Ammonium Salts)
   Step 1: Carboxylic acid reacts with ammonia to form ammonium carboxylate.
   $RCOOH + NH_3 \rightarrow RCOONH_4$
   Step 2: On heating, the ammonium salt dehydrates to form the amide.
   $RCOONH_4 \rightarrow RCONH_2 + H_2O$
2. From Acid Chlorides (Acyl Chlorides)
   Acyl chloride reacts readily with ammonia.
   $RCOCl + NH_3 \rightarrow RCONH_2 + HCl$
3. From Acid Anhydrides
   Acid anhydride reacts with ammonia to give amides.
   $(RCO)_2O + NH_3 \rightarrow RCONH_2 + RCOOH$
4. From Esters
   Ester reacts with ammonia (ammonolysis) to form an amide and alcohol.

$RCOOR' + NH_3 \rightarrow RCONH_2 + R'OH$

5. From Nitriles (Hydrolysis)

   Partial hydrolysis of nitriles under controlled conditions forms amides.

   $RCN + H_2O \rightarrow RCONH_2$

Physical Properties of Amides

1. **State and Appearance**

   Lower aliphatic amides (e.g., methanamide, ethanamide) are colourless liquids. Higher members are crystalline solids with high melting points.

2. **Intermolecular Forces**

   Amides form extensive hydrogen bonding between the $–NH_2$ group and the C=O group. This strong hydrogen bonding is responsible for their high melting and boiling points, higher than comparable hydrocarbons, esters, or ketones.

3. **Solubility**

   Low molecular mass amides (up to about 5 carbons) are soluble in water because they can form hydrogen bonds with water molecules. Higher amides become less soluble due to the increasing hydrophobic alkyl chain.

4. **Boiling and Melting Points**

   a. Primary amides ($R–CONH_2$): Highest boiling points due to two –NH bonds allowing strong hydrogen bonding.

   b. Secondary amides (R–CONHR'): Lower boiling points than primary, since only one –NH group is available.

   c. Tertiary amides (R–CONR'R"): Lowest boiling points because no –NH group is available for hydrogen bonding, though the carbonyl oxygen can still form weak bonds.

5. **Odour and Stability**

   Lower amides are generally odourless and relatively stable compared to amines. Their planar structure due to resonance makes them less reactive towards bases and nucleophiles.

Chemical Properties of Amides

1. **Hydrolysis**

   Amides undergo hydrolysis when heated with acids or bases.

Acidic Hydrolysis: $RCONH_2 + H_2O + HCl \rightarrow RCOOH + NH_4Cl$ or

$$R{-}CONH_2 + H_2O \xrightarrow[\text{heat}]{\text{dil. HCl}} R{-}COOH + NH_4Cl$$

Produces a carboxylic acid and ammonium salt.

Basic Hydrolysis: $RCONH_2 + NaOH \rightarrow RCOONa + NH_3$ or

$$R{-}CONH_2 + H_2O \xrightarrow[\text{heat}]{\text{dil. NaOH}} R{-}COONa + NH_3$$

Produces a carboxylate salt and ammonia.

2. **Dehydration to Nitriles**

   On heating with phosphorus(V) oxide (P_2O_5) or thionyl chloride ($SOCl_2$), amides lose water to form nitriles.

   $RCONH_2 \rightarrow RC{\equiv}N + H_2O$

3. **Reaction with Nitrous Acid**

   Primary amides react with nitrous acid (HNO_2) to release nitrogen gas.

   $RCONH_2 + HNO_2 \rightarrow RCOOH + N_2 + H_2O$

4. **Reduction to Amines**

   Amides can be reduced to amines using lithium aluminium hydride ($LiAlH_4$).

   $RCONH_2 + 4[H]^+ \rightarrow RCH_2NH_2$

5. **Hoffmann Degradation**

   Heating an amide with bromine and aqueous alkali (NaOH or KOH) gives a primary amine with one carbon atom less.

   $RCONH_2 + Br_2 + 4NaOH \rightarrow RNH_2 + Na_2CO_3 + 2NaBr + 2H_2O$

6. **Resonance and Basicity**

   Due to resonance between the nitrogen lone pair and the carbonyl group, amides are neutral compounds (very weak bases compared to amines). They resist protonation under normal conditions.

Applications of Amides

1. **Pharmaceuticals**

   Many medicines are amide derivatives due to their stability and biological activity.

Examples:

a. Paracetamol (acetaminophen): Pain reliever and fever reducer.

b. Lidocaine: A local anaesthetic.

c. Penicillin derivatives: Contain amide linkages crucial for antibacterial activity.

2. **Polymers and Synthetic Materials**

a. Nylons (polyamides): Produced from diamines and dicarboxylic acids, used in textiles, carpets, and ropes.

b. Kevlar: A high-strength aromatic polyamide used in bulletproof vests, helmets, and aerospace materials.

c. Proteins: Naturally occurring polyamides formed by peptide (amide) bonds between amino acids.

3. **Agriculture**

Amide-based compounds serve as herbicides, pesticides, and fungicides.

Example: Diuron (a urea-derived herbicide).

4. **Industrial Uses**

Amides are used as solvents and stabilisers in chemical industries. They act as intermediates in the preparation of dyes, rubber chemicals, and resins. Formamide is a good solvent and plasticizer.

5. **Everyday Products**

Amides are found in

a. Softeners and detergents (as stabilising agents).

b. Cosmetics and shampoos (conditioning agents).

c. Food industry: Certain amides are used as additives and flavouring agents.

6. **Biological Importance**

Peptide bonds (amide linkages) form the backbone of proteins, essential for life. Urea (a simple diamide) is the main nitrogen excretion product in mammals.

ACYL HALIDES

Structure and Bonding of Acyl Halides

Acyl halides are organic compounds. They have the formula **R–COX**, where:

R is a carbon group (like a chain or ring).

CO is the carbonyl group (a carbon joined to oxygen with a double bond).

X is a halogen atom (like chlorine, bromine, fluorine, or iodine).

You can think of acyl halides as coming from **carboxylic acids**. In carboxylic acids, there is an –OH group (like in vinegar). In acyl halides, this –OH is replaced by a halogen atom. The carbon in the **C=O group** is flat (planar) because of its bonding. The bonds around it spread out at about **120°**, like the corners of a triangle.

Nomenclature of Acyl Halides

1. Identify the Parent Carboxylic Acid
   a. Look for the carboxylic acid from which the acyl halide is made.
   b. Find the carbon chain or ring attached to the –COX group.
   c. Count the number of carbon atoms in this chain or identify the ring system.
2. Replace the Ending of the Parent Acid
   Change the ending "–oic acid" to "–oyl halide."
   Example: Ethanoic acid → Ethanoyl chloride
3. Specify the Halogen
   a. Write the name of the halogen present in the acyl halide.
   b. Options: chloride (Cl), bromide (Br), fluoride (F), or iodide (I).

   Examples: Ethanoyl chloride (CH_3COCl), Propanoyl bromide (CH_3CH_2COBr)
4. Add Prefixes for Substituents (if any)
   a. Start numbering the chain from the carbonyl carbon as position 1.
   b. Name and number any groups (like $–CH_3$, –Cl, $–NO_2$) as prefixes.

   Example: 2-methylpropanoyl chloride

Examples of IUPAC Names of some Acyl Halides

1. CH_3COCl: ethanoyl chloride (from ethanoic acid)
2. CH_3CH_2COCl: propanoyl chloride (from propanoic acid)
3. CH_3CH_2COBr: propanoyl bromide (from propanoic acid)
4. CH_2=CHCOCl: propenoyl chloride (from propenoic acid)
5. $CH_3CH_2CH_2CH_2COCl$: pentanoyl chloride (from pentanoic acid)
6. C_6H_5COCl: benzoyl chloride (from benzoic acid)
7. $ClCH_2COCl$: chloroethanoyl chloride (from chloroethanoic acid)

Examples of Common Names of some Acyl Halides

1. CH_3COCl: acetyl chloride (from acetic acid)
2. HCOCl: formyl chloride (from formic acid, though unstable)
3. CH_3CH_2COCl: propionyl chloride (from propionic acid)
4. C_6H_5COCl: benzoyl chloride (from benzoic acid)
5. $CH_3(CH_2)_{16}COCl$: stearoyl chloride (from stearic acid)
6. CH_2=CHCOCl: acryloyl chloride (from acrylic acid)

Preparation of Acyl Halides

Acyl halides are prepared mainly from carboxylic acids by replacing the –OH group with a halogen atom.

1. **From Carboxylic Acids with Halogenating Agents**
   a. With Thionyl chloride ($SOCl_2$): R–COOH + $SOCl_2$ → R–COCl + $SO_{2(g)}$ + $HCl_{(g)}$

      (Most common laboratory method; gaseous by-products make purification easy.)

   b. With Phosphorus trichloride (PCl_3): 3R–COOH + PCl_3 → 3R–COCl + H_3PO_3

   c. With Phosphorus pentachloride (PCl_5): R–COOH + PCl_5 → R–COCl + $POCl_3$ + HCl

2. **Interconversion of Acyl Halides**

   Acid bromides and iodides can be prepared from acid chlorides using NaBr or NaI with a catalyst such as phosphorus:

   a. R–COCl + NaBr → R–COBr + NaCl

   b. R–COCl + NaI → R–COI + NaCl

3. **From Acid Anhydrides**

   $(R–CO)_2O + HX \rightarrow R–COX + R–COOH$

Physical Properties of Acyl Halides

1. Most acyl halides are colourless, fuming liquids with sharp, irritating odours.
2. Aromatic acyl halides such as benzoyl chloride may appear as crystalline solids.
3. They have lower boiling points than carboxylic acids, because they cannot form hydrogen bonds between their own molecules.
4. They are insoluble in water, since they hydrolyse violently, forming the parent carboxylic acid and hydrogen halide (HCl, HBr, etc.).
5. They dissolve well in organic solvents like benzene, ether, and chloroform.
6. Many are denser than water due to the presence of halogen atoms.
7. In moist air, they release dense white fumes of hydrogen halides.

Chemical Properties of Acyl Halides

Acyl halides are very reactive because the carbonyl carbon is strongly electron-deficient, making it easy for nucleophiles to attack.

1. Hydrolysis (reaction with water): $R–COX + H_2O \rightarrow R–COOH + HX$

   Example: $CH_3COCl + H_2O \rightarrow CH_3COOH + HCl$

2. Reaction with Alcohols (Esterification): $R–COX + R'–OH \rightarrow R–COOR' + HX$

   Example: $CH_3COCl + CH_3OH \rightarrow CH_3COOCH_3 + HCl$

3. Reaction with Ammonia: $R–COX + NH_3 \rightarrow R–CONH_2 + HX$

   Example: $CH_3COCl + NH_3 \rightarrow CH_3CONH_2 + HCl$

4. Reaction with Amines: $R–COX + R'–NH_2 \rightarrow R–CONHR' + HX$
5. Friedel–Crafts Acylation (with Aromatic Compounds):

   $R–COCl + C_6H_6 \rightarrow R–CO–C_6H_5 + HCl$ (in the presence of $AlCl_3$ catalyst)

Uses of Acyl Halides

1. Preparation of Esters: Acyl halides react with alcohols to form esters, used in perfumes, flavourings, and solvents.

   $R–COCl + R'–OH \rightarrow R–COOR' + HCl$

2. Preparation of Amides: With ammonia or amines, acyl halides form amides, useful in making plastics, fibres, and medicines.

   $R–COCl + NH_3 \rightarrow R–CONH_2 + HCl$

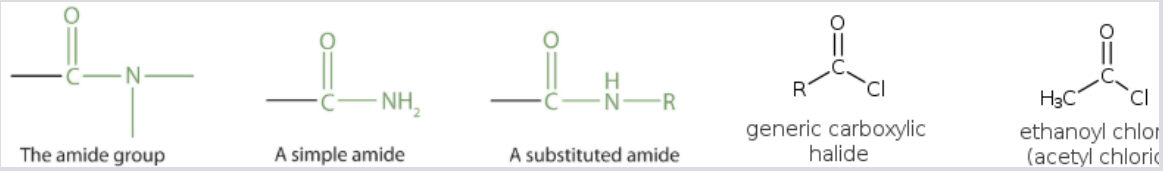
3. Friedel–Crafts Acylation: In the presence of $AlCl_3$ catalyst, acyl halides react with aromatic compounds to form aromatic ketones.

   $R–COCl + C_6H_6 \rightarrow R–CO–C_6H_5 + HCl$ (Important in making dyes, perfumes, and pharmaceuticals.)

4. Intermediates in Organic Synthesis: Acyl halides serve as starting materials in making anilides, acetyl derivatives, agrochemicals, and pharmaceuticals.

Activity 6.1 Amides and Acyl Halides

1. Observe the representations of amides and acyl halides below



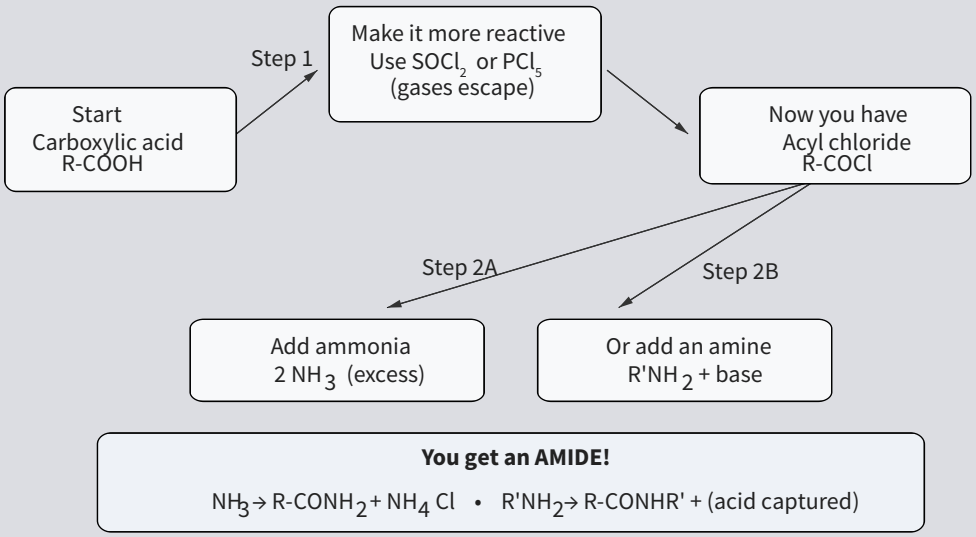
2. Study the flow charts below on the preparation of Amides and Acyl halides

   a. Step-by-Step Preparation of Amides Flow Diagrams

   Two simple routes: make a reactive helper first (acyl chloride) or start from an anhydride.

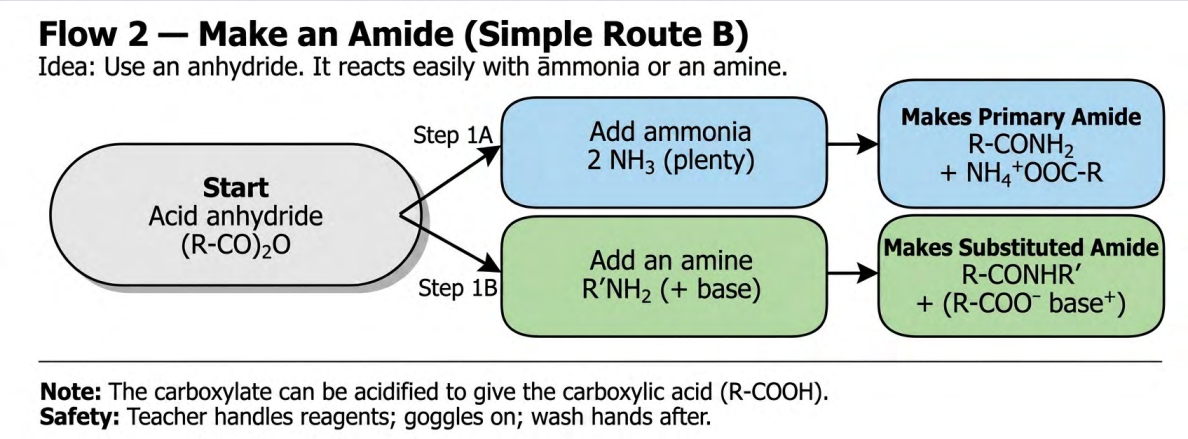
Flow 1 — Make an Amide (Simple Route A)

Idea: Turn the acid into a more reactive form, then add ammonia/amine.



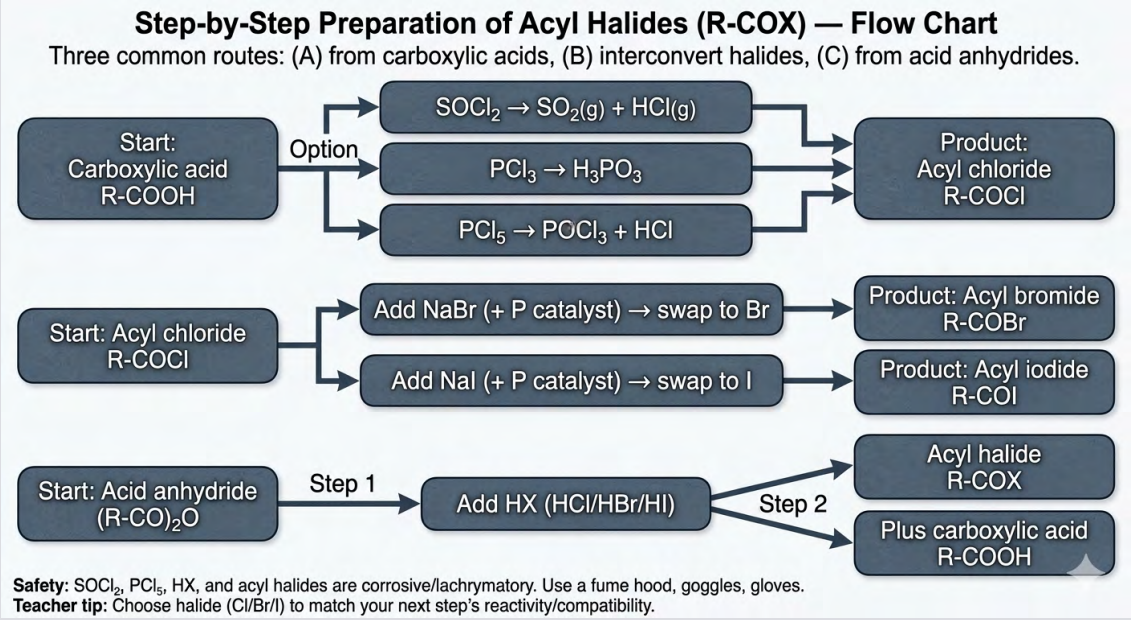
Safety: These steps give off acidic fumes (HCl, SO_2). Teacher handles chemicals; use goggles & good ventilation.

Flow 1: Carboxylic acid → acyl chloride → amide (add NH_3 or $R'NH_2$ + base).



Flow 2: Acid anhydride → amide (direct with NH_3 or amine + base).

a. Step-by-Step Preparation of Acyl Halides — Flow Chart



3. Draw the structures of

a. ethanamide (*acetamide*) — CH_3CONH_2, Benzamide — $C_6H_5CONH_2$

b. Ethanoyl chloride (*acetyl chloride*) — CH_3COCl and Benzoyl chloride — C_6H_5COCl

4. Fill-in-the-Blank Worksheet: Write the correct words or formulas in the blanks. Use standard organic notation (R, R', X).

a. General Formulas

Compound class	General formula (fill in the blank)
Primary amide	R–C O–_____ (write the group)
Secondary amide	R–CO–_____ (NHR′)
Tertiary amide	R–CO–_____ (NR′R″)
Acyl halide (general)	R–CO_____ (X = Cl, Br, I, F)
Acyl chloride	R–CO_____
Acyl bromide	R–CO_____
Benzamide (specific)	C_6H_5–_____

b. For acyl halides: change "–oic acid" to "–_____ _____". (write the two words)

c. For amides: change "–oic acid" to "–_____".

5. Convert each carboxylic acid below to the named derivative. Write the CORRECT derivative name.

a. Methanoic acid → (acyl chloride) _____

b. Ethanoic acid → (amide) _____

c. Propanoic acid → (acyl bromide) _____

d. Benzoic acid → (amide) _____

e. Butanoic acid → (acyl iodide) _____

f. 2-Methylpropanoic acid → (acyl chloride) _____

6. Fill in the missing reactant(s) or product(s). Use HCl/HBr etc. where appropriate.

a. $R–COCl + H_2O \rightarrow$ _____ + _____

b. $R–COCl + CH_3CH_2OH \rightarrow$ _____ + _____

c. $R–COCl + NH_3 \rightarrow$ _____ + _____

d. $R–COCl + CH_3NH_2 \rightarrow$ _____ + _____

e. $R–COOH + PCl_3 \rightarrow$ _____ + _____

f. $R–COOH + SOCl_2 \rightarrow$ _____ + _____ + _____

g. $R–COCl + NaBr \rightarrow$ _____ + _____

h. $C_6H_6 + R–COCl\ (AlCl_3) \rightarrow$ _____ + _____

i. $C_6H_5COCl + NH_3 \rightarrow$ _____ + _____

j. $R–CONH_2$ (amide) + H_2O (acid) → _____ + _____

Activity 6.2 Structures, Names & Reactions of Amides vs Acyl Halides

1. Draw structural formulae for the following five amides: and five acyl halides:

methanamide, ethanamide, propanamide, Butanamide, benzamide, N-methylacetamide, N, N-dimethylformamide.

2. Draw structural formulae for the following five acyl halides ethanoyl chloride, propanoyl bromide, benzoyl chloride, butanoyl chloride, ethanoyl fluoride, Benzoyl chloride.

3. Assign IUPAC names to the structures from questions 1 and 2; exchange sheets with another group for checking.

4. Create a comparison chart (structure & properties differences). Work in pairs or groups. Use your notes/handouts to complete the table.

Feature	Amides	Acyl halides
Functional group (draw/describe)		
General formula		
Element highlight (N vs. halogen)		
Polarity / resonance (brief)		
Hydrogen bonding (self / with water)		
Typical boiling point (relative tendency)		
Water: solubility vs. reaction		
Odour / appearance (simple)		
Common lab preparation (from acids)		
Key reactivity with NH_3 / amines / alcohols		
Example compound (name + formula)		
Safety note (one line)		

5. Record formation reactions from carboxylic acids with reagents/conditions.
6. Fill in the missing reagents/products/conditions. Keep conditions brief (e.g., 'dry solvent', 'cooling', 'reflux').

   a. R–COOH + _____ → R–COCl + _____ + _____ (condition: _____)

   b. 3 R–COOH + _____ → 3 R–COCl + _____ (condition: _____)

   c. R–COOH + ____ → R–COCl + _____ + _____ (condition: _____)

   d. 3 R–COOH + _____ → 3 R–COBr + _____ (condition: _____)

   e. R–COCl + _____ → R–COBr + _____ (condition: _____)

   f. R–COCl + _____ → R–COI + _____ (condition: _____)

   g. R–COCl + 2 _____ → R–$CONH_2$ + _____ (condition: _____)

   h. R–COCl + _____ → R–CONHR′ + _____ (condition: _____)

   i. R–COCl + _____ → R–COOR′ + _____ (condition: _____)

   j. R–COOH + _____ ⇌ R–$COO^-NH_4^+$ → (_____) → R–$CONH_2$ + _____

7. Present one key reaction to your colleagues.

Activity 6.3 Reactions of Amides and Acyl Halides

Formation Reactions — from Carboxylic Acids. Write word and generic symbol equations with reagents and short conditions. Add one safety note.

Route	Word equation	Symbol equation	Reagents/ conditions	Safety note
Acid → Acyl chloride				
Acyl chloride → Amide (NH_3)				
Acyl chloride → N-substituted amide				
Anhydride → Amide				

Activity 6.4 Making Ethyl Ethanoate

Safety rules to observe

1. Wear goggles and gloves; tie back hair; no tasting, only waft smells.
2. Handle concentrated sulphuric acid (H_2SO_4) and hot water with great care.
3. Use a warm water bath (50–60 °C) — no flames.
4. Neutralise spills with sodium bicarbonate, wash hands after.
5. Identify, and collect materials & Equipment to follow the procedure safely
6. **Run the Experiment (under supervision)**
7. Follow the Basic Procedure step by step.

Basic microscale procedure (per group)

1. In a test tube: add 2 ml ethanol + 2 ml ethanoic acid.
2. Add about 5 drops conc. H_2SO_4 (catalyst). Stopper and swirl.
3. Place tube in a 50–60 °C water bath for 10 min, swirling occasionally.
4. Pour into a second tube with 5 ml sodium carbonate solution (~5%). Swirl and vent (CO_2). Check pH with paper; add a little more carbonate if still acidic.
5. Add 5 ml brine (salt water), swirl, and let layers separate.
6. The top layer is ethyl ethanoate. Pipette it into a clean, dry tube.

Observation and calculation

1. Record every observation (colour, bubbles, temperature, smell by wafting, when layers appear)
2. Measure your ester volume (ml).
3. Calculate Percentage Yield using:

$$Percent\ yield = \frac{your\ ester\ volume}{3.3\ ml} \times 100\%$$

Note

(3.3 ml is a classroom "ideal max" for 2 mL ethanol + 2 mL acid. Real student yields are often 20–60%.)

Activity 6.5 Reactions of Amides and Acyl Halides

1. Collect your prompt cards with different reagents (such as water, alcohols, or ammonia) from your teacher.
2. Discuss with someone how amides and acyl halides would react with the reagent on your card. (Use words like: *"fast," "slow," "gives off gas," "forms a product.)*
3. Which reacts faster – amides or acyl halides? Why do you think so?
4. Talk with your partner about where these reactions are useful in everyday life.
5. Join another pair to make a group of four. Share your ideas and listen to theirs.
6. Add new points you had not thought of before.
7. Listen carefully and ask short questions if something is not clear.

Activity 6.6 Presenting Reactions

In small group with your classmates discuss and present on the following.

1. Reaction Mechanisms
   - **a.** Define reaction mechanism

     Key components: reactants, intermediates, transition states and products
   - **b.** State the types of reaction mechanisms:
2. Reduction Reactions
   - **a.** Define reduction reactions
   - **b.** State the types of reduction reactions
3. Chemical Testing Methods
4. Presentation Guidelines
   - **a.** Clear and concise language
   - **b.** Visual aids (diagrams, charts, spectra)
   - **c.** Examples and illustrations
   - **d.** Group participation and discussion
5. Present your work to the whole class

FATS AND OILS

Structure, Properties and Uses of Fats and Oils

Fats and Oils

Fats and oils are part of the lipid family. They are made mostly of carbon and hydrogen, with little oxygen. They do not mix with water (they are hydrophobic). They give us energy and help build parts of living things.

Structure of Fats and Oils

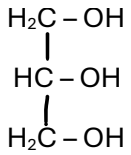
Triglycerides: The Units of Fats and Oils

Fats and oils are made up of triglycerides. A triglyceride = 1 glycerol + 3 fatty acids.

Glycerol

Has 3 carbon atoms in a row. Each carbon has a –OH (hydroxyl) group. Because it has 3 –OH groups, glycerol is called a triol.

Formula: $C_3H_8O_3$.



Glycerol

Fatty acids

Long chains of carbon and hydrogen with a –COOH (carboxyl) group at one end.

Formula: RCOOH (R = long chain).

Chain length usually has 12–20 carbons.

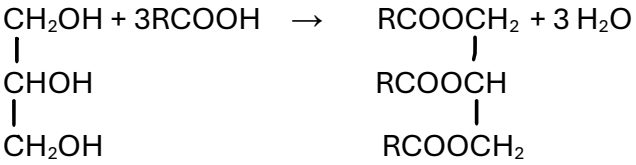
Types of fatty acids

1. Saturated: Only single bonds (C–C). Chains are straight and pack closely. Usually solid at room temperature.
2. Unsaturated: Have one or more double bonds (C=C). Double bonds create bends ("kinks"). Cannot pack tightly. Usually liquid at room temperature.

Reaction to form triglycerides

$C_3H_8O_3 + 3RCOOH \rightarrow (RCOO)_3C_3H_5 + 3H_2O$

Formation of Triglycerides



Glycerol + Fatty Acid → Triglyceride + water

This diagram shows how triglycerides (the building blocks of fats and oils) are formed

Sources of Fats and Oils

Fats and oils come from both animal and plant sources. Animal sources usually give fats (solid at room temperature), while plant sources usually give oils (liquid at room temperature).

Table 6.2: Some sources of fats and oils

Animal Sources (Fats)	Plant Sources (Oils)
Butter (milk/cream)	Olive oil (olives)
Lard (pigs)	Groundnut/Peanut oil
Tallow (cows and sheep)	Palm oil (palm fruits)
Fish oils (cod liver oil)	Coconut oil
	Soybean oil
	Sunflower oil
	Canola oil

Summary

Animal = Fats (solid at room temperature, saturated fatty acids)

Plants = Oils (liquid at room temperature, unsaturated fatty acids)

Physical Properties of Fats and Oils

Fats and oils are members of the lipid family. Their physical properties help us to tell the difference between them. The **Table 6.3** below compares fats and oils:

Table 6.3: Physical properties of fats and oils

Fats	Oils
Solid at room temperature	Liquid at room temperature
Usually colourless or pale (can be white)	Usually colourless, pale yellow or reddish
Little or no odour when pure	Little or no odour when pure
Greasy taste	Greasy or mild flavour
Insoluble in water, soluble in organic solvents	Insoluble in water, soluble in organic solvents
Less dense than water (float on water)	Less dense than water (float on water)
Higher melting point (solid)	Lower melting point (liquid)
Mostly from animal sources	Mostly from plant sources

Summary

1. Fats are solid at room temperature, while oils are liquid.
2. Both are insoluble in water but soluble in organic solvents.
3. Both are less dense than water and float on it.
4. Fats mainly come from animals, while oils mainly come from plants.

Chemical Properties of Fats and Oils

1. **Hydrolysis (Breaking Down Fats and Oils)**

   Fats and oils can be broken into glycerol and fatty acids by enzymes (lipases) during digestion. In the small intestine, bile salts first break fats into small droplets (emulsification), giving enzymes more surface area. Pancreatic lipase then hydrolyses the ester bonds in triglycerides.

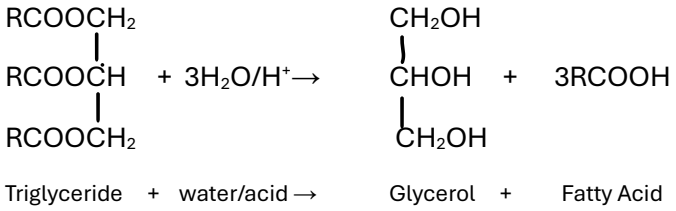
   $(RCOO)_3C_3H_5 + 3H_2O \rightarrow C_3H_8O_3 + 3RCOOH$

   Hydrolysis can also happen in the lab under acidic or basic conditions:

   a. Acidic Hydrolysis

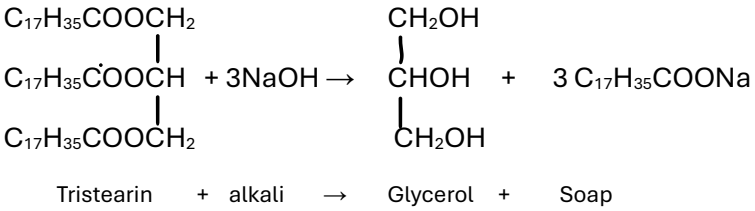
      Uses an acid catalyst (e.g., H_2SO_4). Produces glycerol and fatty acids.

      $CH_3COOCH_2CH_3 + H_2O/(H_2SO_4) \rightarrow CH_3CH_2OH + CH_3COOH$



b. Basic Hydrolysis (Saponification)

Uses a strong base (e.g., NaOH). Produces glycerol and soap (salts of fatty acids).



2. Hydrogenation (Hardening of Oils)

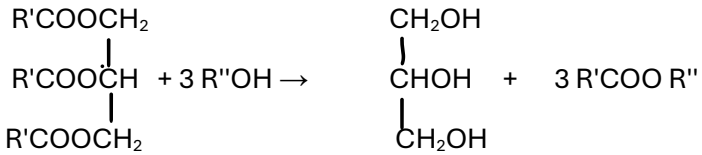
Unsaturated oils (liquid) can be changed into saturated fats (solid) by adding hydrogen across the double bonds. This process requires a catalyst such as nickel, palladium, or platinum at high temperature and pressure. It is used in making margarine.

$C_nH_m + H_2 \rightarrow (\text{Catalyst}) \rightarrow C_nH_{m+2}$

3. Transesterification

Reaction of fats/oils with alcohol to make biodiesel and glycerol. The glycerol backbone is replaced, forming new esters.

$(R'COO)_3C_3H_5 + 3R''OH \rightarrow C_3H_8O_3 + 3R'COOR''$



4. Rancidity (Spoiling of Fats and Oils)

Fats and oils can spoil when exposed to air, light, moisture, or heat. This produces unpleasant odour and taste. Oxidative rancidity occurs in unsaturated fats when oxygen attacks double bonds, forming aldehydes, ketones, and short-chain acids.

Unsaturated Fat + O_2 → Hydroperoxides → Aldehydes + Ketones + Short-chain fatty acids

Uses of Fats and Oils

Fats and oils have many uses in food, industry, medicine, and daily life. The **Table 6.4** below shows how they are applied in different areas:

Table 6.4: Uses of fats and oils

Category	Uses
Food and Nutrition	· Rich source of energy · Supply essential fatty acids · Help absorption of fat-soluble vitamins (A, D, E, K) · Add flavour and texture to food (butter, margarine, cooking oil)
Industrial Uses	· Soap making (saponification) · Cosmetics (creams, lotions, lipsticks) · Candle making · Lubricants for machines · Paints and varnishes (drying oils like linseed oil) · Biodiesel production
Medicine and Health	· Ointments and medicated creams · Fish oils (omega-3) for heart and brain health
Everyday Life	· Cooking and frying foods · Polishes (shoe and floor polish) · Provide warmth and energy storage (blubber, body fat)

Manufacture of Margarine

1. Start with vegetable oil

   Margarine is made from liquid vegetable oil (like soybean oil, sunflower oil or palm oil).

2. Clean and prepare the oil

   The oil is filtered and purified to remove any dirt or unwanted materials.

3. Hydrogenation (hardening the oil)

   A special process called hydrogenation is used.

   Hydrogen gas is bubbled through the oil with a nickel catalyst.

This changes the liquid oil into a semi-solid fat that looks like butter.

4. Blending and mixing

   The hardened oil is mixed with milk or water, salt (for taste) and vitamins (A and D)

   Sometimes colouring and flavouring to make it look and taste like butter.

5. Cooling and packaging

   The mixture is cooled until it becomes smooth and spreadable.

   Finally, it is packed into tubs or wrappers and sent to shops.

   Refer to ***Figure 6.1*** for the Flowchart diagram of Manufacture of Margarine

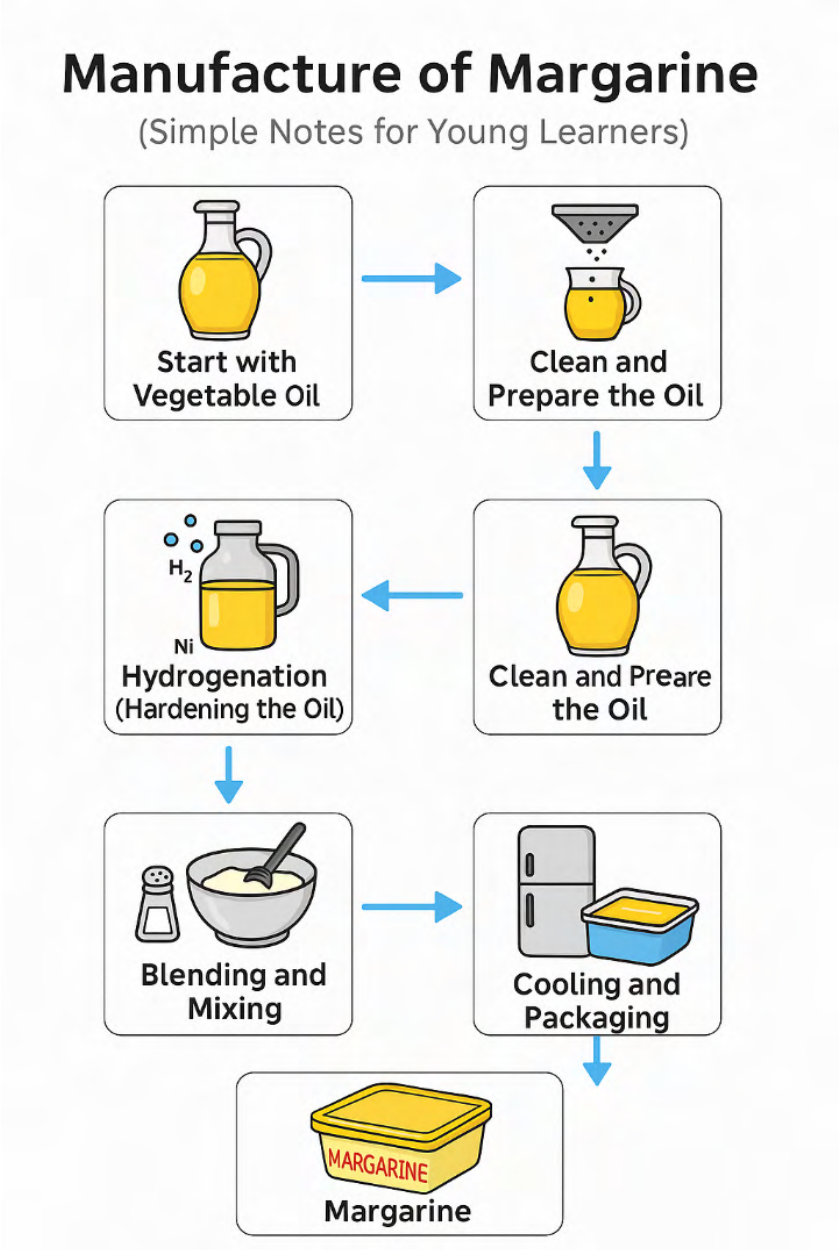


Figure 6.1: Flowchart diagram of Manufacture of Margarine

Preparation of Soap

Soaps are made when oils or fats react with alkalis (bases). This reaction is called saponification, and it also produces glycerol as a by-product. Soap can be made locally or industrially.

Local Production of Soap

Materials

1. Oil (palm oil, coconut oil, shea butter oil, or palm kernel oil)
2. Ashes (from burnt cocoa pods or plantain peels)
3. Water
4. Salt

Steps

1. Dissolve ashes in water and filter. The liquid is the alkali solution.
2. Heat the oil in a pot (don't overheat).
3. Slowly add the alkali solution to the hot oil and stir well.
4. Keep stirring for about 20–30 minutes until the mixture thickens (soap forms).
5. Add salt solution and allow to cool.
6. The soap formed is shaped into bars and packaged.

Industrial Production of Soap

Materials

1. Fats or oils (palm oil, coconut oil, olive oil)
2. Sodium hydroxide or potassium hydroxide
3. Brine (salt solution)
4. Water

Steps

1. Heat the oils in large tanks.
2. Add the alkali (base) carefully and stir. Soap and glycerol are formed.
3. Boil for 1–2 hours while stirring until the mixture thickens.
4. Add brine (salt solution) – soap floats to the top, glycerol stays below.
5. Skim off the soap, remove and purify the glycerol.
6. Wash and bleach the soap to remove colour, smell, and impurities.
7. Add fragrance, preservatives, and colouring for better quality.

8. Mould the soap into shapes and allow it to dry and harden.

Activity 6.7 Exploring Fats and Oils

Work in a group with your classmates for this activity.

1. Write the general molecular formula of a fat (triglyceride) molecule:
2. Use model kits or cutouts to build saturated and unsaturated fats.
3. List the differences in structure and physical properties

a. Structure differences

b. Physical property differences

4. List common sources of fats and oils. Classify them as animal or plant-based.
5. **(a)** Explain why fats and oils are classified as triglycerides

   (b) Write the balanced equation for the formation of a triglyceride
6. Research and presents one reaction of fats and oils
   a. Hydrolysis in alkaline/acidic conditions → formation of glycerol and soap/fatty acids.
   b. Enzymatic hydrolysis → action of lipase enzyme in digestion.
   c. Catalytic hydrogenation → conversion of oils to margarine.
   d. Rancidity → why fats spoil and how to prevent it.

Activity 6.8 Making and Understanding Soap

Work in a group with your classmates for this activity.

Materials

Palm oil or coconut oil, Wood ash (alkali) or caustic soda (NaOH),

Procedure

1. Prepare a small sample of soap by mixing the oil and alkali, following the procedure from the "local production of salt method".
2. Observe the outcomes (texture, hardness, lathering, scent).
3. Discuss and make suggestions to improve quality (e.g., adding perfume, colour, herbs, or improving smoothness).

4. Draw the general structure of a soap molecule. A soap molecule has:
   a. A long tail (hydrocarbon chain) that loves oil and grease(hydrophobic).
   b. A head (–COONa or –COOK) that loves water (hydrophilic).
5. Label the oil-loving part and the water-loving part.
6. Outline the main stages of soap manufacture:
   a. Hydrolysis of fat/oil (Saponification): Fat + Alkali → Glycerol + Soap.
   b. Separation: Soap is separated from glycerol and impurities.
   c. Purification: Soap is washed, dried, and sometimes mixed with additives (perfume, colouring).
7. Give a short oral presentation on observations and improvements.

Activity 6.9 Soap versus Soapless Detergents

1. Observe examples of soapy detergents (traditional soap) and soapless detergents (synthetic detergents) in ***Figure 6.2***.
   a. Discuss how their structures are different.
   b. Discuss how each one removes grease and dirt.



Figure 6.2: Soapy detergents (traditional soap) and soap - less detergents

2. Make a list of advantages of soapless detergents compared to soaps.
   a. Which one works better in hard water?
   b. Which one leaves less residue on clothes and surfaces?
3. Share your ideas with the class.

REVIEW QUESTIONS

1. **a.** State the functional group present in
   - **i.** Esters
   - **ii.** Amides
   - **iii.** acyl halides.

   b.
   - **i.** Write the general formula of an ester.
   - **ii.** Give two everyday uses of esters.

2. Write a balanced chemical equation to show the reaction between:
   - **a.** ethanoic acid and ethanol in the presence of concentrated H_2SO_4.
   - **b.** ethanoyl chloride and ammonia.
   - **c.** Explain briefly why acyl halides are more reactive than amides.

3. **a.** Describe how ethyl ethanoate can be prepared in the laboratory.
   - **b.** State two methods by which amides can be prepared from carboxylic acids.
   - **c.** Compare the solubility of esters and amides in water. Explain your answer in terms of hydrogen bonding.

4. A student prepared an ester in the lab but obtained a very low yield.
   - **a.** Suggest three possible reasons for the low yield.
   - **b.** Propose modifications to improve the yield.
   - **c.** Discuss the importance of alkanoic acid derivatives in everyday life, with at least one example each from the categories below.
     - food and flavours
     - medicine
     - industrial processes.

5. **a.** State the general molecular formula of a triglyceride.
   - **b.** Name the type of chemical bond that joins glycerol to fatty acids in fats and oils.
   - **c.** Mention two physical properties that distinguish fats from oils.

6. **a.** Write a balanced chemical equation for the formation of a triglyceride from glycerol and three fatty acids.
   b. Explain briefly why fats are solid at room temperature while most oils are liquid.
7. **a.** Describe the structure of glycerol and fatty acids that make up fats and oils.
   b. Distinguish between saturated and unsaturated fatty acids using their structural features.
   c. Compare the solubility of fats and oils in water and in organic solvents. Explain your answer.
8. **a.** Discuss three everyday uses of fats and oils, giving one example for each use.
   b. A nutritionist wants to compare the health effects of saturated and unsaturated fats in the human diet.
      i. Explain why excess consumption of saturated fats may be harmful.
      ii. Suggest two benefits of including unsaturated oils in the diet.

Soap and Soapless Detergents

1. **a.** State the functional group present in a soap molecule.
   b. Write the general formula of the hydrophilic head in soap.
   c. List two examples each of
      i. Soapy detergents
      ii. Soapless detergents.
2. **a.** Explain why soaps do not lather well in hard water.
   b. State two differences in the structure of soap and soapless detergents.
3. **a.** Describe how soap can be prepared locally from vegetable oil.
   b. Write a balanced chemical equation for the saponification of a fat.
   c. Compare the cleansing action of soap and soapless detergents, explaining the role of the hydrophilic and hydrophobic parts.
4. **a.** Discuss three everyday uses of soaps and three uses of soapless detergents.

b. A chemist wants to reduce the environmental problems caused by soapless detergents.

i. Identify one major environmental problem caused by some synthetic detergents.

ii. Suggest two solutions to reduce this problem.

SECTION

7

POLYMERS



CHEMISTRY OF CARBON COMPOUNDS

Organic Functional Groups

INTRODUCTION

Polymers are very large molecules that can be natural or man-made. Natural polymers include proteins (from amino acids) and carbohydrates (from glucose), which are important for life. Synthetic polymers, such as polythene, PVC, nylon, and terylene, are made by humans and are widely used in everyday materials.

KEY IDEAS

- **Amino acid:** the small building block (monomer) that makes up proteins.
- **Carbohydrates:** natural polymers made of glucose molecules, providing energy.
- **Glucose:** a simple sugar molecule, the monomer of carbohydrates.
- **Natural polymers:** polymers that occur in living things, such as proteins and carbohydrates.
- **Polymers:** very large molecules made of many repeating units (monomers).
- **Proteins:** natural polymers made of amino acids, important for growth and repair.
- **Synthetic polymers:** man-made polymers like nylon, PVC, and polythene.

POLYMERS

Natural Polymers

Polymers are very big molecules made up of many small, repeating units called monomers. Think of them like a chain made of many identical links. The word polymer comes from Greek: “poly” means *many* and “mer” means *parts*

So, polymers are “many parts joined together.”

How are polymers made?

The making of polymers is called polymerisation. In this process, monomers join together by forming strong bonds, just like beads being joined on a string.

Types of Polymerisations

1. **Addition Polymerisation (Chain-Growth)**

   In this process, *monomers join one after another* to make a long chain. No small molecule is released – everything stays in the chain. Usually happens with monomers that have double bonds **(C=C)** or triple bonds **(C≡C)**.

   Key Features

   a. The chain grows very fast once it starts.

   b. The polymer has the same atoms as the monomer (nothing lost).

   c. Needs careful control of temperature because it gives out heat.

   Example: Making polythene from ethene.

2. **Condensation Polymerisation (Step-Growth)**

   Here, monomers join together but at the same time, a small molecule like water (H_2O), hydrogen chloride (HCl), or ammonia (NH_3) is released. The monomers must have two or more special reactive groups to link up.

   Key Features

   a. The chain grows step by step (slowly).

   b. Small molecules are given off during the reaction.

   c. The final polymer is different from the monomer, because something was lost.

   Example: Making nylon (water is released).

Types of Polymers by Source

1. Natural Polymers: Found in living things, e.g., proteins, carbohydrates, nucleic acids (DNA).
2. Synthetic Polymers: Man-made, e.g., plastics, nylon, polyester, rubber.

Properties of Polymers

The properties of polymers depend on the following

1. The type of monomer used.

2. The arrangement of the monomers.
3. The length of the chains.
4. The interactions between chains.

These factors determine whether a polymer is strong, flexible, soluble, or resistant to heat.

Addition vs. Condensation Polymerisation

Table 7.1: Comparison of addition and condensation polymerisation

	Addition Polymerisation	Condensation Polymerisation
What happens	Monomers add one by one to form a chain.	Monomers join, but a small molecule (like water or HCl) is released.
Speed	Fast once it starts.	Slow, step-by-step growth.
Monomer type	Needs monomers with double (C=C) or triple bonds (C≡C).	Needs monomers with two or more special groups (functional groups).
Atoms lost?	No atoms lost – polymer has same atoms as monomer.	Atoms lost – small molecules (e.g., H_2O) are removed.
Polymer composition	Same as the monomer.	Different from the monomer.
Example	Polythene (from ethene).	Nylon (water is released when formed).

Proteins

Proteins are natural polymers. They are made from small building blocks called amino acids. Many amino acids join together in a chain to form a protein.

Functions of Proteins in the Body

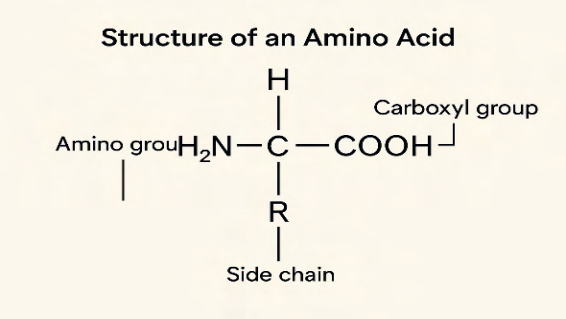
1. Help with growth and repair of body tissues (muscles, skin, hair).
2. Make important enzymes and hormones.
3. Provide some energy when needed.

Example foods rich in protein: eggs, beans, fish, meat, milk, nuts.

Structure of an Amino Acid

Every amino acid has the same basic structure

1. Central Carbon Atom (C) – also called the alpha carbon.
2. Amino Group ($–NH_2$) – this makes the molecule a base.
3. Carboxyl Group (–COOH) – this makes the molecule an acid.
4. Hydrogen Atom (–H) – attached to the central carbon.
5. Side Chain (–R group) – this part is different for each amino acid and gives it special properties.



Special Features

Each amino acid has two important parts

1. Amino group ($–NH_2$)
2. Carboxyl group (–COOH)

They link together like beads on a string to make proteins.

Protein Formation (Peptide Bonds)

When two amino acids join together, the acid part (–COOH) of one reacts with the amino part ($–NH_2$) of the other. In this process, a water molecule (H_2O) is removed, and a peptide bond is formed.

$$H_2N - CHR - COOH + H_2N - CHR' - COOH \rightarrow H_2N - CHR - CO - NH - CHR' - COOH + H_2O$$

This process keeps repeating, so many amino acids link together in a long chain. The chain always grows in one direction from the amino end (N-end) to the acid end (C-end). The order of amino acids in the chain (called the primary structure) is controlled by the genetic code (DNA).

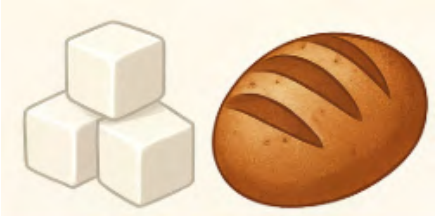
Because there are 20 different amino acids, proteins can be made in almost endless ways. This is why proteins can do many different jobs in living things.

1. Keratin makes your hair and nails strong.
2. Enzymes help break down food in your body.

So, proteins show us how simple building blocks can combine to make complex, useful materials for life.

Carbohydrates

Carbohydrates are food molecules made of carbon, hydrogen, and oxygen. Their name means "carbon with water." They are one of the main groups of substances living things need, along with proteins, fats (lipids), and nucleic acids. Carbohydrates give the body energy. Examples include sugar, bread, rice, and potatoes.



Sugar bread

Classification of Carbohydrates

Carbohydrates are grouped by how many sugar units they have.

1. *Monosaccharides (One sugar unit):* The simplest sugars.

   Examples: Glucose (in fruits), Fructose (in honey), Galactose (in milk).
2. *Disaccharides (Two sugar units):* Made when two simple sugars join together.

   Examples

   a. Sucrose = Glucose + Fructose (table sugar)

   b. Lactose = Glucose + Galactose (milk sugar)

   c. Maltose = Glucose + Glucose (malt sugar)
3. *Oligosaccharides (A few sugar units):* Chains of 3–10 sugars. Often found on cell surfaces and help cells "recognize" each other.

4. *Polysaccharides (Many sugar units):* Very long chains of sugars. Used for energy storage (e.g., starch in plants, glycogen in animals) or for structure (e.g., cellulose in plants, chitin in insects).

Glucose – The Main Building Block of Carbohydrates

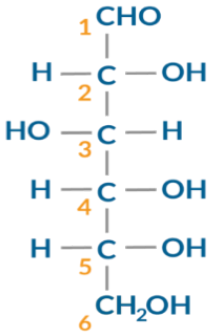
Glucose ($C_6H_{12}O_6$) is a simple sugar (monosaccharide). It is the main source of energy for living things, like plants, animals, and humans.

Structure of Glucose

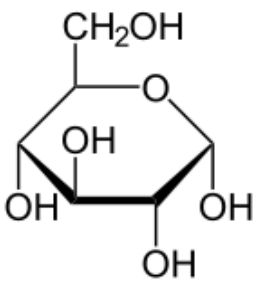
It has 6 carbon atoms usually arranged in a ring.

It has several –OH groups (hydroxyl groups) that make it dissolve easily in water.

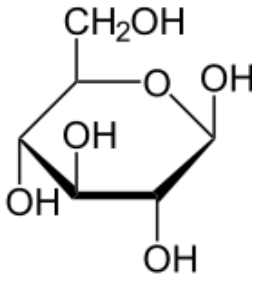
In one form, it has a special –CHO group (aldehyde group).



Glucose can exist in two slightly different ring forms: **α-glucose** and **β-glucose**. This difference is important because it affects how glucose units join to form bigger carbohydrates like starch and cellulose.



α-glucose



β-glucose

Figure 7.1: Structures of α and β glucose

Glycosidic Bonds – How Glucose Units Join

Glucose molecules can join together to make bigger carbohydrates. They do this through a condensation reaction, where a water molecule (H_2O) is removed. The link that holds two glucose units together is called a glycosidic bond.

Example of reaction: $C_6H_{12}O_6 + C_6H_{12}O_6 \rightarrow C_{12}H_{22}O_{11} + H_2O$

The type of glycosidic bond formed (α or β) depends on how the glucose units are joined. This difference (α or β) decides whether the larger carbohydrate will be used for energy storage (like starch or glycogen) or for structure (like cellulose).

Major Glucose Polymers and Their Functions

When many glucose units join together, they form glucose polymers. These are very important for living things.

Table 7.2: Glucose polymers and their functions

Polymer	Where found	Function	Examples
Starch	Plants (rice, yam, potatoes, maize)	Stores energy for plants; source of energy for humans	Rice, bread, maize
Glycogen	Animals (liver, muscles)	Stores energy in animals; quickly broken down to glucose	Human liver and muscles
Cellulose	Plant cell walls	Provides strength and structure; dietary fibre in humans	Wood, cotton, vegetables
Chitin	Exoskeletons of insects, crabs, lobsters	Provides hard protective covering and support	Insects, crabs, lobsters

Structural Diversity from One Building Block

All these polymers are made from the same sugar unit – glucose. But the way glucose units are joined together makes a big difference!

Table 7.3: Structural differences in different glucose polymers

Polymer	Properties	Function
Starch & Glycogen	Easy to break down	Energy storage in plants and animals
Cellulose	Strong, indigestible fibres	Strength and support in plant cell walls
Chitin	Tough, flexible covering	Protective shells in insects, crabs, lobsters

The properties of a big molecule (polymer) depend not only on the building block (glucose), but also on how the blocks are joined together.

Activity 7.1 Testing Carbohydrates

1. Wear goggles and gloves. Always handle test solutions carefully and listen to your teacher's instructions.
2. Test the following solutions
   - Glucose solution
   - Sucrose solution
   - Starch solution
3. Test for Reducing Sugars (Benedict's Test)
   a. Add Benedict's solution to each test tube.
   b. Heat gently in a warm water bath.
   c. Watch the colour changes
      - Blue = No sugar
      - Green/Yellow/Red = Sugar present
4. Test for Starch (Iodine Test)
   a. Add a few drops of iodine solution to the starch sample.
   b. Watch the colour change: Blue-black = Starch present
5. Record Results

   Fill in the table below with your observations

Food Sample	Test Used	Observation (Colour Change)
Glucose solution	Benedict's Test	
Sucrose solution	Benedict's Test	
Starch solution	Iodine Test	

6. Answer these questions
   a. Which samples had sugar?
   b. Which sample had starch?
   c. Why do we need carbohydrates?

Activity 7.2 Carbohydrate Classification

1. Observe the list foods items in ***Figure 7.2,*** list 5 foods that contain carbohydrates.



Figure 7.2: List of food items

2. Compare your list with a partner.
3. On your paper, draw three columns and title them
   a. Monosaccharides (e.g., glucose, fructose)
   b. Disaccharides (e.g., sucrose, maltose, lactose)
   c. Polysaccharides (e.g., starch, cellulose, glycogen)
4. From the food samples, place or draw each food under the most likely group.
5. Write one sentence under each column: "These usually have **1**, **2**, or **many** sugar units."

Test for starch

6. Put a tiny piece or smear of the sample (bread/pasta) on a white tile/paper towel.
7. Ask for one drop of iodine from your teacher. Place it on the sample.
8. Observe: If it turns blue-black, starch is present.
9. Record in your notebook: Food → Iodine colour → Starch present? (Yes/No).
10. In one sentence, explain: **"Polysaccharides → Disaccharides → Monosaccharides."**
11. Your group will receive one carbohydrate to represent
    a. Starch
    b. Sucrose
    c. Lactose
    d. Cellulose
12. Explain Where your assigned carbohydrate is found.

    Can humans digest it easily?

    Why or why not?
13. Present your explanation to the class.

SYNTHETIC POLYMERS (MADE-BY-MAN PLASTICS)

A synthetic polymer is a very large molecule made from many smaller, repeating parts called monomers. These are not found naturally – they are made by scientists in factories through chemical reactions. Most synthetic polymers come from oil or natural gas. Some newer ones are made from plants. Scientists can control how polymers are made by choosing the following

1. The type of monomer
2. How the monomers join
3. How long the chains are
4. Whether the chains are straight, branched, or linked together.

Because of this, polymers can be designed to have the following features.

a. Hard or soft

b. Bendable or stiff
c. Clear or not see-through
d. Waterproof or able to absorb water.

One problem: most synthetic polymers (like plastics) do not break down easily in nature. This leads to plastic pollution.

Classification of Synthetic Polymers

1. **By Structure**
   a. Linear: one straight chain
   b. Branched: main chain with side chains
   c. Cross-linked: chains joined together
   d. Network: lots of cross-links forming a 3D web.
2. **By Heat Behaviour**
   a. Thermoplastics: soften when heated and can be reshaped.(They have weak intermolecular forces that break down under heat)
   b. Thermosets: set hard when heated and cannot be reshaped. (Due to its three-dimensional cross-linked structure, they are rigid and cannot be remelted or reshaped)
   c. Elastomers: rubber-like; can stretch and go back to their original shape when released. (They have weak intermolecular forces allowing them to stretch. They are also lightly cross-linked and cannot be remelted)
3. **By Reaction Type**
   a. Addition Polymers: monomers join without making other small molecules.
   b. Condensation Polymers: monomers join while giving off small molecules (like water).

Addition Polymerisation (Chain Growth)

In addition-polymerisation, monomers with double bonds join to form long chains.

Steps

1. Initiation – a starter (like a free radical) attacks a monomer.
2. Propagation – the chain grows quickly as more monomers join.

3. Termination – the chain stops growing when active ends combine.

Example: Polyethene (Polythene)

Made from ethene (C_2H_4). Reaction: n $CH_2=CH_2 \rightarrow [-CH_2-CH_2-]_n$

Properties

1. Resists acids, bases, and most chemicals.
2. Does not absorb water.
3. Not biodegradable.

Uses

1. Plastic bags, bottles, food containers.
2. Pipes, water tanks, insulation.
3. Car parts, wire coatings.

Example: Polyvinyl Chloride (PVC)

Made from vinyl chloride ($CH_2=CHCl$).

Reaction: n $CH_2=CHCl \rightarrow [-CH_2-CHCl-]_n$

Properties

1. Resists many chemicals.
2. Damaged by heat and sunlight (needs stabilisers).
3. Flame-resistant (self-extinguishing).
4. Not biodegradable.

Uses

1. Pipes, gutters, window frames, doors.
2. Credit cards, bottles, packaging.

Example: Polytetrafluoroethene (PTFE / Teflon)

Made from tetrafluoroethene ($CF_2=CF_2$). Reaction: n $CF_2=CF_2 \rightarrow [-CF_2-CF_2-]_n$

Properties

1. Resists almost all chemicals.
2. Waterproof and non-sticky.
3. Does not break down in sunlight or weather.
4. Not biodegradable.

Uses

1. Non-stick frying pans.
2. Wire and cable coatings.
3. Medical implants and devices.
4. Chemical storage linings.
5. Dental floss coating.

Condensation Polymerisation (Step Growth)

In condensation polymerisation, monomers with special groups react and release a small molecule (like water or HCl). Two different monomers often join together. Molecules get bigger step by step.

General Idea: A–A + B–B → A–B–A–B–A–B + small molecule (like water).

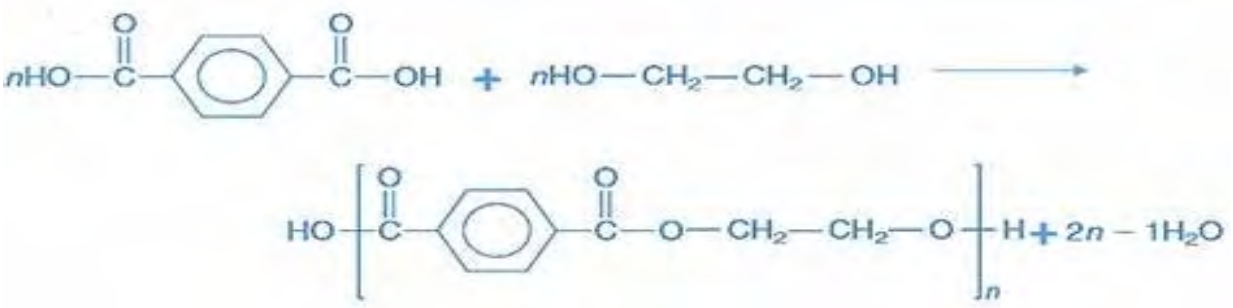
This process makes important materials such as nylon and polyesters.

Types of Condensation Polymers

Polyesters

What are they?

Polyesters are man-made plastics made when two special chemicals (a dicarboxylic acid and a diol) join together. When this happens, water is released, and long chains called polymers are formed.



Properties of Polyesters

1. They do not wear out easily.
2. They resist acids, oils, and many chemicals.
3. They can handle medium heat.
4. They are light in weight.
5. They do not rot away in nature (non-biodegradable).

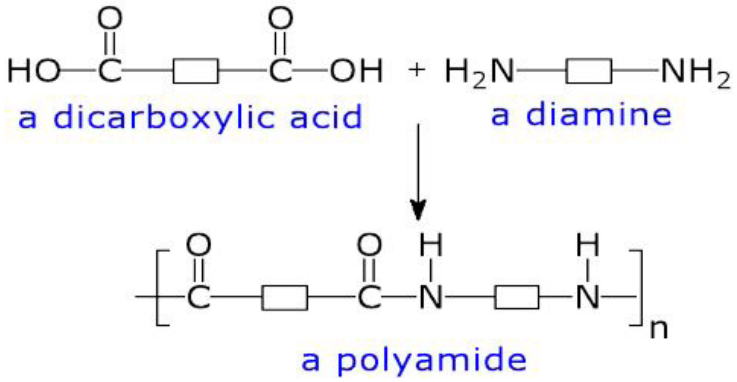
Uses of Polyesters (where we use them)

1. For clothing, sportswear, and other fabrics.
2. For packaging, like PET bottles and food wraps.
3. In cars, for seat belts, airbags, and inside parts.
4. In buildings, for strong roofs and insulation.
5. In hospitals, for stitches and artificial heart valves.
6. In electronics, for tapes and light casings.

Polyamides

What are they?

Polyamides (like nylon) are made when two types of chemicals (dicarboxylic acids and diamines) join together. Water is released, and strong chains with special amide bonds are formed. These bonds allow hydrogen bonds between chains, making them strong.



Properties of Polyamides

1. They are very light materials.
2. They are strong but can also bend easily.
3. They stay flexible even in heat.
4. They are very strong and hard to break.
5. They can handle high temperatures before melting.
6. They resist wearing out, even with rubbing.
7. They do not rust or get damaged by strong chemicals.

Uses of Polyamides

1. In clothing, stockings, sportswear, and swimsuits.
2. In homes, for carpets, curtains, and furniture covers.
3. Outdoors, for parachutes, tents, and climbing ropes.

4. In cars, for tyres, airbags, and engine parts.
5. In machines, for gears, bearings, and cable ties.
6. In wires, for insulation and coverings.
7. In sports, for fishing lines, tennis racquet strings, and bike wheels.
8. In personal care, for toothbrush and hairbrush bristles.

Pollution Problems Caused by Non-Biodegradable Plastics

Plastics that do not rot away (non-biodegradable) cause big problems

1. They stay in the environment for hundreds of years.
2. They fill up landfills and may leak harmful chemicals into the ground.
3. Millions of tons end up in the sea each year.
4. Animals get trapped in plastic waste.
5. Animals eat plastics, which can block their stomachs and cause death.
6. Making plastics uses up lots of oil and produces greenhouse gases.
7. Burning plastics in open air releases dangerous gases.
8. Plastics can block drains and cause floods in cities.
9. When plastics hold water, they can become homes for mosquitoes that spread diseases.

Activity 7.3 Exploring Polymers

Step 1: Group Work (research groups)

Group A: Types of Polymerisations (Addition vs. Condensation).

Group B: Synthetic vs. Natural Polymers.

Group C: Addition Polymers made from Alkenes (e.g., Polyethene).

Group D: Condensation Polymers (e.g., Nylon and Terylene).

Step 2: Research Time

Go to the computer lab and research your group topic using the websites, videos, and templates provided.

You may use the websites below or any relevant textbooks for your research.

a. https://byjus.com/chemistry/difference-between-addition-and-condensation-polymerization/?utm_source=chatgpt.com

b. +https://en.wikipedia.org/wiki/Polymer?utm_source=chatgpt.com

c. https://en.wikipedia.org/wiki/Addition_polymer?utm_source=chatgpt.com

d. https://en.wikipedia.org/wiki/Condensation_polymer?utm_source=chatgpt.com

Step 3: PowerPoint Training

1. Open Microsoft PowerPoint.
2. Create a new slide for your group topic.
3. Show chemical structures and reactions clearly.

   Example: show ethene changing into polyethene.
4. Use colours, arrows, and labels to make your slides easy to understand.

Step 4: Finish your group slides

1. Show them to the teacher before the final presentation.
2. The teacher will check for scientific accuracy (make sure the reactions and properties are correct).

Step 5: Class Presentations

1. Each group will present their work to the class.
2. Use drawings, molecular models, or animations to make the presentation interesting.
3. Be ready to answer simple questions from classmates.

Step 6

Make one big comparison chart on the board or wall.

The chart should include the following.

a. Addition vs. Condensation polymers.

b. Properties (strength, flexibility, biodegradability).

c. Uses in Ghana (e.g., packaging, clothes, building materials).

Activity 7.4 Plastics and the Environment

Step 1: *Observe the different plastic items your teacher shows you (bags, bottles, cups, pipes, rubber bands).*

1. Based on your observation what are plastics?
2. What do we already know about them?
3. Write your ideas together on the board.

Step 2: *Group Investigation*

1. Working in small groups study one type of plastic.
   a. Thermoplastics – soften when heated, can be reshaped.
   b. Thermosets – hard plastics that cannot be reshaped after heating.
   c. Elastomers – rubber-like, can stretch and return to shape.
2. Create an informational poster showing
   a. What your plastic type is.
   b. Its properties.
   c. Examples of everyday use.

Step 3

1. Use the link to watch a short video showing plastic waste in Ghana's rivers and cities. https://youtu.be/uFlKQdf6Klo
2. Alternatively, if no video is available, look at newspaper articles and photographs about plastic pollution.

Step 4: *Take a minute to think about plastic pollution you have seen in your community.*

1. Share your thoughts with a partner.
2. Together, tell the class what you noticed.

Step 5

1. As a class, make one big list of environmental problems caused by non-biodegradable plastics in Ghana.
2. Examples: dirty streets, blocked drains, flooding, dead animals, mosquito breeding, etc.

Step 6

1. In your groups, discuss and debate on the questions below.
   - **a.** What are the advantages of plastic?
   - **b.** What are the disadvantages?
   - **c.** What are possible solutions to plastic pollution?
2. Think about technology (like recycling) and laws/policies (like plastic bans).

Step 7: *Design a short pamphlet (one page) with pictures and messages about*

1. How to use plastics responsibly.
2. How to dispose of plastics properly.

REVIEW QUESTIONS

1. **a.** Define monomer.

b. Define polymer.

2. State two differences between natural and synthetic polymers, giving one example of each.

3. Name the linkage that

a) joins amino acids in proteins,

(b) joins glucose units in starch/cellulose.

4. Classify each item as mono-, di-, or polysaccharide: glucose, sucrose, cellulose, glycogen.

5. State whether protein formation is addition or condensation polymerisation. Give a reason.

6. State whether starch formation from glucose is addition or condensation polymerisation. Give a reason.

7. Write a balanced condensation equation for forming the dipeptide Gly–Ala from glycine and alanine.

8. A hexapeptide is formed from 6 amino acids.

a. How many peptide bonds are present?

b. How many molecules of water are eliminated during its formation?

9. Write the overall formula equation for forming maltose from two glucose molecules.

10. Complete the table (each row)

Polymer	Linkages (major)	Branching?	Main function
Starch			
Glycogen			
Cellulose			

11. Indicate the expected test outcome (Positive/Negative and observed colour/change) for these samples

a. Egg white with biuret test

b. Gari/cooked rice with iodine test

c. Ripe banana juice with Benedict's test.

12. Explain briefly why proteins can show different shapes/functions even though they are all polymers of amino acids.

13. Explain why humans digest starch but not cellulose, although both are polymers of glucose.

14. Show, using symbols, how glucose units form the repeating unit of starch. Hence write the general formula of starch as a polymer of glucose and state how many water molecules are eliminated per n glucose units (use the conventional polymer formula).

15. Compare primary structure of proteins with that of polysaccharides. How does sequence relate to function in proteins versus carbohydrates?

16. A learner claims, "All polymers made from glucose are soluble and sweet." Evaluate this claim with two examples and structural reasons.

17. State two laboratory conditions for hydrolysis of

a. a dipeptide to amino acids

b. maltose to glucose.

18. Investigation design: Test three local foods (e.g., beans, kenkey, groundnuts/rice) for protein and starch. Provide the following

- Aim and hypothesis
- Reagents & safety (biuret, iodine; goggles/gloves)
- Procedure + control
- Results table template (food | biuret result | iodine result | inference)
- Conclusion linking results to macromolecules.

19. Quantitative reasoning: When glucose polymerises to starch, each glucose unit (M_r 180) becomes a repeating unit ($C_6H_{10}O_5$, M_r 162).

a. What mass fraction of the original glucose is lost as water per unit formed?

b. If 5.40 g of glucose are converted completely to starch repeating units, calculate the mass of water evolved.

c. State one assumption/limitation of this calculation for real starch.

20. Case analysis: A cereal label per 100 g shows: Protein 9 g; Carbohydrate 74 g (Sugars 12 g); Fibre 5 g.

- **a.** Classify the carbohydrate portion into likely mono-/di- (sugars) and poly- (starch + non-starch fibre).
- **b.** Predict results for iodine, Benedict's, and biuret tests on an aqueous extract. Justify briefly.
- **c.** Suggest one dietary implication of the polysaccharide fraction.

21. **a.** Define the term synthetic polymer.

b. Give two examples of synthetic polymers and their monomers.

22. Explain briefly how addition polymerisation and condensation polymerisation differ. Illustrate your answer with one example each.

23. **a.** Describe the formation of polyethene from ethene. Write the chemical equation.

- **b.** State three properties of polyethene.
- **c.** Mention three uses of polyethene.

24. Nylon is an important synthetic polymer produced by condensation polymerisation.

- **a.** Name the two types of monomers used in making nylon-6,6.
- **b.** Write a simple equation to represent the formation of nylon.
- **c.** State three properties of nylon.
- **d.** Mention three uses of nylon.

25. Plastics are very useful but have created major environmental problems in Ghana.

- **a.** State four environmental problems caused by non-biodegradable plastics.
- **b.** Suggest three practical solutions that can reduce plastic pollution in Ghana.

ANSWERS TO REVIEW QUESTIONS

Section 1

1. a. pH is a measure of the acidity or alkalinity of a solution. It is defined as the negative logarithm (base 10) of the hydrogen ion concentration: $pH = -\log[H^+]$.

b. The relationship is inverse and logarithmic: as the hydrogen ion concentration $[H^+]$ increases, the pH decreases.

c. i. $pH < 7$: Acidic solution.

ii. $pH = 7$: Neutral solution.

iii. $pH > 7$: Basic (alkaline) solution.

2. a. $pH = -\log[H^+] = -\log(3.2 \times 10^{-4})$

$= -(\log 3.2 + \log 10^{-4})$

$= -(0.51 - 4)$

$= 3.49$

$pOH = 14 - pH$

$= 14 - 3.49 = 10.51$

b. Since $pH < 7$, the solution is acidic.

c. Such acidity is typical of soft drinks or vinegar. It can cause corrosion of metals or tooth enamel if not managed properly.

d. Kw increases with temperature because water dissociation is endothermic. Higher temperature shifts equilibrium toward more H^+ and OH^-, raising Kw.

3. a. $[H^+] = 10^{-5.2} = 6.31 \times 10^{-6}$ mol/dm³

b. If pH decreases to 4.2: New $[H^+] = 10^{-4.2} = 6.31 \times 10^{-5}$ mol/dm³. This is a 10×increase in $[H^+]$.

c. Black coffee or banana may have this pH.

4. a. $[OH^-] = 1.0 \times 10^{-3}$

$pOH = -\log[OH^-]$

$= -\log(1.0 \times 10^{-3}) = 3$

b. pH = 14 - pOH = 14 - 3 = 11

c. At 25 °C, Kw = 1.0×10^{-14} so $-\log[H^+] + -\log[OH^-] = 14$. This relationship holds at 25 °C; at other temperatures the sum equals –log Kw for that temperature.

d. Example: Soap has a pH around 10–11. This basic pH helps break down oils and grease from skin.

5. a. pH difference = 7.4 - 2.5 = 4.9 units

Each unit = 10× difference, so total difference = 104.9 ≈ 79,400×

Lemon juice has 79,400 times more H^+ than blood.

b. Lemon juice is highly acidic and corrosive; blood must remain near neutral to support life.

c. A small change in blood pH disrupts enzyme activity and cell function, leading to serious health issues.

6. a. Agriculture

Medicine

Water treatment

b. i. Soil – pH 6–7 – affects nutrient uptake

ii. Blood – pH 7.4 – supports body function

iii. Drinking water – pH ~7 – ensures safety and taste

c. In farming, wrong pH causes poor plant growth or nutrient deficiency.

7. a. Ammonia solution < Baking soda < Milk < Vinegar < Lemon juice

b. Milk (pH 6.5)

c. i. Red/orange

ii. Blue/light blue

d. Ammonia is caustic and can damage skin due to high alkalinity.

8. a. NaOH reacts with HCl in a 1:1 ratio.

Remaining moles of NaOH = $3.60\times10^{-3} - 5.95\times10^{-4}$

$= 3.005\times10^{-3}$ mol

Concentration of OH^- = 3.005×10^{-3} mol / 1.00 dm^3

= 3.005×10^{-3} mol/dm^3

pOH = $-\log(3.005\times10^{-3}) = 2.52$

pH = 14 - 2.52 = 11.48

b. Direct use of pH = –log[H^+] ignores the neutralization between acid and base. The net amount of unreacted base determines the actual pH.

9. a. [H^+] = $10^{-8.42}$ = 3.80×10^{-9} mol/dm³

pOH = 14 - 8.42 = 5.58

[OH^-] = $10^{-5.58} \approx 2.63 \times 10^{-6}$ mol/dm³

b. To lower pH (make more acidic), farmers can:

Add elemental sulphur or sulphur-containing compounds.

Apply ammonium-based fertilizers (e.g., ammonium sulphate).

10. a. The pH scale is logarithmic, meaning each pH unit represents a tenfold change in [H^+]. It allows very small H^+ values to be expressed as manageable numbers (e.g., 10^{-7} = pH 7), making comparison and communication easier.

b. Cooking: Baking soda (pH ~9) neutralizes acids in recipes and helps cakes rise.

Medicine: Blood must stay near pH 7.4. Even small changes can harm enzymes and bodily functions.

Alternatively:

Gardening: Soil with pH 6–7 is optimal for most crops; pH affects nutrient availability.

Environmental monitoring: Water with pH <6 or >8.5 may harm aquatic life or corrode pipes.

11. a. Initial [H^+] = $10^{-7.8}$ = 1.58×10^{-8} mol/dm³

Final [H^+] = $10^{-6.4}$ = 3.98×10^{-7} mol/dm³

Increase = $3.98 \times 10^{-7} \div 1.58 \times 10^{-8}$ = 25×

b. This increase in acidity may harm fish gills, reduce oxygen intake, disturb plant photosynthesis, and destroy aquatic habitats.

c. Use of organic compost instead of synthetic fertilizers. Establish buffer zones with plants to absorb runoff before reaching water bodies.

Multiple Choice Questions

1. C. Acid dissociation constant
2. C. The negative logarithm of Ka
3. D. They have low pKa values
4. B. 4.74
5. C. Higher Ka = stronger acid
6. B. High Kb, low pKb
7. A. 6
8. B. 4.74
9. B. pH = 4.75
10. B. 6.5

Essay Type Questions

1. A buffer solution is a mixture that resists changes in pH when small amounts of acid or base are added.
2. An acidic buffer contains a weak acid and its salt (conjugate base). For example, CH_3COOH and CH_3COONa.
3. To maintain a stable pH by neutralising added acids or bases.
4. The conjugate base component (e.g., CH_3COO^-) reacts with and absorbs added H^+ ions, forming the weak acid (CH_3COOH).
5. The weak acid (e.g., CH_3COOH) donates H^+ to react with the OH^- ions, forming water and CH_3COO^-.
6. pH = pKa + log([Salt]/[Acid])
7. Step 1: pKa = -log $(1.8 \times 10^{-5}) \approx 4.74$

   Step 2: pH = 4.74 + log (0.1 / 0.2) = 4.74 + log (0.5) = 4.74 - 0.30 = 4.44
8. Buffer solutions maintain a stable pH, which is essential for enzyme activity and overall homeostasis in living organisms.
9. Step 1: $pK_b = -\log (1.8 \times 10^{-5}) \approx 4.74$

   Step 2: pOH = 4.74 + log (0.2 / 0.3) = 4.74 + log (0.667) ≈ 4.74 - 0.18 = 4.56

   Step 3: pH = 14 - pOH = 14 - 4.56 = 9.44

10. Choose a weak acid with pK_a close to 5.0, such as acetic acid ($pK_a \approx 4.74$). Using CH_3COOH and CH_3COONa in nearly equal concentrations will give a buffer close to the desired pH.

11. $pH = pK_a + \log([Salt]/[Acid])$

12. $pH = pK_a + \log([Salt]/[Acid]) = -\log (1.8 \times 10^{-5}) + \log (1) = 4.74 + 0 = 4.74$

13. $5.0 = pK_a + \log (2) \rightarrow pK_a = 5.0 - 0.3010 = 4.699 \rightarrow K_a = 10^{(-4.699)} \approx 2.0 \times 10^{-5}$

14. $pOH = 14 - pH = 4.75 \rightarrow pK_b = -\log (1.8 \times 10^{-5}) = 4.74 \rightarrow 4.75 = 4.74 + \log([Salt]/[Base]) \rightarrow \log([Salt]/[Base]) = 0.01 \rightarrow [Salt]/[Base] = 10^{0.01} \approx 1.023 \rightarrow [Base]/[Salt] \approx 0.978$

Since $pK_a \approx pH$, mix equimolar amounts of acetic acid and sodium acetate.

15. For 250 cm_3 of 0.10 mol dm^{-3} buffer, dissolve 0.025 mol (1.5 g) of acetic acid and 0.025 mol (2.05 g) of sodium acetate in water and dilute to 250 cm^3.

16. Use dihydrogen phosphate ($H_2PO_4^-$) and hydrogen phosphate (HPO_4^{2-}) as the conjugate acid-base pair. The pKa of $H_2PO_4^-$ is 7.2, which is close to 7.4. Use the Henderson-Hasselbalch equation to adjust the ratio [Base]/[Acid] to achieve pH 7.4.

17. Dilution has minimal effect on the pH of a buffer since the ratio [Salt]/[Acid] remains constant. However, buffer capacity decreases as total concentration decreases, making the buffer less effective at resisting pH changes.

Section 2

1. B	**2.** B	**3.** C	**4.** C	**5.** B	**6.** B	**7.** B
8. D	**9.** B	**10.** C	**11.** A	**12.** B	**13.**	**14.** B

Structured type

1. Hydration is the process where water molecules surround and interact with the ions of an ionic compound when it dissolves in water. The polar water molecules stabilize the ions in solution by forming ion-dipole interactions, thereby reducing the energy needed to separate them.

2. Hydrolysis of salts is the chemical reaction between the ions of a dissolved salt and water, resulting in the formation of either acidic or basic solutions. This occurs when the salt comes from a weak acid or a weak base.

3. Salts that produce acidic solutions usually originate from a strong acid and a weak base, where the cation hydrolyses to release H_3O^+ ions. Salts that produce basic solutions originate from a weak acid and a strong base, where the anion hydrolyses to release OH^- ions.

4. NaCl is formed from a strong acid (HCl) and a strong base (NaOH). Both Na^+ and Cl^- do not react with water; hence no hydrolysis occurs, and the solution remains neutral (pH ≈ 7).

5. $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

6. $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$. This produces OH^- ions, making the solution basic.

7. $FeCl_3$ hydrolyses in water to form an acidic solution:

$Fe^{3+} + 3H_2O \rightleftharpoons Fe(OH)_3 + 3H_3O^+$. The release of H_3O^+ ions lower the pH.

8. CO_3^{2-} hydrolyses in water: $CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$. The release of OH^- makes the solution basic.

9. KNO_3 is formed from a strong acid (HNO_3) and a strong base (KOH). Both K^+ and NO_3^- do not hydrolyse in water. Therefore, the solution is neutral (pH ≈ 7).

10. NH_4Cl: $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ (acidic)

CH_3COONa: $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$ (basic)

11. NH_4Cl - acidic (weak base + strong acid)

NaCl - neutral (strong acid + strong base)

CH_3COONa - basic (weak acid + strong base)

$AlCl_3$ - acidic (cation hydrolysis of Al^{3+})

12. The weaker the parent acid or base, the greater the hydrolysis of its salt. For example, NH_4Cl (from weak base NH_3) hydrolyses more than NaCl (from strong acid and base).

13. Materials: 0.1 M solutions of NH_4Cl, NaCl, Na_2CO_3, universal indicator/ pH meter

Steps: Test pH of each solution.

Expected: NH_4Cl (acidic, pH<7), NaCl (neutral, pH=7), Na_2CO_3 (basic, pH>7)

14. Salt hydrolysis affects water pH in ecosystems. Acidic or basic runoff from industrial waste or fertilizers can disturb aquatic life and alter nutrient availability.

15. CH_3COONa dissociates into CH_3COO^- and Na^+. CH_3COO^- undergoes hydrolysis:

$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$. The release of OH^- raises the pH. (Diagram would show CH_3COO^- attracting H from water)

Section 3

1. a. Acid-base indicators are weak acids or bases that exhibit different colours in their dissociated and undissociated forms. The colour change occurs over a narrow pH range as the pH of the solution causes the equilibrium between HIn and In^- to shift.

b. Phenolphthalein is colourless in acidic solutions (pH < 8.3) and turns pink in basic solutions (pH > 10.0). It works by existing in two forms: the undissociated (colourless) form and the dissociated (pink) form, depending on the pH.

2. a. Colour changes of indicators across pH ranges:

i. Methyl orange: Red in acidic solution (pH < 3.1), orange at endpoint, and yellow in basic solution (pH > 4.4).

ii. Phenolphthalein: Colourless in acidic solution (pH < 8.3), pink in basic solution (pH > 10.0).

iii. Bromothymol blue: Yellow in acidic (pH < 6.0), green near neutral, blue in basic (pH > 7.6).

b. Methyl red: Red in acidic solution (pH < 4.4) and yellow in basic solution (pH > 6.2).

3. **a.** Comparing suitability of phenolphthalein and methyl orange:

i. Strong acid vs strong base – Both indicators can be used, but phenolphthalein gives a sharper colour change near the equivalence point (pH ≈ 7).

ii. Weak acid vs strong base – Phenolphthalein is suitable since the equivalence point is basic (pH ≈ 8.7).

iii. Strong acid vs weak base – Methyl orange is more suitable because the equivalence point is acidic (pH ≈ 5.3).

b. Strong acid–strong base titration has a steep rise from low to high pH near pH 7. Weak acid–strong base titration starts at a higher initial pH, with a buffer region before a less steep rise and a basic equivalence point.

4. On a typical strong acid–strong base titration curve, label:

Initial pH (low, e.g. ~1)

Steep vertical rise near the equivalence point

Equivalence point at pH ≈ 7

5. **a.** Phenolphthalein is appropriate for HCl–NaOH titration because the pH at the equivalence point is close to 7 and the indicator's colour change occurs in the pH range 8.3–10.0, just beyond the equivalence point.

b. In a titration between CH_3COOH and NaOH, the equivalence point lies in the basic range (~8.7). Phenolphthalein is preferred over methyl orange because its colour change occurs in the basic range, while methyl orange changes colour too early.

6. **a.** Differences in titration curves

i. Strong acid–strong base: Sharp rise around pH 7, steep vertical section.

ii. Weak acid–strong base: Gradual rise, buffer region, equivalence point in basic range (pH ≈ 8.7).

iii. Strong acid–weak base: Initial pH low, equivalence point in acidic range (pH ≈ 5.3).

iv. Weak acid–weak base: Gradual slope, no sharp vertical section, poor for indicator use.

b. The equivalence point pH helps in choosing an indicator that changes colour near that pH, ensuring accurate detection of the endpoint.

7. a. A weak acid–strong base titration curve should show:

Initial pH around 3–4

A buffer region before the equivalence point

Equivalence point at pH ≈ 8.7

Steep section near equivalence

Buffer region significance: It shows where the weak acid resists pH changes due to partial neutralisation, making it ideal for demonstrating buffering action.

b. A weak acid–weak base titration has no steep vertical region because both components react weakly, resulting in a broad pH range. Indicators don't show a sharp colour change, making pH meter detection more reliable.

Section 4

Part A

1.	C	**2.**	D	**3.**	B	**4.**	A	**5.**	B	**6.**	C	**7.**	B
8.	C	**9.**	D	**10.**	C	**11.**	B	**12.**	A	**13.**	D		

Essay Type Questions

1. a. Mg > Fe > Cu > Au. Based on displacement reactions: Mg displaces Fe^{2+}, Fe displaces Cu^{2+}, and Cu displaces Au^{+}.

b. Oxidation is the loss of electrons. Example: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$ (Iron loses 2 electrons and is oxidised).

c. Reduction is a decrease in oxidation number. Example: $Au^{+}(aq) \rightarrow Au(s)$ (Gold's oxidation number decreases from +1 to 0).

d. The strongest reducing agent is Mg. It donates electrons readily and is oxidised to Mg^{2+}, making it the most powerful electron donor.

e. The strongest oxidising agent is Au^{+}. It accepts electrons easily and is reduced to Au, due to being lowest in the reactivity series.

f. No, a gold coin will not react. Gold is less reactive than magnesium and cannot displace Mg^{2+} from solution.

2. Displacement reactions occur when a more reactive metal displaces a less reactive metal from its salt solution. For example, if you place a piece of zinc into a copper(II) sulphate solution, a reaction occurs and brown copper metal is deposited, while the solution fades. This shows that zinc is more reactive than copper. No reaction occurs if copper is placed in zinc sulphate, confirming copper is less reactive. This method helps compare reactivities of metals by observing if a reaction takes place and the intensity of the reaction.

1. (HCl), stopwatch.

Procedure: Place equal amounts of each metal into separate test tubes. Add the same volume and concentration of dilute HCl to each. Observe the rate and intensity of bubbling (hydrogen gas evolution).

Expected Results: Magnesium reacts vigorously with rapid bubbling. Zinc reacts moderately with steady bubbling. Copper shows little to no reaction. This shows that magnesium is the most reactive, followed by zinc, and copper is the least reactive among the three.

2. Metals high in the reactivity series (e.g., potassium, sodium, calcium) lose electrons easily and react quickly with water to form metal hydroxides and hydrogen gas. For example: $2Na + 2H_2O \rightarrow 2NaOH + H_2\uparrow$. Less reactive metals like copper and silver do not react with water because they have a strong hold on their electrons and cannot displace hydrogen from water. The position of a metal in the reactivity series determines whether it can react with water.

3. Rusting of Iron: $4Fe + 3O_2 + 6H_2O \rightarrow 4Fe(OH)_3$. Iron is oxidised; oxygen is reduced. Importance: causes damage to structures, so rust prevention is important.

Combustion of fuels: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. Carbon is oxidised; oxygen is reduced. Importance: produces energy for homes, industries, and transport.

Respiration in cells: $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$ + energy. Glucose is oxidised to release energy. Importance: provides energy for biological processes in living organisms.

4. Balanced Equation: $4Fe + 3O_2 + 6H_2O \rightarrow 4Fe(OH)_3$. Oxidation: Iron (Fe) loses electrons to form Fe^{3+} ions. Reduction: Oxygen (O_2) gains electrons

to form OH^- ions. This simultaneous process of iron losing electrons and oxygen gaining electrons is a redox reaction that results in rust (hydrated iron (III) oxide).

5. Yes.

Experiment: Copper metal in silver nitrate solution. Materials: Copper strip, silver nitrate solution, beaker.

Procedure: Place copper strip in silver nitrate solution. Leave it for 10 minutes.

Observation: Silver crystals form on copper, and the solution turns blue (due to Cu^{2+} ions). *Explanation:* Copper is oxidised (loses electrons), and silver ions are reduced (gain electrons). *Equation:* $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

6. *Comparison:* Potassium reacts explosively with water and burns in air. Magnesium burns with a bright flame in oxygen and reacts with acids. Zinc reacts slowly with acids and oxygen. Copper does not react with water and reacts slowly with oxygen.

Order of reactivity (most to least): Potassium > Magnesium > Zinc > Copper.

7. *Prediction:* A displacement reaction occurs.

Equation: $Fe + 2AgNO_3 \rightarrow Fe(NO_3)_2 + 2Ag$

Explanation: Iron is more reactive than silver and displaces it from solution. Iron is oxidised (loses electrons), and silver ions are reduced (gain electrons), forming solid silver. This demonstrates the application of the reactivity series.

8. **a.** Identification of Solutions A to E

Observations and Interpretation

Solution	Test with KI	Test with $KMnO_4$	Test with $K_2Cr_2O_7$	pH	Interpretation
A	Brown colour formed	$KMnO_4$ remains purple	$K_2Cr_2O_7$ remains orange	3.5	Oxidises $I^- \rightarrow I_2$ (oxidising agent); not strong enough to reduce $KMnO_4$ or $Cr_2O_7^{2-} \rightarrow$ Likely **$CuSO_4$**

Solution	Test with KI	Test with $KMnO_4$	Test with $K_2Cr_2O_7$	pH	Interpretation
B	No change	$KMnO_4$ decolourised	$Cr_2O_7^{2-}$ turns green	6.5	Reduces $KMnO_4$ and $Cr_2O_7^{2-}$ (reducing agent); neutral pH → Likely **Na_2SO_3**
C	Brown colour formed	$KMnO_4$ decolourised	$Cr_2O_7^{2-}$ turns green	2.0	Oxidising and reducing ability; acidic pH → Likely **H_2O_2**
D	No change	$KMnO_4$ remains purple	$Cr_2O_7^{2-}$ remains orange	7.0	No redox activity; neutral → Likely **NaCl**
E	Brown colour formed	$KMnO_4$ remains purple	$Cr_2O_7^{2-}$ turns green	4.0	Oxidises I^- but reduces $Cr_2O_7^{2-}$; weak reducing agent in slightly acidic → Likely **$FeSO_4$**

Based on the test results and chemical behaviour of oxidising and reducing agents, the five unlabelled solutions were identified as shown in the table below.

Solution	Identified Substance	Justification
A	Copper(II) sulphate ($CuSO_4$)	Oxidises iodide (forms brown I_2) but does not reduce $KMnO_4$ or $Cr_2O_7^{2-}$.
B	Sodium sulphite (Na_2SO_3)	Reduces $KMnO_4$ and $Cr_2O_7^{2-}$ but does not oxidise iodide.
C	Hydrogen peroxide (H_2O_2)	Acts as both oxidising ($I^- \rightarrow I_2$) and reducing agent ($MnO_4^- \rightarrow Mn^{2+}$); acidic pH confirms identity.
D	Sodium chloride (NaCl)	No redox behaviour observed; neutral salt.

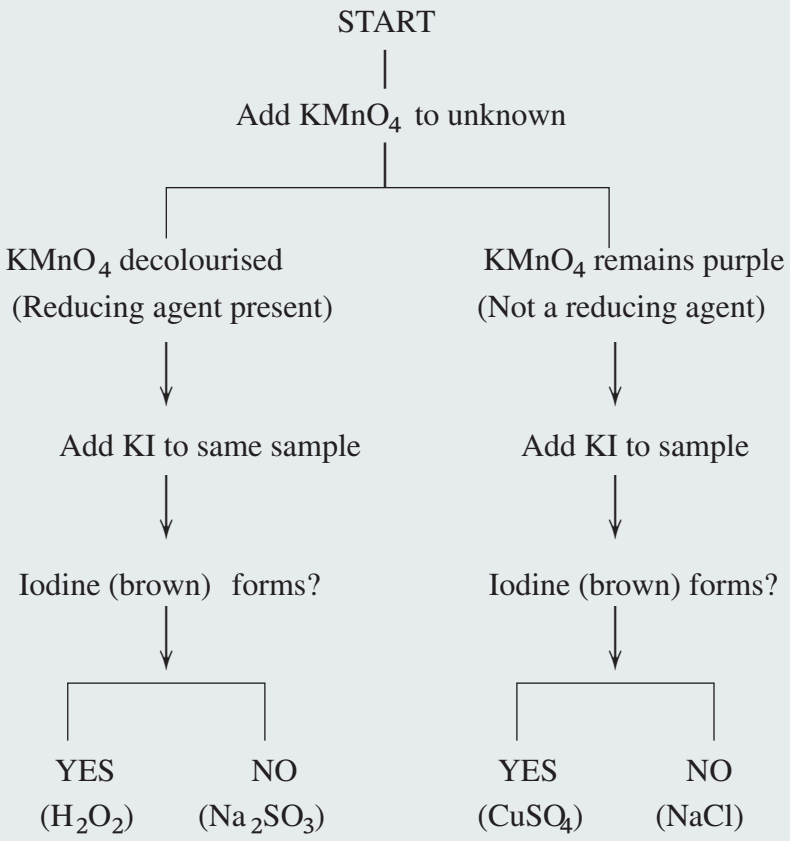
Solution	Identified Substance	Justification
E	Iron(II) sulphate ($FeSO_4$)	Oxidises iodide and reduces $Cr_2O_7^{2-}$; typical behaviour of Fe^{2+} in acidic medium.

b. Flowchart to Identify Solutions Using Two Tests

Two tests were selected for their ability to distinguish between oxidising and reducing agents;

- Potassium permanganate ($KMnO_4$) to detect reducing agents
- Potassium iodide (KI) to detect oxidising agents

Below is a step-by-step flowchart showing how to use these two tests to identify each solution.



9. a. Oxidation: $Mg \rightarrow Mg^{2+} + 2e^-$

Reduction: $O_2 + 4e^- \rightarrow 2O^{2-}$

b. Fe in Fe_2O_3: +3

C: 0 → +4 in CO_2 (oxidised)

Fe: +3 → 0 (reduced)

C is oxidised, Fe is reduced

10. a. Half-reactions

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

$Fe^{2+} \rightarrow Fe^{3+} + e^-$

Multiply Fe equation by 5 and add:

$MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

b. Balanced reaction

$2Cr(OH)_3 + 3ClO^- + 4OH^- \rightarrow 2CrO_4^{2-} + 3Cl^- + 5H_2O$

1. a. Set up test tubes with equal amounts of dilute HCl.

Add equal mass or size of each metal separately.

Observe rate and volume of hydrogen gas evolved.

More bubbles = more reactive metal.

Expected order: Mg > Zn > Cu (Cu shows little/no reaction).

b. $Fe \rightarrow Fe^{2+} + 2e^-$ (oxidation)

$Cu^{2+} + 2e^- \rightarrow Cu$ (reduction)

Brown deposit is copper metal forming on the nail.

2. Redox reactions are used in extracting metals but may produce toxic by-products or acid mine drainage.

To reduce impact: Use greener extraction techniques (e.g., bioleaching), treat wastewater, or recycle metals.

Redox titrations

3. a. Standardisation is the process of determining the exact concentration of a solution using a primary standard.

b. Balanced redox equation: $MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

c. Colour change: From purple to pale pink.

d. Oxidising agent: MnO_4^-

Reducing agent: Fe^{2+}

4. a. Moles of $KMnO_4$ = 0.0200 mol/dm³ × 22.6 cm³ ÷ 1000 = 4.52×10^{-4} mol

b. Moles of Fe^{2+} = 5 × moles of $KMnO_4$ = 5 × 4.52 × 10^{-4} = 2.26 × 10^{-3} mol

c. Concentration of Fe^{2+} = moles ÷ volume in dm^3 = (2.26 × 10^{-3}) ÷ 0.025 = 0.0904 mol/dm^3

d. Safety precautions:

- Wear gloves and goggles to avoid contact with $KMnO_4$ (it stains and is an irritant)
- Use sulphuric acid with care and neutralise spills

5. a. Concordant titres: 22.50 cm^3 and 22.60 cm^3 (within 0.10 cm^3); 21.30 cm^3 is excluded.

b. $KMnO_4$ acts as its own indicator, reducing error from external indicators and simplifying the process.

c. Fe^{2+} can oxidise to Fe^{3+} in air, reducing the amount of Fe^{2+} available for reaction and causing underestimation of concentration.

d. H_2SO_4 ensures acidic medium and does not interfere. HCl is avoided because Cl^- may be oxidised by $KMnO_4$.

6. a. Molar mass of $(NH_4)_2Fe(SO_4)_2 \cdot 6H_2O$ = 392 g/mol. Theoretical concentration = 55.6 ÷ 392 = 0.1418 mol/dm^3. Compare with experimental value and discuss errors (e.g., incomplete dissolution, air oxidation).

b. Possible errors: Incomplete rinsing of burette with $KMnO_4$ → falsely high concentration; Air bubbles in burette tip or parallax error when reading

c. Design experiment:

- Use varying volumes of acid in titration setups
- Keep other variables constant
- Measure and compare average titre values
- Interpret how acid concentration affects redox rate

7. a. Ethanedioic acid (also called oxalic acid) is a weak dicarboxylic acid that acts as a reducing agent in redox titrations.

b. $C_2O_4^{2-}$

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

d. Colour change: From purple (MnO_4^-) to colourless/pale pink (endpoint).

e. Oxidising agent: MnO_4^-; Reducing agent: $C_2O_4^{2-}$

8. a. Moles of $KMnO_4$ = (0.0200 mol/dm³ × 20.0 cm³) / 1000 = 4.00×10^{-4} mol

b. Moles of $C_2O_4^{2-}$ = (5/2) × 4.00×10^{-4} = 1.00×10^{-3} mol

c. Concentration of $C_2O_4^{2-}$ = 1.00×10^{-3} mol / 0.025 dm³ = 0.0400 mol/dm³

d. The reaction is slow at room temperature. Heating increases the rate of reaction to reach the endpoint faster and more accurately.

9. a. If not heated, the reaction proceeds slowly or incompletely, resulting in inaccurate titre values or delayed endpoint.

b. Possible reasons include insufficient heating, poor mixing, or stale reagents.

c. $KMnO_4$ vs Fe^{2+} gives an immediate pink endpoint; $KMnO_4$ vs $C_2O_4^{2-}$ requires heating and has slower colour change.

d. Acidic medium is needed to stabilise MnO_4^- and promote full reduction to Mn^{2+}. Without acid, incomplete reaction occurs.

10. a. Moles = 6.30 g / 126 g/mol = 0.050 mol in 250 cm³ → concentration = 0.050 / 0.250 = 0.200 mol/dm³. Compare with titration result to evaluate purity or error.

b. An incorrectly standardised $KMnO_4$ would yield incorrect titres, leading to an inaccurate calculation of ethanedioate concentration.

c. Design: Set up parallel titrations at 3 temperatures (e.g., 25°C, 60°C, 80°C); use equal volumes and concentrations.

Measure time to endpoint and compare titre volumes. Ensure safe handling of hot liquids.

11. a. Sodium thiosulphate acts as a reducing agent, reacting with iodine to reduce it to iodide ions.

b. $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$

c. Potassium iodide helps release iodine from the oxidising agent by forming I_2 in solution.

d. Starch is used as an indicator. It forms a blue-black complex with iodine. At the endpoint, the blue-black colour disappears.

e. Oxidising agent: I_2; Reducing agent: $S_2O_3^{2-}$

12. a. Moles of $Na_2S_2O_3$ = (0.100 mol/dm³ × 23.5 cm³)/1000 = 2.35×10^{-3} mol

b. Moles of I_2 = (1/2) × moles of $Na_2S_2O_3$ = 1.175×10^{-3} mol

c. Concentration of iodine = 1.175×10^{-3} mol / 0.025 dm³ = 0.0470 mol/dm³

d. Excess KI ensures full liberation of I_2 and prevents reverse reactions.

13. a. $KMnO_4$ has a self-indicating sharp pink endpoint; iodine/starch endpoint is a distinct blue-black to colourless change.

b. Starch forms a stable complex with iodine. Early addition can trap iodine, slowing or hindering the reaction.

c. Iodine may evaporate or be reduced by atmospheric contaminants, leading to a lower titre.

d. Light breaks down iodine or speeds upside reactions, causing loss of iodine and inaccurate results.

14. a. Possible errors: loss of iodine due to evaporation, exposure to light, or incorrect endpoint detection.

b. Set up 2 samples of iodine solution—one in the dark, one in light. Take samples over time and titrate to measure iodine concentration loss.

c. Extract iodine using acid and KI, filter, then titrate with $Na_2S_2O_3$. Adjust for sample mass and ensure reproducibility by multiple trials.

Part B

1. B	**2.** B	**3.** A	**4.** D
5. B	**6.** A	**7.** C	

1.
- Anode (oxidation electrode): Site where oxidation occurs; releases electrons into the external circuit.
- Cathode (reduction electrode): Site where reduction occurs; receives electrons from the external circuit.

- External circuit (wire + load/voltmeter): Pathway for electron flow from anode to cathode; delivers useful electrical work.
- Electrolytes (ionic solutions): Provide ions in each half-cell to sustain half-reactions and carry charge.
- Salt bridge / porous partition: Allows ion migration to maintain neutrality without mixing solutions.

2.
- Cell diagram: Zns. | Zn^{2+}(aq) || Cu^{2+}(aq) | Cus.
  - Anode (oxidation): Zns. → Zn^{2+}(aq) + $2e^-$.
  - Cathode (reduction): Cu^{2+}(aq) + $2e^-$ → Cus..
  - Electron flow: Zn (anode) → Cu (cathode) through wire.
  - Ion flow: Anions to anode, cations to cathode via salt bridge.
  - Purpose of salt bridge: Maintains electrical neutrality, completes circuit.
  - Generates electricity as separated redox reactions force electrons through the external circuit.

3.
- Source of energy: Voltaic – chemical → electrical; Electrolytic – electrical → chemical.
  - Electron flow/polarity: Both flow anode → cathode; Voltaic anode negative, cathode positive; Electrolytic anode positive, cathode negative.
  - Spontaneity: Voltaic – spontaneous; Electrolytic – non-spontaneous.
  - Applications: Voltaic – batteries, fuel cells; Electrolytic – electroplating, metal refining.

4.
- Oxidation (anode): Mgs. → Mg^{2+}(aq) + $2e^-$.
  - Reduction (cathode): Cu^{2+}(aq) + $2e^-$ → Cus..
  - $E°_{cell} = E°_{cathode} – E°_{anode} = (+0.34) – (–2.37) = +2.71$ V.

vSpontaneity: Spontaneous because $E°_{cell} > 0$.

5. $E°(Ag^+/Ag) = +0.80$ V; $E°(Zn^{2+}/Zn) = –0.76$ V

Cell: Zns. | Zn^{2+}(aq) || Ag^+(aq) | Ags.

Identify electrodes:

- Cathode (reduction): $Ag^+ + e^-$ → Ag (higher E°)

- Anode (oxidation): $Zn \rightarrow Zn^{2+} + 2e^-$ (reverse of Zn^{2+}/Zn)

Calculate E°cell:

E°cell = E°cathode − E°anode = (+0.80) − (−0.76) = +1.56 V

Positive, so the cell is feasible under standard conditions.

6. $E°(Cu^{2+}/Cu) = +0.34$ V; $E°(Pb^{2+}/Pb) = -0.13$ V

a. Balanced overall cell reaction

Anode (oxidation): $Pb \rightarrow Pb^{2+} + 2e^-$

Cathode (reduction): $Cu^{2+} + 2e^- \rightarrow Cu$

Overall: Pbs. + Cu^{2+}(aq) → Pb^{2+}(aq) + Cus.

b. Cell potential

E°cell = (+0.34) − (−0.13) = +0.47 V

c. Prediction

E°cell > 0 ⇒ reaction is spontaneous/feasible under standard conditions.

7. $E°(Mg^{2+}/Mg) = -2.37$ V; $E°(Sn^{2+}/Sn) = -0.14$ V; F = 96,500 C mol^{-1}; n = 2

a. Identify electrodes

- Cathode: Sn^{2+}/Sn (−0.14 V) – higher E°, reduced
- node: Mg^{2+}/Mg (−2.37 V) – lower E°, oxidised

b. E°cell

E°cell = (−0.14) − (−2.37) = +2.23 V

c. ΔG° at 298 K

$\Delta G° = -nF\ E°cell = -(2)\ (96{,}500)\ (2.23) = -430{,}390$ J mol^{-1} = −430.39 kJ mol^{-1}

Negative ΔG°, hence spontaneous under standard conditions.

8. Cl_2/Cl^- +1.36 V; Br_2/Br^- +1.09 V

a. Stronger oxidising agent

b. Cl_2 (higher E°, stronger tendency to be reduced).

Overall cell reaction

Cathode (reduction): $Cl_2 + 2e^- \rightarrow 2Cl^-$

Anode (oxidation): $2Br^- \rightarrow Br_2 + 2e^-$

Overall: $Cl_2g. + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2l.$

c. E°cell

E°cell = 1.36 − 1.09 = +0.27 V (spontaneous).

9. $E°(Fe^{3+}/Fe^{2+}) = +0.77$ V; $E°(I_2/I^-) = +0.54$ V; n = 2; F = 96,500 C mol^{-1}

a. E°cell for $2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$

Cathode (reduction): $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (E° = +0.77 V)

Anode (oxidation): $2I^- \rightarrow I_2 + 2e^-$ (reverse of E° = +0.54 V)

E°cell = 0.77 − 0.54 = +0.23 V

b. ΔG°

$\Delta G° = -nF\,E°_{cell} = -(2)\,(96{,}500)\,(0.23) = -44{,}390$ J $mol^{-1} = -44.39$ kJ mol^{-1}

c. Spontaneity

$E°_{cell} > 0$ and $\Delta G° < 0 \Rightarrow$ reaction is spontaneous under standard conditions.

d. Effect of $[Fe^{3+}]$ on cell potential (Nernst argument)

For the reaction as written, the reaction quotient $Q = ([Fe^{2+}]^2 \cdot [I_2]) / ([Fe^{3+}]^2 \cdot [I^-]^2)$.

$E = E° - (0.0592/n) \log Q$ at 298 K.

- Increasing $[Fe^{3+}]$ decreases Q, thus log Q decreases, so E increases → reaction becomes more feasible.
- Decreasing $[Fe^{3+}]$ increases Q, so E decreases → reaction less feasible.

10. a. The standard electrode potential (E°) is the potential difference between a half-cell and the standard hydrogen electrode when both are under standard conditions (298 K, 1 atm gas pressure, 1 mol dm^{-3} ion concentration). Marking Scheme: Correct definition with all three standard conditions.

b. A redox reaction is feasible under standard conditions if E°cell > 0.

c. Electroplating; corrosion prevention using sacrificial anodes.

d. Electrode potentials are expressed for reduction reactions, so more positive E° means greater tendency to gain electrons.

11. **a.** $E°cell = E°cathode - E°anode = (+0.34) - (-0.76) = +1.10\ V \rightarrow$ Reaction is spontaneous.

b. Order of reactivity:

More negative E° = more reactive metal.

Order: Cu < Fe < Al.

c. Corrosion prevention uses:

Using a more reactive metal as sacrificial anode.

Cathodic protection by applying external potential.

12. Fuel cell analysis:

a. Anode: H_2 oxidation (E° = 0.00 V), Cathode: O_2 reduction (E° = +1.23 V) – 2 marks.

b. $E°cell = 1.23 - 0.00 = +1.23\ V$.

c. Positive E°cell means the reaction is feasible –

d. Sacrificial protection:

A more reactive metal with more negative E° is attached to the metal to be protected. It oxidises preferentially, supplying electrons to the protected metal and preventing its oxidation.

13. **a.** Causes for lower EMF:

- Deviation from standard concentrations.
- Temperature changes.
- Polarisation at electrodes.

b) Nernst equation: $E = E° - (0.0592/n) \log Q$; higher [products] or lower [reactants] reduces EMF.

c. Role in rechargeable batteries:

Electrode potentials are used to select electrode materials with suitable voltage difference, reversibility, and stability. Example: Li-ion battery uses Li^+/Li (−3.04 V) and a strong oxidising cathode like $CoO_2/LiCoO_2$ (~+0.6 V) giving ~3.6 V total.

14. **a.** Electrode potential – The potential difference developed between an electrode and its solution of ions at equilibrium under standard conditions.

b. Standard conditions for measuring standard electrode potential

- Temperature of 298 K (25°C)
- Ion concentration of 1 mol dm^{-3}
- Gas pressure of 1 atm (if gas is involved)

c. Applications of standard electrode potentials

- Predicting feasibility/spontaneity of redox reactions
- Calculating cell potentials (E°_{cell})
- Arranging elements in electrochemical series
- Selecting suitable electrodes for cells

15. a. Reactivity order: Most reactive = most negative E° value

E° values: Mg (–2.37 V), Zn (–0.76 V), Cu (+0.34 V)

Increasing reactivity: Cu < Zn < Mg

b. $E^\circ(Ag^+/Ag) = +0.80$ V, $E^\circ(Cu^{2+}/Cu) = +0.34$ V

$E^\circ cell = 0.80 - 0.34 = +0.46$ V (positive) → feasible; Cu will reduce Ag^+ to Ag

c. $E^\circ(K^+/K) = -2.92$ V, $E^\circ(H^+/H_2) = 0.00$ V

H^+ is reduced before K^+ in aqueous solution; potassium cannot be obtained by electrolysis of aqueous KCl

16. a.

i. Diagram: Correct Zn anode, Cu cathode, salt bridge, electron flow from Zn to Cu

ii. Half-cell reactions:

$Zn \rightarrow Zn^{2+} + 2e^-$ (oxidation)

$Cu^{2+} + 2e^- \rightarrow Cu$ (reduction)

Overall: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$)

iii. $E^\circ cell = (+0.34) - (-0.76) = +1.10$ V

iv. Positive E°cell → spontaneous reaction

b. $E^\circ(Zn^{2+}/Zn) = -0.76$ V, $E^\circ(H^+/H_2) = 0.00$ V

$Zn + 2H^+ \rightarrow Zn^{2+} + H_2$, $E^\circ cell = 0 - (-0.76) = +0.76$ V → feasible

$E^\circ(Cu^{2+}/Cu) = +0.34$ V, $E^\circ cell = 0 - 0.34 = -0.34$ V → not feasible

17. a. $Fe^{3+}/Fe^{2+} = +0.77$ V, $Zn^{2+}/Zn = -0.76$ V

b. $E°cell = 0.77 - (-0.76) = +1.53$ V → feasible

c. Industrial application: Galvanisation, sacrificial anode protection

18. $E°(M^{2+}/M) = -0.44$ V, $E°(Cu^{2+}/Cu) = +0.34$ V

a. $E°cell = 0.34 - (-0.44) = +0.78$ V → M displaces Cu

b. $E°(H^+/H_2) = 0.00$ V, $E°cell = 0 - (-0.44) = +0.44$ V → M liberates H_2 gas

c. Positive E°cell values indicate both reactions are feasible

Part C

Electrolytic Cells and Applications

1. **a.** An electrolytic cell is a device in which electrical energy from an external source (DC supply) is used to drive a non-spontaneous chemical reaction.

b. Anode: the positive electrode where oxidation occurs.

Cathode: the negative electrode where reduction occurs.

c. Electroplating of jewelry and cutlery to prevent corrosion and improve appearance.

d. Examples: Sodium chloride solution (NaCl), copper (II) sulphate solution ($CuSO_4$)

2. **a.** DC ensures a constant flow of electrons in one direction so that oxidation and reduction occur at fixed electrodes. With AC, the polarity changes continuously, so electrode reactions would not occur consistently.

b. Cations migrate towards the cathode and are reduced by gaining electrons.

Anions migrate towards the anode and are oxidised by losing electrons.

c. Extraction of aluminium from molten alumina in the Hall–Héroult process.

Manufacture of chlorine, sodium hydroxide, and hydrogen in the chlor-alkali industry.

d. Diagram should show:

- A container with electrolyte
- Positive electrode (anode) and negative electrode (cathode)
- Cations moving to cathode, anions to anode
- Connected DC supply

3. a. Electrolytic cell: Requires external electrical energy; anode is positive (oxidation), cathode is negative (reduction).

Voltaic cell: Generates electrical energy spontaneously; anode is negative (oxidation), cathode is positive (reduction).

b. At the cathode: Hydrogen gas (H_2) is produced (water is reduced preferentially to sodium ions).

At the anode: Chlorine gas (Cl_2) is produced (Cl^- ions oxidized).

Overall: NaOH remains in solution.

c. Electroplating protects metals from corrosion, improves durability, and enhances appearance. For example, steel can be electroplated with chromium or nickel for corrosion resistance and shine.

d. Graphite is inert, does not react with chlorine gas released at the anode, and withstands high temperatures. Metal electrodes might corrode or react with products.

4. Negative impacts: High energy consumption, greenhouse gas emissions, release of chlorine gas harmful to environment.

Positive impacts: Provides essential materials (aluminium, chlorine, hydrogen, NaOH) used in construction, sanitation, and energy.

5. Setup: Hoffman apparatus with dilute sulfuric acid as electrolyte, DC power supply.

Observation: Hydrogen gas collected at cathode, oxygen at anode.

Connection: Hydrogen can be used as a clean fuel in fuel cells.

6. Examples:

- Large-scale green hydrogen production for energy storage.
- Electrochemical CO_2 reduction to fuels.
- Antimicrobial electroplated coatings for medical instruments.

7. Advantages: Produces very pure metals, supports industries (aerospace, construction).

   Disadvantages: Very costly due to high electricity consumption, dependence on fossil fuels increases production cost.

Part D

Faraday's Laws of Electrolysis

1. **a.** First Law: Mass deposited ∝ charge passed.
   b. Formula: $m = (M \times I \times t) / (n \times F)$.
   c. Faraday constant ≈ 96,500 C mol^{-1}.
   d. Michael Faraday.
   e. Cathode – $Cu^{2+} + 2e^{-} \rightarrow$ Cus..
2. **a.** Longer time increases Q=It; m ∝ Q.
   b. Mass Ag ≈ 4.03 g.
3. **a.** Ag has higher M/n (108 vs 31.8 for Cu); more Ag mass.
   b. Al_2O_3 stable, molten electrolysis needs high energy. (3 marks)
   c. Anode – $O^{2-} \rightarrow O_2$ (or CO_2 with C anodes). (3 marks)
4. Mass Na ≈ 0.429 g. (4 marks)
5. Compare Cu plating vs Au refining, link M/n. (5 marks)
6. Cu^{2+} reduced, no replenishment, solution fades. (3 marks)
7. Use First Law with mass, current, time → F → N_A. (4 marks)
8. Q ≈ 4.47×10^3 C. (3 marks)
9. Al requires large charge Q; cost dominated by energy. (4 marks)
10. Au needs less charge than Al (M/n comparison). (5 marks)
11. Measure m, I, t → calculate F → e=F/N_A. (5 marks)
12. Impacts: energy, emissions, effluents; mitigation. (6 marks)
13. Mass Au ≈ 29.4 kg. (5 marks)

Part E

1. a. Corrosion: gradual deterioration of a metal by chemical/electrochemical reaction with its environment (1).

- Rusting: corrosion of iron/steel forming hydrated Fe (III) oxides

b. Conditions necessary for rusting:

- Presence of water (moisture).
- Presence of oxygen.
- Rusting is accelerated by electrolytes such as salts or acids.

c. Equation for rusting

- Oxidation: $Fe \rightarrow Fe^{2+} + 2e^-$
- Reduction: $O_2 + 4e^- + 2H_2O \rightarrow 4OH^-$
- Overall: $4Fe + 3O_2 + 6H_2O \rightarrow 4Fe(OH)_3$
- On drying, $Fe(OH)_3$ dehydrates to form $Fe_2O_3 \cdot xH_2O$ (rust).

2. a. Methods of preventing rusting

- Painting: forms a physical barrier.
- Oiling/Greasing: prevents contact with moisture and air.
- Galvanisation: zinc coating offers barrier and sacrificial protection.
- Alloying: e.g., stainless steel resists rust due to chromium oxide layer.
- Cathodic protection: using more reactive metal such as magnesium.

b. Why galvanisation is better than painting:

- Paint only acts as a barrier; if scratched, corrosion begins immediately.
- Zinc provides both a barrier and sacrificial protection.

c. Why aluminium resists corrosion:

- Forms a thin, tough Al_2O_3 layer which prevents further attack.

3. a. Reasons

- Sea water contains dissolved salts acting as electrolytes, speeding corrosion.

- Moist, salty air provides continuous water and oxygen.

b. Experiment

- Materials: Iron nails, beakers, tap water, salt water.
- Setup: Cup A – nail in tap water; Cup B – nail in salt water.
- Observation: Nail in salt water rusts faster.
- Conclusion: Salt accelerates rusting by increasing conductivity of water.

4. a. Variables

- Independent: environment/medium (dry air, water, etc.)
- Controlled: same nail type/size, same time/volume/temperature/exposure.
- Adequate experimental control/mention of identical nails.

b. Ranking & reason

- Slow→fast: dry air = boiled+oil (tie, 0) < tap water < salt water.
- Reason: more ions/O_2 availability → higher corrosion current.

c. Boiled water + oil

- Boiling removes dissolved O_2/CO_2; oil layer excludes new O_2/water contact; therefore, no cathodic reduction → no rust).

d. Copper touching iron in salt water

- Iron becomes the anode and corrodes faster near the Cu contact.
- Electrons flow Fe→Cu; O_2 reduced on Cu/nearby cathodic areas.
- Justification mentions galvanic couple/standard potentials.

5. a. Effects of corrosion

- Transportation: weakens car bodies, damages ships, compromises aircraft safety.
- Infrastructure: collapse of bridges, damage to pipelines and buildings.
- Economy: high repair/replacement costs, accidents.

b. Large-scale strategies:

- Painting/coating of structures.
- Galvanisation/alloying/alloying metals.
- Cathodic protection.
- Designing structures to avoid water traps.
- Routine inspection/maintenance.

c. Recommendation for pipelines in Ghana:

- Galvanisation or cathodic protection with sacrificial anodes.
- Justification: Pipelines are exposed to moisture; cathodic protection gives continuous defence.

Section 5

1. a. Physical properties

- Good conductors of heat and electricity.
- High melting and boiling points.
- Hard, dense, and strong metals.

b. Chemical properties

- Exhibit variable oxidation states.
- Form coloured compounds.
- Act as catalysts in many reactions.

c. Reason for construction use

- Strong and durable.
- Resistant to high temperatures.
- Can form alloys for extra strength.

2. a. Why compounds are coloured

- Due to electronic transitions (d–d transitions) in partially filled d-orbitals when light is absorbed.

b. Variable oxidation states

- Transition elements have close energy levels between 3d and 4s orbitals.
- Different numbers of d and s electrons can be lost.

- Main group metals usually lose only outermost s electrons.

c. Examples of catalysts

- Iron in the Haber process (NH_3 production).
- Vanadium(V) oxide (V_2O_5) in the Contact process (H_2SO_4 production).
- Nickel in hydrogenation of oils.

3. a. Experiment setup

- Place $CuSO_4$ solution in a test tube – shows blue colour (coloured solution).
- Add ammonia solution dropwise – deep blue complex forms.
- Place H_2O_2 in a second test tube – slow decomposition (few bubbles).
- Add MnO_2 powder – rapid bubbling, release of oxygen gas.

Observations

- $CuSO_4$ solution is blue; with ammonia forms deep blue $[Cu(NH_3)_4]^{2+}$ complex.
- MnO_2 speeds up decomposition of H_2O_2; vigorous bubbling.
- Glowing splint test confirms O_2.

b. Equations

- $2H_2O_2 \rightarrow 2H_2O + O_2$ (catalysed by MnO_2).
- $[Cu(H_2O)_6]^{2+} + 4NH_3 \rightarrow [Cu(NH_3)_4(H_2O)_2]^{2+} + 4H_2O$.

4. a. Electronic configuration links

- Variable oxidation states: energy difference between 3d and 4s orbitals is small; different electrons can be lost.
- Coloured compounds: d–d electron transitions absorb visible light, giving colours.

b. Comparison with main group metals

- Transition metals: variable oxidation states vs. main group: fixed states.
- Transition metals: coloured compounds vs. main group: mostly colourless compounds.

- Transition metals: act as catalysts vs. main group: rarely catalytic.

c. Applications in Ghana

- Mining: manganese and iron ores used for alloys and export.
- Industry: catalysts in fertilizer production and oil refining.
- Medicine: cobalt in vitamin B_{12}, platinum compounds in drugs.
- Construction: iron and steel for buildings, bridges, and vehicles.

5. a. Similarities

- Both are metals and are shiny (lustrous).
- Both conduct heat and electricity.
- Both are malleable and ductile.
- Both form positive ions (cations).

b. Differences

- Transition metals show variable oxidation states; main group metals usually show fixed oxidation states.
- Transition metals form coloured compounds; main group metals usually form colourless compounds.
- Transition metals often act as catalysts; main group metals rarely do.
- Transition metals are generally harder and denser than main group metals.

c. Uses of transition metals

- Iron and steel in construction (bridges, buildings, vehicles).
- Copper in electrical wiring and plumbing.
- Silver and gold in jewellery and coins.
- Platinum in catalytic converters.

6. a. Why transition metals form coloured compounds:

- Transition metals have partially filled d-orbitals.
- Electrons can move between different d-orbital energy levels by absorbing visible light.

- The colour observed is due to the wavelength of light not absorbed.
- Example: Cu^{2+} gives blue solutions, Ni^{2+} gives green solutions.
- Main group metals usually do not have partially filled d-orbitals and form colourless compounds.

b. Reactivity

- Transition metals are less reactive because they have stronger metallic bonds.
- They have higher ionisation energies than reactive main group metals like Na and K.
- Main group metals (like alkali metals) react vigorously with water and acids, releasing hydrogen gas.

c. Everyday uses

- Construction: Iron and steel used in bridges, buildings, and vehicles.
- Technology: Copper used in wiring, silver in electronics.
- Medicine: Cobalt in vitamin B_{12}; platinum in anti-cancer drugs.

7. a. Completed table

- Oxidation states: Transition metals – variable; Main group metals – usually fixed.
- Compound colour: Transition metals – coloured; Main group metals – colourless.
- Catalytic ability: Transition metals – many act as catalysts; Main group metals – rarely catalytic.
- Hardness and strength: Transition metals – generally hard and strong; Main group metals – often softer.

b. Explanation for industry use

- Variable oxidation states allow transition metals to take part in many redox reactions.
- Coloured compounds and complexes useful in analytical chemistry and manufacturing.
- Catalytic properties make them essential in large-scale industrial processes (e.g., Haber process, Contact process).

- Strength and durability make them suitable for construction and transport.

c. Why aluminium is widely used

- It is abundant and relatively cheap.
- Light but strong when alloyed; useful in transport (aircraft, vehicles).
- Resistant to corrosion due to protective oxide layer.
- Good conductor of electricity, used in power cables.

8. a. Comparison

- Oxidation states: Transition metals have variable states; main group metals usually fixed.
- Compound colours: Transition metals form coloured compounds; main group metals form colourless ones.
- Catalytic activity: Transition metals frequently act as catalysts; main group metals rarely do.
- Applications: Transition metals used in construction, industry, technology, medicine; main group metals often used where lightness/reactivity is useful (e.g., Na in street lamps, Al in transport).

b. Economic importance of transition metals in Ghana

- Mining of manganese, iron, and gold provides export revenue.
- Used in construction industries for infrastructure.
- Catalysts in manufacturing and refining processes.
- Employment opportunities in mining and metallurgy.

c. Choice of transition metals for catalysts

- They provide alternative reaction pathways with lower activation energy.
- They have variable oxidation states, making them suitable for redox reactions.
- Main group metals lack these properties, making them poor catalysts.

- Example: Iron used in Haber process, platinum in catalytic converters.

Section 6

1. a. Esters: –COOR

Amides: $–CONH_2$

Acyl halides: –COCl or –COBr

b. i . General formula of ester: R–COOR′

ii. Used as flavouring agents

Used as solvents

Used in perfumes

2. a. $CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$

b. $CH_3COCl + NH_3 \rightarrow CH_3CONH_2 + HCl$ (2 marks)

c. Acyl halides have a good leaving group (Cl^-/Br^-)

Amides have resonance stabilization which reduces reactivity

Explanation of reactivity difference

3. a. Heat ethanol with ethanoic acid in the presence of concentrated H_2SO_4

Distil off ethyl ethanoate

b. Reaction of carboxylic acids with ammonia

Reaction of acyl halides with ammonia

c. Esters are less soluble

Amides are more soluble

Amides form strong hydrogen bonds with water due to –NH group

4. a. Possible reasons: incomplete reaction, loss during distillation, equilibrium not shifted, evaporation

b. Modifications: use excess reactant, reflux longer, remove water formed, purify product

c. Food/Flavors: esters used in flavourings

Medicine: aspirin (ester) or sedatives (amide)

Industrial: polymers, plastics

Discussion/elaboration

5. **a.** The general molecular formula of a triglyceride is $C_3H_5(OOCR)_3$.

b. The bond that joins glycerol to fatty acids is an ester linkage (–COO–).

c. Fats are generally solid at room temperature, while oils are liquid.

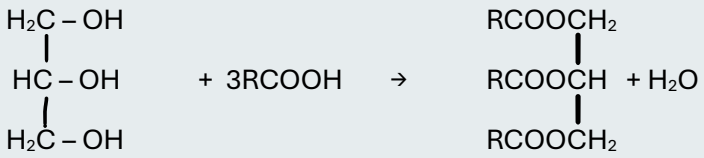
Fats contain more saturated fatty acids, while oils contain more unsaturated fatty acids.

6 (a) Glycerol ($C_3H_5(OH)_3$) reacts with three fatty acids (RCOOH).

The hydroxyl groups of glycerol combine with the carboxyl groups of fatty acids.

This reaction produces one triglyceride ($C_3H_5(OOC–R)_3$) and three water molecules.

Equation: $C_3H_5(OH)_3 + 3RCOOH \rightarrow C_3H_5(OOC–R)_3 + 3H_2O$.



b. Fats mostly contain saturated fatty acids with single carbon–carbon bonds.

These straight hydrocarbon chains pack closely together.

Close packing increases intermolecular forces and raises the melting point, making them solid at room temperature.

Oils contain more unsaturated fatty acids with double bonds.

The double bonds create kinks in the chains, preventing close packing.

This lowers the melting point, so oils remain liquid at room temperature.

7 **a.** Glycerol is a triol with three carbon atoms, each bonded to a hydroxyl (–OH) group.

Fatty acids are long hydrocarbon chains with a terminal carboxyl group (–COOH).

b. Saturated fatty acids contain only single bonds between carbon atoms, giving straight chains.

Unsaturated fatty acids contain one or more carbon–carbon double bonds, giving kinked chains.

c. Fats and oils are insoluble in water but soluble in organic solvents such as ether or chloroform.

They are non-polar molecules and cannot form hydrogen bonds with water.

They dissolve in non-polar organic solvents because of similar intermolecular forces.

8 **a.** As food: Fats and oils provide energy and essential fatty acids (e.g., cooking oil, butter).

In industry: They are used in making soap, cosmetics, and candles (e.g., palm oil in soap-making).

In medicine: They are used in ointments and as sources of omega-3 fatty acids (e.g., fish oil supplements).

b. Saturated fats raise levels of LDL cholesterol in the blood.

This increases the risk of heart disease, stroke, and obesity.

i. Unsaturated oils help reduce LDL cholesterol and increase HDL cholesterol.

ii. They provide essential fatty acids such as omega-3 and omega-6, which support brain and heart health.

Soap and soapless detergents

1. **a.** The functional group in a soap molecule is the carboxylate group ($-COO^-$).

b. The general formula of the hydrophilic head is $-COO^-Na^+$ or $-COO^-K^+$.

c. **i.** Soapy detergents: Sodium stearate, Potassium palmitate.

ii. Soapless detergents: Sodium alkyl sulphate, Sodium alkyl benzene sulphonate.

2. **a.** Hard water contains calcium (Ca^{2+}) and magnesium (Mg^{2+}) ions.

These ions react with soap to form insoluble calcium or magnesium salts (scum).

The scum reduces the availability of soap for lathering and cleaning.

b. Soaps contain a long hydrocarbon chain with a carboxylate head group ($–COO^-Na^+$).

Soapless detergents contain a long hydrocarbon chain with a sulphonate ($–SO_3^-$) or sulphate ($–OSO_3^-$) head group.

Soap molecules are derived from natural fats and oils, while soapless detergents are made from petrochemicals.

3. a. Ashes from burnt plantain peels or cocoa pods are dissolved in water to form alkali solution.

Vegetable oil is heated in a pot, and the alkali solution is gradually added.

The mixture is stirred until it thickens, indicating soap formation.

Salt solution is added to separate soap from glycerol.

The soap is cooled, moulded and packaged.

b. Equation: $CH_2(OOCR)_2CH(OOCR) + 3NaOH \rightarrow CH_2OH–CHOH–CH_2OH + 3RCOONa$.

c. Both soap and soapless detergents have a hydrophilic head and a hydrophobic tail.

The hydrophobic tail attaches to grease or oil on clothes.

The hydrophilic head attaches to water molecules.

This allows grease to be lifted and dispersed in water as micelles.

In hard water, soap forms scum with Ca^{2+} and Mg^{2+} ions, while soapless detergents do not form scum and remain effective.

4. a. Everyday uses of soaps:

- Washing clothes.
- Bathing and personal hygiene.
- Cleaning household utensils.

Everyday uses of soapless detergents:

- Laundry in hard water areas.
- Dishwashing liquids.
- Industrial cleaning products.

b. i. Some synthetic detergents are non-biodegradable and cause water pollution by forming foam on rivers and lakes.

ii. Solutions:

- Produce biodegradable detergents that can break down naturally.
- Use regulations to limit the use of harmful surfactants and encourage eco-friendly alternatives.

Section 7

1. Monomer: small repeating unit. Polymer: large molecule of many repeating units. (1,1)

2. Natural vs synthetic (any 2 differences: source, biodegradability, examples e.g., proteins/starch vs nylon/PE).

3. **a.** **Peptide bond**;

b. **Glycosidic bond** (α in starch/glycogen; β in cellulose). (1,1)

4. Glucose = mono; Sucrose = di; Cellulose = poly; Glycogen = poly.

5. **Condensation** (water eliminated when amino and carboxyl groups react).

6. **Condensation** (water eliminated when glycosidic bond forms).

7. $NH_2CH_2COOH + NH_2CH(CH_3)\ COOH \rightarrow NH_2CH_2CO–NHCH(CH_3)\ COOH + H_2O$.

8. Bonds = 5; water = 5. (1,1)

9. $C_6H_{12}O_6 + C_6H_{12}O_6 \rightarrow C_{12}H_{22}O_{11} + H_2O$.

10. Starch: α-1,4 (and α-1,6 in amylopectin); some branching; energy storage (plants).

Glycogen: α-1,4 & more α-1,6; highly branched; rapid energy store (animals).

Cellulose: β-1,4; unbranched; structural fibre (plant cell walls).

11. **a.** Biuret positive: violet/purple (protein).

b. Iodine positive: blue-black (starch).

c. Benedict's positive: green→orange/red (reducing sugars in ripe banana).

12. Different amino acid sequences (primary structure) → different folding → different active sites/properties. (Any 2–3 correct points).

13. Starch: α-glycosidic bonds; human enzymes (amylases) hydrolyse α-links.

Cellulose: β-1,4 links; humans lack cellulase; H-bonded fibres resist digestion.

14. Repeating unit = $(C_6H_{10}O_5)$; polymer $(C_6H_{10}O_5)_n$; conventional statement: n H_2O eliminated from n glucose in forming the polymer.

15. Proteins: varied 20-amino-acid sequences → high specificity;

Polysaccharides: repeated glucose units; function mainly from linkage/branching pattern rather than sequence variety.

16. Claim false: e.g., cellulose (glucose polymer) is insoluble, not sweet; starch is not sweet; sucrose is sweet but is a disaccharide, not a polymer. Structural reasons (β vs α linkages; crystallinity).

17. Dipeptide hydrolysis: acid (e.g., 6 M HCl, heat) or enzyme protease; maltose: dilute acid + heat or maltase. (Any two correct per item).

18. Clear aim/hypothesis (2); correct reagents & safety (3); stepwise method + control (4); suitable results table (3); justified conclusion (3).

19. **a.** Mass fraction lost as water per unit = 18/180 = 0.10 (10%).

b. Water evolved = 10% of 5.40 g = 0.54 g.

c. Assumption: perfect conversion to internal repeating units; ignores end- groups/side reactions/moisture, etc.

20. **a.** Sugars (12 g) ≈ mono/di; remainder of carb (≈62 g) mostly starch; fibre (5 g) = non- starch polysaccharides (cellulose/hemicellulose).

b. Iodine: + blue-black (starch); Benedict's: + (brick-red/orange) due to reducing sugars (if present; sucrose alone would be −); Biuret: + violet (protein present). (5)

c. Polysaccharides/fibre aid slow energy release/digestive health.

21. **a.** Synthetic polymer = large molecule made by humans through chemical reactions, using monomers.

b. Examples: Polyethene (from ethene), PVC (from chloroethene), Nylon (from diamine + dicarboxylic acid), Terylene (from dicarboxylic acid + diol).

22. Addition: monomers with double bonds join without loss of small molecules. Example: Polyethene from ethene.

- Condensation: monomers with two functional groups join, releasing small molecules like water or HCl. Example: Nylon or polyester formation

23. a. Equation: $nCH_2{=}CH_2 \rightarrow [{-}CH_2{-}CH_2{-}]_n$

b. Properties: light, durable, water-resistant, non-biodegradable, flexible, chemically resistant.

c. Uses: bags, bottles, containers, pipes, tanks, wire coating.

24. a. Monomers: hexamethylenediamine + adipic acid.

b. Equation showing condensation with –COOH and $–NH_2$ groups.

c. Properties: tough, high tensile strength, resistant to wear, resistant to heat, flexible.

d. Uses: ropes, fabrics, carpets, parachutes, tyre cords.

25. a. Environmental problems: blockage of drains → flooding, pollution of rivers/sea, death of animals that eat plastics, breeding grounds for mosquitoes, release of toxic gases when burnt, takes hundreds of years to decompose.

b. Solutions: recycling, bans/restrictions, public education, biodegradable alternatives, improved waste management.

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GLOSSARY

Acid-Base Indicator: A weak acid or base that changes colour depending on the pH of the solution. It exists in two forms—protonated (HIn) and deprotonated (In^-)—each with a distinct colour. Examples: phenolphthalein, methyl orange.

Acidic salt: A salt that makes the solution acidic due to the hydrolysis of its cation (e.g., NH_4^+ from NH_4Cl).

Acyl halides – Very reactive derivatives of carboxylic acids where the –OH group is replaced by a halogen (Cl, Br).

Addition Polymerisation: A type of polymerisation where monomers join directly without the loss of any small molecule.

Alkanoic acids – Organic acids containing the –COOH (carboxyl) group.

Alkanol: An organic compound with one or more hydroxyl (-OH) functional groups attached to an alkane chain; commonly known as alcohols.

Alkyl: A hydrocarbon group derived from an alkane by removing one hydrogen atom, used in various organic compounds.

Amides – Organic compounds formed when the –OH group of a carboxylic acid is replaced by $–NH_2$ or substituted amines.

Amino Acids: The building blocks of proteins that link together to form polypeptides and proteins.

Anhydride: A compound formed from another compound by the removal of water, often used in the production of acids or esters.

Anion hydrolysis: A reaction of a negatively charged ion (from a weak acid) with water to produce OH^-, making the solution basic.

Basic salt: A salt that makes the solution basic due to the hydrolysis of its anion (e.g., CH_3COO^- from CH_3COONa).

Buffer action is the way a buffer keeps the pH steady by neutralizing small amounts of added acid or base.

Buffer Region: A portion of a titration curve where pH changes gradually due to the presence of a weak acid and its conjugate base. It is seen before the equivalence point in titrations involving weak acids or bases.

Buffer solution is solution that resists changes in pH when a small amount of acid or base is added.

Carbohydrates: Polymers of glucose that provide energy in living organisms.

Catalyst – A substance that speeds up a chemical reaction but is not used up in the process (e.g., Fe in the Haber process).

Cation hydrolysis: A reaction of a positively charged ion (from a weak base) with water to produce H_3O^+, making the solution acidic.

Cellulose: A complex carbohydrate that forms the cell walls of plants, providing structural support.

Chitin: A structural carbohydrate found in the cell walls of fungi and the exoskeletons of insects and crustaceans.

Combustion: A chemical reaction that occurs when a substance reacts rapidly with oxygen, producing heat and light, often as flames.

Complex Ion – A central metal ion bonded to surrounding molecules or ions called ligands (e.g., $[Cu\ (NH_3)_4]^{2+}$).

Condensation Polymerisation: A type of polymerisation in which monomers join and release small molecules such as water.

Conductor: A material that allows electricity or heat to flow through it easily, typically metals like copper and aluminium.

Dehydration: A chemical reaction that involves the removal of water, often resulting in the formation of a bond between two molecules.

Derivatives – Compounds obtained from another compound by replacing part of its structure.

Dicarboxylic Acid: An organic acid with two carboxyl (-COOH) groups, which can participate in the formation of polymers.

Diol: An organic compound that contains two hydroxyl (-OH) groups.

Diprotic acid is an acid that can release two hydrogen ions (H^+) per molecule. Example: sulphuric acid (H_2SO_4).

Disaccharide: A sugar formed from two monosaccharides joined together, such as sucrose or lactose.

Ductile: A property of a material that allows it to be stretched or drawn into thin wires without breaking.

Elastomer: A flexible material that can stretch and return to its original shape, often used in rubber products.

Emulsification: The process of mixing two immiscible liquids, such as oil and water, to form a stable mixture called an emulsion.

Equivalence Point: The stage in titration when the moles of titrant added equal the moles of substance in the analyte. It is marked by a sudden pH change or a clear colour change.

Esters – Sweet-smelling organic compounds formed from a carboxylic acid and an alcohol.

Fat: A type of lipid that is solid at room temperature, composed mainly of triglycerides.

Functional Group: A specific group of atoms responsible for the characteristic properties and reactions of a molecule.

Glucose: A simple sugar (monosaccharide) that serves as the building block of carbohydrates.

Glycerol: A colourless, odourless, and sweet-tasting substance that forms the backbone of triglycerides and is used in food and pharmaceuticals.

Glycogen: A polysaccharide that serves as a form of energy storage in animals, made up of glucose molecules.

Glycosidic bond: A type of bond that connects two sugar molecules together in carbohydrates.

Halogen: Any of the elements in Group 17 of the periodic table, including fluorine, chlorine, bromine, iodine, and astatine, known for their reactivity.

Hydration energy: The amount of energy released when one mole of gaseous ions is dissolved in water to form a hydrated ion.

Hydration: The process by which water molecules surround and interact with ions or molecules in solution, stabilizing them.

Hydrogen bonding: A type of weak chemical bond that forms between a hydrogen atom, bonded to a highly electronegative atom, and another electronegative atom.

Hydrogenation: A chemical process that adds hydrogen to unsaturated fats or oils, making them more saturated.

Hydrolysis: A chemical process in which water is used to break down a compound, resulting in the splitting of bonds.

Hydrophobic: A property of molecules that do not interact well with water, often being non-polar and repelling water.

Indicator transition range: The specific pH range over which an indicator changes colour. For instance, phenolphthalein changes from colourless to pink between pH 8.3–10.0.

Ionic product of water (Kw) is the result of multiplying the amounts of H^+ and OH^- ions in water: $Kw = [H^+] [OH^-]$.

Ligand: A molecule or ion that binds to a central metal atom to form a complex, often facilitating chemical reactions.

Lipid: A group of organic compounds, including fats and oils, that are insoluble in water and serve as energy storage and building blocks for cells.

Magnetic: A property of materials that determines their response to magnetic fields; materials can be ferromagnetic, paramagnetic, or diamagnetic.

Main Group Metals – Metals found on the left and right sides of the periodic table (s- and p-block), such as sodium, magnesium, and aluminium, usually with fixed charges.

Malleable – Ability of a metal to be hammered into thin sheets without breaking.

Monomer: A small molecule that can join with others to form a polymer.

Monoprotic acid is an acid that can release one hydrogen ions (H^+) per molecule. Example: hydrochloric acid (HCl)

Monosaccharide: The simplest form of sugar, consisting of a single sugar molecule, such as glucose or fructose.

Neutral salt: A salt that does not undergo hydrolysis in water, typically formed from a strong acid and strong base; its aqueous solution has a pH of 7.

Neutralisation: A chemical reaction in which an acid and a base react to form water and a salt.

Octahedral: A geometric shape with eight faces, commonly used to describe the arrangement of ligands around a central atom in coordination complexes.

Oil: A type of lipid that is liquid at room temperature, also made up of triglycerides.

Oligosaccharide: A carbohydrate made up of a small number (typically 2 to 10) of monosaccharide units.

Orbital: A region in an atom where there is a high probability of finding electrons; these regions can have different shapes (such as s, p, d, or f orbitals).

Oxidation State – The charge on a metal ion, showing how many electrons, it has lost (e.g., Fe^{2+} or Fe^{3+}).

Peptide Bond: A chemical bond that links amino acids together in a protein.

pH is a measure of how acidic a solution is, based on the amount of hydrogen ions present: $pH = -\log[H^+]$.

pH scale is a scale from 0 to 14 used to show how acidic or basic a solution is. $pH < 7$: acidic, $pH = 7$: neutral, $pH > 7$: basic.

pKw is the negative logarithm of the ionic product of water (Kw). At 25°C, pKw = 14.

pOH is a measure of the hydroxide ion (OH^-) concentration in a solution: $pOH = -\log [OH^-]$.

Polyamide: A type of synthetic polymer made from repeating amide groups, commonly found in fabrics like nylon.

Polyester: A synthetic polymer made from ester groups, commonly used in fabrics and bottles.

Polymer: A very large molecule made up of repeating smaller units called monomers.

Polysaccharide: A long chain of sugar molecules bonded together, used for energy storage or structural purposes, such as starch or cellulose.

Protein: A natural polymer made of amino acids, essential for growth and repair of tissues.

Reduction: A chemical reaction that involves the gain of electrons or hydrogen, often resulting in a decrease in oxidationnumber.

Resonance: A concept in chemistry used to describe the delocalization of electrons in molecules that cannot be represented by a single Lewis structure.

Saponification: A chemical reaction that occurs when fats or oils are converted into soap and glycerol by reacting with an alkali.

Saturated: A term used to describe organic compounds that have no double or triple bonds between carbon atoms, meaning they contain the maximum number of hydrogen atoms.

Self-ionisation of water is when water molecules split into hydrogen (H^+) and hydroxide (OH^-) ions naturally: $H_2O \rightleftharpoons H^+ + OH^-$.

Starch: A carbohydrate made of many glucose units, serving as an energy storage in plants.

Subshell: A subdivision of an orbital shell in an atom, characterized by its shape and the types of orbitals it contains (e.g., s, p, d, f).

Synthetic Polymers: Man-made polymers such as polythene, PVC, nylon, and terylene, with diverse industrial and household uses.

Tetrahedral: A geometric shape with four faces, typically used to describe the arrangement of four atoms or groups around a central atom.

Thermoplastic: A type of plastic that can be melted and reshaped multiple times when heated.

Thermoset: A type of plastic that hardens permanently after being shaped and cannot be remelted.

Titration Curve: A graph plotting pH against the volume of titrant added. It illustrates key features like buffer regions and the equivalence point and helps identify the type of titration.

Titration: A method used to determine the unknown concentration of a solution by reacting it with a known volume of another solution, usually an acid or base.

Transesterification: A chemical reaction in which an ester reacts with an alcohol to form a different ester and alcohol, commonly used in biodiesel production.

Transition Elements – Metals found in the middle of the periodic table (Groups 3–12) that often form coloured compounds, variable oxidation states, and complexes.

Triglyceride: A type of fat molecule composed of three fatty acids attached to a glycerol backbone, used for energy storage.

Triprotic acid is an acid that can release three hydrogen ions (H^+) per molecule. Example: phosphoric acid (H_3PO_4).

Unsaturated: A term used to describe organic compounds that contain one or more double or triple bonds between carbon atoms, meaning they have fewer hydrogen atoms than saturated compounds.

This book is intended to be used for the Year Three Chemistry Senior High School (SHS) Curriculum. It contains information and activities to support teachers to deliver the curriculum in the classroom as well as additional exercises to support learners' selfstudy and revision. Learners can use the review questions to assess their understanding and explore concepts and additional content in their own time using the extended reading list provided.

All materials can be accessed electronically from the Ministry of Education's Curriculum Microsite.

[Scan the QR Code to visit the website]

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(GAST)