



MINISTRY OF EDUCATION

Chemistry

for Senior High Schools

TEACHER MANUAL

YEAR THREE



NATIONAL COUNCIL FOR
CURRICULUM & ASSESSMENT
OF MINISTRY OF EDUCATION

MINISTRY OF EDUCATION



REPUBLIC OF GHANA

Chemistry **for Senior High Schools**

Teacher Manual

Year Three



**NATIONAL COUNCIL FOR
CURRICULUM & ASSESSMENT
OF MINISTRY OF EDUCATION**

CHEMISTRY TEACHER MANUAL

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INTRODUCTION

The National Council for Curriculum and Assessment (NaCCA) has developed the third and final set of curriculum materials to support the implementation of the new Senior High School curriculum.

This Year Three Teacher Manual for Chemistry covers all aspects of the content, pedagogy, teaching and learning resources and assessment required to effectively support learners in their final year of Senior High School. All the necessary information for the Student Assessment Portal has been included in the Year Three Teacher Manual to simplify the development of weekly learning plans, meaning that teachers have a single reference document.

The Year Three Teacher Manual has two main objectives:

- To support learners to consolidate and deepen their understanding of the content, pedagogy, and assessment across all three years of the new Senior High School curriculum in preparation for the West Africa Senior Secondary Certificate Examination (WASSCE).
- To equip teachers with the tools and resources needed to deliver revision-focused, assessment-led lessons that build on the knowledge learners have acquired from Year One to Year Three.

In this third year, teachers will be supporting learners to prepare for the WASSCE. Each section of the Teacher Manual is therefore structured around the highest Depth of Knowledge (DoK) levels, allowing learners to build progressively on knowledge acquired from Year One to Year Three. For each week, learning tasks and pedagogy are built around the assessment, ensuring that classroom activities directly prepare learners for the demands of the examination. Two sets of sample examination papers are also provided; one administered by Week 12 as a first semester practice examination, and one by Week 20 as a mock examination, to gauge learners' competencies and identify areas requiring remediation.

To further support teachers in delivering Year Three, NaCCA, working with the Ghana Education Service, has developed a subject specific App for each subject.

The Year Three PLC sessions have been revised in response to findings from the 2025 annual evaluation survey. The survey confirmed that teachers have made strong progress in lesson planning, use of varied instructional methods, ICT integration and modes of assessment. To address the identified gap in explaining concepts clearly using examples familiar to learners, each PLC session now includes a dedicated Content Exploration segment. This ten-minute segment brings the subject group together to work through one challenging problem or concept from the week's Teacher Manual content, not to plan how to teach it, but to ensure that every teacher in the group fully understands it.

The PLC sessions in Year Three also place particular emphasis on connecting the week's content to equivalent topics taught in Years One and Two. Teachers are encouraged to note the recap weeks signposted throughout the Teacher Manual and to use the App to explore how content across all three years links together.

As in previous years, teachers are expected to integrate National Values, Gender Equality and Social Inclusion (GESI), Social and Emotional Learning (SEL), and 21st-century skills into

their learning activities, assessment tasks, and rubrics and mark schemes. The subject-specific App will support teachers in doing this throughout the year.

Ghana Education Service acknowledges the dedication of the teachers who have implemented the new curriculum across Years One and Two and recognises that Year Three represents both the culmination of that effort and a critical moment for learners. The Year Three Teacher Manual and its accompanying resources have been developed to ensure that every learner, regardless of background, ability, or school context, is fully prepared to demonstrate what they know and can do in the WASSCE and to progress confidently to further studies, the world of work, and adult life.

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SECTION 1: pH AND ITS APPLICATIONS TO SOLUTIONS

STRAND 1: PHYSICAL CHEMISTRY

Sub–Strand 2: Equilibria

Learning Outcome: Use the knowledge in pH to distinguish between solutions that are acidic, neutral or basic

Content Standard: Demonstrate knowledge and understanding of the pH scale in determining the acidity and alkalinity of strong and weak acids, bases and aqueous solutions

INTRODUCTION AND SECTION SUMMARY

This section introduces learners to the concepts of pH, pOH, and their applications in acid-base chemistry. Learners will explore the mathematical relationship between ion concentration and pH/pOH, develop skills in calculating these values for various solutions, and understand buffer systems. By the end of this section, learners will be able to interpret the pH scale, perform calculations for strong and weak acids/bases, explain buffer action, and appreciate the significance of pH regulation in biological systems, environmental processes, and industrial applications. This knowledge forms an essential foundation for understanding acid-base equilibria and chemical analysis in real-world contexts.

The weeks covered by this section are:

- Weeks 1 and 2: Explain and undertake calculations on the concepts of pH, and pOH and describe their significance in everyday life.
- Week 3: Explain the incomplete or partial ionisation of weak acids and weak bases and calculate pK_a and pK_b .
- Week 4 and 5: Explain buffer solution, buffer action and their applications in everyday life.

SUMMARY OF PEDAGOGICAL EXEMPLARS

Teachers should employ diverse pedagogical approaches to develop learners' understanding of pH, pOH, and buffer systems. Digital learning should be used to introduce pH and pOH concepts through videos or presentations, providing visual context for these abstract concepts. Teacher-led discussions should explain pH as the negative logarithm of hydrogen ion concentration, deriving the mathematical expression $pH = -\log[H^+]$.

Collaborative learning activities should utilise charts and visual aids to help learners understand the pH scale structure (acidic below 7, neutral at 7, basic above 7). Groups should discuss the ionic product of water and explore relationships between pH, pOH, and pK_w . The problem-solving approach should be employed for calculation practice with strong acids and bases.

For weak acids and bases, talk-for-learning activities should compare and contrast strong and weak electrolytes, explaining ionisation constants (K_a and K_b) and their relationship with pK_a and pK_b . Activity-based learning should focus on calculations involving weak acids and bases. Buffer solutions should be taught through collaborative learning, with mixed-gender groups researching buffer composition, action, and pH calculation using the Henderson-Hasselbalch equation. Experiential learning should involve hands-on preparation of both acidic and basic buffers. Inquiry-based learning should connect buffer concepts to real-world applications in biological systems and industrial processes.

Throughout these activities, teachers should emphasise the practical significance of pH and buffer systems in everyday contexts, including environmental processes, biological systems, and industrial applications.

ASSESSMENT SUMMARY

Indicator	DOK Level	Assessment Strategy and/or Mode
Explain and undertake calculations on the concepts of pH and pOH and describe their significance in everyday life.	Level 3: Strategic Reasoning	Group work
Explain the incomplete or partial ionisation of weak acids and weak bases and calculate pK_a and pK_b .	Level 3: Strategic Reasoning	Questioning
Explain buffer solution, buffer action and their applications in everyday life	Level 4: Extended Thinking	Short Quiz and Group Project work

WEEK 1 & 2

Learning Indicator: Explain and undertake calculations on the concepts of pH, and pOH and describe their significance in everyday life

Recap weeks

Year 1	Week 6: Describe the mole as a unit of the amount of substance. Week 7: Explain the mole concept and its relevance in the preparation of standard solutions.
Year 2	Week 9: Explain Arrhenius, Brønsted-Lowry and Lewis concepts of acids and bases Week 10: Describe the physical and chemical properties of acids and bases.

FOCAL AREA 1: PH AND POH

1. pH and pOH

Understanding whether a substance is acidic or basic is important in many areas of life, from cooking and cleaning to farming and healthcare. One way to describe how acidic or basic a solution is involves using a measurement known as pH. This is a scale that ranges from 0 to 14, showing the acidity or alkalinity of an aqueous solution. The term pH comes from the letter “p,” which represents potential, and “H,” which stands for hydrogen. pH is a measure of the hydrogen potential of a solution. According to Arrhenius and Brønsted-Lowry, acids produce hydrogen ions.

The pH of a substance gives important numerical information about how active the hydrogen ions (H^+) are in a solution. It helps to measure how acidic or basic a substance is, by looking at the balance between hydrogen ions (H^+) and hydroxide ions (OH^-).

If the concentration of H^+ ions is greater than that of OH^- ions, the solution is acidic, and its pH is less than 7.

If the concentration of OH^- ions is greater than that of H^+ ions, the solution is basic, and its pH is greater than 7.

If the concentration of H^+ and OH^- ions are equal, the solution is neutral, with a pH of exactly 7. This means pH is directly related to the ratio of the concentration of Hydrogen ions and hydroxide ions in a solution. It gives us a reliable, scientific way of expressing how acidic or basic a solution is.

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 implies that the higher the hydrogen ion concentration of a solution, the lower the pH value and the lower the hydrogen ion concentration of a solution, the higher its pH. This enables

scientists to conveniently represent very low hydrogen ion concentration in small positive numbers.

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.

$$\text{pOH} = -\log[\text{OH}^-]$$
$$\text{pOH} = \log\left(\frac{1}{[\text{OH}^-]}\right)$$
$$[\text{OH}^-] = 10^{-\text{pOH}}$$

Understanding the Relationships Between pH, pOH and Ion Concentrations

1. Deriving $[\text{H}^+] = 10^{-\text{pH}}$

The definition of pH is:

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

Rearranging:

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

Using the logarithm rule $\log_{10}x = y$ means $x = 10^y$:

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

Therefore:

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

Example: If $\text{pH} = 3$:

$$[\text{H}^+] = 10^{-3} = 0.001 \text{ mol dm}^{-3}$$

2. Deriving $[\text{OH}^-] = 10^{-\text{pOH}}$

The definition of pOH is:

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

Rearranging:

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

Using the logarithm rule $\log_{10}x = y$ means $x = 10^y$:

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

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

Example: If $\text{pOH} = 4$:

$$[\text{OH}^-] = 10^{-4} = 0.0001 \text{ mol dm}^{-3}$$

3. Summary

Quantity	Definition	Rearranged form
Hydrogen ion	$\text{pH} = -\log_{10}[\text{H}^+]$	$[\text{H}^+] = 10^{-\text{pH}}$
Hydroxide ion	$\text{pOH} = -\log_{10}[\text{OH}^-]$	$[\text{OH}^-] = 10^{-\text{pOH}}$

Easy rule to remember: Concentration = $10^{-\text{p-value}}$

Table 1: A table showing the pH and its corresponding concentration of hydrogen and hydroxide ions at 25°C

pH	$[\text{H}^+]$	$[\text{H}^-]$
0	$(10^0) = 1$	0.00000000000001 = (10^{-14})
1	$(10^{-1}) = 0.1$	0.00000000000001 = (10^{-13})
2	$(10^{-2}) = 0.01$	0.00000000000001 = (10^{-12})
3	$(10^{-3}) = 0.001$	0.00000000000001 = (10^{-11})
4	$(10^{-4}) = 0.0001$	0.00000000000001 = (10^{-10})
5	$(10^{-5}) = 0.00001$	0.00000000000001 = (10^{-9})
6	$(10^{-6}) = 0.000001$	0.00000000000001 = (10^{-8})
7	$(10^{-7}) = 0.0000001$	0.00000001 = (10^{-7})
8	$(10^{-8}) = 0.00000001$	0.0000001 = (10^{-6})
9	$(10^{-9}) = 0.000000001$	0.000001 = (10^{-5})
10	$(10^{-10}) = 0.0000000001$	0.0001 = (10^{-4})
11	$(10^{-11}) = 0.00000000001$	0.001 = (10^{-3})
12	$(10^{-12}) = 0.000000000001$	0.01 = (10^{-2})
13	$(10^{-13}) = 0.0000000000001$	0.1 = (10^{-1})
14	$(10^{-14}) = 0.000000000000001$	1 = (10^0)

2. pH scale

The pH scale is a scale that ranges from zero (0) to fourteen (14), which is used to express the extent of acidity or alkalinity of a solution.

On the pH scale a solution that has a pH close to zero (0) is described as very acidic, while a solution with a pH closer to 14 is described as very basic. A solution with a pH of 7 is described as neutral.

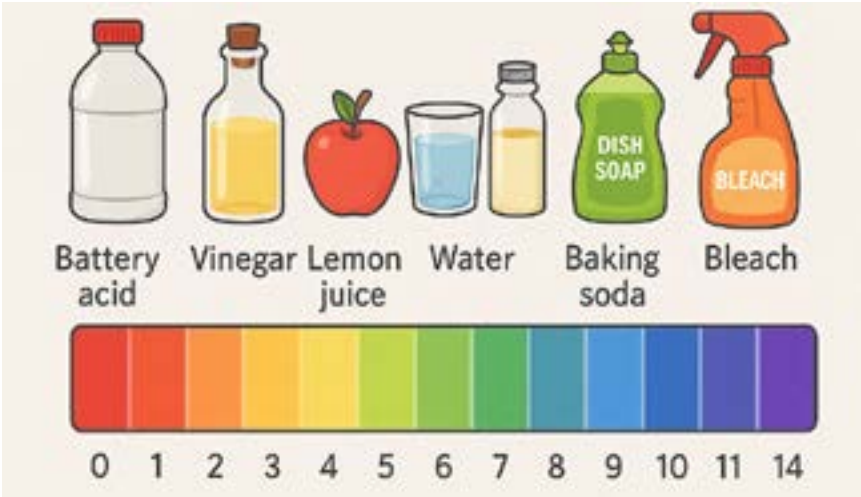


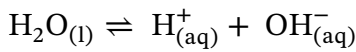
Figure 1: The pH scale showing the pH of household items

FOCAL AREA 2: IONIC PRODUCT OF WATER

1. Meaning of ionic product of water

At the molecular level, water ionises slightly to produce ions, and so it is considered as a weak electrolyte. The ionisation process leads to the formation of hydrogen ions (H⁺) and hydroxide ions (OH⁻). This process is known as the self-ionisation of water.

Ionic product of water, K_w is defined as the product of the molar concentration of the hydrogen ions and hydroxide ions in a solution at a particular temperature.



$$K_c = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$$

$$K_c [\text{H}_2\text{O}] = [\text{H}^+][\text{OH}^-] = K_w$$

K_c [H₂O] is known as the ionic product of water. It is temperature-dependent. At a temperature of 298K;

$$K_w = [\text{H}^+][\text{OH}^-] = 1 \times 10^{-14} \text{ mol}^2\text{dm}^{-6}$$

Considering the self-ionisation of water, [H⁺] = [OH⁻]

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

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

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

$$[\text{H}^+] = 1 \times 10^{-7} \text{ mol/dm}^3$$

For a pure water at 298K, since [H⁺] = [OH⁻]

$$[\text{OH}^-] = 1 \times 10^{-7} \text{ mol/dm}^3$$

The ionic product of water, represented as K_w, is temperature-dependent. As temperature increases, water ionises more, so the value of K_w increases with temperature.

Here are the typical values of K_w at the specified temperatures

Table 1.2: K_w values of water at different temperatures

Temperature/K	Temperature/°C	K _w (mol ² ·dm ⁻⁶)	pK _w = -log K _w
273	0	1.14 × 10 ⁻¹⁵	14.94
293	20	6.81 × 10 ⁻¹⁵	14.17
298	25	1.00 × 10 ⁻¹⁴	14.00
358	85	3.44 × 10 ⁻¹³	12.46
373	100	5.13 × 10 ⁻¹³	12.29

At 298K, there are equal concentration of both hydrogen ions and hydroxide ions in solution. This corresponds to a K_w value of 1.00 × 10⁻¹⁴mol² dm⁻⁶. To maintain the constant, there is an

inverse relationship between [H⁺] and [OH⁻]. When [H⁺] increase, it causes a corresponding decrease in [OH⁻]. Also, when [H⁺] decreases, it causes a corresponding increase in [OH⁻].

2. Relationship between pH, pOH and pK_w

From the ionic product of water,

$$K_w = [\text{H}^+][\text{OH}^-]$$

Taking the negative log of both sides of the equation

$$-\log K_w = -\log([\text{H}^+][\text{OH}^-])$$

Using the law of logarithm, we can expand the right-hand side of the equation as

$$-\log K_w = -\log[\text{H}^+] + -\log[\text{OH}^-]$$

But $-\log K_w = \text{p}K_w$; $-\log [\text{H}^+] = \text{pH}$ and $-\log [\text{OH}^-] = \text{pOH}$

Therefore $\text{p}K_w = \text{pH} + \text{pOH}$

At 298K, the K_w of pure water is taken as 1.00 × 10⁻¹⁴ mol²dm⁻⁶

So, $\text{p}K_w = -\log (1.00 \times 10^{-14}) = 14$

$$\text{pH} + \text{pOH} = 14$$

This relationship helps in calculating pOH when the pH of a solution is known but will only work when the temperature is 298 K. For all other temperatures, the pH + pOH relationship will not equal 14.

Example 1: Calculate the pH of a solution at 298K whose pOH is 10.5.

Solution

$$\text{pH} + \text{pOH} = 14$$

$$\text{pH} = 14 - \text{pOH}$$

If $\text{pOH} = 10.5$

$$\text{pH} = 14 - 10.5$$

$$\text{pH} = 3.5$$

Example 2: Determine the hydrogen ion concentration of a solution at 298K whose pOH is 11.9

Solution

$$\text{pH} + \text{pOH} = 14$$

$$\text{pH} = 14 - \text{pOH}$$

If $\text{pOH} = 11.9$

$$\text{pH} = 14 - 11.9$$

$$\text{pH} = 2.1$$

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

$$[\text{H}^+] = 10^{-2.1}$$

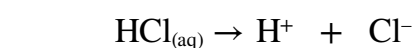
$$[\text{H}^+] = 7.94 \times 10^{-3} \text{ mol/dm}^3$$

3. Calculation of pH and pOH of strong acids and bases

Strong acids dissociate fully or completely in aqueous solution to give out hydrogen ions. Since strong acids dissociate fully in solution, the hydrogen ion concentration is the same as the concentration of monoprotic strong acids (e.g. HCl and HNO₃).

What is the pH of 0.1 mol/dm³ HCl?

From the dissociation of HCl solution, which can be represented as



Initial	0.1	0.0	0.0
---------	-----	-----	-----

Final	0.0	0.1	0.1
-------	-----	-----	-----

$$\text{pH} = -\log (0.1)$$

$$= 1.0$$

Therefore, the pH of 0.1 mol/dm³ HCl is 1.0

This calculation is only valid when the concentration of the monoprotic acid is above the [H⁺] of pure water at 298K (i.e $>1.0 \times 10^{-7}$). In this case the contribution of [H⁺] due to the dissociation of water is negligible and so not considered in calculating the pH.

However, when the concentration of the acid is lower than [H⁺] of pure water at 298K (i.e $<1.0 \times 10^{-7}$), the contribution of the [H⁺] due to the dissociation of water is significant and so it is considered in the calculation of the pH.

Example:

Explain why the pH of $1.0 \times 10^{-8} \text{ mol/dm}^3$ of HCl is not 8.

Solution:

The pH is not 8 because for highly diluted acids, the [H⁺] that determines the pH of the acid is not obtained from only the acid. Water dissociates slightly and contributes extra H⁺ ions. So the [H⁺] from both the dissociation of the acid and water must be considered in the calculating the pH



Initial	10^{-8}	0	0
---------	-----------	---	---

Final	0	10^{-8}	
-------	---	-----------	--



Initial	10^{-8}	0	0
---------	-----------	---	---

Final	0	10^{-8}	
-------	---	-----------	--



At 298K		10^{-7}	10^{-7}
---------	--	-----------	-----------

$$\text{Total } [\text{H}^+] = 10^{-8} + 10^{-7} = 1.1 \times 10^{-7}$$

$$\text{pH} = -\log [\text{H}^+]$$

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

$$\text{pH} = 6.96$$

For a diprotic strong acid like H₂SO₄, 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.

Example:

What is the pH of 0.01mol/dm³ of H₂SO₄ solution?



$$[\text{H}_2\text{SO}_4] = 0.01\text{M}$$

$$[\text{H}^+] = 2 \times 0.01 = 0.02\text{M}$$

$$\text{pH} = -\log 0.02 = 1.70$$

pOH of strong bases

Strong bases also dissociate fully in solution to give hydroxide ions in solution. For monoprotic strong bases the concentration of the hydroxide ion that dissociate in solution is the same as the initial concentration of the base. Examples of monoprotic strong bases are NaOH and KOH

Determine the pOH of 0.2 mol/dm³ NaOH



Initial	0.2	0.0	0.0
---------	-----	-----	-----

Final	0.0	0.2	0.2
-------	-----	-----	-----

$$\text{pOH} = -\log (0.2)$$

$$= 0.7$$

This calculation is only valid when the concentration of the monoprotic base is above the [OH⁻] of pure water at 298K (i.e $>1.0 \times 10^{-7}$). In this instance, the contribution of [OH⁻] due to the dissociation of water is negligible and so not considered in calculating the pOH of the solution.

However, when the concentration of the base is lower than [OH⁻] of pure water at 298K (i.e $<1.0 \times 10^{-7}$), the contribution of the [OH⁻] due to the dissociation of water is significant and so it is considered in the calculation of the pOH.

Example:

Determine the pH of $3.40 \times 10^{-8} \text{ mol/dm}^3$ KOH

Solution:

Bases dissociate in solution to give OH⁻ ions according to the equation



Initial	3.4×10^{-8}	0.0	0.0
---------	----------------------	-----	-----

Final	0.0	3.4×10^{-8}	
-------	-----	----------------------	--



At 298K	10^{-7}	10^{-7}
---------	-----------	-----------

$$\text{Total } [\text{OH}^-] = (3.40 \times 10^{-8}) + (1.00 \times 10^{-7})$$

$$= 1.34 \times 10^{-7}$$
$$\text{pOH} = -\log[\text{OH}^-]$$
$$\text{pOH} = -\log(1.34 \times 10^{-7})$$
$$\text{pOH} = 6.87$$

But $\text{pH} = 14 - \text{pOH}$

$$\text{pH} = 14 - 6.87$$
$$\text{pH} = 7.13$$

4. Significance of the values of pH in everyday life

The values of pH play a big role in many everyday activities and decisions. Knowing the pH of a substance helps to keep living organisms safe, protect the environment and improve the quality of products.

1. Agriculture

- a. Soil pH affects plant growth and crop yield.
- b. Some plants grow best in slightly acidic soil, while others need neutral or alkaline conditions.
- c. Farmers test and adjust soil pH using lime (to increase pH) or sulphur (to lower pH).

2. Medicine and the Human Body

- a. Human blood has a pH of about 7.4. A small change can be dangerous.
- b. Stomach acid (pH 1–2) helps in digestion; antacids are used to reduce excess acidity.
- c. pH is important in urine tests, saliva, and other body fluids to diagnose health conditions.

3. Food Industry

- a. The pH of food affects taste, shelf life, and safety.
- b. Vinegar (acidic) is used to preserve foods by killing harmful microbes.
- c. Baking involves controlling pH for proper rising and texture.

4. Water Treatment

- a. Drinking water must be neutral (pH around 7) to be safe.
- b. Acidic or basic water can corrode pipes or harm aquatic life.
- c. pH is carefully adjusted during water purification.

5. Cleaning Products

- a. Different cleaning agents work best at different pH levels.
- b. Acidic cleaners (e.g. toilet cleaners) remove limescale.
- c. Alkaline cleaners (e.g. bleach, soap) break down grease and stains.

Learning Tasks

- 1. Recap how to write balanced equations for neutralisation reactions taking into consideration the Arrhenius concepts of acids and bases.
- 2. Work in pairs to complete the table below, showing all steps involved in the calculations

pH	[H ⁺] / mol dm ⁻³	pOH	[OH ⁻] / mol dm ⁻³	Acidic/Basic/Neutral
4.5				
		8.4		
	1 × 10 ⁻²			
8.3				
			1 × 10 ⁻⁷	

- 3. Work in pairs to determine the hydrogen ion concentration and pH of solutions made by dissolving the following quantities in water to make 500 cm³ solutions:
 - a. 8.0 g of HCl and
 - b. 4.0 g of NaOH.
- 4. Using textbooks and the internet, research and explain the importance of pH in three areas of daily life, analysing how understanding pH helps us make better decisions about our health, environment, and household practices.

Pedagogical Exemplars

Digital Learning Approach

- 1. Begin the lesson by recapping the concept of acids and bases, asking learners to recall the Arrhenius and Brønsted-Lowry definitions and give examples of common acids and bases from everyday life.
- 2. Second, introduce the learning objectives, explaining that learners will understand how the concentration of hydrogen ions relates to acidity and how this is expressed using the pH scale.
- 3. Show a carefully selected video (5-7 minutes) that explains the concept of pH and pOH ensuring it includes visual representations of the logarithmic scale and real-world applications. (https://youtube/LS67vSI0O5Y?si=NK4s9E_xa5jK8D-i)
- 4. Pause the video at key points to check understanding and emphasise important concepts, particularly the mathematical relationship between hydrogen ion concentration and pH
- 5. After the video, conduct a brief question-and-answer session to assess initial comprehension, addressing any misconceptions immediately.
- 6. Provide learners with a digital or printed worksheet containing simple pH and pOH calculation problems to apply what they have observed in the video.

- Review the worksheet answers as a class, focusing on the mathematical process and common calculation errors.

Talk for Learning Approach

- Begin the lesson by recapping the concept of ion concentration in solutions, reminding learners about molarity and how ions form when acids or bases dissolve in water.
- Write the term “pH” on the board and ask learners what they already know about it, creating a mind map of their prior knowledge.
- Explain that pH is a mathematical way to express the concentration of hydrogen ions in a solution, introducing the formula $\text{pH} = -\log[\text{H}^+]$ and carefully explaining each component.
- Demonstrate how to convert between hydrogen ion concentration and pH using examples on the board, showing both the conversion from concentration to pH and from pH to concentration.
- Use a think-aloud strategy to work through several examples of increasing difficulty, verbalising your thought process for each step.
- Invite learners to suggest hydrogen ion concentrations, then calculate the corresponding pH values together as a class.
- Extend the discussion to pOH, explaining its relationship to pH ($\text{pH} + \text{pOH} = 14$ at 25°C) and demonstrating calculations involving both values.
- Conclude with a discussion of why scientists use the pH scale rather than simply stating hydrogen ion concentration, emphasising the practicality of using smaller numbers and the logarithmic nature of acidity effects.

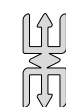
Collaborative Learning Approach

- Begin the lesson by recapping the relationship between hydrogen ion concentration and pH, asking volunteers to explain the formula $\text{pH} = -\log[\text{H}^+]$ in their own words.
- Divide the class into mixed-ability groups of 3-4 learners, ensuring gender balance and inclusion of different learning styles.
- Provide each group with a large pH scale diagram, coloured markers, and cards containing various substances (lemon juice, blood, milk, seawater, etc.) with their pH values.
- Instruct groups to complete their pH scale diagrams by:
 - Clearly marking the acidic (below 7), neutral (exactly 7), and basic (above 7) regions
 - Placing the substance cards at their correct positions on the scale
 - Colour-coding the different regions (e.g., red for acidic, green for neutral, blue for basic)
- Circulate among the groups, asking probing questions about why certain substances have their particular pH values and how this relates to their chemical composition.
- Have each group add to their diagrams by writing the approximate hydrogen ion concentration at various points on the scale, demonstrating the logarithmic relationship.

- Ask groups to prepare a brief presentation explaining one section of the pH scale (acidic, neutral, or basic), including examples of substances in that range and their significance in everyday life.
- Facilitate group presentations, encouraging questions from other learners and highlighting connections between the pH values and the properties of the substances.
- Conclude the lesson with a whole-class discussion about how understanding pH helps us in various contexts, from cooking and gardening to medicine and environmental monitoring.

Key Assessment Level 3: Group Work

- Rainwater collected near Tema industrial area has a pH of 4.5. You may assume the temperature of this rainwater is 298K. As a student researcher working with the Ghana Environmental Protection Agency (EPA),
 - Describe the nature of this rainwater, based on its pH value and explain what this pH level indicates about the chemical composition of the rain.
 - Calculate the hydrogen ion concentration in this rainwater.
 - Compare this to the hydrogen ion concentration in normal rain (pH 5.6) collected from the Kakum National Park.
 - Analyse how this difference affects **any two** different Ghanaian ecosystems: (eg The Volta Lake fisheries, Cocoa farms in the Western Region and The Atewa Forest Reserve) (10 marks)
- You are given three colourless solutions with pH values of 2.5, 7.0, and 11.5. You may assume the temperature of these solutions is 298K.
 - Calculate the $[\text{OH}^-]$ concentration in each solution
 - Predict what would happen when each solution is mixed with a carbonate compound
 - Design a series of tests to identify each solution using locally available materials
 - Explain how the relative concentration of H^+ and OH^- ions influence the properties of each solution. (10 marks)
- The pH of water from a borehole in a village near Tamale was measured to be 5.8. You may assume the temperature of this water is 298K.
 - Calculate the hydrogen ion concentration $[\text{H}^+]$ of this water.
 - Calculate the hydroxide ion concentration $[\text{OH}^-]$ of this water.
 - The community treats this water by filtering it through a mixture of sand and wood ash. Explain the chemical principles behind this treatment method and why it works to make the water safer. (10 marks)



Note

For PLC Sessions 1 and 2 please use the same Instruction Template but choose any assessment strategy of your choice for the other week.

PLC Session 1 & 2

45 minutes	Planning & Reviewing the Learning Plan
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1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week’s concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Plan the week’s key assessment with your app
 - a. The key assessment for the week is **group work**. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the group work (NTS 3k-3q).

10 minutes	Content Exploration
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5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
 - a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step
 - d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
 - a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
 - a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

WEEK 3

Learning Indicator: Explain the incomplete or partial ionisation of weak acids and weak bases and calculate pK_a and pK_b

Recap weeks

Year 1	Week 6: Describe the mole as a unit of the amount of substance. Week 7: Explain the mole concept and its relevance in the preparation of standard solutions.
Year 2	Week 7: Deduce expressions for equilibrium constants in terms of concentrations and partial pressures and perform relevant calculations from appropriate data Week 9: Explain Arrhenius, Brønsted–Lowry and Lewis concepts of acids and bases Week 10: Describe the physical and chemical properties of acids and bases. Week 11: Distinguish between strong and weak acids, strong and weak bases in terms of the extent of dissociation, rate of reaction, electrical conductivity, pH value in aqueous solution and enthalpy change of neutralisation value

FOCAL AREA 1: PK_A AND PK_B

1. Ionisation constant K_a for weak acids

Weak acids are compounds that dissociate partially in aqueous solution to produce hydrogen ions and a negative ion (conjugate base). This means that not all the acid molecules break to form the ions. The dissociation of weak acids is a reversible process that reaches a state of equilibrium where both the acid and its ions exist together in solution.

Unlike strong acids, which dissociate completely in solution, weak acids dissociate partially to produce fewer hydrogen ions which makes them less acidic. The extent to which a weak acid dissociates is measured by its ionisation constant (K_a).

General chemical equation for the dissociation of monoprotic weak acid (HA) is given by:



The equilibrium expression for this reaction is given by

$$K = \frac{[H^+][A^-]}{[HA]}$$

K is the dissociation constant for a weak acid, also known as the acid dissociation constant and written as K_a . It is not affected by changes in concentration, but it is however affected by changes in temperature. Higher K_a values indicates that the weak acid dissociates more in solution and so it corresponds to a stronger weak acid. K_a is only used for weak acids, because strong acids dissociate completely in solution.

Consider a monoprotic weak acid HA, with an initial concentration of “C” mol/dm³. If “x” mol/dm³ of it dissociates at a certain temperature, then the equilibrium expression can be written as:

Equilibrium reaction	$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$		
Initial concentration	C	0	0
Final concentration	C – x	x	x

$$K_a = \frac{[H^+][A^-]}{[HA]}$$
$$K_a = \frac{(x)(x)}{C-x}$$

Since this is a weak acid, the amount of ions that will dissociate is far lower than the initial concentration of the acid. (i.e., $C \gg \gg \gg x$)

Hence $C - x = C$

Therefore, the expression can be written as

$$K_a = \frac{(x)(x)}{C}$$
$$K_a = \frac{(x)^2}{C}$$
$$(x)^2 = K_a \times C$$
$$(x) = \sqrt{K_a \times C}$$

(x) represents the concentration of hydrogen ions that dissociates from a monoprotic acid

This expression is used to determine the concentration of hydrogen ions that dissociate in a solution of a weak acid.

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

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

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

But $-\log[H^+] = \text{pH}$ and $-\log K_a = \text{p}K_a$

$$\text{pH} = \frac{1}{2}\text{p}K_a - \frac{1}{2}\log C$$

Table 3.1: Table showing the K_a and $\text{p}K_a$ values of monoprotic weak acids at a temperature of 298K.

Weak acid	IUPAC Name	Formula	K_a value	$\text{p}K_a$
Phenol	Phenol	C_6H_5OH	1.6×10^{-10}	9.8
Propanoic acid	Propanoic acid	CH_3CH_2COOH	1.3×10^{-5}	4.9
Acetic acid	Ethanoic acid	CH_3COOH	1.8×10^{-5}	4.7
Benzoic acid	Benzenecarboxylic acid	CH_6CH_5COOH	6.4×10^{-5}	4.2
Uric acid		$C_5H_4N_4O_3$	1.3×10^{-4}	3.9
Formic acid	Methanoic acid	$HCOOH$	1.8×10^{-4}	3.7
Nitrous acid	Dioxonitrate(III) acid	HNO_2	4.0×10^{-4}	3.4
Lactic acid	2-Hydroxypropanoic acid	$CH_3CH(OH)COOH$	8.3×10^{-4}	3.1
Hydrofluoric acid	Hydrofluoric acid	HF	6.3×10^{-4}	3.2

Polyprotic weak acids dissociate through more than one step. The number of steps a polyprotic acid will dissociate depends on the number of replaceable hydrogen atoms that are present. For these acids each of the dissociation step has a corresponding dissociation constant.

Table 3.2: A table showing K_a and pK_a values of polyprotic weak acid at a temperature of 298 K

Polyprotic Weak acid	IUPAC Name	Formula	K_a value	pK_a
Carbonic acid	Trioxocarbonate(IV) acid	H_2CO_3	4.3×10^{-7}	6.4
			5.6×10^{-11}	10.3
Phosphoric acid	Tetraoxophosphate(V) acid	H_3PO_4	7.5×10^{-3}	2.1
			6.2×10^{-8}	7.2
			4.8×10^{-13}	12.3

Generally, weak acids have K_a values less than 1.0×10^{-2} in 0.1 mol/dm^3 .

Example;

At a temperature of 298K, 0.02 mol/dm^3 of a weak acid has a dissociation constant of 1.8×10^{-5} . Determine its

- a. Hydrogen ion concentration
- b. pH

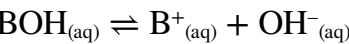
Solution:

- a. From the dissociation expression of weak acids, $[H^+] = \sqrt{K_a \times C}$
If $K_a = 1.8 \times 10^{-5}$
 $C = 0.02 \text{ mol/dm}^3$
 $[H^+] = \sqrt{(1.8 \times 10^{-5}) \times 0.02}$
 $[H^+] = \sqrt{3.6 \times 10^{-7}}$
 $[H^+] = 6.0 \times 10^{-4} \text{ mol/dm}^3$
- b. $pH = -\log[H^+]$
 $pH = -\log(6.0 \times 10^{-4})$
 $pH = 3.2$

2. Ionisation constant K_b for weak bases

Weak bases, on the other hand, are compounds that dissociate partially in solution to produce hydroxide ions and their corresponding cations. The presence of hydroxide ions makes the solution alkaline. At equilibrium, an amount of the undissolved base coexists with the ions formed from its dissociation.

The general chemical equation for the dissociation of a monoprotic weak base (BOH) is given by:



The equilibrium expression for this reaction is given by

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

The equilibrium constant, K_b at a specific temperature is known as the base dissociation constant. Comparing the K_b values of different weak bases, the one with a higher K_b value indicates a stronger base and hence dissociates more.

Consider a monoprotic weak base BOH, with an initial concentration of “C” mol/dm^3 . If “y” mol/dm^3 of it dissociates at a given temperature, then the equilibrium expression can be written as:

Equilibrium reaction	$BOH_{(aq)} \rightleftharpoons B^+_{(aq)} + OH^-_{(aq)}$		
Initial concentration	C	0	0
Final concentration	$C - y$	y	y

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

$K_b = \frac{(y)(y)}{C - y}$

For weak bases, the amount of ions that will dissociate is far lower than the initial concentration of the acid. (ie $C \gg y$)

Hence $C - y = C$

Therefore, the expression can be written as

$K_b = \frac{(y)(y)}{C}$

$K_b = \frac{(y)^2}{C}$

$(y)^2 = K_b \times C$

$y = \sqrt{K_b \times C}$

(y) represents the concentration of hydroxide ions that dissociates from a monoprotic basic

The expression is useful in calculating the concentration of hydroxide ions that dissociate in a solution of a weak base.

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

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

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

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

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

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

Table 3.3: A table showing the K_b and pK_b values of weak bases

Weak base	IUPAC Name	Formula	K_b value	pK_b
Ammonia	Ammonia	NH_3	1.8×10^{-5}	4.7
Methylamine	Methylamine	CH_3NH_2	4.0×10^{-4}	3.4
Ethylamine	Ethylamine	$C_2H_5NH_2$	5.6×10^{-4}	3.3

Pyridine	Pyridine	C ₅ H ₅ N	1.7 × 10 ⁻⁹	8.8
Aniline	Aminobenzene	C ₆ H ₅ NH ₂	4.3 × 10 ⁻¹⁰	9.4

Example:

A weak monoprotic base has a K_b value of 1.74×10^{-5} . If its concentration is 0.1 mol/dm^3 , determine:

- a. pK_b
- b. hydroxide ion concentration

Solution:

- a. $pK_b = -\log K_b$
If $K_b = 1.74 \times 10^{-5}$
Then $K_b = -\log (1.74 \times 10^{-5})$
 $= 4.76$
- b. $[OH^-] = \sqrt{K_b \times C}$
 $[OH^-] = \sqrt{(1.74 \times 10^{-5}) \times 0.1}$
 $[OH^-] = \sqrt{(1.74 \times 10^{-6})}$
 $[OH^-] = 1.32 \times 10^{-3} \text{ mol/dm}^3$

Learning Tasks

- 1. Recap the strength of acids and bases taking into consideration that weak acids and weak bases dissociate partially in aqueous solution to produce few ions
- 2. Write the equilibrium expression for the dissociation of weak acids and use it to derive the ionisation constant, K_a
- 3. Derive an expression to determine the pH of weak acids when the ionisation constant and the initial concentration of the weak acid are known.
- 4. Convert K_a values to pK_a using the expression $pK_a = -\log (K_a)$
 - a. $K_a = 1.8 \times 10^{-5}$ (ethanoic acid)
 - b. $K_a = 7.5 \times 10^{-4}$ (nitrous acid)
 - c. $K_a = 6.3 \times 10^{-10}$ (hydrogen cyanide)

Pedagogical Exemplars

Talk for Learning Approach

- 1. Begin the lesson by recapping the concept of acids and bases, asking learners to recall the definitions and provide examples of common acids and bases they encounter in daily life.
- 2. Introduce the learning objectives, explaining that learners will understand the difference between strong and weak acids/bases and how ionisation constants relate to acid/base strength.

- 3. Create a table on the board with four columns: “Strong Acids,” “Weak Acids,” “Strong Bases,” and “Weak Bases,” and invite learners to suggest examples for each category.
- 4. Use a think-pair-share strategy where learners first individually consider the differences between strong and weak acids, then discuss with a partner, and finally share their ideas with the class.
- 5. Build on learners’ contributions to explain that the key difference lies in the extent of ionisation in aqueous solution, using visual representations to show:
 - a. Strong acids (like HCl) ionise completely:
 $HCl \rightarrow H^+ + Cl^-$ (100%)
 - b. Weak acids (like CH₃COOH) ionise partially:
 $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$ (1-5%)
- 6. Repeat the think-pair-share process for strong versus weak bases, emphasising that:
 - a. Strong bases (like NaOH) dissociate completely:
 $NaOH \rightarrow Na^+ + OH^-$ (100%)
 - b. Weak bases (like NH₃) ionise partially:
 $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ (1-5%)
- 7. Introduce the concept of ionisation constants (K_a and K_b) as a measure of the extent of ionisation, writing the general expressions:
 - a. For weak acids: $K_a = \frac{[H^+][A^-]}{[HA]}$
 - b. For weak bases: $K_b = \frac{[BH^+][OH^-]}{[B]}$
- 8. Explain that K_a and K_b values are typically very small numbers, which is why we often use pK_a and pK_b instead, where: $pK_a = -\log(K_a)$ and $pK_b = -\log(K_b)$
- 9. Demonstrate with examples how larger K_a values indicate stronger acids (more ionisation), while smaller K_a values indicate weaker acids (less ionisation), and similarly for bases with K_b .
- 10. Facilitate a discussion about the relationship between K_a and K_b for conjugate acid-base pairs, introducing the equation $K_a \times K_b = K_w$ and explaining its significance.

Activity-based Learning Approach

- 1. Begin the lesson by recapping the concepts of K_a , K_b , pK_a , and pK_b , asking volunteers to explain these terms and their relationship to acid/base strength.
- 2. Provide learners with a handout containing a reference table of common weak acids and bases with their K_a and K_b values, ensuring the information is clearly presented.
- 3. Demonstrate the step-by-step process for calculating the pH of a weak acid solution using the following approach:
 - a. Write the ionisation equation
 - b. Set up the K_a expression

- c. Make appropriate approximations and solve for $[H^+]$, calculate $pH = -\log[H^+]$
4. Work through a complete example on the board, such as calculating the pH of 0.1 mol/dm^3 acetic acid ($K_a = 1.8 \times 10^{-5}$), verbalising your thought process at each step.
5. Similarly demonstrate the calculation of pOH and pH for a weak base solution, such as 0.1 mol/dm^3 ammonia ($K_b = 1.8 \times 10^{-5}$).
6. Show how to convert between K_a and pK_a (and K_b and pK_b) using examples, emphasising that this is simply an application of the logarithm function.
7. Divide the class into pairs or small groups and distribute a worksheet with problems of increasing difficulty:
 - a. Level 1: Calculate pK_a or pK_b given K_a or K_b values.
 - b. Level 2: Calculate pH of weak acid solutions given concentration and K_a
 - c. Level 3: Calculate pH of weak base solutions given concentration and K_b
 - d. Level 4: Calculate the concentration of a weak acid needed to achieve a specific pH
8. Circulate among the groups as they work, providing guidance and asking probing questions to check understanding and challenge thinking.
9. After sufficient time, review the solutions to the problems as a class, inviting different groups to explain their approach and reasoning for each problem.
10. Conclude with a brief application exercise where learners must determine whether a solution with a given pH would be better buffered by adding a weak acid with a specific pK_a value or a different weak acid with another pK_a value, explaining their reasoning.

Key Assessment Level 3: Questioning

Where necessary assume the temperature of these solutions is 298K

1. Which statement correctly explains why a 0.1 mol/dm^3 solution of HCl has a lower pH than a 0.1 mol/dm^3 solution of CH_3COOH ?
 - A. HCl has a higher molecular weight than acetic acid
 - B. HCl completely ionises in water while acetic acid only partially ionises
 - C. HCl is a binary compound while acetic acid contains carbon
 - D. The chloride ion is a stronger conjugate base than the acetate ion
2. A 0.1 mol/dm^3 solution of a weak acid HA has a pH of 3.5. What percentage of the acid molecules have ionised?
 - A. 3.500%
 - B. 0.100%
 - C. 0.316%
 - D. 0.035%
3. If the K_a of a weak acid is 4.5×10^{-4} , what is its pK_a value?
 - A. 3.35

- B. 3.65
- C. 4.5×10^{-4}
- D. $-\log(4.5 \times 10^{-4})$

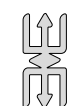
4. A 0.2 mol/dm^3 solution of a weak acid has a pH of 2.8. What is the K_a value for this acid?
 - A. 2.0×10^{-3}
 - B. 3.2×10^{-3}
 - C. 1.3×10^{-5}
 - D. 6.3×10^{-6}
5. If a weak base B has a pK_b of 4.75, what is the pH of a 0.15 mol/dm^3 solution of this base?
 - A. 4.75
 - B. 9.25
 - C. 10.17
 - D. 11.21

PLC Session 3

45 minutes

Planning & Reviewing the Learning Plan

1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week's concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week's content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week's concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Plan the week’s key assessment with your app
- a. The key assessment for the week is **questioning**. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the assessment (NTS 3k-3q).

10 minutes	Content Exploration
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5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
- a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step
 - d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
- a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).

9. Remember to
- a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

WEEK 4

Learning Indicator: Explain buffer solution, buffer action and their applications in everyday life

Recap weeks

Year 1	Week 15: Explain the term solubility and describe the factors that affect the solubility of substances
Year 2	Week 6: Explain the terms reversible and irreversible reactions and dynamic equilibrium. Week 9: Explain Arrhenius, Brønsted–Lowry and Lewis concepts of acids and bases Week 11: Distinguish between strong and weak acids, strong and weak bases in terms of the extent of dissociation, rate of reaction, electrical conductivity, pH value in aqueous solution and enthalpy change of neutralisation value.

FOCAL AREA 1: BUFFER SOLUTIONS AND BUFFER ACTIONS

1. Buffer solutions and their composition

Introduction

In everyday language, a “buffer” is something that absorbs shock or minimises impact. Think about how a cushion buffers you from the hardness of a wooden chair, or how a car’s bumper buffers the impact during a collision.

In chemistry, buffer solutions work in a similar way - they absorb the “shock” of added acids or bases, preventing dramatic changes in pH. Buffer solutions are an important concept in chemistry that help us understand how certain systems maintain pH stability.

A buffer solution is a mixture that resists changes in pH when small amounts of acid or base are added to it, or when dilution occurs.

Buffer Composition

A buffer solution consists of:

1. A weak acid and its conjugate base (acidic buffer), e.g., ethanoic acid (CH_3COOH) and sodium ethanoate (CH_3COO^-)
2. A weak base and its conjugate acid (basic buffer), e.g., ammonia (NH_3) and ammonium chloride (NH_4Cl)

Characteristics of buffer solutions

1. Maintains a relatively constant pH
2. Contains components that can neutralise added H^+ or OH^- ions
3. Most effective when concentrations of weak acid/base and conjugate are approximately equal
4. Works best when pH is within ± 1 unit of the pK_a of the weak acid (or pK_b of the weak base)

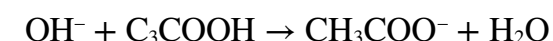
Buffer action

The true power of a buffer lies in its ability to respond to challenges to the system’s pH.

1. **Acidic buffer:** For an acidic buffer (e.g., $\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$), when acid is added, the conjugate base neutralises (absorbs) the added H^+ ions, converting them to the weak acid form, which has much less effect on pH than free H^+ ions would:

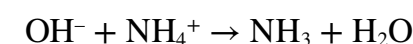


When a base is added, the weak acid neutralises (absorbs) the added OH^- ions, converting them to the conjugate base form and water, preventing the OH^- from significantly raising the pH:



2. **Basic buffer:** For a basic buffer (e.g., $\text{NH}_3/\text{NH}_4^+$), when an acid is added the weak base neutralise the added H^+ ions: $\text{H}^+ + \text{NH}_3 \rightarrow \text{NH}_4^+$

When a base is added, the conjugate acid neutralises the added OH^- ions



Real-life analogies of buffer action

1. Temperature regulation

Our body’s temperature regulation system works similarly to a buffer. When we get too hot, we sweat to cool down. When we get too cold, we shiver to warm up. These responses help maintain our body temperature within a narrow range, just as a buffer maintains pH within a narrow range.

The conjugate base is like sweating, it responds to heat (acid) by absorbing it.

The weak acid is like shivering - it responds to cold (base) by generating more heat (acid).

2. The Cooking analogy

Think about cooking jollof rice, that you accidentally add too much salt, you might add a potato to absorb some of the excess salt or add more rice to dilute it. The potato or additional rice act as “buffers” against the excess salt, maintaining a more balanced flavour and taste.

Similarly, in a buffer solution if you add acid (like adding too much salt), the conjugate base component absorbs it (like the potato absorbs salt)

If you add base (like adding too much water), the acid component counteracts it (like adding more seasoning).

2. Calculation of pH of a buffer solution

Calculation of pH of a buffer solution

The pH of a buffer solution can be calculated using the Henderson-Hasselbalch equation which is derived from the equilibrium expression for a weak acid.

$$\text{For an acidic buffer, } \text{pH} = \text{pK}_a + \log\left(\frac{[\text{conjugate base}]}{[\text{weak acid}]}\right)$$

Where pK_a is the negative logarithm of the acid dissociation constant (K_a)

$$\text{For a basic buffer, } \text{pOH} = \text{pK}_b + \log\left(\frac{[\text{conjugate acid}]}{[\text{weak base}]}\right)$$

Where pK_b is the negative logarithm of the base dissociation constant (K_b)

How to Use the Henderson-Hasselbalch Equation

To calculate the pH of a buffer solution

1. Determine which one is the weak acid (or base) and which is the conjugate base (or acid).
2. Look up the K_a of the weak acid (for an acidic buffer) or K_b for a basic buffer and use it to calculate for the pK_a or pK_b .
3. Calculate the molar concentrations of both components in the buffer solution.
4. Substitute the values into the Henderson-Hasselbalch equation.
5. Solve the equation to find the pH value.

Worked Examples

1. Calculate the pH of a buffer solution containing 0.15 mol/dm³ ethanoic acid (CH₃COOH) and 0.10 mol/dm³ sodium ethanoate (CH₃COONa).

The K_a of ethanoic acid is 1.74×10^{-5}

Solution

- a. Identify components: Weak acid (CH₃COOH) and conjugate base (CH₃COO⁻)
- b. Calculate pK_a : $pK_a = -\log(K_a)$
 $= -\log(1.74 \times 10^{-5})$
 $= 4.76$
- c. [CH₃COOH] = 0.15 mol/dm³ and [CH₃COO⁻] = 0.10 mol/dm³
- d. $pH = pK_a + \log\left(\frac{[CH_3COO^-]}{[CH_3COOH]}\right)$
- e. $pH = 4.76 + \log\left(\frac{[0.10]}{[0.15]}\right)$
 $= 4.76 + \log(0.667)$
 $= 4.76 - 0.18$
 $= 4.58$

The pH of the buffer is 4.58, which is slightly acidic.

2. Calculate the pH of a buffer solution containing 0.12 mol/dm³ methylamine (CH₃NH₂) and 0.18 mol/dm³ methylammonium chloride (CH₃NH₃Cl). The K_b of methylamine is 4.4×10^{-4} .

Solution:

- a. Identify components: Weak base (CH₃NH₂) and conjugate acid (CH₃NH₃⁺)
- b. Calculate pK_b : $pK_b = -\log(K_b)$
 $= -\log(4.4 \times 10^{-4})$
 $= 3.36$
- c. [CH₃NH₂] = 0.12 mol/dm³ and [CH₃NH₃⁺] = 0.18 mol/dm³

- d. Calculate pOH: $pOH = pK_b + \log\left(\frac{[CH_3NH_3^+]}{[CH_3NH_2]}\right)$
 $pOH = 3.36 + \log\left(\frac{0.18}{0.12}\right)$
 $pOH = 3.36 + \log(1.5)$
 $pOH = 3.36 + 0.18$
 $pOH = 3.54$
- e. Convert to pH: $pH = 14 - pOH$
 $pH = 14 - 3.54$
 $pH = 10.46$
 $\therefore$ The pH of this buffer is 10.46, which is basic.

3. Calculate the pH of a buffer solution prepared by mixing 0.20 mol/dm³ ammonia (NH₃) and 0.15 mol/dm³ ammonium chloride (NH₄Cl). The K_b of ammonia is 1.78×10^{-5} .

Solution

- a. Identify components: Weak base (NH₃) and conjugate acid (NH₄⁺)
- b. Calculate pK_b : $pK_b = -\log(K_b)$
 $= -\log(1.78 \times 10^{-5})$
 $= 4.75$
- c. [NH₃] = 0.20 mol/dm³ and [NH₄⁺] = 0.15 mol/dm³
 Calculate pOH: $pOH = pK_b + \log\left(\frac{[NH_4^+]}{[NH_3]}\right)$
 $pOH = 4.75 + \log\left(\frac{0.15}{0.20}\right)$
 $pOH = 4.75 + \log(0.75)$
 $pOH = 4.75 - 0.12$
 $pOH = 4.63$
- d. Convert to pH: $pH = 14 - pOH$
 $pH = 14 - 4.63$
 $pH = 9.37$
 $\therefore$ The pH of this buffer is 9.37, which is basic.

Challenge Question

4. A research team at the Noguchi Memorial Institute for Medical Research in Accra is developing a buffer solution for preserving biological samples collected from various regions of Ghana. They need a buffer with pH 4.5 for preserving certain enzymes found in local medicinal plants.

The team decides to use an ethanoic acid/sodium ethanoate buffer system. They have available 0.20 mol/dm³ ethanoic acid solution and solid sodium ethanoate.

- a. Calculate the ratio of [CH₃COO⁻] to [CH₃COOH] needed to achieve a pH of 4.5.
- b. What mass of sodium ethanoate (CH₃COONa) should be dissolved in 250 cm³ of the 0.20 mol/dm³ ethanoic acid solution to prepare this buffer?

- c. A field researcher accidentally adds 5.0 cm^3 of 0.10 mol/dm^3 HCl to 100 cm^3 of the prepared buffer while processing samples collected from the Kakum National Park. Calculate the new pH of the buffer solution.
[K_a of ethanoic acid is 1.74×10^{-5} and molar mass of $\text{CH}_3\text{COONa} = 82.03 \text{ g/mol}$]

Learning Tasks

- A weak acid HA has a K_a value of 1.8×10^{-5} and a concentration of 0.1 M .
 - Write the expression for the acid dissociation constant K_a
 - Calculate the hydrogen ion concentration $[\text{H}^+]$ at equilibrium.
 - Calculate the pH of this solution.
- What is a buffer solution? Explain the essential components of a buffer and how it maintains a relatively constant pH.
- How does the ratio of [conjugate base]/[weak acid] affect the pH of a buffer solution?
Use the Henderson-Hasselbalch equation to support your explanation.
 - Consider a buffer solution containing 0.15 mol/dm^3 CH_3COOH and 0.10 mol/dm^3 CH_3COONa .
 - Calculate the initial pH of this buffer.
 - If 0.005 moles of HCl are added to 100 cm^3 of this buffer, calculate the new concentrations of CH_3COOH and CH_3COONa .
 - Calculate the new pH after adding the HCl.
Calculate the change in pH (ΔpH).

Pedagogical Exemplars

Collaborative Learning

- Divide class into mixed-gender groups where applicable of four to five learners.
- Assign each group either topic:
 - buffer solutions and behaviour
 - pH calculations.
- Distribute role cards (researcher, note-taker, timekeeper, presenter).
- Establish ground rules for respectful collaboration, emphasising that all learners must:
 - Listen attentively to each member's contributions
 - Value diverse perspectives and approaches
 - Provide constructive feedback rather than criticism
 - Support group members who may be struggling
- Provide each group with reference materials: textbooks, printed articles, approved online resources.

- Groups research their assigned topics:
 - Group A: Buffer composition, examples, and mechanism of action
 - Group B: Henderson-Hasselbalch equation, calculating component concentrations
- Learners take notes individually before sharing with group.
- Teacher moves around to monitor group dynamics, ensuring equitable participation and addressing any social or emotional concerns that arise.
- Groups compile information, prepare presentation materials and present using preferred mode.
- Encourage supportive behaviour during presentations, with audience members showing respect through attentive listening.
- Teacher highlights key points, corrects any misconceptions, and acknowledges positive collaborative behaviours observed.
- Learners complete a quick exit ticket answering:
 - What is a buffer solution?
 - How does a buffer maintain pH?
 - How do you calculate the pH of a buffer?
 - One thing they appreciated about their group's collaboration

Key Assessment Level 3: Group Project

- Calculate the pH of the following solutions:
 - 0.01 mol/dm^3 HCl
 - 0.001 mol/dm^3 NaOH
 - A weak acid HA with concentration 0.1 mol/dm^3 and $K_a = 1.8 \times 10^{-5}$
(10 marks)
- The acetate buffer system consists of acetic acid (CH_3COOH) and sodium acetate (CH_3COONa). The $\text{p}K_a$ of acetic acid is 4.76 at 25°C .
 - Calculate the pH of a buffer solution containing 0.10 mol/dm^3 CH_3COOH and 0.10 mol/dm^3 CH_3COONa .
 - Calculate the pH of the following acetate buffer solutions:
 - 0.15 mol/dm^3 CH_3COOH and 0.10 mol/dm^3 CH_3COONa .
 - 0.10 mol/dm^3 CH_3COOH and 0.15 mol/dm^3 CH_3COONa .
 - Explain how the ratio of $[\text{CH}_3\text{COONa}]$ to $[\text{CH}_3\text{COOH}]$ affects the pH of the buffer solution with reference to the Henderson-Hasselbalch equation.
 - If 0.005 moles of HCl were added to 100 cm^3 of each buffer solution in (b), which would experience the smallest change in pH? Justify your answer with appropriate calculations and chemical equations showing the buffer action.
(20 marks)

3. A phosphate buffer was prepared by mixing 75.0 cm³ of 0.20 moldm⁻³ KH₂PO₄ with 25.0 cm³ of 0.20 moldm⁻³ K₂HPO₄.
- a. Calculate the pH of this buffer solution.
 - b. Explain what would happen to the pH if a small amount of HCl was added to this buffer. [pK_a of H₂PO₄⁻ is 7.21.]



Hint

Assign Group Project to learners and give them reasonable time to complete it. Use the SS App to generate the project as well as the rubric. Learners' projects should be marked immediately and scores entered into the SAP.

PLC Session 4

45 minutes	Planning & Reviewing the Learning Plan
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- Begin the session by recapping lessons from the previous session, focusing on how learning was applied.
- Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week's concepts in a practical way that will make it easier for learners to understand.
 - b. identify three common misconceptions learners may have on the week's content and how to address them using contexts familiar to Ghanaian learners (NTS 3a, 3g and 3m).
- Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week's concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q).
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

- Use your app to plan and administer the mode of assessment for the Student Assessment Portal (SAP).
 - a. The key assessment for the week is **group project**. Chat with your app to plan this assessment by reviewing or developing the assessment items/tasks and marking schemes/rubrics for the group project (NTS 3k-3q).

10 minutes	Content Exploration
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- As a subject group, select one problem or topic/idea/term in the week's content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
 - a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through the problem individually for 3–4 minutes
 - c. the group then works through it together the problem, step by step
 - d. highlight the most common errors made at each step (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not 'How do I teach this?' but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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- Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
 - a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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- Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
- Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
- Remember to
 - a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

WEEK 5

Learning Indicator: Explain buffer solution, buffer action and their applications in everyday life

Recap weeks

Year 1	Week 15: Explain the term solubility and describe the factors that affect the solubility of substances.
Year 2	Week 6: Explain the terms reversible and irreversible reactions and dynamic equilibrium Week 9: Explain Arrhenius, Brønsted-Lowry and Lewis concepts of acids and bases Week 11: Distinguish between strong and weak acids, strong and weak bases in terms of the extent of dissociation, rate of reaction, electrical conductivity, pH value in aqueous solution and enthalpy change of neutralisation value.
Year 3	Week 1-2: Explain and undertake calculations on the concepts of pH, and pOH and describe their significance in everyday life Week 3: Explain the incomplete or partial ionisation of weak acids and weak bases and calculate pK_a and pK_b .

FOCAL AREA 2: PREPARATION OF BUFFER SOLUTIONS

A buffer solution is a chemical system that maintains a relatively constant pH even when small amounts of acid or base are added. This resistance to pH change is a critical property that makes buffers essential in numerous scientific processes and everyday applications.

Buffer solutions consist of either a weak acid and its conjugate base (acidic buffer) or a weak base and its conjugate acid (basic buffer). This carefully balanced partnership allows the solution to neutralise additional H^+ or OH^- ions that might otherwise disrupt the pH equilibrium.

The common types of buffers are the weak acid buffer and weak base buffer. The Henderson-Hasselbalch equation is used in the preparation of buffer solutions. The knowledge of the dissociation constant of the weak acid or weak base and the pH of the buffer helps in determining the ratio of the concentration of the components of the buffer.

1. Preparation of weak acid buffer

A weak acid buffer is made by mixing a weak acid (e.g., ethanoic acid) with the salt of its conjugate base (sodium ethanoate). The quantity of the components that will be required in the preparation process is derived using the Henderson-Hasselbalch equation to calculate the ratio of the concentration of the weak acid to the concentration of the conjugate base (salt).

$$pH = pK_a + \log\left(\frac{[Salt]}{[Acid]}\right)$$

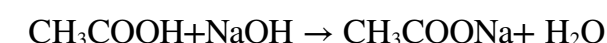
$$\log\left(\frac{[Salt]}{[Acid]}\right) = pH - pK_a$$

$$\frac{[Salt]}{[Acid]} = 10^{(pH - pK_a)}$$

This can be used to determine the ratio of the salt to the acid when the desired pH is known, and the K_a of the weak acid is also known. Once the ratio is known, the buffer can be prepared using three main methods:

- Partial neutralisation of the weak acid with a strong base (such as sodium hydroxide), where part of the acid is converted into the salt of the weak acid (conjugate base), creating the buffer. This requires that fewer moles of the strong base is added. If equal moles of the weak acid and strong base are used, the acid will be completely neutralised and cannot form a buffer. In other words, excess of the weak acid neutralizes a limiting quantity of the strong base. At the end of the reaction, all the strong base would be completely used up and converted to the salt (conjugate base) with the remaining weak acid.

Consider the chemical equation:

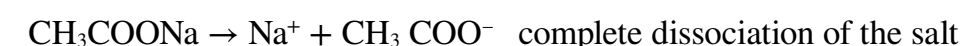
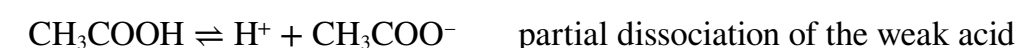


The limiting amount of the strong base is then converted to the salt of the weak acid, leaving a portion of the weak acid unreacted in the reaction vessel. This creates a buffer solution with the weak acid and its salt, which are present in the solution.

- Dissolving a known mass of the salt into a solution of the weak acid, giving a mixture of both components in known concentrations.

In this method, either: a known mass of the salt (conjugate base) is added to a solution of the weak acid, or a known mass of the weak acid is added to a solution of its salt. This gives a buffer with known concentration of both components

The reaction equations can be represented as;



These reactions result in the buffer mixture of CH_3COOH and CH_3COO^-

This gives a buffer with known concentrations of both components.

- Mixing separate solutions of the weak acid and its salt in calculated volumes to achieve the desired concentration ratio.

This is a quick and common method: two solutions are prepared, one containing the weak acid, and the other containing the salt of the conjugate base. These are then mixed in the right volumes to get the desired buffer. Depending on the volume of buffer to be prepared, a certain volume of a standard concentration of the weak acid can be prepared and mixed with another volume of a standard concentration of the salt of the weak acid. Both components of the buffer will co-exist in solution.

2. Preparation of weak base buffer

A weak base buffer solution is prepared by mixing a weak base with the salt of its conjugate acid. Just like weak acid buffers, weak base buffers help resist changes in pH when small amounts of acid or base are added. These types of buffers are useful in chemical and biological systems where a basic (alkaline) environment must be maintained.

To calculate how much of each component (weak base and its salt) is needed, the Henderson-Hasselbalch equation is used, but in terms of pOH instead of pH:

Henderson-Hasselbalch Equation (for bases)

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

This equation can be rearranged as:

$$\log\left(\frac{[\text{Salt}]}{[\text{Base}]}\right) = \text{pOH} - \text{pK}_b$$

$$\frac{[\text{Salt}]}{[\text{Acid}]} = 10^{(\text{pOH} - \text{pK}_b)}$$

This equation is used to determine the ratio of the concentration of the conjugate acid to the concentration of the weak base that is needed for the buffer. When the ratio is derived, it tells how much of the salt and the base are needed to prepare a buffer solution of a desired pOH.

Methods of Preparing Weak Base Buffers

a. Partial Neutralisation of a Weak Base with a Strong Acid

A strong acid (e.g., HCl) is added gradually to an excess of a weak base (e.g., ammonia). The acid reacts (completely neutralised) with part of the base to form the conjugate acid (ammonium ion).



This creates a buffer with both NH_3 and NH_4^+ present.

b. Dissolving a Known Mass of Salt in a Solution of Weak Base

A measured mass of the salt (e.g., NH_4Cl) is dissolved in a prepared solution of ammonia. The concentration and the volume of the ammonia solution is used to determine the number of moles for ammonia. The required moles of the salt is then determined by multiplying the buffer ratio by the moles of ammonia. The number of moles together with the molar mass of the salt is used to find the mass of the salt ($m = nM$). Measure and pour the required volume of ammonia solution into a beaker or flask. Measure the calculated mass of the salt with a chemical balance. Dissolve the salt in the weak base solution and stir until the salt dissolves completely. Transfer the resulting solution into a standard volumetric flask. Add enough distilled water to the required volume. Cork it and shake gently to ensure a uniform buffer solution.

c. Mixing Two Solutions (Weak Base and Salt)

Separate solutions of the weak base and its conjugate acid salt are mixed in the ratio calculated from the Henderson-Hasselbalch equation.

Example: Mix solutions of $0.1 \text{ mol/dm}^3 \text{ NH}_3$ and $0.1 \text{ mol/dm}^3 \text{ NH}_4\text{Cl}$ in appropriate volumes.

3. Applications of buffer in everyday life

The significance of buffers extends far beyond the chemistry laboratory. In biological systems, buffers are vital for maintaining life itself - human blood contains carbonic acid-bicarbonate buffer ($\text{H}_2\text{CO}_3/\text{HCO}_3^-$) and phosphate buffer ($\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$) systems that keep blood pH within the narrow range of 7.35-7.45, essential for proper enzyme function and cellular processes. Without this buffering capacity, even small changes in blood pH could prove fatal.

In consumer products, buffers ensure consistency and effectiveness. Shampoos maintain a specific pH to protect hair and skin, while food products rely on buffers to control acidity for both taste and preservation. Industrial processes, pharmaceutical preparations, and environmental systems all depend on buffer solutions to maintain optimal conditions.

Learning Tasks

- Recap the use of the Henderson-Hasselbalch equation in calculating for the pH of buffer systems
- Outline the steps in the preparation of acetate buffer and list any five (5) apparatus that could be used for the preparation.
 - Determine the ratio of acid to salt using the Henderson-Hasselbalch equation
i.e., $\text{pH} = \text{pK}_a + \log\left(\frac{[\text{Salt}]}{[\text{Acid}]}\right)$
 - Calculate the concentrations of acid and salt
 - Calculate the amount in moles needed for 500 cm^3 solution
 - Explain the effect on buffer capacity if the concentration of either component is changed.
- Preparation of Ammonia-Ammonium Chloride (Weak base) Buffer
 - Determine the pOH and pK_b
From the relation $\text{pH} + \text{pOH} = 14$
And $\text{pK}_b = -\log(K_b)$
 - Determine the ratio of the concentration of weak base to the concentration of the conjugate acid using Henderson-Hasselbalch equation
 - Calculate the concentrations of NH_3 and NH_4^+
 - Calculate the amounts in moles needed for 750 cm^3 solution

Pedagogical Exemplars

Experiential Learning Approach

- First, begin the lesson by recapping the concept of weak acids, weak bases, and their ionisation, asking learners to explain how weak acids only partially dissociate in solution and how this relates to the Henderson-Hasselbalch equation.
- Introduce the learning objectives, explaining that learners will prepare buffer solutions of specific pH values and understand how buffers resist changes in pH when small amounts of acid or base are added.
- Divide the class into mixed-ability groups of 4-5 learners, ensuring each group includes learners with different academic strengths and that roles are assigned to promote equal participation (e.g., materials manager, procedure coordinator, data recorder, etc.).
- Provide each group with the necessary materials:
Weak acid (e.g., acetic acid solution), Salt of the weak acid (e.g., sodium acetate), Weak base (e.g., ammonium hydroxide), Salt of the weak base (e.g., ammonium chloride), pH meter or pH paper, Beakers, measuring cylinders, glass rods, Distilled water, Safety equipment (goggles, gloves).
- Review safety procedures for handling chemicals and using laboratory equipment, emphasising proper disposal methods and the importance of accurate measurements.
- Provide each group with a specific pH target (e.g., pH 4.8, 5.2, 9.3) and the Henderson-Hasselbalch equation:

- a. For acidic buffers: $\text{pH} = \text{pK}_a + \log\left(\frac{[\text{Salt}]}{[\text{Acid}]}\right)$
- b. For basic buffers: $\text{pH} = 14 - \text{pK}_b + \log\left(\frac{[\text{Base}]}{[\text{Salt}]}\right)$
7. Guide groups to calculate the required ratio of acid to salt (or base to salt) needed to achieve their target pH, providing pK_a or pK_b values for their weak acid or base.
8. Have groups prepare their buffer solutions following these steps: Calculate volumes needed based on available concentrations, measure components accurately, mix thoroughly, measure the pH of the resulting solution, make adjustments if necessary to achieve the target pH
9. Once groups have successfully prepared their buffers, have them test the buffer action by:
 - a. Taking three samples of their buffer solution
 - b. Adding a few drops of dilute HCl to one sample
 - c. Adding a few drops of dilute NaOH to another sample
 - d. Leaving the third sample as a control
 - e. Measuring and recording the pH of all three samples
10. Instruct groups to compare their results with other groups that prepared different buffers, discussing which buffers showed the greatest resistance to pH change and why.
11. Facilitate a whole-class discussion about the results, emphasising how the buffer capacity depends on the concentrations of the components and how closely the target pH is to the pK_a of the weak acid (or pK_b of the weak base).

Inquiry-based Learning Approach

1. Begin the lesson by recapping the experiential learning activity, asking learners to explain what buffer solutions are, how they work, and what they observed about buffer action in their laboratory work.
2. Introduce the inquiry question: “Why are buffer systems essential in living organisms and industrial processes?” and explain that learners will investigate real-world applications of buffer systems.
3. Show a brief video or presentation highlighting how pH affects enzyme activity, demonstrating that most enzymes function optimally within narrow pH ranges. (https://youtu.be/_wkuUwA4bdA?si=pfmRjMKp0NnYpmHA)
4. Divide the class into mixed-ability research groups and assign each group a specific buffer system to investigate from one of these categories:
 - a. Biological buffers: blood, intracellular fluid, saliva, etc.
 - b. Industrial buffers: food production, pharmaceuticals, agriculture, etc.
5. Provide research guidelines asking learners to identify:
 - a. The components of their assigned buffer system
 - b. The optimal pH range maintained by the buffer
 - c. What would happen if the buffer system failed?

- d. How the buffer system relates to the Henderson-Hasselbalch equation?
6. Allow groups time to conduct research using provided resources (textbooks, approved websites, scientific journals) and prepare a brief presentation.
7. Facilitate group presentations, with each group explaining their buffer system and its importance, encouraging questions from other learners.
8. After all presentations, lead a comparative discussion highlighting:
 - a. Common features across different buffer systems
 - b. How buffer composition relates to the required pH range
 - c. The consequences of pH disruption in different systems
 - d. How understanding buffer chemistry has led to innovations in medicine and industry
9. Challenge learners to consider a scenario where they must design a buffer system for a specific application (e.g., a sports drink that maintains optimal pH for electrolyte absorption, a soil amendment that stabilises pH for specific crops, or a pharmaceutical formulation).
10. Conclude with a reflective discussion connecting the laboratory experience of preparing buffers to the real-world applications they have investigated, emphasising how theoretical chemistry principles translate to practical solutions for biological and industrial challenges.

NB

1. Throughout both approaches, integrate GESI considerations by ensuring mixed-gender and mixed-ability groupings, providing additional support for learners who need it, and using diverse examples that represent various cultural contexts.
2. Develop SEL skills through collaborative problem-solving, perseverance with challenging calculations, and building confidence through successful experimental work.
3. Build 21st century competencies by connecting buffer chemistry to global challenges in healthcare, food security, and environmental management.

Key Assessment Level 4: Quiz

1. Calculate the pH of a buffer solution prepared by mixing 35.0 cm³ of 0.15 mol/dm³ propanoic acid with 15.0 cm³ of 0.20 mol/dm³ sodium propanoate. Explain how this buffer would respond if 5.0 cm³ of 0.10 mol/dm³ HCl were added, showing all relevant calculations and chemical equations.
 $[\text{K}_a (\text{propanoic acid}) = 1.3 \times 10^{-5}]$ (10 marks)
2. A laboratory technician needs to prepare 500 cm³ of an acetate buffer with pH 4.85 using acetic acid and sodium acetate. Calculate the masses of glacial acetic acid and anhydrous sodium acetate needed to prepare this buffer if the total buffer concentration ($[\text{acid}] + [\text{salt}]$) must be 0.20 mol/dm³. Explain how the buffer capacity would change if the total concentration were doubled while maintaining the same ratio of acid to salt.
 $[\text{C} = 12, \text{H} = 1, \text{O} = 16, \text{Na} = 23, \text{K}_a (\text{acetic acid}) = 1.8 \times 10^{-5}]$ (10 marks)
3. A biochemist needs to prepare 750 cm³ of an ammonia-ammonium chloride buffer with pH 9.30 for enzyme studies. Given that ammonia has a K_b of 1.8×10^{-5} , calculate the masses of ammonium chloride (molar mass 53.49 g/mol) and the volume of 2.0 mol/dm³ ammonia solution needed to prepare this buffer if the total buffer concentration ($[\text{NH}_3] +$

[NH₄⁺]) must be 0.25 mol/dm³. Analyse how the choice of this buffer system rather than a phosphate buffer would affect enzyme stability in the reaction mixture. (10 marks)

PLC Session 5

45 minutes	Planning & Reviewing the Learning Plan
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1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week’s concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



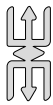
Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Plan the week’s key assessment with your app
 - a. The key assessment for the week is **quiz**. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the quiz (NTS 3k-3q).

10 minutes	Content Exploration
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5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
 - a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step
 - d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
 - a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
 - a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

Section 1 Review

Section 1 Review: pH and Its Applications to Solutions

This section focused on understanding pH, pOH, and buffer systems. During Weeks 1–5, learners have explored the mathematical basis of pH and pOH, developed calculation skills for strong and weak acids/bases, and investigated buffer systems and their applications.

Key Concepts Covered:

pH and pOH Fundamentals: Learners should understand that $\text{pH} = -\log[\text{H}^+]$ and $\text{pOH} = -\log[\text{OH}^-]$, and be able to interpret the pH scale with values below 7 indicating acidic solutions, exactly 7 indicating neutral solutions, and above 7 indicating basic solutions.

Ionic Product of Water: Learners should understand that $\text{pH} + \text{pOH} = 14$ (at 25°C) based on $K_w = 1 \times 10^{-14}$, and be able to calculate pH and pOH for strong acids and bases.

Weak Acids and Bases: Learners should understand the concept of partial ionisation, be familiar with ionisation constants (K_a and K_b), and be able to calculate $\text{p}K_a$, $\text{p}K_b$, pH, and pOH for weak acids and bases.

Buffer Solutions: Learners should understand buffer composition (weak acid/conjugate base or weak base/conjugate acid), explain buffer action (resistance to pH change), calculate buffer pH using the Henderson-Hasselbalch equation, and prepare buffers of specific pH values.

Real-World Applications: Learners should appreciate the significance of pH and buffers in contexts such as environmental processes (acid rain, water quality), biological systems (blood pH regulation, enzyme function), and industrial applications (pharmaceuticals, food production).

MARKING SCHEMES FOR THE KEY ASSESSMENTS

WEEK 1 & 2

Rainwater Analysis (10 marks)

Nature of rainwater based on pH (2 marks)

Expected Answer:

The rainwater with pH 4.5 is acidic (1 mark)

This pH indicates the presence of dissolved acidic oxides such as sulphur dioxide (SO_2) and nitrogen oxides (NO_x) from industrial emissions, which form sulphuric acid and nitric acid in water (1 mark)

Hydrogen ion concentration calculation (2 marks)

Expected Working: $[\text{H}^+] = 10^{-4.5}$

Expected Answer: $[\text{H}^+] = 3.16 \times 10^{-5} \text{ mol/dm}^3$ (2 marks)

Correct formula (1 mark)

Correct calculation (1 mark)

Comparison with normal rain (3 marks)

Expected Working: For normal rain (pH 5.6):

$[\text{H}^+] = 10^{-5.6} = 2.51 \times 10^{-6} \text{ mol/dm}^3$

Comparison ratio: $\frac{3.16 \times 10^{-5}}{2.51 \times 10^{-6}} = 12.6$

Expected Answer: Normal rain $[\text{H}^+] = 2.51 \times 10^{-6} \text{ mol/dm}^3$ (1 mark)

The rainwater from Tema has approximately 12.6 times higher hydrogen ion concentration than normal rain (1 mark)

This indicates significantly higher acidity in the industrial area rainwater (1 mark)

Impact on Ghanaian ecosystems (3 marks)

Expected Answer (any two ecosystems with correct analysis):

Volta Lake fisheries:

Acidic rain lowers the pH of the lake water (0.5 mark)

Affects fish reproduction and can lead to decline in fish populations (0.5 mark)

Disrupts aquatic food chains and reduces biodiversity (0.5 mark)

Cocoa farms in Western Region:

Acidic rain damages cocoa leaves through direct contact (0.5 mark)

Leaches essential nutrients from soil, reducing fertility (0.5 mark)

Can lead to decreased crop yields and economic losses for farmers (0.5 mark)

Atewa Forest Reserve:

Damages foliage and weakens trees, making them susceptible to disease (0.5 mark)

Affects soil microorganisms essential for nutrient cycling (0.5 mark)

Can lead to forest degradation and loss of biodiversity in this critical ecosystem (0.5 mark)

Analysis of Solutions (10 marks)

Calculate $[\text{OH}^-]$ concentration (3 marks)

Expected Working: Using $K_w = [\text{H}^+][\text{OH}^-] = 10^{-14}$ at 25°C

For pH 2.5: $[\text{H}^+] = 10^{-2.5} = 3.16 \times 10^{-3} \text{ mol/dm}^3$

$[\text{OH}^-] = \frac{10^{-14}}{3.16 \times 10^{-3}} = 3.16 \times 10^{-12} \text{ mol/dm}^3$

For pH 7.0: $[\text{H}^+] = 10^{-7} = 1.0 \times 10^{-7} \text{ mol/dm}^3$

$[\text{OH}^-] = \frac{10^{-14}}{1.0 \times 10^{-7}} = 1.0 \times 10^{-7} \text{ mol/dm}^3$

For pH 11.5: $[\text{H}^+] = 10^{-11.5} = 3.16 \times 10^{-12} \text{ mol/dm}^3$

$[\text{OH}^-] = \frac{10^{-14}}{3.16 \times 10^{-12}} = 3.16 \times 10^{-3} \text{ mol/dm}^3$

Expected Answer:

Solution pH 2.5: $[\text{OH}^-] = 3.16 \times 10^{-12} \text{ mol/dm}^3$ (1 mark)

Solution pH 7.0: $[\text{OH}^-] = 1.0 \times 10^{-7} \text{ mol/dm}^3$ (1 mark)

Solution pH 11.5: $[\text{OH}^-] = 3.16 \times 10^{-3} \text{ mol/dm}^3$ (1 mark)

Reaction with carbonate compound (2 marks)

Expected Answer:

Solution pH 2.5 (acidic): Vigorous effervescence as CO₂ gas is produced from the reaction between H⁺ ions and carbonate (1 mark)

Solution pH 7.0 (neutral): No visible reaction or very slight effervescence (0.5 mark)

Solution pH 11.5 (basic): No visible reaction (0.5 mark)

Design tests using local materials (3 marks)

Expected Answer: Any three appropriate tests with clear methodology:

Red hibiscus flower extract as indicator:

turns redder in acidic solution (pH 2.5), remains purple in neutral (pH 7.0), and turns greenish in basic solution (pH 11.5) (1 mark)

Reaction with eggshells (calcium carbonate): bubbling with pH 2.5 solution, minimal or no reaction with others (1 mark)

Soap test: pH 2.5 solution will reduce lathering, pH 7.0 moderate lathering, pH 11.5 enhanced lathering (1 mark)

Alternative: Test with baking soda, aluminium foil, or locally available metal samples with observable differences in reactions

Influence of ion concentration on properties (2 marks)

Expected Answer:

In the acidic solution (pH = 2.5), the high [H⁺] concentration (3.16×10^{-3} mol/dm³) and low [OH⁻] concentration (3.16×10^{-12} mol/dm³) create corrosive properties that can dissolve metals and carbonates (0.5 mark)

In the neutral solution (pH 7.0), equal concentrations of [H⁺] and [OH⁻] (both 1.0×10^{-7} mol/dm³) create balanced properties with minimal reactivity (0.5 mark)

In the basic solution (pH 11.5), the high [OH⁻] concentration (3.16×10^{-3} mol/dm³) creates slippery feeling, bitter taste, and ability to saponify fats (0.5 mark)

The ratio of these ions determines whether a solution acts as an acid, base, or neutral substance in chemical reactions (0.5 mark)

Borehole Water Analysis (10 marks)

Calculate hydrogen ion concentration (3 marks)

Expected Working: [H⁺] = 10^{-5.8}

Expected Answer: [H⁺] = 1.58×10^{-6} mol/dm³ (3 marks)

Correct formula (1 mark)

Correct calculation (1 mark)

Correct units (1 mark)

Calculate hydroxide ion concentration (3 marks)

Expected Working: Using $K_w = [H^+][OH^-] = 10^{-14}$ at 25°C

$$[OH^-] = \frac{10^{-14}}{[H^+]} = \frac{10^{-14}}{(1.58 \times 10^{-6})}$$

Expected Answer: [OH⁻] = 6.33×10^{-9} mol/dm³ (3 marks)

Correct formula (1 mark)

Correct calculation (1 mark)

Correct units (1 mark)

Chemical principles of treatment method (4 marks)

Expected Answer:

Wood ash contains calcium carbonate (CaCO₃), calcium oxide (CaO), and potassium compounds that are basic in nature (1 mark)

When water passes through wood ash, these compounds dissolve to form hydroxide ions: CaO + H₂O → Ca(OH)₂ (1 mark)

The hydroxide ions neutralise the acidity in the water through the reaction:

H⁺ + OH⁻ → H₂O, raising the pH to a more neutral level (1 mark)

The sand acts as a physical filter to remove suspended particles and some pathogens, while the increased pH helps to precipitate some dissolved metals as insoluble hydroxides, making the water safer for consumption (1 mark)

WEEK 3

1. Key: B. HCl completely ionises in water while acetic acid only partially ionises

Explanation: HCl is a strong acid and fully dissociates into H⁺ ions, increasing acidity (lower pH). Acetic acid is a weak acid and only partially ionises.

2. Key: C. 0.316%

Solution: [H⁺] = 10^{-pH}

$$[H^+] = 10^{-3.5} = 3.16 \times 10^{-4} \text{ mol/dm}^3$$

$$\% \text{ ionisation} = \frac{3.16 \times 10^{-4}}{0.1} \times 100 = 0.316\%$$

3. Key: A. 3.35

Solution: pK_a = -logK_a

$$pK_a = -\log(4.5 \times 10^{-4}) = 3.35$$

4. Key: C. 1.3×10^{-5}

Working: pH = $\frac{1}{2}pK_a - \frac{1}{2}\log C$

$$pH + \frac{1}{2}\log C = pK_a$$

$$2.8 + \frac{1}{2}\log 0.2 = \frac{1}{2}pK_a$$

$$2(2.45) = pK_a$$

$$= pK_a$$

$$K_a = 10^{-pK_a}$$

$$K_a = 10^{-4.9}$$

$$K_a = 1.3 \times 10^{-5}$$

5. Key: D. 11.21

$$\text{Solution: } \text{pOH} = \frac{1}{2} \text{pK}_b - \frac{1}{2} \log C$$

$$\text{pOH} = \frac{1}{2}(4.75) - \frac{1}{2} \log(0.15)$$

$$\text{pOH} = 2.79$$

$$\text{pH} = 14 - \text{pOH}$$

$$\text{pH} = 14 - 2.79$$

$$\text{pH} = 11.21$$

WEEK 4

HCl is a strong acid so it fully dissociates: $[\text{H}^+] = 0.01 \text{ mol/dm}^3$ (1 mark).

Correct application of equation: $\text{pH} = -\log[\text{H}^+]$ (1 mark)

Correct answer: $\text{pH} = -\log(0.01) = 2.0$ (1 mark)

NaOH is a strong base so fully dissociates: $[\text{OH}^-] = 0.001 \text{ mol/dm}^3$ (1 mark)

Correct calculation of pOH: $\text{pOH} = -\log[\text{OH}^-] = -\log(0.001) = 3.0$ (1 mark)

Correct application of $\text{pH} + \text{pOH} = 14$, therefore $\text{pH} = 14 - 3 = 11.0$ (1 mark)

Correct setup of equilibrium expression: $K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$ (1 mark)

Appropriate approximation: $[\text{H}^+] = [\text{A}^-]$ and $[\text{HA}] \approx 0.1 \text{ mol/dm}^3$ (1 mark)

Correct substitution: $1.8 \times 10^{-5} = \frac{[\text{H}^+]^2}{0.1}$ (1 mark)

Correct substitution: $[\text{H}^+] = \sqrt{(1.8 \times 10^{-5} \times 0.1)}$

$$= 1.34 \times 10^{-3} \text{ mol/dm}^3$$

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

$$= 2.87 \text{ (1 mark)}$$

a. Use of Henderson-Hasselbalch equation: $\text{pH} = \text{pK}_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right)$ (1 mark)

Correct identification of $\text{pK}_a = 4.76$ (1 mark)

Correct substitution: $\text{pH} = 4.76 + \log \left(\frac{0.10}{0.15} \right)$ (1 mark)

Correct answer: $\text{pH} = 4.76 + \log(1) = 4.76$ (1 mark)

b. i. Correct application of Henderson-Hasselbalch equation (1 mark)

Correct substitution: $\text{pH} = 4.76 + \log \left(\frac{0.10}{0.15} \right)$ (1 mark)

Correct answer: $\text{pH} = 4.76 + \log 0.667$

$$= 4.76 - 0.18$$

$$= 4.58 \text{ (1 mark)}$$

ii. Correct application of Henderson-Hasselbalch equation (1 mark)

Correct substitution: $\text{pH} = 4.76 + \log \left(\frac{0.10}{0.15} \right)$ (1 mark)

Correct answer: $\text{pH} = 4.76 + \log 1.5$

$$= 4.76 + 0.18$$

$$= 4.94 \text{ (1 mark)}$$

c. Henderson-Hasselbalch equation: $\text{pH} = \text{pK}_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right)$ (1 mark)

Explanation that when $[\text{A}^-] = [\text{HA}]$, $\text{pH} = \text{pK}_a$ (1 mark)

Explanation that increasing the ratio $\frac{[\text{A}^-]}{[\text{HA}]}$ increases pH (1 mark)

Explanation that decreasing the ratio $\frac{[\text{A}^-]}{[\text{HA}]}$ decreases pH (1 mark)

d. Buffer with smallest pH change (3 marks)

HCl reacts with CH_3COO^- : $\text{H}^+ + \text{CH}_3\text{COO}^- \rightarrow \text{CH}_3\text{COOH}$ (1 mark)

For buffer (i) (0.15 M CH_3COOH and 0.10 M CH_3COONa):

Initial amounts: 0.015 mol CH_3COOH , 0.010 mol CH_3COONa

After adding 0.005 mol HCl: 0.020 mol CH_3COOH , 0.005 mol CH_3COONa

New concentrations: 0.20 mol/dm³ CH_3COOH , 0.05 mol/dm³ CH_3COONa (1 mark)

$$\text{New pH} = 4.76 + \log \left(\frac{0.05}{0.20} \right)$$

$$= 4.76 + \log(0.25)$$

$$= 4.76 - 0.60$$

$$= 4.16 \text{ (1 mark)}$$

pH change: $\Delta\text{pH} = 4.16 - 4.58$

$$= -0.42 \text{ (absolute change: 0.42) (1 mark)}$$

For buffer (ii) (0.10 M CH_3COOH and 0.15 M CH_3COONa):

Initial amounts: 0.010 mol CH_3COOH , 0.015 mol CH_3COONa

After adding 0.005 mol HCl: 0.015 mol CH_3COOH , 0.010 mol CH_3COONa

New concentrations: 0.15 mol/dm³ CH_3COOH , 0.10 mol/dm³ CH_3COONa (1 mark)

$$\text{New pH} = 4.76 + \log \left(\frac{0.10}{0.15} \right)$$

$$= 4.76 + \log(0.667)$$

$$= 4.76 - 0.18$$

$$= 4.58 \text{ (1 mark)}$$

pH change: $\Delta\text{pH} = 4.58 - 4.94$

$$= -0.36 \text{ (absolute change: 0.36) (1 mark)}$$

Buffer (ii) experiences smaller pH change ($\Delta\text{pH} = 0.36$ vs 0.42) because it has higher initial concentration of base component which can neutralise more added acid (1 mark)

WEEK 5

Determination of the number amount of acid and salt ($n = C \times V \text{ dm}^3$)

Moles of propanoic acid

$$\begin{aligned} n(\text{propanoic acid}) &= 0.15 \text{ mol/dm}^3 \times 0.035 \text{ dm}^3 \\ &= 0.00525 \text{ mol (1 mark)} \end{aligned}$$

Moles of sodium propanoate

$$\begin{aligned} n(\text{sodium propanoate}) &= 0.20 \text{ mol/dm}^3 \times 0.015 \text{ dm}^3 \\ &= 0.00300 \text{ mol (1 mark)} \end{aligned}$$

Calculation of molar concentrations in final solution ($C = \frac{n}{V(\text{Total})}$)

Determination of final volume

$$\begin{aligned} \text{Total volume} &= 0.035 + 0.015 \\ &= 0.050 \text{ dm}^3 \end{aligned}$$

Determination of the final concentration of the acid

$$\begin{aligned} [\text{Acid}] &= \frac{0.00525 \text{ mol}}{0.050 \text{ dm}^3} \\ &= 0.105 \text{ mol/dm}^3 \text{ (1 mark)} \end{aligned}$$

Determination of the final concentration of the salt

$$\begin{aligned} [\text{Salt}] &= \frac{0.00300 \text{ mol}}{0.050 \text{ dm}^3} \\ &= 0.060 \text{ mol/dm}^3 \text{ (1 mark)} \end{aligned}$$

Application of the Henderson-Hasselbalch equation

$$\text{pH} = \text{pK}_a + \log\left(\frac{[\text{salt}]}{[\text{acid}]}\right) \text{ but } \text{pK}_a = -\log K_a$$

$$\begin{aligned} \text{pH} &= -\log(1.3 \times 10^{-5}) + \log\left(\frac{0.060}{0.105}\right) \text{ (1 mark)} \\ &= 4.65 \text{ (1 mark)} \end{aligned}$$

Buffer response to HCl addition

Chemical equation for the addition of HCl



Amount of HCl added

$$\begin{aligned} n(\text{HCl}) &= 0.10 \text{ mol/dm}^3 \times 0.005 \text{ dm}^3 \\ &= 0.0005 \text{ mol (0.5 mark)} \end{aligned}$$

Change in concentration

$$\begin{aligned} \text{New moles of acid} &= 0.00525 + 0.0005 \\ &= 0.00575 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{New moles of salt} &= 0.00300 - 0.0005 \\ &= 0.00250 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Total volume} &= 50.0 + 5.0 \text{ cm}^3 \\ &= 55.0 \text{ cm}^3 \end{aligned}$$

$$= 0.055 \text{ dm}^3$$

$$\begin{aligned} \text{New [acid]} &= \frac{0.00575}{0.055} \\ &= 0.1045 \text{ mol/dm}^3 \text{ (0.5 marks)} \end{aligned}$$

$$\begin{aligned} \text{New [salt]} &= \frac{0.00250}{0.055} \\ &= 0.0455 \text{ mol/dm}^3 \text{ (0.5 marks)} \end{aligned}$$

Change in pH

$$\text{New pH} = \text{pK}_a + \log\left(\frac{[\text{salt}]}{[\text{acid}]}\right)$$

$$\text{New pH} = -\log(1.3 \times 10^{-5}) + \log\left(\frac{0.0455}{0.1045}\right) \text{ (0.5 mark)}$$

$$\text{New pH} = 4.53 \text{ (0.5 marks)}$$

The pH decreased by only 0.12 units despite acid addition, demonstrating buffer action. (0.5 mark)

Preparation of buffer

Calculation of buffer composition

Determine the ratio of salt to acid

$$\text{pK}_a = -\log K_a$$

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

$$\text{pK}_a = 4.74$$

$$\text{pH} = \text{pK}_a + \log\left(\frac{[\text{salt}]}{[\text{acid}]}\right) \text{ (1 mark)}$$

$$4.85 = 4.74 + \log\left(\frac{[\text{salt}]}{[\text{acid}]}\right)$$

$$\log\left(\frac{[\text{salt}]}{[\text{acid}]}\right) = 4.85 - 4.74$$

$$\log\left(\frac{[\text{salt}]}{[\text{acid}]}\right) = 0.11$$

$$\frac{[\text{salt}]}{[\text{acid}]} = 10^{0.11}$$

$$\frac{[\text{salt}]}{[\text{acid}]} = 1.29 \text{ (1 mark)}$$

Determine the concentrations of both salt and acid

$$[\text{acid}] + [\text{salt}] = 0.20 \text{ mol/dm}^3 \text{ (0.5 mark)}$$

$$\frac{[\text{salt}]}{[\text{acid}]} = 1.29 \text{ (0.5 mark)}$$

$$\text{Let } [\text{acid}] = x, \text{ then } [\text{salt}] = 1.29x$$

$$x + 1.29x = 0.20$$

$$2.29x = 0.20$$

$$x = 0.087$$

$$[\text{acid}] = 0.087 \text{ mol/dm}^3 \text{ (0.5 mark)}$$

$$[\text{salt}] = 0.20 - 0.087$$

$$= 0.113 \text{ mol/dm}^3 \text{ (0.5 mark)}$$

Determine the mass needed to prepare the concentrations involved

Moles of acetic acid = Concentration x volume (1 mark)

$$\begin{aligned} n(\text{acetic acid}) &= 0.087 \text{ mol/dm}^3 \times 0.500 \text{ dm}^3 \\ &= 0.0435 \text{ mol (0.5 mark)} \end{aligned}$$

Mass of glacial acetic acid = Molar mass x number of moles

$$\begin{aligned} m(\text{acetic acid}) &= 60.05 \text{ g/mol} \times 0.0435 \text{ mol} \\ &= 2.61 \text{ g (1 mark)} \end{aligned}$$

$$\begin{aligned} \text{Moles of sodium acetate} &= 0.113 \text{ mol/dm}^3 \times 0.500 \text{ dm}^3 \\ &= 0.0565 \text{ mol (0.5 mark)} \end{aligned}$$

$$\begin{aligned} \text{Mass of sodium acetate} &= 82.03 \text{ g/mol} \times 0.0565 \text{ mol} \\ &= 4.63 \text{ g (1 mark)} \end{aligned}$$

Analysis of buffer capacity

Buffer capacity is the amount of acid or base a buffer can neutralise before significant pH change.

Doubling concentration while maintaining the ratio would double the absolute amounts of acid and salt.

This would increase buffer capacity as more acid/base could be neutralised before significant pH change.

The initial pH would remain unchanged as it depends only on the ratio of [salt] to [acid]. (2 mark)

Calculations (7 marks)

Convert K_b of NH_3 to pK_a of NH_4^+ (1 mark)

Use Henderson-Hasselbalch Equation to determine $[\text{NH}_4^+]/[\text{NH}_3]$ ratio (1 mark)

Apply total buffer concentration to find individual concentrations (1.5 marks)

Calculate volume of ammonia solution required (1 mark)

Calculate mass of ammonium chloride required (1.5 marks)

Analysis of Buffer System Choice (3 marks)

Identifies that ammonia buffer maintains pH closer to 9.3, ideal for basic or alkaline enzymes. (1 mark)

Explains phosphate buffer typically stabilises pH ~7.2, which may not support enzyme function in alkaline conditions. (1 mark)

Notes that choosing ammonia buffer avoids possible phosphate interference with enzyme cofactors or precipitation in metal-ion-sensitive reactions. (1 mark)

SECTION 2: HYDRATION AND HYDROLYSIS OF SALTS

STRAND: PHYSICAL CHEMISTRY

Sub-Strand: Equilibria

Learning Outcome: Use your knowledge and understanding of concepts in hydration and hydrolysis to determine the types of salts

Content Standard: Demonstrate knowledge and understanding of concepts of hydration and hydrolysis of salt

INTRODUCTION AND SECTION SUMMARY

This section introduces learners to the concepts of hydration and hydrolysis of salts in aqueous solutions. Learners will explore how water molecules interact with ions during dissolution (hydration) and how certain salts can affect the pH of solutions through reactions with water (hydrolysis). By the end of this section, learners will understand the physical nature of hydration, the factors affecting hydration energy, and be able to predict whether salt solutions will be acidic, basic, or neutral based on their constituent ions. This knowledge is fundamental to understanding solution chemistry and the behaviour of ionic compounds in water.

The week covered by this section is:

- Week 6: Explain hydration and hydrolysis of salts and predict the acidity or alkalinity of aqueous salt solutions

SUMMARY OF PEDAGOGICAL EXEMPLARS

Teachers should employ talk-for-learning and collaborative learning approaches to develop learners' understanding of hydration and hydrolysis. In mixed-ability groups, learners should brainstorm to construct the meaning of hydration, encouraging inclusive participation and knowledge sharing. Through collaborative activities, learners should then write chemical equations showing hydration of various salts and discuss key features of hydration as an exothermic physical process. Groups should analyse how ion size and charge influence hydration energy, developing both conceptual understanding and scientific communication skills.

For hydrolysis concepts, teachers should again use talk-for-learning activities where learners brainstorm the meaning of hydrolysis and distinguish between cation and anion hydrolysis. Collaborative learning should then focus on writing chemical equations illustrating hydrolysis mechanisms and examining why certain salts (those from strong acids and strong bases) do not undergo significant hydrolysis.

Throughout these activities, teachers should use visual aids, demonstrations, and guided questioning to help learners connect theoretical concepts with observable phenomena such as pH changes in salt solutions. Practical demonstrations of pH testing for various salt solutions can reinforce theoretical understanding.

ASSESSMENT SUMMARY

Indicator	DOK Level	Assessment Strategy and/or Mode
Explain hydration and hydrolysis of salts and predict the acidity or alkalinity of aqueous salt solutions	Level 2: Skills & Concepts	Individual class exercise and Mid-semester exams

WEEK 6

Learning Indicator: Explain hydration and hydrolysis of salts and predict the acidity or alkalinity of aqueous salt solutions

Recap weeks

Year 1	Week 20: Describe the types of intermolecular forces and explain how they arise from the structural features of molecules.
Year 2	Week 9: Explain Arrhenius, Brønsted–Lowry and Lewis concepts of acids and bases Week 11: Distinguish between strong and weak acids, strong and weak bases in terms of the extent of dissociation, rate of reaction, electrical conductivity, pH value in aqueous solution and enthalpy change of neutralisation value. Week 12: Explain the meaning of salt and classify salts

FOCAL AREA 1: HYDRATION AND HYDROLYSIS OF SALTS

1. Hydration of salts

When salts interact with water, two distinct but important processes can occur: hydration and hydrolysis. Hydration is primarily a physical process where water molecules surround and bind to ions through ion-dipole interactions, forming hydrated ions. This process does not change the chemical identity of the ions but affects their physical properties and interactions in solution.

Hydration is the process in which water molecules form coordinate bonds with ions in a crystalline salt. $\text{CuSO}_4(\text{s}) + 5\text{H}_2\text{O}(\text{l}) \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O}(\text{s})$

The water molecules that become attached to the ions are called ‘water of crystallisation’ or ‘water of hydration’. The resulting compounds are known as hydrated salts or hydrates.

2. Hydrolysis of salts

Hydrolysis is a chemical reaction in which a salt reacts with water, resulting in the formation of either an acidic or basic solution due to the reaction of the salt’s component ions with water molecules. Hydrolysis occurs when the **anion of a weak acid** or the **cation of a weak base** (or both) reacts with water to produce H_3O^+ or OH^- ions.

Water self-ionises to a small extent: $\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$

The ionic product of water: $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$ at 25°C.

When a salt dissolves in water, its ions can disturb this equilibrium by reacting with water molecules, resulting in either excess H_3O^+ (acidic solution) or excess OH^- (basic solution).

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.

Mechanism of Hydrolysis

The mechanism of hydrolysis depends on the type of ion involved:

Cation Hydrolysis

Cation hydrolysis occurs when the cation of a salt reacts with water molecules, producing hydronium ions (H_3O^+) that make the solution acidic. This process is most significant for two types of cations:

1. Cations that are conjugate acids of weak bases
2. Metal ions with high charge density that strongly polarise coordinated water molecules

Mechanism of Cation Hydrolysis

When a salt containing these cations dissolves in water, the cation acts as a Brønsted-Lowry acid by donating a proton to water:



The equilibrium position of this reaction depends on the strength of the cation as an acid. The weaker the parent base, the stronger its conjugate acid, and the more extensive the hydrolysis.

Example 1: Ammonium Chloride (NH_4Cl)

NH_4^+ is the conjugate acid of the weak base NH_3 (ammonia).

Step 1: Dissociation $\text{NH}_4\text{Cl}(\text{s}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{Cl}^-(\text{aq})$

Step 2: Cation Hydrolysis $\text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$

The production of H_3O^+ ions result in an acidic solution ($\text{pH} < 7$).

Example 2: Aluminium Chloride (AlCl_3)

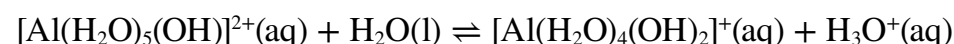
Al^{3+} is a small, highly charged metal ion that strongly polarises water molecules in its hydration sphere.

Step 1: Dissociation $\text{AlCl}_3(\text{s}) \rightarrow \text{Al}^{3+}(\text{aq}) + 3\text{Cl}^-(\text{aq})$

Step 2: Hydration $\text{Al}^{3+}(\text{aq}) + 6\text{H}_2\text{O}(\text{l}) \rightarrow [\text{Al}(\text{H}_2\text{O})_6]^{3+}(\text{aq})$

Step 3: Hydrolysis $[\text{Al}(\text{H}_2\text{O})_6]^{3+}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons [\text{Al}(\text{H}_2\text{O})_5(\text{OH})]^{2+}(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$

The hydrolysis may continue in stepwise fashion:



The production of H_3O^+ ions make the solution acidic ($\text{pH} < 7$).

Anion Hydrolysis

Anion hydrolysis occurs when the anion of a salt reacts with water molecules, producing hydroxide ions (OH^-) that make the solution basic.

The anion are conjugate bases of weak acids

Mechanism of Anion Hydrolysis

When a salt containing these anions dissolves in water, the anion acts as a Brønsted-Lowry base by accepting a proton from water: $\text{Anion} + \text{H}_2\text{O} \rightleftharpoons \text{Conjugate acid} + \text{OH}^-$

The equilibrium position of this reaction depends on the strength of the anion as a base. The weaker the parent acid, the stronger its conjugate base, and the more extensive the hydrolysis.

Example 1: Sodium ethanoate (CH_3COONa)

CH_3COO^- is the conjugate base of the weak acid CH_3COOH (ethanoic acid).

Step 1: Dissociation $\text{CH}_3\text{COONa}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$

Step 2: Anion Hydrolysis $\text{CH}_3\text{COO}^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{COOH}(\text{aq}) + \text{OH}^-(\text{aq})$

The production of OH^- ions result in a basic solution ($\text{pH} > 7$).

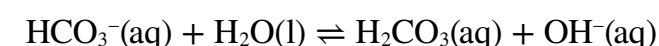
Example 2: Sodium trioxocarbonate(IV) (Na_2CO_3)

CO_3^{2-} is the conjugate base of the weak acid HCO_3^- .

Step 1: Dissociation $\text{Na}_2\text{CO}_3(\text{s}) \rightarrow 2\text{Na}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$

Step 2: Anion Hydrolysis $\text{CO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$

The HCO_3^- ion can undergo further hydrolysis:



Note: The hydrolysis reaction produces both H_2CO_3 and OH^- . The OH^- ions are responsible for the basic nature of the solution. While H_2CO_3 is indeed an acid, it remains largely undissociated in the solution.

The doubly charged carbonate ion has a strong affinity for protons, making it a relatively strong base. The production of OH^- ions result in a basic solution ($\text{pH} > 7$).

Note: Salts derived from strong acids and strong bases do not undergo hydrolysis. Their solutions are neutral ($\text{pH} = 7$) because neither the cation nor the anion has the ability to react with water to any significant extent. Example: Sodium chloride (NaCl)

Dissociation: $\text{NaCl}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$. The two ions do not react with water.

Factors Affecting the Extent of Hydrolysis

Several factors influence the extent of salt hydrolysis, and 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.
2. Charge on the ion: Higher charge generally leads to stronger hydrolysis (especially for metal cations).
3. Size of the ion: For ions with the same charge, smaller size typically results in stronger hydrolysis due to higher charge density.
4. Concentration of the salt solution: More dilute solutions often show proportionally greater hydrolysis.
5. Temperature: Higher temperatures generally increase the extent of hydrolysis.

Learning Tasks

1. Define the term *hydrolysis* in the context of salts. How does it differ from hydration?
2. For each of the following salts (KNO_3 , $\text{CH}_3\text{COONH}_4$, ZnCl_2 and Na_2CO_3), state whether its aqueous solution would be acidic, basic, or neutral. Explain your reasoning.

3. A Form 3 student at Accra Academy conducted an experiment during a practical session at the chemistry laboratory. He was investigating the pH of various salt solutions commonly found in Ghana. He tested four unknown salt solutions with universal indicator and observed the following colours:

Solution	Colour	Approximate pH
A	Green	7
B	Red	3
C	Purple	10
D	Yellow-green	6

The salts he was testing could be:

- a. Potassium chloride (KCl) - commonly found in some fertilisers used in Ghana’s cocoa farms.
- b. Ammonium nitrate (NH₄NO₃) - used in agriculture in the northern regions.
- c. Sodium carbonate (Na₂CO₃) - used in local soap making in Kumasi and
- d. Iron(III) chloride (FeCl₃) - used in water treatment at Ghana Water Company.

Identify which salt corresponds to each solution (A, B, C, and D) and explain your reasoning by discussing the hydrolysis properties of each salt.

Pedagogical Exemplars

Collaborative learning

- 1. Introduce cation hydrolysis (producing acidic solutions) and anion hydrolysis (producing basic solutions). Form small groups of 4-5 learners, establishing ground rules for respectful discussion where all contributions are valued.
- 2. For each group, assign roles: coordinator, recorder, presenter, fact-checker, and timekeeper. Distribute reference materials and emphasise that each role is equally important for group success.
- 3. Groups write and discuss chemical equations for cation hydrolysis using examples like NH₄Cl and FeCl₃. Learners explain why these reactions produce acidic solutions, supporting peers who may be struggling with concepts.
- 4. Groups examine anion hydrolysis using examples like CH₃COONa and Na₂CO₃, discussing why these reactions produce basic solutions. Learners explain concepts in their own words, appreciating different approaches to understanding.
- 5. Groups identify salts that don’t undergo significant hydrolysis (e.g., NaCl, KNO₃) and explain why, using acid-base principles. Learners practice constructive feedback when reviewing each other’s work.
- 6. Each group creates a chart summarising cation hydrolysis, anion hydrolysis, and non-hydrolysing salts with examples and pH predictions. All members contribute, celebrating the unique strengths each brings to the task.

- 7. Groups present their findings, explaining one type of hydrolysis with relevant equations. Classmates practice respectful listening and thoughtful questioning in a supportive atmosphere where mistakes are viewed as learning opportunities.
- 8. The teacher clarifies misconceptions and summarises key points. Learners reflect on new concepts learned, effective collaboration strategies, and areas for improvement. The activity concludes with appreciations for helpful contributions and connections to real-world applications.

Key Assessment: Mid-Semester Examination

- 1. Classify the following as strong acids, weak acids, strong bases, or weak bases: HCl, CH₃COOH, NaOH and NH₃. (4 marks)

- 2. a. Explain the following terms

- i. Hydration of salts
- ii. Hydrolysis of salts

- b. Complete the following table by indicating whether the aqueous solution of each salt would be acidic, basic, or neutral

Salt	Parent acid	Parent base	pH of solution (acidic/basic/neutral)
NaCl			
NH ₄ Cl			
CH ₃ COONa			
AlCl ₃			

- c. Write balanced chemical equations to show:

- i. The hydration of anhydrous copper(II) tetraoxosulphate (IV)
- ii. The hydrolysis of ammonium ions in water
- iii. The hydrolysis of ethanoate ions in water

- d. Explain why a solution of sodium ethanoate (CH₃COONa) is basic, while a solution of ammonium chloride (NH₄Cl) is acidic. (15 marks)

- 3. The pH of different household products varies widely due to the presence of different salts:

- a. Baking soda (NaHCO₃) solution has a pH around 8.5
- b. Table salt (NaCl) solution has a pH around 7.0
- c. Alum [KAl(SO₄)₂.12H₂O] solution has a pH around 3.5.

Explain these observations in terms of salt hydrolysis. (10 marks)



Hint

Conduct **Mid-semester Examination**. It should cover contents taught from Weeks 1 to 5. Refer to the **Table of Specification and Structure of the Paper**. Learners' papers should be marked immediately and scores entered into the SAP.

Structure

Section A: 25 Multiple-choice questions

Section B: 2 Essay Questions

Table of specification for item construction for section A and B

Week	Learning Indicator	Types of questions	DoK Levels				Total
			1	2	3	4	
1& 2	Explain and undertake calculations on the concepts of pH, and pOH and describe their significance in everyday life	MCQs	2	2	2		6
		ESSAY	-	-	1		1
3	Explain the incomplete or partial ionisation of weak acids and weak bases and calculate pKa and pKb	MCQs	2	2	2		6
		ESSAY	-	-	-		
4	Explain buffer solution, buffer action and their applications in everyday life	MCQs	3	2	1		6
		ESSAY					
5	Explain buffer solution, buffer action and their applications in everyday life	MCQs	2	3	2		7
		ESSAY			1		1
		Total	9	9	9		27

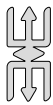
PLC Session 6

45 minutes

Planning & Reviewing the Learning Plan

- Begin the session by recapping lessons from the previous session, focusing on how learning was applied.
- Use your subject specific app to explore the content before planning. Ask the app to
 - explain the week's concepts in a practical way that will make it easier for learners to understand.
 - identify three common misconceptions learners may have on the week's content and how to address them using contexts familiar to Ghanaian learners (NTS 3a, 3g and 3m).
- Use your subject specific app to plan your lesson. Ask the app to

- suggest the most effective method for teaching the week's concepts to learners with diverse backgrounds.
- suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q).
- review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



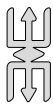
Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the *Values Learning Community Handbook* and link the values for the week in your lesson to ensure that learners are applying what they learn.

- Use your app to plan and administer the mode of assessment for the Student Assessment Portal (SAP).
 - The key assessment for the week is **mid-semester examination**. Chat with your app to plan this assessment by reviewing or developing the assessment items/tasks and marking schemes/rubrics for the mid-semester examination (NTS 3k-3q).

10 minutes	Content Exploration
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- As a subject group, select one problem or topic/idea/term in the week's content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
 - use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - all teachers in the group work through the problem individually for 3–4 minutes
 - the group then works through it together the problem, step by step
 - highlight the most common errors made at each step (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not 'How do I teach this?' but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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- Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
 - Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.

- b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
7.	Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8.	Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9.	Remember to a. transfer your chats into your learning planner template. b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a). c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

Section 2 Review

This section focused on understanding how salts interact with water through the processes of hydration and hydrolysis. During Week 6, learners have explored the physical nature of hydration and how hydrolysis affects solution pH.

Key Concepts Covered

Hydration of Salts: Learners should now understand that hydration involves water molecules surrounding and interacting with ions in solution. They should recognise that hydration is an exothermic physical process and be able to explain how factors such as ion size and charge affect hydration energy.

Hydrolysis of Salts: Learners should understand that hydrolysis occurs when ions from a salt react with water (chemical reaction) to produce H₃O⁺ or OH⁻ ions, affecting solution pH. They should be able to distinguish between:

- 1. Cation hydrolysis (producing acidic solutions)
- 2. Anion hydrolysis (producing basic solutions)
- 3. No significant hydrolysis (neutral solutions)

Predicting Solution pH: Learners should be able to predict whether a salt solution will be acidic, basic, or neutral based on the nature of the constituent ions:

- 1. Salts of strong acids and strong bases → neutral solutions
- 2. Salts of weak acids and strong bases → basic solutions

- 3. Salts of strong acids and weak bases → acidic solutions
- 4. Salts of weak acids and weak bases → pH depends on relative strengths (using K_a and K_b values)

MARKING SCHEMES FOR THE KEY ASSESSMENT

1. Classify the following as strong acids, weak acids, strong bases, or weak bases:

HCl, CH₃COOH, NaOH and NH₃. (4 marks)

- a. Correct classification of HCl as a strong acid (1 mark)
- b. Correct classification of CH₃COOH (ethanoic acid) as a weak acid (1 mark)
- c. Correct classification of NaOH as a strong base (1 mark)
- d. Correct classification of NH₃ (ammonia) as a weak base (1 mark)

2. Explanation of hydration and hydrolysis (15 marks)

- a. (i) Hydration is the process in which water molecules form coordinate bonds with metal ions or are incorporated into the crystal structure of a salt. The water molecules that are bound to the metal ion are called water of hydration or water of crystallisation. Hydrated salts contain a specific number of water molecules as part of their structure, represented in their chemical formula e.g., CuSO₄·5H₂O. (3 marks)

(ii) Hydrolysis is a chemical reaction in which a salt reacts with water, resulting in the formation of either an acidic or basic solution. This occurs when the anion of a weak acid or the cation of a weak base (or both) in the salt interacts with water molecules to produce either hydronium ions (H₃O⁺) or hydroxide ions (OH⁻), thereby affecting the pH of the solution. (3 marks)

b. Complete the following table: (4 marks)

Salt	Parent acid	Parent base	pH of solution (acidic/basic/neutral)
NaCl	HCl (strong acid)	NaOH (strong base)	Neutral (1 mark)
NH ₄ Cl	HCl (strong acid)	NH ₃ (weak base)	Acidic (1 mark)
CH ₃ COONa	CH ₃ COOH (weak acid)	NaOH (strong base)	Basic (1 mark)
AlCl ₃	HCl (strong acid)	Al(OH) ₃ (weak base)	Acidic (1 mark)

- c. Balanced chemical equations (3 marks)
 - i. CuSO₄ + 5H₂O → CuSO₄·5H₂O (1 mark)
 - ii. NH₄⁺ + H₂O ⇌ NH₃ + H₃O⁺ (1 mark)
 - iii. CH₃COO⁻ + H₂O ⇌ CH₃COOH + OH⁻ (1 mark)
- d. Why a solution of CH₃COONa is basic, while a solution of NH₄Cl is acidic. [2 mark]

Explanation that sodium ethanoate contains ethanoate ions from a weak acid that hydrolyse to produce OH^- ions, making the solution basic, while ammonium chloride contains ammonium ions from a weak base that hydrolyse to produce H^+ ions, making the solution acidic. (2 mark)

3. Why the pH of different household products differs (10 marks)

a. Baking soda (NaHCO_3) solution [3 marks]

- Identification that NaHCO_3 is formed from a strong base (NaOH) and a weak acid (H_2CO_3) (1 mark)
- Correct equation for the hydrolysis: $\text{HCO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3 + \text{OH}^-$ (1 mark)
- Explanation that the hydrolysis produces hydroxide ions (OH^-) which makes the solution basic with $\text{pH} > 7$ (1 mark)

b. Table salt (NaCl) solution (3 marks)

- Identification that NaCl is formed from a strong acid (HCl) and a strong base (NaOH) (1 mark)
- Explanation that neither Na^+ nor Cl^- undergoes significant hydrolysis in water (1 mark)
- Explanation that with no significant hydrolysis, the solution remains neutral with $\text{pH} \approx 7$ (1 mark)

c. Alum [$\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$] solution (4 marks)

- Identification that the aluminium ion (Al^{3+}) is a cation of a weak base [$\text{Al}(\text{OH})_3$] (1 mark)
- $\text{Al}^{3+} + 6\text{H}_2\text{O} \rightarrow [\text{Al}(\text{H}_2\text{O})_6]^{3+}$ (hydration) (1 mark)
- $[\text{Al}(\text{H}_2\text{O})_6]^{3+} + 3\text{H}_2\text{O} \rightleftharpoons \text{Al}(\text{H}_2\text{O})_3(\text{OH})_3 + 3\text{H}_3\text{O}^+(\text{aq})$ (hydrolysis) (1 mark)
- Explanation that the hydrolysis produces H_3O^+ ions make the solution acidic with $\text{pH} < 7$ (1 mark)

SECTION 3: ACID-BASE INDICATORS AND TITRATION CURVES

STRAND: PHYSICAL CHEMISTRY

Sub-Strand: *Equilibria*

Learning Outcome: Use the understanding of hydrolysis of salts to decide on suitable indicator as well as plot and interpret titration curves

Content Standard: Demonstrate understanding of concepts of hydrolysis of salt in the choice of indicators based on titration curves

INTRODUCTION AND SECTION SUMMARY

This section introduces learners to acid-base indicators and their applications in analytical chemistry. Acid-base indicators are weak acids or bases whose molecular and ionic forms display different colours, allowing chemists to visually determine titration endpoints.

Learners will explore the chemical principles behind indicator colour changes, understanding how the indicator dissociation constant determines the pH transition range. They will investigate common indicators such as phenolphthalein and methyl orange, learning their appropriate applications in various titration scenarios.

A key focus is determining suitable indicators for different titration types. Learners will analyse how the nature of the reactants affects the equivalence point pH and develop skills to match indicator transition ranges with expected pH changes during titrations.

Learners will also master the interpretation and construction of titration curves. These graphical representations plot pH against volume of titrant added, revealing crucial information about the titration process. Students will identify key features including equivalence points and buffer regions, and use these curves to select appropriate indicators.

Through practical activities and problem-solving exercises, learners will connect theoretical knowledge with laboratory applications used in pharmaceutical analysis, environmental monitoring, and quality control.

The week covered by this section is:

- Week 7: Explain acid-base indicator, determine a suitable indicator for a titration and draw graphs to represent titration.

SUMMARY OF PEDAGOGICAL EXEMPLARS

Teachers should employ diverse pedagogical approaches to help learners understand acid-base indicators and titration curves. Through collaborative learning, mixed gender-based groups will discuss the chemical nature of indicators, their working pH ranges, and selection criteria

for different titration scenarios. This approach promotes inclusive participation and diverse perspectives while developing communication skills.

Exploratory learning activities allow learners to investigate different indicators in laboratory settings. Groups should test various indicators with different acid-base combinations, observing colour changes and determining which indicators provide the most distinct endpoints for specific titrations. This hands-on approach develops practical skills and deepens conceptual understanding.

For activity-based learning, learners should draw titration curves for four different acid-base combinations: strong acid-strong base, strong acid-weak base, weak acid-strong base, and weak acid-weak base. On each curve, they should mark the equivalence point, pH range, and indicate appropriate indicators with their transition ranges. This visualisation exercise helps learners connect theoretical knowledge with practical applications.

Throughout these activities, teachers should emphasise how the nature of the acid and base affects the equivalence point pH, which in turn determines indicator selection. Assessment should focus on learners’ ability to select appropriate indicators and explain their choices based on scientific principles.

ASSESSMENT SUMMARY

Indicator	DOK Level	Assessment Strategy and/or Mode
Explain acid–base indicator, determine a suitable indicator for a titration and draw graphs to represent titration.	Level 3: Strategic Reasoning	Individual class exercise

WEEK 7

Learning Indicator: Explain acid–base indicator, determine a suitable indicator for a titration and draw graphs to represent titration

Recap weeks

Year 1	Week 7: Explain the mole concept and its relevance in the preparation of standard solutions.
Year 2	Week 8: State Le Chatelier’s Principle and apply it to deduce qualitatively, the effects of various factors on the positions of equilibrium and on the value of the equilibrium constant Week 9: Explain Arrhenius, Brønsted–Lowry and Lewis concepts of acids and bases Week 10: Describe the physical and chemical properties of acids and bases. Week 11: Distinguish between strong and weak acids, strong and weak bases in terms of the extent of dissociation, rate of reaction, electrical conductivity, pH value in aqueous solution and enthalpy change of neutralisation value.

FOCAL AREA 1: ACID-BASE INDICATORS

1. Acid-base indicator

An acid-base indicator is a weak organic acid or base that changes colour at a specific pH range due to its structural transformation in acidic and basic media. The general formula for an indicator can be represented as HIn, where the undissociated form (HIn) displays one colour, while the dissociated form (In[−]) exhibits a different colour. The equilibrium involved can be written as: HIn (colour 1) ⇌ H⁺ + In[−] (colour 2). This equilibrium is governed by the indicator’s dissociation constant (K_a): $K_a = \frac{[H^+][In^-]}{[HIn]}$. Taking the negative logarithm:

$$pH = pK_a + \log\left(\frac{[In^-]}{[HIn]}\right)$$

When pH = pK_a, [HIn] = [In[−]], meaning equal concentrations of both forms exist, producing an intermediate colour. The human eye can detect colour changes when the ratio of $\frac{[In^-]}{[HIn]}$ is between 0.1 and 10, corresponding to a pH range of approximately pK_a ± 1.

Common indicators

- 1. Methyl Orange: Changes from red (pH < 3.1) to orange-yellow (pH > 4.4); pK_a ≈ 3.7
- 2. Bromothymol Blue: Changes from yellow (pH < 6.0) to blue (pH > 7.6); pK_a ≈ 7.0
- 3. Phenolphthalein: Colourless (pH < 8.2) to pink-purple (pH > 10.0); pK_a ≈ 9.4
- 4. Universal Indicator: Mixture of indicators providing a continuous spectrum of colours across the pH scale.

Natural indicators from plant sources also exist, such as red cabbage extract, which changes from red in acidic solutions to blue-green in basic solutions. The effectiveness of an indicator depends on its ability to display a sharp, easily observable colour change at the appropriate pH range for a particular titration.

2. Choice of indicator in acid-base titration

The selection of an appropriate indicator for a titration depends primarily on the pH at the equivalence point, which varies according to the strengths of the acid and base involved. The indicator's colour change should occur within the steep portion of the titration curve near the equivalence point.

Strong Acid - Strong Base Titrations

In titrations involving strong acids and strong bases (e.g., HCl and NaOH), the equivalence point occurs around pH 7 (at 25°C). The pH changes rapidly near this point, creating a steep vertical section on the titration curve where a small addition of titrant causes a pH change. This sharp transition means the equivalence point occurs over a very narrow pH range.

Because of this sharp transition, most indicators with transition ranges between approximately pH 3-11 would be suitable choices. Indicators such as phenolphthalein (transition range 8.3-10.0), methyl red (transition range 4.4-6.0), bromothymol blue (transition range 6.0-7.6), methyl orange (pH range: 3.1-4.4) and litmus (transition range 5.0-8.0) would all show clear colour changes at or near the equivalence point.

Weak Acid - Strong Base Titrations

When a weak acid is titrated with a strong base (e.g., CH₃COOH with NaOH), the equivalence point occurs at a pH greater than 7 due to the hydrolysis of the salt formed. The exact pH depends on the K_a of the weak acid. Suitable indicators include: Phenolphthalein (pH range: 8.2-10.0) and thymol blue (pH range: 8.0-9.6)

Strong Acid - Weak Base Titrations

In titrations of a strong acid with a weak base (e.g., HCl with NH₃), the equivalence point occurs at a pH less than 7 due to salt hydrolysis. Suitable indicators include, Methyl orange (pH range: 3.1-4.4) and Methyl red (pH range: 4.4-6.2)

Weak Acid - Weak Base Titrations

When both the acid and base are weak, the pH at the equivalence point depends on the relative strengths of the acid and base (K_a and K_b values). These titrations often have poorly defined equivalence points with gradual pH changes, making indicator selection challenging.

Generally, an indicator with a transition range close to the calculated equivalence point pH should be chosen, but these titrations may not be practical for precise analytical work.

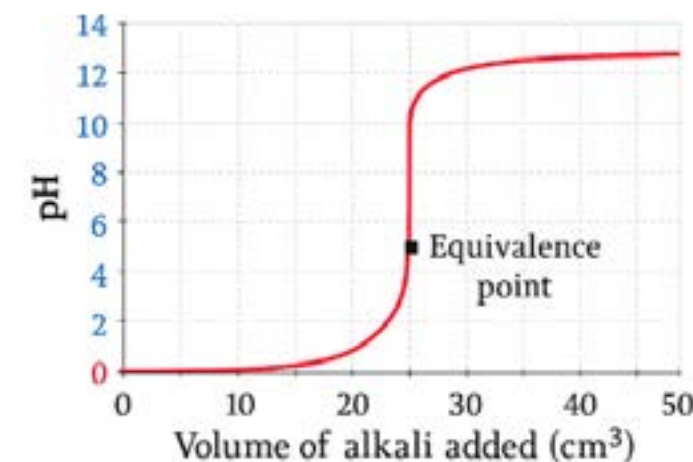
Selection Criteria Summary

1. Calculate or estimate the pH at the equivalence point based on the acid and base strengths
2. Select an indicator whose transition range encompasses this pH
3. Consider the sharpness of the colour change and visibility against the solution background
4. Ensure the indicator does not react with the analyte or titrant
5. For polyprotic acids or bases, consider that multiple equivalence points may exist, requiring different indicators for each endpoint

3. Titration curves

Titration curves are graphical representations of pH versus volume of titrant added during an acid-base titration. These curves provide valuable information about the titration process, including the initial pH, equivalence point, buffer regions, and appropriate indicator ranges.

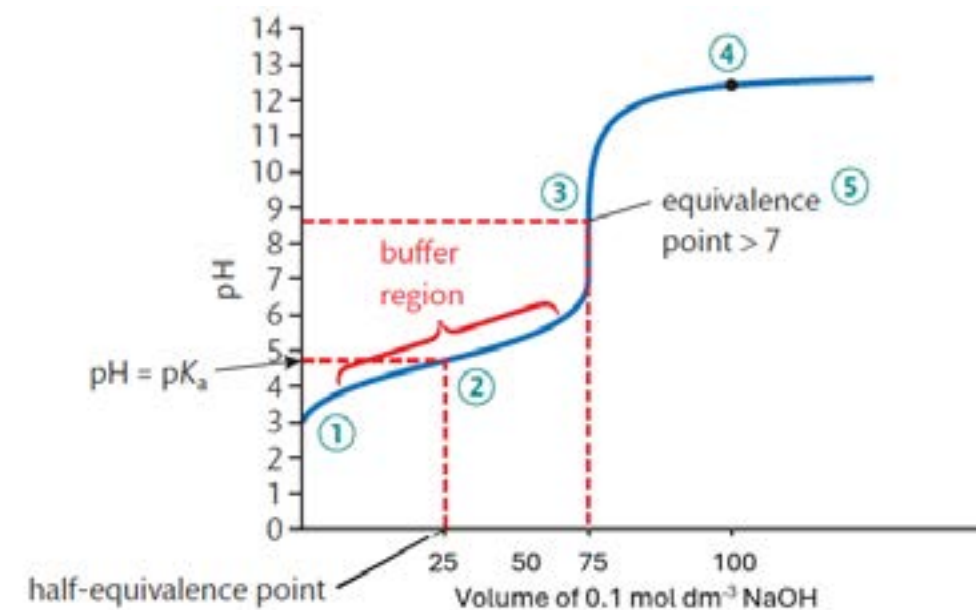
Strong Acid - Strong Base Titration Curve



When titrating a strong acid with a strong base, the curve begins at a low pH (typically 1-2 for 0.1M strong acid). The pH increases slightly during initial base addition. Near the equivalence point, a very steep rise in pH occurs. The equivalence point is at pH 7, after the equivalence point, pH increases more gradually, approaching the pH of the excess base.

The curve is symmetrical around pH 7, and the steep portion typically spans 3-4 pH units over a very small volume range, making indicator selection relatively flexible.

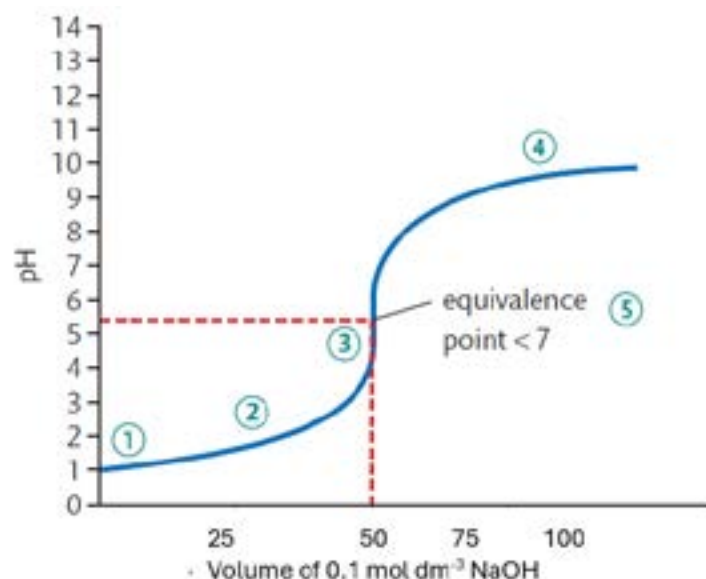
Weak Acid - Strong Base Titration Curve



When titrating a weak acid with a strong base, the curve begins at a higher pH than for a strong acid (typically 3-4 for 0.1M weak acid). The pH rises more gradually initially, creating a buffer region at approximately pH = pK_a . The steepest portion of the curve occurs just before the equivalence point. The equivalence point is above pH 7 (typically 8-10), after the equivalence point, the curve resembles that of a strong acid-strong base titration. The buffer region, where

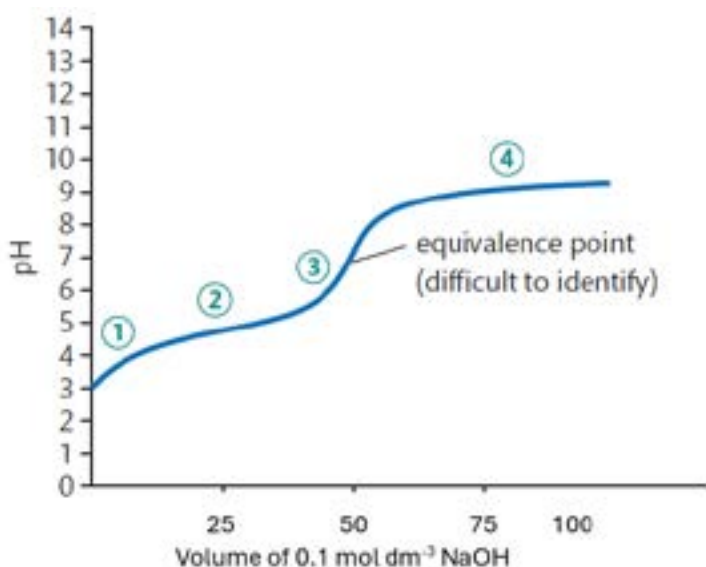
$\text{pH} \approx \text{pK}_a$, is characterised by relatively small pH changes despite significant titrant addition. This region corresponds to when approximately half of the weak acid has been neutralised.

Strong Acid - Weak Base Titration Curve



When titrating a strong acid with a weak base, the curve begins at a low pH typical of the strong acid. The pH increases gradually as base is added, the steepest portion occurs just after the halfway point. The equivalence point is below pH 7 (typically 4-6) and after the equivalence point, a buffer region forms, followed by more gradual pH increases.

Weak Acid - Weak Base Titration Curve



When both the acid and base are weak, the curve has a much less pronounced S-shape. The pH change at the equivalence point is gradual rather than steep. The equivalence point pH depends on the relative strengths of the acid and base. Buffer regions exist both before and after the equivalence point. These titrations generally have poorly defined equivalence points, making them less suitable for analytical determinations.

Drawing Titration Curves

To accurately draw a titration curve:

1. Calculate the initial pH before any titrant is added

2. Calculate the pH after adding small volumes of titrant (more points needed near the equivalence point)
3. Calculate the pH at the equivalence point using appropriate equilibrium considerations
4. Calculate several points after the equivalence point
5. Plot the pH versus volume of titrant added
6. Connect the points with a smooth curve
7. Mark the equivalence point and any buffer regions
8. Indicate suitable indicator ranges on the curve

Using Titration Curves to Select Indicators

The titration curve provides a visual guide for indicator selection:

1. Locate the equivalence point on the curve
2. Identify the steepest portion of the curve near this point
3. Select an indicator whose transition range falls within this steep region
4. Mark the indicator's transition range on the curve to verify its suitability

Learning Tasks

1. Recap calculations of pH of buffer solutions using the Henderson-Hasselbalch equation
 - a. i.e., $\text{pH} = \text{pK}_a + \log \left(\frac{[\text{Salt}]}{[\text{Acid}]} \right)$
 - b. Make a prediction of the outcome based on the physical properties of acids and bases.
2. Calculate the pH of ethanoic acid solution before titration begins.
 - a. Create a data table showing pH values at different volumes of added NaOH.
 - b. Plot points on graph paper to visualise the titration curve.
 - c. Mark and label the initial pH and equivalence point.
 - d. Compare different indicators by shading their pH ranges on the graph.
 - e. Explain why a particular indicator (phenolphthalein) works best for this titration.
 - f. Answer the original assessment question individually.

Pedagogical Exemplars

Collaborative Learning

1. Form diverse groups of 3-4 learners, ensuring gender balance and inclusion of different ability levels.
2. Provide each group with a set of common indicators (phenolphthalein, methyl orange, bromothymol blue) and solutions of varying pH (1-14).

- 3. Third, instruct groups to test each indicator in the different solutions and record the colour changes observed.
- 4. Have learners determine the approximate transition range for each indicator by identifying where colour changes occur.
- 5. Ask groups to create a digital chart or presentation showing their findings, including photographs of colour changes.
- 6. Facilitate a whole-class discussion where groups share their results and compare observations.
- 7. Guide learners to connect their observations to the theoretical concept of indicators as weak acids or bases with different coloured molecular forms.

Inquiry-Based Learning

- 1. Present learners with four titration scenarios: strong acid-strong base, weak acid-strong base, strong acid-weak base, and weak acid-weak base.
- 2. Provide pH data showing how pH changes during each type of titration, without identifying which scenario is which.
- 3. Challenge learners to match each data set to the correct titration type based on initial pH, final pH, and curve shape.
- 4. Ask learners to predict the approximate pH at the equivalence point for each titration type.
- 5. Provide information about common indicators and their transition ranges.
- 6. Have learners select the most appropriate indicator for each titration scenario and justify their choices.
- 7. Conduct demonstration titrations using the selected indicators to verify predictions.
- 8. Guide reflection on the accuracy of predictions and the consequences of choosing inappropriate indicators.

Visual Learning

- 1. Demonstrate the construction of a titration curve using either digital tools or graph paper, emphasising key points to plot.
- 2. Provide learners with titration data for a strong acid-strong base titration.
- 3. Guide learners through plotting their first curve, identifying initial pH, equivalence point, and steep portion.
- 4. Distribute data for the remaining titration types (weak acid-strong base, strong acid-weak base, weak acid-weak base).
- 5. Have learners work in pairs to plot these curves, comparing the shapes and features of each.
- 6. Instruct learners to mark the transition ranges of different indicators on their curves.
- 7. Facilitate analysis of which indicators would be appropriate for each titration type based on where their transition ranges fall relative to the equivalence point.
- 8. Have learners digitally overlay all four curves to create a comprehensive comparison chart.

Experiential Learning

- 1. Demonstrate proper titration technique, emphasising safety considerations and precise measurement.
- 2. Divide learners into mixed-gender, mixed-ability groups and assign each group a different titration type.
- 3. Have each group select an appropriate indicator for their assigned titration based on theoretical knowledge.
- 4. Instruct groups to perform their titrations, recording initial and final burette readings.
- 5. Ask learners to calculate the concentration of their unknown solution using their titration results.
- 6. Have groups repeat their titrations using a different indicator that should theoretically be less appropriate.
- 7. Guide learners to compare results obtained with different indicators and evaluate the impact on accuracy.
- 8. Facilitate a reflection discussion on the connection between theoretical titration curves and practical endpoint detection.

Key Assessment (DoK LEVEL 3)

- 1. A buffer solution is prepared by mixing 50.0 cm³ of 0.20 mol/dm³ CH₃COOH with 25.0 cm³ of 0.30 mol/dm³ CH₃COONa.
[K_a = 1.8 × 10⁻⁵]
 - a. Calculate the pH of this buffer solution. (3 marks)
 - b. How would this buffer solution appear if universal indicator were added? Explain. (2 marks)
- 2. Ethanoic acid (CH₃COOH) is a weak acid with K_a = 1.8 × 10⁻⁵. When 25.0 cm³ of 0.1 mol/dm³ ethanoic acid is titrated with 0.1 mol/dm³ NaOH:
 - a. Calculate the pH of the ethanoic acid solution before any NaOH is added. (3 marks)
 - b. Sketch the titration curve, clearly marking the initial pH and equivalence point. (3 marks)
 - c. Explain why phenolphthalein (pH range 8.3-10.0) would be the most suitable indicator for this titration. (1 mark)
- 3. A student conducts a titration experiment but is unsure which indicator to use. The student has collected the following pH data during the titration:

Volume of base added (cm ³)	pH
0	1.3
5	1.6
10	1.9
15	2.2

20	2.8
22	3.1
24	6.8
24.5	8.7
25	10.3
26	11.4
30	12.0

- a. Plot the titration curve using the data provided. Label the axes and identify the equivalence point on your graph. (4 marks)
- b. Based on your titration curve, determine whether this represents a strong acid-strong base, weak acid-strong base, or strong acid-weak base titration. Justify your answer by analysing the shape of the curve and the pH at the equivalence point. (4 marks)

The student has the following indicators available:

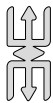
- i. Methyl orange (pH range: 3.1-4.4)
- ii. Bromothymol blue (pH range: 6.0-7.6)
- iii. Phenolphthalein (pH range: 8.3-10.0)
- iv. Alizarin yellow (pH range: 10.1-12.0)

Which indicator would be most suitable for this titration? Explain your reasoning by referring to the titration curve and the pH ranges of the indicators. (4 marks)

PLC Session 7

45 minutes	Planning & Reviewing the Learning Plan
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- 1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
- 2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week’s concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
- 3. Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



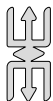
Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

- 4. Plan the week’s key assessment with your app
 - a. The key assessment for the week is *individual class exercise*. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the exercise (NTS 3k-3q).

10 minutes	Content Exploration
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- 5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
 - a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step
 - d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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- 6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
 - a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes

Reflection

7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
 - a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

Section 3 Review

This section has focused on understanding acid-base indicators and their application in titration procedures. Over the past week, learners have explored the chemical principles behind indicators, developed skills in selecting appropriate indicators for various titration scenarios, and mastered the interpretation and construction of titration curves.

Key Concepts Covered

Acid-Base Indicators: Learners should understand that indicators are weak acids or bases with different coloured molecular and ionic forms. They should be familiar with common indicators such as phenolphthalein and methyl orange including their transition pH ranges and colour changes.

Indicator Selection: Learners should be able to select appropriate indicators for different titration types based on the expected pH at the equivalence point. They should understand how the nature of the acid and base affects this selection, particularly for strong acid-strong base, strong acid-weak base, weak acid-strong base, and weak acid-weak base combinations.

Titration Curves: Learners should be proficient in interpreting and drawing titration curves for various acid-base combinations. They should be able to identify key features including initial pH, equivalence point, buffer regions, and endpoints. Most importantly, they should understand how these curves guide indicator selection.

MARKING SCHEMES FOR THE KEY ASSESSMENTS

WEEK 7

- i. Buffer Solution
- i. Calculate the pH of this buffer solution. (3 marks)

Mark Allocation:

- a. Correct calculation of moles of CH_3COOH : $0.050 \text{ dm}^3 \times 0.20 \text{ mol/dm}^3 = 0.010 \text{ mol}$ (1 mark)

- b. Correct calculation of moles of CH_3COONa : $0.025 \text{ dm}^3 \times 0.30 \text{ mol/dm}^3 = 0.0075 \text{ mol}$ (1 mark)

- c. Correct application of Henderson-Hasselbalch equation and final answer:

$$\text{pH} = \text{pK}_a + \log\left(\frac{[\text{CH}_3\text{COONa}]}{[\text{CH}_3\text{COOH}]}\right)$$

$$\text{pH} = -\log(1.8 \times 10^{-5}) + \log(0.0075/0.010)$$

$$\text{pH} = 4.74 + \log(0.75)$$

$$\text{pH} = 4.74 - 0.12$$

$$= 4.62 \text{ (1 mark)}$$

- ii. How would this buffer solution appear if universal indicator were added? Explain. (2 marks)

Mark Allocation:

- i. Correct identification of colour: orange/yellow-orange (1 mark)
- ii. Explanation: The universal indicator would show orange/yellow-orange colour because the pH of the solution is approximately 4.6, which falls in the orange range (4-5) on the universal indicator colour chart (1 mark)

ii. Ethanoic Acid Titration

- a. Calculate the pH of the ethanoic acid solution before any NaOH is added. (3 marks)

Mark Allocation:

1. Recognition that for a weak acid, pH calculation requires solving the equilibrium: $\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+ + \text{CH}_3\text{COO}^-$ (1 mark)

2. Correct setup of the equilibrium expression:

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$\text{Let } x = [\text{H}^+] = [\text{CH}_3\text{COO}^-] \text{ and } [\text{CH}_3\text{COOH}] \approx 0.1 \text{ mol/dm}^3$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.1}$$

$$x^2 = 1.8 \times 10^{-6}$$

$$x = 1.34 \times 10^{-3} \text{ (1 mark)}$$

3. Correct calculation of pH: $\text{pH} = -\log(1.34 \times 10^{-3})$
 $= 2.87$ (1 mark)

- b. Sketch the titration curve, clearly marking the initial pH and equivalence point. (3 marks)

Mark Allocation:

1. Correct shape of curve for weak acid-strong base titration (gradual rise at beginning, steep rise at equivalence point, gradual levelling at end) (1 mark)
2. Correct marking of initial pH at approximately 2.9 (1 mark)

3. Correct marking of equivalence point at 25.0 cm³ with pH above 7 (approximately 8.7-9.0) (1 mark)
- c. Explain why phenolphthalein (pH range 8.3-10.0) would be the most suitable indicator for this titration. (1 mark)

Mark Allocation

Correct explanation: Phenolphthalein is most suitable because its pH transition range (8.3-10.0) includes the pH at the equivalence point of the titration (approximately 8.7-9.0), which ensures a sharp colour change from colourless to pink exactly at the equivalence point (1 mark)

iii. Titration Curve Analysis and Design (17 marks)

i. Plotting the titration curve (4 marks)

1. Correct labelling of x-axis as "Volume of Base Added (cm³)" (1 mark)
2. Correct labelling of y-axis as "pH" (1 mark)
3. Accurate plotting of all data points (1 mark)
4. Correct identification of equivalence point at approximately 24.5 cm³ where the steepest part of the curve occurs (pH around 8.7) (1 mark)

ii. Determining the type of titration (4 marks)

1. Correct identification as a weak acid-strong base titration (1 mark)
2. Justification based on:
 - a. The pH at the equivalence point is above 7 (approximately 8.7) (1 mark)
 - b. The initial pH is not extremely low (starting at 1.3 rather than ~0) (1 mark)
 - c. The buffer region before the equivalence point shows a gradual increase in pH (1 mark)

iii. Choosing the most suitable indicator (4 marks)

- Correct identification of phenolphthalein as the most suitable indicator (1 mark)
- Explanation that:
 - i. The equivalence point occurs at approximately pH 8.7 (1 mark)
 - ii. Phenolphthalein's transition range (8.3-10.0) includes the equivalence point pH (1 mark)
 - iii. The steep portion of the curve passes through phenolphthalein's transition range, ensuring a sharp colour change at the endpoint (1 mark)

SECTION 4: OXIDATION-REDUCTION REACTIONS AND ITS APPLICATIONS

STRAND: PHYSICAL CHEMISTRY

Sub-Strand: *Equilibria*

Learning Outcome: Describe oxidation and reduction reactions and apply their principles to electrochemical cells as well as their importance in everyday life

Content Standard: Demonstrate knowledge and understanding of redox and apply its principles to electrochemical cells as well as their importance in everyday life

INTRODUCTION AND SECTION SUMMARY

This section introduces learners to oxidation-reduction (redox) reactions and their wide-ranging applications. Learners will develop an understanding of redox processes from basic definitions to complex applications in electrochemistry and metallurgy. By the end of this section, learners will be able to explain oxidation and reduction in terms of electron transfer and oxidation number changes, identify oxidising and reducing agents, balance redox equations, perform redox titrations, understand electrochemical and electrolytic cells, and explain corrosion processes. This knowledge forms a critical foundation for understanding many natural and industrial processes, from cellular respiration to battery technology and metal extraction.

The weeks covered by this section are:

- Weeks 8-9: Explain oxidation and reduction reactions and illustrate reactivity of metals experimentally
- Weeks 10-12: Explain the terms oxidising agent, reducing agent, redox half equations and apply the principles in redox titrations.
- Weeks 13-14: Explain how redox is used to produce electricity in storage and fuel cells.
- Weeks 15-16: Explain qualitatively and quantitatively the operations of electrolytic cell as well as their applications in everyday life.
- Week 17: Explain the concept of corrosion of metals and its application in everyday life.

SUMMARY OF PEDAGOGICAL EXEMPLARS

Teachers should employ diverse pedagogical approaches to develop learners' understanding of redox chemistry and its applications. Talk-for-learning activities should introduce the fundamental concepts of oxidation and reduction using multiple definitions (oxygen transfer, hydrogen transfer, electron transfer, and oxidation number changes). Collaborative learning

should then be used to apply these concepts to various reaction types, with groups analysing combination, decomposition, combustion, displacement, disproportionation, photosynthesis, and respiration reactions.

For investigating metal reactivity, exploratory learning approaches should be employed, with mixed-ability or mixed-gender groups designing and performing experiments to establish reactivity patterns. Digital learning resources can supplement these activities with videos demonstrating metal reactivity.

When teaching oxidising and reducing agents, think-pair-share discussions should help learners identify these agents based on oxidation number changes. Exploratory learning should include qualitative tests for oxidising and reducing agents, while talk-for-learning activities should connect these concepts to real-world applications in various industries.

For redox equations, teachers should guide learners through the process of writing half-equations and balancing complex redox reactions in both acidic and basic media. This provides the foundation for redox titrations, which should be taught through inquiry-based learning, with groups designing and performing titrations involving potassium permanganate and iodometric methods. Problem-solving approaches should be used to develop calculation skills for determining analyte concentrations.

Electrochemical cells should be explored through collaborative learning, with gender-based groups investigating factors affecting electrode potentials. Talk-for-learning and digital resources should be used to explain standard electrode potentials and voltaic cell operation. Problem-solving approaches should develop skills in calculating cell potentials and predicting reaction feasibility, while exploratory learning should involve practical investigation of various battery types and fuel cells.

For electrolytic cells, collaborative and activity-based learning should help learners understand the components and processes involved. Talk-for-learning activities should address selective discharge factors, while inquiry-based learning should involve practical investigation of electrolysis products and applications like electroplating.

Finally, corrosion concepts should be taught through think-pair-share discussions, experimental investigations, and talk-for-learning activities that connect to real-world impacts. Throughout all topics, teachers should emphasize connections between theoretical concepts and practical applications, using diverse grouping strategies to promote inclusive participation.

ASSESSMENT SUMMARY

Indicator	DOK Level	Assessment Strategy and/or Mode
Explain oxidation and reduction reactions and illustrate reactivity of metals experimentally	Level 3: Strategic Reasoning	Group work and Hands-on Practical test or Prediction tasks or Sequencing activity
Explain the terms oxidising agent, reducing agent, redox half equations and apply the principles in redox titrations	Level 4: Extended Thinking	Written Assessment, Individual Classroom exercise and Hands-on Practical
Explain how redox is used to produce electricity in storage and fuel cells	Level 4: Extended Thinking	Written, Diagram-based, and Calculation-based

Explain qualitatively and quantitatively the operations of electrolytic cells as well as their applications in everyday life	Level 4: Extended Critical Thinking	Group Research, Problem-solving
Explain the concept of corrosion of metals and its application in everyday life	Level 4: Extended Critical Thinking	Field research and community assessment

WEEKS 8 & 9

Learning Indicator: Explain oxidation and reduction reactions and illustrate reactivity of metals experimentally

Recap weeks

Year 1	Week 17 1. Use the electron configuration of elements to determine their position on the periodic table. 2. Explain how periodic properties change with atomic number and principal quantum number
Year 2	Week 10: Describe the physical and chemical properties of acids and bases. Week 14: Describe the patterns in physical and chemical properties of the period 3 elements.

FOCAL AREA 1: OXIDATION AND REDUCTION (REDOX) REACTIONS

Redox reactions are fundamental chemical processes that occur all around us in Ghana. They play crucial roles in our daily lives, our economy, and our natural environment.

In Ghana's gold mining regions like Obuasi and Tarkwa, redox reactions are essential for extracting gold from ore. The process uses cyanide solutions to oxidise gold, allowing it to be separated from rock.

In northern Ghana, blacksmiths have used redox chemistry for centuries. When they heat iron ore with charcoal in traditional furnaces, reduction reactions convert iron oxide to pure iron for making tools and ceremonial items.

Ghana's primary export crop undergoes redox reactions during fermentation. The oxidation of compounds in cocoa beans develops the characteristic flavour and colour that makes Ghanaian cocoa world-renowned.

In coastal communities like Elmina and Cape Coast, the salt-laden air accelerates the rusting of metal structures, a common oxidation reaction that costs our economy millions of cedis annually in repairs and replacements, etc.

These everyday examples demonstrate that redox reactions are not just theoretical concepts but practical processes that impact our lives in countless ways.

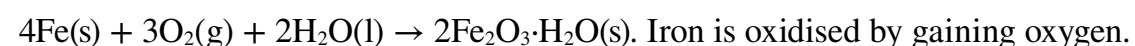
Definitions of Oxidation and Reduction

Oxidation and reduction can be defined in a number of different ways. The definition is often chosen to suit the context of the reaction. Much like how the definitions of acids and bases have changed over time the same is true for oxidation and reduction.

1. Oxidation reactions

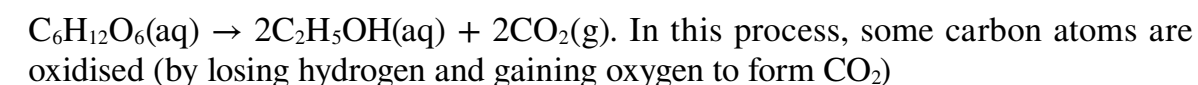
- a. The addition of oxygen to a substance.

Example, rusting of iron is common in coastal Ghana:



- b. The removal of hydrogen from a substance.

Example, fermentation of palm wine:



- c. Oxidation is the loss of electrons from a substance.

Example, $\text{Ca(s)} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$ (Calcium loses two electrons)

- d. Oxidation is an increase in the oxidation number of an element.

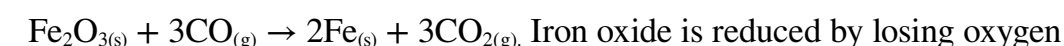
Example, combustion of charcoal $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$

Oxidation number changes: C: 0 $\rightarrow$ +4 (increase)

2. Reduction reactions

- a. The removal of oxygen from a substance.

Example, reduction of iron ore in traditional smelting:



- b. The addition of hydrogen to a substance.

In the production of ammonia for fertilisers $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$, nitrogen gains hydrogen atoms.

- c. Reduction is the gain of electrons by an element. Example, $\text{O}_2 + 4\text{e}^- \rightarrow 2\text{O}^{2-}$ (Oxygen gains electrons - reduction)

- d. Reduction is a decrease in oxidation number of an element.

Example, photosynthesis in trees



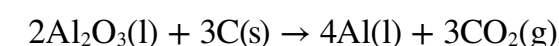
C: +4 (in CO_2) $\rightarrow$ 0 (in $\text{C}_6\text{H}_{12}\text{O}_6$) (decrease - reduction)

3. **Redox reaction:** Combining all four perspectives, a redox reaction can be defined as a reaction in which electrons are transferred from one reactant to another, resulting in changes in oxidation states. Examples of redox reactions include:

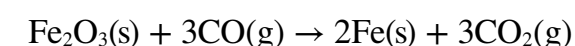
- a. Gold mining and mineral processing:



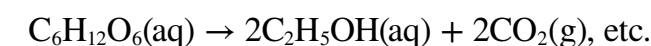
- b. Aluminium production at VALCO:



- c. Blacksmithing:



- d. Ethanol production during fermentation:



Learning Tasks

- Define the following terms in terms of oxygen transfer, hydrogen transfer, electron transfer, and change in oxidation number:
 - Oxidation
 - Reduction
- The extraction of iron in traditional furnaces involves the reduction of iron(III) oxide by carbon: $\text{Fe}_2\text{O}_3 + 3\text{C} \rightarrow 2\text{Fe} + 3\text{CO}$
 - Identify which substance is being oxidised and which is being reduced.
 - Explain this reaction in terms of
 - Oxygen transfer
 - Electron transfer
 - Changes in oxidation number

Pedagogical Exemplars

- Begin with a whole-class discussion using familiar contexts to activate prior knowledge by showing images of rusted metal roofing, jewellery, pictures of traditional iron smelting, etc.
- Ask learners questions like:

What changes do you observe in these materials?

What do you think might be happening at the atomic level?

Has anyone observed similar changes in their community?
- Capture learners' ideas on the board, using their responses as a foundation for introducing formal redox concepts.
- Introduce the four definitions of oxidation and reduction using clear visual aids and one example from their daily life that might involve redox reactions. Circulate, listen to discussions, and note misconceptions to address later.
- Return to whole-class format and work through clear examples for each definition. Example rusting of iron: $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$, methane combustion:

$$\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \text{ etc}$$
- Organise learners into groups of three for collaborative analysis of different reaction types and provide clear instructions.
- Provide each group with worksheets containing five examples of reaction types. Groups must determine which elements are oxidised and which are reduced in each reaction using oxidation numbers as their primary analytical tool
- Every group member must be prepared to explain at least one reaction type. Circulate among groups, providing targeted support and asking probing questions to deepen understanding.
- Bring the class back together for a brief consolidation of learning and ask 2-3 volunteers to share one key insight from their group discussions. Summarise the three approaches to identifying oxidation and reduction.

- Emphasise that all four definitions describe the same fundamental process from different perspectives. Connect the concepts back to the real-world Ghanaian examples from the beginning of the lesson.

Key Assessment Level 3: Group Work

- Explain the relationship between:
 - Loss of electrons and increase in oxidation number
 - Gain of electrons and decrease in oxidation number (10 marks)
- The Obuasi gold mine in Ghana uses cyanide leaching to extract gold from its ore. The process involves the following reaction:

$$4\text{Au} + 8\text{NaCN} + \text{O}_2 + 2\text{H}_2\text{O} \rightarrow 4\text{Na}[\text{Au}(\text{CN})_2] + 4\text{NaOH}$$
 - Determine the oxidation number of gold in the ore and in the complex $\text{Na}[\text{Au}(\text{CN})_2]$.
 - Identify which element is oxidised and which is reduced in this process. (10 marks)
- Why most metals are found in nature as compounds rather than in their elemental form. (10 marks)

FOCAL AREA 2: REACTIVITY OF METALS

The reactivity of metals refers to their tendency to lose electrons and form positive ions (cations). This property determines how readily metals react with other substances such as water, acids, and oxygen.

The Reactivity Series

Metals can be arranged in order of their reactivity, from most reactive to least reactive. Here is an example of how this can be done:

Potassium (K) [**Most Reactive**] > Sodium (Na) > Calcium (Ca) > Magnesium (Mg) > Aluminium (Al) > Zinc (Zn) > Iron (Fe) > Tin (Sn) > Lead (Pb) > Copper (Cu) > Silver (Ag) > Gold (Au) > Platinum (Pt) [**Least Reactive**].

Experimental Determination of Reactivity of Metals

The reactivity of metals can be determined experimentally by observing how they react with water, acids, oxygen, and through displacement reactions. These experiments allow for the construction of a reactivity series based on direct observation.

Required materials

Metal samples (e.g., magnesium, zinc, iron, copper, aluminium), test tubes and test tube rack, dilute HCl, dilute H_2SO_4 , distilled water, bunsen burner, tongs, safety goggles and protective gloves

Solutions of metal salts [e.g., copper(II) tetraoxosulphate(VI), zinc(II) tetraoxosulphate(VI), iron(II) tetraoxosulphate(VI)], pH indicator (e.g., universal indicator paper), stopwatch.

Safety considerations

Wear safety goggles and protective gloves throughout

Handle acids with care

Keep flammable materials away from the Bunsen burner

Some reactions may be vigorous use small samples

Ensure proper ventilation when gases are produced

Experiment 1: Reaction with cold water

Procedure

SAFETY NOTE: STUDENTS SHOULD NOT BE ALLOWED TO TEST ALKALI METALS OR ALKALI EARTH METALS WITH WATER. THIS SHOULD BE A TEACHER DEMONSTRATION ONLY CONDUCTED BEHIND A SAFETY SCREEN IN A TROUGH NOT IN TEST TUBES.

Please refer to the safety guidance linked below:

<https://science.cleapss.org.uk/Resource/PP122-Lithium-Sodium-Potassium-Reacting-with-Water.pdf>

1. Place small samples of each metal in separate test tubes
2. Add 5 cm³ of cold water to each test tube
3. Observe any reactions (bubbling, heat production, colour changes)
4. Test any gas produced with a glowing splint to identify hydrogen
5. Test the solution with universal indicator to detect formation of alkalis

Expected Results

Potassium and sodium: Violent reaction, rapid hydrogen evolution, solution turns alkaline

Calcium: Moderate reaction, steady hydrogen evolution, solution turns alkaline

Magnesium: Slow reaction, few bubbles, slight pH change

Metals below magnesium: No observable reaction with cold water

Experiment 2: Reaction with steam

Note adapt the safety and procedure of handling magnesium <https://science.cleapss.org.uk/Resource/PX059-Magnesium-and-magnesium-compounds.pdf>

Procedure

1. Set up a steam generator using a flask of boiling water
2. Hold each metal sample in tongs and place it in the steam (not group 1 and 2 metals)
3. Observe any reactions or changes to the metal surface

Expected results

Magnesium: Burns brightly in steam, producing magnesium oxide and hydrogen

Aluminium, zinc, iron: React with steam to form metal oxide and hydrogen

Copper and below: No reaction with steam

Experiment 3: Reaction with dilute HCl

Procedure

1. Place small, equal-sized samples of each metal (not group 1 and 2 metals) in separate test tubes
2. Add 5 cm³ of dilute hydrochloric acid to each test tube
3. Observe the rate of hydrogen gas evolution
4. Measure the time taken for the reaction to complete or a fixed volume of gas to be produced, use a gas syringe to measure the volume.

Expected results

Magnesium: Very rapid reaction, vigorous bubbling

Zinc: Moderate reaction, steady bubbling

Iron: Slower reaction, gentle bubbling

Copper and below: No reaction with dilute acid

Experiment 4: Metal displacement reactions

Procedure

1. Prepare solutions of various metal salts (e.g., AgNO₃, CuSO₄, FeSO₄, ZnSO₄)
 2. Place about 5 cm³ of each salt solution in separate test tubes
 3. Add a small piece of each metal (not group 1 and 2 metals) to the solutions of other metals' salts
 4. Observe any colour changes or metal deposition
- <https://www.youtube.com/watch?v=yO9ssl60XAZo>

Expected results

A more reactive metal will displace a less reactive metal from its salt solution

Example: Zinc placed in copper(II) sulphate solution will produce copper metal and zinc sulphate. No reaction will occur when a less reactive metal is placed in a more reactive metal's salt solution

Data Recording: Create a grid showing which metals can displace which other metals from solution, using + for displacement and - for no reaction.

Experiment 5: Reaction with oxygen

Procedure

1. Hold small samples of each metal (not group 1 and 2 metals or Pb) with tongs
2. Heat each metal in a Bunsen flame
3. Observe how readily each metal burns and the appearance of any oxide formed

Expected results

Magnesium: Burns with a bright white flame (the bright flame can cause permanent eye damage, so, do not observe directly).

Zinc: Burns with a blue-green flame when very hot

Iron: Glows and forms black oxide
Copper: Surface darkens but doesn't burn
Silver, gold: No significant reaction
Record: Ease of ignition, brightness of flame, colour of oxide formed, and time taken for visible oxidation.

Analysis and conclusion

Rank the metals in order of reactivity based on each experiment and compare results across different experiments to create a composite reactivity series
Identify any anomalies or unexpected results and suggest explanations and compare your experimental reactivity series with the standard reactivity series.

Learning Tasks

In this investigation, you will explore the relative reactivity of different metals through a series of experiments. By observing how metals react with water, acid, and metal salt solutions, you will construct your own reactivity series and compare it to the established series.

Materials

Metal samples (magnesium ribbon, zinc granules, iron nails, copper wire, aluminium metal), Test tubes (5) and test tube rack, dilute hydrochloric acid (0.5 mol/dm³) distilled water, solutions of copper(II) sulphate, zinc sulphate, and iron(II) sulphate, universal indicator paper, recording sheet.

Safety equipment

goggles and gloves, Stopwatch, Thermometer

Safety precautions

Wear safety goggles and gloves throughout the experiment
Handle acids with care
Report any spills immediately
Wash hands thoroughly after the experiment
Follow proper disposal procedures for all chemicals

Reaction with dilute acid

1. Label 5 test tubes with the names of the metals being tested
2. Place a small, similar-sized sample of each metal in its corresponding test tube
3. Add 5 cm³ of dilute hydrochloric acid to each test tube
4. Immediately start timing and observe the reactions
5. Record your observations in the data table, noting:
 - a. Whether a reaction occurs
 - b. The time it takes for bubbling to begin
 - c. The intensity of the reaction (rate of bubbling)

- d. Any colour changes or other observations
6. Rank the metals from most reactive (1) to least reactive (5) based on your observations.

Reactivity with dilute HCl

Metal	Reaction observed (Yes/No)	Time for bubbling to start (seconds)	Intensity of reaction (1-5 scale)	Other observation	Reactivity Ranking
Magnesium					
Zinc					
Iron					
Aluminium					
copper					

Metal Displacement Reactions

1. Set up 3 rows of test tubes (one row for each metal salt solution)
2. Label each row with the name of the metal salt solution
3. Place 5 cm³ of copper(II) sulphate solution in each test tube of the first row
4. Place 5 cm³ of zinc sulphate solution in each test tube of the second row
5. Place 5 cm³ of iron(II) sulphate solution in each test tube of the third row
6. Add a small piece of each metal to one test tube in each row (except don't add a metal to its own salt solution)
7. Observe for 5-10 minutes and record any changes, particularly:
 - a. Colour changes in the solution
 - b. Formation of a coating on the metal
 - c. Disappearance of the original metal
8. Record your observations in the displacement reaction table

NB: Mark "Y" if displacement occurred (metal deposited or colour change), "N" if no reaction observed.

Displacement Reaction Table

- 1.

Metal	CuSO ₄ solution	ZnSO ₄ solution	FeSO ₄ solution
Magnesium			
Zinc			
Iron			
Aluminium			

Based on your observations from the two tables above, arrange the metals in order of decreasing reactivity (most reactive to least reactive).

- Based on your observations from displacement table, which metals were able to displace other metals from their solutions? What pattern do you notice?
- Combine your results from the two tables to create a final reactivity series. How does your series compare with the standard reactivity series shown below? Standard series: Magnesium > Aluminium > Zinc > Iron > Copper
- Explain any differences between your experimental reactivity series and the standard series. What factors might have affected your results?
- How could you improve this investigation to get more accurate results?
- Based on your findings, predict whether iron would react with:
 - Cold water
 - Steam
 - Magnesium sulphate solution. Explain your reasoning for each prediction.

Teacher demo: Thermite (showing the actual displacement of iron by aluminium).

<https://science.cleapss.org.uk/Resource/SRA026-Thermite-reaction-in-flower-pots.pdf>

<https://www.youtube.com/watch?v=1N50IjDB5Sk>

Pedagogical Exemplars

Exploratory Learning

- Form mixed-ability groups of 3-4 learners. Groups discuss prior knowledge of metal reactivity and brainstorm experiments to test which metals are more reactive than others.
- Groups formalise their experimental plans by selecting metals to test, choosing reactions to investigate (with acid, water, or salt solutions), determining measurements, and creating data tables.
- Groups briefly present their plans, receive safety and design feedback from the teacher, and refine their approach accordingly.
- Groups collect materials and conduct their planned experiments, possible experiments include:
 - Comparing reaction rates of metals with dilute acid
 - Testing which metals react with water
 - Investigating displacement reactions between metals and salt solutions
 - Groups record observations and measurements in their data tables
- Teacher circulates to ensure safety and provide guidance as needed.
- Groups organise their data, rank metals from most to least reactive based on their observations, compare results from different experiments, and create visual representations of their findings.

- Groups discuss observed patterns in metal reactivity, compare results with expectations, consider factors affecting their results, and prepare a summary of their findings.
- Each group shares their experimental approach and key findings. The class compares results across groups to identify consistent patterns in metal reactivity.
- Learners reflect individually on what they learned about metal reactivity, group collaboration, and real-world applications. The teacher connects findings to the established reactivity series.

Key Assessment Level 3: Hands-on Practical test

- In everyday life, copper is often used for water pipes while iron is used for structural supports. Explain why copper is a better choice for water pipes. (10 marks)
- A student tested the reactivity of three metals (X, Y, and Z) by adding them to solutions of their salts. The following observations were recorded:

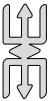
Metal X displaced metal Y from Y’s salt solution

Metal X displaced metal Z from Z’s salt solution

Metal Y displaced metal Z from Z’s salt solution

None of the metals produced hydrogen when added to water at room temperature

Explain these observations in terms of the reactivity series. Identify the possible metals involved (10 marks)



Note

Please use the same instruction template for PLC Sessions 8 and 9, but choose a different assessment strategy for the second week.

PLC Session 9

45 minutes	Planning & Reviewing the Learning Plan
- Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied. - Use your subject specific app to explore the content before planning. Ask the app to - explain the week’s concepts in a practical way that will make it easier for learners to understand - identify two or three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m). - Use your subject specific app to plan your lesson. Ask the app to - suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds. - suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)	

- c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Plan the week’s key assessment with your app
- a. The key assessment for the week is *hands-on practical text*. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the assessment (NTS 3k-3q).

10 minutes	Content Exploration
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5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
- a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
- b. all teachers in the group work through a similar problem individually for 3–4 minutes
- c. the group then works through the problem together step by step
- d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
- a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
- b. After the enactment, ask for structured feedback from the group
- i. was the concept explained in a way that a learner with no prior knowledge could follow?

- ii. were the examples grounded in contexts familiar to Ghanaian learners?
- iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
- a. transfer your chats into your learning planner template.
- b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
- c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

WEEK 10 – 12

Learning Indicator: Explain the terms oxidising agent, reducing agent, redox half equations and apply the principles in redox titrations

FOCAL AREA 1: OXIDISING AND REDUCING AGENTS

Oxidising Agents

An oxidising agent is a substance that accepts electrons in a redox reaction, causing another substance to be oxidised. The oxidising agent itself becomes reduced in the process.

Characteristics of oxidising agents

Gains electrons in redox reactions

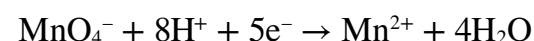
Causes increase in oxidation number of another element

Often contains elements in high oxidation states

Usually contains oxygen or halogens

1. Potassium tetraoxomanganate(VII) (KMnO₄)

KMnO₄ is used in gold mining operations in Obuasi and Tarkwa as part of the extraction process. It helps oxidise impurities in gold ore, allowing for more efficient separation of gold from other minerals. It is also used in water treatment facilities across Ghana to remove iron, manganese, and hydrogen sulphide from drinking water. In acidic solution, manganate(VII) ion acts as a strong oxidising agent:



Visual identification is that the purple solution becomes colourless when reduced to Mn²⁺.

2. Hydrogen Peroxide (H₂O₂)

H₂O₂ is used in textile industries in Tema and Accra for bleaching traditional fabrics like kente and adinkra cloths. It is also commonly found in Ghanaian hospitals and clinics as an antiseptic for wound cleaning. Local palm oil processors sometimes use dilute H₂O₂ to improve the colour and quality of palm oil. H₂O₂ acts as an oxidising agent in many reactions: $\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$. The colourless liquid of H₂O₂ produces oxygen bubbles when decomposing.

3. Dichromate(VI) Ion (Cr₂O₇²⁻):

Cr₂O₇²⁻ is used in leather tanning industries in Kumasi where animal hides are processed. The dichromate(VI) helps in the chromium tanning process, making leather more durable and resistant to heat. It is also used in school laboratories across Ghana for testing the presence of alcohol in organic compounds. In acidic solution:

$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$. The orange solution of Cr₂O₇²⁻ turns green when reduced to Cr³⁺.

Reducing Agents

A reducing agent is a substance that donates electrons in a redox reaction, causing another substance to be reduced. The reducing agent itself becomes oxidised in the process.

Characteristics of reducing agents

1. Loses electrons in redox reactions
2. Causes decrease in oxidation number of another element
3. Often contains elements in low oxidation states
4. Often includes metals, hydrogen, and carbon compounds

Examples of reducing agents

1. Zinc metal (Zn)

Zinc is used extensively in galvanising iron roofing sheets which are common across Ghana, from urban areas like Accra to rural villages. The zinc acts as a sacrificial anode, protecting iron from corrosion. It is also used in local battery production and in brass-making for traditional crafts in Krofrom, Kumasi.

Chemical behaviour: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$

Zinc is a silvery-white metal that dissolves in acids with hydrogen evolution.

2. Sulphur(IV) oxide (SO₂)

Sulphur(IV) oxide is used in food preservation, particularly in preserving dried fruits and vegetables in local markets. In Ghana's cocoa processing industry, sulphur dioxide is sometimes used to prevent unwanted oxidation during processing. It is also used in wine production in the Volta Region to inhibit bacterial growth.

Chemical behaviour: $\text{SO}_2 + 2\text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4 + 2\text{H}^+ + 2\text{e}^-$

Sulphur dioxide is a colourless gas with a pungent smell which turns acidified potassium dichromate(VI) from orange to green.

3. Iron(II) ions (Fe²⁺)

Iron(II) ions are found naturally in groundwater in many parts of Ghana, particularly in the Northern and Upper regions. Iron(II) compounds are used in traditional herbal medicine preparations across Ghana. They are also relevant in Ghana's iron ore mining operations in the Western Region.

Chemical behaviour: $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$

Pale green solutions of Fe²⁺ turn yellow-brown when oxidised to Fe³⁺.

Tests for Oxidising and Reducing Agents

Oxidising Agents

1. Potassium Iodide Test

Add acidified potassium iodide solution to the test substance.

Positive Result: Brown/yellow colour or precipitate (I₂) indicates an oxidising agent. Oxidising agent + 2I⁻ → I₂ + reduced form of oxidising agent.

2. Acidified potassium manganate(VII) Test

Add dilute acidified KMnO₄ solution to the test substance.

Decolourisation of the purple manganate(VII) solution indicates a reducing agent:

$$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$$

3. Litmus paper test

Moisten blue litmus paper and place it near or in contact with the test substance.

Bleaching (turning white) indicates a strong oxidising agent like chlorine or hydrogen peroxide.

Reducing Agents

1. Acidified potassium dichromate(VI) Test

Add acidified potassium dichromate(VI) solution to the test substance.

Colour change from orange to green indicates a reducing agent:

$$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$$

2. Benedict’s or Fehling’s solution test

Add Benedict’s or Fehling’s solution to the test substance and heat.

Formation of a red-brown precipitate indicates a reducing sugar:

$$\text{Cu}^{2+} + \text{reducing agent} \rightarrow \text{Cu}_2\text{O (red precipitate)}$$

3. Silver nitrate test

Add silver nitrate solution to the test substance.

Formation of silver metal (grey-black precipitate) indicates a strong reducing agent:

$$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$$

Applications of Oxidising and Reducing Agents

Oxidising Agents

1. Water Purification: Chlorine and chlorine compounds act as oxidising agents to kill pathogens: $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$
- Ghana Water Company Limited uses chlorine-based oxidising agents to treat drinking water in urban centres like Accra, Kumasi, and Tamale. In rural communities, organisations like Community Water and Sanitation Agency promote the use of chlorine tablets for household water treatment, especially important in areas with high incidence of waterborne diseases.
2. Mining Industry: Cyanide oxidation of gold in the presence of oxygen:
- $$4\text{Au} + 8\text{CN}^- + \text{O}_2 + 2\text{H}_2\text{O} \rightarrow 4[\text{Au}(\text{CN})_2]^- + 4\text{OH}^-$$
- Ghana’s gold mining industry, centred in Obuasi, Tarkwa, and Prestea, relies heavily on oxidising agents in the extraction process. Companies like AngloGold Ashanti and Goldfields Ghana use cyanide leaching processes where oxygen acts as an oxidising agent to dissolve gold from ore.
3. Bleaching in Textile Industry: Sodium chlorate(I) oxidises coloured compounds:

$$\text{NaClO} + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{Cl}^- + 2\text{OH}^-$$

Ghana’s textile industry, including companies in the Tema Industrial Area, uses oxidising agents like hydrogen peroxide and sodium chlorate(I) for bleaching fabrics. Traditional kente weavers in the Volta Region also use mild oxidising agents to prepare cotton threads before dyeing them with traditional dyes.

Reducing Agents

1. Metal Extraction: Carbon monoxide reduces iron(III) oxide in blast furnaces:
- $$\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$$
- Traditional blacksmiths in northern Ghana use charcoal (which produces carbon monoxide) as a reducing agent to smelt iron from local ores. This traditional knowledge has been passed down through generations and is still practised in communities like Wechiau and Yikpabongo.
2. Photography
- Developers like hydroquinone reduce silver ions to metallic silver: $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$
- Despite the digital revolution, traditional film photography is still practised in some photography studios in Ghana, particularly for official documents and special occasions. Photography labs in cities like Accra and Kumasi use reducing agents in the development process. Some art departments in Ghanaian universities also teach traditional darkroom techniques.
3. Food preservation
- Sulphur dioxide prevents oxidation in foods: $\text{SO}_2 + 2\text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4 + 2\text{H}^+ + 2\text{e}^-$
- In Ghana’s food processing industry, sulphur dioxide and sulphites are used as preservatives for dried fruits and vegetables. Local processors in Makola Market and other major markets use these reducing agents to maintain the colour and extend the shelf life of dried foods like mango, pineapple, and garden eggs. The cocoa industry also uses reducing agents to prevent oxidation during fermentation and processing.

Learning Tasks

The Ghana Water Company uses various chemicals to treat drinking water in different regions. Below is data from water treatment plants in three locations:

Location	Treatment Chemical	Initial Water Condition	Final Water Condition
Weija	Chlorine gas (Cl ₂)	Brown water with iron(II) content of 3.5 ppm	Clear water with iron(III) precipitate
Kpong	Potassium manganate(VII) (KMnO ₄)	Water with organic contaminants, purple solution after addition	Clear water with brown precipitate
Tamale	Ozone (O ₃)	Water with bacterial count of 1500 CFU/mL	Water with bacterial count <10 CFU/mL

1. For each treatment chemical, identify whether it acts as an oxidising or reducing agent and write the half-equation for its redox reaction in water treatment.

2. Explain why iron(II) ions in water from Weija are converted to iron(III) during chlorination, and write the balanced redox equation for this process.

Pedagogical Exemplars

Initiating Talk-for-learning

1. Begin by having learners individually reflect on the definitions of oxidising and reducing agents using three approaches (oxygen/hydrogen transfer, electron transfer, and oxidation number change) and analyse a sample reaction.

2. Next, pair learners to share their definitions, refine their understanding, and collaboratively analyse the zinc-copper sulphate reaction and potassium permanganate reaction by calculating oxidation numbers and identifying oxidising and reducing agents.

3. Then facilitate a whole-class discussion where pairs share their definitions and analyses, guiding the class toward comprehensive understanding of oxidising and reducing agents through worked examples on the board.

4. Present a new reaction ($\text{H}_2\text{O}_2 + 2\text{FeSO}_4 + \text{H}_2\text{SO}_4 \rightarrow \text{Fe}_2(\text{SO}_4)_3 + 2\text{H}_2\text{O}$) and have learners apply all three definitions to identify the oxidising and reducing agents, emphasising that “the substance that gets reduced is the oxidising agent; the substance that gets oxidised is the reducing agent.”

5. Conclude with a quick assessment using the iron-oxygen reaction, having learners independently identify the oxidising agent (O_2) and reducing agent (Fe) and display their answers.

6. Throughout the lesson, use consistent colour coding, connect concepts to Ghanaian contexts like gold mining and galvanised roofing, address common misconceptions, and differentiate instruction to support all learners’ understanding of these fundamental redox concepts.

Key Assessment Level 4: Individual Classwork (Flow chart design)

A laboratory in Kumasi received five unlabelled bottles containing clear solutions. Initial tests produced the following results:

Solution	Test with KI	Test with KMnO_4	Test with $\text{K}_2\text{Cr}_2\text{O}_7$	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).

1. 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.

2. Design a flowchart showing how you could systematically identify these five solutions using only two chemical tests. Explain your reasoning.

FOCAL AREA 2: BALANCING OF REDOX REACTIONS

Oxidation Half-Reactions

An oxidation half-reaction represents the portion of a redox reaction where a species loses electrons, resulting in an increase in oxidation number.

Characteristics of oxidation half reaction

1. Electrons appear as products on the right side of the equation

2. Oxidation number of the element being oxidised increases

3. Often involves metals, negative ions, or molecules losing electrons

Examples of reduction half reactions

1. $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$.

2. $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$

3. $\text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O(l)} \rightarrow \text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$

Reduction Half-Reactions

A reduction half-reaction represents the portion of a redox reaction where a species gains electrons, resulting in a decrease in oxidation number.

Characteristics of a reduction half reaction

1. Electrons appear as reactants on the left side of the equation

2. Oxidation number of the element being reduced decreases

3. Often involves non-metals, positive ions, or molecules gaining electrons

Examples of reduction half reactions

1. $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$

2. $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O(l)}$

3. $\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O(l)}$

BALANCING REDOX REACTIONS IN ACIDIC MEDIUM

Systematic approach to balancing redox reactions in acidic medium:

1. Write the unbalanced equation

2. Split into half-reactions (oxidation and reduction)

3. Balance elements other than H and O

- Balance oxygen by adding H_2O
- Balance hydrogen by adding H^+
- Balance charges by adding electrons
- Multiply half-reactions by integers to equalise electrons (Electrons lost = Electrons gained)
- Add half-reactions and cancel common terms
- Verify that atoms and charges balance

Example 1: Balance the redox reaction in acidic medium: $\text{MnO}_4^- + \text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+}$

Step-by-step balancing

- Split into half-reactions:
Reduction half : $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$
Oxidation half : $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+}$
- Balance elements other than H and O:
Both Mn and Fe are already balanced
- Balance oxygen in the reduction half-reaction by adding H_2O :
 $\text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$
- Balance hydrogen by adding H^+ :
 $\text{MnO}_4^- + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$
- Balance charges by adding electrons:
 $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$
 $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$
- Multiply each half – equation by an integer to equalise electrons:
 $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$ ($\times 1$) Reduction half
 $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$ ($\times 5$) Oxidation half
- Add half-reactions:
 $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- + 5\text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} + 5\text{Fe}^{3+} + 5\text{e}^-$
- Cancel common terms (electrons):
 $\text{MnO}_4^- + 8\text{H}^+ + 5\text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} + 5\text{Fe}^{3+}$
- Verify that atoms and charges balance:
Left side: Mn(1),O(4),H(8),Fe(5); Charge: +7
Right side: Mn(1),O(4),H(8),Fe(5); Charge: +7
- Balanced equation:
 $\text{MnO}_4^- + 8\text{H}^+ + 5\text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} + 5\text{Fe}^{3+}$

Example 2: Balance the reaction $\text{Cr}_2\text{O}_7^{2-} + \text{C}_2\text{H}_5\text{OH} \rightarrow \text{Cr}^{3+} + \text{CO}_2$ in acidic medium

Solution

Step-by-step balancing

Unbalanced: $\text{Cr}_2\text{O}_7^{2-} + \text{C}_2\text{H}_5\text{OH} \rightarrow \text{Cr}^{3+} + \text{CO}_2$

- Split into half-reactions:
Reduction: $\text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+}$
Oxidation: $\text{C}_2\text{H}_5\text{OH} \rightarrow \text{CO}_2$
- Balance elements other than H and O:
 $\text{Cr}_2\text{O}_7^{2-} \rightarrow 2\text{Cr}^{3+}$
 $\text{C}_2\text{H}_5\text{OH} \rightarrow 2\text{CO}_2$
- Balance oxygen:
 $\text{Cr}_2\text{O}_7^{2-} \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$
 $\text{C}_2\text{H}_5\text{OH} + 3\text{H}_2\text{O} \rightarrow 2\text{CO}_2$
- Balance hydrogen:
 $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$
 $\text{C}_2\text{H}_5\text{OH} + 3\text{H}_2\text{O} \rightarrow 2\text{CO}_2 + 12\text{H}^+$
- Balance charges:
 $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$
 $\text{C}_2\text{H}_5\text{OH} + 3\text{H}_2\text{O} \rightarrow 2\text{CO}_2 + 12\text{H}^+ + 12\text{e}^-$
- Multiply each half – equation by an integer to equalise electrons:
 $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$ ($\times 2$)
 $\text{C}_2\text{H}_5\text{OH} + 3\text{H}_2\text{O} \rightarrow 2\text{CO}_2 + 12\text{H}^+ + 12\text{e}^-$ ($\times 1$)
- Add half-reactions:
 $2\text{Cr}_2\text{O}_7^{2-} + 28\text{H}^+ + 12\text{e}^- + \text{C}_2\text{H}_5\text{OH} + 3\text{H}_2\text{O} \rightarrow 4\text{Cr}^{3+} + 14\text{H}_2\text{O} + 2\text{CO}_2 + 12\text{H}^+ + 12\text{e}^-$
- Cancel common terms and show effective net amounts of each species:
 $2\text{Cr}_2\text{O}_7^{2-} + 16\text{H}^+ + \text{C}_2\text{H}_5\text{OH} \rightarrow 4\text{Cr}^{3+} + 11\text{H}_2\text{O} + 2\text{CO}_2$
- Verify that atoms and charges balance:
Left side: Cr(4),O(15),H(22),C(2); Charge: +12
Right side: Cr(4),O(15),H(22),C(2); Charge: +12
- Balanced equation:
 $2\text{Cr}_2\text{O}_7^{2-} + 16\text{H}^+ + \text{C}_2\text{H}_5\text{OH} \rightarrow 4\text{Cr}^{3+} + 11\text{H}_2\text{O} + 2\text{CO}_2$

BALANCING REDOX REACTIONS IN BASIC MEDIUM

Systematic Approach

Balance the equation as if in acidic medium (follow steps 1-8 from acidic method)

Add OH^- to both sides equal to the number of H^+ ions

Combine H^+ and OH^- to form H_2O on the appropriate side

Cancel excess water molecules from both sides

Verify that atoms and charges balance

Example 1: Balance the equation in basic medium: $\text{MnO}_4^- + \text{SO}_3^{2-} \rightarrow \text{MnO}_2 + \text{SO}_4^{2-}$

Solution

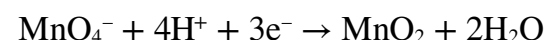
Step-by-step balancing:

Balance as if in acidic medium first:

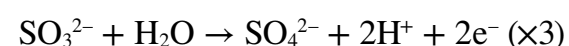
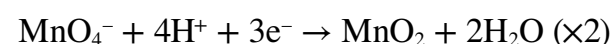
1. Split into half-reactions:



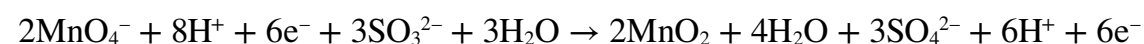
2. Balance elements, oxygen, hydrogen, and charges:



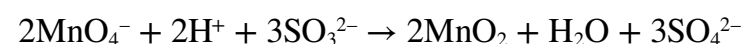
3. Multiply each half – equation by an integer to equalize electrons:



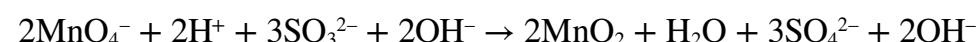
4. Add half-reactions:



5. Cancel common terms and show net amounts:



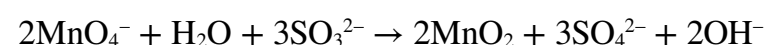
6. Add OH^- to both sides in quantities equal to the net H^+ :



7. Combine H^+ and OH^- to form H_2O :



8. Cancel excess water molecules:

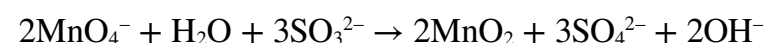


9. Verify that atoms and charges balance:

Left side: Mn(2),O(18),H(2),S(3); Charge: -8

Right side: Mn(2),O(18),H(2),S(3); Charge: -8

10. Balanced equation:



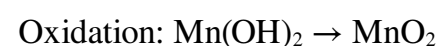
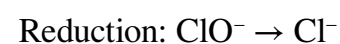
Example 2: Balance the reaction in basic medium: $\text{ClO}^- + \text{Mn}(\text{OH})_2 \rightarrow \text{MnO}_2 + \text{Cl}^-$

Solution

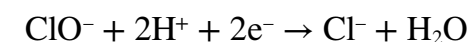
Unbalanced: $\text{ClO}^- + \text{Mn}(\text{OH})_2 \rightarrow \text{MnO}_2 + \text{Cl}^-$

Balance as if in acidic medium first:

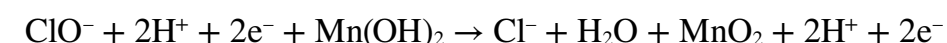
1. Split into half-reactions:



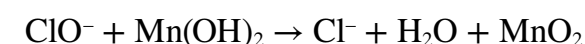
2. Balance elements, oxygen, hydrogen, and charges:



3. These half-reactions have equal electrons, so we can add directly:



4. Cancel common terms:

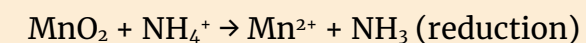
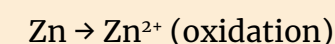


Since there are no H^+ ions in the final equation from step 1, we do not need to add OH^- ions. The equation is already balanced in basic medium.

Learning Task

1. You are explaining how zinc-carbon batteries work to a science club at a senior high school in Accra.

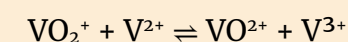
In a zinc-carbon battery, zinc is oxidised and manganese(IV) oxide is reduced in acidic conditions. The unbalanced half-reactions are:



- a. Balance both half-reactions in acidic conditions.
- b. Combine the balanced half-reactions to write the overall reaction in the battery.

2. Ghana is expanding its renewable energy capacity, particularly solar power. A key challenge is energy storage for use during nighttime or cloudy periods. Your research team at the Council for Scientific and Industrial Research (CSIR) is investigating redox flow batteries as a potential solution.

A vanadium redox flow battery utilises the following unbalanced redox reaction:



- a. Identify which species is being oxidised and which is being reduced during:
The charging process
The discharging process
- b. Write and balance the half-reactions for both processes in acidic medium.
- c. In hot climates like Ghana's, battery efficiency can be affected by side reactions. One potential side reaction in vanadium flow batteries is:
 $\text{VO}_2^+ + \text{V}^{2+} + \text{H}_2\text{O} \rightarrow \text{V}_2\text{O}_3 + 2\text{H}^+$
- d. Balance this equation using the half-reaction method and explain how this side reaction would affect battery performance.

Pedagogical Exemplars

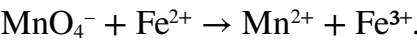
Initiating Talk for learning

1. Show images of rusting iron roofs and gold mining in Ghana; have learners discuss what these processes have in common chemically, ensuring equal participation from boys and girls.
2. Present simple examples like $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ and $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$; demonstrate how to identify which species is oxidised or reduced by tracking oxidation numbers.
3. Think aloud while splitting a redox reaction like $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$ into component half-equations, emphasising that electrons lost must equal electrons gained.
4. Present step-by-step approach for balancing half-equations in acidic and basic media; demonstrate with simple examples, checking for understanding at each step.
5. Organise mixed-gender, mixed-ability groups; provide worksheets with progressively challenging problems; circulate to support all learners.
6. Have groups share solutions using random selection; address misconceptions; create a collaborative summary of balancing steps.
7. Assign contextually relevant redox reactions that appeal to diverse interests; offer differentiated support and extension problems.
8. Have students self-assess their confidence and complete sentence stems about key concepts; discuss connections to future topics and careers.
9. Discuss how redox reactions help solve challenges in Ghanaian communities; highlight applications across traditionally gendered domains.
10. Task learners with identifying redox reactions in their homes or communities and explaining them using half-equations, connecting chemistry to cultural knowledge.

Key Assessment Level 4: Prediction Task

1. Your community’s water sometimes has a reddish-brown colour due to iron compounds. The local water treatment plant uses potassium permanganate (KMnO_4) to solve this problem.

The unbalanced reaction that occurs during treatment is:



Complete the table below for this reaction:

Element	Change in oxidation number	Oxidised or reduced
Mn		
Fe		

- a. Write the balanced half-reaction for the reduction reaction
- b. Write the balanced half-reaction for the oxidation. reaction
- c. Combine these half-reactions to write the balanced overall equation.

- d. After treatment, the water is still slightly coloured. Suggest two possible reasons for this problem and recommend specific solutions based on your understanding of the chemistry involved.

FOCAL AREA 3: REDOX TITRATIONS

Redox titration is an analytical method based on electron transfer reactions. Unlike acid-base titrations which involve proton transfer, redox titrations involve the exchange of electrons between an oxidising agent and a reducing agent. This electron transfer forms the basis for quantitative analysis of many substances important in industrial, environmental, and pharmaceutical applications.

The endpoint is reached when stoichiometrically equivalent amounts of oxidising and reducing agents have reacted. The equivalence point represents the exact point where the number of electrons lost equals the number gained. Redox indicators or self-indicators signal the endpoint through visible colour changes

Types of Redox Titrations

1. Permanganate titrations: utilising the strong oxidising power of MnO_4^-
2. Dichromate titrations: using $\text{Cr}_2\text{O}_7^{2-}$ as the oxidising agent
3. Iodometric/iodimetric titrations: based on the redox chemistry of iodine

Recording and Interpreting Burette Readings

Titration table format:

Burette readings/ cm^3	1	2	3
Final reading	24.50	24.20	24.30
Initial reading	0.00	0.00	0.00
Titre	24.50	24.20	24.30

Concordant or consistent titres

Concordant titres are successive titrations that agree within $\pm 0.2 \text{ cm}^3$. At least two concordant titres should be obtained and the average of the concordant titres is used for calculations.

Example from the above titration, average titre = $\frac{24.20+24.30}{2} \text{ cm}^3$
=24.25 cm^3

Rough Titration

The first titration is often considered a rough titration to approximate the endpoint. Add KMnO_4 quickly until near the expected endpoint, then dropwise.

Use this to guide more precise subsequent titrations. Rough titration results may be excluded from calculations

Standardisation of potassium tetraoxomanganate(VII), KMnO_4 solutions

Potassium tetraoxomanganate(VII) (KMnO_4) cannot be used as a primary standard because:

1. It often contains impurities (manganese(IV) oxide)
2. It decomposes slowly on standing
3. It reacts with organic matter and dust
4. It cannot be dried without decomposition

Therefore, KMnO_4 solutions must be standardised against reliable primary standards.

Preparation of Potassium tetraoxomanganate(VII) Solution (0.02M)

1. Dissolve approximately 3.2g potassium tetraoxomanganate(VII) in 1 dm³ of distilled water
2. Heat gently to about 60°C and stir for one hour.
3. Allow to stand for at least 24 hours to decompose impurities.
4. Filter through a glass funnel (never paper). Alternatively, use glass wool or a glass crucible and store in a clean amber bottle away from light and dust

Standardisation of KMnO_4 solution using $(\text{NH}_4)_2\text{SO}_4 \cdot \text{FeSO}_4 \cdot 6\text{H}_2\text{O}$

Titration equation: $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$

Preparation of standard $(\text{NH}_4)_2\text{SO}_4 \cdot \text{FeSO}_4 \cdot 6\text{H}_2\text{O}$ solution

1. Weigh accurately about 39.2g of $(\text{NH}_4)_2\text{SO}_4 \cdot \text{FeSO}_4 \cdot 6\text{H}_2\text{O}$
2. Dissolve in approximately 500 cm³ distilled water containing 5 cm³ concentrated sulphuric acid.
3. Transfer quantitatively to a 1 dm³ volumetric flask.
4. Dilute to the mark with freshly boiled and cooled distilled water.

Standardisation Procedure

1. Pipette 25 cm³ of the iron(II) solution into a conical flask
2. Add 10-15 cm³ of 1M sulphuric acid
3. Titrate with KMnO_4 solution until a faint permanent pink colour appears
4. Repeat for concordant results and present your results in a tabular form.

STANDARDISATION USING SODIUM ETHANEDIOATE ($\text{Na}_2\text{C}_2\text{O}_4$) or ETHANEDIOIC ACID ($\text{H}_2\text{C}_2\text{O}_4$)

Titration equation: $2\text{MnO}_4^- + 5\text{C}_2\text{O}_4^{2-} + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 10\text{CO}_2 + 8\text{H}_2\text{O}$

1. Weigh accurately about 6.7g $\text{Na}_2\text{C}_2\text{O}_4$
2. Dissolve in about 250 cm³ distilled water
3. Transfer into 1 dm³ volumetric flask and top up with distilled to reach the mark.

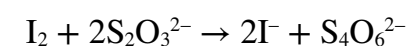
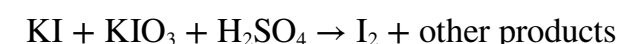
Standardisation Procedure

1. Pipette 25 cm³ of the ethanedioate solution to a conical flask
2. Add 20-30mL of 1M sulphuric acid

3. Heat to 60-70°C (not boiling)
4. Titrate hot solution with potassium tetraoxomanganate(VII) until a faint permanent pink colour appears
5. Add the first 1-2 cm³ slowly (reaction is initially slow)
6. Repeat for concordant results.

STANDARDISATION USING SODIUM THIOSULPHATE (INDIRECT METHOD)

This is a two-step process involving:

**Preparation of standard KIO_3 solution (0.02M)**

1. Dry primary standard grade KIO_3 at 110°C for 1 hour
2. Cool in a desiccator
3. Weigh accurately about 4.28g (MW = 214.0 g/mol)
4. Dissolve in distilled water and dilute to 1L in a volumetric flask

Preparation of Sodium Thiosulphate Solution (0.1M)

1. Dissolve about 24.82g of sodium thiosulphate pentahydrate (MW = 248.2 g/mol) in 1dm³ of freshly boiled and cooled distilled water
2. Add 0.1g of sodium carbonate or a few drops of trichloromethane as preservative and store in an amber bottle.

STANDARDISATION OF $\text{Na}_2\text{S}_2\text{O}_3$ USING KIO_3

1. Pipette 25 cm³ of potassium iodate(V) solution into a conical flask
2. Add about 2g of potassium iodide (solid)
3. Add 10mL of 1M sulphuric acid
4. Titrate the liberated iodine with sodium thiosulphate until pale yellow
5. Add starch indicator and continue until blue colour disappears
6. Repeat the procedure two more times and calculate the average titre.

Learning Tasks**1. Materials Required**

Solution A: 0.05 mol/dm³ potassium tetraoxomanganate(VII) solution

Solution B: Iron(II) solution of unknown concentration

Dilute sulphuric acid (1M)

Burette (50 cm³)

Pipette (25 cm³)

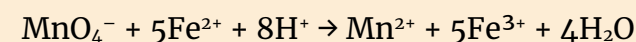
Conical flasks (250 cm³)

White tile

Retort stand and clamp

a. Procedure

1. Set up the titration apparatus and put solution A (potassium tetraoxomanganate(VII)) into the burette.
2. Using a pipette, transfer exactly 25.0 cm³ of solution B into a clean conical flask.
3. Add approximately 10 cm³ of dilute sulphuric acid to the conical flask and swirl gently to mix.
4. Place the conical flask on a white tile beneath the burette.
5. Titrate the mixture against solution A until a permanent pale pink colour appears, indicating the endpoint.
6. Record your burette readings in a titration table.
7. Repeat the titration to obtain at least two concordant results and calculate the average titre. The balanced equation for the redox reaction is:



b. Using your results and the information provided, calculate:

- i. The concentration of solution B in mol/dm³ of Fe²⁺.
- ii. Solution B was prepared by dissolving 71.5 g of ammonium iron(II) sulphate hydrate, (NH₄)₂Fe(SO₄)₂·xH₂O, per dm³ of solution.
Calculate the value of x (the number of water molecules of crystallisation) in the salt.
- iii. A groundwater sample from a community near a mining site contains iron that needs to be quantified. The water appears clear when first collected but develops a reddish-brown colour after standing.

Explain the chemical transformations occurring in the water and how your method accounts for these changes.

[Relative atomic masses: Fe = 55.0, S = 32.0, O = 16.0, H = 1.0, N = 14.0]

2. A mining company in Obuasi has discovered a new potential source of iron ore. Before mining operations begin, they need to determine the iron content of samples from different locations at the site. As the analytical chemist, you are tasked with this analysis using potassium tetraoxomanganate(VII) titration.

A 2.50g sample of the iron ore was dissolved in hydrochloric acid, and all iron was converted to iron(II) ions. The solution was filtered, and the filtrate was diluted to 250.0 cm³ in a volumetric flask. A 25.0 cm³ aliquot of this solution was titrated with 0.02 mol/dm³ potassium tetraoxomanganate(VII) solution, requiring 18.40 cm³ to reach the endpoint.

- a. Write a balance redox reaction for the reaction that took place.
- b. Calculate the percentage by mass of iron in the ore sample.
- c. The company considers an ore economically viable if it contains at least 30% iron by mass. Based on your analysis and knowledge of redox reactions, evaluate whether this ore deposit should be developed.

Pedagogical Exemplars

1. Inquiry- based Learning

a. Laboratory Preparation

First, prepare all laboratory workstations with the necessary equipment (burettes, pipettes, conical flasks, measuring cylinders) and pre-made solutions (potassium tetraoxomanganate(VII), iron(II), sodium ethanedioate, iodine/potassium iodide, sodium thiosulphate). Ensure all solutions are clearly labelled with “Standard” or “Unknown” designations.

b. Group Formation

Next, divide the class into mixed-ability groups of 3-4 learners, ensuring gender balance where possible. Assign complementary roles within each group based on learners’ strengths (equipment manager, data recorder, analyst, reporter).

c. Procedure Distribution

Distribute the prepared titration procedure sheets for each titration type:

- i. potassium tetraoxomanganate(VII) vs iron(II)
- ii. potassium tetraoxomanganate(VII) vs ethanedioate,
- iii. iodine/potassium iodide vs sodium thiosulphate) and have each group select one titration type to investigate first.

d. Safety Briefing

Conduct a safety briefing demonstrating proper handling of solutions, correct burette reading technique, and waste disposal procedures. Highlight specific safety concerns for each titration type, such as the staining properties of potassium tetraoxomanganate(VII) and the need for proper temperature control with ethanedioate titrations.

e. Materials Distribution

After the safety briefing, distribute the appropriate standard and unknown solutions to each group.

f. First Investigation

Then, have groups set up their equipment and conduct their first titration investigation following the provided procedure.

- i. For potassium tetraoxomanganate(VII) vs iron(II) groups, monitor their technique and observation of the pink endpoint colour change.
- ii. For potassium tetraoxomanganate(VII) vs ethanedioate groups, ensure proper heating of solutions to 60-70°C and observation of reaction rate differences with temperature.
- iii. For iodine/potassium iodide vs sodium thiosulphate groups, ensure proper starch indicator use and recognition of the distinctive colour changes.

g. Second Rotation

- Next, rotate groups to a different titration type, distributing the appropriate procedure sheet and solutions for the second titration. Encourage groups to apply insights from their first investigation to improve their technique and data collection.
- h. Third Rotation**
- After completing the second titration, rotate groups to their third and final titration type, distributing the appropriate procedure sheet and solutions. Challenge groups to work with increasing independence and precision.
- i. Assessment Strategy**
- Throughout the process, conduct formative assessment by observing laboratory skills, data recording accuracy, analytical thinking, and group collaboration. For summative assessment, evaluate the quality of data analysis, and error identification.
- 2. Problem solving approach**
- a.** Begin by having learners organise their titration data from the three redox titrations they performed
potassium tetraoxomanganate(VII) vs iron(II),
potassium tetraoxomanganate(VII) vs ethanedioate, and
iodine/potassium iodide vs sodium thiosulphate).
 - b.** Guide learners to review the balanced chemical equations for each titration performed. Focus their attention on understanding the electron transfer processes that occur in each reaction system.
 - c.** Have learners identify the crucial mole ratios from each balanced equation (e.g.,1:5 for $\text{MnO}_4^-:\text{Fe}^{2+}$,2:5 for $\text{MnO}_4^-:\text{C}_2\text{O}_4^{2-}$,and 1:2 for $\text{I}_2:\text{S}_2\text{O}_3^{2-}$). Emphasise that these ratios are essential for accurate concentration calculations.
 - d.** Instruct learners to calculate the concentration of their unknown solutions using the appropriate formula and compare with other groups.
 - e.** Guide learners to check their calculations for mathematical accuracy, correct units, and appropriate significant figures.
 - f.** Challenge learners to identify specific experimental errors that might have affected their results, such as endpoint determination challenges, burette reading imprecision, or temperature variations.
 - g.** Guide learners to interpret their calculated concentrations within relevant contexts, relating to water quality standards, standardisation processes, or analytical applications.
 - h.** Throughout these steps focus on learners ability to correctly apply stoichiometric principles, perform accurate calculations, identify sources of error, and meaningfully interpret their results in context.

Key Assessment: Practice Paper 1

- 1.** Write the half-equations for the reaction between potassium tetraoxomanganate(VII) and sodium ethanedioate in acidic solution.
- Explain how these half-equations help you determine the stoichiometric ratio needed for your calculations.

- 2.** Calculate the concentration of an iron(II) solution if 25.0 cm³ required 20.5 cm³ of 0.020 mol/dm³ KMnO₄ solution for complete reaction.
- Show your working and explain each step.
- 3.** When performing an iodine-thiosulphate titration, explain why starch indicator should be added only when the solution has turned pale yellow rather than at the beginning of the titration.
- 4.** During a redox titration practical session, a learner noticed that their calculated concentration for an iron(II) solution was significantly higher than expected. The learner had followed the correct procedure but had used glassware that had previously contained potassium dichromate solution and had been rinsed only with distilled water. Analyse how this could affect the results, explaining the redox chemistry involved, and suggest appropriate laboratory practices to prevent such errors.



Hint
Conduct *Practice Paper 1*. It should cover contents taught from Year 1 to 3. Refer to the **Table of Specification and Structure of the Paper** at the end of this section. Learners' papers should be marked immediately and scores entered into the SAP.

TABLE OF TEST SPECIFICATION (PAPER 2, 1 & 3)

Strands	Sub- strands	Learning Outcomes	Content Standards	Learning Indicators	Levels of Assess- ment SECTION A				Total	Levels of Assess- ment SECTION B				Total
					L 1	L 2	L 3	L 4		L1	L2	L3	L4	
Physical Chemis- try														
	Matter and its Properties	6	4	30	1		1		2	2	3	2	1	7
	Equilibria	7	7	22	1	1			2	2	4	2		8
Sys- tematic Chem- istry of the Ele- ments	Periodic- ity	4	4	8	1	1			2	2	2	1	1	5
	Bonding	3	3	7		1			1	1	2	1		4

Chem-istry of Carbon Com-pounds	Charac-terization of Organ-ic Com-pounds	2	2	3	1				1	1	1	1	1	3
	Organic Function-al Groups	4	4	11	1	1			2	2	2	1	5	
No of Qs					5	4	1		10	10	14	8	3	35
MARKS					10	12	3		25	19	41	25	15	100
Strands	Sub-strands		Learning Out-comes	Content Stand-ards		Learning Indica-tors		Levels of Assess-ment				To-tal		
Physical Chemistry								1	2	3	4			
	Matter and its Properties		6	4		30		8	7	4	2	21		
	Equilibria		7	7		22		4	3	4		11		
Systematic Chemis-try of the Elements	Periodicity		4	4		8		1	3	1		5		
	Bonding		3	3		7			2	1		3		
Chemistry of Carbon Com-pounds	Characterization of Organic Compounds		2	2		3			2			2		
	Organic Functional Groups		4	4		11		2	3	2	1	8		
								15	20	12	3	50		
Strands	Sub-strands		Learning Out-comes	Content Stand-ards		Learning Indica-tors		Levels of As-sessment				Total		

Physical Chemistry					1	2	3	4	
	Matter and its Properties	6	4	30					
	Equilibria	7	7	22					
Systematic Chemistry of the Elements	Periodicity	4	4	8					
	Bonding	3	3	7	1	2			
Chemistry of Carbon Compounds	Characterization of Organic Compounds	2	2	3					
	Organic Functional Groups	4	4	11					
					1	2			50



Note

Please use the sample instruction template for PLC Sessions 1 for PLC 10 and 11 but choose a different assessment strategies for each of these weeks.

PLC Session 12

45 minutes	Planning & Reviewing the Learning Plan
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1. Begin the session by recapping lessons from the previous session, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week’s concepts in a practical way that will make it easier for learners to understand.
 - b. identify three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q).
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Use your app to plan and administer the mode of assessment for the Student Assessment Portal (SAP).
- d. The key assessment for the week is **practice paper 1**. Chat with your app to plan this assessment by reviewing or developing the assessment items/tasks and marking schemes/rubrics for the practice paper (NTS 3k-3q).

10 minutes	Content Exploration
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5. As a subject group, select one problem or topic/idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
- a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through the problem individually for 3–4 minutes
 - c. the group then works through it together the problem, step by step
 - d. highlight the most common errors made at each step (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
- a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
- a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

WEEK 13 - 14

Learning Indicator: Explain how redox is used to produce electricity in storage and fuel cells

Recap weeks

Year 1	Week 8: Use IUPAC nomenclature to name inorganic compounds, write the formulae of compounds based on the laws of chemical combination and write balanced chemical equations.
Year 2	Week 8: State Le Chatelier's Principle and apply it to deduce qualitatively, the effects of various factors on the positions of equilibrium and on the value of the equilibrium constant

FOCAL AREA 1: ELECTROCHEMICAL CELLS

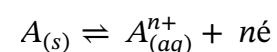
Electrochemistry is the branch of chemistry that studies the relationship between electrical and chemical processes, particularly how chemical reactions can generate electricity and how electrical energy can drive chemical changes. It deals with interconversion of electrical and chemical energy. Many chemical changes can be related to the electrons that move from one species to another. Often, this electron exchange can be captured to do electrical work external to the chemical system (storage battery, fuel cell). Other times, electrical energy can be used to bring about chemical change (electrolysis, battery charging, etc.).

1. Introduction to electrochemical cells

An electrochemical cell is a device used to generate electrical energy by means of chemical reactions. Electricity is generated when electrons move from one entity to another. This transfer of electrons can be achieved when atoms in their natural form losses electrons to form ions and the electrons are transferred to another species. Elements which are most likely to lose electrons are metals. When a metal plate is dipped into a solution which contains its ions, the metal may interact with the aqueous solution of its ions in three basic ways

- The metal atom may lose electron(s) to form its ion(s) which will then leave metal lattice and dissolve in solution
- The metal ions which are present in the solution may gain electron(s) to form a metal atom which crystallises out of the solution
- There will be no net change of electrons between the metals and its ions in solution.

This generates an equilibrium between the metal and its ions in solution. The equilibrium reaction can be represented as



The equilibrium position will vary depending on the ease with which the metal plate loses electrons. A metal like sodium, which loses electrons easily, will develop a negative charge, whereas a metal like silver, which does not lose electrons easily, will acquire a less negative charge. When two metal plates which are dipped in a solution of their ions are connected using a conducting wire and the solutions are also connected by a porous membrane to ensure that ions flow between them, electrons will flow from the metal plate which is more reactive to the

metal plate which is less reactive through the wire. The flow of electrons generates an electric current that can be harnessed to perform useful work, with the magnitude of the potential difference (voltage) being equal to the difference between the two electrode potentials under the given conditions. The device which generates electric current by this means is referred to as the electrochemical cell.

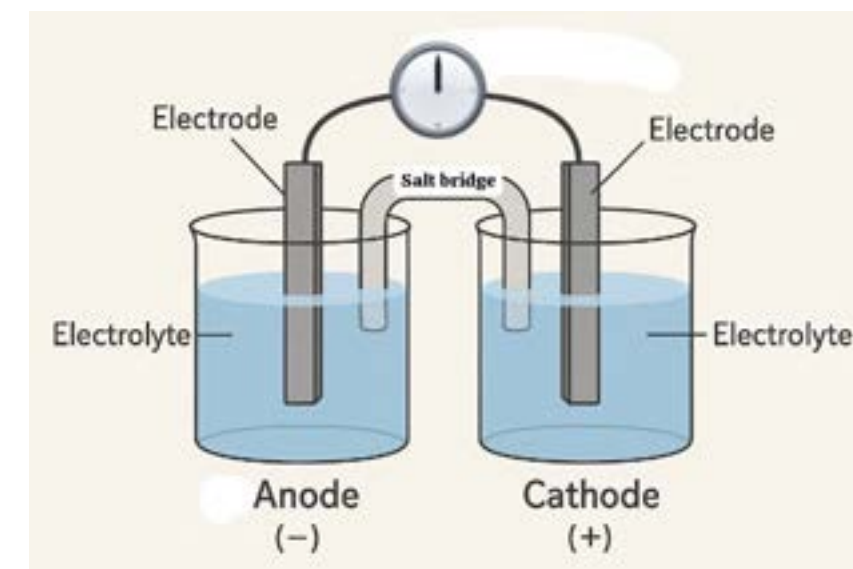


Figure 13.1: Diagram of an electrochemical cell (Voltaic cell) showing its components

Composition of an electrochemical cell

Electrodes

Electrodes are conductors where electron transfer occurs. There are two types of electrodes; the anode and the cathode.

Anode

It is conventionally placed on the left hand side in diagrams of electrochemical cells. It is the site where oxidation occurs (loss of electrons). It is designated as the negative electrode. Electrons flow from the anode to the external circuit. Metal atoms (A) lose electrons to become cations (A^{n+}) which accumulates in solution. To balance the high concentration of cations at the anode, anions are attracted from the porous membrane (salt bridge) into the solution.

Equation for reaction at the Anode: $A_{(s)} \rightarrow A^{n+} + ne^{-}$

Cathode

It is conventionally positioned on the right-hand side of a galvanic cell diagram, creating a visual balance with the anode. The cathode serves as the crucial site where reduction occurs (the gaining of electrons), making it the electrochemical destination for the electron journey. In galvanic cells, it is designated as the positive electrode, attracting electrons like a magnetic beacon through the external circuit.

As electrons arrive at the cathode from the external circuit, they create a negative charge at the electrode surface. This negative charge causes an attractive force of attraction for positively charged cations in the surrounding electrolyte solution. When these cations reach the cathode surface, they accept the electrons reduction reaction. For metal cations (M^{n+}), this electron transfer transforms them from dissolved ions into solid metal atoms that deposit onto the

cathode surface, sometimes creating visible crystalline growth patterns as the reaction progresses.

Equation for reaction at the Cathode: $M^{n+} + ne^- \rightarrow M_{(s)}$

The continuous reduction process at the cathode gradually depletes the concentration of cations in the solution. To maintain electrical neutrality in this half-cell as cations are removed from solution, positively charged ions flow through the porous membrane or salt bridge from the anode half-cell into the cathode solution. Simultaneously, anions from the cathode half-cell migrate in the opposite direction—toward the anode—completing the internal ionic circuit that balances charges throughout the system.

Electrolytes

The electrodes are dipped into two separate solutions which are referred to as the electrolytes. Electrolytes are ionic solutions that conduct electricity through ion movement. They are usually made up of a solution of salts, acids or bases that dissociate into ions. Electrolytes complete the internal circuit by allowing ions to flow and provide ions for electrode reactions which maintains charge balance. In a galvanic cell, the cations in the solution of the electrolyte at the cathode, picks up the awaiting electrons and transforms into metallic atoms (reduction). This reduces the concentration of the cations which is compensated for by the cations that move through the porous membrane that connects the two electrolytes. Meanwhile at the anode, metal atoms from the electrode lose electrons and enter the electrolyte solution as cations. The electrolyte is where these newly formed cations dissolves during the oxidation process. This chips away the metallic components of the electrode which increases the concentration of the cations which are then balanced by the migration of the anions through the salt bridge.

Salt bridge

The salt bridge functions as the essential ionic connector between the two half-cells of a galvanic cell, maintaining electrochemical balance while preventing direct solution mixing. Typically appearing as a U-shaped glass tube or filter paper strip, it contains a gel of agar or gelatin infused with inert salts like potassium chloride or potassium nitrate.

This semi-permeable structure serves as a selective ionic conductor, allowing ions to pass while preventing bulk solution flow that would contaminate the distinct chemical environments of each half-cell. As the cell operates, critical imbalances develop that the salt bridge must address: the cathode accumulates negative charge as cations are consumed during reduction, while the anode develops positive charge as metal atoms oxidise to form cations.

The salt bridge responds with precisely coordinated ion movement. Anions migrate from the bridge toward the anode half-cell, neutralising excess positive charge from accumulating metal cations. Simultaneously, cations travel from the bridge toward the cathode half-cell, balancing the excess negative charge left as cations are depleted through reduction. This bidirectional movement maintains electrical neutrality in both half-cells, preventing the charge build up that would otherwise halt the cell's operation.

Without this 'ionic highway', the galvanic cell would quickly cease functioning as opposing electric fields would develop, stopping electron flow through the external circuit. The salt bridge thus allows oxidation and reduction to proceed unhindered at their respective electrodes.

Electrode potential

The concentration of electrons on the electrodes is measured by a quantity called the electrode potential. This can be read by a high resistance voltmeter or a potentiometer. Electrode potential

measures an electrode's tendency to gain or lose electrons when contacting its ions in solution. It results from the dynamic equilibrium at the electrode-electrolyte interface where continuous exchange occurs between metal atoms and ions. When a metal contacts its ion solution, either oxidation (metal atoms losing electrons to become ions) or reduction (metal ions gaining electrons to become atoms) predominates, determining the electrode's potential. The difference between the electrode potential of the two electrodes in a galvanic cell is called the electromotive force (e.m.f.) of the cell. It is measured in volts. It is impossible to measure the absolute potential of a single electrode in isolation. Potential difference can only be measured between two electrodes. In order to compare the electrode potentials of elements, the Standard Hydrogen Electrode (SHE) is used as the universal reference point. It is assigned an electrode potential of zero from which all others are measured. It should be noted that both positive and negative electrode potentials are possible and that when calculating the potential difference this should be considered (and watch out for).

e.g Sodium has an electrode potential of -2.71 V, whilst Gold has a value of +1.5 V. Therefore the potential difference would be $1.5 - (-2.71) = 4.21$ V. The cathode is always listed first as this is the direction of electron flow to. This means the potential difference for a cell which will work spontaneously is always positive.

2. Factors that affect electrode potential values

The tendency of an electrode to gain or lose electrons is influenced by several key factors. These factors can significantly alter its value in practical application

a. Nature of the electrode

Different metals have inherently different tendencies to lose electrons, which is influenced by their atomic structure and physical properties. Reactive metals like potassium (K) and sodium (Na) have strongly negative (compared to the SHE) electrode potentials because they easily lose electrons to form positive ions (K^+ , Na^+). This makes them good reducing agents. On the other hand, noble metals like gold (Au) and platinum (Pt) have positive (compared to the SHE) potentials because they do not lose electrons easily. They are more likely to gain electrons or remain in their elemental state, making them good oxidising agents and very stable in aqueous environments. This is determined by the electron configuration, atomic radius, and ionisation energy. This is very key in predicting the direction of electron flow in an electrochemical cell and in the use of the electrochemical series.

b. Concentration of Electrolyte

The Nernst equation quantifies how electrode potential varies with ion concentration. It provides a mathematical relationship between the electrode potential and the ion concentration:

$$E = E^\circ - \frac{RT}{nF} \ln \left(\frac{[\text{Reduced form}]}{[\text{Oxidised form}]} \right)$$

Where:

E = Actual electrode potential under non-standard conditions (volts)

E° = Standard electrode potential (volts)

R = Universal gas constant (8.314 J/mol·K)

T = Temperature in Kelvin

n = Number of electrons transferred in the half-reaction

F = Faraday constant (96,500 C/mol)

Q = Reaction quotient (ratio of concentrations of products to reactants, each raised to power of stoichiometric coefficient)

Higher concentrations of metal ions in solution make reduction more favourable, shifting the potential more positive. Conversely, dilute solutions result in more negative potentials. This relationship allows concentration cells to generate voltage from concentration differences alone.

c. Temperature of the electrolyte

Higher temperatures affect the equilibrium position and thus the electrode potential. For endothermic reactions, potential increases with temperature; for exothermic reactions, potential decreases. This is why batteries often perform poorly in cold conditions. Temperature increases will also increase the reaction kinetics of the battery.

d. Pressure of gaseous elements

For half-reactions involving gases, changes in pressure alter the position of equilibrium according to Le Chatelier's principle. Increased gas pressure makes reduction more favourable, shifting the potential more positive. According to Le Chatelier's principle higher pressure drives the equilibrium toward the side with fewer gas molecules.

3. Standard hydrogen Electrode (SHE)

It is a half-cell in which hydrogen gas, at a temperature of 298K and a pressure of 1 atm is bubbled around an inert platinum electrode into a 1 mol/dm³ [H⁺] solution. It serves as a reference point for the measurements of the standard electrode potential of other elements. It is assigned electrode potential of zero volts.

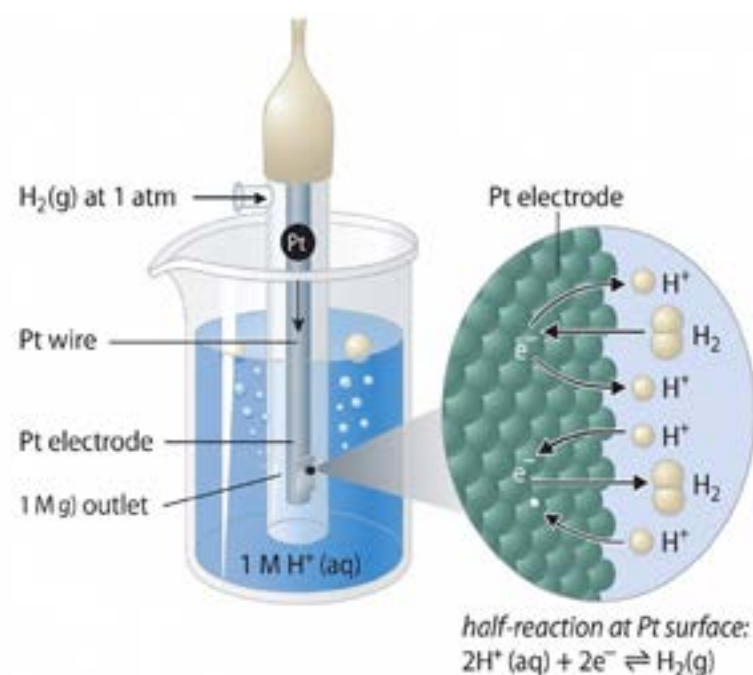


Figure 13.2: A standard hydrogen electrode.

4. Standard electrode potential

The standard electrode potential of an element is a fundamental electrochemical measurement that quantifies the tendency of a chemical species to undergo reduction under standard conditions. It represents the potential difference developed between an electrode and its ions in solution when all components are in their standard states. The standard electrode potential is measured relative to the standard hydrogen electrode. It is expressed as a reduction half equation such as $A_{(aq)}^{n+} + ne^- \rightleftharpoons A_{(s)}$ $E^\circ = \pm x \text{ V}$

The standard electrode potentials roughly range between $\pm 3.0 \text{ V}$. The more negative the value the value the more likely an oxidation reaction will occur. On the other hand, the more positive the value, the more likely a reduction reaction will occur.

Uses of Standard Electrode Potential (SEP)

- Used to predict the emf of an electrochemical cell
- It is used to predict the spontaneity of a redox reaction
- Used to compare the reducing ability of elements

Learning Tasks

- Review oxidation and reduction as electron transfer processes and write balanced equations to illustrate each.
- In pairs or groups, identify and compare components of an electrochemical cell by observing diagrams or videos and creating a comparison chart for both electrodes.
- Construct a simple voltaic cell using zinc and copper materials; measure voltage, label components, and write half and overall cell reactions.
- Explain the development of electrode potential and describe the standard hydrogen electrode with a labelled diagram.
- Analyse factors affecting electrode potentials and apply standard electrode potential values to predict changes in cell behaviour and feasible electrochemical pairs.
- Differentiate between primary and secondary cells and describe their practical uses in everyday applications.

Pedagogical Exemplars

Collaborative learning

- In gender-based groups where applicable (or mixed groups if preferred), guide learners to describe how a potential is developed at the interface of an electrode and the aqueous solution of its ions
- Prompt groups to consider what happens when a metal is placed in a solution of its ions.
- Have learners research and discuss the factors that affect its value:
 - nature of electrode
 - temperature of electrolyte

- c. concentration of electrolyte
 - d. pressure of gas species
4. Guide learners to create a comprehensive chart showing all factors and their effects on electrode potential.
5. Address any misconceptions and highlight key points

Talk-for-learning

1. Explain why it is impossible to obtain an absolute electrode potential for a single half – cell.
2. Show a diagram of the standard hydrogen electrode and assist learners to describe the standard hydrogen electrode in their own words.
3. Show a diagram of a simple electrochemical cell and guide learners to identify the various component and standard conditions of the cell. Sum up the reactions in the electrochemical cell using a cell diagram.
4. Explain that electrode potential is generated between two half-cells. Describe how to measure standard electrode potentials relative to the standard hydrogen electrode as a reference point.
5. Describe how standard electrode potentials are used to construct and represent voltaic cells by writing and drawing of:
- i. Metals or non-metals in contact with their ions in aqueous solution
 - ii. Ions of the same element in different oxidation states
6. Model how to write cell notation for a simple voltaic cell and demonstrate the connection between the notation and the physical setup.
7. Provide learners with examples of both metal/metal ion cells and variable oxidation state cells
8. Engage learners to practice writing cell notations for different combinations

Digital learning

1. Use an appropriate video or simulation to illustrate and explain the operation of a voltaic cell https://youtu.be/qpFC_Ecu_yQ?si=6ZHRyCsb2yK6W7PE
2. Explain what students should focus on during the video/simulation and distribute viewing guide with specific questions to answer
3. Pause at strategic points to check understanding and highlight key concepts
4. Engage learners by asking targeted questions and allow them to make predictions before continuing.
5. Connect the digital visualisation to previously learned concepts

Key Assessment (DoK LEVEL 2)

1. Use the following standard electrode potential values to answer the questions that follows

Half-reactions	Standard Electrode Potential (E°/V)
$Q^+_{(aq)} + e^- \rightarrow Q$	-2.72
$X^{2+}_{(aq)} + 2e^- \rightarrow X$	+1.10
$Y^{2+}_{(aq)} + 2e^- \rightarrow Y$	-2.10
$Z^+_{(aq)} + e^- \rightarrow Z$	+0.40

- a. Arrange the species in order of decreasing chemical reactivity and give reasons for your answer.
- b. A cell is constructed using $Y^{2+} \rightarrow Y$ and $Z^+ \rightarrow Z$ half-cells,
 - i. Draw an electrochemical cell to represent this reaction
 - ii. Write the overall cell reaction
 - iii. Calculate the emf of the cel

[15 marks]

2. Consider the following standard electrode potentials



- a. Calculate the standard cell potential for a cell constructed using $Fe^{3+} \rightarrow Fe^{2+}$ and $Cr^{2+} \rightarrow Cr$ half-cells. Write the cell notation and determine whether the reaction is feasible under standard conditions. (4 marks)
- b. For the reaction: $2Fe^{2+} + Cl_2 \rightarrow 2Fe^{3+} + 2Cl^-$
 - i. Calculate the standard cell potential
 - ii. Calculate the Gibbs free energy change (ΔG°) at 25°C
 - iii. Determine the equilibrium constant for this reaction

[F = 96,500 C/mol]

[10 marks]

3. A student is investigating different metal combinations for constructing a voltaic cell. They have access to zinc, copper, iron, and lead metals along with their respective sulphate solutions.
- Explain the essential components of a voltaic cell and how it produces electricity through redox reactions. (4 marks)
 - Given the following standard electrode potentials:





- Determine which combination of metals would produce the highest cell potential.
- Write the half-equations and overall cell reaction for this optimal combination. (5 marks)
- Analyse how each of the following factors would affect the electrode potential values in your selected cell, and explain the electrochemical principles involved:
 - Increasing the temperature
 - Changing the surface area of the electrodes
 - Using a porous barrier instead of a salt bridge (6 marks)

FOCAL AREA 2: CALCULATIONS OF STANDARD ELECTRODE POTENTIALS

1. Calculation of standard cell potential

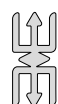
The standard electrode potential of a galvanic/voltaic cell is the difference between the standard reduction potentials of the cathode and anode under standard conditions (298 K, 1 atm, and 1 mol/dm³ concentrations).

Formula:

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

Where:

- E°_{cell} = standard cell potential
- E°_{cathode} = standard reduction potential of the cathode (reduction half-reaction)
- E°_{anode} = standard reduction potential of the anode (oxidation half-reaction)



Note:

The standard electrode potentials are always quoted as reductions. The more positive E° value is assigned to the cathode, while the more negative value is assigned to the anode.

Example:

Calculate the standard emf for the cell: $\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$

$$E^\circ(\text{Zn}^{2+} \rightarrow \text{Zn}) = -0.76 \text{ V}, E^\circ(\text{Cu}^{2+} \rightarrow \text{Cu}) = +0.34 \text{ V}$$

$$E^\circ_{\text{cell}} = 0.34 - (-0.76) = +1.10 \text{ V}$$

2. Predicting the feasibility of a reaction using standard cell potential and Gibbs free energy

A redox reaction is feasible (spontaneous) if the calculated standard cell potential (E°_{cell}) is positive. This can be quantitatively confirmed using the relationship between Gibbs free energy (ΔG°) and E°_{cell} :

$$\Delta G^\circ = -nFE^\circ_{\text{cell}}$$

Where:

- ΔG° = change in Gibbs free energy (J)

- n = number of moles of electrons exchanged

- F = Faraday's constant (96,500 C/mol)

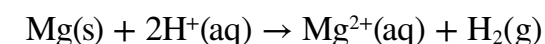
- E°_{cell} = standard cell potential (V)

If $\Delta G^\circ < 0$, the reaction is spontaneous.

If $\Delta G^\circ > 0$, the reaction is non-spontaneous.

Example:

Determine the feasibility of the redox reaction:



$$\text{Given: } E^\circ(\text{Mg}^{2+} \rightarrow \text{Mg}) = -2.37 \text{ V}, E^\circ(\text{H}^+ \rightarrow \text{H}_2) = 0.00 \text{ V}$$

$$E^\circ_{\text{cell}} = 0.00 - (-2.37) = +2.37 \text{ V}$$

$$\Delta G^\circ = -(2)(96,500)(2.37) = -457,530 \text{ J} = -457.53 \text{ kJ (Spontaneous)}$$

3. Experiment to determine the emf of a voltaic cell

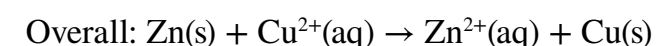
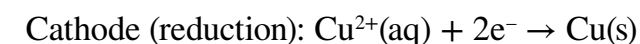
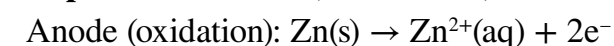
Materials:

- Zinc and copper strips
- Zinc sulphate and copper sulphate solutions (1 mol/dm³)
- Salt bridge (e.g., KNO₃ or NH₄NO₃ in agar gel)
- Voltmeter
- Wires and clips
- Two beakers

Procedure

- Prepare 1 mol/dm³ ZnSO₄ and 1 mol/dm³ CuSO₄ solutions in separate beakers.
- Immerse Zn in ZnSO₄ and Cu in CuSO₄.
- Connect the two electrodes with a wire and voltmeter.
- Connect the beakers using a salt bridge.
- Observe and record the voltage.

Expected Reaction (Daniel Cell):



$$\text{Expected emf} = 1.10 \text{ V}$$

4. Applications of principles of electrode potentials.

Standard electrode potentials have many practical applications in chemistry, industry, and daily life. Some of these include:

- a. Predicting Redox Reactions: They are used to determine the direction and feasibility of redox reactions.
- b. Designing Batteries: It helps in selecting suitable electrode materials to construct efficient voltaic cells (e.g., dry cells, lead-acid batteries, lithium-ion cells).
- c. Electroplating and Metal Purification: They are used in electroplating (e.g., silver-plating, gold-plating) and purification of metals (e.g., copper).
- d. Corrosion Prevention: They are used to select sacrificial anodes for cathodic protection of some metals.
- e. Biochemical Applications: In bio electrochemical systems like enzyme-based biosensors and energy generation in microbial fuel cells. This principle is used in glucose meters for diabetes management, where the emf signal correlates with glucose levels in blood.
- f. Analytical Chemistry: Redox titrations and potentiometric methods use standard electrode potentials to determine endpoint detection and analyte concentrations.

Learning Tasks

1. Write standard reduction half-reactions and use E° values to identify species being oxidised and reduced.
2. Calculate standard cell potential (E°_{cell}) and determine spontaneity using formulas and decision flowcharts.
3. Link E°_{cell} to Gibbs free energy (ΔG°) by calculating values and interpreting spontaneity through a relationship table.
4. Design and build voltaic cells, predicting electron flow and potential based on chosen materials.
5. Compare different battery types by analysing their structure, materials, reactions, uses, and limitations.
6. Evaluate hydrogen-powered vehicles against alternative energy systems, highlighting advantages and challenges.

Pedagogical Exemplars

Problem solving approach

1. Review the concept of standard electrode potentials and then display the electrochemical series showing standard reduction potentials
2. Clarify that all potentials are for reduction half-reactions under standard conditions
3. Demonstrate the systematic process for calculating standard cell potential:
 - a. Identify the two half-reactions involved
 - b. Write both as reduction half-reactions with their corresponding E° values
 - c. Determine which species will be oxidised (more negative E°) and which will be reduced (more positive E°)
 - d. Reverse the oxidation half-reaction (and its sign)

- e. Apply the formula: $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$

4. Explain the significance of the calculated value
5. Guide learners to practice problems using scaffolding worksheet
6. Introduce the relationship between E°_{cell} and reaction spontaneity:
 - a. Positive E°_{cell} means the reaction is spontaneous
 - b. Negative E°_{cell} means the reaction is non-spontaneous
7. Introduce the equation relating cell potential to Gibbs free energy: $\Delta G^\circ = -nFE^\circ_{\text{cell}}$ and explain each term
 - a. ΔG° = Standard Gibbs free energy change (J/mol)
 - b. n = Number of moles of electrons transferred in the balanced equation
 - c. F = Faraday constant (96,500 C/mol)
 - d. E°_{cell} = Standard cell potential (V)
8. Explain the significance of the negative ΔG° (spontaneous process)
9. Provide learners with a set of redox reactions and guide them to:
 - a. Calculate the standard cell potential
 - b. Determine the number of moles of electrons transferred
 - c. Calculate the Gibbs free energy change
 - d. Predict whether the reaction is feasible under standard conditions

Note: Limitations of predictions made using standard cell potentials in terms of kinetic energy and concentrations are not required.

Exploratory Learning

1. Review safety considerations when working with electrolytes and metals
2. Assist learners in groups to set up voltaic cells using available resources or simulations
3. Measure emf values for different cell combinations
4. Guide learners to record all observations and measurements
5. Compare experimental values with theoretical calculations
6. Analyse sources of error and discrepancies
7. Transition to discussing commercial applications of electrochemical principles
8. Provide an overview of different types of modern batteries and their key characteristics
9. Apply the principles of electrode potentials to modern storage cells:
 - a. Lead-acid battery (wet cell)
 - b. Alkaline cell (dry)
 - c. Lithium-ion battery
 - d. Nickel-cadmium battery

- e. Fuel cells
- 10. Provide a structured worksheet for learners to practice writing:
 - b. Anode half-reactions for each battery type
 - c. Cathode half-reactions for each battery type
 - d. Overall balanced reactions

Digital Learning

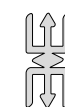
1. Show selected videos on storage cell technologies and pause at strategic points to check understanding
2. Have learners observe videos or diagrams on storage and fuel cells and how scientists in car industry are developing fuel cell vehicles fuelled by
 - a. Hydrogen gas
 - b. Hydrogen-rich fuels and their advantages over conventional diesel powered vehicles (STEM)

<https://youtu.be/wWOVzSmc72E?si=js6PBhbng0eYz6hN>
<https://youtu.be/Rwy6WozMbt8?si=y3f7y-sFzwO3bKOF>
<https://youtu.be/SRt1KOP9a2g?si=6yYaa2wRIbumpaXP>
3. Guide learners to organise information from the observations to create a concept map showing
 - a. Fuel cell components
 - b. Electrochemical reactions
 - c. Energy transformations
 - d. Advantages over conventional batteries

Key Assessment Level 4: Individual Class Exercise & Practical

1. Ghana's gold mining industry faces challenges with environmental contamination from cyanide and mercury used in extraction processes. A team of researchers at the University of Ghana is investigating electrochemical methods to recover gold from mining waste while reducing environmental impact.
The team has access to the following metals and solutions:
Zinc metal and Zn^{2+} solution ($E^\circ = -0.76\text{V}$)
Copper metal and Cu^{2+} solution ($E^\circ = +0.34\text{V}$)
Iron metal and Fe^{2+} solution ($E^\circ = -0.44\text{V}$)
Gold ions in solution ($\text{Au}^{3+} \rightarrow \text{Au}$, $E^\circ = +1.50\text{V}$)
Carbon electrodes
2. Design an electrochemical system to recover gold from mining waste solutions using electrodeposition. Your design must:
 - a. Use available materials to create an effective electrochemical cell
 - b. Calculate the standard cell potential for your proposed system
 - c. Determine the Gibbs free energy change to prove process feasibility

- d. Explain why your design would be more environmentally sustainable than current methods
- 3. Develop a complete electrochemical analysis of your recovery system including:
 - a. Half-reactions at each electrode with correct balancing
 - b. Overall cell reaction with states of matter
 - c. Cell diagram with proper notation
 - d. Energy efficiency calculations based on theoretical and practical considerations
 - e. Predictions for gold recovery rates with supporting calculations



Note

Please use the same instruction template for PLC Sessions 13 and 14, but choose a different assessment strategy for the second week.

PLC Session 13 & 14

45 minutes

Planning & Reviewing the Learning Plan

1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week's concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week's content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week's concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).
4. Plan the week's key assessment with your app
 - a. The key assessment for the week is **individual class exercise & practical**. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the assessments (NTS 3k-3q).



Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

10 minutes	Content Exploration
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5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
- a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step
 - d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
- a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
- a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

WEEK 15 – 16

Learning Indicator: Explain qualitatively and quantitatively the operations of electrolytic cell as well as their applications in everyday life

Recap weeks

Year 1	Week 6: Describe the mole as a unit of the amount of substance.
Year 2	Week 9: Explain Arrhenius, Brønsted-Lowry and Lewis concepts of acids and bases Week 11: Distinguish between strong and weak acids, strong and weak bases in terms of the extent of dissociation, rate of reaction, electrical conductivity, pH value in aqueous solution and enthalpy change of neutralisation value. Week 12: Explain the meaning of salt and classify salts

FOCAL AREA 1: ELECTROLYSIS

Introduction to Electrolysis

Electrolysis is a process that uses electrical energy to drive a non-spontaneous chemical reaction, specifically the decomposition of a chemical compound.

During electrolysis, electrical current passes through an electrolyte (an ionic substance that conducts electricity when molten or in solution), causing chemical changes at the electrodes.

Unlike spontaneous redox reactions that occur in galvanic cells (batteries), electrolysis requires an input of energy. This makes it fundamentally a process that converts electrical energy into chemical energy, storing that energy in the form of chemical bonds.

The term “electrolysis” comes from the Greek words “elektron” (electricity) and “lysis” (breaking down or decomposition), accurately describing the process of using electricity to break chemical bonds.

Components of an Electrolytic Cell

An electrolytic cell consists of several essential components, each serving a specific function in the electrolysis process:

1. **Electrodes:** The electrodes are the solid conductors through which electric current enters and leaves the electrolyte. They serve as the sites for the oxidation and reduction reactions.

Anode (Positive Electrode): Connected to the positive terminal of the power supply and is a site of oxidation reactions (loss of electrons). Anions are attracted to the anode. In some cases, the anode itself may participate in the reaction (active anode). Common anode materials include graphite, platinum, titanium coated with precious metals, and various metals depending on the application

Cathode (Negative Electrode): Connected to the negative terminal of the power supply and a site of reduction reactions (gain of electrons). Cations are attracted to the cathode and may collect the reduced products (e.g., metal deposits). Common cathode materials include various metals, graphite, and mercury (historically)

2. **Electrolyte:** The electrolyte is the medium that conducts electricity through the movement of ions. It must contain or form ions to function in an electrolytic cell.

Examples of electrolytes include: molten ionic compounds (e.g., molten NaCl at $>801^{\circ}\text{C}$), aqueous solutions of ionic compounds (e.g., NaCl dissolved in water), ionic liquids (room temperature molten salts), and solutions of acids or bases (which provide H^+ or OH^- ions).

3. **Power source:** The power source provides the electrical energy necessary to drive the non-spontaneous electrolysis reactions.

Direct current (DC) is required (alternating current would reverse the reactions) and the voltage must exceed the decomposition potential of the electrolyte. Current determines the rate of reaction according to Faraday's laws. Industrial electrolysis often uses very high currents (thousands of amperes).

Cell Container: The container holds the electrolyte and supports the electrodes in their proper positions.

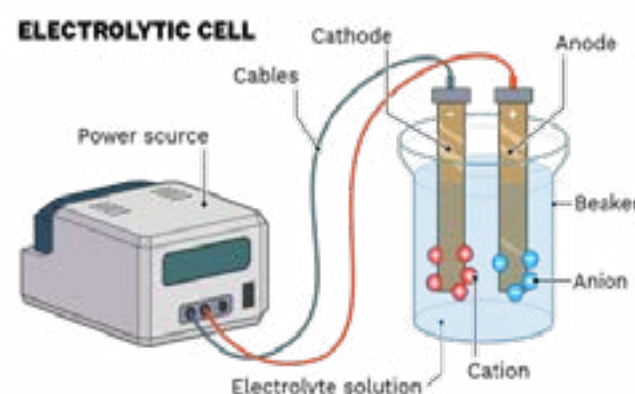


Figure 15.1: A basic electrolytic cell showing key components

Factors that affect selective discharge of species during electrolysis

During electrolysis, different ions in the electrolyte compete to be discharged at the electrodes. This is so when the electrolyte contains more than one anion and or cation. Understanding which species will be preferentially discharged is crucial for predicting the products of electrolysis. Several key factors determine this selective discharge process.

1. Position in the electrochemical series

The electrochemical series ranks ions according to their standard electrode potentials, indicating their tendency to be reduced or oxidised.

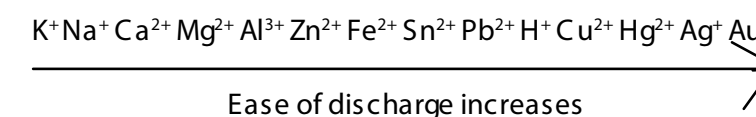
At the Cathode (Negative Electrode), species with more positive (or less negative) reduction potentials are preferentially reduced. Cations lower in the electrochemical series (less reactive metals) are discharged more readily.

At the Anode (Positive Electrode), species with more negative (or less positive) oxidation potentials are preferentially oxidised, anions lower in the electrochemical series are discharged more readily. E.g. In a solution containing Cu^{2+} and Zn^{2+} ions with inert electrodes:

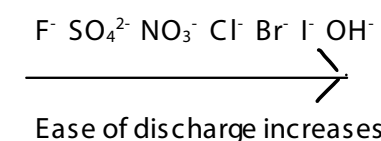


Copper ions will be preferentially discharged at the cathode because their reduction potential is more positive than that of zinc ions.

Electrochemical Series for Cations

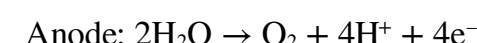
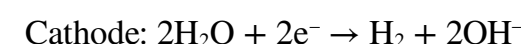


Electrochemical Series for Anions

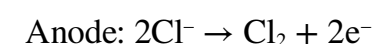
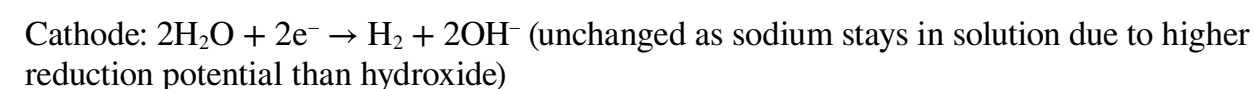


2. **Concentration of ions:** The concentration of ions significantly affects which species will be discharged, particularly in aqueous solutions where water itself can be electrolysed. Higher ion concentrations increase the likelihood of their discharge. In very dilute solutions, water electrolysis may dominate even if standard potentials favour the dissolved ions, e.g.

Dilute NaCl solution:



Concentrated NaCl solution:



This explains why chlorine gas is produced during the electrolysis of concentrated sodium chloride solution but not in dilute solutions.

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.

3. **Nature of Electrodes:** The material composition of electrodes significantly influences which species will be discharged through several mechanisms. Inert electrodes (e.g., platinum, graphite) do not participate in the reaction. Active electrodes (e.g., copper, zinc) may themselves be oxidised or reduced

Products of Electrolysis

The products formed during electrolysis depend on several factors including the nature of the electrolyte, electrode materials, concentration, etc. Let us explore the various products formed during electrolysis of different substances and provide specific examples to illustrate the principles involved.

1. **Electrolysis of molten compounds in inert electrodes:** When molten ionic compounds are electrolysed, they decompose into their constituent elements because there are no other ions (eg from water) present to compete for discharge at the various electrodes.

a. Molten sodium chloride (NaCl)

Ions present: Na^+ and Cl^-

Electrode Reactions:

Cathode: $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$

Anode: $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$

Overall: $2\text{NaCl} \rightarrow 2\text{Na} + \text{Cl}_2$

Products:

Sodium metal (at cathode) and Chlorine gas (at anode)

b. Molten lead(II) bromide (PbBr_2)

Ions present : Pb^{2+} and Br^-

Electrode Reactions:

Cathode: $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$

Anode: $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$

Overall: $\text{PbBr}_2 \rightarrow \text{Pb} + \text{Br}_2$

Products: Lead metal (at cathode) and bromine vapour (at anode)

c. Molten aluminium oxide (Al_2O_3)

Ions present: Al^{3+} and O^{2-}

Electrode Reactions:

Cathode: $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$

Anode: $2\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-$

Overall: $2\text{Al}_2\text{O}_3 \rightarrow 4\text{Al} + 3\text{O}_2$

Products: Aluminium metal (at cathode) and oxygen gas (at anode)

2. Electrolysis of aqueous solutions with inert electrodes (graphite or platinum)

In aqueous solutions, both the dissolved ions and water molecules can be discharged. The products depend primarily on the position of ions in the electrochemical series and their concentration.

a. Dilute sodium chloride solution (NaCl)

Ions present: Na^+ , Cl^- , H^+ and OH^- (from H_2O)

Electrode Reactions:

Cathode: $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$

Anode: $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$

Overall: $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$

Products: Hydrogen gas (at cathode) and oxygen gas (at anode).

b. Concentrated sodium chloride solution (NaCl)

Ions present: Na^+ , Cl^- , H^+ and OH^- (from H_2O)

Cl^- is preferentially discharged at the anode because of its higher concentration

Electrode Reactions:

Cathode: $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$

Anode: $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$

Overall: $2\text{NaCl} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2 + \text{Cl}_2$

Products: Hydrogen gas (at cathode) and chlorine gas (at anode)

The solution becomes alkaline due to the formation of NaOH

c. Copper(II) sulphate solution (CuSO_4)

Ions present: Cu^{2+} , SO_4^{2-} , H^+ and OH^- (from H_2O)

Electrode Reactions:

Cathode: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

Anode: $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$

Overall: $2\text{CuSO}_4 + 2\text{H}_2\text{O} \rightarrow 2\text{Cu} + 2\text{H}_2\text{SO}_4 + \text{O}_2$

Products: Copper metal (at cathode) and oxygen gas (at anode)

The solution becomes acidic due to H^+ formation

The blue colour of the solution fades out since Cu^{2+} ions are being removed and deposited at the cathode

d. Dilute sulphuric acid (H_2SO_4)

Ions present: SO_4^{2-} , H^+ and OH^-

Electrode Reactions:

Cathode: $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$

Anode: $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$

Overall: $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$

Products: Hydrogen gas (at cathode) and oxygen gas (at anode)

e. Silver nitrate solution (AgNO_3)

Ions present: Ag^+ , NO_3^- , H^+ and OH^- (from H_2O)

Electrode Reactions:

Cathode: $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$

Anode: $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$

Overall: $4\text{AgNO}_3 + 2\text{H}_2\text{O} \rightarrow 4\text{Ag} + 4\text{HNO}_3 + \text{O}_2$

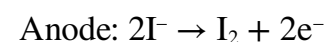
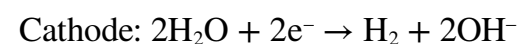
Products: Silver metal (at cathode) and oxygen gas (at anode)

The solution becomes acidic due to H^+ formation

f. Concentrated potassium iodide solution (KI)

Ions present: K^+ , I^- , H^+ and OH^-

Electrode Reactions:



Products: Hydrogen gas (at cathode) and iodine (at anode), which may appear as a brown solution. The solution becomes alkaline due to OH^- formation

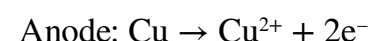
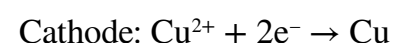
3. Electrolysis with active electrodes

When active electrodes are used, the electrode material itself may participate in the reactions, affecting the products formed.

a. Copper sulphate solution with copper electrodes

Ions present: Cu^{2+} , SO_4^{2-} , H^+ and OH^-

Electrode Reactions:



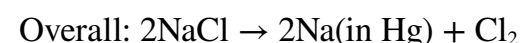
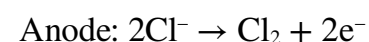
Overall: No net chemical change, copper transfers from anode to cathode

Products: Copper metal deposits at the cathode and copper dissolves from the anode. The concentration of copper sulphate remains approximately constant.

b. Sodium chloride solution with mercury cathode

Ions present: Na^+ , Cl^- , H^+ and OH^- (from H_2O)

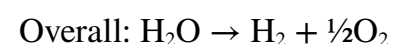
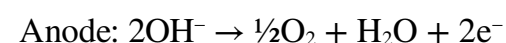
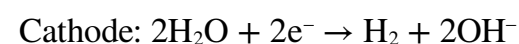
Electrode Reactions:



Products: Sodium amalgam (at mercury cathode) and chlorine gas (at anode)

c. Electrolysis of water using nickel-based electrodes

Electrode Reactions:



Products: Hydrogen gas (at cathode) and oxygen gas (at anode).

4. Electrolysis of copper(II) Sulphate Solution

a. Materials

Copper(II) sulphate solution (0.1 mol/dm^3)

Two graphite electrodes

DC power supply (6-12V)

Connecting wires with alligator clips

250 mL beaker

Universal indicator (optional)

Electronic balance (if measuring mass changes)

b. Procedure

- Pour approximately 150 mL of copper sulphate solution into the beaker
- Connect the graphite electrodes to the DC power supply
- Place the electrodes in the solution, keeping them separated
- Turn on the power supply and set to 6V
- Observe both electrodes for 10-15 minutes

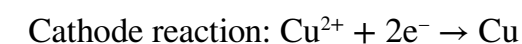
Record all observations including colour changes and gas formation (this gas can be collected and tested for, but students should **not** collect it as they will be exposing themselves to copper sulphate solution)

c. Expected Observations

Cathode (negative electrode): Reddish-brown deposit forms (copper metal). The blue colour of the solution fades near the cathode.

Anode (positive electrode): Small bubbles form (oxygen gas). The solution may become slightly acidic near the anode

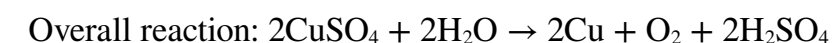
d. Explanation



Copper ions are reduced to copper metal. Copper is deposited because Cu^{2+} has a more positive reduction potential than H^+



Water is oxidised to produce oxygen gas. SO_4^{2-} ions are not discharged due to their higher oxidation potential.



5. Electrolysis of sodium chloride solution

a. Materials

Sodium chloride solution (0.1 mol/dm^3)

Two graphite electrodes

DC power supply (6-12V)

Connecting wires with alligator clips

250 cm³ beaker

Universal indicator

Gas collection apparatus (test tubes)

Wooden splint

b. Procedure

- i. Pour approximately 150 cm³ of 0.1 mol/dm³ NaCl solution into a beaker
- ii. Add a few drops of universal indicator
- iii. Connect the graphite electrodes to the DC power supply
- iv. Place the electrodes in the solution
- v. Turn on the power supply and set to approximately 6V
- vi. Observe colour changes around each electrode
- vii. Collect any gases produced using inverted test tubes
- viii. Test the gas from the cathode with a burning splint

c. Expected Observations**Dilute NaCl Solution**

Cathode: Hydrogen gas bubbles form (positive “pop” test), solution turns purple/blue (alkaline)

Anode: Oxygen gas bubbles form, solution turns red/orange (acidic)

Discussion of expected results from Concentrated NaCl Solution (which should not be done by students):

Cathode: Hydrogen gas bubbles form (positive “pop” test), solution turns purple/blue (alkaline)

Anode: Greenish-yellow gas forms (chlorine), bleaches blue litmus paper, distinctive smell.

d. Explanation**Dilute NaCl Solution**

Cathode: $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$

Anode: $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$

Overall: $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$

Concentrated NaCl Solution

Cathode: $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$

Anode: $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$

Overall: $2\text{NaCl} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2 + \text{Cl}_2$

6. Electrolysis of Dilute sulphuric acid**a. Materials**

Dilute sulfuric acid (0.5 mol/dm³)

Two graphite electrodes

DC power supply (6-12V)

Connecting wires with alligator clips

Hoffman voltmeter (if available)(do not attempt to collect the gas using submersion of gas collection tubes as this will expose students to acid)

Matches or wooden splints

Stopwatch

b. Procedure

- i. Fill the Hoffman voltmeter with dilute sulfuric acid
- ii. Connect the electrodes to the DC power supply
- iii. Turn on the power supply and set to approximately 6V
- iv. Observe gas formation at both electrodes
- v. Allow the electrolysis to proceed until sufficient gas has collected in both tubes of the Hoffman.
- vi. Measure the volume of gas in each tube
- vii. Test the gas from the cathode with a burning splint
- viii. Test the gas from the anode with a glowing splint

c. Expected Observations

Cathode: Hydrogen gas collects (positive “pop” test with burning splint)

Anode: Oxygen gas collects (relights a glowing splint)

d. Explanation

Cathode reaction: $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$

Hydrogen ions from the acid are reduced to hydrogen gas

Anode reaction: $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$. Water is oxidised to produce oxygen gas

Overall reaction: $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$

Learning Tasks

Your school chemistry club is learning about how metals are extracted using electricity (electrolysis). Your teacher has asked you to investigate this topic.

1. Name three factors that affect how well electrolysis works when extracting metals.
2. For each factor:
 - a. Explain how it affects the electrolysis process in simple terms
 - b. Describe how you could control this factor in a school laboratory experiment
 - c. What basic equipment would you need to measure or monitor this factor?
3. If you were demonstrating electrolysis to a lower class at your school's science fair, which factor would be easiest to show and explain? Give reasons for your choice.

Pedagogical Exemplars

Collaborative Learning: Components of an Electrolytic Cell

- 1. Form diverse mixed-ability groups of 4-5 learners.
- 2. Provide multiple learning resources (digital simulations, printed diagrams, videos).
- 3. Have groups create labelled diagrams identifying key components:
 - a. Anode (positive electrode)
 - b. Cathode (negative electrode)
 - c. Electrolyte
 - d. Power supply or source
 - e. Direction of electron flow
- 4. Each member contributes to presentations, ensuring all learners participate.
- 5. Facilitate respectful feedback addressing misconceptions.

Activity-Based Learning: Banking Analogy Role-Play

- 1. Set up “Anode Bank” and “Cathode Bank” in the classroom.
- 2. Assign diverse roles (bank tellers, customers/ions, bank manager).
- 3. Explain how the activity represents an electrolytic cell system.
- 4. Distribute “electron tokens” to participants.
- 5. At Anode Bank: Customers deposit electron tokens (oxidation).
- 6. At Cathode Bank: Customers receive electron tokens (reduction).
- 7. Guide reflection on how this models electron flow in electrolytic cells.
- 8. Have learners document the process and create diagrams.

Talk-for-Learning: Selective Discharge and Products

- 1. Begin with individual reflection on factors affecting ion discharge.
- 2. Guide discussion on key factors:
 - a. Position in electrochemical series
 - b. Concentration of ions
 - c. Nature of electrodes
 - d. Applied voltage
 - e. Temperature
- 3. Demonstrate predicting electrolysis products with examples.
- 4. Have learners predict products for different scenarios:
 - a. Molten salts (e.g., NaCl)
 - b. Dilute salt solutions
 - c. Concentrated salt solutions

- d. Systems with inert vs. active electrodes

- 5. Encourage learners to explain their reasoning using half-equations.

Inquiry-Based Learning: Circus Activity on Electrolysis

- 1. Conduct comprehensive safety briefing.
- 2. Have groups design experiments for investigating:
 - a. Electrolysis of copper(II) sulphate solution
 - b. Electrolysis of sodium chloride solution (dilute only)
 - c. Electrolysis of dilute sulphuric acid (in a Hoffman apparatus)
- 3. Set up experiment stations with clear instructions.
- 4. Have groups rotate through stations, recording observations.
- 5. Guide learners to test for products formed and write balanced equations.
- 6. Facilitate collaborative analysis of results.
- 7. Connect findings to industrial applications and environmental considerations.

Key Assessment Level 4: Investigation

- 1. Complete the table below by identifying the products formed at each electrode during the electrolysis of the following substances using inert electrodes: (4 marks)

Electrolyte	Product at Anode	Product at Cathode
Molten NaCl		
Dilute H ₂ SO ₄		

- 2. When dilute sodium sulphate solution is electrolysed: (4 marks)
 - a. Write the half-equations for the reactions occurring at both electrodes. (2 marks)
 - b. Explain why the solution around the anode becomes acidic while the solution around the cathode becomes alkaline. (2 marks)
- 3. A learner conducted an experiment to investigate the electrolysis of copper (II) sulphate solution. (6 marks)

The learner set up two different electrolytic cells:

Cell A: Copper electrodes with copper (II) sulphate solution

Cell B: Carbon electrodes with copper (II) sulphate solution

The learner passed the same current through both cells for 20 minutes and made the following observations:

Cell	Anode Observation	Cathode Observation	Solution Colour Change
A	Mass decreased	Mass increased	Mass increased
B	Gas bubbles	Copper deposit	Became paler

- a. Explain the observations for Cell A in terms of the electrochemical processes occurring. (2 marks)
 - b. Explain why the observations for Cell B differ from those for Cell A. (2 marks)
 - c. If the learner wanted to determine the electrochemical equivalent of copper from this experiment, which cell would be more appropriate to use? Justify your answer. (2 marks)
4. Design a comprehensive investigation to explore how different factors affect the efficiency of electrolysis in aqueous solutions. (6 marks)

Your investigation should:

- a. Identify at least three key variables that could affect electrolysis efficiency
- b. Propose specific experiments to isolate and test each variable
- c. Describe how you would measure and calculate efficiency
- d. Explain how you would control confounding variables
- e. Discuss how your findings could be applied to improve industrial electrolysis processes
- f. Consider limitations of your investigation and suggest improvements.

FOCAL AREA 2: APPLICATIONS OF ELECTROLYSIS

Electroplating is the process of coating one metal with a thin layer of another metal using electrolysis. The object to be plated serves as the cathode, while the coating metal serves as the anode, with a solution containing ions of the coating metal as the electrolyte.

The object to be plated (cathode) is thoroughly cleaned to remove dirt, grease, and oxides

The coating metal (anode) and the object are immersed in an electrolyte containing ions of the coating metal.

When current flows, metal ions from the solution are reduced and deposited on the cathode

Simultaneously, metal atoms from the anode are oxidised, replenishing the solution with metal ions.

Practical applications of electrolysis

1. Jewellery production (gold, silver plating)
2. Corrosion protection (zinc plating on steel/galvanising)
3. Decorative finishes (chrome plating)
4. Electronics (copper and gold plating on circuit boards)
5. Automotive parts (chrome bumpers, trim)

Experiment: Copper electroplating

Materials

Copper sulphate solution (0.1M)

Copper anode (strip or wire)

Object to plate (key, coin, metal item)

DC power supply (3-6V)

Connecting wires with clips

Beaker or plastic container

Sandpaper or steel wool

Degreasing agent (dish soap)

Vinegar solution (for cleaning)

Mass balance

Procedure

1. Clean the object to be plated: Rub with fine sandpaper to remove oxides
 - a. Wash with soap solution to remove grease
 - b. Rinse with water
 - c. Dip briefly in vinegar solution
 - d. Rinse again with water
2. Set up the electroplating cell:
 - a. Fill container with copper sulphate solution
 - b. Connect the object to the negative terminal (cathode)
 - c. Connect copper strip to the positive terminal (anode)
 - d. Ensure both are immersed but not touching
3. Run the electroplating process:
 - a. Apply 3-4V DC for 15-30 minutes
 - b. Observe the gradual formation of copper layer
 - c. Periodically check the quality of deposition
4. Complete the process:
 - a. Remove the plated object
 - b. Rinse thoroughly with water
 - c. Dry completely
 - d. Examine the quality and coverage of plating
5. Observations and Analysis:
 - a. Copper should deposit as a reddish-brown layer on the object
 - b. The copper anode should show signs of erosion
 - c. The quality of plating depends on cleanliness, current, and time
 - d. Uneven current distribution leads to uneven plating

Extraction of aluminium by electrolysis

Aluminium is the most abundant metal in the Earth's crust, yet it is never found in its free state due to its high reactivity with oxygen. Traditional reduction methods using carbon are inadequate because aluminium has a greater affinity for oxygen (reduction potential) than carbon does. Electrolysis becomes necessary for extraction, but pure alumina has an extremely high melting point of 2072°C, making direct electrolysis impractical. This challenge led to the development of the Hall-Héroult process, which revolutionised aluminium production.

Step 1: Bauxite to Alumina (Bayer Process)

1. Bauxite ore is crushed and ground to increase surface area.
2. The crushed ore is mixed with hot sodium hydroxide solution under pressure.
3. Aluminium oxide dissolves to form sodium aluminate solution.
4. Insoluble impurities (iron oxides, silica) are filtered out as "red mud."
5. The filtered solution is cooled to precipitate aluminium hydroxide.
6. The aluminium hydroxide is heated to approximately 1100°C to form pure alumina (Al_2O_3).

Step 2: Electrolytic Cell Setup

1. A steel container is lined with carbon (graphite and pitch) which serves as the cathode.
2. Carbon anodes are suspended from above into the cell.
3. Cryolite (Na_3AlF_6) is added to the cell and melted at approximately 960°C.
4. Alumina is dissolved in the molten cryolite (typically 2-8% concentration).
5. The carbon anodes are connected to the positive terminal of a DC power supply.
6. The carbon-lined cell is connected to the negative terminal.

Step 3: Electrolysis Process

1. High-amperage direct current (typically 100,000-150,000 amperes) passes through the cell.
2. Aluminium ions are reduced at the cathode: $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$.
3. Oxide ions are oxidised at the carbon anode: $2\text{O}^{2-} + \text{C} \rightarrow \text{CO}_2 + 4\text{e}^-$.
4. Molten aluminium forms and sinks to the bottom of the cell due to its higher density.
5. The molten aluminium is periodically tapped off using vacuum siphons.
6. Carbon anodes are gradually consumed and regularly replaced.
7. Fresh alumina is continuously added to maintain the concentration in the electrolyte.

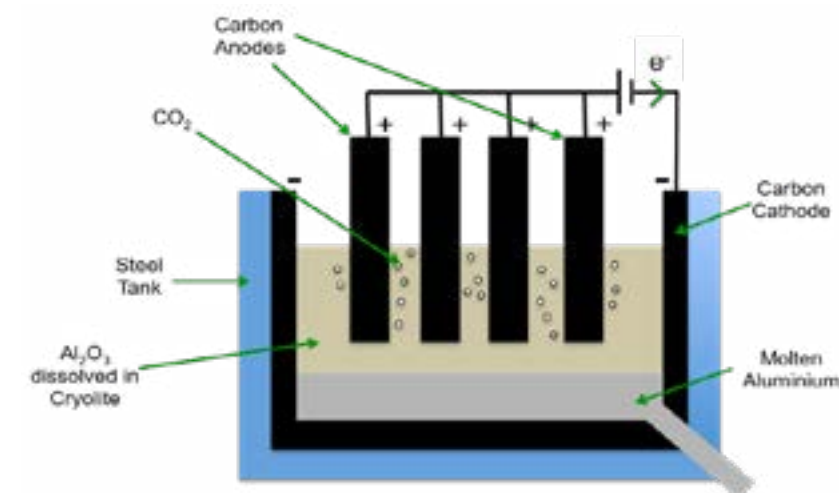


Figure 15.2: Hall-Héroult Electrolysis Process (From Alumina to Pure Aluminium)

Extraction of gold by electrolysis

Gold is often found in its native state but is typically mixed with other metals such as silver, copper, and platinum group metals. Many applications, particularly in electronics and jewellery, require gold of very high purity. Traditional purification methods like cupellation can achieve limited purity levels, but electrolytic refining allows for the production of extremely pure gold. This high-purity requirement led to the development of the Wohlwill process.

Step 1: Preparation

1. Impure gold (typically 95-98% pure) from preliminary refining processes is collected.
2. The impure gold is melted and cast into anodes of appropriate size and shape.
3. Thin sheets of high-purity gold are prepared to serve as cathodes.
4. An electrolyte solution of gold chloride and hydrochloric acid is prepared.
5. A temperature-controlled tank is set up and maintained at 60-70°C.

Step 2: Electrolysis Setup

1. The impure gold anodes are suspended in the electrolyte solution.
2. The pure gold cathodes are positioned in the same solution.
3. Proper spacing between anodes and cathodes is ensured.
4. The anodes are connected to the positive terminal of a DC power supply.
5. The cathodes are connected to the negative terminal.
6. The tank is covered to minimise evaporation and contamination.

Step 3: Electrolysis Process

1. Controlled direct current at appropriate current density passes through the cell.
2. Gold dissolves from the anode: $\text{Au} \rightarrow \text{Au}^{3+} + 3\text{e}^-$.
3. Gold ions are reduced at the cathode: $\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$.
4. Pure gold deposits gradually build up on the cathode surface.
5. Impurities either remain in solution (if more reactive than gold) or form "anode slime" at the bottom (if less reactive).

- The process continues until anodes are substantially dissolved.
- Cathodes are removed, washed, and the pure gold deposit is melted and cast.
- Anode slime is collected for recovery of valuable metals (silver, platinum, etc.).

Faraday's Laws of Electrolysis

Faraday's laws of electrolysis, formulated by Michael Faraday in 1833-1834, establish the quantitative relationships between electrical charge and chemical change during electrolysis. These fundamental principles form the cornerstone of electrochemistry, connecting the worlds of electricity and chemical reactions in precise mathematical terms. Faraday's meticulous experiments with various electrolytes led to these revolutionary laws that not only advanced scientific understanding but also paved the way for numerous industrial applications from metal extraction to modern battery technology.

First Law

The mass of a substance produced at an electrode during electrolysis is directly proportional to the quantity of electricity (charge) passed through the electrolyte.

Mathematical expression: $m \propto Q$ or $m = k \times Q$

Where:

m = mass of substance produced (g)

Q = electrical charge transferred (coulombs)

k = electrochemical equivalent of the substance

Since $Q = I \times t$ (current $\times$ time): $m = k \times I \times t$

Second Law

For the same quantity of charge, the masses of different substances produced at electrodes are proportional to their chemical equivalent weights (atomic mass divided by the number of electrons transferred).

Mathematical Expression: $m_1/m_2 = E_1/E_2$

Where:

m_1, m_2 = Masses of substances

E_1, E_2 = Chemical equivalent weights

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)

Worked examples

- A gold refinery in Obuasi processes gold from local mines. During the final purification stage, they use the Wohlwill electrolytic process with a gold chloride solution. If they need to deposit 250 grams of pure gold onto the cathode, calculate the time required when using a current of 15 amperes.

Solution:

Using Faraday's law: $m = \frac{(M \times I \times t)}{(n \times F)}$

Rearranging to find t: $t = \frac{(m \times n \times F)}{(M \times I)}$

Where:

m = mass of gold = 250 g

M = molar mass of gold = 197.0 g/mol

I = current = 15 A

n = electrons transferred per ion = 3 ($\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$)

F = Faraday constant = 96,500 C/mol

$$t = \frac{250 \times 3 \times 96,500}{197.0 \times 15}$$

$$t = \frac{72,375,000}{2,955}$$

$$t = 24,492 \text{ seconds} = 6.80 \text{ hours}$$

Therefore, it would take approximately 6 hours and 48 minutes to deposit 250 grams of gold using a current of 15 amperes.

- The VALCO aluminium smelter in Ghana operates electrolytic cells for aluminium production. In one cell, a current of 125,000 amperes flows through molten cryolite containing dissolved alumina. Calculate the mass of aluminium produced in this cell over a 24-hour period if the cell operates at 92% efficiency.

Solution:

Using Faraday's law with efficiency factor: $= \frac{(M \times I \times t \times \eta)}{(n \times F)}$

Where:

m = mass of aluminium produced

M = molar mass of aluminium = 27.0 g/mol

I = current = 125,000 A

t = time = $24 \times 60 \times 60 = 86,400$ seconds

η = efficiency = 92% = 0.92

n = electrons transferred per ion = 3 ($\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$)

F = Faraday constant = 96,500 C/mol

$$m = \frac{27.0 \times 125,000 \times 86,400 \times 0.92}{3 \times 96,500}$$

$$m = \frac{270,864,000,000 \times 0.92}{289,500}$$

$$m = \frac{249,195,000,000}{289,500}$$

$$m = 860,777 \text{ g} = 860.8 \text{ kg}$$

Therefore, approximately 860.8 kg of aluminium would be produced in the cell over a 24-hour period at 92% efficiency.

Learning Tasks

1.

A metal plating workshop in Accra needs to deposit 50.0 grams of silver onto decorative items. Using Faraday’s laws of electrolysis, calculate:

a.

The quantity of electricity (in coulombs) required for this deposition

b.

Show all your work and explain each step of your calculation

(Use: $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$, atomic mass of Ag = 107.9 g/mol, Faraday constant = 96,500 C/mol)
2.

Coastal regions of West Africa experience accelerated corrosion of metal structures due to salt air and humidity. Analyse which metals would be most suitable as protective anodes in this environment:

a.

List three potential metals for protective anodes

b.

Explain why each metal would or would not be suitable based on its position in the electrochemical series

c.

Recommend the best option for coastal environments, considering availability and cost.

Pedagogical Exemplars

Inquiry-Based Learning: Electroplating Demonstration

1.

Set up 3-4 electroplating stations with copper sulphate solution, copper anodes, metal objects for plating, and DC power supplies.
2.

Conduct safety briefing on handling chemicals and electrical equipment.
3.

Have mixed-ability groups design their electroplating procedure.
4.

Groups prepare objects (clean, degrease, acid dip, rinse) before connecting as cathodes.
5.

Learners run electroplating for 15-20 minutes while recording observations.
6.

Groups examine plated objects and discuss factors affecting quality.
7.

Connect to real-world applications in Ghana (jewellery, electronics, preservation).
8.

Each group presents their results and suggests improvements.

Talk-for-Learning: Metal Extraction and Faraday’s Laws

A. Extraction of Aluminium and Gold

1.

Show samples/images of raw materials and refined metals from Ghana.
2.

Guide discussion on the Hall-Héroult process for aluminium extraction.
3.

Explain the Wohlwill process for gold refining.
4.

Facilitate learner-led discussion on environmental impacts of mining in Ghana.
5.

Connect industrial processes to Ghana’s economic development.

B. Faraday’s Laws of Electrolysis

1.

Introduce Faraday’s First and Second Laws with clear explanations.
2.

Present the combined equation: $m = \frac{M \times I \times t}{n \times F}$
3.

Work through example calculations as a class.
4.

Have learners practice calculations in mixed-ability groups.
5.

Discuss industrial applications in Ghanaian context.
6.

Assess understanding through targeted questions.

Key Assessment Level 4: Computations

1.

A silver-plating facility in Kumasi needs to coat traditional Ashanti jewellery with a uniform layer of silver. They want to deposit a silver layer that is 0.015 mm thick on a collection of items with a total surface area of 0.25 m². The density of silver is 10.5 g/cm³. Calculate the current required to complete this plating process in 3 hours.
2.

Ghana has been exploring ways to reduce corrosion of metal structures in coastal areas where salt air accelerates deterioration. Cathodic protection using sacrificial anodes is one potential solution.

a.

Compare and contrast galvanic cells and electrolytic cells in the context of preventing corrosion of steel structures in Ghana’s coastal regions.

b.

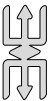
Explain the electrochemical principles behind cathodic protection

c.

Analyse which metals would be most suitable as sacrificial anodes in Ghana’s coastal environment.

d.

Evaluate the environmental and economic implications of using this technology compared to traditional anti-corrosion methods.



Note

Please use the same instruction template for PLC Sessions 15 and 16, but choose a different assessment strategy for the second week.”

PLC Session 15 & 16

45 minutes

Planning & Reviewing the Learning Plan

1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week’s concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Plan the week’s key assessment with your app
 - a. The key assessment for the week is *investigation and computations*. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the assessments (NTS 3k-3q).

10 minutes

Content Exploration

5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
 - a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step
 - d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?

- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes

Enactment

6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
 - a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes

Reflection

7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
 - a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

WEEK 17

Learning Indicator: Explain the concept of corrosion of metals and its application in everyday life

FOCAL AREA 1: CORROSION OF METALS

Ghana's tropical climate, extensive coastline, and developing industrial landscape create perfect conditions for metal deterioration. From the salty breeze of coastal cities to the alternating wet and dry conditions of the northern regions, corrosion challenges are widespread and economically significant. Everyday structures suffer, with iron roofing sheets replacing traditional thatch only to deteriorate rapidly without proper protection.

Historical buildings like Cape Coast and Elmina Castles require constant maintenance against coastal elements. Understanding corrosion science is therefore essential for sustainable development by mastering the electrochemical processes involved and appropriate prevention techniques.

Corrosion is the deterioration of a metal as a result of chemical reactions between it and the surrounding environment. It is an electrochemical process where the metal is oxidised, changing from a metallic state to a non-metallic state, usually an oxide, hydroxide, or salt.

The general electrochemical corrosion process involves two half-reactions:

Anodic Reaction (Oxidation): $M \rightarrow M^{n+} + ne^-$ Where M represents a metal atom, M^{n+} is the metal ion, and n is the number of electrons (ne^-) transferred.

Cathodic Reaction (Reduction): In acidic solutions: $2H^+ + 2e^- \rightarrow H_2$ and in neutral or alkaline solutions: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

For example, when zinc corrodes in an acidic environment:

Anodic reaction: $Zn \rightarrow Zn^{2+} + 2e^-$

Cathodic reaction: $2H^+ + 2e^- \rightarrow H_2$

Overall reaction: $Zn + 2H^+ \rightarrow Zn^{2+} + H_2$

Electrochemical nature of corrosion

Most metal corrosion occurs through electrochemical processes involving the transfer of electrons. These processes require:

1. An anode (where oxidation occurs and metal atoms lose electrons)
2. A cathode (where reduction occurs and electrons are consumed)
3. An electrolyte (a conducting solution that allows ion movement)
4. An electrical connection between the anode and cathode

Factors affecting corrosion rate

1. Nature of the metal: Reactivity, purity, surface condition, and metallurgical structure.
2. Environmental conditions: Temperature, humidity, presence of pollutants, pH, and oxygen concentration.

3. Electrolyte concentration: Higher electrolyte concentrations generally increase corrosion rates.
4. Temperature: Higher temperatures typically accelerate corrosion reactions.
5. Oxygen concentration: Oxygen is often the oxidising agent in corrosion processes.
6. pH value: Acidic environments generally accelerate corrosion of most metals.

Rusting

Rusting is a specific form of corrosion that affects iron and its alloys. It produces iron (III) oxide ($Fe_2O_3 \cdot nH_2O$), 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 ($Fe_2O_3 \cdot nH_2O$).

The rusting process involves several steps:

1. Iron loses electrons at anodic sites: $Fe \rightarrow Fe^{2+} + 2e^-$
2. Electrons reduce oxygen at cathodic sites:
 $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ (in acidic conditions)
 $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$ (in neutral/basic conditions)
3. Iron(II) ions are further oxidised to iron(III):
 $4Fe^{2+} + O_2 + 4H^+ \rightarrow 4Fe^{3+} + 2H_2O$
4. Iron(III) ions form hydrated iron(III) oxide (rust):
 $Fe(3+) + 3OH^- \rightarrow Fe(OH)_3 \rightarrow Fe_2O_3 \cdot nH_2O$ (rust)

Experiment to Investigate the Conditions Needed for Rusting

Materials

Clean iron nails

Test tubes with stoppers

Test tube rack

Boiled distilled water (oxygen-free)

Tap water

Anhydrous calcium chloride

Oil

Labels

Procedure

Setup A:

1. Place a clean iron nail in a test tube with boiled distilled water (oxygen-free).
2. Add a layer of oil on top to prevent oxygen from dissolving and stopper the tube.

Setup B: Place a clean iron nail in a test tube with tap water (contains dissolved oxygen) and stopper the tube.

Setup C: Place a clean iron nail in a dry test tube with anhydrous calcium chloride (to keep the environment dry) and stopper the tube.

Setup D: Place a clean iron nail in an empty test tube exposed to air and stopper the tube.

Label all test tubes and observe for 1 week.

Expected Results

Setup A: No rusting (water present but no oxygen)

Setup B: Rusting occurs (both water and oxygen present)

Setup C: No rusting (oxygen present but no water)

Setup D: Minimal or no rusting (oxygen present, but limited moisture)

Conclusion

Both oxygen and water are necessary for rusting to occur. If either is absent, rusting will not take place.

Experiment to investigate the conditions that affect rate of rusting

Materials

Iron nails of uniform size

Test tubes with stoppers

Test tube rack

Sodium chloride solution (various concentrations)

Distilled water

Dilute acids (HCl)

Tap water

Thermometer

Water bath

Iron filings and iron nails (different surface areas)

Labels

Procedure

1. Effect of Electrolytes

Place identical iron nails in separate test tubes containing:

- Distilled water
- 1% NaCl solution
- 5% NaCl solution
- 10% NaCl solution

Observe for 3-5 days and record the extent of rusting.

2. Effect of pH

Place identical iron nails in separate test tubes containing:

- Neutral water (pH 7)
- Slightly acidic solution (pH 5)
- Acidic solution (pH 3)

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.

- Keep one at room temperature,
- One in a refrigerator and
- One in a warm water bath.

Observe for 3-5 days and record the extent of rusting.

4. Effect of Surface Area

- Place an iron nail in one test tube
- Place an equivalent mass of iron filings in another test tube.
- Add equal volumes of salt water to both.

Observe for 3-5 days and record the extent of rusting.

Expected Results

- Electrolytes: Higher salt concentration leads to faster rusting
- pH: Lower pH (more acidic) accelerates rusting
- Temperature: Higher temperature increases the rate of rusting
- Surface Area: Larger surface area (iron filings) results in faster rusting

Conclusion: The rate of rusting increases with:

- Higher concentration of electrolytes
- Lower pH (more acidic conditions)
- Higher temperature
- Greater surface area of the metal

Experiment to investigate ways of preventing rusting

Materials

Clean iron nails

Test tubes with stoppers

Test tube rack

Paint

Oil

Zinc strips

Magnesium strips

Copper strips

Grease

Galvanised nails

Salt water solution (electrolyte)

Labels

Procedure

1. Physical barriers:

- Coat different iron nails with paint, oil and grease
- Leave one nail uncoated as a control.
- Place each in separate test tubes with salt water.

2. Sacrificial protection:

- Connect clean iron nails to zinc strip, magnesium strip and copper strip
- Place each pair in separate test tubes with salt water.
- 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.

Observe all setups for 1 week and record observations.

Expected results

- Physical Barriers: Nails coated with paint, oil, or grease show little to no rusting
- 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)
- 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.

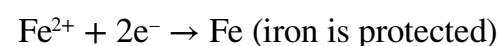
Redox and non-Redox methods of preventing rusting

Redox Methods (Electrochemical protection)

1. Sacrificial Protection (Cathodic Protection)

Process: A more reactive metal (sacrificial anode) is connected to the iron. The reactive metal oxidises preferentially, protecting the iron. Examples: zinc (galvanising), magnesium anodes, aluminium anodes.

Chemical Principle: The sacrificial metal has a more negative electrode potential than iron, making it more likely to lose electrons.



2. Impressed current cathodic protection

Process: An external DC power source provides electrons to the metal being protected.

Applications: Underground pipelines, ship hulls, water heaters.

The advantage of this method is that it can protect large structures over long distances

3. Alloying

Process: Creating alloys that are more resistant to corrosion. Examples:

Stainless steel (iron with chromium and nickel)

Weathering steel (contains copper, chromium, and nickel)

Mechanism: Forms a passive oxide layer that protects the underlying metal

Non-Redox Methods (Barrier protection)

1. Surface Coating

- Painting: Paint creates a physical barrier preventing contact with water and oxygen
- Oil/Grease: Provides a water-repellent layer
- Plastic Coating: Durable barrier protection
- Powder Coating: Electrostatic application of dry powder followed by curing
- Enamel: Glass-like coating fused to metal surface

- Passivation: Formation of a thin inert layer on the metal surface, e.g. chromate conversion coating, phosphate coating, etc.

Impact of rusting and corrosion in everyday life

Financial costs

- Significant expenses are incurred for maintenance and replacement of corroded infrastructure.
- Corroded equipment operates at reduced efficiency, leading to economic losses.
- Preventive measures and corrosion-resistant materials often increase initial costs of projects.

Industrial impacts

- Corrosion shortens the lifespan of industrial equipment, requiring premature replacement.
- Corrosion by-products can contaminate products, reducing quality and market value.
- Energy losses occur due to reduced efficiency of corroded equipment and systems.

Safety concerns

- Corrosion causes weakening of bridges, buildings, and critical infrastructure.
- Failure of critical components in vehicles and machinery can occur due to corrosion.
- Metal structures may collapse due to stress corrosion cracking, endangering lives.

Transportation safety

1. Corrosion of vehicle brake lines, chassis, and structural components compromises road safety.
2. Aircraft component failures due to corrosion pose serious aviation safety risks.
3. Railway track and train component degradation from corrosion threatens rail transportation safety.

Utility infrastructure

1. Water pipe leakage and contamination result from corroded water distribution systems.
2. Gas pipeline failures due to corrosion can lead to dangerous leaks and potential explosions.
3. Power outages are frequently caused by corroded electrical connections in the power grid.
4. Contamination of drinking water with heavy metals can occur from corroded pipes.

Daily life

1. Household appliance failures
2. Vehicle maintenance costs
3. Deterioration of metal roofing and construction materials
4. Water supply contamination from corroded pipes

Learning Tasks

1. Learners will test common household items to prevent rusting on iron nails.
Instructions:
 - a) In pairs, select three household items (e.g., cooking oil, paint, petroleum jelly)
 - b) Coat iron nails with each substance
 - c) Place the coated nails and one uncoated nail (control) in cups of salt water
 - d) Observe and record changes every day for 5 days
 - e) Create a poster showing your results with before/after pictures
2. Learners will create a basic galvanic cell using fruits or vegetables.
Instructions:
 - a) Create a simple cell using a lemon/potato and two different metals (copper and zinc)
 - b) Use the cell to light a small LED or power a simple digital clock
 - c) Draw a labelled diagram showing how your cell works
 - d) Explain how this relates to corrosion in metals.

Pedagogical Exemplars**Initiating Talk-for-Learning**

1. Begin by showing learners two metal objects one rusted and one un-rusted.

Ask: “What happened to this one? Why not the other?”

2. Think-Pair-Share Activity:

Think (1 minute): learners individually write definitions of corrosion and rusting

Pair (2 minutes): Partners compare definitions and refine them

Share (3 minutes): Pairs present their definitions to class

3. Show images of the statue of Liberty (copper corrosion) and rusted vehicles, asking learners to identify differences.
4. Divide class into teams. Each team member adds one step to the electrochemical equation for rusting, passing a marker to the next person.
5. Compare rusting to “metal decay” using fruit decay time-lapse video, highlighting similarities in oxidation processes. (<https://youtu.be/2jeYXGJyY9k>)

Experiential Learning

1. Provide mystery setups with rusted and non-rusted nails. learners must determine what conditions were different.
2. Groups compete to create conditions that produce rust fastest on identical nails.
3. Present a crime scene where corrosion caused a structural failure. Learners gather evidence to determine causes.
4. Groups design, test, and present the most effective rust prevention method with limited materials.
5. Each group tests one variable (salt, temperature, etc.) and creates a visual graph of results to share.

Talk-for-Learning Reflection

1. Gallery Walk: Groups post experiment results and prevention methods; learners rotate, adding sticky-note questions/comments.
2. Rust Prevention Advertisement: Teams create 60-second commercials for their most effective prevention method.
3. Real-Life Impact: Set up stations with photos of corrosion impacts (bridges, cars, pipes). Learners rotate, discussing economic and safety implications.
4. Concept Mapping Challenge: Teams race to create the most comprehensive concept map connecting rusting processes, prevention, and impacts.
5. Video Analysis: Show a short video on corrosion (<https://youtube/P5ra9BMtYLo>), having learners complete a structured note-taking guide with “What surprised you most?” reflection.

Key Assessments Level 4: Poster Presentation

1. Create an educational poster about the impacts of corrosion in your community. Your poster should include:
 - a) Three examples of corrosion with pictures/drawings

- b) Explanation of the type of corrosion in each example
 - c) Two negative impacts for each example (economic, safety, etc.)
 - d) One prevention method for each example
2. Conduct a simple experiment to test how salt concentration affects rusting rate. Write a 1-2 page report including:
- a) Your hypothesis
 - b) Materials and procedure
 - c) Results with a simple data table and graph
 - d) Conclusion explaining what happened and why.

PLC Session 17

45 minutes	Planning & Reviewing the Learning Plan
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1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
- a. explain the week’s concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to
- a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



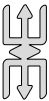
Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Plan the week’s key assessment with your app
- a. The key assessment for the week is **poster**. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the poster presentation (NTS 3k-3q).

10 minutes	Content Exploration
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5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
- a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step
 - d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
- a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
- a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

Section 4 Review

This section has focused on understanding oxidation-reduction (redox) reactions and their diverse applications. During Weeks 8–17, learners have progressed from fundamental redox concepts to complex applications in electrochemistry, metallurgy, and corrosion science.

Key Concepts Covered

Oxidation and Reduction Fundamentals: Learners should understand multiple definitions of **oxidation and reduction**: as oxygen/hydrogen transfer, electron transfer, and changes in oxidation numbers. They should be able to identify oxidation and reduction processes in various reaction types including combination, decomposition, combustion, displacement, disproportionation, and biological processes.

Metal Reactivity: Learners should understand the reactivity series of metals, be able to predict displacement reactions, and explain reactivity patterns in terms of electron-donating tendency.

Oxidising and Reducing Agents: Learners should be able to identify oxidising and reducing agents in reactions, understand their characteristics, perform qualitative tests for their presence, and recognise their applications in various industries.

Redox Equations: Learners should be proficient in writing half-equations for oxidation and reduction processes and balancing complex redox equations in both acidic and basic media.

Redox Titrations: Learners should understand the principles of redox titrations, be able to perform titrations involving potassium permanganate and iodometric methods, and calculate analyte concentrations from experimental data.

Electrochemical Cells: Learners should understand how voltaic cells generate electricity through redox reactions, factors affecting electrode potentials, the standard hydrogen electrode reference system, and how to calculate cell potentials and predict reaction feasibility.

Batteries and Fuel Cells: Learners should understand the principles and applications of various battery types (lead-acid, alkaline, lithium-ion, nickel-cadmium) and fuel cells, including their half-reactions and environmental considerations.

Electrolytic Cells: Learners should understand the principles of electrolysis, factors affecting selective discharge, predict electrolysis products for various electrolytes, and understand quantitative aspects through Faraday's laws.

Applications of Electrolysis: Learners should understand practical applications including electroplating and metal extraction, with awareness of environmental implications.

Corrosion: Learners should understand corrosion as an electrochemical process, the specific conditions required for rusting, factors affecting corrosion rates, and both redox and non-redox methods for corrosion prevention.

MARKING SCHEMES FOR THE KEY ASSESSMENTS

WEEK 8 & 9

FOCAL AREA 1

(10 marks in total)

Loss of electrons and increase in oxidation number (5 marks)

Definition of oxidation number (1 mark)

Statement that loss of electrons results in a more positive charge on the atom (1 mark)

Explanation that when an atom loses electrons, its oxidation number increases (1 mark)

Correct mathematical relationship: if an atom loses n electrons, its oxidation number increases by n (e.g., $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$). (2 marks)

Gain of electrons and decrease in oxidation number (5 marks)

Statement that gain of electrons results in a more negative charge on the atom (1 mark)

Explanation that when an atom gains electrons, its oxidation number decreases (1 mark)

Correct mathematical relationship: if an atom gains n electrons, its oxidation number decreases by n (e.g., $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$). (2 marks)

Explanation that reduction involves gain of electrons and decrease in oxidation number. (1 mark)

(10 marks in total)

Determine the oxidation number of gold in the ore and in the complex $\text{Na}[\text{Au}(\text{CN})_2]$ (6 marks)

Identification that gold in the ore (Au) has oxidation number 0 (1 mark)

Recognition that Na has oxidation number +1 in the complex (1 mark)

Recognition that CN has an overall oxidation number of -1 in the complex (1 mark)

Correct application of the rule that the sum of oxidation numbers equals the overall charge of the complex (1 mark)

Calculation: $\text{Na}[\text{Au}(\text{CN})_2]: +1 + x + 2(-1) = 0$, where x is the oxidation number of Au (1 mark)

Correct determination that gold has oxidation number +1 in the complex $\text{Na}[\text{Au}(\text{CN})_2]$ (1 mark)

Identify which element is oxidised and which is reduced in this process (4 marks)

Identification that gold is oxidised (1 mark)

Explanation that gold's oxidation number increases from 0 to +1, indicating oxidation (1 mark)

Identification that oxygen is reduced (1 mark)

Explanation that oxygen's oxidation number decreases from 0 in O_2 to -2 in NaOH, indicating reduction (1 mark)

Why most metals are found in nature as compounds rather than in their elemental form. (10 marks)

Understanding of Metal Reactivity (3 marks)

2 marks: Basic explanation of metal reactivity and tendency to form positive ions, but without detailed reference to the reactivity series or electron transfer.

1 mark: Simple statement that metals are reactive without substantive explanation.

Thermodynamic Stability (3 marks)

2 marks: Basic understanding that metal compounds are more stable than elemental metals, but without reference to energy considerations or with minor misconceptions.

1 mark: Simple statement about stability without explanation.

Environmental Factors and Examples (2 marks)

1 mark: Basic mention of environmental factors or examples without clear connection between them.

Award up to 1 mark for exceptional answers that include relevant equations, specific ore examples, or discussion of extraction methods as they relate to the stability of metal compounds.

FOCAL AREA 2

Copper Vs. Iron for Water Pipes

Corrosion Resistance (3 marks)

3 marks: Detailed explanation of copper's superior corrosion resistance compared to iron, including discussion of copper's protective oxide layer that prevents further corrosion, versus iron's tendency to form rust that flakes off and continues the corrosion process.

2 marks: Basic explanation of copper's better corrosion resistance without detailed mechanism.

1 mark: Simple statement that copper resists corrosion better than iron without explanation.

0 marks: No mention of corrosion or incorrect explanation.

Malleability and Workability (3 marks)

3 marks: Clear explanation that copper is more malleable and ductile than iron, making it easier to form into pipes, bend around corners, and join without leaks.

1 mark: Basic mention of copper being easier to work with without specific properties.

0 marks: No mention of workability properties.

Antimicrobial Properties (3 marks)

3 marks: Discussion of copper's natural antimicrobial properties that help prevent bacterial growth in water systems, with explanation of how this improves water quality.

1 mark: Simple mention of antimicrobial properties without elaboration.

0 marks: No mention of antimicrobial properties.

Interpretation of Observations (5 marks)

5 marks: Clear deduction that X is the most reactive metal, followed by Y, with Z being the least reactive ($X > Y > Z$). Includes explanation that more reactive metals can displace less reactive metals from their salt solutions.

3 marks: Correct ordering of reactivity ($X > Y > Z$) without detailed explanation of displacement reactions.

2 mark: Partial ordering with errors or incomplete explanation.

0 marks: Incorrect interpretation of the reactivity order.

Identification of Possible Metals (5 marks)

5 marks: Accurate identification of possible metals that fit the observations, with justification based on the reactivity series. For example: X could be zinc or iron; Y could be lead or copper; Z could be silver or copper (if Y is not copper). Explanation that they must be below hydrogen in the reactivity series since none produce hydrogen with water.

3 marks: Reasonable suggestions for possible metals with minor errors in justification.

1 mark: Some correct metal suggestions without proper justification.

0 marks: Incorrect **metal suggestions that don't match the observed reactivity pattern.**

WEEK 10 - 12

FOCAL AREA 1

Identification of Solutions (15 marks)

Solution A (3 marks)

Correct identification: Copper(II) sulphate (CuSO_4) (1 mark)

Justification: Forms brown colour with KI due to oxidation of I^- to I_2 (Cu^{2+} acts as an oxidising agent) (1 mark)

Further justification: Does not react with KMnO_4 or $\text{K}_2\text{Cr}_2\text{O}_7$ as Cu^{2+} is already in a high oxidation state and cannot be further oxidised by these oxidising agents (1 mark)

Solution B (3 marks)

Correct identification: Sodium sulphite (Na_2SO_3) (1 mark)

Justification: Decolourises KMnO_4 because SO_3^{2-} acts as a reducing agent, reducing Mn^{7+} to Mn^{2+} (1 mark)

Further justification: Reduces $\text{K}_2\text{Cr}_2\text{O}_7$ (orange) to Cr^{3+} (green) as SO_3^{2-} is oxidised to SO_4^{2-} ; does not react with KI as SO_3^{2-} is not a strong enough oxidising agent to oxidise I^- (1 mark)

Solution C (3 marks)

Correct identification: Iron(II) sulphate (FeSO₄) (1 mark)

Justification: Forms brown colour with KI as Fe³⁺ (formed by partial oxidation of Fe²⁺ in acidic solution) oxidises I⁻ to I₂ (1 mark)

Further justification: Decolourises KMnO₄ as Fe²⁺ is oxidised to Fe³⁺; reduces K₂Cr₂O₇ to Cr³⁺ (green) as Fe²⁺ acts as a reducing agent (1 mark)

Solution D (3 marks)

Correct identification: Sodium chloride (NaCl) (1 mark)

Justification: Shows no reaction with any of the test reagents (1 mark)

Further justification: Cl⁻ is neither a strong oxidising agent nor a reducing agent under these conditions, and Na⁺ is stable in its oxidation state (1 mark)

Solution E (3 marks)

Correct identification: Hydrogen peroxide (H₂O₂) (1 mark)

Justification: Forms brown colour with KI as H₂O₂ oxidises I⁻ to I₂ (1 mark)

Further justification: Does not decolourise KMnO₄ as the concentration of H₂O₂ is likely low or the pH condition favours the stability of KMnO₄; reduces K₂Cr₂O₇ to Cr³⁺ (green) as H₂O₂ can act as a reducing agent in certain conditions (1 mark)

Flowchart Design (10 marks)

Example of a Flowchart Approach:

First Test: KMnO₄ test

Solutions that decolourise KMnO₄ (B and C): These contain reducing agents (1 mark)

Solutions that do not decolourise KMnO₄ (A, D, and E): These either contain oxidising agents or are unreactive (1 mark)

Second Test: KI test

For solutions that decolourised KMnO₄ (B and C):

If brown colour forms with KI: Solution C (FeSO₄)(1 mark)

If no reaction with KI: Solution B (Na₂SO₃)(1 mark)

For solutions that did not decolourise KMnO₄ (A, D, and E):

If brown colour forms with KI: Solutions A and E(1 mark)

If K₂Cr₂O₇ turns green: Solution E (H₂O₂)(1 mark)

If K₂Cr₂O₇ remains orange: Solution A (CuSO₄)(1 mark)

If no reaction with KI: Solution D (NaCl) (1 mark)

FOCAL AREA 2

Part 1: Oxidation number changes (6 marks)

Element	Change in oxidation number	Oxidised or reduced
Mn	+7 to +2 (change of -5)	Reduced
Fe	+2 to +3 (change of +1)	oxidised

Marking criteria:

- Correct identification of Mn oxidation number change: +7 to +2 (1 mark)
- Correct identification of Fe oxidation number change: +2 to +3 (1 mark)
- Correct numerical change for Mn: -5 (1 mark)
- Correct numerical change for Fe: +1 (1 mark)
- Correct identification of Mn as being reduced (1 mark)
- Correct identification of Fe as being oxidised (1 mark)

Part 2: Balanced half-reaction for permanganate reduction (4 marks)

Correct answer: $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

Marking criteria:

- Correct reactants (MnO₄⁻) and products (Mn²⁺) (1 mark)
- Correct electron balance (5e⁻) (1 mark)
- Correct hydrogen ion balance (8H⁺) (1 mark)
- Correct water balance (4H₂O) (1 mark)

Part 3: Balanced half-reaction for Iron(II) oxidation (2 marks)

Correct answer: $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$

Marking criteria:

- Correct reactants (Fe²⁺) and products (Fe³⁺) (1 mark)
- Correct electron balance (e⁻) (1 mark)

Part 4: Balanced overall equation (3 marks)

Correct answer: $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$

Marking criteria:

- Correct balancing of Mn and Fe (1 mark)
- Correct balancing of oxygen atoms using water molecules (1 mark)
- Correct balancing of hydrogen atoms using H⁺ (1 mark)

Part 5: Problem analysis and solutions (5 marks)

Possible reasons for persistent colouration (any two):

Incomplete oxidation of iron(II) to iron(III)

Explanation: Insufficient KMnO_4 dosage relative to Fe^{2+} concentration (1 mark)

Solution: Increase KMnO_4 dosage based on water testing to ensure complete oxidation (1 mark)

Formation of colloidal iron(III) hydroxide

Explanation: Fe^{3+} forms colloidal $\text{Fe}(\text{OH})_3$ that remains suspended in water (1 mark)

Solution: Add coagulants (e.g., aluminium sulphate) to promote flocculation and settling of iron hydroxide particles (1 mark)

Presence of organic matter complexing with iron

Explanation: Organic compounds form coloured complexes with iron that resist oxidation (1 mark)

Solution: Implement pre-treatment with activated carbon or enhanced coagulation to remove organic matter before KMnO_4 addition (1 mark)

Inadequate filtration after treatment

Explanation: Oxidised iron particles not effectively removed by filtration (1 mark)

Solution: Improve filtration systems or add sand filtration step after sedimentation (1 mark)

Presence of other metal ions

Explanation: Other transition metals (e.g., manganese) causing colouration (1 mark)

Solution: Implement specific treatment for other metals or adjust pH to optimise removal of multiple metal species (1 mark)

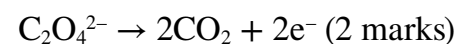
Note: Award marks for any two well-explained reasons with appropriate solutions (maximum 5 marks total for this section).

FOCAL AREA 3

Half-Equations for KMnO_4 and Sodium Ethanedioate (8 marks)

Half-Equations (4 marks)

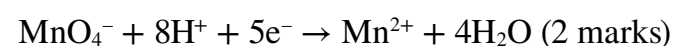
Correct oxidation half-equation:



1 mark for correct reactant and product

1 mark for correct electron balance

Correct reduction half-equation:



1 mark for correct reactants and products

1 mark for correct electron and hydrogen balance

Explanation of Stoichiometric Ratio (4 marks)

Identification that the oxidation half-equation produces 2 electrons per ethanedioate ion (1 mark)

Identification that the reduction half-equation requires 5 electrons per permanganate ion (1 mark)

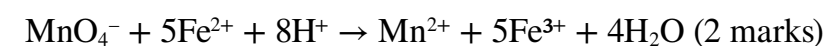
Explanation that the lowest common multiple of 2 and 5 is 10, requiring 5 ethanedioate ions to provide 10 electrons and 2 permanganate ions to accept 10 electrons (1 mark)

Clear conclusion that the stoichiometric ratio of MnO_4^- to $\text{C}_2\text{O}_4^{2-}$ is 2:5 (1 mark)

Iron(II) Concentration Calculation (10 marks)

Balanced Equation (2 marks)

Correct balanced equation:



1 mark for correct reactants and products

1 mark for correct balancing

Calculation Steps (6 marks)

Correct calculation of moles of KMnO_4 used (2 marks)

$$n(\text{KMnO}_4) = c \times V$$

$$= 0.020 \text{ mol/dm}^3 \times \frac{20.5}{1000} \text{ dm}^3$$

$$= 4.1 \times 10^{-4} \text{ mol}$$

Correct use of the stoichiometric ratio from the balanced equation (2 marks)

$$n(\text{Fe}^{2+}) = 5 \times n(\text{KMnO}_4)$$

$$= 5 \times 4.1 \times 10^{-4} \text{ mol}$$

$$= 2.05 \times 10^{-3} \text{ mol}$$

Correct calculation of Fe^{2+} concentration (2 marks)

$$C(\text{Fe}^{2+}) = \frac{n}{V}$$

$$= \frac{2.05 \times 10^{-3} \text{ mol}}{0.025 \text{ dm}^3}$$

$$= 0.082 \text{ mol/dm}^3$$

Explanation of Steps (2 marks)

Clear explanation of how the mole ratio from the balanced equation is applied (1 mark)

Logical explanation of the unit conversions and final concentration calculation (1 mark)

Starch Indicator in Iodine Titration (6 marks)

Explanation (6 marks)

Explanation that starch forms a deep blue-black complex with iodine (1 mark)

Identification that at high iodine concentrations at the beginning of the titration, the starch-iodine complex can be irreversibly adsorbed onto the starch molecules (1 mark)

Explanation that this adsorption can lead to delayed or indistinct endpoint (1 mark)

Explanation that adding starch when the solution is pale yellow ensures there is just enough iodine to form the complex but not enough to cause adsorption problems (1 mark)

Description of how this practice improves endpoint detection precision (1 mark)

Reference to the fact that the pale yellow colour indicates that most of the iodine has been reduced but enough remains to react with starch for a visible endpoint (1 mark)

Analysis of Contaminated Glassware (11 marks)

Identification of the Problem (3 marks)

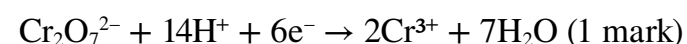
Recognition that residual potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$) remained in the glassware (1 mark)

Explanation that potassium dichromate is a strong oxidising agent (1 mark)

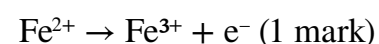
Identification that dichromate ions can oxidise Fe^{2+} to Fe^{3+} (1 mark)

Redox Chemistry Explanation (4 marks)

Correct half-equation for dichromate reduction:



Correct half-equation for iron oxidation:



Explanation that each dichromate ion can oxidise 6 iron(II) ions (1 mark)

Explanation that this pre-oxidation reduces the amount of Fe^{2+} available to react with KMnO_4 during the titration, leading to a lower titre value and thus a calculated concentration higher than the actual value (1 mark)

Laboratory Practice Recommendations (4 marks)

Recommendation to clean glassware with an appropriate cleaning solution (e.g., dilute acid followed by thorough rinsing) (1 mark)

Recommendation to use separate, dedicated glassware for different types of titrations or analyses (1 mark)

Suggestion to rinse glassware with a small portion of the solution to be used in it before the actual experiment (1 mark)

Recommendation to verify cleanliness of glassware through appropriate quality control procedures (e.g., blank titrations) (1 mark)

WEEK 13 - 14

FOCAL AREA 1

a. Q, Y, Z, X

Correct arrangement of species in order of decreasing reactivity (2 marks)

Reason

From the standard electrode potential values given the reactivity of the species decreases in that order because the more negative E° values indicate stronger reducing agents (more reactive metals). Q has the most negative E° value (-2.72V), making it the most reactive, followed by Y with E° value of -2.10V. Z and X have positive E° values, making them less reactive, with Z ($E^\circ = +0.40\text{V}$) being more reactive than X ($E^\circ = +1.10\text{V}$). (2 marks)

Construction of electrochemical cell

Drawing of electrochemical cell.

Correct labelling of anode (Y electrode) and cathode (Z electrode) (1 mark)

Correct representation of the electrodes in 1.0 mol/dm^3 solution of their respective ions. (1 mark)

Inclusion of salt bridge with correct orientation (1 mark)

Correct indication of electron flow in the external circuit (from Y to Z electrode) (1 mark)

Writing the Overall Cell Reaction

Correct half-equation at anode: $\text{Y} \rightarrow \text{Y}^{2+} + 2\text{e}^-$ (1 mark)

Correct half-equation at cathode: $2\text{Z}^+ + 2\text{e}^- \rightarrow 2\text{Z}$ (1 mark)

Correct balanced overall equation: $\text{Y} + 2\text{Z}^+ \rightarrow \text{Y}^{2+} + 2\text{Z}$ (1 mark)

Calculation of emf of the cell

Recognition that $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ (1 mark)

Correct substitution: $E^\circ_{\text{cell}} = +0.40\text{V} - (-2.10\text{V})$ (1 mark)

Correct calculation: $E^\circ_{\text{cell}} = +2.50\text{V}$ (2 marks)

wrong unit, no score

a. Correct identification of half-reactions (1 mark)

Calculation: $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 0.77 - (-0.91) = 1.68 \text{ V}$ (1 mark)

Cell notation: $\text{Cr}|\text{Cr}^{2+}||\text{Fe}^{3+}|\text{Fe}^{2+}|\text{Pt}$ (1 mark)

Conclusion that reaction is feasible as E°_{cell} is positive (1 mark)

b. i. $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 1.36 - 0.77 = 0.59 \text{ V}$ (1 mark)

ii. $\Delta G^\circ = -nFE^\circ = -2 \times 96,500 \times 0.59 = -113,870 \text{ J/mol} = -113.87 \text{ kJ/mol}$ (2 marks)

iii. $\Delta G^\circ = -RT\ln K$, therefore $\ln K = \Delta G^\circ/(-RT)$ (1 mark) $K = e^{(113,870/(8.314 \times 298))}$ (1 mark) $K = 1.26 \times 10^{20}$ (1 mark)

a. Identification of essential components: two electrodes, electrolyte solutions, salt bridge/porous barrier, external circuit (2 marks)

Explanation that voltaic cells convert chemical energy to electrical energy through spontaneous redox reactions (1 mark)

Explanation of electron flow from anode (oxidation) to cathode (reduction) (1 mark)

b. Correct identification that Zn and Cu would produce highest potential (1 mark)

Calculation: $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 0.34 - (-0.76) = 1.10 \text{ V}$ (1 mark)
Half-equation at Zn electrode: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ (oxidation) (1 mark)
Half-equation at Cu electrode: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ (reduction) (1 mark)
Overall cell reaction: $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$ (1 mark)

- c. i. Temperature analysis: Explanation that increased temperature generally increases cell potential by increasing reaction kinetics; discussion of equilibrium effects on cell potential (2 marks)
- ii. Surface area analysis: Explanation that surface area affects reaction rate but not the standard potential; discussion of current vs. potential distinction (2 marks)
- iii. Porous barrier analysis: Explanation of how ion migration might be affected, potential for increased internal resistance, and impact on cell performance over time (2 marks)

FOCAL AREA 2

Mark scheme

Design an electrochemical system to recover gold

Anode: Zinc metal in the solution of its ions

Cathode: Carbon electrode in a solution containing gold ions in solution.

Calculation of standard cell potential

Half-reactions:

Cathode (reduction): $\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au(s)}$ $E^\circ = +1.50 \text{ V}$

Anode (oxidation): $\text{Zn(s)} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ $E^\circ = -0.76 \text{ V}$ (as a reduction potential)

Calculating the cell potential correctly: $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = +1.50 \text{ V} - (-0.76 \text{ V}) = +2.26 \text{ V}$

Calculation of Gibbs free energy

For the balanced reaction: $2\text{Au}^{3+} + 3\text{Zn(s)} \rightarrow 2\text{Au(s)} + 3\text{Zn}^{2+}$

Number of electrons transferred: $n = 6$ (2 gold atoms $\times$ 3e^- per atom) Faraday constant: $F = 96,500 \text{ C/mol}$

$\Delta G^\circ = -nFE^\circ_{\text{cell}}$ $\Delta G^\circ = -(6 \text{ mol e}^-)(96,500 \text{ C/mol})(2.26 \text{ V})$ $\Delta G^\circ = -1,307,940 \text{ J/mol} = -1,308 \text{ kJ/mol}$

This large negative ΔG° confirms the process is highly spontaneous and thermodynamically favourable.

Environmental Sustainability Advantages

This electrochemical approach offers significant environmental benefits compared to conventional cyanide and mercury extraction methods:

Elimination of toxic chemicals: No cyanide or mercury is required, eliminating the risk of these toxins entering water systems and soil

Closed-loop system: The process can be designed as a contained system with minimal emissions or effluents

Selective recovery: Electrodeposition is highly selective for gold due to its reduction potential, reducing the need for additional purification steps

Energy efficiency: The spontaneous galvanic reaction requires minimal external energy input compared to energy-intensive conventional processes

Recyclable materials: The zinc anode can be regenerated through electrolysis using renewable energy, creating a sustainable cycle

Complete Electrochemical Analysis

Half-Reactions at each electrode

Cathode (reduction): $\text{Au}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Au(s)}$

Anode (oxidation): $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$

Overall balanced reaction with states of matter:

$2\text{Au}^{3+}(\text{aq}) + 3\text{Zn(s)} \rightarrow 2\text{Au(s)} + 3\text{Zn}^{2+}(\text{aq})$

Cell diagram

$\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) \parallel \text{Au}^{3+}(\text{aq}) \mid \text{Au(s)}, \text{C}$

WEEK 15 – 16

FOCAL AREA 1

Products of electrolysis with inert electrodes (4 marks)

Electrolyte	Product at Anode	Product at Cathode
Molten NaCl	Chlorine gas (Cl_2)	Sodium metal (Na)
Dilute H_2SO_4	Oxygen gas (O_2)	Hydrogen gas (H_2)

Award 1 mark for each correct product (4 marks total)

Electrolysis of dilute sodium sulphate solution (4 marks)

Half-equations (2 marks):

Anode: $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$ (1 mark)

Cathode: $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$ (1 mark)

Explanation for pH changes (2 marks):

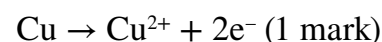
At the anode, water is oxidised to produce oxygen gas and hydrogen ions (H^+), causing the solution to become acidic (1 mark)

At the cathode, water is reduced to produce hydrogen gas and hydroxide ions (OH^-), causing the solution to become alkaline (1 mark)

Investigation of copper(II) sulphate electrolysis (6 marks)

Explanation for Cell A observations (2 marks):

At the anode, copper metal is oxidised to copper(II) ions:



At the cathode, copper(II) ions are reduced to copper metal:

$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$. This results in copper transferring from anode to cathode with no net change in solution concentration or colour (1 mark)

Explanation for Cell B differences (2 marks):

At the carbon anode, water is oxidised instead of copper:



With copper still being deposited at the cathode but no copper entering the solution at the anode, the copper(II) ion concentration decreases, causing the solution to become paler (1 mark)

Cell selection for determining electrochemical equivalent (2 marks):

Cell B would be more appropriate (1 mark)

Justification: In Cell B, the mass of copper deposited at the cathode directly corresponds to the amount of charge passed through the circuit without being affected by simultaneous anode dissolution. This allows for a direct measurement of the relationship between charge and mass of copper deposited. (1 mark)

Investigation design (6 marks)

Award marks based on the sophistication and completeness of the investigation design:

Identification of variables (1-2 marks):

Clear identification of at least three relevant variables (e.g., temperature, concentration, electrode material, current density, pH, electrode spacing)

Explanation of how each variable might affect electrolysis efficiency

Justification for selection of these variables based on electrochemical principles

Experimental methodology (1-2 marks):

Detailed description of specific experiments for each variable

Appropriate control of confounding variables

Clear explanation of how each experiment isolates the variable being tested

Consideration of replication and statistical validity

Measurement and calculation of efficiency (1 mark):

Precise definition of efficiency (e.g., current efficiency, energy efficiency)

Appropriate methods for measuring relevant quantities (mass, gas volume, energy input)

Clear calculation procedures with consideration of sources of error

Application and limitations (1 mark):

Thoughtful discussion of how findings could improve industrial processes

Realistic assessment of limitations in the investigation

Scientifically sound suggestions for improvements

Consideration of scale-up challenges

FOCAL AREA 2

Silver Plating Calculation (10 marks)

Step 1: Calculate the volume of silver needed (3 marks)

1 mark for correct formula: Volume = thickness × surface area

1 mark for correct conversion of thickness from mm to m

1 mark for correct calculation: Volume = $3.75 \times 10^{-6} \text{ m}^3 = 3.75 \text{ cm}^3$

Step 2: Calculate the mass of silver (2 marks)

1 mark for correct formula: Mass = volume × density

1 mark for correct calculation: Mass = 39.375 g

Step 3: Apply Faraday's law to calculate current (5 marks)

1 mark for correct formula rearrangement: $I = \frac{m \times n \times F}{M \times t}$

1 mark for correct identification of all variables

1 mark for correct conversion of time to seconds

1 mark for correct substitution of values

1 mark for correct final answer with units: 3.26 amperes

Cathodic Protection (10 marks)

Comparison of galvanic and electrolytic cells (3 marks)

1 mark for correct description of galvanic cells

1 mark for correct description of electrolytic cells

1 mark for relevant comparison in the context of corrosion prevention in coastal Ghana

Electrochemical principles of cathodic protection (3 marks)

1 mark for explaining corrosion as an electrochemical process

1 mark for explaining how cathodic protection prevents oxidation of the protected metal

1 mark for describing electron flow in the protection system

Analysis of suitable sacrificial anodes (4 marks)

1 mark for identifying at least two suitable metals with correct reasoning

1 mark for explaining the importance of electrode potential in selecting anodes

1 mark for considering Ghana's coastal environment (saltwater, temperature, etc.)

1 mark for justifying the final selection with reference to cost, availability, or effectiveness.

WEEK 17

Educational poster on the impacts of corrosion

Examples and Identification (6 marks)

Three clear examples of corrosion with visuals (3 marks)

Correct identification of corrosion type for each example (3 marks)

Impact Analysis (6 marks)

Two relevant impacts for each example (3 marks)

Explanation of why each impact matters to the community (3 marks)

Appropriate prevention method for each example (3 marks)

Presentation and Communication (2 marks)

Clear organisation and visual appeal (1 mark)

Correct use of scientific terminology (1 mark)

Experimental Report (15 marks)

Hypothesis and Planning (3 marks)

Clear, testable hypothesis (1 mark)

Appropriate materials list (1 mark)

Logical procedure with controlled variables (1 mark)

Data Collection and Presentation (5 marks)

Complete data table with appropriate measurements (2 marks)

Accurate graph with labelled axes (2 marks)

Observations of visual changes (1 mark)

Analysis and Conclusion (7 marks)

Correct interpretation of results (2 marks)

Scientific explanation of why salt affects rusting (2 marks)

Connection to electrochemical process (1 mark)

Identification of one possible source of error (1 mark)

Real-world application example (1 mark)

SECTION 5: PHYSICAL AND CHEMICAL PROPERTIES OF THE TRANSITION ELEMENTS**STRAND: SYSTEMATIC CHEMISTRY OF THE ELEMENTS****Sub-Strand: Periodicity****Learning Outcome:** *Predict and describe the physical and chemical properties of the transition elements as well as their uses in everyday life***Content Standard:** Demonstrate knowledge and understanding of the pattern of physical and chemical properties of the transition elements as well as their uses. Recognise the colour production, catalytic properties and many applications of transition elements**INTRODUCTION AND SECTION SUMMARY**

This section introduces learners to the fascinating world of transition elements, exploring their unique physical and chemical properties that distinguish them from main group elements. Learners will investigate the characteristic behaviours of transition metals, including their variable oxidation states, coloured compounds, catalytic properties, and complex formation. By the end of this section, learners will be able to identify transition elements on the periodic table, explain their distinctive properties through practical experiments, understand the formation and naming of complex compounds, compare transition metals with main group metals, and appreciate their widespread applications in everyday life. This knowledge forms an essential foundation for understanding advanced inorganic chemistry, catalysis, coordination chemistry, and the technological applications that make transition metals indispensable in modern society, from construction and electronics to medicine and environmental remediation.

The weeks covered by this section are:

- Weeks 18 and 19: State and describe properties of transition elements and perform experiment to investigate some properties of transition elements
- Week 20: Outline the similarities and differences between transition metals and main group metals and everyday uses of transition metals

SUMMARY OF PEDAGOGICAL EXEMPLARS

This section encourages learners to actively explore the physical and chemical properties of transition elements through hands-on, collaborative, and inquiry-based learning. Mixed-ability groups should engage in discussions to identify transition elements, describe their characteristics—such as variable oxidation states, paramagnetism, complex ion formation, coloured compounds, and catalytic abilities—and compare these with main group metals.

Inquiry-based and exploratory activities such as the catalytic decomposition of hydrogen peroxide using manganese(IV) oxide and the formation of soluble Cu^{2+} complexes in dilute

ammonia help learners observe and explain abstract concepts. Learners should model and draw the shapes of complex compounds using everyday materials and apply IUPAC rules to name them.

Talk-for-learning strategies should allow all learners to describe transition metal properties and discuss their practical uses, such as in catalysts, pigments, and electronics. Approaching proficiency learners may explain findings orally or use diagrams, while high-performing learners should be challenged to apply knowledge to real-world contexts and complex naming tasks.

This learner-centred approach ensures participation from all learners, promotes deep understanding of periodicity, and connects chemistry to everyday life through relevant, practical experiences.

ASSESSMENT SUMMARY

Indicator	DOK Level	Assessment Strategy and/or Mode
State and describe properties of transition elements and perform experiment to investigate some properties of transition elements	Level 3: Strategic Reasoning	Group work
Outline the similarities and differences between transition metals and main group metals and everyday uses of transition metals	Level 3: Strategic Reasoning	Group work

WEEK 18 – 19

Learning Indicator: State and describe properties of transition elements and perform experiment to investigate some properties of transition elements

Recap weeks

Year 1	Week 4: State the main postulates of Bohr's planetary theory and explain the importance of the quantum numbers to the electron structure of the atom Week 17: Use the electron configuration of elements to determine their position on the periodic table.
Year 2	Week 18: Predict the shape and bond angles for species and distinguish between sigma and pi bonds. Week 19: Explain hybridisation and relate it to various structures

FOCAL AREA 1: PROPERTIES OF 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. They occupy the d-block (Groups 3-12) of the periodic table. While zinc, cadmium, and mercury (Group 12) have a full d¹⁰ configuration and technically should be considered post-transition metals, they are traditionally included in discussions of transition elements.

The general electronic configuration is [noble gas]d¹⁻¹⁰ns¹⁻², with some exceptions like Cr [Ar]3d⁵4s¹ and Cu [Ar]3d¹⁰4s¹ due to the stability of half-filled and completely filled d-orbitals.

Physical Properties

- 1. Metallic Character:** All transition elements exhibit typical metallic properties. They are good conductors of heat and electricity due to the presence of free electrons. Their metallic character is less pronounced than alkali metals but more than non-metals.
- 2. Melting and Boiling Points:** Transition metals generally have high melting and boiling points due to strong metallic bonding. The strong bonding results from the involvement of d-electrons in metallic bonding in addition to the s-electrons. Tungsten has the highest melting point (3422°C) among all metals. The melting points generally increase to a maximum in the middle of the series (Group 6) and then decrease.
- 3. Density:** Transition metals typically have high densities due to their small atomic radii and close packing of atoms in their crystal lattices. Osmium (22.6 g/cm³) and iridium (22.4 g/cm³) are the densest elements. Density generally increases across the period to a maximum and then decreases.
- 4. Hardness and Strength:** Most transition metals are hard and strong, making them useful for structural applications. Chromium is exceptionally hard and is used for hardening steel. The hardness is due to the strong metallic bonds formed by the d-electrons.
- 5. Malleability and Ductility:** Despite their hardness, transition metals are generally malleable (can be hammered into sheets) and ductile (can be drawn into wires). This property arises from the non-directional nature of metallic bonding.

- Magnetic Properties:** Many transition metals exhibit magnetic properties due to unpaired electrons in their d-orbitals. Iron, cobalt, and nickel are ferromagnetic (strongly attracted to magnets). Manganese and chromium are paramagnetic (weakly attracted to magnets). Zinc, with a full d-orbital, is diamagnetic (slightly repelled by magnets).
- Alloy Formation:** Transition metals readily form alloys with other metals. This property is extensively used in manufacturing steel (iron with carbon and other elements), brass (copper and zinc), bronze (copper and tin), and various specialised alloys for specific applications.

Occurrence

Transition metals occur in various forms in nature. Some less reactive metals like gold and platinum are found in their native state. Most transition metals occur as oxides (Fe_2O_3 , MnO_2) or sulfides (CuFeS_2 , ZnS). Ghana has significant deposits of gold and manganese, which contribute to its mining industry.

Characteristic properties of transition elements

1. Variable oxidation states

Transition metals exhibit multiple oxidation states due to the availability of both (n-1)d and ns electrons for bonding. This property distinguishes them from main group elements, which typically show fewer oxidation states.

The energy difference between (n-1)d and ns orbitals is small, allowing electrons from both to participate in bonding. The d-orbitals are incompletely filled, making electrons available for bonding. The promotion energy required to involve additional electrons in bonding is relatively low.

Titanium shows oxidation states of +2 (TiO), +3 (Ti_2O_3), and +4 (TiO_2). Vanadium exhibits oxidation states of +2 (VO), +3 (V_2O_3), +4 (VO_2), and +5 (V_2O_5).

Chromium demonstrates oxidation states of +2 (CrCl_2), +3 (Cr_2O_3), and +6 ($\text{K}_2\text{Cr}_2\text{O}_7$, CrO_3).

Manganese displays an impressive range with +2 (MnCl_2), +3 (Mn_2O_3), +4 (MnO_2), +6 (K_2MnO_4), and +7 (KMnO_4).

Iron commonly shows +2 (FeCl_2) and +3 (FeCl_3) states.

Copper exhibits +1 (Cu_2O) and +2 (CuO) oxidation states.

Applications

Variable oxidation states are utilised in redox titrations (KMnO_4 , $\text{K}_2\text{Cr}_2\text{O}_7$), catalysis in industrial processes, electrochemical cells and batteries, and in creating pigments and dyes with distinctive colours.

2. Paramagnetism

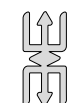
Paramagnetism is the property of materials to be weakly attracted to magnetic fields. In transition metals, paramagnetism arises from the presence of unpaired electrons in the d-orbitals. Unlike ferromagnetic materials (like iron) that retain magnetisation after the external field is removed, paramagnetic materials only exhibit magnetic behaviour when placed in an external magnetic field.

The paramagnetic character of transition metals is directly related to their electronic configuration. According to Hund's rule, electrons occupy orbitals singly with parallel spins before pairing occurs. Since transition metals have partially filled d-orbitals, many of them contain unpaired electrons. Each unpaired electron has a magnetic moment due to its spin, contributing to the overall paramagnetic behaviour of the atom or ion.

The degree of paramagnetism varies across the transition series based on the number of unpaired electrons, e.g.

- Scandium (Sc^{3+}): $[\text{Ar}]3d^0$ - No unpaired electrons, diamagnetic
- Titanium (Ti^{3+}): $[\text{Ar}]3d^1$ - One unpaired electron, weakly paramagnetic
- Vanadium (V^{3+}): $[\text{Ar}]3d^2$ - Two unpaired electrons, paramagnetic
- Chromium (Cr^{3+}): $[\text{Ar}]3d^3$ - Three unpaired electrons, strongly paramagnetic
- Manganese (Mn^{2+}): $[\text{Ar}]3d^5$ - Five unpaired electrons, maximum paramagnetism
- Iron (Fe^{3+}): $[\text{Ar}]3d^5$ - Five unpaired electrons, strongly paramagnetic
- Iron (Fe^{2+}): $[\text{Ar}]3d^6$ - Four unpaired electrons, paramagnetic
- Cobalt (Co^{2+}): $[\text{Ar}]3d^7$ - Three unpaired electrons, paramagnetic
- Nickel (Ni^{2+}): $[\text{Ar}]3d^8$ - Two unpaired electrons, paramagnetic
- Copper (Cu^{2+}): $[\text{Ar}]3d^9$ - One unpaired electron, weakly paramagnetic
- Zinc (Zn^{2+}): $[\text{Ar}]3d^{10}$ - No unpaired electrons, diamagnetic

Manganese(II) with five unpaired electrons shows the highest paramagnetic character among the first-row transition metal ions.



Note

Diamagnetic materials are repelled by magnetic fields. This property arises when all electrons in an atom or molecule are paired, resulting in no net magnetic moment.

Ferromagnetic materials exhibit strong attraction to magnetic fields and can retain magnetisation after the external field is removed. This occurs because the magnetic moments of atoms align parallel to each other within microscopic regions called domains.

3. Formation of coloured compounds

Most transition metal compounds are coloured due to d – d electron transitions when absorbing visible light. This property contrasts with the generally colourless or white compounds of main group elements.

When light strikes a transition metal compound, electrons jump between d-orbitals of different energies, absorbing specific wavelengths and transmitting/reflecting complementary colours. The energy difference between d-orbitals determines which wavelengths are absorbed. In transition metal ions, the five d-orbitals are split into different energy levels when surrounded by ligands. The energy gap between these levels corresponds to wavelengths in the visible spectrum. When light passes through the compound, wavelengths matching this energy gap are absorbed. The observed colour is the complementary colour to the absorbed wavelength. For example, if blue-green light is absorbed, the compound appears red. Examples:

- Scandium compounds (Sc^{3+}) are colourless due to d^0 configuration, with no d-d transitions possible.
- Titanium compounds (Ti^{3+}) appear purple due to d^1 configuration. Vanadium compounds (V^{3+}) show green colour with d^2 configuration.
- Chromium compounds (Cr^{3+}) can be green or violet with d^3 configuration - $[\text{Cr}(\text{H}_2\text{O})_6]^{3+}$ is violet, while $[\text{Cr}(\text{H}_2\text{O})_5\text{Cl}]^{2+}$ is green.
- Manganese compounds (Mn^{2+}) are pale pink with d^5 configuration.
- Iron(II) compounds (Fe^{2+}) appear pale green with d^6 configuration, while iron(III) compounds (Fe^{3+}) are yellow-brown with d^5 configuration.
- Cobalt compounds (Co^{2+}) can be pink or blue with d^7 configuration - $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$, etc.

Applications

Coloured compounds are used for metal ion identification in qualitative analysis, as pigments and dyes (Cr_2O_3 for green, Fe_2O_3 for red), in gemstones (ruby: Cr^{3+} in Al_2O_3 , emerald: Cr^{3+} in beryl), as indicators in titrations (KMnO_4), and in colorimetric analysis for metal determination.

4. Formation of complex compounds

Transition metals readily form coordination complexes with ligands (electron pair donors). This property is much more pronounced in transition metals than in main group elements.

Small, highly charged metal ions attract electron pair donors. Vacant d-orbitals are available for accepting electron pairs. Transition metals have the ability to accommodate various coordination numbers.

a. Coordination compound/complex

A coordination compound consists of a central metal atom/ion bonded to one or more ligands (atoms ions or molecules). The bond formed is a coordinate covalent bond (both electrons provided by ligand).

b. Coordination number

Coordination number refers to the number of ligands (atoms, ions or molecules) directly bonded to a central metal atom or ion in a coordination compound or complex. Common coordination numbers are 2, 4, and 6. The arrangement of ligands around the metal creates specific geometries.

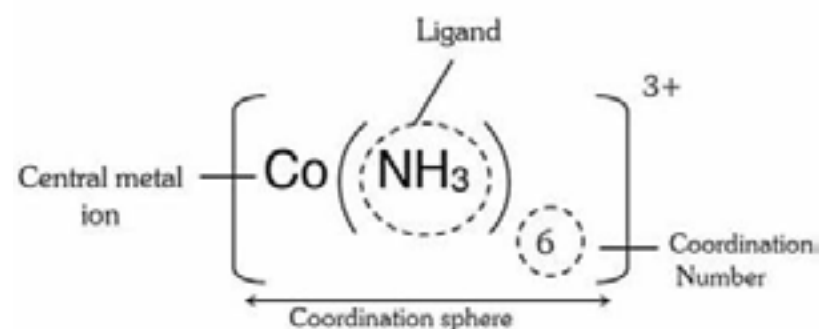


Figure 18.1: Coordination compound

c. Relationship between coordination number and geometry

Coordination number and geometry are closely related:

- A coordination number of 2 typically results in linear geometry e.g., $[\text{Ag}(\text{NH}_3)_2]^+$.
- A coordination number of 4 can lead to tetrahedral e.g., $[\text{Zn}(\text{NH}_3)_4]^{2+}$ or square planar e.g., $[\text{Ni}(\text{CN})_4]^{2-}$ geometry.
- A coordination number of 6 usually results in octahedral geometry e.g., $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$.

Applications

Complex formation is utilised in qualitative analysis for metal ion identification, catalysis in industrial and biological systems, biological systems (haemoglobin, chlorophyll, vitamin B_{12}), industrial processes (metal extraction, electroplating), and medicine (cisplatin for cancer treatment).

5. Catalytic Properties

Transition metals and their compounds are excellent catalysts due to their ability to change oxidation states and form temporary bonds with reactants. This property is widely exploited in industrial processes.

Transition metals have the ability to adsorb reactants on their surface, activating them. They can form intermediate compounds with lower activation energy. Transition state stabilisation occurs through d-orbital interactions.

Advantages of use of transition metal catalyst

- Transition metal catalysts increase reaction rates at lower temperatures.
- They improve selectivity for desired products.
- They allow reactions to occur under milder conditions.
- They can be recovered and reused, especially heterogeneous catalysts.
- They reduce energy requirements for industrial processes.

Applications

- Iron catalyses the Haber process $\text{N}_2 + \text{H}_2 \rightleftharpoons \text{NH}_3$
- Nickel is used for hydrogenation of unsaturated fats.
- Vanadium(V) oxide catalyses the contact process ($\text{SO}_2 \rightleftharpoons \text{SO}_3$)
- Platinum and rhodium are used in catalytic converters to reduce pollutants, etc.

Experiments to show catalytic behaviour

Decomposition of Hydrogen peroxide using MnO_2

Materials

Hydrogen peroxide solution

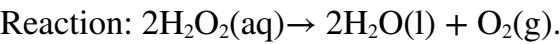
Manganese(IV) oxide powder,

Test tubes

Spatulas
Wooden splints,
Safety goggles and gloves for protection.

Procedure

- a. Place 5 cm³ hydrogen peroxide in a test tube and observe the slow decomposition (few bubbles).
- b. Then add a half spatula of MnO₂ powder and observe rapid oxygen evolution.
- c. Finally, test the gas with a glowing splint to confirm it is oxygen.



Explanation: MnO₂ catalyses the decomposition by providing a surface for the reaction and by temporarily changing oxidation states. The MnO₂ accepts electrons from H₂O₂, forming an intermediate that then decomposes, regenerating the MnO₂. This lowers the activation energy for the decomposition reaction.

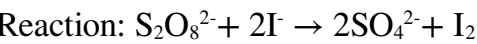
Iron(III) catalysed reaction between KI and Na₂S₂O₈

Materials

Potassium iodide solution,
Sodium persulphate solution,
Iron (III) chloride solution,
Starch solution,
Test tubes, and a stopwatch.

Procedure

- a) Begin by mixing equal volumes of KI and Na₂S₂O₈ solutions with starch.
- b) Observe the slow development of blue colour (I₂-starch complex).
- c) Repeat the experiment with a few drops of FeCl₃ solution added.
- d) Compare the reaction rates using the stopwatch.



Explanation: Iron(III) ions catalyse the reaction through a two-step process.



The iron cycles between Fe³⁺ and Fe²⁺ oxidation states, providing a lower energy pathway for the reaction. This demonstrates the importance of variable oxidation states in transition metal catalysis.

Learning Tasks

The following table shows some transition elements with their electronic configurations:

Element	Symbol	Atomic number	Ground state electron configuration
Titanium	Ti	22	[Ar]3d ² 4s ²
Vanadium	V	23	[Ar]3d ³ 4s ²
Chromium	Cr	24	[Ar]3d ⁵ 4s ¹
Manganese	Mn	25	[Ar]3d ⁵ 4s ²
Iron	Fe	26	[Ar]3d ⁶ 4s ²
Cobalt	Co	27	[Ar]3d ⁷ 4s ²

- 1. Analyse the electronic configuration of chromium and explain why it differs from the expected [Ar]3d⁴4s² configuration. Extend your analysis to explain any other anomalies in the electronic configurations of first-row transition elements.
- 2. For each element in the table, determine:
 - i. The maximum possible oxidation state
 - ii. The most stable oxidation state(s)
 - iii. The number of unpaired electrons in each common oxidation stateWhether each oxidation state would be paramagnetic or diamagnetic

Pedagogical Exemplars

Talk-for-Learning Approach: Understanding transition elements

- a. Form mixed-ability groups of 4-5 learners
- b. Task groups to identify first-row transition elements (Sc to Zn)
- c. Have learners develop a definition of transition elements
- d. Assign each group one property to investigate in depth:
 - i. Variable oxidation states
 - ii. Paramagnetism
 - iii. Complex ion formation
 - iv. Coloured compounds formation
 - v. Catalytic abilities.
- e. Reorganise into jigsaw groups where learners teach their property to others
- f. Conduct whole-class discussion to consolidate understanding and the key discussion points should include:
 - i. Location of transition elements in the periodic table
 - ii. Electronic configuration and d-orbitals
 - iii. How each property relates to electronic structure
 - iv. Real-world applications in Ghana (mining, manufacturing).

Inquiry-Based Activity: Catalytic Behaviour Experiments

Guide learners through the following procedure:

- a. Place 5 cm³ of hydrogen peroxide solution in a test tube
- b. Observe and record any changes occurring (slow bubbling)
- c. In a second test tube, place another 5 cm³ of hydrogen peroxide solution
- d. Add a small amount of manganese(IV) oxide powder
- e. Immediately observe and record the changes (rapid bubbling)
- f. Light a wooden splint, blow out the flame leaving it glowing
- g. Insert the glowing splint into the test tube with the rapid reaction
- h. Record observations (splint should rekindle)
- i. Measure and compare reaction rates with and without catalyst
- j. Ask learners to record:
 - i. The speed of gas evolution with and without MnO₂
 - ii. The test for oxygen gas (glowing splint test)
 - iii. The chemical equation for the reaction.

Key Assessment Level 3: Group Work

A learner is investigating the catalytic properties of transition metals in the decomposition of hydrogen peroxide. The learner sets up an experiment using three different catalysts: manganese(IV) oxide, iron(III) chloride, and copper(II) sulphate. The following data is collected for the volume of oxygen produced over time:

Time (seconds)	Volume of O ₂ with MnO ₂ (cm ³)	Volume of O ₂ with FeCl ₃ (cm ³)	Volume of O ₂ with CuSO ₄ (cm ³)
0	0	0	0
30	42	15	28
60	78	29	52
90	98	42	70
120	110	53	85
150	116	63	95
180	118	72	102

- 1. Construct a graph to represent this data and use it to analyse the relative effectiveness of these three transition metal compounds as catalysts.
- 2. Based on your understanding of transition metal properties, propose a scientific explanation for the differences in catalytic activity observed.

- 3. Design a follow-up experiment to investigate whether the catalytic activity is related to the variable oxidation states of these transition metals. Include:
 - a. Your hypothesis
 - b. Key variables to control
 - c. Experimental procedure
 - d. Expected results if your hypothesis is correct

FOCAL AREA 2: COMPLEXES OF TRANSITION ELEMENTS

Experiment on applications of complexes (Copper complex formation)

Materials

Copper (II) tetraoxosulphate(VI) solution (0.1 mol/dm³)
Ammonia solution (2 mol/dm³)
Hydrochloric acid (concentrated, 12 mol/dm³)
Sodium hydroxide solution (2 mol/dm³)
Test tubes (clean, borosilicate glass)
Dropper pipettes
Glass stirring rods
White background (paper or tile) for colour observation
Safety equipment (goggles, gloves, lab coat)

Procedure

- 1. Place 2 cm³ of copper (II) tetraoxosulphate(VI) solution in a clean test tube. Observe the pale blue colour of the hexaaquacopper (II) complex [Cu(H₂O)₆]²⁺. Record the exact shade of blue.
- 2. Add sodium hydroxide solution dropwise (approximately 10 drops), swirling after each addition, until a light blue precipitate forms completely. This is copper(II) hydroxide: Cu²⁺(aq) + 2OH⁻(aq) → Cu(OH)₂(s)
- 3. Continue adding sodium hydroxide in excess (about 5 more drops). Note any changes in the precipitate (it may darken slightly due to partial dehydration).
- 4. To this mixture, add ammonia solution dropwise while swirling, until the precipitate completely dissolves. Note the formation of an intensely deep blue solution of tetraamminediaquacopper(II) complex:
Cu(OH)₂(s) + 4NH₃(aq) → [Cu(NH₃)₄(H₂O)₂]²⁺(aq) + 2OH⁻(aq)
- 5. To a fresh 2 mL sample of copper(II) tetraoxosulphate(VI) solution in a new test tube, carefully add concentrated hydrochloric acid dropwise (about 15-20 drops) until a colour change to yellow-green is observed. This indicates the formation of tetrachlorocuprate(II) complex: [Cu(H₂O)₆]²⁺(aq) + 4Cl⁻(aq) → [CuCl₄]²⁻(aq) + 6H₂O(l)
- 6. Observations: The colour changes observed demonstrate the significant effect of different ligands on the d-orbital splitting in copper complexes. The pale blue hexaaqua complex

changes to deep blue with ammonia ligands because ammonia is a stronger field ligand than water, increasing the crystal field splitting energy.

Applications

These reactions form the basis for qualitative analysis of copper in unknown samples. The deep blue $[\text{Cu}(\text{NH}_3)_4]^{2+}$ complex is particularly distinctive and can detect copper at concentrations as low as 0.02 mg/dm^3 , making it valuable for environmental testing.

Shapes of some common complex compounds

The arrangement of ligands around a metal ion (coordination geometry) determines a complex's key properties. These structures linear, tetrahedral, square planar, or octahedral result from the metals electronic configuration and ligand interactions. Each geometry creates a unique environment affecting colour, magnetism, and reactivity. Understanding these patterns allows chemists to design compounds for specific applications in catalysis, medicine, and materials science.

1. Linear Complexes (Coordination number 2)

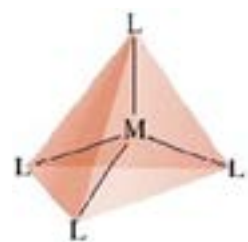
Linear complexes have a 180° bond angle between the ligands, creating a straight line arrangement around the central metal atom. This geometry results from the metal using sp hybridisation, with two hybrid orbitals oriented at 180° to each other.



Examples: $[\text{Ag}(\text{NH}_3)_2]^+$, $[\text{AuCl}_2]^-$, $[\text{Cu}(\text{CN})_2]^-$, etc.

2. Tetrahedral complexes (Coordination number 4)

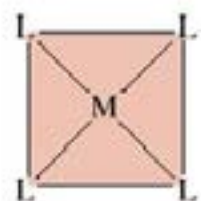
Tetrahedral complexes have four ligands arranged at the corners of a tetrahedron, with bond angles of 109.5° . This arrangement minimises ligand-ligand repulsions and results from sp^3 hybridisation of the metal orbitals.



Example: $[\text{Zn}(\text{NH}_3)_4]^{2+}$, $[\text{CoCl}_4]^{2-}$, $[\text{CuCl}_4]^{2-}$, $[\text{FeCl}_4]^-$, etc.

3. Square Planar Complexes (Coordination number 4)

Square planar complexes have four ligands arranged in a flat square around the central metal, with 90° bond angles. This geometry typically results from dsp^2 hybridisation and is favoured by certain d^8 metal ions.



Example: $[\text{Zn}(\text{NH}_3)_4]^{2+}$, $[\text{CoCl}_4]^{2-}$, $[\text{CuCl}_4]^{2-}$, $[\text{FeCl}_4]^-$, etc.

4. Octahedral Complexes (Coordination number 6)

Octahedral complexes have six ligands arranged at the corners of an octahedron, with all bond angles at 90° . This is the most common geometry for transition metal complexes and results from d^2sp^3 hybridisation.



Examples: $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$, $[\text{Co}(\text{NH}_3)_6]^{3+}$, $[\text{Cr}(\text{H}_2\text{O})_6]^{3+}$, $[\text{Fe}(\text{CN})_6]^{4-}$, etc.

Naming of Complex Compounds

Coordination compound nomenclature provides a precise system for describing complex metal-ligand structures. These IUPAC rules encode crucial information about the central metal, its oxidation state, and the identity, number, and arrangement of coordinating ligands. Unlike simple compound naming, this specialised system accounts for the unique features of coordination complexes, including different ligand types and possible isomers. Mastering this nomenclature is essential for clear communication in chemistry research, education, and industry, allowing chemists to systematically describe both existing compounds and newly synthesised complexes with clarity and precision.

1. Identify the Central Metal Atom and Its Oxidation State

- Determine the central metal atom in the complex
- Calculate its oxidation state by balancing the overall charge
(Metal oxidation state) + (Sum of ligand charges) = (Overall complex charge)
- The oxidation state will be written as a Roman numeral in parentheses.

2. Identify and Classify All Ligands

- List all atoms or groups directly bonded to the central metal
- Classify each ligand as:
 - Anionic (negatively charged): Cl^- , CN^- , OH^- , etc.
 - Neutral: H_2O , NH_3 , CO , etc.

3. Name the Ligands According to IUPAC Rules

- For anionic ligands, change the ending to “-o”, e.g.
Chloride $\rightarrow$ chloro,
Cyanide $\rightarrow$ cyano,
Hydroxide $\rightarrow$ hydroxo
- For neutral ligands, use specific names, e.g.
 $\text{H}_2\text{O} \rightarrow$ aqua,
 $\text{NH}_3 \rightarrow$ ammine (note the double “m”)

CO → carbonyl

For abbreviated ligands like ethylenediamine (en), use the full name in the complex.

4. Add Numerical Prefixes for Multiple Identical Ligands

- Use the prefixes di-, tri-, tetra-, penta-, hexa-, etc.
- For complex ligands that already contain these prefixes, use bis-, tris-, tetrakis-, etc.

5. Arrange Ligand Names in Alphabetical Order

Alphabetise based on the first letter of the ligand name, ignoring prefixes e.g. “diaqua” is alphabetised under “a” for aqua, not “d” for di.

6. Name the Metal

For cationic or neutral complexes, use the element name, for anionic complexes, change the ending to “-ate”, e.g. Iron (ferrate), Copper (cuprate), Cobalt (cobaltate), etc.

Some metals have special “-ate” names, e.g. Gold (aurate), Silver (argentite), Lead (plumbate).

7. Indicate the Oxidation State

Write the oxidation state as a Roman numeral in parentheses, e.g. Iron(III), Copper(II), except when the metal has only one common oxidation state (like Zn^{2+}), the number may be omitted.

8. Add Counter Ions (if present)

For complex salts, name the cation first, then the anion. Include numerical prefixes (di-, tri-, etc.) for multiple counter ions, e.g., potassium hexacyanoferrate(III).

Application of IUPAC rules in naming complexes

1. $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$

- Central metal: Cobalt with +3 oxidation state
- Ligands: Six NH_3 (ammine)

Name: Hexaamminecobalt(III) chloride.

2. $\text{K}_3[\text{Fe}(\text{CN})_6]$

Central metal: Iron with +3 oxidation state

Complex is anionic (ferrate)

Ligands: Six CN^- (cyano)

Counter ion: Three K^+ (potassium)

Name: Potassium hexacyanoferrate(III).

3. $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$

Central metal: Platinum with +2 oxidation state

Ligands: Two NH_3 (ammine) and two Cl^- (chloro)

Name: Diamminedichloroplatinum(II).

4. $[\text{Fe}(\text{CO})_5]$

Name : Pentacarbonyliron(0).

5. $\text{K}_4[\text{Fe}(\text{CN})_6]$

Name: Potassium hexacyanoferrate(II).

6. $[\text{Co}(\text{NH}_3)_4\text{Cl}_2]^+$

Name: Tetraamminedichlorocobalt(III) ion.

7. $[\text{Co}(\text{C}_2\text{O}_4)_2(\text{NH}_3)_2]^-$

Name: Diamminebis(oxalato)cobaltate(III) ion.

Pedagogical Exemplars

Exploratory Learning: Solubilising Insoluble Compounds

- Prepare solutions of copper(II)tetraoxosulphate(VI) , silver trioxonitrate(V), and zinc sulfate
- Form insoluble precipitates using appropriate reagents (NaOH or NaCl)
- Add excess ligand solution (ammonia or sodium hydroxide) drop-wise
- Observe dissolution of precipitates as complex ions form
- Record colour changes and other observations
- Write balanced equations for all reactions
- Discuss applications of complex formation in real-world contexts.

Collaborative Learning: Modelling Complex Geometries

- Form mixed-ability groups and assign different complex compounds
- Distribute modelling materials (commercial kits or alternative materials)
- Construct 3D models showing correct geometry:
 - Tetrahedral (109.5° bond angles)
 - Square planar (90° and 180° bond angles)
 - Octahedral (90° and 180° bond angles)
- Draw 2D representations of the 3D structures
- Name each complex using IUPAC nomenclature rules
- Present models to class, explaining geometry and naming
- Conduct a gallery walk for peer feedback and discussion
- Connect geometries to physical and chemical properties of complexes.

Key Assessment: Practice Paper 2

A learner performed an experiment with different metal salts and recorded the following observations:

Experiment	Initial solution	After adding aqueous NaOH	After adding excess aqueous NH ₃
1	Blue solution (Cu ²⁺)	Blue precipitate	Deep blue precipitate
2	Pale green solution (Ni ²⁺)	Green precipitate	Purple solution
3	Colourless solution (Zn ²⁺)	White precipitate	Colourless solution

Analyse these results and answer the following questions

- Write balanced chemical equations for all the reactions occurring in each experiment
- Explain why the precipitates dissolve when excess ammonia is added. Include the structure of one complex ion formed in your answer.
- The formula of the copper-ammonia complex is [Cu(NH₃)₄]²⁺. Draw its structure showing the correct geometry, and explain why this complex has a deeper blue colour than the original copper solution.

Hint

Conduct *Practice Paper 2*. It should cover contents taught from Year 1 to 3. Refer to the *Table of Specification and Structure of the Paper* at the end of this section. Learners' papers should be marked immediately and scores entered into the SAP.

TABLE OF SPECIFICATION FOR PAPER 2, 1 AND 3

Strands	Sub-strands	Learning Outcomes	Content Standards	Learning Indicators	Levels of Assessment SECTION A				Total	Levels of Assessment SECTION B				Total
					L1	L2	L3	L4		L1	L2	L3	L4	
Physical Chemistry														
	Matter and its Properties	6	4	30	1		1		2	2	3	2	1	7
	Equilibria	7	7	22	1	1			2	2	4	2		8

Systematic Chemistry of the Elements	Periodicity	4	4	8	1	1			2	2	2	1	1	5
	Bonding	3	3	7		1			1	1	2	1		4
Chemistry of Carbon Compounds	Characterization of Organic Compounds	2	2	3	1				1	1	1	1	1	3
	Organic Functional Groups	4	4	11	1	1			2	2	2	1		5
No of Qs					5	4	1		10	10	14	8	3	35
MARKS					10	12	3		25	19	41	25	15	100

Strands	Sub-strands	Learning Outcomes	Content Standards	Learning Indicators	Levels of Assessment				Total
Physical Chemistry					1	2	3	4	
	Matter and its Properties	6	4	30	8	7	4	2	21
	Equilibria	7	7	22	4	3	4		11
Systematic Chemistry of the Elements	Periodicity	4	4	8	1	3	1		5
	Bonding	3	3	7		2	1		3
Chemistry of Carbon Compounds	Characterization of Organic Compounds	2	2	3		2			2
	Organic Functional Groups	4	4	11	2	3	2	1	8
					15	20	12	3	50

Strands	Sub-strands	Learning Out-comes	Content Stand-ards	Learning Indica-tors	Levels of As-sessment				To-tal
					1	2	3	4	
Physical Chemistry	Matter and its Properties	6	4	30		1	2		
	Equilibria	7	7	22					
Systematic Chemistry of the Elements	Periodicity	4	4	8					
	Bonding	3	3	7					
Chemistry of Carbon Com-pounds	Characterization of Organic Compounds	2	2	3					
	Organic Functional Groups	4	4	11					
						1	2		50



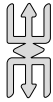
Note

Please use the PLC instruction template for PLC Sessions 1 for PLC 19, but choose an assessment strategy for that week.

PLC Session 18 & 19

45 minutes	Planning & Reviewing the Learning Plan
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1. Begin the session by recapping lessons from the previous session, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week’s concepts in a practical way that will make it easier for learners to understand.
 - b. identify three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q).
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



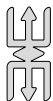
Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Use your app to plan and administer the mode of assessment for the Student Assessment Portal (SAP).
 - d. The key assessment for the week is **practice paper 2**. Chat with your app to plan this assessment by reviewing or developing the assessment items/tasks and marking schemes/rubrics for the practice paper (NTS 3k-3q).

10 minutes	Content Exploration
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5. As a subject group, select one problem or topic/idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
 - a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through the problem individually for 3–4 minutes
 - c. the group then works through it together the problem, step by step
 - d. highlight the most common errors made at each step (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
 - a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

- | | |
|-----------|------------|
| 5 minutes | Reflection |
|-----------|------------|
7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
 8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
 9. Remember to
 - a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

WEEK 20

Learning Indicator: Outline the similarities and differences between transition metals and main group metals and everyday uses of transition metals

Recap weeks

Year 1	Week 17: Use the electron configuration of elements to determine their position on the periodic table.
Year 2	Week 15: Describe the patterns in physical and chemical properties of the period 3 elements.

FOCAL AREA 1: SIMILARITIES AND DIFFERENCES BETWEEN TRANSITION METALS AND MAIN GROUP METALS

Similarities between Transition Metals and Main Group Metals

Despite occupying different regions of the periodic table and exhibiting distinctive chemical behaviours, transition metals and main group metals share fundamental characteristics:

1. Physical Properties
 - a. Metallic Lustre: Both transition and main group metals exhibit the characteristic shiny appearance of metals due to their ability to reflect light.
 - b. Electrical Conductivity: Both groups are good conductors of electricity due to their delocalised electrons that can move freely through the metallic lattice.
 - c. Thermal Conductivity: Both efficiently transfer heat energy through the movement of free electrons and vibrations in their crystal lattice.
 - d. Malleability and Ductility: Both can be hammered into sheets (malleability) and drawn into wires (ductility) due to the non-directional nature of metallic bonding.
 - e. Solid State: At room temperature, elements from both groups exist as solids (with the exception of mercury, a transition metal).
2. Chemical Properties
 - a. Metallic Character: Both groups show metallic behaviour, tending to lose electrons to form positive ions (cations).
 - b. Oxide Formation: Both form basic or amphoteric oxides (though transition metal oxides show more variation).
 - c. Reaction with Acids: Elements from both groups can react with acids to produce hydrogen gas (though with varying reactivity).
 - d. Thermal Stability: Compounds of both groups generally show reasonable thermal stability.

3. Electronic Structure

- a. **Metallic Bonding:** Both groups exhibit metallic bonding characterised by a “sea” of delocalised electrons.
- b. **Presence of Valence Electrons:** Both have valence electrons that can participate in bonding or be removed to form ions.

Differences between Transition Metals and Main Group Metals**1. Electronic Configuration**

- a. **Transition Metals:** Have partially filled d-orbitals in their ground state or in one of their common oxidation states; general configuration is $(n-1)d^{1-10}ns^{1-2}$.
- b. **Main Group Metals:** Have completely filled d-orbitals (if present) and partially filled s-orbitals in their outermost shell; valence electrons are in s and p orbitals only.

2. Oxidation States

- a. **Transition Metals:** Exhibit multiple oxidation states due to the involvement of both d and s electrons in bonding; e.g., iron can form Fe^{2+} and Fe^{3+} ions.
- b. **Main Group Metals:** Typically show one or two common oxidation states corresponding to the loss of their valence electrons; e.g., sodium typically forms only Na^+ .

3. Complex Formation

- a. **Transition Metals:** Readily form coordination complexes with ligands due to their vacant d-orbitals that can accept electron pairs.
- b. **Main Group Metals:** Generally less prone to form complex ions, though some (like aluminium) can form limited complexes.

4. Coloured Compounds

- a. **Transition Metals:** Often form coloured compounds and complex ions due to d-d electronic transitions.
- b. **Main Group Metals:** Typically form colourless compounds (when the anion is also colourless) as they lack d-d transitions.

5. Catalytic Activity

- a. **Transition Metals:** Excellent catalysts (especially Fe, Ni, Pt, Pd) due to their ability to change oxidation states and provide adsorption sites.
- b. **Main Group Metals:** Generally show limited catalytic activity.

6. Magnetic Properties

- a. **Transition Metals:** Many exhibit paramagnetic or ferromagnetism due to unpaired electrons in d-orbitals.
- b. **Main Group Metals:** Typically diamagnetic (repelled by magnetic fields) due to paired electrons.

7. Density and Hardness

- a. **Transition Metals:** Generally have higher densities and hardness due to smaller atomic radii and stronger metallic bonding.
- b. **Main Group Metals:** Typically less dense and softer, especially those in Groups 1 and 2.

8. Melting and Boiling Points

- a. **Transition Metals:** Generally have higher melting and boiling points due to stronger metallic bonding.
- b. **Main Group Metals:** Often have lower melting and boiling points, particularly the alkali and alkaline earth metals.

9. Reactivity with Water/Acids

- a. **Transition Metals:** Show variable reactivity; many are relatively unreactive with water and dilute acids (e.g., copper, silver).
- b. **Main Group Metals:** Those in Groups 1 and 2 are typically highly reactive with water and acids.

Uses of Transition Metals in Everyday Life**1. Construction and Infrastructure**

- a. Iron and steel are used to build bridges, skyscrapers, and transportation systems due to their strength and durability.
- b. Titanium is used in aircraft frames and spacecraft components because it combines light weight with exceptional strength.
- c. Chromium is used in stainless steel production and chrome plating to provide corrosion resistance for automotive parts and household fixtures.

2. Electronics and Technology

- a. Copper is used in electrical wiring and circuit boards due to its excellent electrical conductivity.
- b. Gold is used in electrical connectors and high-end electronics because it doesn't corrode and maintains reliable connections.
- c. Silver is used in printed circuit boards and photovoltaic cells due to having the highest electrical conductivity of all metals.

3. Healthcare and Medicine

- a. Titanium is used to make artificial joints and dental implants because it's biocompatible and doesn't trigger immune responses.
- b. Platinum is used in chemotherapy drugs and pacemaker electrodes due to its biocompatibility and electrical properties.
- c. Silver is used in antimicrobial wound dressings and medical instruments because of its ability to kill bacteria.

4. Catalysts in Industrial Processes

- Iron is used as a catalyst in the Haber process to produce ammonia for fertilisers.
- Vanadium(V) oxide is used as a catalyst in the Contact process to manufacture sulfuric acid.
- Platinum and palladium are used in catalytic converters to reduce harmful emissions from vehicles.

5. Energy Production and Storage

- Cobalt is used in lithium-ion batteries for phones, laptops, and electric vehicles.
- Nickel is used in rechargeable batteries and in some solar panel technologies.
- Vanadium is used in redox flow batteries for large-scale energy storage systems.

6. Household Items

- Stainless steel (containing iron, chromium, and nickel) is used in cutlery, cookware, and kitchen appliances due to its corrosion resistance.
- Copper is used in plumbing systems and cookware because of its thermal conductivity and antimicrobial properties.
- Zinc is used to galvanise steel for trash cans, nails, and roofing to prevent rusting.

7. Coinage and Currency

- Nickel, copper, and zinc are used in most modern coins due to their durability and appropriate value.
- Gold and silver are used in commemorative coins and as investment metals due to their rarity and value.

8. Pigments and Colourants

- Titanium dioxide is used as a white pigment in paints, plastics, and paper.
- Cobalt compounds are used to produce blue glass, ceramics, and paints.

Learning Tasks

- Create flashcards with the name, symbol, and common oxidation states of 10 transition metals.
- Design an experiment that would demonstrate the difference in reactivity between zinc (transition metal) and sodium (main group metal) with water and dilute acids.
- Compare and contrast the oxidation states of chromium and potassium. Explain why chromium exhibits multiple oxidation states while potassium typically shows only one.
- Research and prepare a presentation on how the unique properties of transition metals make them essential in developing technologies for renewable energy.

Pedagogical Exemplars

Digital Learning Approach

- Start the lesson by reviewing the organisation of the periodic table, prompting learners to identify the locations of transition metals and main group metals while recalling basic metallic properties studied earlier.
- Conduct a quick whole-class brainstorming activity about the metals learners encounter daily. Have learners identify examples of both transition and main group metals from their experience. Record responses visibly for reference throughout the lesson.
- Guide groups to explore the properties of main group elements and record their findings on a comparison chart.
 - Compare the hardness of transition metals with main group metals using prior knowledge, textbook images, or sample materials (if available) to decide which metals are harder.
 - Compare the reactions with water and dilute acids. Note the differences in reactivity, release of gases or any heat observed.
 - Identify metals with multiple charges (e.g., Fe^{2+} and Fe^{3+}) and those with only one charge (e.g., Na^+).
 - Research on the metals that can form complex ions with ligands like water or ammonia. Note which type of metal forms complex ions more readily.
 - Dissolve a small amount of salts of different metals in distilled water and place the resulting solution on a white tile. Observe the colour of the solutions that form and record it against its corresponding metal.
 - Research on chemical reactions that involves the use of catalysts. Take note of the metals that are used as catalyst. Use MnO_2 to break down H_2O_2 if available.
- Guide learners in groups to discuss the uses of transition metals in everyday life in areas of construction, technology, medicine, transportation and household items.

Key Assessment Level 3: Group Work

- Copper, a transition metal, and magnesium, a main group metal, are both used in various applications. Analyse the chemical and physical properties of these two metals that make them suitable for their respective uses.

For your analysis:

- Compare and contrast their hardness and reactivity with water and acids.
 - Explain how their electronic configurations influence these properties.
 - Evaluate which properties make each metal particularly valuable for specific applications.
- The Ghana Atomic Energy Commission is researching materials for solar energy applications, while the Council for Scientific and Industrial Research (CSIR) is exploring catalysts for agricultural processing.

- i. Compare and contrast the catalytic properties of transition metals versus main group metals, and explain why the CSIR might prefer transition metals for developing catalysts for cassava processing.
- ii. Analyse how the hardness and reactivity with acids of different metals would affect their suitability for solar panel frames that must withstand Ghana’s tropical climate conditions.
- iii. The government of Ghana is promoting value addition to mineral resources. Propose and justify a specific application where the unique properties of transition metals (variable oxidation states, complex formation, or coloured compound formation) could be utilised to create higher-value products from Ghana’s mineral resources.

PLC Session 20

45 minutes	Planning & Reviewing the Learning Plan
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1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week’s concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Plan the week’s key assessment with your app
 - a. The key assessment for the week is **group work**. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the group work (NTS 3k-3q).

10 minutes	Content Exploration
------------	---------------------

5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
 - a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step
 - d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
 - a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
 - a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).

- c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

Section 5 Review

This section reviews all the lessons taught for weeks 18–20. This section discussed the physical and chemical properties of transition elements, covering their characteristic properties, complex formation, and comparisons with main group metals. Learners explored variable oxidation states, paramagnetism, coloured compounds, and catalytic abilities through discussions and practical experiments. It is encouraged that portions of the sections which require a review of concepts be administered to learners as presentation tasks. Learners should be provided with adequate learning materials such as periodic tables, transition metal compounds, hydrogen peroxide, manganese(IV) oxide, copper salts, ammonia solution, and molecular modelling kits to enable them to investigate catalytic behaviour, complex formation, and structural geometry. The practical demonstrations of transition metal properties, particularly the catalytic decomposition of hydrogen peroxide and the formation of coloured complex compounds, provide excellent opportunities for connecting theoretical concepts to observable phenomena and everyday applications of these versatile elements.

MARKING SCHEMES FOR THE KEY ASSESSMENTS

WEEK 18 – 19

FOCAL AREA 1

Catalytic analysis

- Graph construction and analysis (5 marks)
 - Award 1 mark for appropriate graph type selected (line graph) with correct axes labels and units
 - Award 1 for all three data sets accurately plotted with different colours/symbols and a legend
 - Award 1 mark for appropriate scale used allowing clear visualisation of differences
 - Award 1 mark for correct identification of MnO_2 as most effective catalyst, followed by CuSO_4 , then FeCl_3
 - Award 1 mark for analysis of reaction rates (initial rates and/or time to reach plateau) with reference to specific data points
- Scientific explanation (5 marks)
 - Award 1 mark for explanations that links catalytic activity to fundamental properties of transition metals
 - Award 1 mark for discussion of variable oxidation states and electron transfer capabilities
 - Award 1 mark for specific explanation of how MnO_2 functions as a catalyst (e.g., $\text{Mn}^{4+} \rightarrow \text{Mn}^{2+} \rightarrow \text{Mn}^{4+}$ cycle)

- Award 1 mark for comparison of the redox potentials or stability of different oxidation states of the three metals
- Award 1 mark for reference to surface effects, complex formation, or other relevant factors affecting catalytic efficiency.

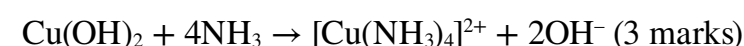
3. Experimental design (6 marks)

- Award 1 mark for clear, testable hypothesis relating catalytic activity to variable oxidation states
- Award 1 mark for identification of independent variable (different oxidation states of same metal or compounds with metals in different oxidation states)
- Award 1 mark for identification of dependent variable (rate of oxygen production) and how it will be measured.
- Award 1 mark for comprehensive list of controlled variables (temperature, concentration, surface area, etc.)
- Award 1 mark for logical, sequential experimental procedure with safety considerations
- Award 1 mark for reasonable prediction of expected results with scientific justification.

FOCAL AREA 2

(Total: 20 marks)

1. Chemical Equations (5 marks)



3-4 marks: Equations with minor errors or imbalances

1-2 marks: Partially correct equations with significant errors

2. (i) Explanation of Dissolution (5 marks)

Formation of complex ions with ammonia as ligand (2 marks)

Shift in equilibrium due to complex formation (1 mark)

Correct structure of at least one complex ion (2 marks)

3-4 marks: Good explanation with minor omissions

1-2 marks: Basic explanation with significant gaps

(ii) Structure and Colour Explanation (5 marks)

Correct square planar of $[\text{Cu}(\text{NH}_3)_4]^{2+}$ or octahedral structure of Cu^{2+} (2 marks)

Explanation that NH_3 is a stronger field ligand than H_2O (1 mark)

Connection between ligand field strength and colour intensity (2 marks)

3-4 marks: Mostly correct structure and partial explanation

1-2 marks: Structure with major errors and minimal explanation

WEEK 20**i. (10 marks)**

Criteria	Description	Marks
Comparison of hardness	- Correct identification that copper (transition metal) is generally harder than magnesium (main group metal) - Explanation linking hardness to metallic bonding and d-orbital involvement in transition metals - Discussion of crystalline structure differences.	2
Comparison of reactivity	- Accurate description of copper's low reactivity with water and acids compared to magnesium's high reactivity - Explanation of reactivity trends related to standard electrode potentials - Discussion of observable reaction differences (e.g., hydrogen evolution rates).	2
Electronic configuration analysis	- Correct electronic configurations for both metals - Explanation of how d-orbitals in copper affect its properties - Analysis of how valence electrons influence chemical behaviour.	2
Application suitability	- Thoughtful evaluation linking copper's properties to applications (e.g., electrical wiring, plumbing) - Thoughtful evaluation linking magnesium's properties to applications (e.g., lightweight alloys, flares) - Discussion of how specific properties determine application suitability.	3
Scientific communication	- Use of appropriate chemical terminology - Logical structure and coherence - Evidence-based reasoning.	1

ii. (10 marks)

Criteria	Description	Marks
Catalytic properties comparison	- Comprehensive explanation of transition metals' catalytic properties due to d-orbitals - Contrast with main group metals' limited catalytic abilities - Analysis of how variable oxidation states facilitate catalysis - Specific application to cassava processing (e.g., starch modification, waste treatment) - Reference to locally available transition metals in Ghana.	3

Materials suitability analysis	- Analysis of corrosion resistance requirements in Ghana's climate - Evaluation of different metals' reactivity with acidic rain-water - Discussion of hardness and structural integrity needs - Connection between chemical properties and long-term durability - Consideration of cost and availability in the Ghanaian context.	3
Value-addition proposal	- Creative yet scientifically sound proposal - Clear connection to specific properties of transition metals - Feasibility within Ghana's industrial capacity - Economic viability consideration - Environmental sustainability analysis - Specific reference to Ghanaian resources (e.g., manganese from Nsuta, gold from Obuasi) - Detailed justification based on chemical principles.	3
Scientific communication	- Use of appropriate chemical terminology - Logical structure and coherence.	2

SECTION 6: ALKANOIC ACID DERIVATIVES AND LIPIDS

STRAND: CHEMISTRY OF CARBON COMPOUNDS

Sub-Strand: Organic Functional Groups

Learning Outcome: Explain structure, preparation, properties of alkanolic acid derivatives and lipids as well as their uses in everyday life.

Content Standard: Demonstrate knowledge and understanding of the concepts of organic chemistry to describe the structure and properties of alkanolic acid derivatives and lipids.

INTRODUCTION AND SECTION SUMMARY

This section introduces learners to alkanolic acid derivatives and lipids, exploring the diverse family of organic compounds derived from carboxylic acids and their biological counterparts. Learners will investigate the structure, preparation methods, chemical properties, and practical applications of key derivatives including alkyl alkanates (esters), amides, and acyl halides. The section extends to examine fats and oils as naturally occurring lipids, their structural features, physical and chemical properties, and their significance in biological systems and commercial applications. By the end of this section, learners will be able to identify the functional groups of alkanolic acid derivatives, explain the reaction mechanisms involved in their formation, compare their relative reactivities, understand the relationship between structure and properties of fats and oils, and appreciate their widespread importance in contexts ranging from food science and nutrition to pharmaceuticals, cosmetics, and industrial processes. This knowledge forms an essential foundation for understanding the chemistry of biomolecules, organic synthesis pathways, and the technological applications of these versatile compounds in everyday life.

The weeks covered by this section are:

1. Weeks 21: Explain the structure, preparation and properties of alkanolic acid derivatives and their uses in everyday life.
2. Week 22: Explain structure, properties and uses of fats and oils in everyday life.

SUMMARY OF PEDAGOGICAL EXEMPLARS

This section requires learners to actively explore the structures, preparation, and uses of alkanolic acid derivatives and lipids through collaborative, experiential, and digital learning approaches. Mixed-ability and gender-inclusive groupings are essential in engaging learners to model, name, and write structural formulae of alkanolic acid derivatives like esters, amides, and acid chlorides. Learners investigate reactions using appropriate reagents and conditions, applying IUPAC rules and chemical equations.

Hands-on activities, such as preparing ethyl ethanoate and making soap from local materials, should dominate the learning experience. These provide opportunities to apply concepts like esterification, hydrolysis, and saponification. Through exploratory learning, learners test and analyse chemical properties, while discussions clarify key concepts like fat structures, unsaturation (iodine number), and catalytic hydrogenation.

Digital resources (videos, online research, presentations) help learners understand complex reactions and mechanisms, while talk-for-learning activities enable them to discuss detergent types and grease-removal mechanisms.

Approaching proficiency learners should be supported through oral explanations and visual models, while highly proficient learners extend learning to reaction mechanisms, equation writing, and industrial applications of esters and detergents. This inclusive and activity-based approach encourages deep conceptual understanding and real-life application of organic chemistry.

ASSESSMENT SUMMARY

Indicator	DOK Level	Assessment Strategy and/or Mode
Explain the structure, preparation and properties of alkanolic acid derivatives and their uses in everyday life.	Level 3: Strategic Reasoning	Individual Classwork
Explain structure, properties and uses of fats and oils in everyday life	Level 3: Strategic Reasoning	Individual Classwork

WEEK 21

Learning Indicator: Explain the structure, preparation and properties of alkanolic acid derivatives and their uses in everyday life

Recap weeks

Year 2	week 23: Explain the structure, nomenclature, preparation, properties and uses of alkanols as well as their characteristics test. Week 24: Explain the structure, nomenclature, preparation, properties and uses of alkanolic acids as well as their characteristic test.
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FOCAL AREA: ALKYL ALKANOATES (ESTERS)

Alkyl alkanoates (esters) are organic compounds formed when the alkyl group (R') of an alkanol replaces the ionisable hydrogen atom in an alkanolic acid (RCOOH) to form RCOOR', where R is H atom or an alkyl group and R' is an alkyl group.

Structure and Functional Group

Alkyl alkanoates contain $\text{R}-\overset{\text{O}}{\parallel}{\text{C}}-\text{O}-\text{R}'$ as their functional group, where R is hydrogen atom or alkyl group and R' is an alkyl group. The R and R' may be the same or different.

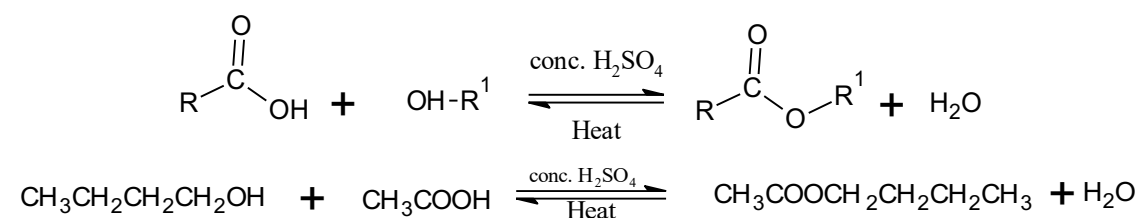
The functional group contains:

1. The polar carbonyl group (C=O)
2. C—O single bonds attached to the same carbon atom.

The general molecular formula for alkyl alkanoates is $\text{C}_n\text{H}_{2n}\text{O}_2$ where $n = 2, 3, 4, 5$, etc.

The first member of the alkyl alkanoate series corresponds to $n = 2$ and has the molecular formula $\text{C}_2\text{H}_4\text{O}_2$.

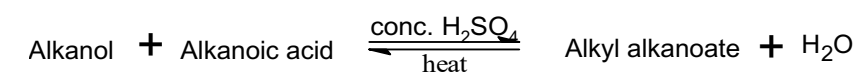
When an alkanol reacts with an alkanolic acid, the products formed are alkyl alkanoate and water:



For example, when butan-1-ol and ethanoic acid react, butyl ethanoate and water are produced:

Esterification

The formation of an alkyl alkanoate by the reaction between an alkanol and an alkanolic acid is known as esterification. This reaction is extremely slow and reversible at room temperature and is catalysed by a high concentration of hydrogen ions (typically from concentrated H_2SO_4).



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.

Esterification reaction	Neutralisation reaction
An alkanolic 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
Molecules react, hence very slow	Ions react in solution, hence very fast

Sources of Alkyl Alkanoates

- a. Animal and vegetable fats and oils are alkyl alkanoates with high relative molecular masses.
- b. Simple alkyl alkanoates are present in fruits of plants and are responsible for the natural sweet smell of flowers.
- c. Many have pleasant smells and are largely responsible for the fragrance of flowers and flavours of fruits.
- d. Pentyl ethanoate ($\text{CH}_3\text{COOC}_5\text{H}_{11}$) has the smell of bananas and is called banana oil.
- e. Octyl ethanoate ($\text{CH}_3\text{COOC}_8\text{H}_{17}$) has the smell of oranges.

IUPAC Nomenclature of Alkyl alkanoates

The name of an alkyl alkanoate is derived from the corresponding alkanol and alkanolic acid from which it is formed:

1. The alkanol portion (R') is named first as an alkyl group
2. The alkanolic acid portion (RCOO-) is named as an alkanoate

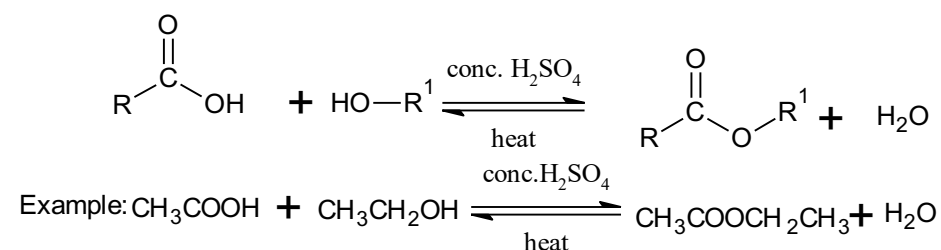
Alkanol	Alkanolic acid	Ester	IUPAC name of the alkyl alkanoate
Methanol	Methanoic acid	HCOOCH_3	Methyl methanoate
Methanol	Ethanoic acid	$\text{CH}_3\text{COOCH}_3$	Methyl ethanoate
Ethanol	Methanoic acid	$\text{HCOOCH}_2\text{CH}_3$	Ethyl methanoate
Ethanol	Ethanoic acid	$\text{CH}_3\text{COOCH}_2\text{CH}_3$	Ethyl ethanoate
methanol	Propanoic acid	$\text{CH}_3\text{CH}_2\text{COOCH}_3$	Methyl propanoate

Preparation of alkyl alkanoates

1. Esterification (Fischer Esterification)

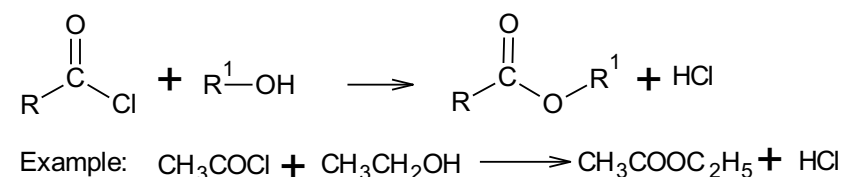
This method involves the direct reaction between alkanolic acid and alkanol in the presence of a strong acid catalyst, typically concentrated H_2SO_4 or HCl . The reaction is an equilibrium process that requires either excess alcohol or alkanolic acid.

The reaction is slow at room temperature and requires heating.



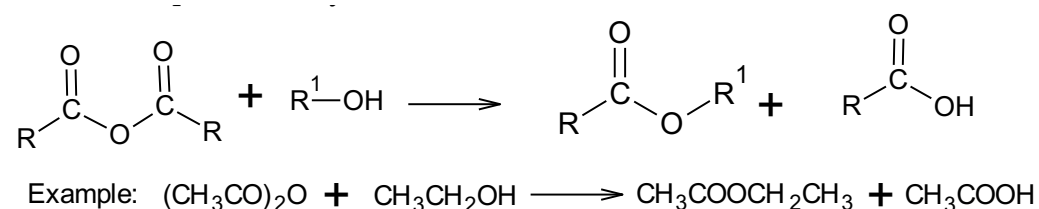
2. Reaction of Acyl Halides with Alcohols

Acyl halides (acid chlorides or bromides) are highly reactive derivatives of alkanolic acids. When they react with alcohols, they readily form alkyl alkanoates with the elimination of hydrogen halide (HCl or HBr). The reaction occurs rapidly even at room temperature and does not require heating.



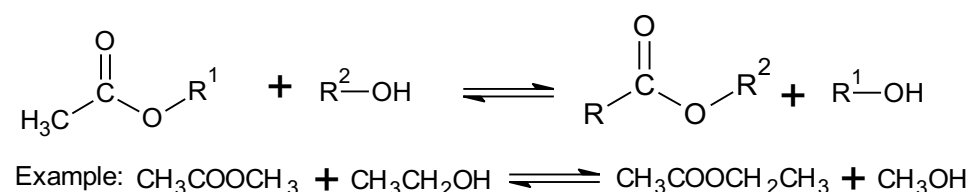
3. Reaction of Acid Anhydrides with Alcohols

Acid anhydrides are less reactive than acyl halides but more reactive than alkanolic acids themselves. They react with alcohols to form alkyl alkanoates and the corresponding alkanolic acid. The reaction proceeds at moderate temperatures and does not require a catalyst.



4. Transesterification

This method involves the exchange of the alkoxy group of an existing alkyl alkanoate with another alcohol. The reaction requires an acid or base catalyst (commonly H_2SO_4 , HCl, or alkoxides) and is an equilibrium process.



Physical Properties of Alkyl Alkanoates

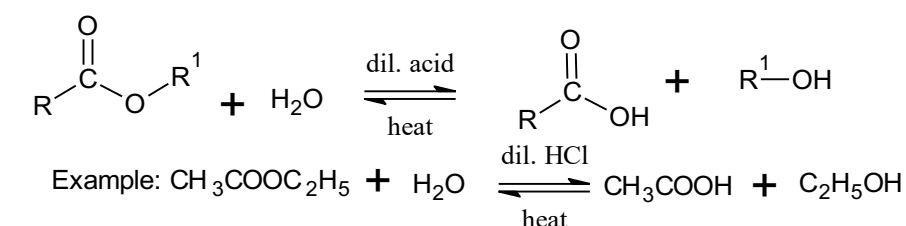
- Physical State and Appearance:** Alkyl alkanoates are typically colourless liquids at room temperature. Lower members have low viscosity while higher members (above C16) exist as waxy solids.
- Odour:** These compounds possess characteristic pleasant, fruity aromas. Examples include pentyl ethanoate (banana smell), octyl ethanoate (orange smell), and ethyl butanoate (pineapple smell).

- Boiling Points:** Alkyl alkanoates have lower boiling points than corresponding alkanolic acids due to absence of hydrogen bonding. Boiling points increase with molecular weight. Ethyl ethanoate boils at 77°C while ethanoic acid boils at 118°C despite the ester's higher molecular weight.
- Melting Points:** Melting points increase with molecular weight but remain relatively low compared to other organic compounds of similar size. The ester group creates some molecular rigidity, but alkyl chains adopt multiple conformations that hinder efficient crystal packing.
- Density:** Most alkyl alkanoates have densities slightly less than water (0.85-0.95 g/cm³). Density increases slightly with molecular weight but tends to plateau for higher members. Branched alkyl alkanoates have lower densities than straight-chain counterparts.
- Solubility:** Water solubility decreases dramatically as carbon chain length increases. Methyl ethanoate dissolves in water to about 25% by weight, while octyl ethanoate is practically insoluble. All alkyl alkanoates dissolve readily in organic solvents like alcohols, ethers, and hydrocarbons.

Chemical reactions of alkyl alkanoates

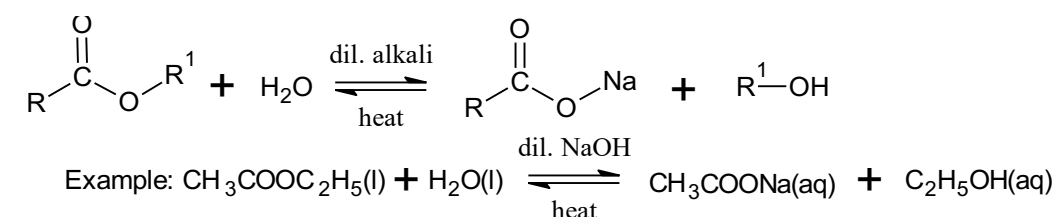
1. Acid hydrolysis of alkyl alkanoates:

This is the reverse of the esterification reaction. In the presence of aqueous acid and heat, alkyl alkanoates undergo hydrolysis to form the parent alkanolic acid and alkanol.



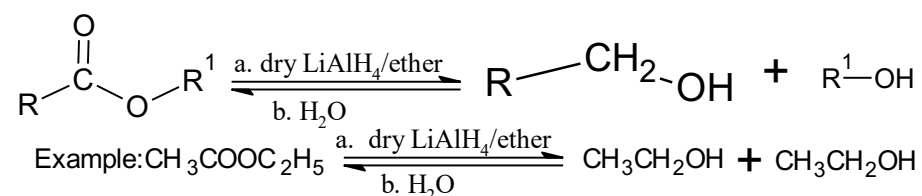
2. Base-catalysed hydrolysis (Saponification):

In the presence of aqueous base (usually NaOH or KOH), alkyl alkanoates undergo irreversible hydrolysis to form the carboxylate salt and alcohol. This reaction is called saponification because it is the basis for soap production from fats and oils.



3. Reduction:

With LiAlH_4 (Lithium Aluminium Hydride): Lithium aluminium hydride is a powerful reducing agent that can reduce alkyl alkanoates to primary alcohols. The reaction produces two alcohols, one from the acid portion and one from the alcohol portion of the ester.



AMIDES

Structure and Bonding

Amides contain the functional group $-\text{CONH}_2$ and are derived from alkanolic acids. They are classified as primary ($\text{R}-\text{CONH}_2$), secondary ($\text{R}-\text{CONHR}'$), or tertiary ($\text{R}-\text{CONR}'\text{R}''$) amides depending on the substitution on the nitrogen atom. The amide group has a planar structure due to resonance between the nitrogen lone pair and the carbonyl group. This resonance stabilisation makes amides more stable than would be expected from simple structure considerations.

The nitrogen atom in amides is sp^2 hybridised, unlike in amines where it is sp^3 hybridised, which explains the planar geometry around nitrogen in amides.

Nomenclature

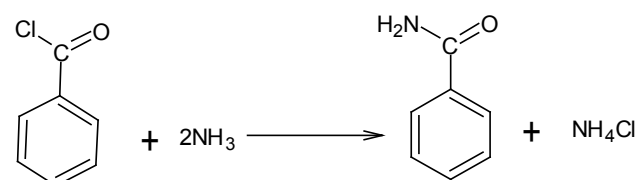
In the IUPAC system, amides are named by replacing the “-oic acid” ending of the parent alkanolic acid with “-amide.” For example, CH_3COOH (ethanoic acid) becomes CH_3CONH_2 (ethanamide), and $\text{C}_6\text{H}_5\text{COOH}$ (benzoic acid) becomes $\text{C}_6\text{H}_5\text{CONH}_2$ (benzamide). For N-substituted amides, the substituents on the nitrogen atom are indicated by using the prefix “N-” followed by the name of the substituent. For example, $\text{CH}_3\text{CONHCH}_3$ is named N-methylethanamide, and $\text{CH}_3\text{CON}(\text{CH}_3)_2$ is named N,N-dimethylethanamide.

Common names are also frequently used, especially for simpler amides. For instance, CH_3CONH_2 is commonly called acetamide (from acetic acid), and HCONH_2 is called formamide (from formic acid).

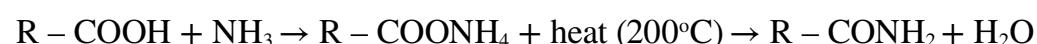
Preparation Methods of Amides

Amides can be prepared through several methods, each with specific reaction conditions and mechanisms:

1. From Acyl Halides (Schotten-Baumann Reaction): Acyl halides react rapidly with ammonia to form amides. The reaction proceeds through nucleophilic attack of ammonia on the carbonyl carbon, forming a tetrahedral intermediate, followed by elimination of the halide ion. For example, benzoyl chloride reacts with ammonia to form benzamide and ammonium chloride:



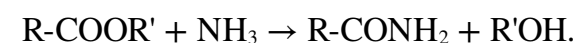
2. From Alkanolic acids: Direct conversion of alkanolic acids to amides requires heating with ammonia at high temperatures (around 200°C) in a pressure vessel.



3. From Acid Anhydrides: Acid anhydrides react with ammonia to form amides and ammonium salts of the corresponding acid. For example, acetic anhydride reacts with ammonia to form acetamide and ammonium acetate:



4. From Esters (Aminolysis): Esters react with ammonia to form amides and alcohols. The reaction involves nucleophilic attack by ammonia on the carbonyl carbon, formation of a tetrahedral intermediate, and elimination of the alkoxy group:



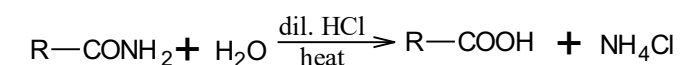
Physical Properties

1. Physical State: Lower amides like methanamide are liquids at room temperature, while higher amides such as ethanamide and propanamide exist as solids.
2. Boiling Points: Amides have significantly higher boiling points than corresponding acids of similar molecular weight due to hydrogen bonding; for example, methanamide boils at 210°C , ethanamide at 221°C , and propanamide at 213°C .
3. Melting Points: The melting points of amides are generally high compared to other organic compounds of similar molecular weight and tend to decrease with increasing chain length.
4. Solubility: Lower amides ($\text{C}_1\text{--}\text{C}_4$) are highly soluble in water, with solubility decreasing as carbon chain length increases; amides with more than six carbon atoms are practically insoluble in water, though all amides dissolve well in polar organic solvents.
5. Odour: Lower amides have distinctive, often unpleasant odours, while higher amides are generally odourless.
6. Density: Amides generally have densities greater than water, with density decreasing as molecular weight increases in a homologous series.

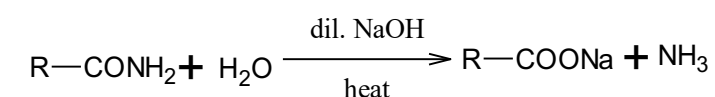
Chemical Properties

Amides undergo various chemical reactions, though they are generally less reactive than other alkanolic acid derivatives like acyl halides and esters. Their chemistry is dominated by reactions at the carbonyl group and transformations of the amide functional group:

1. Hydrolysis: Amides can be hydrolysed under both acidic and basic conditions, though the reaction is slow compared to other alkanolic acid derivatives.
 - i. In acidic hydrolysis, the amide reacts with water in the presence of an acid catalyst to form an alkanolic acid and an ammonium salt:

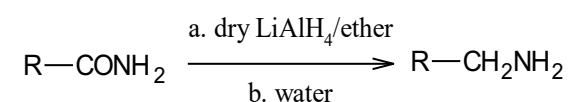


- ii. In basic hydrolysis, the amide reacts with hydroxide ions to form a carboxylate salt and ammonia:



This reaction typically requires $1\text{--}6 \text{ mol/dm}^3$ NaOH or KOH at $80\text{--}100^\circ\text{C}$.

- Reduction: Amides can be reduced to amines using strong reducing agents like lithium aluminium hydride (LiAlH_4) or through catalytic hydrogenation.



Applications of Amides

Amides find extensive applications across various industries due to their unique properties and chemical reactivity:

- Pharmaceuticals:** Amides are essential components in numerous medications, including paracetamol for pain relief, lidocaine for local anesthesia, penicillins as antibiotics, and barbiturates as sedatives.
- Polymers:** The amide linkage forms the backbone of important polymers like nylon used in textiles and engineering plastics, Kevlar employed in bulletproof vests, and polyacrylamide utilised in water treatment and laboratory procedures.
- Solvents:** Amide-based solvents such as DMF (N,N-dimethylformamide), DMAc (N,N-dimethylacetamide), and NMP (N-methyl-2-pyrrolidone) are valued in industry for their high boiling points, excellent solvation properties, and ability to dissolve a wide range of compounds including polymers.
- Pesticides and Herbicides:** Many agricultural chemicals contain amide functional groups, including propanil for rice cultivation, alachlor as a pre-emergence herbicide, and carbaryl as an insecticide.
- Surfactants and Plasticisers:** Alkanolamides serve as foam stabilisers in personal care products, while amide-based compounds like N-butylbenzenesulfonamide function as plasticisers to improve flexibility in plastic materials.
- Biological Systems:** Amide bonds are fundamental to life, forming the peptide linkages in proteins, appearing in essential biomolecules like urea and neurotransmitters, and playing crucial roles in DNA and RNA structure.

ACYL HALIDES

Structure and Bonding

Acyl halides are organic compounds with the general formula $\text{R}-\text{COX}$, where R is an alkyl or aryl group and X is a halogen atom (typically chlorine, bromine, fluorine, or iodine). They are derivatives of carboxylic acids where the $-\text{OH}$ group of the carboxylic acid is replaced by a halogen atom.

The structure has a planar geometry around the carbonyl carbon due to sp^2 hybridisation, with bond angles of approximately 120° .

Nomenclature of Acyl Halides

- Identify the parent carboxylic acid
 - Determine the carbon chain or ring structure attached to the $-\text{COX}$ group
 - Count the number of carbon atoms in this chain or identify the ring system

- Replace the ending of the parent acid
 - Change “-oic acid” ending to “-oyl halide”
 - For example: ethanoic acid $\rightarrow$ ethanoyl chloride
- Specify the halogen
 - Indicate which halogen is present (chloride, bromide, fluoride, or iodide)
 - For example: ethanoyl chloride (for $-\text{CH}_3\text{COCl}$), propanoyl bromide (for $\text{CH}_3\text{CH}_2\text{COBr}$)
- Add prefixes for any substituents
 - Number the carbon chain starting with the carbonyl carbon as position 1
 - Name and number any substituents as prefixes

For example: 3-methylbutanoyl chloride $\{\text{CH}_3\text{CH}(\text{CH}_3)\text{CH}_2\text{COCl}\}$

Examples of IUPAC Names of some Acyl Halides

- CH_3COCl : ethanoyl chloride (from ethanoic acid)
- $\text{CH}_3\text{CH}_2\text{COCl}$: propanoyl chloride (from propanoic acid)
- $\text{CH}_3\text{CH}_2\text{COBr}$: propanoyl bromide (from propanoic acid)
- $\text{CH}_2=\text{CHCOCl}$: propenoyl chloride (from propenoic acid)
- $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{COCl}$: pentanoyl chloride (from pentanoic acid)
- $\text{C}_6\text{H}_5\text{COCl}$: benzoyl chloride (from benzoic acid)
- ClCH_2COCl : chloroethanoyl chloride (from chloroethanoic acid)

Examples of Common Names of some Acyl Halides

- CH_3COCl : acetyl chloride (from acetic acid)
- HCOCl : formyl chloride (from formic acid, though unstable)
- $\text{CH}_3\text{CH}_2\text{COCl}$: propionyl chloride (from propionic acid)
- $\text{C}_6\text{H}_5\text{COCl}$: benzoyl chloride (from benzoic acid)
- $\text{CH}_3(\text{CH}_2)_{16}\text{COCl}$: stearoyl chloride (from stearic acid)
- $\text{CH}_2=\text{CHCOCl}$: acryloyl chloride (from acrylic acid)

Preparation Acyl Halides

Acyl halides are typically prepared from alkanolic acids or their derivatives through several methods:

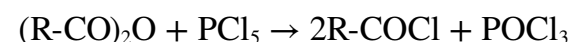
- Reaction of carboxylic acids with thionyl chloride in the presence of heat ($40\text{--}80^\circ\text{C}$)

$$\text{R}-\text{COOH} + \text{SOCl}_2 \rightarrow \text{R}-\text{COCl} + \text{SO}_2 + \text{HCl}$$
- Reaction of carboxylic acids with phosphorus pentachloride at $50\text{--}100^\circ\text{C}$

$$\text{R}-\text{COOH} + \text{SOCl}_2 \rightarrow \text{R}-\text{COCl} + \text{SO}_2 + \text{HCl}$$
- Reaction of carboxylic acids with phosphorus trichloride at $70\text{--}100^\circ\text{C}$

$$3\text{R}-\text{COOH} + \text{PCl}_3 \rightarrow 3\text{R}-\text{COCl} + \text{H}_3\text{PO}_3$$

4. Reaction of acid anhydrides with phosphorus pentachloride at 50-80°C:

**Physical Properties**

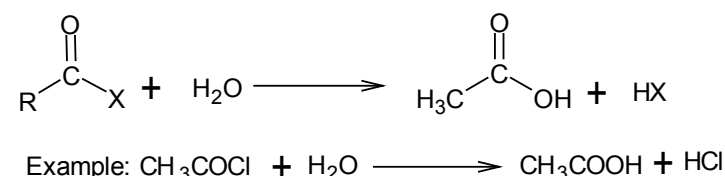
Acyl halides exhibit distinctive physical properties that reflect their molecular structure and reactivity:

- Physical State:** Lower acyl chlorides (C₁-C₉) are colourless, fuming liquids at room temperature, while higher acyl chlorides (C₁₀+) are low-melting solids.
- Boiling Points:** Acyl halides have higher boiling points than corresponding alkyl halides but lower than their parent alkanolic acids. For acyl halides of the same carbon count, the boiling point increases in the order: R-COF < R-COCl < R-COBr < R-COI, reflecting the increasing molecular weight and Van der Waals forces.
- Melting Points:** Acyl halides generally have low melting points for small molecules, which increase with molecular weight. For example, acetyl chloride has a melting point of -112°C, propanoyl chloride -94°C, and hexanoyl chloride -87°C.
- Density:** Acyl halides typically have densities greater than 1 g/cm³. For instance, acetyl chloride has a density of 1.11 g/cm³, and benzoyl chloride 1.21 g/cm³.
- Odour:** Acyl halides have a sharp, pungent, irritating smell and cause lachrymation (tearing) due to their reaction with moisture in the eyes. They are also respiratory irritants due to hydrolysis in moist tissues.

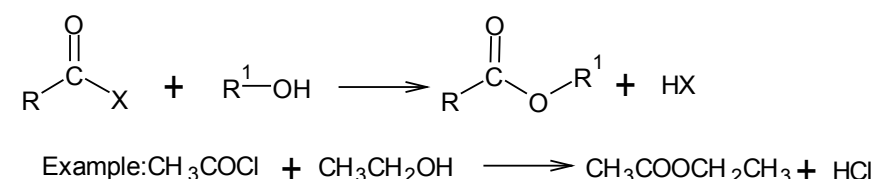
Chemical Properties

Acyl halides are among the most reactive alkanolic acid derivatives and undergo a wide range of nucleophilic substitution reactions. Their high reactivity is due to the good leaving group ability of the halide ion and the strong polarisation of the carbonyl group:

- a. **Hydrolysis:** Acyl halides react rapidly with water to form alkanolic acids and hydrohalic acids:



- b. **Alcoholysis (Esterification):** Acyl halides react with alcohols to form esters and hydrohalic acids:



This reaction is used industrially in the production of aspirin (acetylsalicylic acid).

- c. **Reduction with Metal Hydrides:** Acyl halides can be reduced to primary alcohols using strong reducing agents like lithium aluminium hydride (LiAlH₄):

**Uses of Acyl halides**

- Ester Synthesis:** Acyl halides react with alcohols to produce esters used in perfumes, flavours, and pharmaceuticals like aspirin.
- Amide Production:** They react with ammonia or amines to form amides essential for pharmaceuticals like paracetamol and in the production of nylon.
- Chemical Analysis:** Acyl halides are used to derivatise compounds for analytical techniques like gas chromatography.
- Dye Production:** They serve as intermediates in the synthesis of azo and anthraquinone dyes for the textile industry.
- Agricultural Chemical Synthesis:** Acyl halides are precursors for herbicides, insecticides, and fungicides.
- Catalyst Preparation:** Acyl halides are used to synthesise organometallic catalysts for industrial processes.

Learning Tasks

- Working individually or in pairs, complete the following activities to explore the structures and naming of amides and acyl halides.
 - Draw the general structures of the following alkanolic acid derivatives and label all functional groups clearly
 - Primary, secondary, and tertiary amides
 - Acyl chlorides
 - Practice naming the following compounds according to IUPAC rules:
 - CH₃CH₂COCl
 - C₆H₅CONH(CH₃)
 - C₆H₅COCl
 - Draw the structures for the following compounds:
 - Benzoyl chloride
 - N-methylethanamide
 - 3-methylbutanoyl bromide
- Research paracetamol (acetaminophen):
 - Draw its structure and identify the amide group
 - Explain how it works as a pain reliever
 - Describe how its structure relates to its function
- Design a synthetic route to prepare benzocaine (ethyl 4-aminobenzoate) starting from 4-nitrobenzoic acid. Your synthesis must include an acyl chloride intermediate.
 - Draw each intermediate compound
 - List reagents and conditions for each step
 - Explain why each step is necessary

Pedagogical Exemplars

Digital Learning

1. Use pre-downloaded or teacher-prepared PowerPoint slides showing structures and 3D representations of amides and acyl halides.
2. Distribute printed handouts with key information about alkanolic acid derivatives that would normally be researched online.
3. Show offline video clips (if available) or use detailed diagrams to demonstrate preparation methods for amides and acyl halides.
4. Have learners draw molecular structures on paper and compare with printed reference images of correct structures.
5. Provide printed worksheets with fill-in-the-blank activities about properties and nomenclature of the derivatives.
6. Use physical molecular model kits (if available) instead of digital modelling software.
7. Conduct a written quiz rather than a digital one to check understanding of basic structures and nomenclature.

Collaborative Learning

1. Form groups of 3-4 learners with mixed abilities, ensuring each group has members who researched different derivatives.
2. Have each group draw structural formulae for five examples of amides and five examples of acyl halides.
3. Direct groups to assign IUPAC names to each structure drawn and check each other's work.
4. Ask groups to create a comparison chart showing how amides and acyl halides differ in structure and properties.
5. Have groups discuss and record the formation reactions of amides and acyl halides from carboxylic acids, specifying reagents and conditions.
6. Ask each group to present one key reaction to the class.
7. Facilitate a brief class discussion to address common misconceptions identified during presentations.

Exploratory Learning

1. Provide groups with a basic procedure for preparing ethyl ethanoate and necessary safety guidelines.
2. Ask each group to identify the materials and equipment needed for their experiment.
3. Have groups predict observations and possible challenges before starting the experiment.
4. Direct learners to conduct the experiment under supervision, recording all observations.
5. Guide learners to test the product's properties (odour, solubility) and compare with expected results.
6. Have groups calculate percentage yield of their product.

7. Ask each group to suggest one improvement to the experimental procedure.

Talk-for-Learning

1. Pair learners and provide a set of prompt cards with different reagents (water, alcohols, ammonia, etc.).
2. Ask pairs to discuss how amides and acyl halides would react with each reagent.
3. Direct pairs to compare the reactivity of amides versus acyl halides and explain differences.
4. Have learners discuss real-world applications of these reactions.
5. Ask pairs to join another pair to share and refine their ideas.
6. Select several groups to share key insights from their discussions with the whole class.
7. Summarise the main points and clarify any misconceptions.

Collaborative Learning (Presentations)

1. Divide the class into small mixed-gender or mixed-ability groups of 4-5 learners.
2. Assign each group a specific focus area: reaction mechanisms, reduction reactions, or testing methods.
3. Provide basic templates and resources for creating effective presentations.
4. Give groups 20-30 minutes to prepare a 3-5 minute PowerPoint presentation on their assigned topic.
5. Have each group present their findings to the class, ensuring all members participate.
6. After presentations, facilitate a whole-class discussion on synthetic pathways and applications.
7. Conduct a brief exit ticket assessment to check understanding of key concepts covered.

Key Assessment Level 3: Individual Classwork

1. Identify the basic structures and preparation of amides and acyl halides.
 - a. Draw the general structure of an amide and an acyl halide. Circle and label the functional group in each. (4 marks)
 - b. Name the following compounds according to IUPAC rules:
 - a. $\text{CH}_3\text{CH}_2\text{CONH}_2$
 - b. CH_3COCl
 - c. $\text{C}_6\text{H}_5\text{CONH}(\text{CH}_3)$ (5 marks)
 - c. Write a balanced chemical equation for the preparation of ethanoyl chloride from ethanoic acid using thionyl chloride (SOCl_2). (4 marks)
2. Analyse and connect concepts related to amides and acyl halides in real-world contexts.
 - a. Paracetamol is a common pain reliever that contains an amide group. Draw the structure of paracetamol, identify the amide group, and explain how its structure contributes to its function as a medication. (6 marks)

- b. Design a synthetic route to prepare the local anesthetic benzocaine (ethyl 4-aminobenzoate) starting from 4-nitrobenzoic acid. Your synthesis must include an acyl chloride intermediate. Explain each step and the required reagents and conditions. (8 marks)

PLC Session 21

45 minutes	Planning & Reviewing the Learning Plan
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1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
- a. explain the week’s concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to
- a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



Note

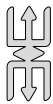
- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Plan the week’s key assessment with your app
- a. The key assessment for the week is **individual classwork**. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the classwork (NTS 3k-3q).

10 minutes	Content Exploration
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5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
- a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step

- d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
- a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
- a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

WEEK 22

Learning Indicator: Explain structure, properties and uses of fats and oils in everyday life

Recap weeks

Year 1	Week 24: Distinguish between organic and inorganic compounds and classify organic compounds.
Year 2	Week 19: Explain hybridisation and relate it to various structures. Week 20: Apply the concept of hybridisation to explain sigma and pi bond of organic compounds and explain their structure and bonding

FOCAL AREA 1: FATS AND OILS

Structure, properties and uses of fats and oils

Fats and oils are key members of the lipid family—organic compounds made predominantly of carbon, hydrogen, and a small proportion of oxygen. They are non-polar molecules, making them hydrophobic (water-repelling), and they serve as important energy sources and structural materials in living systems.

In chemistry, fats and oils are classified as triglycerides, formed through a condensation reaction between one molecule of glycerol (a triol) and three fatty acid molecules. This reaction results in the formation of ester linkages, which are key features of lipid structure.

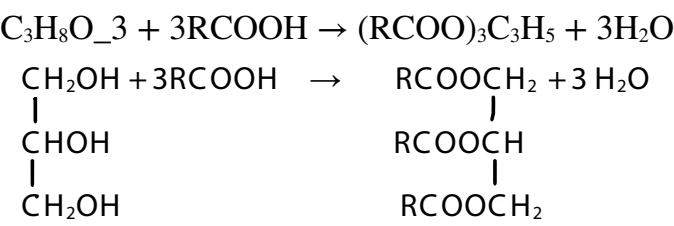
Fats are generally solid at room temperature and are composed mostly of saturated fatty acids with single carbon–carbon bonds. Oils, on the other hand, are liquid at room temperature and contain a higher proportion of unsaturated fatty acids, which have one or more carbon–carbon double bonds. These double bonds create kinks in the hydrocarbon chains, preventing close packing and resulting in a lower melting point. Fats are derived mainly from animal sources like butter and lard while oils are derived from plant sources such as olive and canola oil.

Structure of Fats and Oils

The fundamental units of fats and oils are triglycerides. Triglycerides are triesters formed from one glycerol molecule bonded to three fatty acids. Glycerol has a simple molecular structure consisting of a three-carbon unit with three hydroxyl (-OH) groups and serves as the backbone of triglycerides. The carbon atoms are arranged in a straight chain and each carbon atom is bonded to a hydroxyl group. This structure makes glycerol a triol (a molecule with three alcohol groups) with a chemical formula (C₃H₈O₃) or (CH₂OHCHOHCH₂OH). Fatty acids are molecules that consist of a long hydrocarbon chain and a carboxylic acid group (-COOH) at one end. The chemical formula of fatty acid group can be written as RCOOH where R represents a long hydrocarbon chain. The length of the hydrocarbon chain can be the same (forming a simple triglyceride) or vary (forming a mixed triglyceride) but it typically contains an even number of carbon atoms (usually between 12 and 20).

If all the bonds between carbon atoms in the chain are single bonds, the fatty acid is classified as saturated. These fatty acids are usually straight and can closely pack together, resulting in them often being solid at room temperature. If there is one or more double bonds between carbon atoms in the chain, the fatty acid is considered unsaturated. These double bonds

introduce kinks in the chain, which hinders tight packing and results in a liquid state at room temperature.



Sources

Fats and oils are naturally occurring substances found in both plant (fruits and seeds) and animal (tissues) sources. The tissues of marine animals such as fishes provide fish oil and cod-liver oil while the fruit and seeds of plants such as cocoa provide butter fat, palm oil, coconut oil etc. These sources provide essential energy and nutrients in the human diet and are widely used in cooking, food processing, and various industrial applications.

Table 23.1: The table describes the sources of fats and oils and their predominant fatty acids

Source	Common Name	IUPAC Name of Predominant Fatty acids
Beef	Beef fat	Tristearin
Pork	Lard	Tristearin
Chicken	Chicken fat	Tristearin
Butter	Butter fat	Butyrate
Olive oil	Olive oil	Oleic acid
Canola oil	Canola oil	Erucic acid
Sunflower oil	Sunflower oil	Linoleic acid
Soybean oil	Soybean oil	Triglycerides
Coconut oil	Coconut oil	Lauric acid
Avocado oil	Avocado oil	Oleic acid
Almond oil	Almond oil	Oleic acid
Flaxseed oil	Flaxseed oil	Alpha linolenic acid

Physical Properties

Fats are non-crystalline solids at room temperature while oils are liquids.

Many fats and oils are colourless, odourless and tasteless in their pure state. The characteristic colour, odour or flavour associated with them are due to foreign substances that are added. Example beta-carotene provides the characteristic yellow colour to butter, while butyric acid and diacetyl contribute to its rich and creamy taste.

They are insoluble in water due to the presence of long hydrocarbon chains (fatty acids) that are non-polar. These non-polar molecules do not interact favourably with polar molecules such as water due to the lack of hydrogen bonding between them.

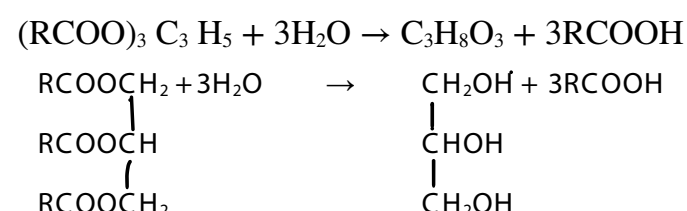
They have relative densities less than one, which makes them lighter than water. Typically, fats have a relative density ranging from 0.85 to 0.95 (Butter is 0.86) while oils have relative densities between 0.90 and 0.93 (olive oil is 0.91). This causes them to float in water, forming a layer at the upper surface.

They readily form emulsions when agitated with water in the presence of emulsifiers such as soap, gelatin or lecithin. Emulsions help to create a wide range of stable products. Example, a stable oil-in-water emulsion (mayonnaise) is formed when lecithin in egg yolk is mixed vigorously with oil and vinegar.

They are poor conductors of heat and electricity therefore serve as excellent insulators for the animal body, cooking and transformers.

Chemical Properties

1. **Hydrolysis** – Fats and oils are easily hydrolysed by enzymes (lipases) in the digestive track of animals to give fatty acid and glycerol which is crucial for the digestion and absorption of dietary fats. Before hydrolysis, fats and oils are emulsified by bile salts in the small intestine, increasing the surface area for enzyme action. Pancreatic lipase, secreted by the pancreas, acts on the emulsified fats, hydrolysing the ester bonds in triglycerides. The general reaction for the hydrolysis of a triglyceride can be represented as:



This reaction can occur under different conditions and can be classified into two main types: acidic hydrolysis and basic hydrolysis (saponification).

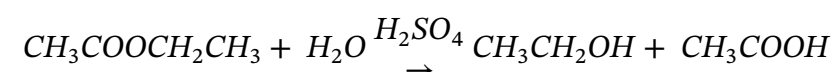
Acidic hydrolysis: It occurs in the presence of an acid catalyst.

Mechanism:

The acid donates a proton (H^+) to the carbonyl oxygen of the ester, increasing its electrophilicity. Water acts as a nucleophile, attacking the carbonyl carbon, leading to the formation of a tetrahedral intermediate.

The intermediate collapses, releasing the alcohol and regenerating the acid.

An example of an acid catalysed hydrolysis can be represented by the equation



Basic hydrolysis: This is commonly known as saponification reaction. It occurs in the presence of a strong base, resulting in the formation of soap (salt of fatty acids) and glycerol.

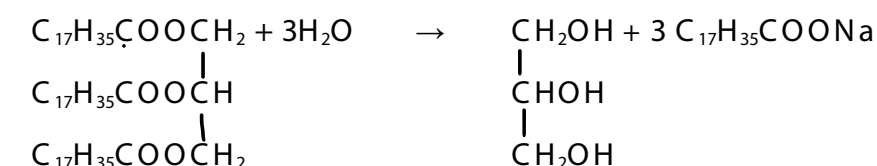
Mechanism:

The strong base (e.g. NaOH) deprotonates the water, generating a hydroxide ion (OH^-) that acts as a strong nucleophile.

The hydroxide ion attacks the carbonyl carbon of the ester, forming a tetrahedral intermediate.

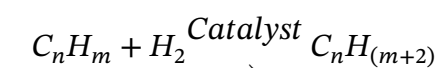
The intermediate collapses, resulting in the formation of a carboxylate salt and glycerol.

An example of a basic hydrolysis can be represented by the equation



2. **Hydrogenation** – This is the chemical process of adding hydrogen to unsaturated fats or oils to form saturated fats. The process of hydrogenation requires a catalyst such as nickel, palladium or platinum under high temperatures and pressures. Unsaturated fats are typically liquid at room temperature. Hydrogenation solidifies these fats into saturated fats which are solid at room temperature. Therefore, hydrogenation is also referred to as the hardening of oils.

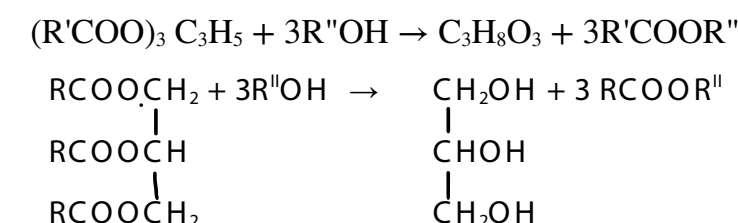
The general chemical reaction for hydrogenation can be represented as:



This indicates that hydrogen is added across the double bonds in unsaturated fats.

3. **Transesterification** – This is the process of esterification of fatty acids of fats or oils by the action of alcohol. In this reaction the glycerol backbone is removed where there is an exchange of the organic functional group R' of an ester with the organic functional group R'' of an alcohol. This reaction is essential in the production of biodiesel.

The general chemical reaction for transesterification can be represented as:



4. **Rancidity** – It is the chemical decomposition of fats, oils, or other lipids due to complete or incomplete oxidation or hydrolysis when exposed to light, air or moisture. This results in an unpleasant taste and odour. Oxidative rancidity occurs in triglycerides containing unsaturated acid. The acid produced by hydrolysis oxidises the double bonds forming short-chain offensive aldehydes, ketones and short-chain fatty acids. The process of rancidity can be accelerated by light exposure, heat, metal ions (especially copper and iron) and enzyme activity.

The equation for rancidity can be represented as:

Unsaturated Fat + Oxygen $\rightarrow$ Hydroperoxides $\rightarrow$ Aldehydes + Ketones + Short-chain Fatty Acids

Rancidity can be prevented through methods such as adding natural or synthetic antioxidants, using vacuum or nitrogen-flushed packaging, storing fats in cold conditions, hydrogenating unsaturated fats (with caution due to trans fats), and using metal chelators like citric acid to inhibit oxidation.

Uses

Fats and oils are essential components in various aspects of daily life, serving numerous functions across multiple industries. These lipid substances are not only crucial in culinary practices, where they enhance flavour and texture, but they also play significant roles in personal care, pharmaceuticals, and industrial applications.

1. Cooking and Food Preparation

Fats and oils (cooking oils like frytol) are used for frying and making stews.

Flavour enhancers like butter and lard are used to add richness and flavour to dishes.

Emulsifiers like egg yolk helps to blend water and fats in dishes and sauces (e.g., mayonnaise).

Fats like butter provide texture and moisture in cakes, cookies and pastries.

Fats such as shortening, butter and margarine helps to retain moisture and prolong the freshness of baked products.

2. Shelf life

Coconut oil contains lauric acid, which has strong antimicrobial and anti-fungal properties effective against food-borne pathogens. Local food fried in coconut oil include agbelikakro.

Palm oil contains tocotrienols and carotenoids which have antioxidant properties which reduces microbial activity by limiting oxidative stress in food products. Local food prepared with palm oil include eto and kontomire stew.

Olive oil is rich in phenolic compounds that inhibit the growth of bacteria and fungi, helping preserve foods like vegetables and cheeses.

Clove oil contains eugenol, a powerful natural antimicrobial agent that is effective against bacteria, yeasts, and moulds. Local food that contains cloves include shito and waakye stew.

Rosemary oil possesses both antioxidant and antimicrobial properties, useful in extending the shelf life of meats and oils.

3. Cosmetic and Personal Care Products

Shea butter contains stearic and oleic acids which is used as a moisturiser in lotions, creams and balms to hydrate the skin.

Coconut oil contains lauric acid which have moisturising properties. It is common in body lotions, lip balms and hair conditioners.

Avocado oil contains vitamins A, D and E which nourishes dry, sensitive or damaged skin.

Sunflower seed oil is high in linoleic acid and vitamin E that helps to maintain the skin barrier and suitable for sensitive skin. Found in most baby products and lotions.

4. Industrial Application

They are used in the manufacture of soap.

Unsaturated oils serve as a raw material for the production of margarine.

5. Pharmaceutical Application

They are used in transporting fat-soluble vitamins (A, D, E, K).

They are used as food mainly to provide energy.

They serve as a source of essential fatty acids.

Fats stored in the body of animals protects them from cold.

Manufacture of Margarine

Margarine is produced from highly refined vegetable oils through the process of catalytic hydrogenation of the alkene functional groups of unsaturated fatty acids.

Procedure

1. Large quantities of oils such as palm oil and groundnut oil are heated to about 200°C to remove undesirable material.
2. Hydrogen gas is passed through the hot oil at about 1 – 5 atmospheric pressure in the presence of an activated nickel catalyst. This hardens the oil to a desirable level as it becomes progressively saturated by controlling the extent of hydrogenation.
3. The hardened oil is emulsified with milk.
4. The resulting product is mixed with measured amounts of salt, vitamins, flavours, colour, preservatives and kneaded to a required texture. Margarine is used as a substitute for butter.
5. It is then packaged for sale.

Preparation of soap

Soaps are produced by the reaction of fatty acids and soluble bases, with glycerol as a by-product. This process is called saponification. Sodium hydroxide and potassium hydroxide are the common alkalis used. Soap can be produced locally and industrially.

Local Production of soap

Materials

1. Oils (Common options include palm oil, palm kernel oil, coconut oil or shea butter oil)
2. Ashes (from burnt cocoa pods or plantain peels)
3. Water
4. Salt

Steps

1. The ashes are dissolved in sufficient amount of water and the resulting solution filtered using fine mesh or cloth to remove any undissolved particles. The resulting filtrate is the alkali solution.
2. The desired volume of oil is heated in a separate pot until it reaches its boiling point. Care is taken not to overheat the oil which can degrade it.
3. The filtered alkali solution is then added gradually to the boiling oil and stirred continuously. This helps to evenly distribute the alkali throughout the oil.
4. The mixture is continuously stirred for approximately 20 – 30 minutes. This helps to initiate the saponification process which forms the soap and glycerol. The mixture will begin to thicken and change in texture, indicating that soap is being formed.

- Once the soap mixture has thickened, a prepared saturated sodium chloride solution is slowly added to the mixture and allowed to cool.
- The glycerol by-product is not properly removed in most cases (even though this has been improved in certain areas).
- The soap formed is moulded into desired shape and packaged.

Industrial Production of soap

Materials

- Fats or oils (Common sources include palm oil, coconut oil, and olive oil)
- Bases (Sodium hydroxide or potassium hydroxide)
- Brine solution (concentrated salt solution)
- Water

Steps

- The selected fats or oils are placed in a large pot and heated until it reaches its boiling point. This facilitates the saponification reaction, allowing for better mixing with the base.
- A pre-measured base is dissolved in a small amount of water and then carefully added to the boiling oil. The mixture is stirred continuously to ensure uniform distribution of the base throughout the boiling fats or oil. This reaction forms crude soap and glycerol as byproducts through a process known as saponification.
- The mixture is allowed to continue boiling for a specified duration (usually around 1-2 hours), while stirring occasionally. The mixture thickens as the reaction progresses, indicating the formation of soap.
- Brine is slowly added to the resulting mixture to salt-out the crude soap from the glycerol. This causes the soap to separate and float to the top.
- The crude soap is skimmed off from the top of the mixture and the remaining glycerol is extracted and purified.
- The soap is washed multiple times with water to remove any remaining salt and glycerol impurities and then bleached. This step is repeated to remove undesirable colours, odour and other impurities.
- Additives such as preservatives, fragrances and colourants are then added and mixed thoroughly with the soap. This gives a high quality soap called neat soap.
- The neat soap is skimmed off any excess liquid and moulded into various shapes. The moulded soap is allowed to dry and harden for a specified period.

Learning Tasks

- Working in groups or in pairs, create a labelled diagram and reaction summary:
 - Draw and label the structural formula of a triglyceride, indicating the glycerol backbone and three fatty acid chains.

- Write the general chemical equation for the formation of a triglyceride from glycerol and fatty acids.
 - Show the reverse (hydrolysis) reaction, identifying the role of acidic or basic conditions.
- Research and construct a comparative table to analyse:
 - The structural differences between saturated and unsaturated fatty acids.
 - How these differences affect the physical state (solid vs. liquid) at room temperature.
 - Common animal and plant sources of each type
 - Develop a concept map showing the interconnections between:
 - Fats, oils, glycerol, and fatty acids.
 - Chemical reactions such as hydrolysis, hydrogenation, rancidity, and transesterification.
 - Their applications in food, cosmetics, pharmaceuticals, and industry.
 - Model or simulate a process flow that:
 - Outline and illustrate the steps involved in either local or industrial production of soap.
 - Highlight the reagents used, intermediate changes, and final products (soap and glycerol).
 - Explain how brine aids in separating soap during industrial saponification.
 - List and explain at least four physical and chemical properties of fats and oils that make them valuable in:
 - Food preparation and preservation.
 - Personal care and cosmetics.
 - Pharmaceutical or industrial applications.
 - Investigate rancidity:
 - Explain the process of oxidative rancidity with a focus on the chemical changes involved.
 - Suggest four practical methods to prevent rancidity, relating each to its chemical action.
 - Provide real-life food examples (e.g., butter, palm oil) where rancidity is commonly observed.

Pedagogical Exemplars

Collaborative Learning

- Form gender-based or mixed-ability groups and task each group with writing the general molecular formula of a fat (triglyceride) molecule.
- Provide molecular model kits or cutouts for groups to construct and compare saturated and unsaturated fat structures; have them identify structural and physical property differences (e.g., melting point, state at room temperature).

3. Ask groups to list common sources of fats and oils, categorising them as animal- or plant-based.
4. Guide each group to explain why fats and oils are classified as triglycerides and to write a balanced chemical equation for the formation of a triglyceride from glycerol and three fatty acid molecules.
5. Introduce the concept of iodine number, then have each group apply it to assess the degree of unsaturation in a given fat sample.
6. Assign subtopics to each group to research and present on the reactions of fats and oils:
 - a. Hydrolysis using alkaline or acidic conditions.
 - b. Enzymatic hydrolysis.
 - c. Catalytic hydrogenation (e.g., margarine production).
 - d. Rancidity.

Experiential Learning

1. In small mixed-ability groups, prepare soap using locally available materials (e.g., palm oil, coconut oil, wood ash or caustic soda), and discuss observed outcomes and suggestions for improving quality (e.g., texture, scent, lathering ability).
2. Guide learners to draw the generalised structure of a soap molecule (sodium or potassium salt of a fatty acid).
3. Outline and discuss the main steps in soap production: fat hydrolysis (saponification), separation, and purification.

Talk-for-Learning

1. Pair learners and ask them to compare soapy detergents (e.g., traditional soap) with soapless detergents (e.g., synthetic detergents), focusing on their chemical structure and grease-removal mechanisms.
2. Have pairs discuss and list the advantages of soapless detergents over soapy ones (e.g., effectiveness in hard water, less residue).
3. Facilitate a class-wide discussion where pairs share insights and clarify misconceptions related to detergent action and environmental impact.

Key Assessment Level 3: Individual Classwork

1. Which of the following best explains why oils remain liquid at room temperature while fats are solid?
 - A. Oils have more carbon atoms than fats that helps them to slide over each other
 - B. Oils are composed of longer fatty acid chains which decreases their melting point
 - C. Oils have a higher proportion of unsaturated fatty acids that prevent tight molecular packing
 - D. Oils contain glycerol which makes them flow instead of fatty acids

2. A student is given two unknown lipid samples—one solid and one liquid at room temperature. Which method would best help the student determine the degree of unsaturation?
 - A. Melting point test
 - B. Solubility in water
 - C. Iodine number test
 - D. Reaction with sodium carbonate
3. Which reaction pathway accurately describes the saponification of a triglyceride?
 - A. $\text{alcohol} + \text{acid} \rightarrow \text{Triglyceride} + \text{soap}$
 - B. $\text{Triglyceride} + \text{enzyme} \rightarrow \text{ketone} + \text{ester}$
 - C. $\text{Triglyceride} + \text{water} \rightarrow \text{alcohol} + \text{soap}$
 - D. $\text{Triglyceride} + \text{base} \rightarrow \text{glycerol} + \text{soap}$
4. A food scientist hydrogenates vegetable oil. What structural and physical changes should they expect in the final product?
 - A. Unsaturated fats become saturated and the oil solidifies
 - B. Formation of ketones and increased rancidity
 - C. Addition of double bonds and increased solubility in water
 - D. Production of aldehydes and decrease in melting point
5. During a lab test, a student observes a pungent smell from a stored oil sample. The most likely explanation is:
 - A. The oil was completely hydrogenated during storage
 - B. The oil contains a high number of iodine atoms
 - C. The oil underwent oxidation causing rancidity
 - D. The oil was mixed and emulsified with soap
6. Which structural feature of fatty acids determines whether the resulting triglyceride is solid or liquid at room temperature?
 - A. Number of glycerol units
 - B. Presence of ester groups
 - C. Length of the hydrocarbon chain
 - D. Degree of saturation in the hydrocarbon chains
7. Which of the following best explains why fats and oils float when added to water?
 - A. They are polar molecules with strong hydrogen bonding
 - B. Their relative density is greater than that of water
 - C. Their long hydrocarbon chains make them hydrophilic
 - D. They have lower relative density due to non-polar nature

8. A chemistry student wants to prepare soap using local methods. Which pair of materials should they combine to initiate the saponification process?

A. Glycerol and vinegar
B. Coconut oil and sodium chloride
C. Palm kernel oil and ash solution
D. Water and essential oil
9. What process is used to produce biodiesel from fats and oils?

A. Hydrogenation
B. Hydrolysis
C. Transesterification
D. Saponification
10. Why are fats and oils considered poor conductors of heat and electricity?

A. They contain high levels of water.
B. They are composed of non-polar molecules.
C. They have low molecular weight.
D. They contain a high number of double bonds.

PLC Session 22

45 minutes	Planning & Reviewing the Learning Plan
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1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to

a. explain the week’s concepts in a practical way that will make it easier for learners to understand
b. identify two or three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to

a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.

- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

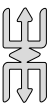
4. Plan the week’s key assessment with your app

a. The key assessment for the week is **individual classwork**. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the classwork (NTS 3k-3q).

10 minutes	Content Exploration
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5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then

a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
b. all teachers in the group work through a similar problem individually for 3–4 minutes
c. the group then works through the problem together step by step
d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).

a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
b. After the enactment, ask for structured feedback from the group

i. was the concept explained in a way that a learner with no prior knowledge could follow?
ii. were the examples grounded in contexts familiar to Ghanaian learners?
iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
 - a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

Section 6 Review

This section reviews all the lessons taught for weeks 21–22. This section discussed alkanolic acid derivatives and lipids, covering the structure, preparation, properties, and uses of alkyl alkanolates, amides, and acyl halides, as well as fats and oils. Learners explored reaction mechanisms, IUPAC nomenclature, synthetic pathways, and practical applications of these compounds. It is encouraged that portions of the sections which require a review of concepts be administered to learners as presentation tasks. Learners should be provided with adequate learning materials such as molecular model kits, laboratory reagents for esterification experiments, and digital resources for visualising reaction mechanisms to enable them to investigate the relationship between structure and properties of these compounds. Special attention should be given to the practical soap-making activity using local materials, which connects theoretical organic chemistry to everyday applications and provides opportunities for discussing the science behind cleaning agents and sustainable production methods.

MARKING SCHEMES FOR THE KEY ASSESSMENTS

WEEK 21

1. Identify the basic structures and preparation of amides and acyl halides.
 - a. Draw the general structure of an amide and an acyl halide. Circle and label the functional group in each. (4 marks)

For amide:
Correct structure with R-CO-NH₂ format (1 mark)
Correctly circled amide group (-CONH₂) (1 mark)

For acyl halide:
Correct structure with R-CO-X format (where X = halogen, typically Cl) (1 mark)
Correctly circled acyl halide group (-COX or -COCl) (1 mark).
 - b. Name the following compounds according to IUPAC rules: (5 marks)
 - i. CH₃CH₂CONH₂: Correct answer: Propanamide (1 mark)
Award ½ mark for propionamide (common name)

- ii. CH₃COCl, Correct answer: Ethanoyl chloride (2 marks)
Award 1 mark for acetyl chloride (common name)
- iii. C₆H₅CONH(CH₃), Correct answer: N-methylbenzamide (2 marks)
Award 1 mark for N-methyl benzamide (spacing error)
Award 1 mark for methylbenzamide (missing N- prefix)

- c. Write a balanced chemical equation for the preparation of ethanoyl chloride from ethanoic acid using thionyl chloride (SOCl₂). (4 marks)

Correct equation: CH₃COOH + SOCl₂ → CH₃COCl + SO₂ + HCl (4 marks)

Award marks as follows:

Correct reactants (1 mark), Correct products (1 mark), Correct formula for each compound (1 mark), Balanced equation (1 mark).

2. Analyse and connect concepts related to amides and acyl halides in real-world contexts.
 - a. Paracetamol is a common pain reliever that contains an amide group. Draw the structure of paracetamol, identify the amide group, and explain how its structure contributes to its function as a medication. (6 marks)

Structure and identification:

Correct structure of paracetamol (4-hydroxyacetanilide or N-(4-hydroxyphenyl)acetamide) (2 marks)

Correctly identified amide group (-NHCOCH₃) (1 mark)

Explanation of structure-function relationship (3 marks):

Mention of how the amide group contributes to stability in the body (1 mark)

Explanation of how the structure allows it to inhibit prostaglandin synthesis or COX enzymes (1 mark)

Discussion of how the structure allows it to cross the blood-brain barrier or reach its site of action (1 mark)

Alternative valid points:

Explanation of how hydrogen bonding affects solubility and absorption (1 mark)

Discussion of how the structure relates to metabolism in the liver (1 mark)

Comparison with related compounds like aspirin (1 mark).

- b. Design a synthetic route to prepare the local anesthetic benzocaine (ethyl 4-aminobenzoate) starting from 4-nitrobenzoic acid. Your synthesis must include an acyl chloride intermediate. Explain each step and the required reagents and conditions. (8 marks)

Synthetic route (4 marks):

Step 1: Conversion of 4-nitrobenzoic acid to 4-nitrobenzoyl chloride (acyl chloride intermediate) (1 mark)

Step 2: Conversion of 4-nitrobenzoyl chloride to ethyl 4-nitrobenzoate (1 mark)

Step 3: Reduction of nitro group to amine to form benzocaine (1 mark)

Logical sequence of steps (1 mark)

Reagents and conditions (4 marks):

Step 1: SOCl_2 or PCl_5 or $(\text{COCl})_2/\text{DMF}$, heat ($40\text{--}80^\circ\text{C}$) (1 mark)

Step 2: Ethanol, pyridine or base, $0\text{--}25^\circ\text{C}$ (1 mark)

Step 3: H_2/Pd or Fe/HCl or Sn/HCl or Na_2S , appropriate conditions (1 mark)

Correct conditions for each step (temperature, solvent, catalysts) (1 mark)

OR Alternative valid synthesis:

Step 1: Conversion of 4-nitrobenzoic acid to 4-nitrobenzoyl chloride (1 mark)

Step 2: Reduction of nitro group to amine to form 4-aminobenzoyl chloride (1 mark)

Step 3: Reaction with ethanol to form benzocaine (1 mark)

Logical sequence with appropriate reagents and conditions (1 mark)

WEEK 22

1. C
2. C
3. D
4. A
5. C
6. D
7. D
8. C
9. C
10. B

SECTION 7: POLYMERS

STRAND: CHEMISTRY OF CARBON COMPOUNDS

Sub-Strand: Organic Functional Groups

Learning Outcome: Explain the structure, preparation, properties of polymers and their uses in everyday life

Content Standard: Demonstrate knowledge and understanding of the concepts of organic chemistry to describe the structure and properties of polymers

INTRODUCTION AND SECTION SUMMARY

This section introduces learners to the fascinating world of polymers, exploring both natural and synthetic macromolecules that form the building blocks of life and countless modern materials. Learners will investigate the fundamental concept of polymerisation, distinguishing between addition and condensation mechanisms, before examining important natural polymers including proteins and carbohydrates as essential biomolecules. The section then transitions to synthetic polymers, where learners will explore the formation, properties, and applications of key industrial polymers including polythene, polyvinyl chloride, polytetrafluoroethene, nylon, and terylene. By the end of this section, learners will be able to classify polymers based on their structure and formation mechanisms, explain how monomer structure influences polymer properties, understand the relationship between molecular structure and material characteristics, evaluate the practical applications of various polymers in everyday contexts, and critically assess the environmental challenges posed by non-biodegradable plastics. This knowledge forms an essential foundation for understanding materials science, biochemistry, industrial chemistry, and the pressing environmental considerations surrounding plastic use and disposal in contemporary society.

The weeks covered by this section are:

- Weeks 23: Explain polymers and describe proteins and carbohydrates as natural polymers of amino acid and glucose respectively.
- Week 24: Explain the formation, properties and uses of synthetic polymers as well as their environmental impact.

SUMMARY OF PEDAGOGICAL EXEMPLARS

This section requires hands-on activities that immerse learners in the structure, preparation, properties, and everyday applications of polymers. Learners should explore concepts through collaborative learning—modelling amino acid structures, discussing protein formation, and testing for biomolecules. In mixed-ability groups, learners brainstorm and classify carbohydrates, describe the properties of proteins and carbohydrates, and relate structure to function. Experiential learning should focus on testing for reducing sugars and starch, emphasising observation, recording, and explanation of results.

Talk-for-learning strategies should encourage oral explanations of polymerisation types and classification of carbohydrates, deepening conceptual understanding. Digital learning tasks, such as internet-based research and group presentations on synthetic and natural polymers, polymer structures, and plastic pollution, offer a platform for learner-led discovery.

All learners should be supported to participate actively. Approaching proficiency learners may use diagrams and make oral presentations, while high-performing learners extend their learning to equations of polymer formation, protein structures, and the implications of polymer use in environmental sustainability. This inclusive, inquiry-driven pedagogy promotes both knowledge and real-world problem-solving.

ASSESSMENT SUMMARY

Indicator	DOK Level	Assessment Strategy and/or Mode
Explain polymers and describe proteins and carbohydrates as natural polymers of amino acid and glucose respectively.	Level 3: Strategic Reasoning	Individual Classwork
Explain the formation, properties and uses of synthetic polymers as well as their environmental impact.	Level 3: Strategic Reasoning	Group Research

WEEK 23

Learning Indicator: Explain polymers and describe proteins and carbohydrates as natural polymers of amino acid and glucose respectively

Recap weeks

Year 1	Week 20: Describe the types of intermolecular forces and explain how they arise from the structural features of molecules
Year 2	Week 19: Explain hybridisation and relate it to various structures Week 20: Apply the concept of hybridisation to explain sigma and pi bond of organic compounds and explain their structure and bonding Week 23: Explain the structure, nomenclature, preparation, properties and uses of alkanols as well as their characteristics test Week 24: Explain the structure, nomenclature, preparation, properties and uses of alkanolic acids as well as their characteristic test.

FOCAL AREA 1: NATURAL POLYMERS

Polymers are large molecules composed of many repeating subunits called monomers that are chemically bonded together in a chain-like structure. The term “polymer” comes from the Greek words “poly” meaning many and “meros” meaning parts. These macromolecules can be visualised as long chains or networks made up of smaller, identical building blocks linked together through chemical bonds.

The process of forming polymers, called polymerisation, occurs when monomers react with each other to form covalent bonds. This can happen through different mechanisms:

1. Addition Polymerisation (Chain-Growth Polymerisation)

Addition polymerisation involves the sequential addition of monomer units to a growing polymer chain without the elimination of any by-products. This mechanism typically involves unsaturated monomers containing carbon-carbon double bonds ($C = C$) or triple bonds ($C \equiv C$).

Characteristics of Addition Polymerisation

1. Rapid Reaction: Chain growth occurs quickly once initiated
2. High Molecular Weight: Achieved early in the reaction
3. Polymer Composition: Identical to monomer composition (no loss of atoms)
4. Reaction Conditions: Often requires precise temperature control due to exothermic nature
5. Active Centres: Growth occurs only at chain ends with active centres
6. Monomer Requirements: Typically requires unsaturated monomers

2. Condensation Polymerisation (Step-Growth Polymerisation)

Condensation polymerisation involves the reaction between functional groups on different monomers, resulting in the formation of a polymer with the elimination of small molecules

such as water (H₂O), hydrogen chloride (HCl), or ammonia (NH₃). This mechanism typically requires bifunctional or multifunctional monomers.

Characteristics of Condensation Polymerisation

- 1. Gradual Growth: Molecular weight increases slowly throughout the reaction
- 2. By-product Formation: Small molecules are eliminated during bond formation
- 3. High Conversion Requirement: High conversion (>99%) needed for high molecular weight
- 4. Polymer Composition: Different from monomer composition due to elimination of small molecules
- 5. Functional Group Preservation: Reactive groups remain at chain ends
- 6. Monomer Requirements: Requires monomers with two or more functional groups.

Polymers can be classified based on their source, natural polymers, derived from living organisms (proteins, carbohydrates, nucleic acids) and synthetic polymers manufactured artificially (plastics, synthetic fibres, synthetic rubbers)

The properties of polymers depend on several factors including the type of monomers, the arrangement of these monomers, the length of the polymer chains, and the interactions between chains. These factors determine characteristics such as strength, flexibility, solubility, and thermal behaviour.

Table 1.1: Comparison of Addition and Condensation Polymerisation

Feature	Addition polymerisation	Condensation polymerisation
Monomers	Usually one type with double bonds	Two or more types with functional groups
By-products	none	Small molecules (H ₂ O, HCl, etc.)
Reaction Rate	Rapid	gradual
Molecular Weight Growth	Rapid	gradual
Active Sites	Only at chain ends with active centres	At all functional groups
Reaction Time	Short	Longer
examples	Polyethylene, PVC, polystyrene	Polyesters, polyamides, polyurethanes

Proteins

Proteins are biological polymers made up of amino acid monomers linked together by peptide bonds. These remarkable macromolecules serve as the molecular machinery of living cells, performing diverse functions including catalysis, transport, structural support, movement, defence, and regulation.

Amino Acids (The Monomers of Proteins)

Amino acids are organic compounds that serve as the building blocks of proteins. Each amino acid contains:

- 1. A central carbon atom (α-carbon)
- 2. An amino group (-NH₂)
- 3. A carboxyl group (-COOH)
- 4. A hydrogen atom
- 5. A variable side chain (R group) that gives each amino acid its unique properties.

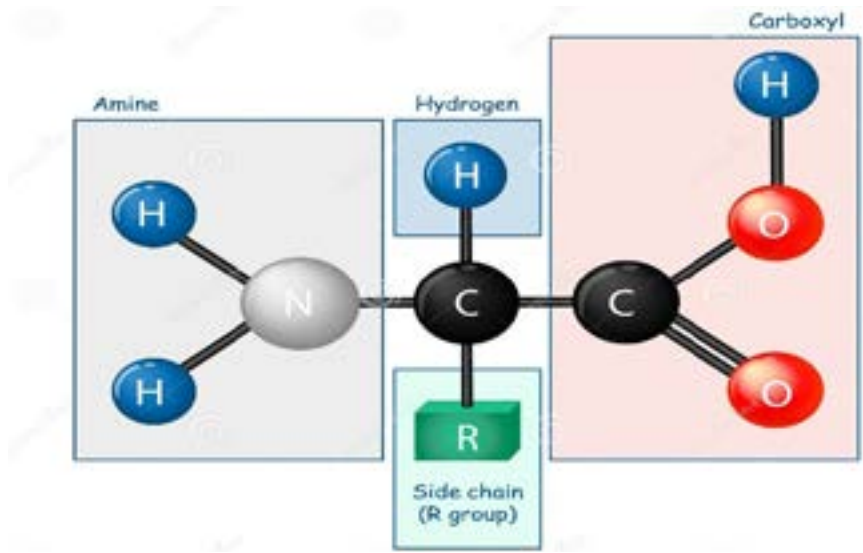
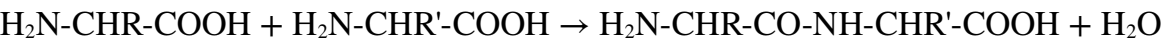


Figure 23.1: Structure of amino acid

There are 20 standard amino acids used by cells to synthesise proteins, each with distinct chemical properties determined by their side chains. These range from hydrophobic (water-repelling) to hydrophilic (water-attracting), acidic to basic, and small to large.

Protein formation through peptide bonds

During protein synthesis, amino acids polymerise through a condensation reaction where the carboxyl group of one amino acid reacts with the amino group of another, forming a peptide bond with the elimination of a water molecule:



This process continues as more amino acids are added to the growing chain, always proceeding from the N-terminus (amino end) to the C-terminus (carboxyl end). The specific sequence of amino acids in a protein, known as its primary structure, is determined by the genetic code.

The diversity of protein functions in living organisms stems from the almost limitless combinations possible with 20 different amino acid monomers. From the keratin in your hair to the enzymes that digest your food, proteins demonstrate how the specific arrangement of simple building blocks can create macromolecules with highly specialised functions.

Carbohydrates

Carbohydrates are organic compounds composed of carbon, hydrogen, and oxygen atoms, typically with the general formula C_x(H₂O)_y. They are one of the four major classes of biomolecules (along with proteins, lipids, and nucleic acids) and serve as the primary source of

energy for most organisms. The name “carbohydrate” comes from their chemical composition, which can be viewed as “carbon with water” (carbon-hydrate).

Classification of Carbohydrates

Carbohydrates are classified into four main groups based on their complexity:

1. **Monosaccharides:** Simple sugars that cannot be further hydrolysed; include glucose, fructose, and galactose; general formula $(\text{CH}_2\text{O})_n$ where $n = 3-7$.
2. **Disaccharides:** Formed by two monosaccharides joined by a glycosidic bond; examples include sucrose (glucose + fructose), lactose (glucose + galactose), and maltose (glucose + glucose).
3. **Oligosaccharides:** Short chains of 3-10 monosaccharide units; often found attached to proteins or lipids on cell surfaces; play important roles in cell recognition.
4. **Polysaccharides:** Long polymers containing hundreds to thousands of monosaccharide units; serve as energy storage (starch, glycogen) or structural components (cellulose, chitin).

Glucose (The Primary Monomer of Carbohydrates)

Glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) is a simple sugar or monosaccharide that serves as the primary energy source for many organisms. Its structure includes:

1. Six carbon atoms arranged in a ring (typically)
2. Multiple hydroxyl (-OH) groups that make it soluble in water
3. An aldehyde group (in its open-chain form).

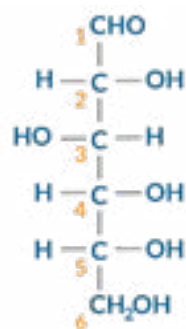


Figure 23.2: Structure of glucose

In its cyclic form, glucose can exist as either α -glucose or β -glucose, depending on the orientation of the hydroxyl group at carbon-1 (the anomeric carbon). This distinction becomes crucial in polysaccharide formation.

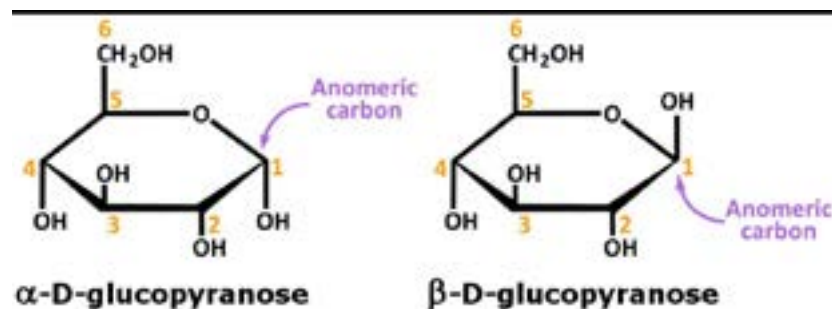
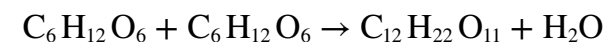


Figure 23.3: Structures of α and β glucose

Glycosidic Bonds: Linking Glucose Monomers

Glucose molecules polymerise through condensation reactions that form glycosidic bonds. These bonds connect the anomeric carbon of one glucose molecule to a hydroxyl group of another glucose molecule, with the elimination of a water molecule:



The specific type of glycosidic bond formed (α or β) and the carbon atoms involved in the linkage determine the structural and functional properties of the resulting polysaccharide.

Major glucose polymers and their functions

Several important polysaccharides are composed primarily of glucose:

1. **Starch:** A storage polysaccharide in plants, consisting of:
 2. **Amylose:** Linear chains of glucose linked by α -1,4-glycosidic bonds
 3. **Amylopectin:** Branched chains with α -1,4-glycosidic bonds and α -1,6 branches at branch points
 4. **Glycogen:** The storage polysaccharide in animals, similar to amylopectin but more highly branched, allowing for rapid mobilisation of glucose when energy is needed.
 5. **Cellulose:** A structural polysaccharide in plant cell walls, composed of glucose units linked by β -1,4-glycosidic bonds. These bonds create a linear, unbranched structure that allows cellulose chains to form hydrogen bonds with adjacent chains, creating strong fibres that provide structural support to plants.
 6. **Chitin:** A structural polysaccharide found in fungal cell walls and arthropod exoskeletons, composed of modified glucose units (N-acetylglucosamine) linked by β -1,4-glycosidic bonds.

Structural diversity from a common monomer

The remarkable diversity in structure and function among glucose polymers demonstrates how variations in the polymerisation process, specifically, the type of glycosidic linkage (α vs. β) and the pattern of branching can produce polymers with dramatically different properties from the same monomer:

- a. Starch and glycogen are readily digestible energy storage molecules
- b. Cellulose forms strong, indigestible fibres that provide structural support
- c. Chitin creates tough, flexible protective coverings

This illustrates a fundamental principle in polymer chemistry; the properties of a polymer depend not only on the monomers used but also on how these monomers are connected.

Learning Tasks

1. Create a concept map that illustrates the relationship between:
 - a. Polymers and monomers
 - b. Natural and synthetic polymers
 - c. Specific examples of each type.
2. Draw the general structure of an amino acid, labelling:

- a. The amino group
 - b. The carboxyl group
 - c. The R group (side chain)
 - d. The alpha carbon.
3. Write a short essay comparing and contrasting proteins and carbohydrates as natural polymers. Your essay should address:
 - a. The monomers that make up each polymer
 - b. The type of polymerisation involved in their formation
 - c. How their structures contribute to their diverse functions
 - d. Specific examples of each polymer and their roles in biological systems.

Pedagogical Exemplars

Collaborative Learning

1. Begin the lesson by displaying molecular models of common amino acids and proteins. Explain that today's collaborative activity will explore the building blocks of life through hands-on modelling and discussion.
2. Organise learners into mixed-ability groups of 3-4. Distribute molecular modelling kits, drawing materials, and reference sheets to each group.
3. Have each group model and draw the molecular structures of assigned amino acids (glycine, alanine, serine, cysteine, and lysine). Instruct learners to identify the key functional groups on each amino acid (amino group, carboxyl group, and unique side chain).
4. Then, direct learners to apply IUPAC nomenclature rules to name each modelled amino acid. Have them create a reference chart showing both the structures and systematic names.
5. Next, demonstrate how to model the formation of a dipeptide using two amino acids. Have learners identify the peptide bond and water as a by-product, emphasising this as an example of condensation polymerisation.
6. After this, facilitate a discussion on protein primary structure. Have learners explain how the sequence of amino acids determines protein function and create a simple diagram showing a polypeptide chain.
7. Then, introduce protein hydrolysis by having learners model the reverse process of peptide bond formation. Discuss the three types of hydrolysis (acid, base, and enzymatic) and their applications.
8. Next, present the common tests for proteins (Biuret, Ninhydrin, and Xanthoproteic tests). Explain the chemical basis for each test and what positive results look like.
9. Then, transition to carbohydrates by having learner's model glucose structure and identify its functional groups. Discuss how these molecules form the basis of various carbohydrates.
10. Finally, have each group compile and present their findings, focusing on one aspect of amino acids, proteins, or carbohydrates. Encourage other groups to ask questions and provide feedback.

Experiential Learning: Testing Carbohydrates

Lesson Introduction

Begin the lesson by showing learners samples of different sugars and carbohydrate-rich foods. Explain that today they will investigate the chemical properties of carbohydrates through hands-on testing.

1. Begin with a safety briefing, demonstrating proper handling of reagents and distributing safety equipment (goggles, gloves). Emphasise the importance of careful observation and recording.
2. First, guide learners in preparing solutions of different carbohydrates: glucose, fructose, sucrose, and starch. Label each solution clearly.
3. Then, demonstrate the Fehling's test procedure. Have learners add Fehling's solution to each sample, heat in a water bath, and record observations. Explain that a brick-red precipitate indicates a reducing sugar.
4. Next, have learners perform Benedict's test on the same samples. Instruct them to add Benedict's solution, heat in a water bath, and observe colour changes from blue to green/yellow/red depending on the concentration of reducing sugar.
5. Then, carefully guide learners through the Tollens' test. Demonstrate how to prepare Tollens' reagent safely, add it to each sample, and warm gently to observe for silver mirror formation.
6. After completing the tests, direct learners to write balanced equations for the reactions observed and identify the functional group responsible (aldehyde group). Discuss why some sugars test positive and others negative.
7. Next, have learners complete a results table showing test outcomes for each carbohydrate and draw conclusions about their reducing properties.
8. Finally, lead a reflection discussion on sources of error, the effectiveness of different tests, and how these observations relate to carbohydrate structures.

Talk-for-Learning: Carbohydrate Classification

Lesson Introduction

Begin the lesson by displaying various food items rich in different types of carbohydrates (fruits, bread, pasta, honey). Explain that today's discussion will focus on classifying carbohydrates and understanding their properties.

1. Begin with a think-pair-share activity. Have learners individually list examples of carbohydrates they know, then pair up to compare lists, and finally share findings with the class.
2. Then, guide a classification discussion. Define the criteria for classifying carbohydrates as monosaccharides (glucose, fructose), disaccharides (sucrose, maltose, lactose), and polysaccharides (starch, cellulose, glycogen). Create a classification chart with examples.
3. Next, focus on starch as a glucose polymer. Explain its components (amylose and amylopectin), the α -1,4 and α -1,6 glycosidic linkages, and the difference between linear and branched structures. Demonstrate the iodine test for starch (blue-black colour).

- 4. Then, explain the hydrolysis process of complex carbohydrates. Discuss how polysaccharides break down to disaccharides and eventually to monosaccharides through acid hydrolysis, enzymatic hydrolysis, and base hydrolysis.
- 5. Next, organise a brief debate on carbohydrate digestion. Assign different groups to represent different carbohydrates and have them explain how their carbohydrate is processed during digestion. Discuss why humans can digest starch but not cellulose.
- 6. After this, facilitate the creation of a collaborative concept map showing carbohydrate classifications, structural relationships, hydrolysis pathways, and testing methods.
- 7. Finally, conclude with a summary where each learner shares one key insight from the discussion. Address any misconceptions or questions and connect the discussion to everyday examples of carbohydrates in food and nutrition.

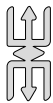
Key Assessment Level 3: Individual Classwork

- 1. Define the term “polymer” and explain what a monomer is. (2 marks)
- 2. Name the specific monomers that make up each of the following polymers: (2 marks)
 - a. Proteins
 - b. Carbohydrates (specifically starch and cellulose)
- 3. Compare and contrast addition polymerisation and condensation polymerisation. Which type occurs in the formation of proteins and carbohydrates? Explain your answer. (3 marks)
- 4. A learner claims that “all natural polymers are formed through condensation reactions.” Evaluate this claim using your knowledge of protein and carbohydrate formation. Provide specific chemical evidence to support your evaluation. (3 marks)

PLC Session 23

45 minutes	Planning & Reviewing the Learning Plan
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- 1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
- 2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week’s concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week’s content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
- 3. Use your subject specific app to plan your lesson. Ask the app to
 - a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
 - b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
 - c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



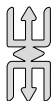
Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

- 4. Plan the week’s key assessment with your app
 - a. The key assessment for the week is **individual classwork**. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the classwork (NTS 3k-3q).

10 minutes	Content Exploration
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- 5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
 - a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step
 - d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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- 6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
 - a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.
 - b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
7.	Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8.	Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9.	Remember to - transfer your chats into your learning planner template. - read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a). - integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

WEEK 24

Learning Indicator: Explain the formation, properties and uses of synthetic polymers as well as their environmental impact

Recap weeks

Year 1	Week 20: Describe the types of intermolecular forces and explain how they arise from the structural features of molecules
Year 2	Week 19: Explain hybridisation and relate it to various structures Week 20: Apply the concept of hybridisation to explain sigma and pi bond of organic compounds and explain their structure and bonding Week 23: Explain the structure, nomenclature, preparation, properties and uses of alkanols as well as their characteristics test Week 24: Explain the structure, nomenclature, preparation, properties and uses of alcanoic acids as well as their characteristic test.

FOCAL AREA 1: SYNTHETIC POLYMERS

A synthetic polymer is a large molecule composed of many repeating structural units (monomers) that has been artificially created through chemical reactions.

Most synthetic polymers are created from petroleum-based chemicals (derived from oil and natural gas), though some newer ones use plant-based materials.

Scientists can control exactly how these polymers are made by adjusting:

1. The type of monomer used
2. How the monomers connect
3. The length of the chains
4. Whether the chains branch or cross-link.

This control allows manufacturers to create materials with very specific properties.

They can be hard or soft, flexible or rigid, transparent or opaque, water-resistant or water-absorbent.

Unlike many natural polymers (like cotton or wool), most synthetic polymers do not break down easily in nature, which creates environmental challenges with plastic pollution.

Classification synthetic polymers

1. By structure, polymers may be linear (consisting of a single continuous chain), branched (having side chains attached to the main chain), cross-linked (where different chains are connected by covalent bonds), or network (extensively cross-linked in three dimensions).
2. By thermal behaviour, polymers are categorised as thermoplastics (which soften when heated and can be remoulded), thermosets (which form irreversible chemical bonds when cured), or elastomers (rubber-like materials that can stretch and return to their original shape).

- By reaction mechanism, polymers are primarily divided into addition polymers and condensation polymers.

The properties of synthetic polymers depend on their molecular structure, molecular weight, degree of crystallinity, and intermolecular forces. These properties include mechanical strength, elasticity, thermal stability, chemical resistance, optical characteristics, and electrical properties. The diversity of these properties explains why polymers have found applications in virtually every industry.

Classification of synthetic polymers base on reaction mechanism

1. Addition Polymerisation

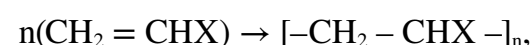
Addition polymerisation, also known as chain-growth polymerisation, is a process in which unsaturated monomer molecules add to one another to form a polymer without the elimination of any smaller molecules. The process typically involves monomers containing carbon-carbon double bonds.

Mechanism of Addition Polymerisation

The addition polymerisation process generally occurs in three distinct stages:

- Initiation:** This first stage involves the creation of an active centre, typically a free radical, cation, or anion. In free radical polymerisation, for example, an initiator molecule decomposes to form free radicals that attack the pi (π) bond of a monomer molecule.
- Propagation:** During this stage, the active centre reacts with another monomer molecule, adding it to the growing chain while transferring the active centre to the newly added unit. This process repeats rapidly, adding many monomer units to the chain in a short time.
- Termination:** The growth of the polymer chain stops when the active centre is destroyed. This can occur through various mechanisms such as combination (two growing chains joining together) or disproportionation (hydrogen transfer between two growing chains).

The general equation for addition polymerisation can be represented as:



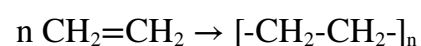
where X represents a substituent group that varies depending on the specific monomer.

Examples of Addition Polymers

Polyethylene

Polyethene (polyethylene) is formed through addition polymerisation of ethene (C_2H_4) under specific conditions. The reaction involves free radical polymerisation or catalytic polymerisation.

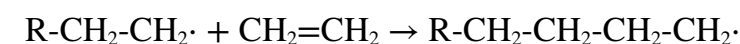
The polymerisation reaction occurs when ethene (C_2H_4) molecules undergo repeated bonding to form a long-chain polymer.



Reaction mechanism

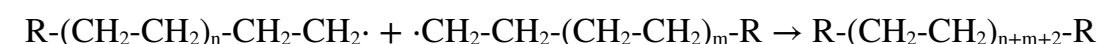
- Initiation:** A free radical initiator ($\text{R}\cdot$) reacts with ethene:

$$\text{R}\cdot + \text{CH}_2=\text{CH}_2 \rightarrow \text{R}-\text{CH}_2-\text{CH}_2\cdot$$
- Propagation:** The growing chain reacts with another ethene molecule:



This step repeats thousands of times

- Termination:** Two free radicals combine, stopping the reaction:



Properties of Polyethene:

- It has excellent chemical resistance to acids, bases, and most aqueous solutions.
- It is inert to most chemicals at room temperature.
- It is susceptible to oxidation at higher temperatures.
- It is not biodegradable.
- It is resistant to water absorption.

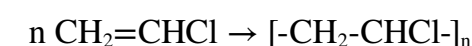
Uses of Polythene (Polyethylene)

Polythene is one of the most widely used synthetic polymers, known for its lightweight, durability, flexibility, and water resistance. It is used in various industries for practical applications, for example:

- packaging industry for plastic bags, food containers, wraps, and bottles.
- food storage for their moisture resistance.
- construction and infrastructure for pipes and fittings (especially HDPE) in plumbing systems.
- water tanks and insulation materials.
- protective coatings for cables and wires.
- the automotive industry for lightweight materials in fuel tanks, seats, and interior components.
- protective coverings for cables and wiring, etc.

Polyvinyl Chloride (PVC).

Polyvinyl chloride is produced by the addition polymerisation of vinyl chloride monomer (VCM).



Mechanism of polyvinyl chloride formation

- Initiator step:** (I) decomposes to form free radicals ($\text{I}\cdot$)

$$\text{I}\cdot + \text{CH}_2=\text{CHCl} \rightarrow \text{I}-\text{CH}_2-\text{CHCl}\cdot$$
- Propagation step:**

$$\text{I}-\text{CH}_2-\text{CHCl}\cdot + \text{CH}_2=\text{CHCl} \rightarrow \text{I}-\text{CH}_2-\text{CHCl}-\text{CH}_2-\text{CHCl}\cdot$$

This step repeats many times
- Termination step:**

Occurs by combination or disproportionation of growing chains

Properties of Polyvinyl Chloride

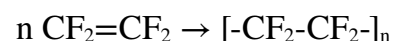
1. It has excellent resistance to acids, bases, and many solvents.
2. It is vulnerable to certain organic solvents like ketones and chlorinated hydrocarbons.
3. It is susceptible to degradation by heat and UV light, requiring stabilisers.
4. It is inherently flame-retardant due to chlorine content (self-extinguishing).
5. It is not biodegradable.

Uses of Polyvinyl Chloride

1. It is used for building materials such as window frames, doors, siding, and rain gutters.
2. It is used for pipes and fittings for water supply, drainage, and sewage.
3. It is used for electrical conduit and junction boxes.
4. It is used for credit cards and other plastic cards.
5. It is used for blister packaging for pharmaceuticals.
6. It is used for bottles for non-food applications.

Polytetrafluoroethylene (PTFE)

Polytetrafluoroethylene, commonly known by the brand name Teflon, is produced by the addition polymerisation of tetrafluoroethene monomers.



Mechanism of Polytetrafluoroethylene formation

1. Initiation: Persulphate decomposes to form sulphate radical anions
 $\text{SO}_4^{2-} \rightarrow \text{SO}_4^{\cdot -} + \text{CF}_2=\text{CF}_2 \rightarrow \text{SO}_4-\text{CF}_2-\text{CF}_2^{\cdot}$
2. Propagation:
 $\text{SO}_4-\text{CF}_2-\text{CF}_2^{\cdot} + \text{CF}_2=\text{CF}_2 \rightarrow \text{SO}_4-\text{CF}_2-\text{CF}_2-\text{CF}_2-\text{CF}_2^{\cdot}$
 This step repeats many times
3. Termination: Primarily by combination of growing chains

Properties of Polytetrafluoroethylene

1. It has exceptional chemical resistance to virtually all chemicals except molten alkali metals and elemental fluorine.
2. It has the highest resistance to acids, bases, and solvents among all plastics.
3. It is extremely hydrophobic surface (water contact angle $>110^\circ$).
4. It is highly inert and non-reactive.
5. It is not biodegradable.
6. It is resistant to weathering and UV radiation.

Uses of PTFE

1. It is used in industrial applications for non-stick coatings for cookware.
2. It is used as lubricants for mechanical parts.
3. It is used for lining for chemical storage tanks.
4. It is used in electrical applications for high-performance wire and cable insulation.
5. It is used for microwave components.
6. It is used in medical applications for vascular grafts.
7. It is used for surgical implants.
8. It is used for medical device components.
9. It is used for dental floss coating.

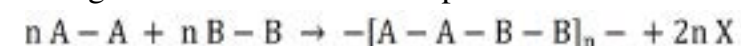
Condensation Polymerisation

Condensation polymerisation, also referred to as step-growth polymerisation, involves reactions between functional groups of monomers with the elimination of small molecules such as water, alcohol, or hydrogen chloride. Unlike addition polymerisation, condensation reactions typically involve two different functional groups that react with each other.

Mechanism of Condensation Polymerisation

In condensation polymerisation, bifunctional or polyfunctional monomers react to form dimers, which can further react with monomers or other oligomers to form progressively larger molecules. The reaction proceeds in a step-wise manner, with each step being essentially the same chemical reaction.

The general reaction can be represented as:



where A–A and B–B represent difunctional monomers with complementary functional groups, and X represents the small molecule eliminated during the reaction.

Unlike addition polymerisation, molecular weight increases gradually throughout the reaction.

Types of Condensation Polymers

Polyesters

Polyesters are synthetic polymers formed through condensation polymerisation of a dicarboxylic acid and a diol (two hydroxyl groups). The reaction eliminates water as a by-product, leading to polymer formation.

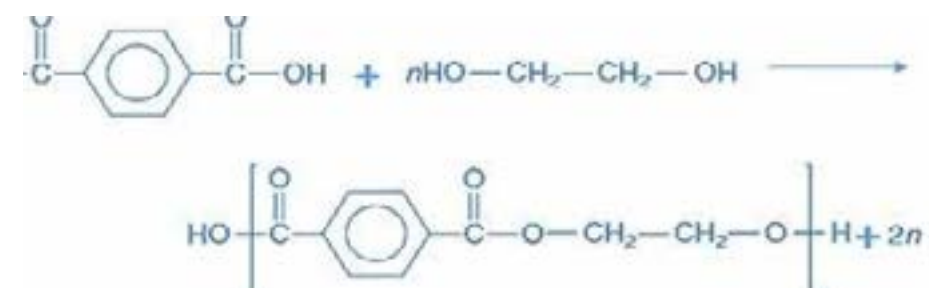


Figure 24.1: Formation of Polyesters through condensation polymerisation

Properties of Polyesters

1. They are resistant to wear and tear.
2. They are resistant to acids, oils, and solvents.
3. They can withstand moderate temperatures.
4. They are lightweight, making them ideal for textiles and packaging.
5. They are non-biodegradable.

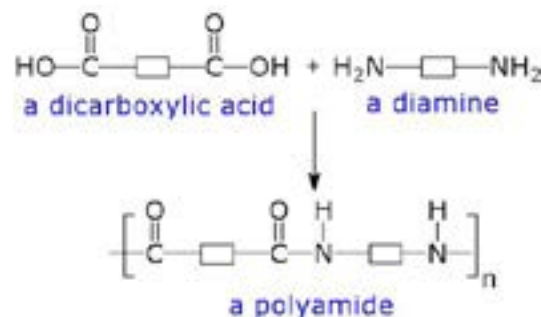
Uses of Polyesters

Polyesters are widely used due to their durability and versatility:

1. They are used in textiles for clothing, sportswear, and industrial fabrics.
2. They are used in packaging for PET bottles, containers, and food wraps.
3. They are used in the automotive industry for seat belts, airbag fabrics, and interior materials.
4. They are used in construction for fibreglass-reinforced roofing and insulation.
5. They are used in medicine for artificial heart valves and surgical sutures.
6. They are used in electronics for insulating tapes and lightweight casings.

Polyamides

Polyamides, exemplified by nylon, are synthesised through condensation polymerisation in which difunctional monomers such as dicarboxylic acids and diamines react by forming amide linkages with the elimination of water molecules, creating polymer chains with regularly spaced amide groups that enable hydrogen bonding between adjacent strands.



Polyamides come in various forms, each tailored to specific applications and performance requirements, with PA 6-6 and PA 6 being the most common types widely used in everyday products due to their excellent balance of strength, flexibility, and durability that makes them suitable for a broad range of uses.

Properties of polyamides

1. They have a low density, making them very lightweight materials that are easy to produce, develop, and install.
2. They have exceptional flexural strength, enabling them to bend and adhere to any place and improving assembly in areas where rigid metal structures cannot fit.
3. They remain flexible even after prolonged exposure to high temperatures.
4. They have high tensile strength and impact strength compared to many other polymers.
5. They have a wide temperature range, with aliphatic polyamides having melting temperatures around 500-550K, giving exceptional heat tolerance.

6. They have exceptional resistance to wear and abrasion, ensuring extensive durability in all contact areas.
7. They provide excellent corrosion resistance against caustic compounds, making them appropriate for harsh environments like chemical processing plants, refineries, and wastewater treatment plants.

Uses of polyamides

1. They are used in textiles for clothing, stockings, sportswear, and swimwear.
2. They are used in home furnishings for carpets, upholstery, and curtains.
3. They are used in outdoor equipment for parachutes, tents, and climbing ropes.
4. They are used in automotive applications for tire cords, airbags, and engine components.
5. They are used in mechanical components for gears, bearings, bushings, and cable ties.
6. They are used in electrical applications for wire insulation and cable jacketing.
7. They are used in sporting goods for fishing lines, tennis racquet strings, and bicycle wheels.
8. They are used in personal care products for toothbrush bristles and hairbrushes.

Pollution problems caused by non-biodegradable plastics

1. They persist in the environment for hundreds to thousands of years due to their non-biodegradable nature.
2. They accumulate in landfills, taking up valuable space and potentially leaching additives into soil and groundwater.
3. They contribute to marine pollution, with an estimated 8 million metric tons entering oceans annually.
4. They can entangle and harm wildlife, with marine animals particularly vulnerable to entanglement in plastic waste.
5. They can be ingested by animals who mistake them for food, causing intestinal blockages and starvation.
6. They contribute to greenhouse gas emissions during production, with the plastic industry accounting for approximately 4-8% of annual global oil consumption.
7. They can release harmful chemicals when burned in open air, contributing to air pollution.
8. They can block waterways and drainage systems, worsening flooding in urban areas.
9. They can serve as breeding grounds for mosquitoes and other disease vectors when they collect water.

Learning Tasks

1. Collect at least 5 different plastic items commonly found in Ghanaian households (water sachets, nylon ropes, bottles, containers, etc.).

For each item:

- a. Identify the type of polymer using recycling codes or physical tests.

- b. Record the physical properties (flexibility, transparency, density) and note its common uses in Ghana.
- c. Create a chart that classifies these polymers as either addition or condensation polymers.
- d. Draw the monomer structure for each polymer and write the chemical equation showing how the polymer forms.
2. Research and create an information sheet explaining:
 - a. Why certain polymers are non-biodegradable (relate to their chemical structure).
 - b. The typical degradation time for each polymer type.
 - c. The specific environmental hazards each polymer presents in Ghana's climate.
 - d. Create a visual display showing the polymer waste cycle in Ghana, from production to disposal.

Pedagogical Exemplars

Digital Learning

1. Begin by dividing learners into four research groups focusing on: (a) polymerisation types, (b) synthetic vs. natural polymers, (c) addition polymers from alkenes, and (d) condensation polymers (nylon and terylene).
2. Schedule a guided computer lab session where learners research their topics using provided reliable sources and templates. In case of internet unavailability, provide relevant textbooks and printed resources for each group.
3. Circulate during research to ensure learners correctly understand key concepts like monomer structures, reaction mechanisms, and polymer properties.
4. Teach necessary PowerPoint skills, emphasising how to properly represent chemical structures and polymerisation reactions.
5. Have groups submit draft presentations for feedback, focusing on scientific accuracy of polymer formation processes.
6. Organise presentation sessions where each group explains their polymer topic to classmates, using molecular models or animations to illustrate concepts.
7. Synthesise all presentations by creating a class reference chart comparing addition and condensation polymers, their properties, and applications in Ghana.

Collaborative Learning

1. Begin with a whole class brainstorming session to define plastics, using actual samples of different polymers for learners to examine.
2. Have small groups investigate specific plastic types (thermoplastics, thermosets, and elastomers) and create informational posters about their properties.
3. Show a brief video on plastic pollution in Ghana's water bodies and urban areas to contextualise environmental concerns. If video equipment is unavailable, use newspaper articles and photographs.

4. Conduct a "Think-Pair-Share" activity where learners discuss plastic pollution they've observed in their communities.
5. Guide learners to collaboratively compile a comprehensive list of environmental problems caused by non-biodegradable plastics in Ghana.
6. Have groups debate the advantages and disadvantages of different solutions to plastic pollution, considering both technological approaches and policy changes.
7. Conclude with learners creating informational pamphlets about responsible plastic use and disposal for distribution in their community.

Key Assessment Level 3: Group Research

1. The Ghana Water Company Limited is considering replacing aging metal pipes with plastic ones to improve water distribution in urban areas while reducing costs and preventing corrosion. They have narrowed their choices to PVC and HDPE pipes.
 - a. Identify the monomers used to produce each of these polymers.
 - b. Compare the chemical and physical properties of PVC and HDPE, explaining how these properties relate to their molecular structures.
 - c. Evaluate which polymer would be most suitable for Ghana's water supply system, considering factors such as resistance to Ghana's tropical climate, durability during the dry and rainy seasons.
2. The accumulation of plastic waste in Ghana's water bodies, including the Volta Lake and coastal areas, has become a significant environmental concern.
 - a. Describe the difference between addition and condensation polymerisation, giving one example of a polymer produced by each method.
 - b. Explain why most synthetic polymers found in Ghana's environment (such as sachet water bags and food packaging) are non-biodegradable, relating this to their chemical structure.
 - c. Design an investigation to compare the properties of conventional plastic bags used for "pure water" sachets with a potential biodegradable alternative that could be manufactured locally in Ghana.

PLC Session 24

45 minutes

Planning & Reviewing the Learning Plan

1. Begin the session by recapping lessons from the previous year/session/week, focusing on how learning was applied.
2. Use your subject specific app to explore the content before planning. Ask the app to
 - a. explain the week's concepts in a practical way that will make it easier for learners to understand
 - b. identify two or three common misconceptions learners may have on the week's content and how to address them using contexts familiar to Ghanaian learners. (NTS 3a, 3g and 3m).
3. Use your subject specific app to plan your lesson. Ask the app to

- a. suggest the most effective method for teaching the week’s concepts to learners with diverse backgrounds.
- b. suggest learning activities, differentiation and essential questions based on the method of teaching (NTS 3a-3q)
- c. review your prepared learning plan and address any challenging issues you may have identified in your subject area (NTS 3a-3c).



Note

- Take note of the recap topics linked to Years 1 and 2 while preparing for your lesson.
- Refer to the Values Learning Community Handbook and link the values for the week in your lesson to ensure that learners are applying what they learn.

4. Plan the week’s key assessment with your app
- a. The key assessment for the week is **group research**. Chat with your app to plan this assessment, indicating what you and your learners will do before, during and after the research (NTS 3k-3q).

10 minutes	Content Exploration
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5. As a subject group, select one problem or idea/term in the week’s content that is challenging (not a routine procedural question, but one that requires conceptual understanding) and then
- a. use the app, ask it to solve the problem step by step and explain the reasoning behind each step (you may be assisted by a teacher who is strong in the content area)
 - b. all teachers in the group work through a similar problem individually for 3–4 minutes
 - c. the group then works through the problem together step by step
 - d. highlight the most common errors made at each step and discuss (NTS 1c, 1e, 2c, 2e, 3e-3j).



Note

- This part is not just about ‘How do I teach this?’ but how do I teach it for learners to understand?
- If the group still shares the same content gap, use the app: ask it to solve the problem step by step and explain the reasoning behind each step.
- Problems/ideas discussed should be sourced from the Teacher Manual content, not invented separately. This keeps the content exploration directly connected to what teachers will deliver.

30 minutes	Enactment
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6. Enact a teaching activity based on your learning plan (NTS 1a-1b, 1e-1g, 2c, 2e, 2f, 3c-3j and 3o).
- a. Colleagues observe with specific attention to clarity of concept explanation, use of familiar Ghanaian examples and how misconceptions are handled if they arise.

- b. After the enactment, ask for structured feedback from the group
 - i. was the concept explained in a way that a learner with no prior knowledge could follow?
 - ii. were the examples grounded in contexts familiar to Ghanaian learners?
 - iii. did any misconceptions surface and if so, how were they addressed?

5 minutes	Reflection
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7. Reflect and share your views on the session. One thing that worked well, what did not work well and what you will do differently. (NTS 1a, 1b).
8. Identify a critical friend to observe and provide feedback on your lesson in relation to the PLC Session conducted (NTS 1f, 1g).
9. Remember to
- a. transfer your chats into your learning planner template.
 - b. read your Teacher Manual and related Learner Material in preparation for the next session. Take note of the recap weeks, which are linked to Years 1 and 2 (NTS 3a).
 - c. integrate National Values, GESI, SEL, and 21st century skills into your learning activities, assessment tasks, and rubrics/mark schemes.

Section 7 Review

This section reviews all the lessons taught for weeks 23–24. This section discussed polymers, covering both natural polymers (proteins and carbohydrates) and synthetic polymers (polythene, PVC, PTFE, nylon, and terylene). Learners explored polymerisation mechanisms, molecular structures of amino acids, protein formation, carbohydrate classification, and the environmental impact of non-biodegradable plastics. It is encouraged that portions of the sections which require a review of concepts be administered to learners as presentation tasks. Learners should be provided with adequate learning materials such as molecular model kits, chemical reagents for testing procedures (Fehling’s solution, Benedict’s solution, Tollen’s reagent), and digital resources for visualising polymer structures to enable them to investigate the relationship between monomer structure and polymer properties, and to critically evaluate the benefits and challenges of plastics in modern society.

MARKING SCHEMES FOR THE KEY ASSESSMENTS

WEEK 23

1. Define the term “polymer” and explain what a monomer is. (2 marks)
- a. 1 mark: Polymer: A large molecule (macromolecule) composed of many repeating subunits that are joined by covalent bonds.
 - b. 1 mark: Monomer: A small, reactive molecule that can join with other identical molecules to form a polymer; the basic building block or subunit of a polymer.
 - c. 1 mark: Partial definition of either polymer or monomer with some inaccuracy or incompleteness.

2. Name the specific monomers that make up each of the following polymers: proteins and carbohydrates (specifically starch and cellulose) 2 marks
 - a. 1 mark: Proteins: Amino acids
 - b. 1 mark: Carbohydrates (starch and cellulose): Glucose (or α -D-glucose for starch and β -D-glucose for cellulose)
 - c. 1 mark: Correct identification of only one monomer type.
3. Compare and contrast addition polymerisation and condensation polymerisation. Which type occurs in the formation of proteins and carbohydrates? Explain your answer. (3 marks)
 - a. 1 mark: Clear comparison of addition polymerisation (monomers add together without loss of atoms/no by-products) versus condensation polymerisation (monomers join with elimination of small molecules like water).
 - b. 1 mark: Correct identification that condensation polymerisation occurs in the formation of both proteins and carbohydrates.
 - c. 1 mark: Explanation that includes specific chemical evidence: In protein formation, the carboxyl group of one amino acid reacts with the amino group of another, eliminating water to form a peptide bond. In carbohydrate formation, hydroxyl groups on glucose molecules react, eliminating water to form glycosidic bonds.
 - d. 2 marks: Correct comparison of polymerisation types and identification of condensation polymerisation, but incomplete explanation.
4. A learner claims that “all natural polymers are formed through condensation reactions.” Evaluate this claim using your knowledge of protein and carbohydrate formation. Provide specific chemical evidence to support your evaluation. (3 marks)
 - a. 1 mark: Clear evaluation of the claim, noting that while many important natural polymers (including proteins and carbohydrates) are formed through condensation polymerisation, the claim is not universally true for all natural polymers.
 - b. 1 mark: Specific chemical evidence for condensation in protein formation: The peptide bond forms between the carboxyl group ($-\text{COOH}$) of one amino acid and the amino group ($-\text{NH}_2$) of another, releasing a water molecule (H_2O).
 - c. 1 mark: Specific chemical evidence for condensation in carbohydrate formation: The glycosidic bond forms between the hydroxyl groups ($-\text{OH}$) of glucose molecules with the elimination of water (H_2O).
 - d. 2 marks: Correct evaluation with specific evidence for only one polymer type, OR Partial evaluation with some evidence for both polymer types.
 - e. 1 mark: Basic evaluation of the claim without specific chemical evidence, OR Some relevant chemical information but no clear evaluation.
 - f. 0 marks: No correct evaluation or irrelevant answer.

WEEK 24

Question 1

- a. Identify the monomers used to produce each of these polymers. (4 marks)

- i. PVC: Vinyl chloride/chloroethene ($\text{CH}_2=\text{CHCl}$) (2 marks)
- ii. HDPE: Ethene/ethylene ($\text{CH}_2=\text{CH}_2$) (2 marks)
- b. Compare the chemical and physical properties of PVC and HDPE, explaining how these properties relate to their molecular structures. (8 marks)

Comparison of properties (4 marks, 1 mark for each valid comparison of properties):

 - Density: PVC (1.3 to 1.45 g/cm^3) is denser than HDPE (0.94 to 0.97 g/cm^3)
 - Chemical resistance: PVC has excellent resistance to acids and bases but is vulnerable to some organic solvents; HDPE has excellent chemical resistance to most solutions
 - Flexibility: PVC can be rigid or flexible (with plasticisers); HDPE is more rigid.
 - Thermal properties: PVC decomposes around 180°C ; HDPE melts at 130 - 137°C without decomposition
 - Strength: PVC has higher tensile strength than HDPE
 - UV resistance: PVC degrades under UV light; HDPE has better UV resistance

Relation to molecular structure (4 marks):

 - PVC: The presence of chlorine atoms creates polarity in the molecule, leading to stronger intermolecular forces and higher density/strength (2 marks)
 - HDPE: The linear structure with minimal branching allows chains to pack closely, increasing crystallinity and strength, while the non-polar C-C and C-H bonds contribute to chemical resistance (2 marks)
- c. Evaluate which polymer would be most suitable for Ghana’s water supply system, considering factors such as resistance to Ghana’s tropical climate, durability during the dry and rainy seasons. (8 marks)
 - i. Evaluation of suitability (6 marks):

(Climate considerations)

 - i. Effect of high temperatures (1 mark)
 - ii. UV exposure during dry season (1 mark)
 - iii. Humidity and rainfall effects (1 mark)

(Durability considerations)

 - Resistance to soil conditions in Ghana (1 mark)
 - Mechanical strength requirements (1 mark)
 - Expected lifetime in local conditions (1 mark)
 - iii. Recommendation with justification (2 marks):

(Clear recommendation of either PVC or HDPE) (1 mark)

Scientific justification based on polymer properties and Ghana’s specific conditions (1 mark)

Accept well-reasoned arguments for either polymer if properly justified. Most likely recommendation, e.g.

3. HDPE may be preferred for its better resistance to UV radiation, flexibility to handle ground movement, and lack of additives that might leach in hot conditions
4. PVC might be preferred for its strength, lower cost, and established usage history

Question 2

1. Description: Process where monomers with double bonds add together without the loss of any atoms/small molecules (1 mark)
2. Mechanism: Involves initiation, propagation, and termination steps (1 mark)
3. Example: Any correct example such as polyethylene, polypropylene, PVC, polystyrene (1 mark)

e. Condensation polymerisation (3 marks):

1. Description: Process where monomers combine with the elimination of small molecules (usually water) (1 mark)
2. Mechanism: Reaction between functional groups on different monomers (1 mark)
3. Example: Any correct example such as nylon, polyester, bakelite, polyurethane (1 mark)

- e. Explain why most synthetic polymers found in Ghana's environment (such as sachet water bags and food packaging) are non-biodegradable, relating this to their chemical structure. (6 marks)

Explanation of non-biodegradability (6 marks):

1. Carbon-carbon backbone is resistant to microbial attack (1 mark)
2. Lack of functional groups that can be targeted by enzymes (1 mark)
3. High molecular weight prevents cellular uptake by microorganisms (1 mark)
4. Hydrophobic nature repels water and water-soluble enzymes (1 mark)
5. Specific example relating to sachet water bags (polyethylene) structure (1 mark)
6. Specific example relating to food packaging (could be PP, PS, or PET) structure (1 mark)

- f. Design an investigation to compare the properties of conventional plastic bags used for "pure water" sachets with a potential biodegradable alternative that could be manufactured locally in Ghana. (8 marks)

a. Investigation design (8 marks):

1. Clear statement of aim and hypothesis (1 mark)
2. Identification of variables (independent, dependent, controlled) (1 mark)
3. Materials and equipment required (1 mark)
4. Step-by-step procedure that is replicable (2 marks)
5. Methods to test relevant properties (tensile strength, water permeability, biodegradability) (2 marks)
6. Safety considerations and risk assessment (1 mark)

NB: Award full marks for a comprehensive, well-designed investigation that specifically addresses the Ghanaian context and is feasible with locally available resources.