MINISTRY OF EDUCATION

CHEMISTRY
for Senior High Schools

TEACHER MANUAL

Year 2

NaCCA
NATIONAL COUNCIL FOR
CURRICULUM & ASSESSMENT
OF MINISTRY OF EDUCATION

MINISTRY OF EDUCATION

REPUBLIC OF GHANA

Chemistry

for Senior High Schools

Teacher Manual

Year Two

NaCCA
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 a new Senior High School (SHS) curriculum which aims to ensure that all learners achieve their potential by equipping them with 21st Century skills, competencies, character qualities and shared Ghanaian values. This will prepare learners to live a responsible adult life, further their education and enter the world of work.

This is the first time that Ghana has developed an SHS Curriculum which focuses on national values, attempting to educate a generation of Ghanaian youth who are proud of our country and can contribute effectively to its development.

This Teacher Manual for Chemistry is a single reference document which covers all aspects of the content, pedagogy, teaching and learning resources and assessment required to effectively teach Year Two of the new curriculum. It contains information for all 24 weeks of Year Two including the nine key assessments required for the Student Transcript Portal (STP).

Thank you for your continued efforts in teaching our children to become responsible citizens.

It is our belief that, if implemented effectively, this new curriculum will go a long way to transforming our Senior High Schools and developing Ghana so that we become a proud, prosperous and values-driven nation where our people are our greatest national asset.

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SECTION 1: ENERGY CHANGES

Strand: Physical Chemistry

Sub-Strand: Matter and its properties

Learning Outcome: *Describe the different energy changes that occurs as a result of reorganisation of atoms.*

Content Standard: Demonstrate knowledge and understanding that energy changes occur in chemical reactions as a result of reorganisation of atoms at different rates.

Hint

- ***Assign Portfolio Assessment*** *for the academic year by Week 3. Portfolio should be submitted by Week 22.* ***See Appendix A*** *of this Section and Teacher Assessment Manual and Toolkit pages 22-25 for more information on how to organise a portfolio assessment.*
- *Assign* ***Group Project*** *in Week 2. This should be submitted after week 10. Refer to Teacher Assessment Manual and Toolkit pages 27-29 for how to conduct Project-Based Assessment. Refer to* ***Appendix B*** *which has been provided at the end of the section for the structure and rubrics of the group project.*

INTRODUCTION AND SECTION SUMMARY

This section spans from weeks one to three. In this section learners will be taken through the complexities of enthalpy change, standard enthalpy changes, experimental determination of enthalpy change, and Hess's law of constant heat summation. In addition, strong emphasis will be placed on developing essential 21st-century skills such as critical thinking, problem-solving, and communication. Utilisation of digital learning tools will enhance the learning experience and provide a platform for interactive engagement. Furthermore, gender equality and social inclusion will be actively promoted ensuring that every learner feels valued and empowered. By using an interdisciplinary approach, learners will not only master key chemical concepts but also see the interconnectedness of these concepts with other disciplines such as Physics

The following are the weeks to be covered in the section

Week 1**:** *Standard Enthalpy Changes*

Week 2: *Calorimetry*

Week 3: *Hess's Law*

SUMMARY OF PEDAGOGICAL EXEMPLARS

In delivering the lessons on enthalpy change, standard enthalpy changes, experimental determination of enthalpy change, and Hess's law of constant heat summation, a variety of pedagogical approaches will be employed. The lessons will involve engaging demonstrations to visually illustrate key concepts, fostering collaborative learning where learners work together to deepen their understanding. Activity-based learning exercises will provide hands-on experiences to solidify learning outcomes, while "talk for learning" discussions will encourage active participation and communication amongst learners. Enquiry-based learning will prompt critical thinking and exploration of real-world applications. Additionally, digital learning tools will be integrated to enhance learners' engagement and provide interactive learning opportunities. Through these diverse pedagogical methods, learners will be actively involved in the learning process, ensuring a comprehensive and impactful experience.

ASSESSMENT SUMMARY

This summary outlines the types of assessments to be conducted, how they will be given, and what data will be recorded in the transcript for continuous assessment purposes. To facilitate this, a range of formative assessment methods will be employed, including short answers, oral and written presentations, lab reports, manipulation of equipment, safety procedure processes, measurements, demonstrations, and homework assignments. Projects and group activities tailored to cover levels one, two, three, and four of the Depth of Knowledge (DOK) framework will assess learners' comprehension and application of concepts while developing critical thinking, communication, and creativity. Class exercises, lab reports, measurements, and group tasks during and after each lesson will allow learners to demonstrate their grasp of concepts, with homework assignments offering practical application opportunities. Practical group activities such as equipment manipulation, demonstrations, and measurements will also be utilised for assessment purposes.

WEEK 1

Learning Indicator: *Explain the various standard enthalpy changes that occur and carry out calculations to determine them*

FOCAL AREA 1: ENTHALPY CHANGE

A system is the part of the Universe under study. A boundary separates a system from its surroundings (part of the Universe outside the system). Based on the nature of interactions a system has with its surroundings, there are three types of systems; open, closed and isolated. An open system exchanges matter and heat with its surroundings, example heating water in an open beaker. A closed system allows for the exchange of heat but not matter with its surroundings, example is heating water in a covered beaker. Isolated system does not allow for the exchange of matter or heat with its surroundings, an example of this is the well-lagged calorimeter.

To measure the energy changes that occur in chemical reactions, chemists usually use a related thermodynamic quantity called enthalpy (H). The enthalpy of a system is defined as the sum of its internal energy (U) and the product of its pressure (P) and volume (V).

Mathematically, $H = U + PV$

Because internal energy, pressure, and volume are all state functions, enthalpy is also a state function. A state function is a property of a system that depends on its present state or whose value does not depend on the path taken to reach that specific value. At constant-pressure, enthalpy is equal to the amount of heat transferred to the system in the process. The enthalpy change of a chemical reaction can be calculated from the enthalpy changes of formation of all products and reactants. Enthalpy change (ΔH) is therefore defined as:

$$\Delta \boldsymbol{H} = \boldsymbol{H}_{products} - \boldsymbol{H}_{reactants}$$

Endothermic and Exothermic reactions

In the study of energy changes that accompanies chemical reactions, the most important quantity is the enthalpy of reaction (ΔH_{rxn}). When a system releases heat to its surroundings, the enthalpy of the system decreases, so ΔH_{rxn} is negative. Also, if heat flows from the surroundings into the system, the enthalpy of the system increases, so ΔH_{rxn} is positive.

An endothermic reaction is a reaction in which heat is absorbed from the surrounding into the system. In endothermic reactions, the total energy of the system is greater than the total energy of the surroundings. The difference is the heat taken into the system from the surroundings.

$\Delta H = positive$

Example: Decomposition of calcium carbonate

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ $\Delta H = +177.8$ kJ.

An exothermic reaction is a reaction in which heat is released from the system to the surroundings. The system ends up with a lower heat content than the surrounding.

$\Delta H = negative$

Example: Combustion of methane

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\Delta H = -890.4$ kJ

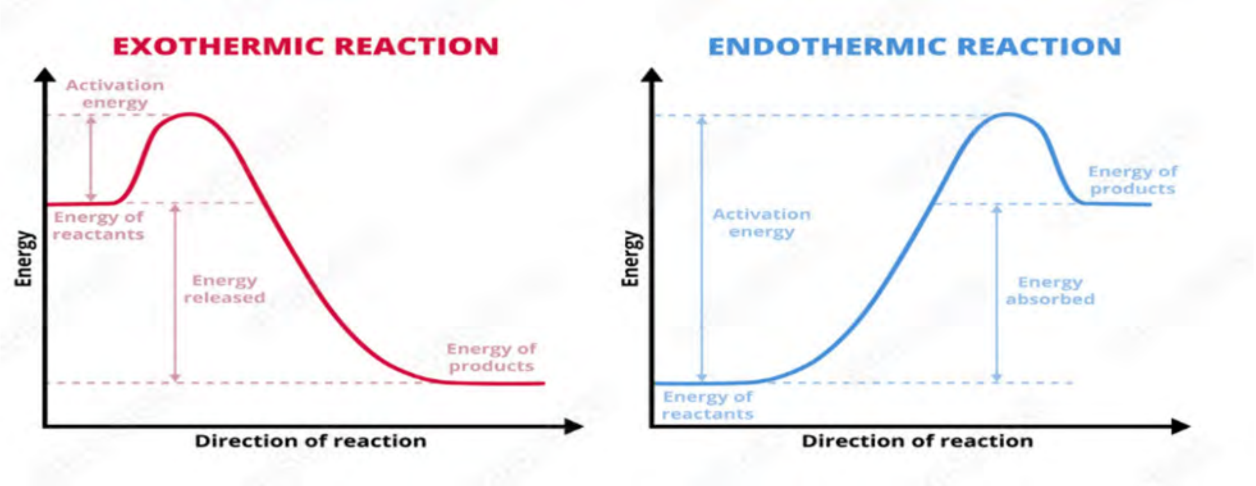


Figure 1.1: *Energy profile diagram for exothermic and endothermic reactions*

Learning Tasks

1. Define what a system is in the context of thermodynamics.
2. Discuss how enthalpy changes during chemical reactions (exothermic vs. endothermic).
3. Differentiate between exothermic and endothermic reactions and provide real-life examples of each type of reaction.

PEDAGOGICAL EXEMPLARS

Demonstration

Introduce the concept of chemical systems and the terms surrounding, open, closed, and isolated systems.

1. Demonstrate with hot water in a bottle
   - a. Measure initial temperature.
   - b. Predict, observe and describe changes over time.
2. Discuss and explain terms using the demonstration
   - a. Chemical system: hot water in the bottle.

  b. Surrounding: air and materials around the bottle.
  c. Open, closed, and isolated systems.

3. Use a simple, relatable examples and encourage observation, description, and questioning.
4. Conclude and relate to real-life scenarios
5. Use real life chemical reactions which clearly show Exothermic reactions (combustion: https://edu.rsc.org/experiments/the-whoosh-bottle-demonstration/708.article).
6. Use real life chemical reactions which clearly show Endothermic reactions (https://edu.rsc.org/experiments/endothermic-solid-solid-reactions/739.article)

Collaborative learning

In small-mixed ability groupsHave groups discuss and explain the concepts chemical energy, enthalpy and enthalpy change during chemical reaction.

1. Encourage learners to question where the energy is coming from for an exothermic reaction.
2. Encourage learners to question where the energy is coming from for an endothermic reaction.
3. Encourage groups to share their understanding with the class.
4. Facilitate a class discussion to explain that chemical reactions are accompanied by enthalpy change that can be exothermic or endothermic and give examples.
5. Have learners watch video to reinforce the concept of endothermic and exothermic reactions.

Activity-based learning

1. Introduce the concept of enthalpy profile diagrams and their importance in representing energy changes in chemical reactions.
2. Provide a simple reaction example (e.g., combustion of methane) and ask learners to draw an energy profile diagram for an exothermic reaction.
3. Have learners label the diagram with the following: Reactants, products, activation energy (E_a), enthalpy change (ΔH), etc.
4. Repeat step 2 for an endothermic reaction (e.g., photosynthesis).
5. Allow learners to share their diagrams with peers and discuss any questions or clarifications.

6. Encourage learners to colour-code their diagrams to distinguish between exothermic and endothermic reactions.

KEY ASSESSMENTS

1. **Level 1:** Identify the basic difference between exothermic and endothermic reactions in terms of energy transfer.
2. **Level 2:** Describe a system in thermochemistry and differentiate between an open, closed, and isolated system in terms of energy exchange.
3. **Level 2:** Use the table to answer this question

Reaction	Starting temperature/ °C	Final temperature/ °C
X	22	37
Y	25	18
Z	21	25

Decide whether each reaction is endothermic or exothermic, explain how you could tell.

   a. Which reaction has the largest energy change?

4. **Level 3**
   a. Design an experiment to determine if a given reaction is exothermic or endothermic and interpret the observed energy changes.
   b. When hydrochloric acid reacts with ammonium hydroxide in a beaker, the temperature goes up. $HCl\ (aq) + NH_4OH\ (aq) \rightarrow NH_4Cl\ (s) + H_2O\ (l)\ \Delta H = –53.4 kJ/mol$
   c. Draw the energy profile diagram and state whether the reaction is endothermic or exothermic, explain your answer.

FOCAL AREA 2: STANDARD ENTHALPY CHANGES

The enthalpy change of a reaction depends on the reaction conditions. It is therefore necessary to specify standard conditions for the measurement of enthalpy changes. These are taken to be atmospheric pressure (1 atm or 101.3 kPa) and room temperature (298K or 25°C). Enthalpy changes measured under standard conditions are known as standard enthalpy changes. During these chemical changes, the pressure is kept constant (this may mean that for some reactions, especially those involving gases, the volume is allowed to change freely). The enthalpy change for a reaction is the energy change measured under conditions at constant pressure. The standard enthalpy change for a reaction is the energy change measured under standard conditions of 101.3 kPa and a temperature of 298K. The enthalpy change also depends on the mass (or moles) of substance used. It is therefore necessary to specify the mass (or moles) of reactants used. Enthalpy changes are conventionally measured in $kJmol^{-1}$. It is a measure of the heat exchange (positive or negative) during a chemical reaction under standard conditions. It is denoted by $\Delta H°$.

Standard enthalpy change includes

Standard State: The standard state refers to the most stable form of an element or compound at a defined temperature and pressure. For elements, the standard state is often the most common form at room temperature; for compounds, it is typically the pure substance in its most stable form.

Standard Conditions: Standard conditions are usually at a temperature of 25°C (298 K) and a pressure of 1 atm pressure (101.3 kPa).

Standard concentration: Standard concentration is often set at 1.0 mol/dm^3 (molar concentration) for aqueous solutions. This value simplifies calculations and comparisons, as it represents a concentration commonly used in laboratory settings. It makes it easier to apply stoichiometric ratios and coefficients to determine the quantities of reactants and products involved in a reaction.

A negative standard enthalpy change ($\Delta H° < 0$) indicates an exothermic reaction, where heat is released to the surroundings. A positive standard enthalpy change ($\Delta H° > 0$) denotes an endothermic reaction, where heat is taken in from the surroundings.

The standard enthalpy change can be calculated based on the difference in the standard enthalpies of the products and reactants. It is a useful thermodynamic quantity for comparing the energy changes of different reactions.

Standard enthalpy change plays a crucial role in thermochemistry, helping chemists understand and quantify the energy changes associated with chemical reactions. It provides valuable insights into the heat flow of reactions and is utilised in various aspects of chemical analysis and industry.

Types of standard enthalpy changes

1. Standard enthalpy change of reaction

It is the quantity of heat involved when any valid stoichiometric molar quantities of reactants combine to form products under standard conditions. Any heat change measured under standard conditions is described as a standard enthalpy change of reaction. It is dependent on the stoichiometry of the balanced chemical equation. Changes in the balanced coefficients of reactants and products in any chemical reaction will change the value of its enthalpy changes accordingly. It has a unit of J/mol (or kJ/mol).

For example

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) \qquad \Delta H^\circ_{rxn} = -92.38\ kJ/mol$$

$$\tfrac{1}{2}N_2(g) + \tfrac{3}{2}H_2(g) \rightarrow NH_3(g) \qquad \Delta H^\circ_{rxn} = -46.19\ kJ/mol$$

$$3N_2(g) + 9H_2(g) \rightarrow 6NH_3(g) \qquad \Delta H^\circ_{rxn} = -277.14\ kJ/mol$$

Notice here this is the same reaction, the standard enthalpy of reaction changes based upon how much of the substances react. The value is not constrained in any way.

2. Standard enthalpy change of formation

Enthalpy of formation is the heat change that accompanies the formation of a compound from its constituent elements at a specific temperature and pressure. The standard enthalpy change of formation of a compound is the enthalpy change that results when one mole of the compound is formed from its elements under standard conditions (298K and 101.3 kPa), all reactants and products being in their standard states. It is given the symbol ΔH°_f

E.g. $C(s) + 2H_2(g) \rightarrow CH_4(g)$,

ΔH°_f $(CH_4) = -74.8\ kJmol^{-1}$.

The standard enthalpy of formation of methane is $-74.8\ kJmol^{-1}$.

The standard enthalpy of formation of all elements in their standard states is **zero**.

Using oxygen as an example, it exists in the most stable form as a diatomic molecule (O_2) in the gaseous form as compared to the other allotropic forms such as ozone (O_3). As such $\Delta H^\circ_f\left(O_{2(g)}\right) =$ but $\Delta H^\circ_f\left(O_{3(g)}\right) = 142.2\ kJ/mol$ while $\Delta H^\circ_f\left(O_{(g)}\right) = 249.4\ kJ/mol$. The most stable form of carbon under standard conditions is graphite, therefore ΔH°_f ($C_{(graphite)}) = 0$.

The table below describes some elements in their standard state

Substance	ΔH°_f	**Substance**	ΔH°_f
$O_{2(g)}$	0	$C_{(graphite)}$	0
$N_{2(g)}$	0	$S_{(rhombic)}$	0
$H_{2(g)}$	0	$Ag_{(s)}$	0
$F_{2(g)}$	0	$Ca_{(s)}$	0
$I_{2(g)}$	0	$Cu_{(s)}$	0
$Cl_{2(g)}$	0	$Hg_{(l)}$	0

Standard enthalpies of formation values help in calculating the standard enthalpy of reaction (ΔH°_{rxn}), defined as the enthalpy of a reaction carried out under standard conditions. For example, consider the hypothetical reaction

$aA + bB \rightarrow cC + dD$

where a, b, c, and d are stoichiometric coefficients. For this reaction, ΔH°_{rxn} is given by

$$\Delta H^{\circ}_{rxn} = [c\Delta H^{\circ}_{f}\,(C) + d\Delta H^{\circ}_{f}\,(D)] - [a\Delta H^{\circ}_{f}\,(A) + b\Delta H^{\circ}_{f}\,(B)]$$

We can generalise

$$\Delta H^{\circ}_{rxn} = \sum n\,\Delta H^{\circ}_{f}\left(\text{products}\right) - \sum m\,\Delta H^{\circ}_{f}\left(\text{reactants}\right)$$

where m and n are the stoichiometric coefficients for the reactants and products, respectively,

and Σ (sigma) means "the sum of."

Standard enthalpy of formation values also give a measure of the stability of compounds upon their formation. Compounds with negative ΔH°_{f} values are more stable than compounds with positive ΔH°_{f} values.

3. Standard enthalpy change of combustion

Enthalpy of combustion is the quantity of heat given out when a substance combusts completely in excess oxygen at a specific temperature and pressure. The standard enthalpy of combustion of a substance is defined as the enthalpy change that occurs when one mole of a substance is combusted completely in oxygen under standard conditions. (298K and 101.3 kPa), all reactants and products being in their standard states.

It is given the symbol ΔH°_{c}

Eg $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$, $\Delta H = -285.8\ kJmol^{-1}$.

The standard enthalpy of combustion of hydrogen is -285.8 $kJmol^{-1}$.

Notice here that fractional values for stoichiometric coefficients can be used to balance the equation. This ensures only 1 mole of the substance is combusted.

Knowledge of ΔH°_{c} are applied in many fields including nutrition, industries, transportation and research.

Respiration is a biological combustion process whereby food nutrients are combusted in the body to liberate energy for daily activities. The quantity of energy produced per meal depend greatly on the nature of food eaten. Carbohydrates and fats are the main energy giving foods. In nutritional science, the quantity of energy these food substances provide are measured in Calories. Unused energy from carbohydrates are stored in the body as fats, excess production and storage of fat from over-consumption of carbohydrates can lead to obesity.

In industries enthalpy of combustion helps to determine the efficiency of power plants or engines by finding the quotient between work output and ΔH_c°.

It allows for comparison between different fossil fuels. Fuels with higher ΔH_c° values are generally more efficient (per kg) than those with lower values.

It is also used to estimate the difference in the energy content of allotropes.

4. Standard enthalpy change of neutralisation

The standard enthalpy change of neutralisation is the enthalpy change when solutions of an acid and an alkali react together under standard conditions to produce 1 mole of water. It is assigned the symbol ΔH_n°.

When one mole of a monobasic acid reacts with one mole of a monobasic base under standard conditions the amount of energy released is almost the same as the enthalpy of formation of water from its ions.

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} \qquad \Delta H = -\ 57.5$ kJ/mol

$HNO_{3(aq)} + KOH_{(aq)} \rightarrow KNO_{3(aq)} + H_2O_{(l)} \qquad \Delta H = -\ 57.3$ kJ/mol

Enthalpy of formation of water from its ions is represented by the equation;

$\mathrm{H}^+_{(aq)} + \mathrm{OH}^-_{(aq)} \rightarrow H_2O_{(l)} \qquad \Delta H = -\ 57.4$ kJ/mol

Strong acids and strong bases ionise completely in aqueous solution to give off hydrogen ions and hydroxyl ions therefore they tend to have higher enthalpies of neutralisation than the weak acid and weak bases (which do not fully ionise).

5. Standard enthalpy of solution and hydration.

The standard enthalpy of solution is the enthalpy change when one mole of an ionic substance is dissolved in excess water under standard conditions such that no further addition of water produces a further change in energy. It is assigned the symbol ΔH_{soln}°.

e.g., $\mathrm{NaCl\ (s)} + H_2O_{(l)} \rightarrow \mathrm{Na^+\ (aq)} + \mathrm{Cl^-\ (aq)} \qquad \Delta H = -\ 3.9$ kJ/mol

Standard enthalpy of hydration is the enthalpy change when one mole of gaseous ions is dissolved in water to give one mole of aqueous ions and a solution of infinite dilution. It can also be defined as the energy evolved when one mole of a gaseous ion is surrounded, coordinated and stabilised by a sheath of water molecules under standard conditions.

e.g., $\mathrm{Na^+\ (g)} + H_2O_{(l)} \rightarrow \mathrm{Na^+\ (aq)} \qquad 406$ kJ/mol

Learning Tasks

1. Identify the standard conditions for the determination of enthalpy changes in chemical reactions.

2. Write a balanced equation to represent the standard enthalpy of formation of the following compounds water, carbon dioxide and ammonia.
3. Identify instances where the standard enthalpy of reaction and formation are the same.

PEDAGOGICAL EXEMPLARS

Talk-for-learning

a. Introduce the concept of standard enthalpy changes.

b. Ask learners to share what they know about enthalpy and chemical reactions.

c. Define and explain each term:

  i. Standard enthalpy change of reaction ($\Delta H^\circ r$)

  ii. Standard enthalpy change of formation ($\Delta H^\circ f$)

  iii. Standard enthalpy change of combustion ($\Delta H^\circ c$)

  iv. Standard enthalpy change of neutralisation ($\Delta H^\circ neut$)

  v. Standard enthalpy change of solution ($\Delta H^\circ sol$)

  vi. Standard enthalpy change of hydration ($\Delta H^\circ hyd$)

d. Use examples to illustrate each concept and encourage learners to ask questions and clarify doubts.

e. Provide sample calculations for each type of standard enthalpy change.

f. Have learners work in pairs to practice calculations and circulate to assist and facilitate discussion.

g. Discuss the significance of standard enthalpy changes in various fields (e.g., chemistry, physics, engineering) and explore real-life examples and applications (e.g., energy production, materials science).

KEY ASSESSMENTS

Level 1: What is meant by the term standard enthalpy of combustion?

Level 2

1. Explain why the standard enthalpy of neutralisation of a weak monobasic is less than the standard enthalpy of neutralisation of a strong monobasic acid.
2. Discuss what happens to the temperature of warm water when it is poured into a closed system.

Hint

The recommended mode of assessment for week 1 is ***discussion****. Use the level 2 question 2 as a sample question.*

WEEK 2

Learning Indicator: *Determine the various enthalpy changes as well as the calorific value of common foods and fuels used in everyday life through calorimetry*

FOCAL AREA: EXPERIMENTAL DETERMINATION OF ENTHALPY CHANGE

Experimental determination of heat of combustion of alcohol

Procedure

1. Place a volume of ethanol in a lamp and dip a wick into it through the cork of the bottle.
2. Weigh the lamp together with the ethanol.
3. A pre-weighed mass of water is transferred into a tin can and placed above the lamp.
4. The initial temperature of the water in the can is measured and noted.
5. The lamp positioned below the can of water is lighted to heat the water.
6. The water is stirred frequently to evenly distribute the heat.
7. The rise in temperature of the water is recorded.
8. The flame is put out and the lamp with the remaining ethanol is reweighed.

Data

Initial mass of the lamp + alcohol = x g

Mass or volume of water in the can = z g

Initial temperature of the water in the can = a °C

Final temperature of the water in the can = b °C

Final mass of the lamp + alcohol = y g

Calculation

Note

The quantity of heat produced by to burn the wick is used to raised the z g of water from a temperature of a°C to b°C.

Assuming all the heat lost by the burning of alcohol is gained by the water to raise its temperature then, Heat lost by burning alcohol = Heat gained by water

- Q (alcohol) = Q (water)

But $Q = mc\Delta T$

m(water) = z g

c(water) = 4.184 $Jg^{-1}K^{-1}$

$\Delta T = (b - a)K$

$Q = z(4.184)(b - a)$

Therefore 4.184z(b – a) J is gained by the mass of water.

Mass of alcohol that produced that amount of energy = ($y - x$)

This mass ($y - x$) can be converted to moles by dividing with the molar mass of the fuel (alcohol).

Moles of alcohol burnt = $\frac{\textit{mass of fuel used}}{\textit{Molar mass of fuel}}$

Molar mass of alcohol (C_2H_5OH) = 46 g/mol = $\frac{y-x}{46}$

If $\left(\frac{y-x}{46}\right)$ mol of alcohol produces 4.184z(b – a) J of heat

$1 \text{ mol} = \frac{46}{y-x} \times 4.184z\left(b - a\right)$

Experimental Discussion Points

1. What is the significance of stirring the water regularly?
2. Compare the value gained from your experiment to the official laboratory value, what do you notice?
3. Why are the values different?
4. Calculate the % difference between the two values.
5. How could this experiment be improved to reduce the % difference between the values.
6. Would the results of the experiment be different if the room temperature was colder? Explain your reasoning.

Standard enthalpy of Neutralisation

It is the heat released when one mole of an acid is neutralised by an alkali to produce one mole of water under standard conditions.

Experiment Title: Determination of the Standard Enthalpy of Neutralisation for a Strong Acid-Base Reaction.

Objective: To measure the standard enthalpy change ($\Delta H°$) for the neutralisation reaction between a strong acid (e.g., hydrochloric acid) and a strong base (e.g., sodium hydroxide) under standard conditions.

Materials

Strong acid (e.g., 1M HCl)

Strong base (e.g., 1M NaOH)

Calorimeter

Thermometer

Graduated cylinders

Stirring rod

Distilled water

Pipettes

Procedure

1. Measure equal volumes of the strong acid and strong base solution (e.g., 50 mL) using graduated cylinders.
2. Record the initial temperatures of both solutions.
3. Place the acid solution in the calorimeter and record its temperature.
4. Add the base solution to the acid in the calorimeter and immediately start stirring.
5. Record the highest temperature reached during the reaction.
6. Calculate the temperature change (ΔT) based on the initial and final temperatures.
7. Repeat the experiment at least three times to ensure accuracy and consistency.
8. Take an average of the value to determine your Enthalpy change.

Calculate the standard enthalpy of neutralisation using the formula: $\Delta H° = q / n$, where q is the heat absorbed or released during the reaction, and n is the number of moles of the limiting reactant.

Calculations

1. Use the heat capacity of the calorimeter to correct for heat loss or gain to the surroundings.

2. Determine the number of moles of the reactants based on the concentration and volume used.
3. Calculate the enthalpy change per mole of water formed in the reaction.

Results and Analysis

Calculate the average standard enthalpy of neutralisation based on the replicates.

Compare the experimental value to the theoretical enthalpy change for the neutralisation of a strong acid and base (about -57 kJ/mol). Why are the values different? Calculate the % difference between the two values.

Discuss any sources of error and potential improvements in the experimental procedure.

1. What is the significance of stirring the water regularly?
2. How could this experiment be improved to reduce the % difference between the values.
3. Would the results of the experiment be different if the room temperature was colder? Explain your reasoning.

Conclusion

The experiment provides valuable insights into the energy changes associated with the neutralisation of a strong acid and base, leading to the determination of the standard enthalpy of neutralisation under controlled conditions. The obtained value can be compared to the theoretical value to assess the accuracy of the experimental method.

Standard enthalpy of solution

It is the heat absorbed or released when one mole of an ionic substance is dissolved in excess water such that no further addition of water produces further heat change under standard conditions.

Experiment Title: Determination of the Standard Enthalpy of Dissolution for a Solute

Objective: To determine the standard enthalpy change ($\Delta H°$) for the dissolution of a specific solute (e.g., sodium chloride) in water and calculate the heat energy associated with the dissolution process.

Materials

Solute (e.g., sodium chloride)

Distilled water

Calorimeter

Thermometer

Graduated cylinders

Stirring rod

Pipettes

Balance

Safety goggles and protective gear

Procedure

1. Measure a known mass of the solute (e.g., sodium chloride) using the balance and record the mass accurately.
2. Pour a measured volume of distilled water into the calorimeter and record the initial temperature of the water.
3. Add the solute to the water in the calorimeter and stir gently to ensure complete dissolution.
4. Monitor and record the highest temperature reached during the dissolution process.
5. Repeat the experiment with different solute masses to ensure reproducibility of results.
6. Calculate the heat absorbed or released during the dissolution process using the temperature change (ΔT), calorimeter constants, and specific heat capacity of water.
7. Determine the moles of solute dissolved based on the mass and molar mass of the solute.
8. Calculate the standard enthalpy of dissolution using the formula: $\Delta H^\circ = q / n$, where q is the heat absorbed or released during the process, and n is the number of moles of the solute.
9. Calculate the average value for the enthalpy of dissolution across your entire process.

Calculations

1. Correct for heat losses to the surroundings by using the heat capacity of the calorimeter.
2. Determine the number of moles of the solute dissolved in the water.

3. Calculate the standard enthalpy of dissolution per mole of the solute.

Results and Analysis

1. Calculate the average value of the standard enthalpy of dissolution based on replicates of the experiment.
2. Compare the experimental value with the theoretical or literature value for the specific solute. Why are the values different? Calculate the % difference between the two values.

Discuss any sources of error and potential improvements in the experimental procedure.

1. What is the significance of stirring the water regularly?
2. How could this experiment be improved to reduce the % difference between the values.
3. Would the results of the experiment be different if the room temperature was colder? Explain your reasoning.

Analyse the energy changes associated with the dissolution process and discuss the enthalpy of dissolution in relation to the solubility of the solute.

Conclusion: By conducting this experiment to determine the standard enthalpy of dissolution for a solute in water, students can gain practical insights into the thermodynamics of solution processes. Understanding the heat energy changes associated with dissolution provides valuable information for predicting solubility trends and quantifying the energetics of the dissolution process.

Learning Tasks

1. Draw an experimental set-up to determine the enthalpy change of combustion of a named fuel.
2. Draw an experimental set-up to determine the enthalpy change of neutralisation of a named reaction.
3. Draw an experimental set-up to determine the enthalpy change of solution of a named solute.

PEDAGOGICAL EXEMPLARS

1. **Inquiry-based learning:** In small mixed-ability groups, design and carry out circus experiments to determine:
    - **a.** enthalpy change of combustion of alcohols and food substances, such as maize or groundnut.
    - **b.** enthalpy of neutralisation of acid and base (HCl and NaOH).
    - **c.** enthalpy of solution (NH_4Cl and $CaCl_2$ solutes)
    - **a.** use the relationship, $\Delta H = mc\Delta T$ to calculate enthalpy change from the experimental results.
    - **b.** interpret the results and evaluate the experiment.
2. **Digital Learning**

    Watch video or use virtual laboratory to reinforce the measurement of energy content of food and fuels.

KEY ASSESSMENTS

Level 1 List five apparatus needed to investigate the standard heat of combustion of butane.

Level 4

Methanol, ethanol, propan – 1 – ol and butan-1-ol are members of alkanol series: Plan and design an experiment to determine the enthalpy of combustion of each under laboratory conditions. Your plans should include;

- **a.** Stating the hypothesis of the trend in ΔHc of the compounds.
- **b.** Stating the dependent and independent variables.
- **c.** Stating two variables to be controlled.
- **d.** Briefly describing the procedure
- **e.** Stating three sources of error for the experiment
- **f.** Outlining how these sources of error could be minimised.

Analyse the calorific values of 5 selected food groups: Fruits, Vegetables, Grains, Proteins, and fats/oils and 5 selected fuels: Gasoline, Coal, Diesel, Wood, and Natural gas. Compare and highlight their calorific value and energy density and potential applications of these values to the foods and the fuels.

Report your findings in the format below:

- **a.** Record and organise data in a table
- **b.** Collect data on the calorific values of foods and the fuels

c. Calculate the energy density of each food and fuel

d. Compare and contrast the calorific values of foods and fuels

e. Discuss the implications of energy density on food choice and fuel selection

f. A 3-page comprehensive report detailing the research, calculations and analysis and representation of the calorific values

g. A 10-minute presentation summarising key findings

Hint

*The recommended mode of assessment for week 2 is **project work (group)**. Use the level 4 question 2 as a sample question. Refer to Appendix B at the end of this section for further information on how to go about the project.*

WEEK 3

Learning Indicators

1. *Explain Hess's law of constant heat summation and apply it to construct energy cycle diagrams as well as calculate relevant enthalpy changes*
2. *Explain bond energy and apply it to determine endothermic and exothermic reactions*

FOCAL AREA 1: HESS'S LAW OF CONSTANT HEAT SUMMATION

Many compounds cannot be directly synthesised from their elements. In some cases, the reaction proceeds too slowly, or produce substances other than the desired compound. In these cases, ΔH_f° can be determined by an indirect approach, which is based on Hess's law of heat summation.

Hess's law states that the enthalpy change for a reaction that occurs in a series of steps is equal to the sum of the enthalpy changes of the individual steps. Hess's law is valid because enthalpy is a state function.

The enthalpy change can be measured by calorimetry for many, but not all, chemical processes. For a hypothethical reaction

$A + B \rightarrow D$

Whether the reaction moves directly into completion to form the stable product D or goes through an unstable intermediate to form C before it finally forms D, the overall energy for the reaction will always be the same.

$A + B \rightarrow C$ $\Delta H1$

$C + B \rightarrow D$ $\Delta H2$

Then $A + 2B \rightarrow D$ $\Delta H3$

$\Delta H3 = \Delta H1 + \Delta H2$

Consider, for example, the oxidation of carbon to form carbon monoxide.

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$

The primary product of the reaction of carbon and oxygen is CO_2, even if a deficiency of oxygen is used. As soon as CO is formed, it reacts with O_2 to form CO_2. Because the reaction cannot be carried out in a way that allows CO to be the sole product, it is not possible to measure the change in enthalpy for this reaction by calorimetry.

The enthalpy change for the reaction forming CO(g) from C(s) and $O_2(g)$ can be determined indirectly, however, from enthalpy changes for other reactions for which values of ΔH_{rxn}° can be measured.

The calculation is based on **Hess's law**, which states that if a reaction is the sum of two or more other reactions, ΔH°_{rxn} for the overall process is the sum of the ΔH°_{rxn} values of those reactions.

The oxidation of C(s) to $CO_2(g)$ can be viewed as occurring in two steps: first the oxidation of C(s) to CO(g) (Equation 1) and second the oxidation of CO(g) to $CO_2(g)$ (Equation 2). Adding these two equations gives the equation for the oxidation of C(s) to $CO_2(g)$ (Equation 3).

Equation 1: $C(s) + 1/2\ O_2(g) \rightarrow CO(g)$ $\Delta H^\circ_{rxn}1 = ?$

Equation 2: $CO(g) + 1/2\ O_2(g) \rightarrow CO_2(g)$ $\Delta H^\circ_{rxn}2 = -283.0$ kJ/mol

Equation 3: $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H^\circ_{rxn}3 = -393.5$ kJ/mol

Hess's law says that the enthalpy change for the overall reaction ($\Delta rH^\circ 3$) will equal the sum of the enthalpy changes for reactions 1 and 2 ($\Delta H^\circ_{rxn}1 + \Delta H^\circ_{rxn}2$). Both $\Delta H^\circ_{rxn}2$ and $\Delta H^\circ_{rxn}3$ can be measured, and these values are then used to calculate the enthalpy change for reaction 1.

$\Delta H^\circ_{rxn}3 = \Delta H^\circ_{rxn}1 + \Delta H^\circ_{rxn}2$

$-393.5\text{ kJ/mol} = \Delta H^\circ_{rxn}1 + (-283.0\text{ kJ/mol})$

$\Delta H^\circ_{rxn}1 = -110.5\text{ kJ/mol}$

Applications of Hess's law of constant heat summation

(i) This law provides a practical method for computing the enthalpy change of a reaction from the enthalpy changes of other reactions.

(ii) Hess's law is used to calculate the reaction enthalpies even for those reactions which are otherwise difficult to carry out in the laboratory.

(iii) According to Hess's law the thermochemical equations may be treated as algebraic equations.

Rules for Manipulating Thermochemical Equations

1. For any exothermic reaction the reverse is an endothermic reaction. Therefore, when an equation is reversed or written in the opposite direction, the sign of its enthalpy must also be reversed or negated. To illustrate, the reverse of the equation: $A + B \rightarrow C\ \Delta H = +x$ J/mol is the following equation: $C \rightarrow A + B\ \ \Delta H = -x$ J/mol
2. Substances can be cancelled from both sides of an equation only if the substance is in an identical physical state.
3. If all the coefficients of an equation are multiplied by a factor, the value of its enthalpy must be multiplied by the same factor.
4. To illustrate, the factors of an equation: $A + B \rightarrow C\ \Delta H = +x$ J/mol

When an equation is multiplied by 2: $2A + 2B \rightarrow 2C\ \Delta H = 2\ (+x\ J/mol)$

5. When equations are added, the enthalpy changes are also added accordingly with proper signs

$A + B \rightarrow C\ \Delta H1$

$C + B \rightarrow D \quad \Delta H_2$

Then $A + 2B \rightarrow D \quad \Delta H_3 = \Delta H_1 + \Delta H_2$

Example

Apply Hess's law to calculate the enthalpy change (ΔH°_{rxn}) of the following reaction:

C (graphite) + 2 H_2(g) → CH_4(g)

Data given:

(i) C (graphite) + O_2(g) → CO_2(g) $\quad \Delta H^\circ_f$ (i) = – 393.5 kJ mol^{-1}

(ii) H_2 (g) + $\frac{1}{2}$ O_2(g) → H_2O(l) $\quad \Delta H^\circ_f$ (ii) = – 285.8 kJ mol^{-1}

(iii) CH_4 (g) + 2 O_2(g) → CO_2(g) +2 H_2O(l) $\quad$ (iii) = – 890.3 kJ mol^{-1}

Solution

Our aim C (graphite) + 2 H_2(g) → CH_4(g) ΔH°_{rxn} (A) = ?

$\Delta H^\circ_{rxn} = \Delta H^\circ_f$(products) – ΔH°_f(reactants)

(ΔH°_{rxn}) = -890.3 kJ/mol

ΔH°_fproduct = ΔH°_f (CO_2) + 2 x ΔH°_f (H_20)

= -393.5 + 2(-285.8)

= - 965.1

ΔH°_freactants = x + 0

-890.3 = -965.1 – (x)

X = -965.1 + 890.3

= - 74.8 kJ/mol

Alternatively, using the rearrangement approach,

Maintain equation (i)

Multiply equation (ii) by 2

Reverse equation (iii) we get

(iv) $2H_2$ (g) + O_2(g) → $2H_2O$(l) $\quad \Delta H^\circ_{rxn}$ (iv) = – 2×285.8 kJ mol^{-1} = -571.6 kJ mol^{-1}

(v) CO_2(g) + H_2O(l) → CH_4 (g) + 2 O2(g) ΔH°_{rxn} (v) = + 890.3 kJ mol^{-1}

On adding equations (i), (iv) and (v) along with corresponding ΔH°_{rxn} values, equation (vi) can be obtained.

(vi) $[C\ (\text{graphite}) + O_2(g)] + [2H_2\ (g) + O_2(g)] + [CO_2(g) + 2\ H_2O(l)] \rightarrow CO_2\ (g) + 2\ H_2O(l) + [CH_4(g) + 2O_2(g)]$

$$\Delta H^{\circ}_{rxn}\ (vi) = \Delta H^{\circ}_{rxn}\ (iv) + \Delta H^{\circ}_{rxn}\ (v) + \Delta H^{\circ}_{rxn}\ (i)$$

$$= (-\ 393.5 + (-\ 2\times 285.8 + 890.3)\ \text{kJ mol}^{-1}$$

$$= -74.8\ \text{kJ mol}^{-1}$$

On cancelling the common terms on the two sides of equation (vi), we get equation (vii)

(vii) $C\ (\text{graphite}) + 2H_2\ (g) \rightarrow CH_4(g)\ \Delta H^{\circ}_{rxn}$ (vii) $= -74.8\ \text{kJ mol}^{-1}$

Learning Tasks

1. Explain the principles that underpin Hess's law and give its significance in thermochemistry.
2. Manipulate a set of chemical equations to obtain a desired equation.
3. Find the ΔH for the reaction below, given the following reactions and subsequent ΔH values:

$2\ CO_2(g) + H_2O(g) \rightarrow C_2H_2(g) + \frac{5}{2}\ O_2(g)$

$C_2H_2(g) + 2H_2(g) \rightarrow C_2H_6(g)$ $\Delta H = -94.5$ kJ

$H_2O(g) \rightarrow H_2(g) + \frac{1}{2}O_2\ (g)$ $\Delta H = 71.2$ kJ

$C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$ $\Delta H = -283$ kJ

PEDAGOGICAL EXEMPLARS

1. **Activity-based learning**

   State Hess' Law and apply it to construct simple energy cycle diagrams.

2. **Inquiry-based learning**

   Carry out calculations involving determination of enthalpy changes that cannot be found by direct experiment using thermochemical equations, energy cycles and given data. Examples: Enthalpy change of reaction from enthalpy change of combustion and enthalpy change of formation.

KEY ASSESSMENTS

Level 1: State Hess's law of constant heat summation.

Level 2: Given the following information

$P_4(s) + 3O_2(g) \rightarrow P_4O_6(s)$ **ΔH** $= -1640.1$ kJ/mol

$P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s) \qquad \Delta H = -2940.1$ kJ/mol

what is the value of $\Delta H(rxn)$ for $P_4O_6(s) + 2O_2(g) \rightarrow P_4O_{10}(s)$?

Level 3: Consider the following date $C(s) + O_2(g) \rightarrow CO_2(g)$ ΔH = -394 kJmol^{-1} $H_2(g) + 1/2O_2(g) \rightarrow H_2O(l)$ ΔH = –286 kJmol^{-1} $C_4H_6(g) + 5/2O_2(g) \rightarrow 4CO_2(g) + 3H_2O(l)$ ΔH = -254 kJmol^{-1}

Determine the standard enthalpy of formation of C_4H_6.

FOCAL AREA 2: BORN – HABER CYCLES

Born – Haber cycles are energy cycles drawn for ionic compounds using Hess' law. Ionic compounds are formed when constituent elements go through a series of steps for which most of the energy changes can be measured directly except the lattice energy. Lattice energy is determined from the enthalpy data by using a Born-Haber cycle. It relates lattice energies of ionic compounds to ionization energies, electron affinities, sublimation energy, bond dissociation energy and its enthalpy of formation.

Consider the reaction between lithium and fluorine:

$Li(s) + \frac{1}{2}(g) \rightarrow LiF(s)$

The standard enthalpy change for this reaction is -594.1 kJ/mol. The sum of enthalpy changes for the steps is equal to the enthalpy change for the overall reaction (-594.1 kJ/mol), formation of LiF from its elements can be traced through five separate steps. This pathway helps to analyse the energy changes of ionic compound formation, with the application of Hess's law.

1. Convert solid lithium to lithium vapor (the direct conversion of a solid to a gas is called sublimation):

$Li(s) \rightarrow Li(g) \quad \Delta H° = +155.2$ kJ/mol

The energy of sublimation for lithium is 155.2 kJ/mol.

2. Dissociate $\frac{1}{2}$ mole of F_2 gas into separate gaseous F atoms:

$\frac{1}{2}F_2(g) \rightarrow F(g)$ $\Delta H° = +75.3$ kJ/mol

The energy needed to break the bonds in 1 mole of F_2 molecules is +150.6 kJ.

Here we are breaking the bonds in half a mole of F_2, so the enthalpy change is $\frac{150.6}{2}$, or 75.3 kJ.

3. Ionize 1 mole of gaseous Li atoms:

$Li(g) \rightarrow Li^+_{(g)} + e^- \qquad \Delta H° = +520$ kJ/mol

This process corresponds to the first ionization of lithium.

4. Add 1 mole of electrons to 1 mole of gaseous F atoms.

   $F(g) + e^- \rightarrow F^-_{(g)}$ $\Delta H° = -328$ kJ/mol

5. Combine 1 mole of gaseous Li^+ a nd 1 mole of F^- to form 1 mole of solid LiF:

   $Li^+(g) + F^-(g) \rightarrow LiF(s)$ $\Delta H° = ?$

   The reverse of step 5,

   $\text{energy} + LiF(s) \rightarrow Li^+(g) + F^-(g)$

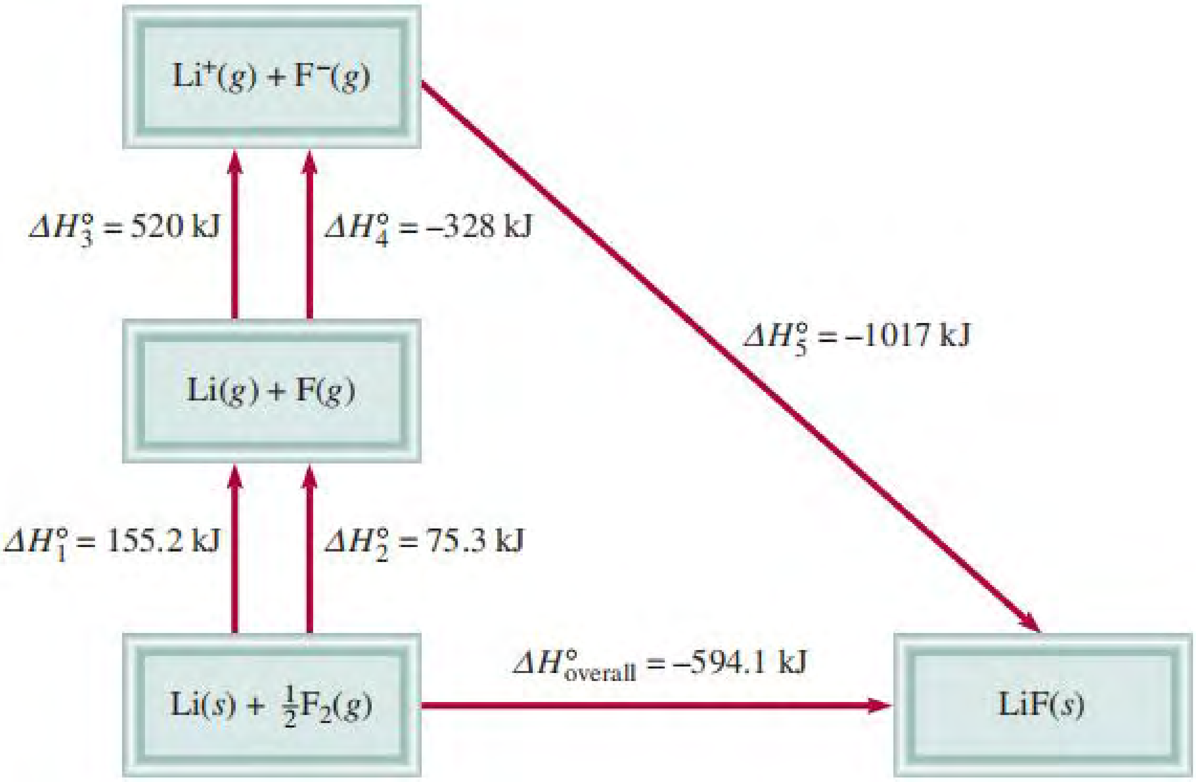
   According to Hess' law

   $\Delta H^o_{(overall)} = \Delta H^o_1 + \Delta H^o_2 + \Delta H^o_3 + \Delta H^o_4 + \Delta H^o_5$

   $-594.1 = 155.2 + 75.3 + 520 - 328 + x$

   $x = -594.70 - 422.5$

   $x = -1016.6\ kJ/mol$



Learning Tasks

1. Draw the Born-Haber cycle diagram for the formation of sodium fluoride.
2. Write an equation to represent the enthalpy of atomisation potassium.
3. Calculate the lattice energy of caesium chloride (CsCl) using the following data:

   IE (Cs) = + 376 kJ/mol, EA (Cl) = - 349.0 kJ/mol, $\Delta Hf°[Cs(g)] = +76.50$ kJ/mol, $\Delta Hf°[Cl(g)] = +121.7$ kJ/mol, and $\Delta Hf°[CsCl(s)] = -442.8$ kJ/mol.

PEDAGOGICAL EXEMPLARS

1. **Inquiry-based learning**

   Use Born-Haber cycle to determine ionization energy and electron affinity, lattice energy, dissociation energy, atomisation energy, sublimation energy, etc.

2. **Digital Learning**

   Watch videos to reinforce concept.

KEY ASSESSMENT

Level 2

1. Explain the term lattice energy.
2. Draw Born-Haber cycle for the formation of NaCl(s).
3. Use the following data to calculate for the lattice energy of sodium chloride:

   $\Delta H_f(NaCl) = - 411$ kJ/mol $\Delta H_{atom}(Na) = +107$ kJ/mol $\Delta H_{atom}(Cl) = +122$ $kJmol^{-1}$ IE(Na) = + 494 $kJmol^{-1}$ E.A. (Cl) = -349 kJ/mol Calculate lattice energy of NaCl(s).

FOCAL AREA 3: BOND ENERGY

In chemical reactions, one or more pure substances are changed to one or more other chemical substances. Generally chemical bonds are broken in the reactants whiles chemical bonds are formed in the products. The enthalpy change for any chemical reaction predominantly results from these bonds breaking and bond forming processes. Energy is required to break a chemical bond while energy is given out to form a chemical bond. Bond breaking is an endothermic process (absorbs energy) and bond formation is an exothermic process (releases energy). Therefore, a knowledge of individual bond energies can help to estimate the overall enthalpy change in a chemical reaction.

Bond enthalpy is the average energy required to break one mole of a particular kind of covalent bond in a molecule in the gaseous state into its ions.

Standard bond dissociation energy is the energy absorbed to break one mole of a particular type of covalent bond in a molecule in the gaseous state into free radicals under standard conditions.

The enthalpy of atomisation is the enthalpy change when a compound's bonds are broken and the component atoms are separated into single atoms. It can also be defined as the energy change that occurs when the chemical bonds in a compound is broken to form one mole of gaseous atoms as products.

The bond enthalpies of simple diatomic molecules such as H_2, O_2 and Cl_2 are usually measured spectroscopically but for complex molecules, thermochemical data is used to calculate bond enthalpies using Hess's law.

The experimentally determined bond enthalpy of the diatomic hydrogen molecule, for example, is $H_2(g) \rightarrow H(g) + H(g)$ $\Delta H° = +\ 436.4$ kJ/mol

According to this equation, breaking the covalent bonds in 1 mole of gaseous H_2 molecules requires 436.4 kJ of energy.

For the less stable chlorine molecule, $Cl_2(g) \rightarrow Cl(g) + Cl(g)$ $\Delta H° = +\ 242.7$ kJ/mol

Bond enthalpies can also be directly measured for heteronuclear diatomic molecules, such as HCl, $HCl(g) \rightarrow H(g) + Cl(g)$ $\Delta H° = +\ 431.9$ kJ/mol

as well as for molecules containing multiple bonds:

$O_2(g) \rightarrow O(g) + O(g)$ $\Delta H° = +\ 498.7$ kJ/mol

The table below shows the bond enthalpies of some diatomic molecules

Bond	Bond Enthalpy (kJ/mol)
H - H	+436.4
O = O	+498.7
Cl - Cl	+242.7
H - O	+460

$$\Delta H^{\circ}_{rxn} = \sum \Delta H^{o}_{bonds\ broken} - \sum \Delta H^{o}_{bonds\ formed}$$

Example

Calculate the enthalpy of reaction when one mole of water is formed from hydrogen gas and oxygen gas.

[Bond enthalpies in kJ/mol are H – H = +436, O = O = +498, H – O = +464]

Solution

$$H_2 + \tfrac{1}{2}O_2 \rightarrow H_2O$$

Bonds broken: one mole of (H – H) and half of (O = O)

$= 436 + \frac{1}{2}\left(498\right)$

$= +685$

Bonds formed: H – O – H which is twice (H – O)

$= 2(464)$

$= +928$

$$\Delta H^{\circ}_{rxn} = \sum \Delta H^{o}_{bonds\ broken} - \sum \Delta H^{o}_{bonds\ formed}$$

$$\Delta H^{\circ}_{rxn} = 685 - 928$$

$$\Delta H^{\circ}_{rxn} = -\ 243\ \text{kJ}/mol$$

Learning Tasks

1. Explain bond energy and it give its importance in studying chemical reactions.
2. Calculate the enthalpy of reaction for the combustion of methane as described in the equation

   $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

   Given the following bond enthalpies C – H (414 kJ/mol), O = O (498 kJ/mol), C = O (799 kJ/mol), H – O (460 kJ/mol).
3. Compare the bond enthalpies of different chemical bonds between the carbon atoms and relate it to their bond strength.

PEDAGOGICAL EXEMPLARS

1. **Talk-for-learning approach**

   Using think-pair-share approach, distinguish between bond dissociation energy and bond energy.

2. **Activity-based learning**

   a. Calculate the enthalpy change of reaction from bond energy values.

   b. Discuss bond energy as a measure of energy content, hence bond strength.

   c. Discuss why the amount of energy released per unit mass is a factor in choosing fuel used at home, transport and industry. Is this the only consideration when deciding upon a fuel?

KEY ASSESSMENTS

Level 1: Define bond enthalpy.

Level 2: Distinguish between bond dissociation energy and bond energy.

Level 3

1. Consider the following bond enthalpies

Bond	BDE (kJmol⁻¹)
H – H	436
C – C	348
C = C	612
N ≡ N	944
N – H	388

   a. Write a balanced chemical equation for the formation one mole of NH_3.

   b. Use the bond enthalpy data to calculate for the enthalpy of formation of NH_3.

2. Study the case below and answer the question that follow.

Case

EdinChem's industry has been producing and exporting ethanol for 10 years now. Recently the management team has identified some inconsistencies in the heat content measurement for the ethanol they produce. As a new quality control officer, investigate and suggest 3 ways to stop the inconsistencies in the measurements.

Hint

The recommended mode of assessment for week 3 is ***case study****. Use the level 3 question 2 as a sample question. Refer to the Teacher Assign Portfolio Assessment for the academic year by Week 3. Portfolio should be submitted by Week 22. See Appendix A of this Section and Teacher Assessment Manual and Toolkit pages 22-25 for more information on how to organise a portfolio assessment.*

Section 1 Review

This section introduced learners to the foundational principles of chemistry, specifically focusing on the complex realm of thermodynamics and the reorganisation of atoms that underlies energy transformations. By exploring energy changes and their corresponding profiles, students gain a deeper understanding of the dynamic processes at play within chemical systems. The explanation of standard enthalpy changes and methods for experimentally determining these values equips learners with the tools to quantify energy variations accurately.

Furthermore, the section looked into the application of Hess's Law of constant heat summation, Born-Haber cycles, and bond energies in calculating the enthalpies of various reactions. These concepts provide students with a systematic approach to analysing and predicting energy changes in chemical reactions, fostering a more comprehensive comprehension of the underlying principles of thermodynamics.

APPENDIX A: INDIVIDUAL PORTFOLIO

The Purpose

The purpose of the annual portfolio is to document and assess a learner's progress over the academic year. Learners understanding why they are creating the portfolio is crucial. It serves as a comprehensive collection of work that reflects their achievements, skills, and areas for growth. The portfolio can be used for various purposes, including parent-teacher discussions, personal reflection, etc.

Personal Information

Name, contact information, a brief introduction or personal statement that outlines their goals and aspirations and a brief story about the life and discoveries of two prominent chemists as role models is required.

Table of Contents

To enhance navigation through the portfolio, include a clear table of contents that outlines all sections and elements. This will help readers easily find specific information about your work.

Work Samples

Choose a variety of work samples that demonstrate your skills and growth. This should include ***projects, practical works, or reports, class exercises, mid semester and end of semester examinations and others that highlight your learning***. Also include creative works such as ***writing samples, or any other creative outputs***. Aim for a balanced representation of your abilities, highlighting both high-quality work and examples that illustrate your development over time.

Reflective Pieces

Reflection is key to understanding your learning journey. For each piece included in the portfolio: Write a brief reflection on what you learned from the experience, discuss challenges faced and how you overcame them. Highlight any feedback received from peers or instructors.

Organisation

Structure your portfolio into clear sections based on themes or subject areas. Common categories include Academic achievements, creative projects, community service and personal development. This organisation helps convey a coherent narrative about your educational experience.

Use Digital Tools for Presentation (If applicable)

Consider using digital platforms to create your portfolio, as they allow for easy updates and sharing. Tools like online portfolio makers can help you design an appealing layout and incorporate multimedia elements such as videos or links to online projects.

Review and Revise Regularly

Your portfolio should be a living document. Regularly review and update it with new work and reflections throughout the year. This ensures it remains current and accurately reflects your ongoing development.

Prepare for Submission

If the portfolio is intended for submission (e.g., academic transcript), ensure it meets any specific guidelines provided by institutions regarding format, length, and content requirements.

Feedback

Before finalising your portfolio, seek feedback from teachers, mentors, or peers. Their insights can provide valuable perspectives on how to improve the presentation and content of your portfolio. By following these steps, learners can create a comprehensive yearly portfolio that effectively highlights their academic journey, personal growth, and readiness for future challenges or opportunities. Offer targeted feedback alongside scores to help learners understand strengths and areas for improvement

How to Administer

Regular Updates: Encourage learners to update their portfolios periodically throughout the year.

Guidance and Support: Provide guidance on what types of work to include and how to reflect on their learning. Provide the rubric to learners before they begin compiling their portfolios to clarify expectations. Use checklists to ensure all required items are collected.

A well-organised portfolio should include:

Section	Description
Title Page	Name and title of the portfolio
Table of Contents	Outline of sections included
Personal Information	Basic details about the learner
Monthly Summary	Overview of learning experiences
Work Samples	Selected artifacts that demonstrate learning
Reflections	Insights into personal growth and learning
Future Goals	Aspirations for the upcoming year

Scoring Rubrics for portfolio

Criteria	Excellent (4marks)	Very Good (3marks)	Good (2marks)	Fair (1mark)
Content Quality	*The portfolio contains 5 or more relevant work samples or documentation (e.g., Assignments, projects, reflections, practical work, end of semester examination*	*The portfolio contains 4 relevant work samples or documentation (eg. Assignments, projects, reflections, practical work, end of semester examination, etc.)*	*The portfolio contains 3 relevant work samples or documentation (eg. Assignments, projects, reflections, practical work, end of semester examination, etc.)*	*The portfolio contains 2 or less relevant work samples or documentation (e.g. Assignments, projects, reflections, practical work, end of semester examination, etc.)*
Reflection	*Reflections are profound, demonstrating deep understanding and future application of learning and contains explanation of all the 4 points* *(e.g. What did I learn? How did I learn? What did I find challenging or surprising? What will I do differently next time?)*	*Reflections are insightful and connect well with the work presented but contains 3 of the points* *(e.g., What did I learn? How did I learn? What did I find challenging or surprising? What will I do differently next time?)*	*Reflections are present but lack depth or connection to the learning process but contain 2 of the points.* *(e.g. What did I learn? How did I learn? What did I find challenging or surprising? What will I do differently next time?)*	*Reflections do not connect to the work (e.g. What did I learn? How did I learn? What did I find challenging or surprising? What will I do differently next time?)*

Criteria	Excellent (4marks)	Very Good (3marks)	Good (2marks)	Fair (1mark)
Organization	*The portfolio is exceptionally organised, enhancing readability and presentation quality. (e.g. table of content provided and well organised, portfolio items are presented in orderly manner, clear headings and labels are used to separate sections and items)*	*The portfolio is well-organised with clear sections and logical flow. (e.g. table of content provided and well organised, portfolio items are presented in orderly manner, but label and headings are not clear)*	*Some organisation is evident, but sections may be unclear or inconsistent. (e.g. table of content provided but not well organised, portfolio items are presented in orderly manner, but label and headings are not clear)*	*The portfolio is poorly organised, making it difficult to navigate (e.g. table of content provided and but not well organised, portfolio items are not presented in orderly manner, also the label and headings are not clear)*
Presentation	*Presentation exceeds standards with less than 4 errors; highly professional and engaging layout.*	*Presentation meets standards with 4 errors; visually appealing layout.*	*The presentation has several errors that detract from overall quality.*	*The presentation quality is low; numerous errors in spelling and grammar.*

Total Score = 15 Marks

APPENDIX B: GROUP PROJECT WORK

Structure

Give guidelines to learners by week 2 of the first semester on the nature, structure, and the duration of the group project work. Assign learners to groups based on class size, gender, background, etc. Give them the reference sources and all the necessary materials required. Finally, distribute the printed questions to all group leaders and assure them of the support systems available.

Task

Analyse the calorific values of 5 selected food groups: fruits, vegetables, grains, proteins, and fats/oils and 5 selected fuels: gasoline, coal, diesel, wood, and natural gas. Compare and highlight their calorific value and energy density and potential applications of these values to the foods and the fuels.

Report your findings in the format below:

a) Record and organise data in a table

b) Collect data on the calorific values of foods and the fuels

c) Calculate the energy density of each food and fuel

d) Compare and contrast the calorific values of foods and fuels

e) Discuss the implications of energy density on food choice and fuel selection

f) A 3-page comprehensive report detailing the research, calculations and analysis and representation of the calorific values

g) A 10-minute presentation summarising key findings

Table 1.1: Calorific Value and Energy Density of some Foods

Food	Calorific Value (CV)kcal/100g	Energy Density (ED) kcal/g
Fruits	40-60	0.40-0.60
Vegetables	20-40	0.20-0.40
Grains	80-120	0.80-1.20
Proteins	120-180	1.20-1.80
Fats and oils	800-1000	8.00-10.00

SOURCE: Ghana Food Composition Table

Note

The exact values may vary depending on the specific type, cooking method and ripeness of each food.

Energy Density is calculated by dividing the Calorific Value by 100

Table 1.2: Calorific Value and Energy Density of some Fuels

Fuel	Calorific Value MJ/kg	Energy Density MJ/L
Gasoline	44.4-45.4	44.4-45.8
Coal	24-32	40-50
Diesel	45.4-46.2	45.5-46.5
Natural Gas	53.6-55.5	53.6-55.5
Wood	16-20	10-15

SOURCE: US Energy Information Administration (EIA)

Note

The exact values may vary depending on the specific type, grade, and moisture content of each fuel.

Energy Density is calculated by dividing the Calorific Value by density of the fuel

How to administer

1. Provide clear guidelines for developing the project and how to assess it, design and plan the project that aligns with the learning outcomes, etc.
2. Establish assessment criteria and scoring rubrics
3. Communicate project requirements to learners, etc.

Feedback

Share the scoring rubrics with the learners and discuss overall performance of learners highlighting learners' strengths, weaknesses, and areas of improvement, etc.

Marking Scheme for the Group Project Work Assessment

Introduction

Define calorific value

*Calorific value in food refers to the **amount of energy** that the **body obtains** from consuming a particular food item **(2marks)**.*

*The calorific value in fuel also refers to the **amount of energy released** when a **unit of the fuel** is completely **burnt (2marks)***

Significance of caloric value on food

1. *Helps to determine the **energy requirement** of individuals based on their age, sex, weight and activity level*
2. *Helps individuals to **manage their weight** by balancing energy intake and expenditure*
3. *Helps in **planning diets** for individuals with specific nutritional needs, such as pregnant women, athletes, or people with certain medical conditions*
4. *Enable consumers to make **informed choices** about their diets*

(Any 2 × 1marks)

Significance of caloric value on fuel

1. *Helps to determine the **energy efficiency** of fuels, which is crucial for optimising energy production and consumption*
2. *Enables us to **compare different fuels,** to allow one to make informed decision about fuel selection and usage*
3. *Can influence the **environmental impact** of fuel combustion*
4. *Is essential for designing and optimising engines*

(Any 2 × 1marks)

Objectives of the analysing calorific value

1. *To provide accurate information for decision making in various fields such as nutrition, energy production, and environmental sustainability*
2. *To optimise the use of resources such as food and fuel, to minimise waste and maximise efficiency*
3. *To provide the basis for research and development of new products, fuels, and energy technologies*

(Any 2× 1marks)

Table 1: *Calorific Value and Energy Density of some Foods*

Food	Calorific Value (CV)kcal/100g	Energy Density (ED) kcal/g
Fruits	*40-60*	*0.40-0.60*
Vegetables	*20-40*	*0.20-0.40*
Grains	*80-120*	*0.80-1.20*
Proteins	*120-180*	*1.20-1.80*
Fats and oils	*800-1000*	*8.00-10.00*

SOURCE: Ghana Food Composition Table

Total = 1 × 10 = 10marks

Note

The exact values may vary depending on the specific type, cooking method and ripeness of each food.

Energy Density is calculated by dividing the Calorific Value by 100

Table 2: *Calorific Value and Energy Density of some Fuels*

Fuel	Calorific Value MJ/kg	Energy Density MJ/L
Gasoline	44.4-45.4	44.4-45.8
Coal	24-32	40-50
Diesel	45.4-46.2	45.5-46.5
Natural Gas	53.6-55.5	53.6-55.5
Wood	16-20	10-15

SOURCE: US Energy Information Administration (EIA)

Total = 1 × 10 = 10marks

Note

The exact values may vary depending on the specific type, grade, and moisture content of each fuel.

Energy Density is calculated by dividing the Calorific Value by density of the fuel

Comparative analysis of Calorific Value and Energy Density

1. *High CV and ED foods tend to be high in fats and calories*
2. *Moderate CV and ED foods are often balanced in micronutrients*
3. *Low CV and ED foods are typically high in fibre and water content*
4. *Protein content is moderately correlated with CV but weakly correlated with ED*
5. *Air pollution is high when using fuels with high CV and ED*

Any 3 × 1= 3marks

Conclusion

1. *For a healthy living eat foods with low/moderate CV and ED*
2. *For a clean environment use fuels with low/moderate CV and ED*

2 × 1 = 2 marks

Class presentation

Rubrics for scoring a group presentation task

Criteria	Excellent (4 Marks)	Very Good (3 Marks)	Good (2 Marks)	Fair (1 Mark)
Presentation	*The presentation portrays all 4 of these points* *Audible voice, simple language, coherence, and flow of ideas*	*The presentation portrays 3 of these points* *Audible voice, simple language, coherence, and flow of ideas*	*The presentation portrays 2 of these points* *Audible voice, simple language, coherence, and flow of ideas*	*The presentation portrays 1 of these points* *Audible voice, simple language, coherence, and flow of ideas*
Teamwork	*Exhibit 4 of these Contributing to the group.* *Respecting the views of others* *Tolerating others* *Everyone participates,* *Taking responsibility*	*Exhibit 3 of these Contributing to the group.* *Respecting the views of others* *Tolerating others* *Everyone participates* *Taking responsibility*	*Exhibit 2 of these Contributing to the group.* *Respecting the views of others* *Tolerating others* *Everyone participates* *Taking responsibility*	*Exhibit 1 of these Contributing to the group.* *Respecting the views of others* *Tolerating others* *Everyone participates* *Taking responsibility*

Total = 45marks

SECTION 2: CHEMICAL KINETICS

Strand: Physical Chemistry

Sub-Strand: Matter and its properties

Learning Outcome: *Use the concept of chemical kinetics to explain that chemical reactions proceed at different rates depending on their conditions and design experiments to investigate these conditions.*

Content Standard: Demonstrate knowledge and understanding that energy changes occur in chemical reactions as a result of reorganisation of atoms at different rates.

INTRODUCTION AND SECTION SUMMARY

In this section, learners will be taken through Chemical kinetics under Physical chemistry which covers week four and five of the chemistry curriculum. In week four, learners will be assisted to explain and measure the rate of reactions through suitable experimental procedures, investigating the varying factors that affect reaction rates. By week five, learners will describe the collision theory, linking it to the effects on reaction rates and construct rate equations from experimental data. Through analytical thinking and problem-solving, learners will interpret graphs to deduce the order of reactions and apply their knowledge to solve rate expression problems. Throughout these weeks, 21st-century skills, digital learning tools, and a focus on gender and social inclusion will be woven into series of engaging lessons to ensure a comprehensive learning experience for all learners.

The weeks covered by the section are:

Week 4: *rates of reaction and factors which affect it*

Week 5: *collision theory*

SUMMARY OF PEDAGOGICAL EXEMPLARS

In delivering lessons on rate of reaction and factors affecting reactions, a variety of pedagogies will be employed to enhance engagement and understanding of learners. Through collaborative learning, learners will utilise think-pair-share to explain the term rate of reaction effectively. An inquiry-based approach will be used to determine suitable procedures for measuring reaction rates, while experiential learning will involve designing and conducting experiments to investigate the impact of concentration, surface area, temperature, and catalysts on reaction rates. Talk-for-learning discussions will analyse data and draw conclusions on the effects of various parameters on reaction rates and explore real-world applications. Activity-based learning will utilise sketching and using Maxwell-Boltzmann energy distribution curves to explain the effects of temperature and catalysts on reaction rates. Also, problem-solving approach will be implemented to deduce rate expressions, sketch reaction graphs, and predict the order of reactions. Through these diverse pedagogical methods, students will actively engage in the learning process and develop a deep understanding of the concepts.

ASSESSMENT SUMMARY

To facilitate comprehensive assessment of learners on rate of reaction and factors affecting reactions, a variety of formative assessment methods will be utilised, including short answers, oral and written presentations, lab reports, equipment manipulation, safety procedures, graph plotting, deducing rate expressions from experimental data, measurements, demonstrations, and homework assignments. Projects and group activities will be employed to assess learners' comprehension and application of concepts, tailored to cover levels 1, 2, 3, and 4 of the Depth of Knowledge (DOK) framework while emphasising critical thinking, communication, and creativity. Class exercises, lab reports, measurements, and group tasks will allow learners to demonstrate their grasp of concepts, while homework assignments and practical group activities will provide opportunities for practical application of knowledge in real-world contexts.

WEEK 4

Learning Indicators

1. *Explain the term rate of reaction and describe the suitable experimental procedures for measuring rates of reactions*
2. *Perform experiments to investigate the factors that affect the rate of reaction and analyse the data from the rate experiments*

FOCAL AREA 1: RATE OF REACTION

Chemical kinetics is the area of chemistry concerned with the speed or rate at which a chemical reaction occurs. Kinetics refers to the rate of a reaction, or the reaction rate. A knowledge of reaction rates is useful in drug design, in pollution control, and in food processing.

Rate of a chemical reaction is defined as the change in concentration of a reactant or product per unit time.

Rate can also be defined as the amount of reactants consumed (disappearance, expenditure, depletion, etc.) per unit time or the amount of products formed (appearance, creation, etc.) per unit time.

Rate $= \frac{\Delta[A]}{\Delta t}$, where A is the reactant or product being considered, and the square brackets indicate concentration in mol/dm^3. The symbol Δ indicates a change in a given quantity and a change can be positive (increase) or negative (decrease), thus leading to a positive or negative reaction rate by this definition. However, for convenience, rate is always defined as a positive quantity.

Consider the hypothetical reaction: $aA + bB \rightarrow cC + dD$

$$\text{rate} = -\frac{1}{a}\frac{\Delta[A]}{\Delta t} = -\frac{1}{b}\frac{\Delta[B]}{\Delta t} = \frac{1}{c}\frac{\Delta[C]}{\Delta t} = \frac{1}{d}\frac{\Delta[D]}{\Delta t}$$

Example 1

Consider the thermal decomposition of gaseous N_2O_5 to NO_2 and O_2 through the following equation: $2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$.

Write expressions for the reaction rate in terms of the rates of change in the concentrations of the reactant and each product with time.

Solution

Step 1: Because O_2 has the smallest coefficient in the balanced chemical equation for the reaction, define the reaction rate as the rate of change in the concentration of O_2 and write that expression.

Step 2: The balanced chemical equation shows that two moles of N_2O_5 decompose to produce one mol of O_2 and four moles of NO_2. The molar ratios of O_2 to N_2O_5 and to NO_2 are thus 1:2 and 1:4, respectively. This means that the rate of change of $[N_2O_5]$ and $[NO_2]$ must be divided by its stoichiometric coefficient to obtain equivalent expressions for the reaction rate.

$$\text{rate} = \frac{\Delta[O_2]}{\Delta t} = \frac{\Delta[NO_2]}{4\Delta t} = -\frac{\Delta[N_2O_5]}{2\Delta t}$$

Example 2

Consider the thermochemical equation: $2O_3(g) \rightarrow 3O_2(g)$.

a. How is the rate at which ozone disappears related to the rate at which oxygen appears in the reaction?

If the rate at which O_2 appears is $6.0 \times 10^{-5}\, moldm^{-3}$ at a particular instant, at what rate is O_3 disappearing at this same time the rate at which oxygen appears in the reaction.

Solution

$2O_3(g) \rightarrow 3O_2(g)$

a. $\text{rate} = -\dfrac{\Delta[O_3]}{2 \times \Delta t} = \dfrac{\Delta[O_2]}{3 \times \Delta t}$

$$= \gg \frac{\Delta[O_3]}{\Delta t} = \frac{2}{3} \times \frac{\Delta[O_2]}{\Delta t},$$

b. But the rate of appearance of $O_2\left(\dfrac{\Delta[O_2]}{\Delta t}\right)$ is $6.0 \times 10^{-5}\,\text{moldm}^{-3}$

$= \gg$ the rate which O_3 disappears $\left(\dfrac{\Delta[O_3]}{\Delta t}\right) = \dfrac{2}{3} \times 6.0 \times 10^{-5}$

$$= 4.0 \times 10^{-5}\ \text{moldm}^{-3}$$

Exercise

The contact process is used in the manufacture of sulphuric acid. A key step in this process is the reaction of SO_2 with O_2 to produce SO_3 through the equation: $2SO_2(g)+O_2(g)\rightarrow 2SO_3(g)$. Write expressions for the reaction rate in terms of the rate of change of the concentration of each species. If the rate of consumption of SO_2 in this process in 3×10^{-4} moldm^{-3} calculate the rate of consumption of oxygen and the rate of appearance of SO_3.

Initial rate of a reaction

The initial rate of reaction is the rate at which the reagents are first brought together. Initial rate can be obtained either experimentally or graphically. To experimentally determine the initial rate, reagents must be brought together and the rate of reaction measured as quickly as possible.

Initial rate experiments often involve measuring the time taken for some identifiable event to happen early in a reaction. This could be the time required for 5 cm^3 of gas to be produced, for a small, measurable amount of precipitate to form, or for a colour change to occur.

Average rate of reaction

An average rate of reaction is simply a change in reactant (or product) concentration over a period of time. The average rate of reaction can be determined by choosing two points in time say t_1 *and* t_2 and measuring the concentration of the reactants (C_1 *and* C_2) at those points.

The average rate of reaction for the reaction (a.v.g rate) $= \frac{\Delta[\text{reactants}]}{\Delta t}$

$$\text{a.v.g rate} = \frac{C_2 - C_1}{t_2 - t_1}$$

The average rate of reaction can also be expressed from a balanced chemical equation taking into consideration the stoichiometric coefficients. Consider the hypothetical chemical reaction:

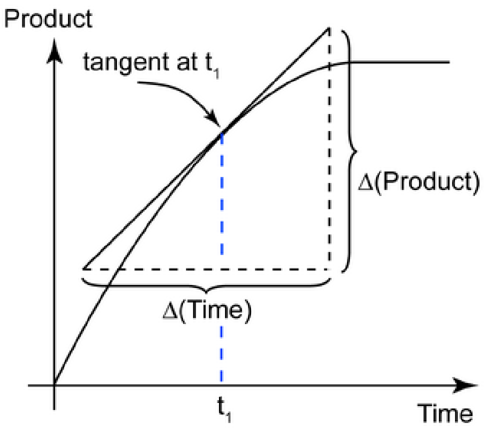
$$aA + bB \rightarrow cC + dD$$

$$\text{a.v.g rate} = = -\frac{1}{a}\frac{\Delta[A]}{\Delta t} = -\frac{1}{b}\frac{\Delta[B]}{\Delta t} = \frac{1}{c}\frac{\Delta[C]}{\Delta t} = \frac{1}{d}\frac{\Delta[D]}{\Delta t}$$

Instantaneous rate of a reaction

The instantaneous rate of a reaction is the reaction rate at any given point in time. As the period of time used to calculate an average rate of a reaction becomes shorter and shorter, the average rate approaches the instantaneous rate.

Suppose we have a reaction that forms a product A. A graph of the concentration [A] against time t is usually not a straight line but a curve. The instantaneous rate of reaction is the slope of the line (the tangent to the curve) at any time (t).



Instantaneous rate $= \frac{\Delta[P]}{\Delta t}$(where Δt is small)

Measurement of rate of reaction by colour change

Sodium thiosulfate solution reacts with dilute hydrochloric acid according to the reaction:

$Na_2S_2O_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) + S(s)$

The sulphur forms a cloudy yellow-white precipitate during the reaction. The time taken for this to achieve a given cloudiness provides a way to measure the reaction time.

Procedure

1. Using a measuring cylinder, add 50 cm^3 of dilute sodium thiosulfate solution to a conical flask.
2. Place the conical flask on a piece of paper with a black cross drawn on it.
3. Using a different measuring cylinder, add 10 cm^3 of bench dilute hydrochloric acid to the conical flask. Immediately swirl the flask to mix its contents, and start a stop clock.
4. Look down through the reaction mixture. When the cross can no longer be seen, record the time on the stop clock.
5. Measure and record the temperature of the reaction mixture, and clean the apparatus
6. Repeat steps 1 to 5 with different starting temperatures of sodium thiosulfate solution.
7. Plot a graph of reaction rate on the vertical axis and temperature (°C) on the horizontal axis and draw a curve of best fit.

Measurement of rate of reaction by changes in volume of gas evolved

Calcium carbonate reacts with dilute hydrochloric acid:

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$. The volume of carbon dioxide gas produced can be measured using a gas syringe.

Procedure

1. Support a gas syringe with a stand, boss and clamp.
2. Using a measuring cylinder, add 50 cm^3 of dilute hydrochloric acid to a conical flask.
3. Add 0.4 g of calcium carbonate to the flask. Immediately connect the gas syringe and start a stop clock.
4. Every 10 seconds, record the volume of gas produced.
5. When the reaction is complete, clean the apparatus.
6. Repeat steps 1 to 5 with different concentrations of hydrochloric acid.
7. Plot a graph of volume of gas (cm^3) on the vertical axis and time (s) on the horizontal axis

8. Draw a curve of best fit.

Learning Tasks

1. Define the term rate of reaction and explain its units.
2. Explain the difference between instantaneous and average rate of reaction.
3. Describe a simple experiment to measure the rate of reaction (e.g., mixing baking soda and vinegar).

PEDAGOGICAL EXEMPLARS

1. Collaborative learning approach

- **a.** Ask learners to define the term rate of reaction in their own words and encourage them to consider what factors might affect the rate.
- **b.** Pair learners to share their definitions and ideas and ask them to discuss and agree on a common definition.
- **c.** Encourage them to brainstorm ways to express the rate (initial, average, instantaneous). Have pairs share their definitions and ideas with the class.
- **d.** Encourage whole-class discussion to reach a consensus on the definition and ways to express the rate.
- **e.** Provide guidance on how to express the rate (e.g., equations, graphs). Offer examples to illustrate each type of rate.
- **f.** Encourage learners to create their own examples. Observe learner participation during the think-pair-share activity.

2. Inquiry-based approach

- **a.** Divide learners into small mixed-ability groups and assign each group a specific method for measuring rates of reaction (e.g., colour change, gas evolution, precipitate formation, mass change, appearance/disappearance of product).
- **b.** Ask groups to discuss and determine:
    - i. The procedure for measuring the rate of reaction using their assigned method.
    - ii. The equipment and materials needed.
    - iii. Potential sources of error and how to minimise them.
    - iv. How to calculate the rate of reaction from their measurements.
- **c.** Encourage groups to consider the advantages and limitations of their assigned method.

d. Provide learners with access to a virtual laboratory simulation (e.g., online chemistry lab software or YouTube videos) and encourage them to take notes on the procedures and any limitations or potential sources of error.

e. Ask each group to present their procedure and findings to the class and encourage comparison and discussion of the different methods.

f. Facilitate a class discussion to highlight the advantages and limitations of each method.

g. Observe learner participation during the group discussion and virtual laboratory activities and assess their understanding through a written quiz or class discussion.

KEY ASSESSMENTS

Level 1: Define the term rate of reaction.

Level 2: Compare and contrast the use of changes in volume of gas evolved verses formation of precipitate to measure rate of reaction.

Level 3: Design an experiment to measure the rate of reaction using colour change.

FOCAL AREA 2: FACTORS THAT AFFECTS RATE OF REACTIONS

Reaction rates can vary greatly over a large range of time scales. Some reactions can proceed at explosively fast rates while others can occur at a slow rate over many years.

1. **Temperature**: Increasing temperature increases the rate of all chemical reactions. This is because at higher temperature, the average kinetic energy of the particles increases, so more reacting particles have energy equal to or greater than the activation energy of the reaction. Also at higher temperatures, the reacting particles travel at higher speed, the frequency of collision between the reacting particles increases causing the frequency of effective collision to increase and as such more products are formed per unit time and hence the rate of reaction increases.

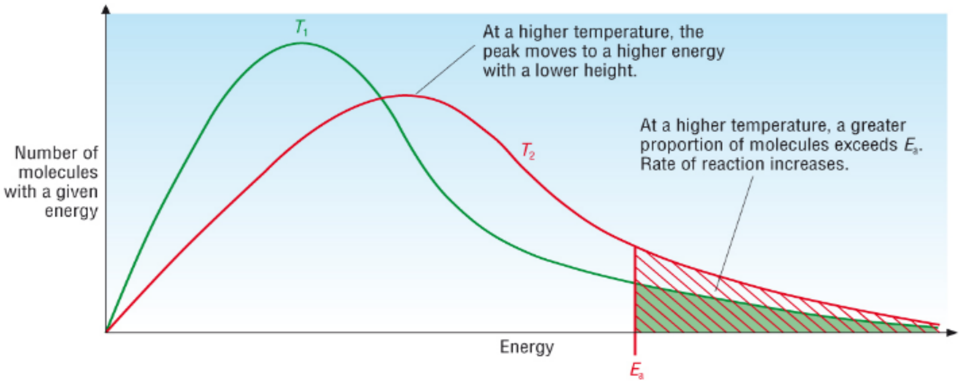


Figure 2.1: *Effect of temperature on the kinetic energy distribution of molecules in a sample*

This explains why food can spoil quickly when left on the kitchen counter. However, the lower temperature inside of a refrigerator slows that process so that the same food remains fresh for days.

Gas burners, hot plates, and ovens are often used in the chemistry laboratory to increase the speed of reactions that proceed slowly at ordinary temperatures.

2. **Concentration:** Increasing the concentration of one or more reactants increase the rate of reaction. When the concentration of the reactants is increased, there is a higher chance that collisions among the particles will increase, more collisions mean a higher reaction rate.

3. **Available surface area and nature of the reactants:** Surface area is the measure of how much area a reacting particle is exposed. Most reactions involve two or more reactants whose constituent atoms, ions or molecules must collide for the reaction to occur. That is why reactions are rarely carried out in the solid state, but often occur in liquid solutions or in the gas phase where the reactants are able to intermingle and their constituent atoms, molecules or ions can collide with each other easily.

   For the same mass, many small particles have a greater total surface area than one large particle. The more surface contact between reactants, the higher the rate of the reaction. The less surface contact, the lower the reaction rate.

Note

It is also important to note that not all reactions depend on surface area. If both reactants are gases or liquids that mix together, then surface area is not a factor.

4. **Catalyst:** A catalyst is a substance that speeds up the rate of a chemical reaction without being used up (consumed) in the reaction. A catalyst works by providing an alternative route for the reaction to occur. This alternative route has a lower activation energy. This increases the proportion of reacting particles with energy greater than the activation energy. As a result, the frequency of effective collision increases. More products are formed on per unit time and the rate of reaction is higher.

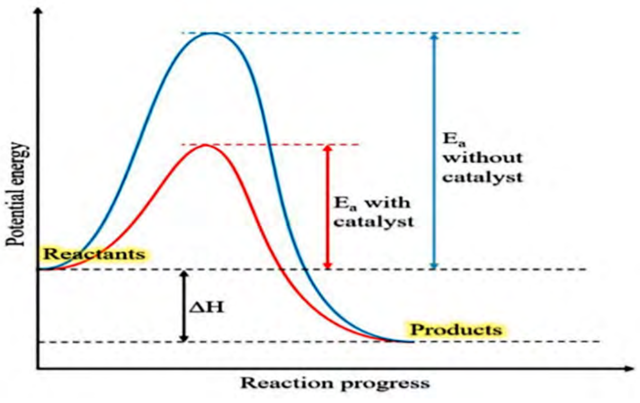


Figure 2.2: *Effect of catalyst on the rate of a reaction*

5. **Pressure:** Increasing pressure increases the rate of chemical reactions whose reactants and products are in their gaseous state. Increasing pressure means either forcing the same number of gaseous particles into a smaller volume, or increasing the number of particles in a given volume. In both cases the particles are now closer together and this increases the frequency of collisions per unit time thereby increasing the reaction rate.

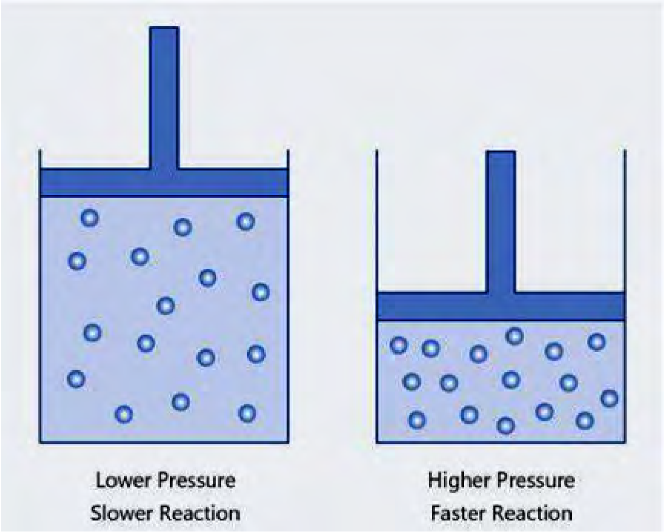


Figure 2.3: *Effect of pressure on the rate of a reaction*

Experiments to investigate the factors that affect rate of reactions

1. Experiment to investigate the effect of temperature on the rate of reaction

Procedure

a. Dilute hydrochloric acid is heated to a set temperature using a water bath

b. Add the dilute hydrochloric acid into a conical flask

c. Add a strip of magnesium and start the stopwatch

d. Stop the time when the magnesium fully reacts and disappears

e. Repeat at different temperatures and compare results

Result

With an increase in the temperature, the rate of reaction will increase.

This is because the particles will have more kinetic energy than the required activation energy, therefore more frequent and effective collisions will occur, increasing the rate of reaction.

2. Experiment to investigate the effect of concentration on rate of reaction

Procedure

a. Measure 50 cm^3 of sodium thiosulfate solution into a conical flask.

b. Measure 5 cm^3 of dilute hydrochloric acid into a measuring cylinder.

c. Draw a cross on a piece of paper and put it underneath the flask.

d. Add the acid into the flask and immediately start the stopwatch.

e. Look down at the cross from above and stop the stopwatch when the cross can no longer be seen.

f. Repeat using different concentrations of sodium thiosulfate solution (mix different volumes of sodium thiosulfate solution with water to dilute it).

Result

With an increase in the concentration of a solution, the rate of reaction will increase, this is because there will be more reactants in a given volume, allowing more frequent and effective collisions, increasing the rate of reaction

3. Experiment to investigate the effect of Surface Area on the rate of reaction

Procedure

a. Add dilute hydrochloric acid into a conical flask.

b. Use a delivery tube to connect this flask to a measuring cylinder upside down in a bucket of water (downwards displacement).

c. Add magnesium ribbon to the conical flask and quickly put the bung back into the flask.

d. Measure the volume of gas produced in a fixed time using the measuring cylinder.

e. Repeat with different size pieces of magnesium ribbon (The same total mass of magnesium must be used).

Result

Smaller pieces of magnesium ribbon cause an increase in the surface area of the solid, so the rate of reaction will increase. This is because more surface area of the particles will be exposed to the other reactant so there will be more frequent and successful collisions, increasing the rate of reaction.

4. Experiment to investigate the effect of catalyst on the rate of reaction

Procedure

a. Add hydrogen peroxide into a conical flask.

b. Use a delivery tube to connect this flask to a measuring cylinder upside down in a tub of water (downwards displacement).

c. Add the catalyst manganese (IV) oxide into the conical flask and quickly place the bung into the flask.

d. Measure the volume of gas produced in a fixed time using the measuring cylinder.

e. Repeat experiment without the catalyst of manganese (IV) oxide and compare results.

Result

Using a catalyst will increase the rate of reaction. The catalyst will provide an alternative pathway with a lower activation energy so that more colliding particles will have the required activation energy to react. This will allow more frequent and effective collisions thereby increasing the rate of reaction.

Application of the factors that affect rate of reactions in everyday life

1. **Temperature**

   Food Preservation: Cold temperatures slow down spoilage reactions. That is why perishables are refrigerated to keep bacteria and their enzymes from working as quickly as they could. Storing fruits and vegetables in the fridge to extend their freshness.

2. **Concentration**

   Cleaning: Household cleaning products often contain concentrated solutions. Diluting them with water changes their effectiveness and reaction rates.

3. **Catalysts**

   Chemical Industry: Catalysts are essential for efficient production. They speed up reactions, allowing manufacturers to meet demand while maintaining safety standards.

4. **Surface area**

   Medicine and Pills: Crushed or powdered medications dissolve more quickly in the stomach. Their larger surface area speeds up their absorption into the blood stream.

5. **Pressure**

Gases at higher pressure occupy a smaller volume therefore the chances of a collision between particles increases. This means that the likelihood of a successful collision goes up and then the reaction rate goes up.

Learning Tasks

1. State at least three factors that affects rate of chemical reaction.
2. Describe how the following factors affect reaction rates:
   - a. Reactant concentration
   - b. Physical state of reactants and surface area
   - c. Temperature
   - d. Presence of a catalyst
   - e. Pressure
3. Explain how surface area affects the rate of dissolving a sugar cube versus powdered sugar.
4. The first step in the production of nitric acid is the combustion of ammonia: $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$. Write the equations that relate the rates of consumption of the reactants and the rates of formation of the products.

PEDAGOGICAL EXEMPLARS

1. Experiential Learning

In small mixed-ability groups and using circus activities, assist learners to design, and perform experiments to investigate the effect of the following on rate of reaction:

- **a.** change in concentration (using gas syringe to measure the volume of CO_2 from HCl and $CaCO_3$ reaction with time).
- **b.** changes in surface area (using balance to measure loss of mass or using gas syringe to measure volume of CO_2 produced from HCl and $CaCO_3$ reaction with time).
- **c.** changes in temperature or concentration with respect to time (Sulphur-clock experiment).
- **d.** catalyst (using gas syringe to measure the volume of oxygen gas produced from decomposition of H_2O_2 by MnO_2).
- **e.** Provide additional support for struggling learners and offer extensions for advanced learners, such as designing additional experiments or researching real-world applications.

2. **Talk-for- learning**
   - **a.** In small mixed-ability and mixed-gender groups where applicable, guide learners to analyse experimental data and graphs of reactions and deduce patterns and hence the effect of the parameters on the rate of reaction. Encourage active listening and respectful dialogue.
   - **b.** Have learners discuss in their group's practical applications of each outcome on everyday life (Example, surface area and safety of grain mill factories, why a glowing splint rekindles when it is put in a bottle of oxygen gas, why smoking is forbidden in areas where bottled oxygen is in use, burning of charcoal, the use of antioxidants as competitive inhibitors to preserve food etc). Circulate and facilitate discussions to keep the conversation on track.

KEY ASSESSMENTS

Level 1: State at least three factors that affects rate of chemical reactions.

Level 2

1. Discuss the rationale behind using powdered versus solid reactants.
2. State and explain three factors that affect rate of chemical reactions.

Level 3

1. Consider the spoilage of milk, a perishable item that undergoes a chemical reaction involving the enzyme lactase. Explain how refrigeration slows down the spoilage process, using concepts related to rate of reaction, activation energy. Include a graphical representation of the effect of temperature on the rate of reaction.
2. Consider the following reaction in aqueous solution:

   $5Br^{-}(aq) + BrO^{3-}(aq) + 6H^{+}(aq) \rightarrow 3Br_2(aq) + 3H_2O(l)$. If the rate of disappearance of $Br^{-}(aq)$ at a particular moment during the reaction is 3.5×10^{-4} [[OMML-EQ-163]], what is the rate of appearance of $Br_2(aq)$ at that moment?

Hint

The recommended mode of assessment for week 4 is ***short quiz****. Use the level 2 question 2 as a sample question.*

WEEK 5

Learning Indicators

1. *Describe the collision theory and use it to explain qualitatively, the effects of the factors that affect the rate of reaction*
2. *Construct rate equations from experimental data, analyse graphs to deduce the order of reactions and solve problems involving the rate expression*

FOCAL AREA 1: COLLISION THEORY

The collision theory is a set of principles based around the idea that reactants form products when they collide with one another, but this only happens when the collisions have enough kinetic energy and are properly aligned to cause a reaction.

The collision theory states that

1. For particles to react with each other, they must collide in the correct orientation and with energy greater than or equal to the activation energy of the reaction.
2. The activation energy is the minimum energy that the reacting particles must possess for a successful collision to take place.
3. Every reaction has a different specific activation energy.
4. If the collision results in the reactants changing into products, it is said to be an effective (or successful) collision.

Explanation of the factors that affect rate of reactions using the collision theory

1. **Temperature:** Increasing temperature increases the rate of all chemical reactions. This is because the particles will have more kinetic energy than the required activation energy and this will lead to more frequent and effective collisions per unit time. A higher proportion of the reactants will have their kinetic energy greater than the activation energy of the reaction, the higher the fraction of molecules with their kinetic energy greater than the activation energy of the reaction, the higher the rate of the reaction.
2. **Available surface area and nature of reactant:** the larger the surface area of the reactant the faster the rate of reaction. This is because more surface area of the particles will be exposed to the other reactants, producing a higher number of collisions per time. Of these collisions the number of successful ones will increase as the total number of collisions itself has increased. Higher number of collisions per unit time increases the rate of the reaction.

3. **Concentration:** increasing concentration increasing the rate of reaction. This is because there will be more reactant particles per unit volume of reaction vessel, allowing more frequent and effective collisions per unit time, increasing the rate of reaction. For a gaseous reaction, increasing the pressure has the same effect as the same number of particles will occupy a smaller space, increasing the concentration

   If you double the number of particles, you will double the number of collisions per second.

   The number of collisions is proportional to the number of particles present.

4. **Catalyst:** Different processes require different types of catalysts but they all work on the same principle of providing a different pathway for the reaction to occur. Catalyst provides a new reaction pathway with a lower activation energy such that more reactants will have their kinetic energy greater than the activation energy of the reaction. This means a higher proportion of the reactant particles have energy greater than the activation energy and will result in more effective collisions per unit time.

Maxwell-Boltzmann energy distribution curve

The effect of temperature on rate of reaction can be shown in a Maxwell-Boltzmann distribution curve. A typical curve looks like this:

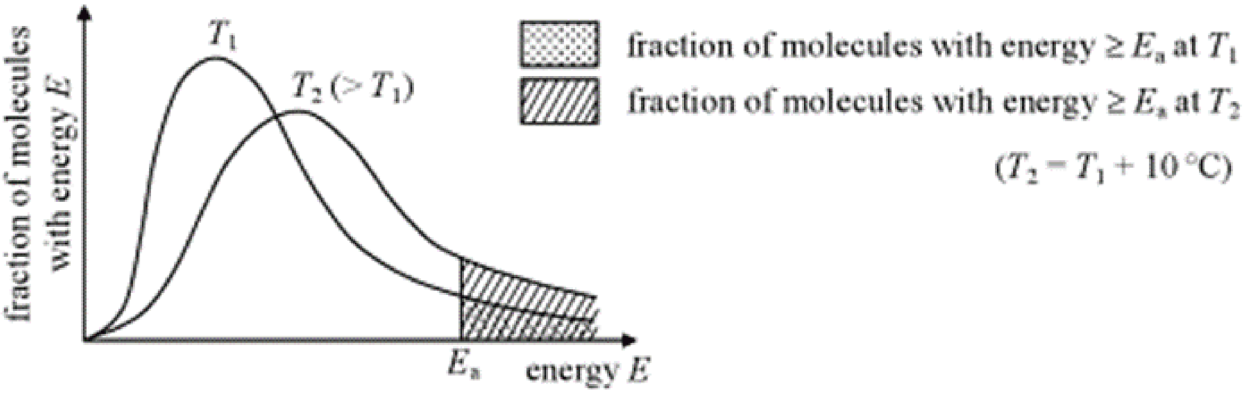


Figure 2.4: *Maxwell-Boltzmann energy distribution curve*

The area under the curve represents the total number of molecules.

There are more particles with energy greater than the activation energy at T_2. In drawing the curve at a higher temperature, the height of the curve is reduced so that the area under the curve remains constant.

Increasing the temperature means an increase in the kinetic energy and this will result in many collisions per unit time. Also, a greater fraction of the particles has energy above the activation energy (as seen in the Boltzmann above) meaning more of the collisions will result in a reaction. This means there will be more frequent effective collisions thereby increasing the rate of the reaction.

Learning Tasks

1. Identify and list the factors that affect rate of reactions according to collision theory.
2. Explain how the Maxwell-Boltzmann energy distribution curve relates to collision theory.
3. Compare and contrast the effects of the two factors on the rate of reaction, using graphs and numerical examples to support your argument.

PEDAGOGICAL EXEMPLARS

1. Collaborative learning

a. Divide the class into small groups mix-ability groups and ask each group to brainstorm and list all the interactions they can think of that occur in a chemical reacting system (e.g., atoms/molecules colliding, bonds forming/breaking, energy transfer). Encourage wild and creative ideas.

b. Ask each group to share their list with the class and write all the interactions on the board or a large piece of paper.

c. Facilitate a class discussion to reach a consensus on the most important interactions.

d. Introduce the collision theory, highlighting how it explains the interactions in a chemical reacting system.

e. Write the main points on the board.

f. Use learner-generated examples from the brainstorming session to illustrate each point.

g. Ask each group to create a concept map or diagram illustrating the collision theory, using the main points and examples from the class discussion.

2. Activity-Based learning

a. Divide learners into small mixed-ability groups and assign each group a scenario:

  i. Particles with different energies and orientations.

  ii. Particles with different surface areas.

  iii. Particles with different concentrations.

b. Ask them to role-play the collision theory, simulating how particles interact and react. Encourage learners to discuss and observe how different factors affect the collisions and reactions.

c. Provide learners with scenarios related to the factors that affect the rate of reaction.

d. Ask them to predict and explain, using the collision theory, how each factor will qualitatively affect the rate of reaction. Encourage them to use diagrams and sketches to support their explanations.

e. Introduce the Maxwell-Boltzmann energy distribution curve and ask learners to sketch the curve and explain how it relates to the collision theory.

f. Facilitate a whole-class discussion to Summarise the key points and takeaways**.**

KEY ASSESSMENTS

Level 1: What is the collision theory, and how does it explain the rate of a chemical reaction?

Level 2: Use the Maxwell-Boltzmann energy distribution curve to explain why an increase in temperature increases the rate of a chemical reaction.

Level 3: Compare and contrast the effects of increasing concentration and increasing temperature on the rate of a chemical reaction, using collision theory and the Maxwell-Boltzmann energy distribution curve to support your answer.

FOCAL AREA 2: RATE EQUATIONS AND ORDER OF REACTIONS

1. Rate Equation

The rate equation or rate law for a chemical reaction is an expression that provides a relationship between the rate of the reaction and the concentration of the reactants participating in it. Consider the hypothetical chemical reaction given by:

$aA + bB \rightarrow cC + dD$

Where a, b, c and d are the stoichiometric coefficients of the reactants and products, the rate equation for the reaction is given by: $Rate\ \alpha\ [A]^x[B]^y = \gg Rate = k[A]^x[B]^y$

Where; [A] and [B] denotes the concentrations of reactants A and B. 'x' and 'y' denote the orders with respect to reactants A and B (which may or may not be equal to their stoichiometric coefficients a and b

The proportionality constant 'k' is the rate constant of the reaction.

2. Order of reaction

Order of a reaction with respect to each reactant is defined as the exponent to which the concentration term of that reactant in the rate law is raised.

Consider the general reaction: $aA + bB \rightarrow cC + dD$ and suppose the rate expression for this reaction is: $rate = k[A]^x[B]^y$, The order with respect to A and B are x *and* y respectively.

a. Overall order of a reaction

The overall order of a reaction is the sum of the exponents to which the concentrations of all the reactants in the rate law are raised. It is found by adding up the orders of all the reactants involved in the reaction. Using the example above, *overall order of the reaction* $= x + y$

b. Zero order reaction

For a chemical reaction in the form $aA \rightarrow product$, if the order with respect to A is a zero order, the rate law for the reaction is:

$rate = k[A]^0 = \gg rate = k(1) = k.$

This means that for a zero-order reaction, the rate is constant. It does not change with concentration. The rate is independent of the concentration of the reactants, meaning an increase in concentration of the reactant will not translate to an increase in rate of the reaction. It is worth noting that a zero-order reaction will obviously change rate when all of the reactant is exhausted (the reaction cannot continue at the constant rate past completion).

c. First-order reaction

For a chemical reaction in the form $aA \rightarrow product$, if the order with respect to A is a first order, $rate = -\frac{\Delta[A]}{\Delta t} = k[A]$, For a first order reaction, doubling the concentration of a first order reaction causes the rate to double.

d. Second order reaction

For a chemical reaction of the form $aA \rightarrow product$, if the order with respect to A is a second order, $rate = -\frac{\Delta[A]}{\Delta t} = k[A]^2$, doubling the concentration of a second order reaction causes the rate to quadruple.

e. Half-life of a reaction

The time required for a reactant to reach half its original concentration is called the half-life of a reactant and is designated by the symbol $\mathbf{t_{\frac{1}{2}}}$. For first order reaction, $t_{\frac{1}{2}} = \frac{0.693}{k}$.

For a second order reaction, $t_{\frac{1}{2}} = \frac{1}{[A]_0}$.

3. Rate determining step

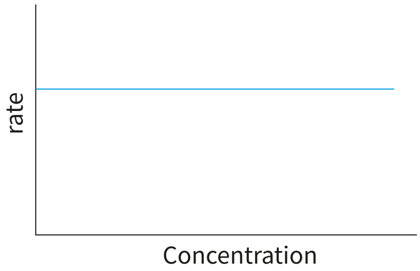
Reactions may occur all at once or through several discrete steps. Each of these processes is known as an elementary reaction. The slowest step among those elementary reaction steps determines the rate of reaction. Rate determining step can be determined by setting up reaction mechanisms, that is the sequence of events that describes the actual process by which reactants become products.

Rate determining step is therefore defined as the slowest step in a sequence of reactions steps leading to the formation of a product.

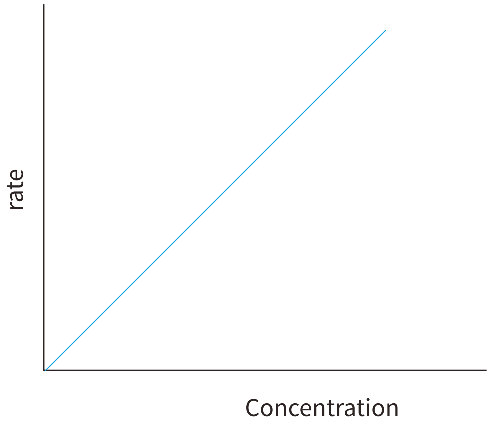
Graphical Determination of Reaction Order

In chemical kinetics, the order of a reaction describes how the rate of the reaction depends on the concentration of reactants. Graphical analysis allows us to determine the order of a reaction with respect to a specific reactant. Here's how it's done:

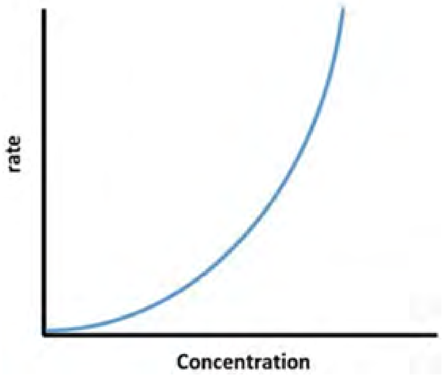
1. **Experimental Data Collection:** Conduct an investigation where the concentration of the reactant is monitored over time. Record data points of concentration at different time intervals.
2. **Plotting the Concentration vs. Time Graph:** Create a graph with concentration (usually on the y-axis) against time (on the x-axis). The resulting curve represents the change in concentration as the reaction progresses.
3. **Drawing Tangents:** Draw tangents (straight lines) to the curve at various points. These tangents represent the instantaneous rates of reaction at those concentrations.
4. **Calculating Tangent Gradients:** Calculate the gradient (slope) of each tangent line. The gradient at a specific concentration corresponds to the rate of reaction at that concentration.
5. **Interpreting the Graph Shape:** The shape of the rate against concentration graph provides clues about the reaction order:
    a. **Zero Order:** A straight horizontal line indicates zero order with respect to the reactant.



    b. **First Order:** A linear (straight-line) graph suggests first order.



c. **Second Order:** A curved graph (concave upward) indicates second order.



4. Calculations involving the rate law expression

Example 1

Consider data for the reaction: $A + B \rightarrow products$

Experiment	[A] / moldm^{-3}	[B] / moldm^{-3}	Rate / moldm^{-3} s^{-1}
1	0.001	0.002	5×10^{-4}
2	0.002	0.002	10×10^{-4}
3	0.001	0.004	20×10^{-4}

Determine;

a. The rate law

b. The specific rate constant k

Solution

a. **exponents separately**

Rate = $k[A]^x[B]^y$, Substituting data from experiment 1 and 2 into the general rate equation,

$5 \times 10^{-4} = k(0.001)^x(0.002)^y$……………….equation 1.

$10 \times 10^{-4} = k(0.002)^x(0.002)^y$……………….equation 2.

Divide equation 2 by1

$$\frac{10 \times 10^{-4}}{5 \times 10^{-4}} = \frac{k(0.002)^x(0.002)^y}{k(0.001)^x(0.002)^y}$$

$$2 = 2^x = \gg x = 1$$

also, substituting data from experiment 1 and 3 into the rate law expresion

$5 \times 10^{-4} = k(0.001)^x(0.002)^y$……………….equation 3

$10 \times 10^{-4} = k(0.001)^x(0.004)^y$………………..equation 4

Divide equation equation 4 by 3

$$\frac{20 \times 10^{-4}}{5 \times 10^{-4}} = \frac{k(0.001)^x(0.004)^y}{k(0.001)^x(0.002)^y}$$

$$4 = 2^y = \gg 2^2 = 2^y = \gg y = 2$$

Step two: Combine the exponents into a rate law

$\text{Rate} = k[A]^x[B]^y = \gg \text{rate} = k[A][B]^2$

b. Substitute one data set to find k

Substituting data from experiment one into the rate equation

$\text{rate} = k[A][B]^2$

$$= \gg 5 \times 10^{-4} = k(0.001)(0.002)^2$$

$$= \gg k = \frac{5 \times 10^{-4}}{(0.001)(0.002)^2}$$

$$= \gg k = 1.25 \times 10^5$$

$$\text{Units of } k = \frac{\text{moldm}^{-3}\,\text{s}^{-1}}{(\text{moldm}^{-3})(\text{moldm}^{-3} \times \text{moldm}^{-3})}$$

$$k = \text{mol}^{-2}\,\text{dm}^{-6}\,\text{s}^{-1}$$

$$\therefore k = 1.25 \times 10^5\ \text{mol}^{-2}\,\text{dm}^{-6}\,\text{s}^{-1}$$

Learning Tasks

1. Define activation energy (Ea) and its role in collision theory.
2. Create a table listing factors that influence reaction rates (e.g., concentration, surface area, catalysts) and for each factor and explain how it impacts the rate of a chemical reaction.
3. Compare the reaction rates at different temperatures using the Maxwell-Boltzmann distribution. Analyse how the curve changes when temperature increases or decreases.
4. Present a graph of concentration vs. time for a reaction and ask learners to determine the order of reaction based on the slope of the graph.

PEDAGOGICAL EXEMPLARS

1. Collaborative learning

a. Divide learners into mixed-gender groups where applicable and ensure each group has a mix of strengths and abilities. Encourage learners to accept diverse views and ideas from group members.

  b. Ask each group to brainstorm and come up with a definition for rate constant (k), order of reaction (n) and rate Determining Step (RDS). Encourage creative ideas.

  c. Ask each group to share their definition with the class.

  d. Facilitate a class discussion to reach standards definition of the terminologies.

  e. Encourage learners to identify what they learned from group members and what they would like to learn more about.

  f. Provide guidance and support to learners who may need more explanations.

2. Problem solving approach

  a. Provide learners with experimental data for a reaction and ask them to deduce the rate expression. Encourage learners to work in pairs and share their answers with the class.

  b. Ask learners to sketch graphical representations for:

    i. Zero-order reactions

    ii. First-order reactions

    iii. Second-order reactions and have them work in pairs, then share their answers with the class.

  c. Provide learners with data for a first-order reaction and ask them to calculate the half-life and use it to predict the order of reaction, encourage learners to work in pairs and share their answers with the class.

  d. Provide learners with various problems involving rate of reaction, such as:

    i. Finding the rate constant (k)

    ii. Determining the order of reaction

    iii. Calculating the half-life

  e. Have them work in pairs to solve the problems and share their answers with the class.

KEY ASSESSMENTS

Level 1: Why does increasing the concentration of a solution increase the rate of reaction?

Level 2

1. Why do catalysts increase the rate of reaction?
2. Give any 2 of the main ideas of the collision theory of matter in chemistry.

Level 3: Describe the Maxwell-Boltzmann distribution curve and explain how you can estimate the fraction of particles with sufficient energy to react.

Hint

The recommended mode of assessment for week 5 is ***class exercise****. Use the level 2 question 2 as a sample question.*

Section 2 Review

In this section, learners deepened their knowledge and understanding of the concept of chemical kinetics, which explores the varying rates at which chemical reactions occur under different conditions. Key terms in chemical kinetics were introduced, and learners engaged in practical experiments to measure reaction rates accurately. The collision theory was elaborated upon, serving as a foundation to explain how certain factors influence reaction rates qualitatively. Students learned to construct rate equations based on experimental data, analyse graphs to determine the order of reactions, and solve problems relating to rate expressions. Through this comprehensive approach, learners gained a thorough grasp of chemical kinetics, paving the way for a greater appreciation of reaction mechanisms and their applications.

SECTION 3: DYNAMIC EQUILIBRIUM

Strand: Physical Chemistry

Sub-Strand: Equilibria

Learning Outcome: *Explain that dynamic equilibrium is attained when the rate of the forward and backward reactions is equal and this principle has industrial applications.*

Content Standard: Demonstrate understanding that a balance of opposing reactions occurs in chemical equilibrium systems to attain equilibrium and this principle is applied in chemical industrial processes.

Hint

Remind learners of ***Mid-semester examination*** *in week 6. Refer to the Appendix C for aTable of Specification to guide you to construct your test items.*

INTRODUCTION AND SECTION SUMMARY

In week six, learners will explore reversible and irreversible reactions through a video and apply these concepts across various processes. They will use graphical methods to understand dynamic equilibrium and conduct experiments with reagents like anhydrous copper (II) tetraoxosulphate (VI).

Week seven will see learners in mixed-ability groups deducing equilibrium constant expressions using the law of mass action and discussing its significance. They will perform calculations to determine equilibrium concentrations and pressures and equilibrium constants and relate K_p to K_c using the ideal gas equation.

During week eight, students will identify factors affecting chemical equilibrium, applying Le Chatelier's principle to predict system shifts under different stresses. They will also discuss the impact of these factors on equilibrium and write down conditions to optimise reaction outcomes.

The weeks covered by the section are:

Week 6: *Dynamic equilibrium*

Week 7: *Equilibrium expressions*

Week 8: *Le Chatelier's principle*

SUMMARY OF PEDAGOGICAL EXEMPLARS

Week six introduces Participatory Learning, where students will analyse reversible and irreversible reactions through video analysis and graphical methods to understand dynamic equilibrium. This week also incorporates Exploratory Learning, encouraging learners to conduct experiments with reagents like anhydrous copper (II) tetraoxosulphate (VI).

Week seven focuses on Activity-Based Learning, with mixed-ability groups deducing equilibrium constant expressions and discussing the significance of these constants. Learners will engage in calculations to determine equilibrium concentrations and constants, and establish the relationship between K_p and K_c.

Week eight employs the Talk-for-learning approach, identifying factors affecting chemical equilibrium and applying Le Chatelier's principle to predict system shifts. Additionally, Activity-Based Learning will also be used to guide learners to optimise reaction conditions for desired outcomes. This approach aims to deepen understanding through active participation and collaborative exploration.

ASSESSMENT SUMMARY

To facilitate comprehensive assessment of learners on dynamic chemical equilibrium and its application in industrial processes, a variety of formative assessment methods will be utilised, including short answers, oral and written presentations, identifying factors that affect equilibrium, deducing equilibrium constant expressions, graph plotting, predicting the yield of products, measurements, demonstrations, and homework assignments. Projects and group activities will be employed to assess learners' comprehension and application of concepts, tailored to cover levels 1, 2, 3, and 4 of the Depth of Knowledge (DOK) framework while emphasising critical thinking, communication, and creativity. Class exercises, measurements, and group tasks will allow learners to demonstrate their grasp of concepts, while assignments and practical group activities will provide opportunities for practical application of knowledge in real-world contexts.

WEEK 6

Learning Indicator: *Explain the terms reversible and irreversible reactions and dynamic equilibrium*

FOCAL AREA: DYNAMIC EQUILIBRIUM

1. Reversible reactions

Some reactions go to completion, where the reactants are used up to form the product and the reaction stops when all of the reactants are used up. However, in some reactions when the reactants react to form the products, the products can themselves react with each other or decompose and form the reactants again. In this case the reaction can occur in both directions with the forward reaction forming the products and the reverse reaction forming the reactants.

A reaction is said to be reversible if under certain conditions of temperature and pressure, the forward and reverse reactions occur simultaneously. Reversible reactions are indicated by placing two half arrows ($\rightleftharpoons$) pointing in opposite directions between the reactants and products.

2. Irreversible reactions

Most chemical reactions occur only in one direction. They are called irreversible reactions. For example, when carbon is combusted in air to form carbon dioxide the reaction goes only in one direction i.e. in the direction of formation of carbon dioxide:

$C(s) + O_2(g) \rightarrow CO_2(g)$.

Irreversible reactions are indicated by placing an arrow ($\rightarrow$) pointing in one direction between the reactants and products.

3. Dynamic equilibrium

Many chemical reactions do not go to completion but instead attain a state of chemical equilibrium.

Consider the reaction $NO_2(g) + NO_2(g) \rightarrow N_2O_4(g)$. The reactant NO_2 is a dark brown gas and the product N_2O_4 is a colourless gas. When reactant is placed in an evacuated sealed glass vessel at 258 ^{0}C, the initial dark brown colour decreases in intensity as it is changed to the product N_2O_4 which is colourless. However, even over a long period of time, the contents of the reaction vessel do not become colourless, instead the intensity of the brown colour eventually becomes constant, which means that the concentration of NO_2 is no longer changing.

This observation is a clear indication that the reaction appears to have stopped. In fact, the system has reached its equilibrium the state where the concentrations of all reactants and products remained constant with time.

The reaction could better be described with the following equation:

$$NO_2(g) + NO_2(g) \rightleftharpoons N_2O_4(g)$$

Chemical equilibrium is a state in which the rates of the forward and reverse reactions are equal and the concentrations of the reactants and products remain constant. Equilibrium is a dynamic process where the conversions of reactants to products and products to reactants are always on going, although there is no net change in the number of reactant and product molecules.

Dynamic equilibrium can therefore be defined as the steady state of a reversible reaction occurring in a closed system.

Characteristics of dynamic equilibrium

1. The equilibrium can be achieved if the reversible reaction occurs in a closed vessel.
2. At equilibrium, both the reactants and product are present.
3. The forward and backward reactions move in the opposite direction with equal rates.
4. The equilibrium state can be approached from either direction.
5. The concentration of reactants and products, along with their temperature, pressure, and colour, remains unchanged.

Applications of dynamic equilibrium

1. Chemical systems: Rechargeable batteries lithium-ion batteries operate based on chemical equilibrium. The reversible reactions between lithium ions and electrode materials store and release electrical energy.
2. Biological Systems: Many biological processes operate in dynamic equilibrium. Glucose levels in the body remain relatively stable despite its continuous use. The liver stores glucose as glycogen and releases it into the bloodstream as needed, thereby maintaining dynamic equilibrium.
3. Environmental Processes: Processes like photosynthesis, respiration, and decomposition contribute to dynamic equilibrium in carbon cycling, etc.

Learning Tasks

1. Provide learners with a list of everyday processes (e.g., melting ice, boiling water, rusting of iron, etc.) and ask them to classify each process as either reversible or irreversible. Have learners discuss their reasoning behind each classification.
2. Show learners a concentration-versus-time graph for a reversible reaction. Ask them to identify when the equilibrium is established. Have them discuss how the graph changes for irreversible reactions.

3. Discuss how these processes maintain balance:
   a. Haemoglobin binding and releasing oxygen in blood.
   b. Water evaporation and condensation.

PEDAGOGICAL EXEMPLARS

1. **Participatory Learning**
   - **a.** In a whole inclusive class, show a short video illustrating reversible and irreversible reactions.
   - **b.** Define reversible reactions and explain that they can proceed in both forward and reverse directions.
   - **c.** Provide examples of reversible reactions such as photosynthesis/respiration, acid-base neutralisation and use simple equations to illustrate the forward and reverse reactions.
   - **d.** Introduce the concept of dynamic equilibrium, where forward and reverse reactions occur at equal rates and use a graphical method (e.g., graphing software or a whiteboard) to illustrate the concept and emphasise that equilibrium is not static, but dynamic.
   - **e.** Use examples from the previous section to illustrate dynamic equilibrium in different contexts.
2. **Exploratory Learning**
   - **a.** Guide learners conduct an experiment on the reversible reaction between anhydrous copper(II) tetraoxosulphate (VI) and water or any other workable reagent.
   - **b.** Emphasise observation, recording, and analysis of data.
   - **c.** Relate the experiment to real-life scenarios or applications.

KEY ASSESSMENTS

Level 1: What is a reversible reaction?

Level 2: What do you understand by the term dynamic equilibrium?

Level 3: Describe how the liver maintains a dynamic equilibrium in the release of sugar into the bloodstream.

Hint

The recommended mode of assessment for week 6 is **Mid-semester examination.** *Refer to the Appendix C for aTable of Specification to guide you to construct your test items.*

WEEK 7

Learning Indicator: *Deduce expressions for equilibrium constants in terms of concentrations and partial pressures and perform relevant calculations from appropriate data*

FOCAL AREA: EQUILIBRIUM CONSTANT EXPRESSIONS

1. Law of mass action

The equilibrium state of a reversible reaction is characterised by its equilibrium constant, which is related to the equilibrium concentrations of all the reactants and products.

The law of mass action states that, the rate at which a substance reacts is proportional to its active mass (molar concentration) and the overall rate of the reaction is proportional to the product of active masses (molar concentrations) of all the reacting substances.

Consider the equilibrium reaction $aA + bB \rightleftharpoons cC + dD$,

The rate of the forward reaction $r_f = k_f[A]^a[B]^b$ and

the rate of the reverse reaction $r_b = k_b[C]^c[D]^d$. Where r_f and r_b are the rate constants of the forward and backward reactions.

At equilibrium rate of forward reaction = rate of backward reaction, that is $r_f = r_b$

$= \gg k_f[A]^a[B]^b = k_b[C]^c[D]^d$

Rearranging the equations $\frac{k_f}{k_b} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$.

$\frac{k_f}{k_b}$ is equilibrium constant of the reaction. It is also called the equilibrium concentration constant and the symbol given is K_c, $= \gg Kc = \frac{[C]^c[D]^d}{[A]^a[B]^b}$.

2. Equilibrium constant expression in terms of concentrations (K_c) and partial pressures (K_p)

In general, for a reversible reaction such as $\boldsymbol{aA + bB \rightleftharpoons cC + dD}$ at equilibrium, the following ratio of concentration terms always remains constant at a given temperature.

$$\boldsymbol{Kc = \frac{[C]^c[D]^d}{[A]^a[B]^b}}$$

The above relation is known as the law of equilibrium, where $\boldsymbol{K_c}$ is the equilibrium concentration constant.

In case of gases their partial pressures can also be used in place of molar concentrations. The new equilibrium constant $\boldsymbol{K_p}$ is called the equilibrium pressure constant.

For the general reaction $\boldsymbol{aA(g) + bB(g) \rightleftharpoons cC(g) + dD(g)}$ in the gaseous phase

$$K_p = \frac{P_C^c \times P_D^d}{P_A^a \times P_B^b}$$

Example 1

The following reaction is at equilibrium $N_2 + 3H_2 \rightleftharpoons 2NH_3$

The K_c at a particular temperature is 13.7. If the equilibrium concentration N_2 is 1.88 moldm^{-3} and that of NH_3 is 6.62 moldm^{-3}, calculate the equilibrium concentration of H_2.

Solution

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

From the balance equilibrium reaction, $Kc = \frac{[NH_3]^2}{[N_2]\,[H_2]^3}$

Substituting for the known equilibrium concentrations and the K_c, this becomes

$$13.7 = \frac{(6.62)^2}{(1.88)\,[H_2]^3}$$

$$= \gg [H_2]^3 = \frac{(6.62)^2}{(1.88 \times 13.7)}$$

$$= \gg [H_2]^3 = 1.502112$$

$$= \gg [H] = \sqrt[3]{1.502112} \quad = 1.15$$

$\therefore$ *the concentration of H is* 1.15 *moldm*$^{-3}$

Example 2

Calculate the K_p for the reaction below given the equilibrium partial pressures of NO_2 is 0.664 atm and that of N_2O_4 is 1.09 atm

$2\,NO_2(g) \rightleftharpoons N_2O_4(g)$

Significance of equilibrium constants K_c and K_p

Kc represents the equilibrium constant based on molar concentrations of reactants and products. It helps predict the direction of a reaction at equilibrium.

a. If $K_c > 1$, the reaction favours products (the rate of the forward reaction is higher than the rate of the reverse reaction).

b. If $K_c < 1$, the reaction favours reactants (the rate of the forward reaction is lower than the rate of the reverse reaction).

c. When $K_c = 1$, the system is at equilibrium (the rate of the forward reaction is equal to the rate of the reverse reaction).

Real-world applications of K_c and K_p

a. **Ammonia Synthesis:** In the Haber process, ammonia (NH_3) is synthesised from nitrogen (N_2) and hydrogen (H_2) gases. The equilibrium expression involves K_p or

K_c. Optimizing reaction conditions (temperature, pressure) to maximise ammonia yield relies on understanding these constants.

b. **Airbag Deployment:** The decomposition of sodium azide (NaN_3) in airbags generates nitrogen gas (N_2). The equilibrium constant helps engineers design airbag systems for rapid and safe inflation.

c. **Petroleum Refining:** Reactions involving hydrocarbons, such as cracking and reforming, occur at equilibrium. K_p and K_c guide refinery operations to achieve desired product ratios, etc.

In summary, K_c and K_p provide understanding into the position of equilibrium and help us understand how reactants and products coexist in a reversible reaction.

3. Relationship between K_c and K_p using ideal gas equation

The relationship between K_c and K_p for a particular reaction follows the fact that for an ideal gas, $PV = nRT$, dividing both sides by volume, $\left(V\right) \Rightarrow\gg P = \left(\frac{n}{V}\right)RT$

$= \gg P = [\ ]RT$, where $[\ ] =$ concentration.

Consider the equilibrium reaction $aA(g) + bB(g) \rightleftharpoons cC(g) + dD(g)$

Equilibrium constant Kp in terms of partial pressures

$$K_p = \frac{P C c \times P_D^d}{P_A^a \times P_B^b} \text{........................equation 1}$$

Substituting $P = [\ ]RT$ into equation **1**

$$K_p = \frac{([C]RT)^c \times ([D]RT)^d}{([A]RT)^a \times ([B]RT)^b}$$

$$K_p = \frac{[C]^c \times RT^c \times [D]^d \times (RT)^d}{[A]^a \times RT^a \times [B]^b \times (RT)^b}$$

$$K_p = \frac{[C]^c \times [D]^d \times (RT)^{c+d}}{[A]^a \times [B]^b \times (RT)^{a+b}}$$

$$K_p = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b} \times \frac{(RT)^{c+d}}{(RT)^{a+b}}$$

$$K_p = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b} \times (RT)^{(c+d)-(a+b)}$$

but Kc expresion for the same equilibrium reaction, $K_c = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$

$$= \gg K_p = K_c \times \left(RT\right)^{(c+d)-(a+b)}$$

Note

a and b are the coefficients of reactants A and B and c and d are the coefficients of products C and D.

$Let\ (c + d) - (a + b) = \Delta n$(change in moles of gas)

Where $(c + d)$ = total number of moles of the products in the equilibrium equation

and $(a + b)$ = total number of moles of the reactants in the equilibrium equation.

$=\gg K_p = K_c(RT)^{\Delta n}$

Example 1

Write the relationship between K_c *and* K_p for each of the following equilibrium reactions:

a. $2A(g) + 3B(g) \rightleftharpoons C(g) + 2D(g)$

b. $2NO(g) + Cl_2(g) \rightleftharpoons 2\ NOCl(g)$

c. $N_2 + 3H2 \rightleftharpoons 2NH_3$

Solution

a. $2A + 3B \rightleftharpoons C + 2D$

$K_p = K_c(RT)^{\Delta n}$

Δn (change in number of moles) = total number of moles of products – total number of moles of reactants.

$\Delta n = (1+2) - (2+3)$

$= 3-5$

$\Delta n = -2$

$K_p = K_c(RT)^{-2}$

$\therefore K_p = \frac{K_C}{(RT)^2}$

b. $2NO(g) + Cl_2(g) \rightleftharpoons 2\ NOCl(g)$

Δn (change in number of moles) = total number of moles of products – total number of moles of reactants.

$\Delta n = (2) - (2+1)$

$= -1$

$K_p = K_c(RT)^{-1}$

$\therefore K_p = \frac{K_C}{RT}$

c. $N_2 + 3H_2 \rightleftharpoons 2NH_3$

Δn (change in number of moles) = total number of moles of products – total number of moles of reactants.

$$\Delta n = (2) - (1 + 3)$$

$$= -2$$

$$K_p = K_c(RT)^{-2}$$

$$\therefore K_p = \frac{K_C}{(RT)^2}$$

Exercise 1

Use the equilibrium reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ to answer the questions that follows:

a. Write an equilibrium expression K_c for the reaction

b. Derive K_p expression for the reaction in terms of K_c

c. Given that K_c is $3 \times 10^3\,dm^3\,mol^{-1}$ at room temperature, calculate the value of K_p, take $R = 8.314\ Jmol^{-1}K^{-1}$

Note

The equilibrium constants K_c and K_p have units unless the units cancel, that is when the sum of the powers in the numerator and the denominator are the same.

4. K_c and K_p for homogeneous and heterogeneous equilibrium reactions

Many equilibrium reactions may involve more than one phase, those reactions are called **heterogeneous equilibrium** reactions whilst those that have one phase like the examples shown in exercise 1 above are **homogeneous equilibrium** reactions.

Example the thermal decomposition of calcium carbonate in commercial preparation of lime occurs by a reaction involving both solid and gas phase:

$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$, by the law of mass action;

$K_c = \dfrac{[CaO] \times [CO_2]}{[CaCO_3]}$, however, experimental results has shown that the position of a heterogenous equilibrium does not depend on the amount of pure solids or liquids present. The reason is that the concentrations of pure solid and liquids cannot change.

For that matter, $K_c = \dfrac{[CO_2] \times C_1}{C_2}$ and since C_1 *and* C_2 are constants, the expression can be re-writing as $K_c = [CO_2]$. This reasoning also applies to K_p.

Example

Write the K_c and K_p expressions for the following processes

a. The decomposition of solid phosphorus pentachloride (PCl_5) to liquid phosphorus trichloride (PCl_3) and chlorine gas (Cl_2)

b. Deep blue solid copper (II) sulphate pentahydrate is heated to drive off water vapour to form white solid copper (II) sulphate.

Solutions

a. $PCl_5(s) \rightleftharpoons PCl_3(l) + Cl_2(g)$

$Kc = [Cl_2]$ and $Kp = P_{Cl_2}$

b. $CuSO_4.5H_2O \rightleftharpoons CuSO_4(s) + 5H_2O(g)$

$K_C = [H_2O]^5$ and $K_P = P_{(H_2O)}{}^5$

5. Calculations involving K_{sp} of sparingly soluble salts

The solubility product constant (K_{sp}) is an equilibrium constant that describes the solubility of sparingly soluble compounds. It is generally used for a sparingly soluble solute that does not totally dissolve. It signifies the extent to which the solute can go into the solution. The more soluble a solute, the higher the K_{sp} value. A solute is soluble if more than 1 g dissolves in 100 cm^3 of water.

When a soluble ionic compound is added to water, it disassociates into its constituent ions and goes into the solution. There is an equilibrium between the undissociated solid compound and the ionic solution.

Consider the dissolution of a substance A_xB_y in water. Its equilibrium reaction equation is given as: $A_xB_y(s) \rightleftharpoons xA^{y+}(aq) + yB^{x-}(aq)$

The solubility product constant, $K_{sp} = [A^{y+}]^x[B^{x-}]^y$

To calculate K_{sp} for this salt, one must raise the concentrations of the products to their stoichiometric coefficients and multiply them. The solid reactant $A_xB_y(s)$ is not included in the expression because its concentration does not change during the reaction. Hence, K_{sp} denotes the maximum extent that a solid can dissolve in a solution. Since it is the product of the ion concentrations present in a saturated solution of an ionic compound, K_{sp} is called the solubility product constant.

Example 1

Calcium fluoride (CaF_2) is slightly soluble in water and dissolves as follows:

$CaF_2(s) \leftrightharpoons Ca^{2+}(aq) + 2F^-(aq)$

The concentration of Ca^{2+} in a saturated solution of CaF_2 is 1.5×10^{-4} moldm^{-3}. What is the solubility product constant of CaF_2?

Solution

Given that the $[Ca^{2+}] = 1.5 \times 10^{-4}$ moldm^{-3} $= \gg [F^-] = 2 \text{ x } 1.5 \times 10^{-4}$ moldm^{-3} $= 3 \text{ x } 10^{-4}$ moldm^{-3}

$= \gg K_{sp} = [Ca^{2+}]\,[F^-]^2$

$= \gg K_{sp} = (1.5 \times 10^{-4})\,(3 \text{ x } 10^{-4})^2$

$= \gg K_{sp} = 1.35 \times 10^{-11}$

$\therefore$ the solubility product constant of $Ca\,F_2 = 1.35 \times 10^{-11}$

Example 2

Estimate the solubility of Ag_2CrO_4 in pure water if the solubility product constant for silver chromate is 1.1×10^{-12}.

Solution

$Ag_2CrO_4(s) \leftrightharpoons 2Ag^+(aq) + CrO_4^{\,2-}(aq)$

$K_{sp} = \left[Ag^+\right]^2\left[CrO_4^{\,2-}\right]$

Let the solubilty of $Ag_2CrO_4 = x\ moldm^{-3}$

$[Ag^+] = 2x$ and $\left[CrO_4^{\,2-}\right] = x$

Substituting the concentration of the ions into the K_{sp} expression

$K_{sp} = (2x)^2(x)$

$1.1 \times 10^{-12} = (4x^2)(x)$

$1.1 \times 10^{-12} = 4x^3$

$x^3 = \frac{1.1 \times 10^{-12}}{4}$

$x^3 = 2.75 \times 10^{-13}$

$x = \sqrt[3]{2.75 \times 10^{-13}}$

$x = 6.5 \times 10^{-5}$ moldm^{-3}

$\therefore$ The solubility of Ag_2CrO_4 in pure water is 6.5×10^{-5} moldm^{-3}

Example 3

Calculate the solubility of barium sulphate solution given that the solubility product constant for barium sulphate is 1.1×10^{-10}.

Solution

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

$K_{sp} = [Ba^{2+}][SO_4^{\ 2-}]$

$K_{sp} = 1.1 \times 10^{-10}$,

Note that there are two forms of $SO_4^{\ 2-}$, the one from $BaSO_4$ and the one from solution

let the amount of $BaSO_4$ that dissolved in the sodium sulphate = x $moldm^{-3}$

$=\gg$ Initial $[Ba^{2+}] = 0$ and final $[Ba^{2+}] = x$

Also initial $[SO_4^{\ 2-}] = 0.020\ moldm^{-3}$ from the Na_2SO_4 solution

final $[Ba^{2+}] = 0.020 + x$

$K_{sp} = (x)(0.020 + x)$

$=\gg 1.1 \times 10^{-10} = 0.020x + x^2$

$=\gg x^2 + 0.020x - 1.1 \times 10^{-10} = 0$

solving the quadratic equation will give $x = 5.5 \times 10^{-9}$ *and* -0.020.

The concentration of Ba^{2+} *can not negative so we pick* $x = 5.5 \times 10^{-9}$

$=\gg x = 5.5 \times 10^{-9}\ moldm^{-3}$

$\therefore$ *the solubilty of* $BaSO_4\ 5.5 \times 10^{-9}\ moldm^{-3}$

Learning Tasks

1. How would you know whether a system has reached the equilibrium state or not?
2. Give the meaning of the term **dynamic** in the context of a dynamic equilibrium.
3. In an experiment carried out at 298 K, 4.0 mol of NOCl were placed in a 2 litre flask and after the equilibrium was reached 1.32 mol of NO were formed. Calculate K_c at 298 K for the reaction $2NOCl\ (g) \leftrightharpoons 2NO\ (g) + Cl_2\ (g)$.

PEDAGOGICAL EXEMPLARS

1. Activity-Based Learning

a. Introduce the concept of the law of mass action in a whole inclusive class.

b. Divide learners into mixed-ability groups and ask them to deduce the expression for the equilibrium constant (K_c or K_p) using the law of mass action. Encourage groups to share their answers and discuss any differences.

c. Ask learners to think about the significance of the equilibrium constant and have them share their thoughts with a partner. Encourage pairs to share their answers with the class.

d. Provide learners with experimental data and initial quantities of reactants and ask them to perform calculations to determine.

a. Equilibrium concentrations or partial pressures.

b. Equilibrium constant (K_c or K_p). And encourage learners to work in pairs and share their answers with the class.

e. Use the ideal gas equation to establish the relationship between K_p and K_c and ask learners to perform calculations to demonstrate this relationship.

f. Introduce the concept of sparingly soluble salts and ask learners to perform calculations involving K_{sp}, given the solubility of a salt. Encourage learners to work in pairs and share their answers with the class.

KEY ASSESSMENTS

Level 1: What is the significance of the magnitude of equilibrium constant?

Level 2: Calculate the solubility product constant for lead (II) chloride, if 50.0 mL of a saturated solution of lead (II) chloride was found to 0.2207 g of lead (II) chloride dissolved in it.

Level 3

1. Samples of sulphur dioxide, oxygen and sulphur trioxide were added to a flask of volume 1.40 dm^3 and allowed to reach equilibrium at a given temperature. The flask contained 0.0550 mol of sulphur dioxide and 0.0720 mol of sulphur trioxide at equilibrium. The K_c has the numerical value of 27.9 under these conditions.
   - **a.** Calculate the amount, in moles, of oxygen gas in this equilibrium mixture.
   - **b.** The experiment in (a) was repeated with the same amounts of sulphur dioxide, oxygen and sulphur trioxide at the same temperature but in a smaller flask.

   The mixture was allowed to reach equilibrium. State the effect, if any, of using a smaller flask on the value of K_c.

2. You are a laboratory technician tasked with determining the solubility of the three salts namely silver chloride, Calcium carbonate and Copper (II) sulphate. Which of the salts is most soluble in water?

Hint

*The recommended mode of assessment for week 7 is **simulation**. Use the level 3 question 2 as a sample question. Refer to the Teacher Assessment Manual and Toolkit page 135 for additional information on how to go about this assessment mode.*

WEEK 8

Learning Indicators

1. *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*
2. *Apply the concepts of equilibrium and rates to industrial processes (Haber and Contact Process)*

FOCAL AREA 1: LE CHATELIER'S PRINCIPLE AND FACTORS THAT AFFECT EQUILIBRIUM POSITIONS

Le Chatelier's Principle

A system at equilibrium responds when a change is made that disturbs the system. Such a change might be a change in temperature, volume or pressure of the system, addition or removal of chemical species to or from the system, or addition of a catalyst. How a system responds to changes in such conditions is important as it may allow chemists to optimise reaction conditions to obtain the maximum amount of a desired product. Le Chatelier's Principle states that when a stress is applied to a chemical system at equilibrium, the equilibrium will shift to relieve the stress.

Factors that affect equilibrium positions

a. **Concentration**: Alteration of the concentration of a reaction mixture can be linked to the removal or addition of reactant or product. It can as well be linked to addition of entities that react with either the reactant or the product.

Addition of a reactant increases its concentration. The side of the reaction that consumes the added reactant will be favoured.

For a hypothetical equation at equilibrium which is represented by the equation, $A + B \rightleftharpoons C + D$. When more of the A is added to the mixture, some of the B that is still present in the mixture will react with the added A and equilibrium will shift to the right. This will reduce the concentration of A and produce more of the products.

Likewise, when the concentration of C is increased, the added concentration will be consumed by the quantity of D present in the mixture. This will shift the equilibrium to the left, producing more of the products.

On the other hand, when the concentration of a reactant (B) is reduced or when an entity that reacts with the reactant is added (reduces the quantity of that reactant in solution), more of the products (C and D) will react to form more of the entity that was reduced. So the equilibrium will shift to the left.

Similarly, a decrease in the concentration of the products will cause more of the reactants to react forming back more of the reduced product to restore equilibrium.

The equilibrium position of a chemical reaction shifts away from the side that has its entities been added in order to reduce its concentration or shift towards the side that has its entities removed so as to increase its concentration.

b. Pressure and Volume

From the equation $P = n{-}RT$

An increase in volume reduces pressure while a decrease in volume increases pressure. Changing the volume of a mixture of reacting gases changes molar concentrations and partial pressures of the gases, so changes in volume have some effect on the position of equilibrium.

Changes in pressure ordinarily do not affect the concentrations of reacting species in aqueous solution because liquids and solids are virtually incompressible. On the other hand, concentrations of gases are greatly affected by changes in pressure.

For a homogenous mixture of gases at equilibrium which can be represented by the hypothetical equation

$\mathrm{xA} \rightleftharpoons \mathrm{yB}$

where x and y are the stoichiometric values of the chemical equation at equilibrium

An increase in pressure (decrease in volume) favours the net reaction that decreases the total number of moles of gases. The nature of the effect depends on the total number of moles on either side of the equation.

A decrease in pressure (increase in volume) favours the side of the equation that has greater number of moles.

For reactions where there are equal number of moles for both reactants and products, a change in pressure (either increase or decrease) will have no effect on the position of the equilibrium.

The total pressure on a system can also be increased by addition of inert gases (example Helium). When an inert gas is added to a homogenous gaseous system at equilibrium, it increases the total gas pressure and decreases the mole fractions of both reactant(s) and product(s) but the partial pressure of each entity remain the same. Therefore, the presence of an inert gases will not affect the equilibrium position of the system.

c. Temperature

A change in concentration, pressure, or volume may alter the equilibrium position, that is, the relative amounts of reactants and products, but it does not change the value of the equilibrium constant. Only a change in temperature can alter the equilibrium constant. Prediction on the effect of temperature on equilibrium relies on thermochemical equation, whether the equation is endothermic or exothermic.

For endothermic reactions, heat is absorbed into the system and so heat is regarded as a reactant

$A + B \rightleftharpoons C \; \Delta H = +x \; KJ/mol$

$A + B + x \rightleftharpoons C$

When temperature is increased in an endothermic reaction, the reacting molecules absorb the heat which favours the forward reaction (shift from left to right) decreasing the concentration of the reactants and increasing the concentration of the products. A decrease in temperature conversely, favours the side that produces more heat. In this case the backward reaction will be favoured.

For exothermic reactions, heat is released from the system and so heat is regarded as a product

$D + E \rightleftharpoons F \; \Delta H = -y \; KJ/mol$

$D + E \rightleftharpoons F + y$

When temperature is increased in an exothermic reaction, favours the backward reaction which reduces the quantity of heat that is added to the system. That is equilibrium shifts from right to left decreasing the quantity of products and increasing the concentration of the reactants. A decrease in temperature on the other hand, favours the forward reaction since that side produces more heat to restore the equilibrium.

d. Catalyst

A catalyst increases the rate of a chemical reaction by providing an alternative pathway with a shorter activation energy for both the forward and backward reactions to the same extent. Therefore, the presence of a catalyst does not alter the equilibrium constant neither does it affect the equilibrium position of an equilibrium system.

Learning Tasks

1. State Le Chateliers principle.
2. The reaction $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$ is endothermic with $K_c = 5.67$ at 1500 °C. How will the amount of CO(g) at equilibrium be affected by
   - a. adding more $H_2O(g)$,
   - b. lowering the pressure by increasing the volume of the container,
   - c. raising the temperature of the reaction mixture.
   - d. adding a catalyst to the system.

e. Which of these will change the equilibrium constant and in which direction (increase or decrease) will K_c change?

PEDAGOGICAL EXEMPLARS

1. **Talk-for-learning approach**
   a. Identify the factors that affect chemical equilibrium and use Le Chatelier's principle to predict the direction of the shifting from various possible stress (these should include changes in concentration, temperature, pressure, volume and catalyst) on a system at equilibrium.
   b. In groups discuss and predict the effect of each of the factors on the position of equilibrium as well as the equilibrium constant.
2. **Activity-Based learning**
   Individually, identify and write down a set of conditions that will optimise a desired outcome of a reaction, such as a product.

KEY ASSESSMENTS

Level 1: State Le Chatelier's principle

Level 2: Explain how changes in temperature affect the equilibrium position of an endothermic reaction. Provide an example to support your explanation.

Level 3: The process for the production of ammonia is represented by the following equilibrium reaction equation:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$; $\Delta H = -92$ kJ/mol Identify the factors that could increase the yield of ammonia

FOCAL AREA 2: APPLICATION OF THE CONCEPTS OF EQUILIBRIUM AND RATES TO INDUSTRIAL PROCESSES

The concept of chemical equilibrium, factors that affect equilibrium and Le Chatelier's principle has helped many industries to maximise their yield. Chemists in key industries study the reaction steps and optimise the favourable conditions that increase the production of their desired products.

The manufacture of ammonia through the Haber process is a typical industrial process that applies the principles of equilibrium. The thermochemical equation for the production of ammonia from its constituent elements is represented as

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -92$ kJ/mol

This implies that the reaction is exothermic in nature. Because this Haber process sets a homogenous gaseous equilibrium, factors such as lowering temperature, increasing pressure and addition of iron catalysts are strategically manipulated to increase the yield of ammonia.

The contact process is another industrial method that relies on the principles of equilibrium. This involves a series of steps that leads to the production of tetraoxosulphate (VI) acid which is used in the laboratory as well as car batteries. The process basically involves five stages, which are;

a. Production of sulphur (IV) oxide from sulphur

b. Purification of sulphur (IV) oxide

c. Catalytic oxidation of sulphur (IV) oxide to sulphur (VI) oxide

d. Conversion of sulphur (VI) oxide to oleum

e. Dilution of oleum to tetraoxosulphate (VI) acid

Catalytic oxidation of SO_2 to SO_3 is the main chemical reaction in the entire process which takes place within a special unit known as the contact tower. In the contact tower, the clean dried gaseous mixture of SO_2 and air are passed through a tower that is loosely packed with vanadium (V) oxide or platinum catalyst. This is a reversible exothermic reaction that drives the production of SO_3 through the process described below:

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H = -\ 196\ kJ/mol$

The equilibrium between $SO_{2,}$ O_2 and SO_3 is carefully controlled to maximise the yield of SO_3 which is the intermediate for the production of H_2SO_4. Chemists skilfully manipulate factors such as low temperature, high pressure and the presence of V_2O_5 to shift the equilibrium position in favour of the formation of SO_3.

Industrialisation has brought about the use of different types of machines which depends on diverse sources as their fuel. Depending on the type or nature of fuel used, some harmful gases or pollutants are expelled into the environment which include carbon (II) oxide, hydrocarbons nitrogen (IV) oxide and nitrogen (II) oxide. Some of the primary sources of these air pollutants are incomplete combustion of fossil fuels in vehicle, combustion of coal in industries, emissions from petroleum evaporation during refuelling or from fuel storage tanks, high temperature combustion in diesel engines as well as combustion of fossil fuels in power plants.

These pollutants are converted to less harmful substances such as carbon (IV) oxide, water and nitrogen gas with the help of a **catalytic converter.** Knowledge of equilibrium are essential in ensuring in ensuring that these reactions proceed efficiently to completion.

The essential part of the catalytic converter with regards to equilibrium is the honeycomb structure where these harmful gases are converted to less harmful gases. It is basically made up of two steps, the reduction catalytic process and the oxidation catalytic process. The thermochemical equation of the processes can be represented as;

Reduction catalytic process

$2NO(g) \rightleftharpoons N_2(g) + O_2(g)\ \Delta H = +x$ kJ/mol (Pt catalyst)

$2NO_2(g) \rightleftharpoons N_2(g) + 2O_2(g)\ \Delta H = +x$ kJ/mol

Oxidation catalytic process

$2CO(g) + O_2(g) \rightleftharpoons 2CO_2(g)\ \Delta H = +x$ kJ/mol (Pt, Pd/Rh)

$C_xH_y\ (g) + O_2(g) \rightleftharpoons CO_2(g) + H_2O(g)\ \Delta H = +x$ kJ/mol

As the engine starts the fuel is burnt to produce more of the poisonous gases which then increases in concentration. According to Le Chatelier's principle, an increase in the concentration of reactant molecules in an equilibrium process favours the forward reaction which drives the productions of the less harmful gases. This is an endothermic process which requires heat. To solve this, catalytic converters are strategically positioned close to the engine in the exhaust system in order to benefit from the high temperatures of the exhaust gases when the engine starts. This proximity ensures that the catalysts reach their optimal operating temperatures quickly and efficiently while inducing the equilibrium condition of high temperatures that favours the formation of the environmentally friendly gases (such as carbon (IV) oxide, nitrogen gas and water).

In petroleum refining, equilibrium concepts play a crucial role in transforming hydrocarbons into valuable products such as petrol (gasoline), diesel and other fossil fuels through catalytic cracking and reforming.

In catalytic cracking, long-chain hydrocarbons are broken down into shorter chain hydrocarbons by the use of catalysts. Equilibrium concepts help determine the optimal conditions that drive the formation of a desired product. It requires bonds breaking so this process is an endothermic process that can be represented as;

$C_{12}H_{26}(g) \rightleftharpoons C_6H_{14}(g) + C_7H_{16}(g)\ \Delta H = +x$ kJ/mol

To optimise the production of the desired short-chain hydrocarbons, a high temperature should be maintained. This is provided by a furnace and transferred to the cracking unit through heat exchangers.

In catalytic reforming, Le Chatelier's principle is applied to optimise the conversion of hydrocarbon molecules into high-octane petrol (gasoline). In the context of catalytic reforming, chemist adjust the process conditions to favour the desired reactions (high-octane petrol) and increase product yield. Reforming involves the rearrangement of hydrocarbon molecules to produce petrol of high-octane ratings, aromatic hydrocarbons, hydrogen gas, ethylene, propylene as well as cyclic hydrocarbons. Petrol of high-octane ratings are much desired in the automotive industry due to its ability to resist knocking in internal combustion engines (it has anti-knocking properties).

One of the key processes in the reforming process is dehydrogenation of naphthene's and paraffin to form aromatic compounds. This is an endothermic reaction which requires heat in order to favour the formation of the aromatic compounds and hydrogen gas. The equation for this process can be represented as;

$C_6H_{12}(g) \rightleftharpoons C_6H_6(g) + 3H_2(g)\ \Delta H = +\ x$ kJ/mol

Equilibrium concepts are also applied in water treatment particularly during the ion exchange used for softening water. This is the process whereby hardness-causing ions such as calcium and magnesium are removed from water. Within the ion-exchange process particularly during the regenerative process calcium ions are set in equilibrium with chloride ions according to the equation;

$Ca^{2+}_{(aq)} + 2Cl^{-}_{(aq)} \rightleftharpoons CaCl_{2(s)}$

This step is characterised by the addition of brine (concentrated NaCl) which disturbs the equilibrium by increasing the chloride ions. According to Le Chatelier's principle an increase in the concentration of reactants favours the formation of products. This then ensure that more calcium ions and magnesium ions are removed to combine with the increased chloride ions to produce calcium chloride solids. The exchange of sodium ions for calcium and magnesium ions enhances the water softening efficiency.

Learning Task

List three areas that apply the concepts of equilibrium Research and present case studies in the pharmaceutical industry the utilises the concept of chemical equilibrium, highlighting strategies used to maximise yield and efficiency.

PEDAGOGICAL EXEMPLARS

1. **Exploratory approach:** Investigate from the internet and other sources about how Le Chatelier's principle is applied in industrial processes.
2. **Collaborative Learning:** Give group presentations on how chemists carry out reactions economically to get the highest yields in the shortest possible time using the Haber and Contact processes in the production of ammonia and sulphuric acid.

KEY ASSESSMENT

Level 2: Identify and explain at least three specific strategies employed by the ammonia manufacturing firms to enhance the yield of ammonia.

Level 3: The process for the production of ammonia is represented by the following equilibrium reaction equation $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$; $\Delta H = -92$ kJ/mol If more reactants are added to the system, predict the direction in which the equilibrium will shift based on the principles of Le Chatelier. Provide a detailed explanation for your prediction.

Hint

The recommended mode of assessment for week 8 is ***essay type****. Use the level 2 question as a sample question.*

Section 3 Review

Throughout this section, learners were taken through the concept of dynamic equilibrium in chemistry. Understanding that dynamic equilibrium is achieved when the rates of forward and backward reactions are equal was a key focus. Graphical methods were employed to illustrate and reinforce this notion of dynamic equilibrium. Students were tasked with deriving expressions for equilibrium constants based on concentrations and partial pressures, enabling them to perform calculations with precision. They also explored the relationship between K_c and K_p through data analysis. Le Chatelier's principle was introduced, allowing students to qualitatively deduce and predict the effects of various factors on the position of equilibrium and the entities involved in equilibrium reactions. Additionally, learners examined real-world industrial applications of Le Chatelier's principle, further solidifying their understanding of dynamic equilibrium in practical contexts

APPENDIX C: MID-SEMESTER EXAMINATION

Structure

This mid-semester exam will comprise 40 multiple-choice questions (MCQ). Learners should attempt all questions by choosing correct options. The exam should cover lessons learned from weeks 1-5. Time allowed for this paper is 50minutes

Resources

Stationery, Printer, Scannable answer sheets calculators, pencils, etc

Sample question

MCQ

A decrease in the equilibrium constant value has **one** of the following effects on equilibrium. It

A. favours the forward reaction.

B. favours the reverse reaction.

C. has no effect on the equilibrium.

D. increases the rate of attaining equilibrium.

How to Administer

1. Prepare examination questions and materials
2. Inform learners of examination details (date, time location, etc)
3. Distribute examination materials (questions, answer sheets, etc)
4. Monitor learners to prevent cheating and maintain a quiet and secure environment

Table of Specification

Week	Focal Area	Type of Question	DoK Level				Total
			1	2	3	4	
1	Enthalpy Change and Standard Enthalpy Changes	Multiple Choice	3	2	3	-	8
2	Experimental Determination of Enthalpy Change	Multiple Choice	3	4	3	-	10
3	Hess' Law of Constant Heat Summation	Multiple Choice	2	3	3	-	8

Week	Focal Area	Type of Question	DoK Level				Total
			1	2	3	4	
4	Rate of Reaction	Multiple Choice	2	4	3	-	9
5	Collision Theory	Multiple Choice	2	2	1	-	5
	Total		**12**	**15**	**13**		**40**

SECTION 4: ACIDS, BASES AND SALTS

Strand: Physical Chemistry

Sub-Strand: Equilibria

Learning Outcome: *Apply your knowledge in acids and bases to classify and describe substances as acids and bases as well as determine the concentration of analyte through titration.*

Content Standard: Demonstrate knowledge and understanding of the concepts and properties of acids and bases to classify substances as well as determine concentration of analyte through titration.

Hint

- *Remind learners of the end of semester examination in week 12. Refer to Appendix D at the end of this section for Table of specification to guide you to construct the items.*
- *Assign Individual Project work to learners in week 14 to be submitted in week 22. Refer to Appendix E at the end of this section for sample of materials and scoring rubrics.*

INTRODUCTION AND SECTION SUMMARY

This section covers weeks nine to fourteen of the chemistry curriculum, learners will study the fundamental principles of acids and bases, exploring various theories such as Arrhenius, Bronsted-Lowry, and Lewis. They will also study the physical and chemical properties of acids and bases, as well as how they are used in daily life. Additionally, the section encompasses the concept of salts, including its definition, classification, preparation methods, and practical uses. Moreover, learners will be introduced to acid-base titration techniques, including simple titration, back titration, and double indicator titration. Teachers can highlight the interdisciplinary nature of this section by connecting it to subjects like biology (pH balance in living organisms), environmental science (acid rain impact), and even math (calculating molarity in titration experiments).

The weeks covered by the section are:

Week 9: *Explain Arrhenius, Bronsted-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.*

Week 12: *Explain the meaning of salt and classify salts.*

Week 13/14: *Determine the quantity of analyte in a solution using acid-base titration.*

SUMMARY OF PEDAGOGICAL EXEMPLARS

In section four of the curriculum, a variety of interactive and experiential pedagogies are employed to facilitate a comprehensive understanding of acids and bases. Collaborative learning encourages learners to work together to grasp the concepts and limitations of acids and bases. Activity-based learning engages learners in exploring conjugate acid-base pairs through hands-on tasks. Exploratory learning enables learners to investigate the physical properties of acids and bases, including pH levels, texture, and reactions with indicators like litmus paper. Experiential learning involves conducting experiments to analyse the chemical properties of acids and bases. Talk for learning is utilised to discuss real-life applications of neutralisation reactions. Inquiry-based learning aids in qualitatively explaining the varying strengths of acids and bases. Through a mix of talk for learning, experiential learning, and exploratory learning, students also delve into salt classifications and undertake practical experiments on titration methods.

ASSESSMENT SUMMARY

To facilitate comprehensive assessment of learners on the concepts of acids and bases, a variety of formative assessment methods will be utilised, including short answers, oral and written presentations, lab reports, equipment manipulation, safety procedures, graph plotting, measurements, demonstrations, and homework assignments. Projects and group activities will be employed to assess learners' comprehension and application of concepts, tailored to cover levels 1, 2, 3, and 4 of the Depth of Knowledge (DOK) framework while emphasising critical thinking, communication, and creativity. Class exercises, lab reports, measurements, and group tasks will allow learners to demonstrate their grasp of concepts, while homework assignments and practical group activities will provide opportunities for practical application of knowledge in real-world contexts.

WEEK 9

Learning Indicator: *Explain Arrhenius, Bronsted-Lowry and Lewis concepts of acids and bases*

FOCAL AREA: CONCEPTS OF ACIDS AND BASES

Acids and bases are two important classes of chemicals found in household products, medical products and industrial uses. These two sets of chemicals were initially differentiated based on the criteria of taste and feel. Acids are found in fermented products and are responsible for their sour taste while bases have a soapy feel between the fingers.

Arrhenius concept of acids and bases

Arrhenius was the first scientist to give a deeper chemical understanding of the properties of acids and bases. He described an acid as a substance that produces hydrogen ions in water (aqueous solution). Using HCl as an example, in water, the molecules of HCl break up (dissociate) to give off its ions.

$\mathrm{HCl} \rightarrow H^+ + Cl^-$

Therefore hydrochloric acid contains hydrogen ions. All other aqueous solutions of acids also contain hydrogen ions which gives them their acidity.

A base, according to Arrhenius, is a compound that releases hydroxide ions in water (aqueous) solution. Using NaOH as an example, in aqueous solution NaOH molecules dissociate to give off OH^- ions.

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

Arrhenius' definition of acids and bases are limited because they apply only to aqueous solutions.

Bronsted-Lowry concept of acids and bases

A broader definition was given by two scientists Bronsted and Lowry that the essential feature of an acid – base reaction lies in the transfer of hydrogen ions from one compound to the other. It does not require acids and bases to be in aqueous solution.

A hydrogen ion (H^+) is formed when a hydrogen atom (${}^1_1H$) loses its electron. Therefore H^+ is just like a proton. They defined an acid as any substance that acts as a proton donor and a base as a proton acceptor. Such substances are referred to as Bronsted-Lowry acid and Bronsted-Lowry base respectively.

$H_2O + \mathrm{HCl} \rightarrow H_3O^+ + Cl^-$

Here HCl is acting as a Bronsted-Lowry acid, and H_2O is acting as a Bronsted-Lowry base.

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

Here NH_3 is acting as a Bronsted-Lowry base, and H_2O is acting as a Bronsted-Lowry acid.

This means that in a Bronsted – Lowry acid – base reaction, the acid donates protons to the base.

Not all hydrogen present in a compound can be donated and so not all compounds with hydrogens can be classified as Bronsted – Lowry acids. For example, methane (CH_4) shows no acidic properties. The measure of acidity of hydrogen atom bonded to an element depends on polarity and lone pairs of electrons on the element to which the hydrogen atom is bonded to. This means that elements of group 16 (VI A) and group 17 (VII A) when bonded to hydrogen forms a Bronsted – Lowry acid. Examples of Bronsted – Lowry acids that are common in most chemistry laboratories include HCl, HNO_3, H_2SO_4, CH_3COOH, H_3PO_4, H_2CO_3.

The measure of basicity of a compound depends on one or more lone pairs in order for it to receive a proton. Elements of group 15 (V A) and group 16 (VI A) usually form bases. The basic compounds that have the group 16 elements usually have that element deprotonated and hence negatively charged. Examples of common Bronsted – Lowry bases are NaOH, KOH, NH_3, C_5H_5N, Na_2CO_3.

Based on the number of protons that can be donated by a compound, Bronsted – Lowry acid can be grouped into two main categories. Compounds which can donate only proton are called monoprotic acids. Examples include HCl, HNO_3, CH_3COOH. Compounds which donate more than one proton are generally referred to as polyprotic acids. Those that donate two protons such as H_2SO_4, H_2CO_3, H_2S are exclusively called diprotic acids while those that donate three protons such as H_3PO_4 are referred to as triprotic acids.

Similarly, Bronsted – Lowry bases are generally grouped into monoprotic and polyprotic bases depending on the number of protons that can be accepted. Those that accept only one proton are referred to as the monoprotic bases, examples are NaOH, KOH, NH_3 and C_5H_5N. Exclusively, those that accept two protons such as CO_3^{2-} and S^{2-} are called diprotic bases.

Certain compounds behave differently based on the nature of environment they find themselves in a chemical reaction. A typical reaction is the one that takes place when an acid or a base is dissolved in water.

When gaseous hydrogen chloride is dissolved in water, the hydrogen chloride donates its hydrogen to the water.

$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$

In this reaction the HCl is acting as Bronsted – Lowry acid by donating its hydrogen to the H_2O in order to form a Cl^- while H_2O acts as a Bronsted – Lowry base by accepting the hydrogen ion to form a hydroxonium ion.

On the other hand, when gaseous ammonia is dissolved in water, hydrogen ions in water are donated to the ammonia compound.

$NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$

In the dissolution of ammonia in water, water acts as a Bronsted – Lowry acid by donating its hydrogen moiety to the ammonia in order to form OH^- which the NH_3 acts as a Bronsted – Lowry base by accepting the hydrogen ions to form an ammonium ion.

In this case, water is acting as an acid on reacting with NH_3 and also as a base when reacting with HCl. Compounds that can act as either an acid or a base are called amphiprotic compounds.

Conjugate Acid – base Pairs

The concept of conjugate acid – base pair is an extension of the definition of an acid and a base based on Bronsted and Lowry. A conjugate is the species formed when a proton is either removed or added to a compound. When a Bronsted – Lowry acid loses a proton, it forms a conjugate base of that acid. The original acid and the conjugate base form a conjugate acid – base pair which differ by one proton. Conversely, a conjugate acid is formed when a proton is added to a Bronsted – Lowry base.

Every Bronsted – Lowry acid has its conjugate base and there is a conjugate acid for every Bronsted – Lowry base.

$H_2O(l) + HCl(g) \rightarrow H_3O^+(aq) + Cl^-(aq)$

Bronsted - Lowry base + Bronsted – Lowry acid → Conjugate acid + Conjugate base

$NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$

Bronsted Lowry base + Bronsted – Lowry acid →Conjugate acid + Conjugate base

Lewis's concept of acids and bases

Arrhenius and Bronsted – Lowry definitions of acids assume that an acid must contain or produce hydrogen ions. According to Bronsted – Lowry an acid must lose proton(s) and a base must be able to accept proton(s). One characteristic property about bases is that they should have unshared pair of electrons to accept the proton. Lewis gave a more general definition of acids and bases. According to Lewis, an acid is an atom, ion or molecule that can accept a pair of electrons to form a covalent bond while a base is an atom, ion or molecule that can donate a pair of electrons to form a covalent bond. A proton acts as a Lewis acid when it reacts with ammonia in which it accepts a lone pair of electrons from the ammonia to form a dative covalent bond in ammonium ion.

$H^+(aq) + :NH_3(aq) \rightarrow [H\text{–}NH_3]^+(aq)$ or $[NH_4]^+(aq)$.

A Lewis acid does not necessarily need a hydrogen.

A Lewis acid-base reaction, therefore, is one that involves the donation of a pair of electrons from one species to another. These reactions do not necessarily produce salt and water.

In the reaction between BF_3 and NH_3, which is another typical Lewis acid-base reaction, NH_3 donates an electron pair to BF_3 which results in the formation of a B – N covalent bond.

$NH_3 + BF_3 \rightarrow F_3B\text{-}NH_3$

Learning Tasks

1. Define an acid according to Arrhenius.
2. State the limitations of using Arrhenius concept in defining an acid.
3. Use chemical equation to explain the concept of Bronsted – Lowry acid and bases.
4. Write a typical Bronsted – Lowry acid – base reaction and determine the conjugate acid – base pairs present.

PEDAGOGICAL EXEMPLARS

1. Collaborative Learning

With the aid of charts or videos, brainstorm to find answers to the following tasks:

a. Arrhenius theory of acids and bases and give their strengths and limitations

i. Introduce the Arrhenius theory of acids and bases, explaining the fundamental principles of how acids release H^+ ions and bases release OH^- ions in aqueous solutions.

ii. Use visual aids such as diagrams, charts or interactive simulations to demonstrate the behaviour of acids and bases according to Arrhenius theory.

iii. Divide learners into small mixed ability groups of 3-5 members each to promote collaboration and peer learning.

iv. Engage learners in their groups to discuss and analyse the strengths of Arrhenius theory of acids and bases. Encourage students to brainstorm and share their ideas within their groups.

v. Assign each group to research and identify the strengths and limitations of the Arrhenius theory. Provide resources such as textbooks, articles, or online materials to facilitate their research.

   vi. Instruct each group to prepare a short presentation highlighting the strengths and limitations they have identified. Encourage them to use visual aids like charts, chemical equations or videos to enhance their presentation.

   Allocate time for each group to present their findings to the rest of the class. Encourage active participation and questions from the audience to promote knowledge sharing.

b. Bronsted-Lowry theory of acids and bases and give their strengths and limitations.

   i. Explain Bronsted-Lowry theory of acids and bases to learners in that acids are proton donors and bases are proton acceptors.

   ii. In mixed ability groups guide learners to identify group of compounds that have dissociable hydrogen atoms and give the types of Bronsted-Lowry acids (simple acids, oxoacids and hydrated cations) and bases.

   iii. Provide examples of acid-base reactions and demonstrate how proton transfer occurs in these reactions. Emphasise the importance of proton transfer in determining whether a substance acts as an acid or a base.

   iv. Discuss the strengths of the Bronsted-Lowry theory, such as its ability to explain a wider range of acid-base reactions compared to the Arrhenius theory. Highlight how the Bronsted-Lowry theory is more versatile and applicable to non-aqueous solutions and complex chemical systems.

   v. Address the limitations of the Bronsted-Lowry theory, such as its inability to explain reactions involving species that do not involve proton transfer. Discuss situations where the theory may not fully account for the behaviour of certain acids and bases.

c. Lewis theory of Acids and Bases and give their strengths and limitations.

   i. Explain that Lewis theory of acids and bases focus on the electron pair behaviour in chemical reactions. Mention that according to the Lewis theory, an acid is a substance that can accept a pair of electrons, while a base is a substance that can donate a pair of electrons.

   ii. Provide examples of Lewis acids, such as metal cations (e.g. Al^{3+}), hydrogen ion (H^+), and electron-deficient molecules like BF_3. Explain why these species can accept an electron pair.

   iii. Give examples of Lewis bases, such as ammonia (NH_3), hydroxide ion (OH^-), and water (H_2O). Discuss why these species can donate an electron pair.

iv. Illustrate Lewis acid-base reactions using simple chemical equations. For example, show how ammonia (NH_3) acts as a Lewis base by donating a pair of electrons to a proton (H^+) to form ammonium ion (NH_4^+). Explain that Lewis acid-base reactions involve the formation of coordinate covalent bonds, where both electrons in the bond come from the Lewis base.

2. **Activity-Based Learning**
   1. Provide a brief explanation of what constitutes a conjugate acid-base pair and why it is important in understanding acid-base chemistry.
   2. Design an interactive activity that actively engages the learners to identify conjugate acid-base pairs in a given acid-base reaction. This could involve a hands-on experiment, a group discussion, or a simulation.
   3. Present a series of acid-base reactions to the learners, ensuring that each reaction clearly shows the transfer of a proton from the acid to the base. The reactions can progress from simple to more complex reactions in order to challenge learners.
   4. Guide the learners through the process of identifying the conjugate acid-base pairs in each reaction. Encourage them to focus on the species that gain or lose a proton during the reaction and explain how they form conjugate pairs.
   5. Promote collaborative learning by encouraging the learners to work together in pairs or groups to discuss and analyse the reactions. Encourage peer-to-peer teaching and discussion to deepen understanding.
   6. Provide the learners with additional acid-base reactions to analyse independently. Ask them to identify the conjugate acid-base pairs in each reaction and explain their reasoning.

Watch video to reinforce the concepts.

KEY ASSESSMENTS

Level 1: What is an Arrhenius Acid?

Level 2

1. Classify each of the following species in aqueous solution as a Bronsted-Lowry acid or base: (a) HBr (b) NO_2 (c) HCO_2
2. Identify the Bronsted-Lowry acid and its conjugate base in the following reactions
   - **a.** $HSO_4^- + OH^- \rightarrow H_2O + SO_4^{2-}$
   - **b.** $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$
   - **c.** $NH_3 + H_2S \rightarrow HS^- + NH_4^+$

3. Identify Bronsted-Lowry acids and bases in the reaction below:

$HCl + H_2O \rightleftharpoons H_3O + Cl-$

Hint

The recommended mode of assessment for week 9 is ***peer/self assessment****. Use the level 2 question 3 as a sample question. Refer to the Teacher Assessment Manual and Toolkit page 143 for additional information on how to carry out this assessment mode.*

WEEK 10

Learning Indicator: *Describe the physical and chemical properties of acids and bases*

FOCAL AREA: PHYSICAL AND CHEMICAL PROPERTIES OF ACIDS AND BASES

Physical properties of acids and bases

Physical Properties of acids

i. They have a sharp sour taste.

ii. They change the colour of blue litmus paper to red.

iii. They change the colour of Methyl Orange/Yellow to Pink.

iv. Acidic substances convert Phenolphthalein from deep pink to colourless.

v. Have a wet feel.

vi. Conduct electricity.

vii. Have pH values less than seven.

Physical Properties of bases

i. They have a bitter taste.

ii. They change the colour of red litmus paper to blue.

iii. They change the colour of Methyl Orange/Yellow to yellow.

iv. Basic substances convert Phenolphthalein from colourless to pink.

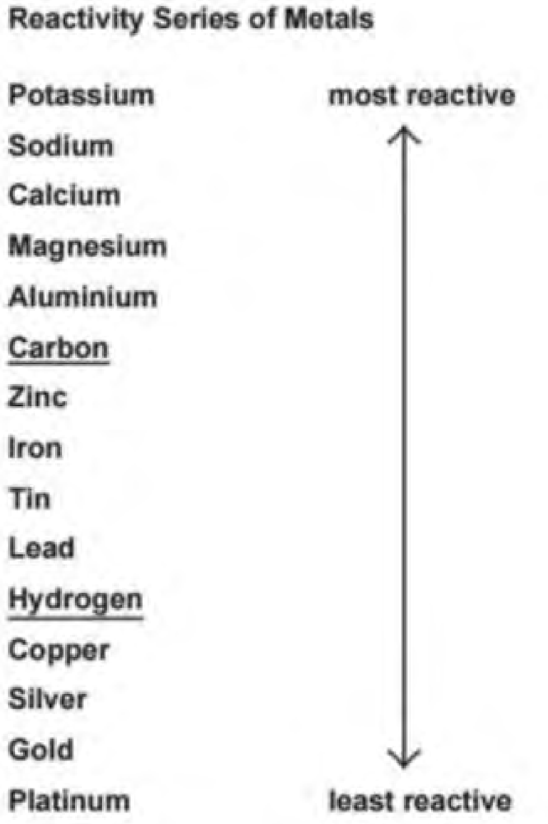
v. Feel slippery between the fingers.

vi. Aqueous solutions of bases conduct electricity.

vii. Have pH values greater than seven.

Chemical properties of acids and bases

Chemical properties of acids



i. Acids react with certain metals, (that are above hydrogen in the electrochemical series, common examples are shown below) to produce hydrogen gas.

A typical reaction is that between hydrochloric acid and magnesium:

$2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$

ii. Acids react with carbonates and bicarbonates, such as Na_2CO_3, $CaCO_3$ and $NaHCO_3$, to produce carbon dioxide gas.

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

iii. Acids react with bases to produce salt and water. This reaction is called neutralisation.

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

Chemical properties of bases

i. Soluble bases (alkalis) react with aqueous salt solution of heavy metals to produce a gelatinous precipitate (ppt) of the metallic hydroxide

$3NaOH(aq) + AlCl_3(aq) \rightarrow Al(OH)_3(s) + 3NaCl(aq)$

ii. Bases are warmed with ammonium salt to give off ammonia gas

$NaOH(aq) + NH_4Cl(s) \rightarrow NaCl(aq) + H_2O(l) + NH_3(g)$

iii. Bases neutralise acids to form salt and water

$$KOH(aq) + HCl(aq) \rightarrow KCl(s) + H_2O(l)$$

Classification of household substances as acids and bases

Classification of household items are based on their physical and chemical properties. These are primarily qualitative tests that can be done outside the laboratory. The common tests include

a. Taste (Do not taste any chemical in the laboratory)

Acids are substances that taste sour. Unripe oranges taste very sour because their acidic content are higher than the ripe ones.

Bases on the other hand tastes bitter.

b. Litmus paper test

Acidic substances or their solutions turn blue litmus paper red.

Bases substances or their solutions turns red litmus paper blue.

c. The pH scale

Solutions of acids have pH values less than 7.

Solutions of bases have pH values greater than 7.

d. Universal indicator test

The universal indicator is a mixture of different indicators that change colour based on the pH level of a particular substance. It can be used to test whether a substance is acidic, neutral or basic by matching the colour change with a colour code that corresponds to the pH value.

When the universal indicator is mixed with an acidic solution, the resulting solution becomes yellow, beige, orange, pink or red. When mixed with a basic solution, the resulting solution becomes dark green, turquoise, pale blue, blue, dark blue, violet or purple. Neutral solutions, turn green.

Application of neutralisation reactions in real life

a. The chemistry behind the mechanism of antacid medications involves neutralisation reactions to relieve symptoms of acidity and heartburn. The stomach naturally produces hydrochloric acid (HCl) to aid in the digestion of food. Excess of stomach acid can lead to conditions such as acid reflux, heartburn, and indigestion. Antacid medications contain basic compounds, such as magnesium hydroxide ($Mg(OH)_2$), aluminium hydroxide ($Al(OH)_3$), calcium carbonate ($CaCO_3$), or sodium bicarbonate ($NaHCO_3$). These basic compounds work by neutralising the excess stomach acid. When an antacid is ingested, the basic compounds react with the hydrochloric acid in the stomach through a neutralisation reaction. The reaction can be represented as follows, taking magnesium hydroxide ($Mg(OH)_2$) as an example:

$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$

In this reaction, magnesium hydroxide (antacid) reacts with hydrochloric acid to form magnesium chloride (a salt) and water. The formation of water (H_2O) helps to neutralise the acidic environment in the stomach, providing relief from symptoms. The neutralisation of stomach acid by antacids results in the pH of the stomach becoming less acidic. As a result, the discomfort caused by acid reflux and heartburn is alleviated, providing temporary relief to individuals experiencing these symptoms. Depending on the severity of symptoms, individuals may take antacid medications as needed or as directed by a healthcare provider. It is important to follow the recommended dosage and not exceed the prescribed duration of use to avoid potential side effects.

b. The treatment of wastewater also involves the application of neutralisation reactions to adjust the pH of acidic or basic wastewaters to a neutral level before it is discharged into the environment. Wastewaters generated from various industrial processes, municipal sources, or agricultural activities can contain acidic or basic components. These components can be harmful to the environment if discharged untreated. The pH of water is a measure of its acidity or alkalinity. Wastewaters with low pH values or high pH values need to be neutralised before being released into water bodies.

In wastewater treatment plants, neutralisation reactions are used to adjust the pH of wastewaters. For acidic wastewaters, bases such as calcium hydroxide (slake lime, $Ca(OH)_2$) or calcium oxide (quicklime, CaO) are added to neutralise the acidity. For basic wastewaters, acids such as sulphuric acid (H_2SO_4) or hydrochloric acid (HCl) may be used to neutralise the alkalinity. The neutralisation reactions in wastewater treatment can be represented as follows, taking the example of neutralising acidic wastewater with calcium hydroxide:

$2HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O$

In this reaction, hydrochloric acid (HCl) reacts with calcium hydroxide to form calcium chloride (a salt) and water. The addition of the base calcium hydroxide helps raise the pH of the acidic wastewater to a more neutral level. By adjusting the pH of wastewaters to a neutral level through neutralisation reactions, the harmful effects of acidic or basic effluents on aquatic ecosystems are minimised. This process helps prevent damage to aquatic life, water quality, and ecosystems downstream.

c. Application of neutralisation reaction is also seen in the field of agriculture. Soil acidicty affects plant growth and nutrient availability. Agricultural lime (calcium carbonate) is added to acidic soils to neutralise the acidic components present in the soil. The reaction can be represented as follows:

$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2O(l) + CO_2(g)$

In this reaction, calcium carbonate reacts with sulphuric acid present in the acidic soil to form calcium sulphate, water, and carbon dioxide. As a result, the acidity of the soil is neutralised, and the pH of the soil is increased towards a more neutral level. This process helps to optimise plant growth and crop yield.

d. In the area of cleaning, neutralisation reactions are utilised to remove acidic or basic residues from surfaces. When acidic cleaners neutralise alkaline residues or when alkaline cleaners neutralise acidic residues, they help in effectively cleaning the surfaces. The chemical equations involved in these neutralisation reactions are as follows:

1. Neutralisation reaction of an acidic cleaner (e.g., citric acid) with alkaline residues:

$C_6H_8O_7(aq) + 3NaOH(aq) \rightarrow 3H_2O(l) + Na_3C_6H_5O_7(aq)$

Citric acid + sodium hydroxide → water + sodium citrate

2. Neutralisation reaction of an alkaline cleaner (e.g., sodium hydroxide) with acidic residues:

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

Sodium hydroxide + hydrochloric acid → water + sodium chloride

In these examples, acidic residues are neutralised by alkaline cleaners through the reaction between citric acid and sodium hydroxide, resulting in the formation of water and sodium citrate. On the other hand, alkaline residues are neutralised by acidic cleaners through the reaction between sodium hydroxide and hydrochloric acid, leading to the formation of water and sodium chloride. These neutralisation reactions help in removing unwanted residues from surfaces and ensuring effective cleaning.

e. Another industry that applies the concept of neutralisation reaction is the food and beverages industry. A typical example is the production of fermented foods like yogurt where pH adjustment plays a critical role in ensuring food safety and quality. Neutralisation reactions are employed to adjust the pH levels during fermentation processes.

During the fermentation of yogurt, lactic acid bacteria convert lactose (milk sugar) in milk into lactic acid through a microbial process. This production of lactic acid leads to a decrease in pH, making the environment more acidic.

To stabilise the pH and control the fermentation process, a base like calcium hydroxide (lime water) can be used to neutralise the excess lactic acid. The neutralisation reaction between lactic acid and calcium hydroxide is as follows:

$$2C_3H_6O_3 + Ca(OH)_2 \rightarrow Ca\left(C_3H_6O_3\right)_2 + 2H_2O$$

lactic acid + calcium hydroxide → calcium lactate + water

In this reaction, the lactic acid is neutralised by calcium hydroxide, forming calcium lactate and water. Calcium lactate is a neutral salt that stabilises the pH of the yogurt mixture and helps control the acidity level.

Learning Tasks

1. List two physical properties each for acids and bases.
2. Investigate and compare the strengths of various acids $\left(HCl, CH_3COOH, H_2SO_4\ etc\right)$ by measuring pH and observing reactions with metals.
3. Perform neutralisation reaction between acids and bases, observing temperature changes and pH shifts using suitable indicators.
4. Perform the reaction of acids with carbonates or hydrogen carbonates and test for the gas that will be evolved.
5. Investigate the effect of acids on basic oxides and record any observations from the experiment.
6. Design an experiment to examine how bases interact with acidic oxides.
7. Explore the reaction of bases with ammonium salts to liberate ammonia gas.
8. Analyse the pH levels of various household substances, create a detailed classification chart based on acidity levels, and predict potential reactions when mixed.
9. Investigate the reactivity of acids and bases by examining their reactions with metals, carbonates, and indicators like litmus paper.

PEDAGOGICAL EXEMPLARS

1. Exploratory Learning

a. Introduce the concept of acids and bases using simple definitions and examples. Provide additional support for struggling learners by offering visual aids and simplified explanations.

b. Use a virtual lab or simulation to explore pH levels, acid-base indicators, and physical properties.

c. Conduct simple activities, such as testing the pH of household substances using pH strips or litmus paper.

d. Explore the feel and behaviour of acids and bases using simple activities, such as: Testing the feel of acids and bases (e.g., vinegar, baking soda). Observing the behaviour of indicators (e.g., litmus paper, pH strips).

- **e.** Have learners collect and analyse data on the physical properties of acids and bases. Provide scaffolding techniques, such as graphic organisers and guided questions.
- **f.** Have learners work in small groups to share and discuss their findings. Facilitate peer support by encouraging learners to ask each other questions and share their expertise.
- **g.** Have learners reflect on their learning and identify areas for improvement.
- **h.** Have learners create posters or presentation, summarizing their findings. Offer choices in the mode of presentations to accommodate different learning styles and abilities.

2. Experiential Learning

- **a.** Undertake experiments on acids such as:
  - i. Reaction with metals like Zn, Fe and Mg
  - ii. Reaction with bases (Neutralisation).
  - iii. Reaction with trioxocarbonate (IV), (CO_3^{2-}) or hydrogen trioxocarbonate (IV), (HCO_3^-)
  - iv. Reaction with basic oxides.
- **b.** Undertake experiments on bases such as:
  - i. Reaction with acidic oxides.
  - ii. Reaction with ammonium salts to liberate ammonia gas
  - iii. Reaction with acids.
  - iv. Classify household chemical, fruits, vegetables etc as acids or bases

3. Talk-for-Learning

Discuss application of neutralisation in treating insect bites, stomach indigestion, treating soil and lake acidity.

KEY ASSESSMENTS

Level 2

1. A pH meter was used to check the pH level of four solutions labelled as A, B, C and D as described in the table below

Solution	A	B	C	D
pH	7	1	13	3

- **a.** Which of the solution(s) will taste sour?

b. Identify the solution that will have the same pH as a dilute Na_2CO_3.

c. Which solution is the most acidic?

d. Which solution will have no effect on litmus paper?

2. A pH meter was used to check the pH level of five solutions labelled as A, B, C, D, and E. The result obtained is tabulated below. Use the information to answer the questions that follows:

Solution	A	B	C	D	E
pH	12	3	7	5	1

a. Which of the solution(s) will taste sour?

b. Identify the solution that will have similar pH as a dilute Na_2CO_3

c. Which solution is most acidic?

Hint

*The recommended mode of assessment for week 10 is **test of practical**. Use the level 2 question 2 as a sample question.*

WEEK 11

Learning Indicator: *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: STRENGTH OF ACIDS AND BASES

Differences between strong acids and weak acids

A strong acid can be distinguished from a weak acid by taking into consideration their pH value, electrical conductivity and reactivity.

1. **pH:** pH is the logarithmic value of the inverse of the hydrogen or hydronium ion concentration.

   Consider the reaction $HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$;

   For a strong acid, the equilibrium lies far to the right meaning that almost all the original acid (HA) is dissociated at equilibrium producing a higher concentration of H_3O^+/H^+ ions in solution. The higher concentration of the H_3O^+/H^+ ions in solution will result in a lower pH value.

   On the other hand, for a weak acid, the equilibrium lies far to the left, meaning the original acid (HA) dissociates partially or incompletely in aqueous solution producing a lower concentration of H_3O^+/H^+ ions in solution with most of the original acid remaining as HA (The extent of dissociation is minimal). The lower the concentration of the H_3O^+/H^+ ions in solution, the higher the pH value.

2. **Electrical conductivity:** The tendency of a substance to conduct electric current is due to the movement of ions in solution. Strong acids are known for their complete dissociation into their corresponding ions producing high concentration of ions in solution making them highly conductive. Weak acids on the other hand due to their incomplete or partial dissociation produces lower concentration of ions makes them less conductive.

3. **Reactivity:** Strong and weak acids of the same concentrations react differently with reactive metals. This is because the concentration of H_3O^+/H^+ ion is higher in strong acids compared to weak acids. The higher H_3O^+/H^+ ion concentration of the acid the more hydrogen gas is produced in a shorter time. On the other hand, the lower the H_3O^+/H^+ ion concentration of the acid the less hydrogen gas is produced in a shorter time.

4. **Enthalpy change of neutralisation value ($\Delta H_{neut.}$):**

   The enthalpy change (ΔH) of neutralisation of strong acids and strong bases are almost the same, this is because strong acids and strong bases dissociates completely in solution and so no chemical bonds will be broken when a base is added. Energy is released during the formation water molecules causing the reaction to be exothermic, e.g.

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)\ \Delta H = -57.1\ kJ\,mol^{-1}$

The enthalpy change of neutralisation involving weak acids are smaller in magnitude compared to that of strong acids and strong bases. The reason is weak acids/bases partially dissociates, which means not all hydrogen ions are free to react with hydroxide ions. Energy is required to break the bonds in weak acid molecules to produce hydrogen ions.

For example, consider the weak acid:

$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}\ \Delta H = +1.0\ kJ\,mol^{-1}$

At equilibrium, only a small amount of CH_3COOH molecules are dissociated. Most hydrogen ions are still bonded to CH_3COOH molecules. The forward reaction is endothermic with a positive ΔH as seen above.

When a strong base such as NaOH is added, the OH^- ions of the base causes more CH_3COOH molecules to be deprotonated according to the equation:

$CH_3COOH_{(aq)} + OH^-_{(aq)} \rightleftharpoons H_2O_{(l)} + CH_3COO^-_{(aq)}\ \Delta H = -56.1\ kJ\,mol^{-1}$

The overall enthalpy of this neutralisation reaction is less negative compared to the reaction between HCl and NaOH above because 1.0 kJ is absorbed for every mole of CH_3COOH deprotonated.

Strong acid	Weak acid
HCl	H_2CO_3
HNO_3	H_2S
HBr	HCN
H_2SO_4	H_3PO_4
$HClO_3$.	HF

Differences between strong bases and weak bases

Just like acids, the relative strength of a base is attributed to the extent of its dissociation in aqueous solutions. For solutions with the concentration, stronger bases dissociate to a greater extent and so produces higher hydroxide ion (OH^-) concentrations than do weaker bases.

The position of the equilibrium for a strong base is so far to the right that it can be represented as an irreversible reaction. $NaOH(s) + H_2O(l) \rightarrow Na^+(aq) + OH^-(aq)$

The equilibrium position for weak bases on the other hand lies far to the left and the reaction is a reversible reaction. $NH_3(l) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.

The relative strength of bases can also be compared using their pH and electrical conductivity similar to those of the acids.

1. **pH:** The pH of a strong base solution is above seven, indicating its alkaline nature. Strong bases have a higher concentration of hydroxide ions which results in a higher pH value. On the other hand, weaker bases have lower concentrations of hydroxide ions and thus lower pH values.
2. **Electrical conductance:** Hydroxide ions are responsible for electrical conductivity in aqueous solutions of bases, a higher concentration of hydroxide ions indicates a stronger electrical conductivity. Weak bases on the other hand have lower electrical conductivity due to the fewer hydroxide ions concentration in solution.
3. **Reactivity:** Strong bases exhibit high reactivity and can react vigorously with acids unlike weak bases, leading to the formation of salt and water. However, this reactivity also makes them corrosive and potentially dangerous if not handled with caution.

Strong base	Weak base
NaOH	NH_3
KOH	$Fe(OH)_3$
RbOH	$Pb(OH)_2$
CsOH	$Zn(OH)_2$
$Mg(OH)_2$	$Al(OH)_3$
$Ca(OH)_2$	$Cu(OH)_2$

Learning Tasks

1. What are the differences between weak and strong acids and weak and strong bases in terms of dissociation?
2. Using a circus activity, investigate acid or base strengths based on the following methods:
   a. conductivity measurement or using the acids or bases as electrolyte with electric bulb in a circuit.
   b. pH measurement using a pH meter or universal indicator method.

   c. use of heat of neutralisation value.

3. Discuss basicity of acids and relate it to the strength of acids.

PEDAGOGICAL EXEMPLARS

Inquiry-based Learning

1. Introduce acids and bases using visual aids, interactive simulations, and real-life examples.
2. Explain weak vs. strong acids/bases with analogies and hands-on activities.
3. Divide learners into mixed-ability groups for a circus of stations exploring acid/base strengths.
4. At the Conductivity Measurement Station, provide differentiated tasks and guiding questions.
5. At the pH Measurement Station, offer scaffolded support and extension activities.
6. At the Heat of Neutralisation Station, use visual instructions and encourage data analysis.
7. Provide additional support through peer tutoring and step-by-step guides.
8. Allow groups to present their findings in various formats with clear rubrics.
9. Define basicity using interactive models and comparative charts.
10. Encourage discussions with think-pair-share and real-life connections.

KEY ASSESSMENTS

Level 1: What is the main difference between strong acids and weak acids in terms of dissociation?

Level 3

1. Compare the acid strength of tetraoxosulphate (VI) acid (H_2SO_4) and trioxonitrate(V) acid (HNO_3). Provide reasons for your answer.
2. Demonstrate that strong acids conduct electricity better than weak acids when in aqueous solution.

Hint

*The recommended mode of assessment for week 11 is **observation**. Use the level 3 question 2 as a sample question.*

WEEK 12

Learning Indicator: *Explain the meaning of salt and classify salts*

FOCAL AREA: SALTS

A salt is an ionic compound that is formed when one or more hydrogen ions of an acid are replaced by metallic or ammonium ions. These ionic compounds play a very important role in various chemical processes ranging from everyday use to industrial use.

A salt is therefore formed when the hydrogen ions in an acid are replaced wholly or partially, directly or indirectly by a metal or ammonium radical.

The general equation for salt formation is: Acid + Base (metal or ammonium) → salt + water.

Classification of salts

There are many types of salts, each with unique properties and uses. Some common types of salts include:

1. **Normal Salts:** Normal salts are salts obtained by complete neutralisation of an acid with a base.

   $NaOH + HCl \rightarrow NaCl + H_2O$

   Other examples of simple salts include $NaNO_3$, Na_2SO_4, Na_2CO_3, etc.

2. **Acidic Salts:** Acid salts are derived from the partial replacement of hydrogen ions of an acid by a metal:

   $NaOH + H_2SO_4 \rightarrow NaHSO_4 + H_2O$.

   An acidic salt can also be formed as a result of the reaction between a weak base and a strong acid. An example of an acidic salt is ammonium chloride (NH_4Cl), which is formed from the neutralisation between hydrochloric acid (strong acid) and ammonia (weak base): $HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$. Other examples include, $NaHSO_4$, KH_2PO_4, $CuSO_4$, etc.

3. **Basic Salts:** Basic salts are formed by the partial replacement of hydroxide ions in a diacidic or triacidic base with an acid radical.

   Example: $Pb(OH)_2 + HCl \rightarrow Pb(OH)Cl + H_2O$

   Basic salts result from the reaction between a weak acid and a strong base. An example of a basic salt is Sodium ethanoate (CH_3COONa), which is formed by the neutralisation of ethanoic acid (a weak acid) with Sodium hydroxide (a strong base):

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$, Other examples of basic salts include Potassium cyanide (KCN), Sodium sulphide (Na_2S), etc.

4. **Double salts:** Double salts are salts made from equimolar aqueous mixture of two salts. They contain two different cations in different oxidation states. A double salt in which the cation are in +1 and +3 oxidation states are called alums.

   Examples $FeNH_4(SO4)_2$ [Ammonium iron (III) bis tetraoxosulphate (VI)], $Fe(NH_4)_2(SO_4)_2.6H_2O$ [diammonium iron (II) bis tetraoxosulphate (VI) hexahydrate, $KAl(SO_4)_2.12H_2O$ [aluminium potassium bis tetraoxosulphate (VI) dodecahydrate], etc. The later commonly called potash alum is used as mordant in dyeing and in tanning leather goods. It coagulates and precipitates dust particles in water purification processes.

5. **Complex salt:** A complex salt is a compound which consists of a central metal atom having coordination bonds with ligands around it. It is not completely dissociated into its ions just like the case of double salts when added to water. Instead, they remain as complex structures. $[Ni(CO)_4]$, $[Cr(NH_3)_6]Cl_3$, etc.

6. **Hydrated salts:** Hydrated salts are crystalline salt molecules that are loosely attached to a certain number of water molecules. This occurs when salt is exposed to water or a humid environment. In such conditions, water permeates the salt, leading to the process of hydration. When water molecules are absorbed into the salt, the structure and properties of the salt undergo partial changes and as a result, its usability is also altered. Examples of hydrated salts include: calcium tetraoxosulphate (VI) dihydrate, magnesium tetraoxosulphate (VI) heptahydrate, etc.

7. **Deliquescent salts:** Deliquescent salts are those that absorb water from the atmosphere and form a solution. Examples of deliquescent salts include: sodium trioxonitrate (IV), Calcium chloride, sodium hydroxide, iron (II) chloride, magnesium chloride, etc.

8. **Hygroscopic salts:** Hygroscopic salts are salts that absorb water from the atmosphere but not necessarily to the point of dissolution. Examples of hygroscopic salts include: anhydrous copper (II) tetraoxosulphate (VI), anhydrous cobalt (II) chloride, etc.

9. **Efflorescent salts:** Efflorescent salts/compounds are those that lose their water of crystallisation to the atmosphere on their own. Examples of efflorescence salts include: sodium trioxocarbonate (IV) decahydrate, sodium tetraoxosulphate (VI) decahydrate, etc.

Preparing soluble salts

There are two methods of preparing a solution salt: Reacting an acid and a solid metal (insoluble base or insoluble carbonate) and reacting a dilute acid and alkali (soluble base).

Reacting an acid and a solid metal (insoluble base or insoluble carbonate)

1. **Materials Needed**

   Dilute sulphuric acid (H_2SO_4)

   Insoluble base or insoluble carbonate (e.g., copper (II) oxide)

   Beaker

   Bunsen burner

   Stirring rod

   Evaporating basin

   Glass rod

   Filter paper

2. **Procedure**
   - **a.** Add dilute sulphuric acid into a beaker.
   - **b.** Heat the beaker using a Bunsen burner flame. Take care not to boil the solution if a water bath (60 degrees C) is available use this instead.
   - **c.** Gradually add the insoluble metal, base, or carbonate (e.g., copper (II) oxide) to the warm dilute acid. Stir continuously until the copper (II) oxide stops disappearing and a suspension of the base forms in the acid.
   - **d.** Filter the mixture into an evaporating basin to remove any excess base.
   - **e.** Heat the solution in the evaporating basin to evaporate water, making it saturated. Take care not to over boil this solution as it will begin to spit when the water content is very low.
   - **f.** Check if the solution is saturated by dipping a cold glass rod into it. If crystals form on the end of the rod, it indicates saturation.
   - **g.** Leave the filtrate in a warm place to allow crystallisation.
   - **h.** Once crystals have formed, decant any excess solution carefully.
   - **i.** Allow the crystals to dry completely (you can also blot them with filter paper)

Reacting a dilute acid and alkali (soluble base)

1. **Materials Needed**

   Alkali (e.g., sodium hydroxide solution)

   Dilute acid (e.g., hydrochloric acid)

   Methyl orange indicator

   Conical flask

   Burette

   Evaporating basin

2. **Procedure**

   a. Use a pipette to measure the alkali (e.g., sodium hydroxide solution) into a conical flask.

   b. Add a few drops of methyl orange indicator to the conical flask. The indicator will change colour during the titration.

   c. Fill the burette with the dilute acid (e.g., hydrochloric acid).

   d. Record the initial volume of acid in the burette.

   e. Slowly add the acid from the burette to the conical flask while swirling the mixture until the indicator changes from yellow (alkaline) to orange/red (neutral) as the acid reacts with the alkali.

   f. Record the final volume of acid in the burette once the colour change occurs and calculate the volume of acid added.

   j. Add the same volume of acid to the same volume of alkali (without the indicator) and heat (Take care not to over boil this solution as it will begin to spit when the water content is very low) the resulting solution in an evaporating basin to partially evaporate the water, leaving a saturated solution.

   g. Crystals will start forming on the sides of the basin, leave the solution to crystallise further.

   h. Decant any excess solution carefully and allow the crystals to dry completely.

Uses of salts

1. **Medicine and Health**

   Saline, a mixture of salt (usually sodium chloride) and water, is commonly used in medical settings. It serves several purposes:

   a. Rehydration: Saline is administered intravenously to rehydrate patients who are dehydrated due to illness, surgery, or other medical conditions.

   b. Wound Cleaning: Saline is used to clean wounds, helping to remove debris and prevent infection.

2. **Industry**
   a. Desalination: Salts play a crucial role in desalination processes. Seawater contains dissolved salts, and desalination removes these impurities to produce fresh water. Reverse osmosis and distillation are common methods.
   b. Chemical Production: sodium chloride (table salt) is a raw material for chlorine production. Chlorine is essential for water treatment, plastics, and other chemical processes. Soda Ash (sodium carbonate) is used in glassmaking, detergents, and paper production.

3. **Agriculture**
   a. Fertilizers: Salts provide essential nutrients to plants. Common examples include: potassium nitrate (KNO_3) is a source of potassium and nitrogen. Ammonium sulphate $(NH_4)_2SO_4$ provides nitrogen and sulphur.
   b. Animal Nutrition: Salt licks or mineral blocks are given to livestock to ensure they receive necessary minerals like sodium and iodine.
   c. Pest and Disease Control: Some salts are used to control pests and fungal diseases in crops.

4. **De-Icing:** During winter, rock salt is spread on roads and sidewalks to melt ice and improve traction. It lowers the freezing point of water, preventing ice formation, etc.

Learning Tasks

1. Identify at least three uses of salt.
2. Differentiate between neutral, acidic, and basic salts.
3. Discuss at least three specialised uses of salts in agriculture, health, and other fields.

PEDAGOGICAL EXEMPLARS

1. **Talk-for-learning**
   a. Begin the lesson by introducing the topic: What is salt?
   b. Explain that brainstorming encourages creative thinking and diverse perspectives.
   c. Ask learners to share their initial thoughts on salt. Encourage them to think freely without judgment.
   d. Emphasise that unconventional ideas are valuable.

e. Divide learners into small mixed-ability groups, ensuring a range of learning abilities and needs in each group.

f. Assign each group a specific task related to the meaning and uses of salt, tailored to their collective ability level.

g. Have learners record their ideas in books and present to the whole class. Offer options for presentation formats, such as posters, infographics, etc.

h. Discuss and expand upon the ideas presented by each group.

i. Whilst maintaining the ability groupings, assign each group again with specific type of salt (e.g., normal, acidic, complex, etc.) and have them explore and discuss the characteristics of their assigned salt type with the help of textbooks and the internet and have them do a presentation of their choice.

j. Provide additional support or scaffolding for struggling learners, such as graphic organisers or sentence starters.

k. Clarify any misconceptions or gaps in their understanding and offer positive feedback after each presentation.

2. Experiential Learning

Put learners in small groups and explain the purpose of the lesson (to prepare salts and explore their uses)

For preparing Soluble Salts

1. Choose an acid (e.g., hydrochloric acid) and an insoluble reactant (e.g., metal oxide or carbonate).
2. Add dilute acid to a beaker.
3. Gradually add powdered reactant, stirring until effervescence occurs. Continue until some unreacted powder remains (excess).
4. Filter the mixture to collect the soluble salt solution.

Preparing Insoluble Salts

1. Mix two aqueous solutions of different soluble salts.
2. The insoluble salt forms as a precipitate.
3. Filter to separate the precipitate.
4. Wash with distilled water to remove traces of other solutions.
5. Use discussion to explore everyday uses of salts (e.g., table salt, baking soda, Epsom salt)

3. Field trip or Excursion

- **a.** Arrange the trip to a salt mine or mining facility.
- **b.** Inform learners about safety guidelines, appropriate attire, and expectations.
- **c.** Take learners to the salt mining location and have learners observe the processes involved
- **d.** Back in the classroom, encourage learners to share their experiences with the whole class.
- **e.** Connect it to classroom learning and real-world contexts.

KEY ASSESSMENTS

Level 1: Define of a salt.

Level 2: Explain how neutralisation reactions lead to the formation of salts.

Level 3: Compare and contrast acidic salts and basic salts. How do their properties differ?

Hint

The recommended mode of assessment for Week 12 is ***end of semester examination****. Refer to Appendix D at the end of this section for a Table of specification to guide the setting of the items.*

WEEKS 13 AND 14

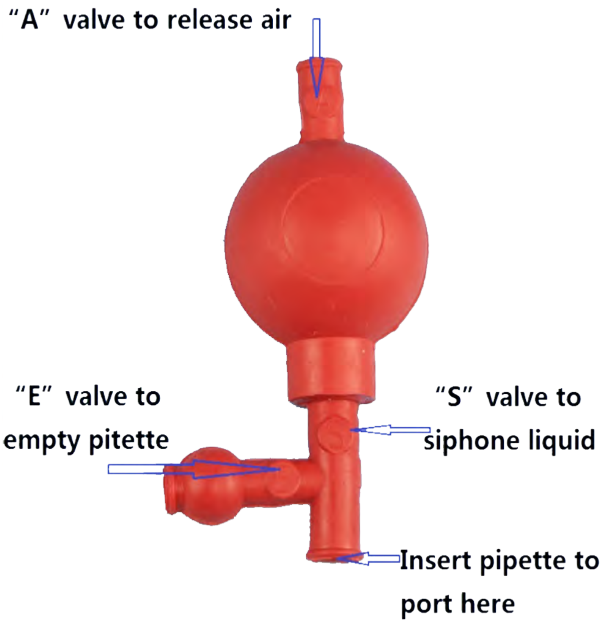
Learning Indicator: *Determine the quantity of analyte in a solution using acid – base titration*

FOCAL AREA: ACID – BASE TITRATION

Quantitative studies of some parameters of acid-base neutralisation reactions are usually carried out using a technique called titration. Titrations is a technique used to determine the concentration of an acid or a base by comparing it to the concentration of another acid or base using an acid-base indicator. The process involves adding a certain volume of a standard solution to a specific volume of a solution whose concentration is not known until neutralisation occurs.

The apparatus that are used for a titration process are;

1. **Conical flask**: It is also known as the Erlenmeyer flask. It is a conical shaped vessel with a narrow neck that allows for easy mixing and swirling of solutions. It is used as the receiving vessel for the analyte after it has been pipetted. Its transparent nature allows for easy identification of colours that develops when the acid-base indicator is added.
2. **Beaker**: It is usually used to hold excess solution or waste during the titration process.
3. **Titrant**: It is the solution whose concentration is accurately known and usually placed in the burette.
4. **Analyte**: It is the solution whose parameters are to be determined in a titration experiment. It is usually in a conical flask and an acid-base indicator added to it to develop a particular colour.
5. **Burette**: This is a graduated apparatus which is used to measure the volume of a titrant dispensed into an analyte during titration. It has to be rinsed with water and later rinsed with a small portion of the solution to be used. It must be ensured that the tap of the burette is functional. The required solution is transferred into the burette with the help of a funnel. Air bubbles must be removed by opening the tap. To reduce errors, always remove the funnel from the burette after filling the burette to prevent any excess solution from flowing into the burette and affecting the accuracy of the titration. All burette readings must be taken to two decimal places.
6. **Pipette**: The volume of the analyte is measured accurately into a conical flask with the help of a pipette. It is filled and dispensed into the conical flask with the help of a pipette bulb or filler. The pipette has two openings (the mouth and the tip).



Steps in using the pipette bulb

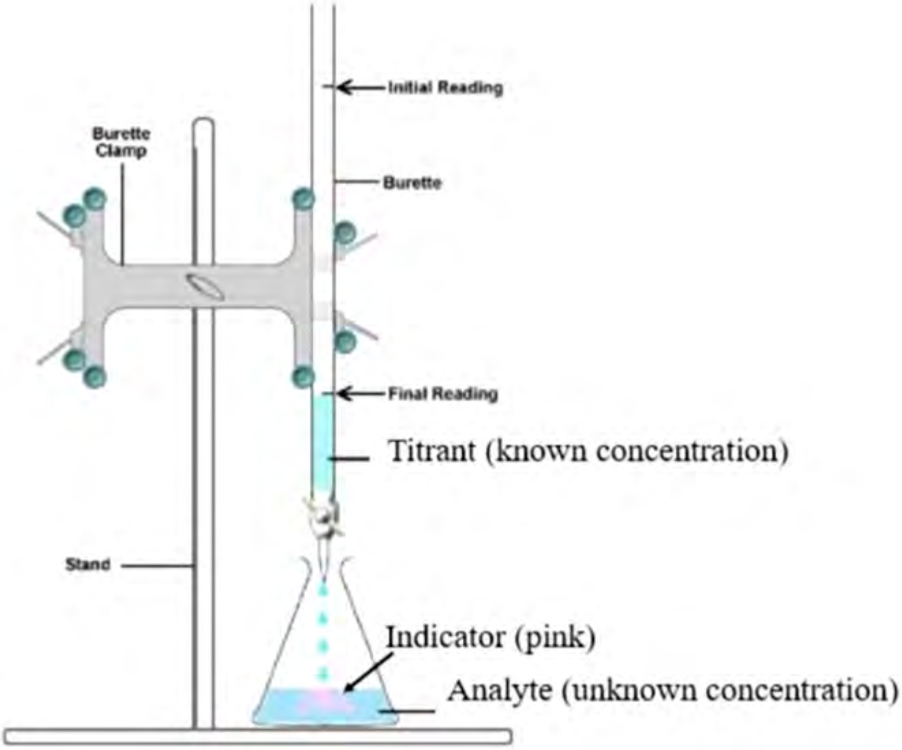
a. Insert the mouth of the pipette into the bottom of the pipette bulb.

b. Release air from the pipette bulb by squeezing the nob lettered "A" (Air valve) while simultaneously squeezing the bulb.

c. Insert the tip of the pipette into the solution to be dispensed.

d. Siphon the liquid into the pipette to the desired level by squeezing the valve lettered "S" (Syphon valve) which is just below the bulb.

e. Empty the pipette by squeezing the nob labelled "E" (Empty valve)



7. **Funnel**: It is a cone-shaped device with a wide mouth and a narrow stem placed at the open end of the burette where the titrant is poured through it to fill the burette.

   It ensures that the burette is filled accurately and without spills. It also helps to minimise wastage of solution during the transfer process.

8. **White tile**: It is placed at the base of the retort stand as a background on which the conical flask is positioned to facilitate the detection of the endpoint colour. It enhances visibility of the colour change by matching it against the colour that develops.

9. **Retort stand**: It is a laboratory device made up of three parts (the base, the rod and the clamp) that is used to securely hold the burette in position. It allows for precise control over the addition of titrants to the burette. It also enables the burette to be positioned directly above the conical flask which ensures accurate delivery of the titrant.

10. **Acid-base indicator**: It is a weak organic acid or base that changes colour according to the pH of the medium. It is used to signal the endpoint of the neutralisation process. During titration the indicator is added to the analyte in the conical flask which will develop a particular colour. As the titrant is added to the coloured analyte, the colour of the analyte remains unchanged at the initial stage due to unequal amount of OH^- and H^+. At some point, where the amount of OH^- and H^+ are exactly equal (equivalence point), the slightest volume of the titrant that will be added changes the resultant solution from acidic to basic or from basic to acidic which is characterised by a permanent colour change.

Phenolphthalein and methyl orange are two common examples of acid-base indicators used in the laboratory. Phenolphthalein is pink in basic solution and colourless in acidic solution. Thus when added to the acid solution in a flask it shows no colour. When a base is added to an acidic solution that has phenolphthalein in it, it shows no colour until the base is in slight excess then it reacting solution becomes pink in colour. This means phenolphthalein signals the endpoint by a colour change from colourless to pink. Similarly, methyl orange indicates the endpoint by a colour change. It shows a red colour in acidic medium and yellow in the basic medium.

Most indicators do not change colour at a particular pH. They do so over a range of pH from two to three units. This is called the pH range which is different for various indicators. Therefore, a pH range is the working range within which an acid-base indicator changes from one colour to another.

The table below gives the pH ranges of some acid-base indicators

Indicator	Colour change (acid to base)	pH range	pKa
Methyl orange	Red to orange	3.1 – 4.4	3.46
Methyl red	Red to yellow	4.4 – 6.0	5.00
Litmus	Red to blue	5.0 – 8.0	
Bromothymol blue	Yellow to blue	6.0 – 7.6	7.10
Phenolphthalein	Colourless to pink	8.3 – 10.0	9.4

YouTube links for procedures relating to titration (as these can be quite difficult for non-specialists):

a. https://www.youtube.com/watch?v=vn3Rx3g1VPk	
b. https://www.youtube.com/watch?v=Pvzqiz8wqJ4	
c. https://www.youtube.com/watch?v=rLc148UCT2w	

Principles of Acid–Base Indicators

1. Phenolphthalein is a weak organic acid indicator which exists as an undissociated compound in equilibrium with its conjugate base and hydrogen ions.

In solution, a weak-acid indicator (HIn) dissociates according to the equation;

$HIn \rightleftharpoons H^{+} + In^{-}$

Where HIn is the acidic form of the indicator which has a particular colour and In^{-} is the symbol of the anion part (conjugate base) of the indicator which has a different colour.

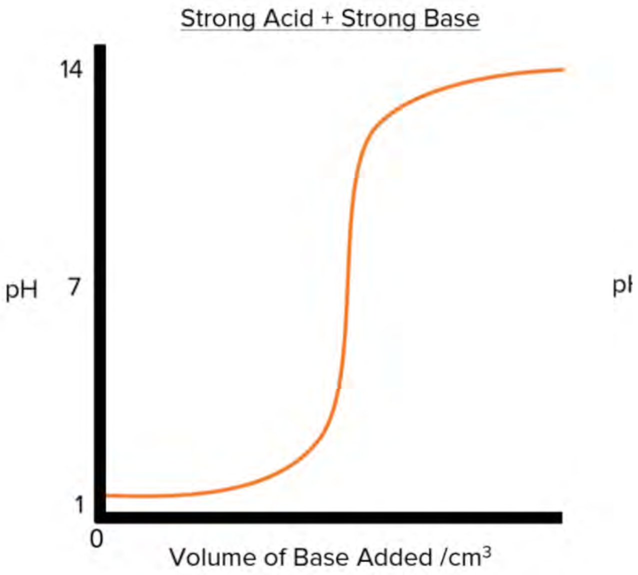
Both HIn and In^{-} are present in solution because it is a reversible reaction. The colours displayed result from the fact that HIn and In^{-} are different colours.

In acidic solutions, more H^{+} ions are present which shifts the equilibrium position to the left (according to Le Chatelier's Principle). In this instance, any In^{-} ions that are present act as Brønsted-Lowry bases and accept protons from the acid. The backward reaction is favoured and the indicator is then present in largely non-ionized form, HIn. The indicator has its acid-indicating colour.

On the contrary, a basic solution introduces hydroxyl ions (OH^-) into the equilibrium reaction. The hydroxyl (OH^-) ions which is in excess in the solution combine with the H^+ ions produced by the indicator. This reaction then reduces the concentration of the H^+ ions in solution. The forward reaction is favoured and the indicator molecules further ionize to offset the loss of H^+ ions. The indicator is present largely in the form of its anion, In^-. The solution now displays its base-indicator colour.

2. Weak organic base indicators exist as an undissociated molecule in equilibrium with its conjugate acid and hydroxyl ions according to the equation;

$MOH \rightleftharpoons M^+ + OH^-$



In the acidic solution, the H^+ (hydrogen ions) produced by the acid reacts with the OH^- (hydroxyl ions) which is the conjugate acid of the indicator which reduces its concentration at equilibrium. From the idea of Le Chatelier's Principle when the concentration of the product (OH^-) is reduced, more of the reactants would have to dissociate to form back the product. And so the forward reaction is favoured which leads to the development of the acid-indicator colour.

In the basic solution, more of the OH^- will be present (thus from the base as well as the dissociation of the indicator). Therefore, the backward reaction is favoured and the base-indicator colour is then observed.

Choice of a suitable indicator

The choice of a suitable indicator for a particular acid-base titration depends on the nature of the acid and the base involved in the titration. Acid-base titration can be grouped into the following categories based on the strength of the reacting entities as follows:

a. A strong acid with a strong base

b. A weak acid with a strong base

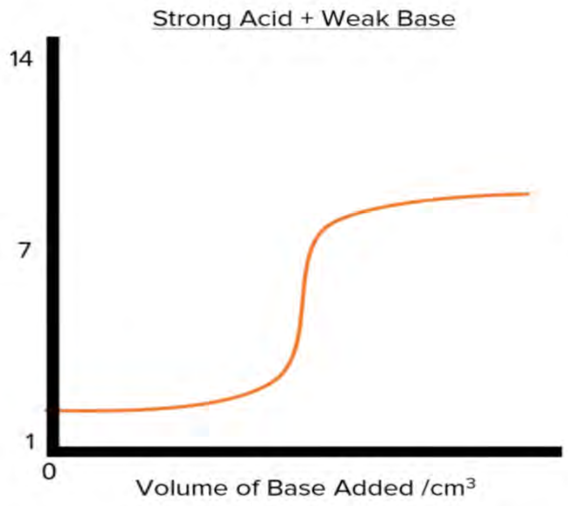
c. A strong acid with a weak base

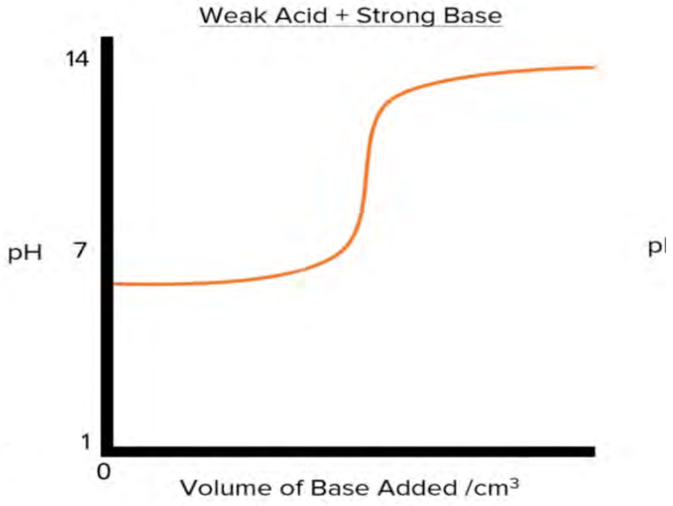
a. **For a strong acid and a strong base,** the titration can be represented on a graph (titration curve) where the volume of the titrant is the y-axis and the pH of the resulting solution is plotted on the x-axis. The curve will be relatively smooth and very steep near the equivalence point.

Because of this, a small change in titrant volume near the equivalence point results in a large pH change and many indicators would be appropriate (for instance litmus, phenolphthalein, methyl Orange or bromothymol blue).

b. For a weak acid and a strong base or

c. For a string acid and weak base





If one reagent is a weak acid or base and the other is a strong acid or base, the titration curve is irregular, and the pH shifts less with small additions of titrant near the equivalence point. These are shown in the diagrams above. Here the indicator will need to be chosen such that the equivalence point pH matches with that of the colour change of the indicator. For example, methyl-orange can be used with strong acids and Phenolphthalein can be used with strong bases.

If both the acid and base are weak then Bromothymol blue would be a good choice.

The following terms are associated with titration

1. **Titrant**: It is a standard solution which is used to determine the properties of another solution (analyte) by reacting with it. It is usually placed in a burette during the titration process.

   Characteristics of titrant

   a. The concentration of a titrant in titration is accurately known and precise.

  - **b.** It is a pure substance as far as is possible.
  - **c.** It should be able to react completely with the analyte
  - **d.** The reaction between the titrant and analyte should have a clear and measurable endpoint which indicates the completion of the reaction and characterised by a sharp colour change.

2. **Analyte:** It is a solution of an unknown concentration that is reacted with a titrant. It is the sample within the titration process whose properties are being analysed or measured. It is pipetted into a conical flask during the titration process.
3. **Equivalence point:** It is a point in the titration process at which all the titrant reacts with the analyte. At this point the reaction is stoichiometrically complete with exact proportion of titrant and analyte. It is a theoretical point which is independent of the experimental conditions or indicators.
4. **End point:** It is the pH of the reaction mixture at the point of neutralisation which is indicative of a sharp colour change depending on the nature of indicator.
5. **Titre value:** It is the volume of the titrant which reacts with the given volume of the analyte at the end point where a colour change is attained. It is often determined as the difference between the initial and final burette readings for a given set of titration.
6. **Consistent titre values:** They are a set of titre values which are close to each other (differ by $\pm 0.20\, cm^3$)
7. **Average titre:** It is the average of consistent titre values obtained from a titration experiment.
8. **Indicator:** They are organic dyes or plant extracts that changes colour due to changes in the pH of a certain medium.

Types of acid-base titration

1. Simple or Direct acid-base titrations

In the direct titration, a sample is titrated directly with a suitable standard solution. It involves the use of only one acid-base indicator. It is usually used in the standardisation of an acid or a base in volumetric analysis. Standardisation is the process of determining the exact amount of substance present in a specific volume of solution by reacting it with a primary standard solution. The amount of reagent consumed to the endpoint together with mole ratio of the reacting molecules are critically manipulated to determine the properties of the unknown sample in the titration.

Steps

1. Set up the titration apparatus by clamping a burette to the retort stand. Rinse the burette with distilled water, then rinse with the titrant and fill it with the titrant.

It must be ensured that the burette is properly calibrated and the initial volume set. Record the initial volume of the titrant.

2. Pipette a known volume (usually 20 cm^3 or 25 cm^3) of the analyte into a conical flask.
3. Add few drops (2 to 3) of an appropriate indicator to the analyte in the conical flask, swirl uniformly to develop a uniform colour and record the colour formed.
4. Carefully add the titrant from the burette to the analyte in the conical flask until a permanent colour change is observed (endpoint is reached).
5. Tabulate the initial burette readings, final burette readings and the volume of titrant used (all to two decimal places).
6. Repeat the titration process to obtain consistent titre values and calculate the average titre used.
7. With the help of the balanced chemical equation between the titrant and the analyte, determine the mole ratio and use it to calculate the concentration of the analyte.

Worked Example

A solution of 0.2 mol/dm^3 of HCl was titrated against 25 cm^3 of NaOH using methyl orange as indicator. The results obtained in the titration experiment are as follows

Titrant: 0.2 mol/dm^3 HCl

Analyte: NaOH of unknown concentration

Capacity of Burette: 50 cm^3

Capacity of pipette: 25 cm^3

Indicator used: 2 drops of methyl orange

Colour change: Yellow to red

Titration Table

Burette readings/cm^3	I	II	III
Final reading	12.80	25.30	12.60
Initial reading	0.00	12.80	0.00
Titre value	12.80	12.50	12.60

Calculate

a. The average titre values

b. The molar concentration of NaOH

c. The mass concentration of NaOH

Solution

a) Average titre $= \frac{12.60 + 12.50}{2}$

$= 12.55\ \text{cm}^3$

b) From the titration reaction

$HCl + NaOH \rightarrow NaCl + H_2O$

The mole ratio is given by

$\frac{n(HCl)}{n(NaOH)} = \frac{1}{1}$

$n(HCl) = n(NaOH)$

But $n = C \times V$

Where n is the amount of substance in moles

C is the concentration of the substance in mol/dm^3

V is the volume of solution used in dm^3

$\Rightarrow C_{HCl} \times V_{HCl} = C_{NaOH} \times V_{NaOH}$

$C_{HCl} = 0.2\ \text{mol dm}^{-3}$

V_{HCl}, average titre $= 12.55\ \text{cm}^3$

$C_{NaOH} = ?$

V_{NaOH}, pipetted volume $= 25\ \text{cm}^3$

$0.2 \times 12.55 = C_{NaOH} \times 25$

$C_{NaOH} = \frac{0.2 \times 12.55}{25}$

$C_{NaOH} = 0.10\ \text{mol dm}^{-3}$

c) Mass concentration, $\rho = C \times M$

Where C is the molar concentration in moldm^{-3}

M is the molar mass of the substance

$C_{NaOH} = 0.10\ \text{mol dm}^{-3}$

$M_{NaOH} = 23 + 16 + 1$

$= 40\ \text{g mol}^{-1}$

$\rho_{NaOH} = 0.10 \times 40$

$= 4.0\ \text{g dm}^{-3}$

Aside from standardisation of solutions, direct acid-base titrations have many applications. They can also be used to determine:

1. The percentage purity of a substance
2. The number or number of moles of ions present in a solution.
3. The molar mass of a substance
4. The solubility of a substance.

Back or Indirect titration

In back titrations an unknown substance (A) which cannot be determined directly by titration is reacted with an excess amount of a standard reagent (B) where all of A is neutralised leaving excess B. The remaining portion of B which is left after reaction A is then titrated against another standard reagent. The actual sample of B which is neutralised by A is then obtained by subtracting the quantity of B left after reaction with A, from the original quantity of B. In this way the amount of A that reacted can be calculated.

Two common examples that involves the principles of back titration are;

1. **Determination of acetylsalicyclic acid composition in tablets of aspirin**

   Aspirin is insoluble in water but soluble in an alkaline solution (NaOH) and warmed to ensure complete reaction. The excess of the NaOH that remains after the reaction is back titrated with a standard acid solution (HCl) using phenolphthalein as an indicator. From the results of the titration the quantity of NaOH that reacted with the aspirin can be determined and by extension the amount of aspirin in a tablet.

Worked Example 1

A new brand of aspirin (Ecotrin) was suspected to contain low doses of acetylsalicyclic acid. To verify this, a chemist crushed two tablets of aspirin and mixed with 25 cm^3 of 0.1 $moldm^{-3}$ NaOH. The mixture was warmed and transferred to a 250 cm^3volumetric flask. A solution of 0.05 $moldm^{-3}$ H_2SO_4 was titrated against 25 cm^3 portions of a solution of Sodium hydroxide using phenolphthalein as an indicator. The results obtained in the experiment are as follows:

Table of Results

Burette reading/ cm^3	A	B	C
Volume of acid used	8.90	8.70	8.60

Calculate

The average titre values

a. Moles of H_2SO_4 that reacted

b. Moles of excess NaOH present in the solution.

c. Moles of NaOH that reacted with acetylsalicyclic acid

d. Mass of acetylsalicyclic acid in one tablet.

e. If the mass of each tablet is 2g, determine the percentage mass of pure acetylsalicyclic acid in each tablet.

[C=12, H=1, O=16, Na=23, S=32]

Solution

$$\text{Average titre} = \frac{8.70 + 8.60}{2}$$

$$= 8.65 \text{ cm}^3$$

i. $n_{H_2SO_4} = C \times V$

$C_{H_2SO_4} = 0.05 \text{ moldm}^{-3}$

$V_{H_2SO_4} = 0.00865 \text{ dm}^3$

$n_{H_2SO_4} = 0.05 \times 0.00865$

$= 0.000433$

$= 4.33 \times 10^{-4}$ mol

ii. From the balanced equation

$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

The mole ratio for the titration is

$$\frac{n_{NaOH}}{n_{H_2SO_4}} = \frac{2}{1}$$

$n_{NaOH} = 2 \times n_{H_2SO_4}$

But $n_{H_2SO_4} = 4.33 \times 10^{-4}$ mol

$n_{NaOH} = 2 \times 4.33 \times 10^{-4}$

$= 8.65 \times 10^{-4}$ mol

Therefore, moles of NaOH that was present in the 25 $cm^3$and reacted with the acid is 8.65×10^{-4} mol.

If 25 cm^3 of NaOH contains 8.65×10^{-4} mol

Then 250 cm^3 of NaOH will contain y mol

$$y = \frac{250}{25} \times 8.65 \times 10^{-4}$$

$$= 8.65 \times 10^{-3} \text{ mol}$$

iii. From the relation n = CV

Initial concentration of NaOH before it was mixed with the crushed tablets = 1.0 mol dm^{-3}.

Volume of the solution that was mixed with the crushed tablets = 25 cm^3 = 0.025 dm^3

$n_{NaOH} = 1 \times 0.025$

$= 0.025$ mol

If 0.025 mol of NaOH was mixed with the crushed tablets and $8.65 \text{ x } 10^{-3}$mol remained, then,

Mole of NaOH that reacted with the acetylsalicyclic acid = 0.025 – 0.00865

= 0.01635 mol

iv. Acetylsalicyclic acid reacts with sodium hydroxide according to the equation

$C_9H_8O_4 + NaOH \rightarrow C_9H_7O_4Na + H_2O$

The mole ratio is given by

$$\frac{n_{NaOH}}{n_{C_2}} = \frac{2}{1}$$

$$n_{H_2SO_4} = \frac{n_{NaOH}}{2}$$

$$= \frac{0.01637}{2}$$

$$= 0.008185 \text{ mol}$$

but $m = n \times M$

$M(C_9H_8O_4) = (12 \times 9) + (1 \times 8) + (16 \times 4)$

$= 180 \text{ gdm}^{-3}$

$m = 0.008185 \times 180$

$= 1.47$ g

Two tablets contain 1.47 g

Each tablet contains $\frac{1.47}{2} = 0.735$ g

vi. If the mass of each tablet is 2g

$$\% \text{ (mass)} = \frac{\text{mass of pure acetylsalicyclic acid}}{\text{mass of tablet}}$$

$$\% \text{ (mass)} = \frac{0.735}{2}$$

$$\% \text{ (mass)} = 36.7$$

Back titration has many applications, some of which include the estimation of:

a. The $CaCO_3$ content in a certain quantity of limestone or egg shell

b. The content of ammonia in an ammonium salt

c. The $Ca(OH)_2$ content in a sample of soil.

2. Double-indicator titration

It is used to determine the concentration of individual components of a mixture that reacts differently with a titrant using more than one indicator. The reaction of the components of the mixture with the titrant takes place in stages where one forms an acidic product while the other forms a product that is alkaline. Indicators that have their working ranges around the products that will be formed are used to identify the endpoint of the various stages of the reaction. Methyl orange and phenolphthalein are usually used to identify the volume of the titrant that is used at the stages.

There are two methods of double-indicator titration

a. Continuous method

b. Discontinuous method.

A typical example of double indicator titration which can be illustrated using these two methods is the reaction between a strong acid (HCl) and a mixture of a diprotic and monoprotic bases $(Na_2CO_3/NaHCO_3)$.

a. ***Continuous Method***

i. The titration apparatus is set-up. The titrant which is usually the acid is placed in the burette.

ii. 20 or 25 cm^3 of the basic mixture is pipetted into a conical flask and about two drops of phenolphthalein added. The mixture is swirled gently to develop a uniform colour after which it is placed under the burette.

iii. Slowly add the titrant from the burette to the coloured solution in the conical flask until a colour change develops. Record the volume of HCl used for the phenolphthalein endpoint as $x\ cm^3$. The equation for the reaction with the phenolphthalein indicator is represented as;

$$Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$$

iv. Add about two drops of methyl orange indicator to the resulting solution from the initial titration.

v. Continue the titration until a second colour change is developed. Record the new volume of HCl used for the methyl orange endpoint as $y\ cm^3$. The equation for the methyl orange indicator is represented as;

$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$$

vi. In the first part of the titration, all Na_2CO_3 in the mixture reacted with the HCl to produce $NaHCO_3$, which adds to the original $NaHCO_3$ that was present. At the methyl orange stage, the converted $NaHCO_3$ together with the original $NaHCO_3$ reacts with HCl and are all converted to NaCl, H_2O and CO_2.

vii. To determine the volume of HCl that reacted with the original $NaHCO_3$ in the sample the first titre value ($x\ cm^3$) is subtracted from the second titre ($y\ cm^3$). These individual values are then used to calculate parameters of the components present in the mixture.

Worked Example

A mixture of Na_2CO_3 *and* $NaHCO_3$ reacted with 0.1 moldm^{-3} HCl.

The results obtained when phenolphthalein indicator was used is

Burette reading/cm^3	A	B	C
Volume of HCl used (x)	10.00	10.50	10.00

The titration continued with methyl orange being added to the resulting solution. The results obtained at the methyl orange endpoint is

Burette reading/cm^3	A	B	C
Volume of HCl used (y)	19.00	19.30	19.00

a. Write a balanced chemical equation to represent the reaction that took place at the;

i. Phenolphthalein endpoint

ii. Methyl orange endpoint

b. Calculate the concentration in moldm^{-3} of,

i. Na_2CO_3

ii. $NaHCO_3$

In the mixture that was used for the titration

b. Determine the mass concentration of;

  i. Na_2CO_3

  ii. $NaHCO_3$

c. Determine the percentage composition of Na_2CO_3 in the mixture.

Solution

a. The balanced chemical equation for

  i. Phenolphthalein endpoint is represented by:

  $Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$

  ii. Methyl orange endpoint is represented by:

  $NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$

b. From the titration reaction

  i. The average titre value at the phenolphthalein endpoint is given by

  $$x = \frac{10.00 + 10.00}{2} = 10.00\ cm^3$$

  10.00 cm^3 of HCl reacted with half of the Na_2CO_3 present in the mixture.

  If 0.5 of Na_2CO_3 required 10 cm^3

  Then 1.0 of Na_2CO_3 will require $= \frac{1.0}{0.5} \times 10.00$

  $= 20.00\ cm^3$

  The complete reaction of HCl with Na_2CO_3 that requires 20.00cm^3 of the titrant (HCl) is represented by the chemical equation:

  $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$

  Mole ratio for the complete reaction is

  $$\frac{n(Na_2CO_3)}{n(HCl)} = \frac{1}{2}$$

  But n = CV

  $$\frac{C_{Na_2CO_3} \times V_{Na_2CO_3}}{C_{HCl} \times V_{HCl}} = \frac{1}{2}$$

  $$C_{Na_2CO_3} \times V_{Na_2CO_3} = \frac{1}{2} \times C_{HCl} \times V_{HCl}$$

  $$C_{Na_2CO_3} = \frac{1}{2} \times \frac{C_{HCl} \times V_{HCl}}{V_{Na_2CO_3}}$$

  C_{HCl}, concentration of HCl = 0.1moldm^3

  V_{HCl}, volume of HCl for complete reaction = 20.00 cm^3

$V_{Na_2CO_3}$, pipetted volume of the mixture used for the titration $= 25.00\ cm^3$

$C_{Na_2CO_3}$, concentration of Na_2CO_3 in the mixture $= C_{Na_2CO_3} = \frac{1}{2} \times \frac{0.1 \times 20.00}{25.00}$

$= 0.0400\ mol\,dm^{-3}$

Alternatively

10 cm^3 of HCl was required at the phenolphthalein endpoint and can be represented by the equation:

$Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$

The mole ratio for this reaction is given by

$$\frac{n(Na_2CO_3)}{n(HCl)} = \frac{1}{1}$$

$$\frac{C_{Na_2CO_3} \times V_{Na_2CO_3}}{C_{HCl} \times V_{HCl}} = \frac{1}{1}$$

$$C_{Na_2CO_3} \times V_{Na_2CO_3} = C_{HCl} \times V_{HCl}$$

$$C_{Na_2CO_3} = \frac{C_{HCl} \times V_{HCl}}{V_{Na_2CO_3}}$$

$$C_{Na_2CO_3} = \frac{0.1 \times 10.00}{25.00}$$

$$= 0.0400\ mol\,dm^{-3}$$

ii. The average titre value at the methyl orange endpoint is given by

$$y = \frac{19.00 + 19.00}{2} = 19.00\ cm^3$$

19.00 cm^3 of HCl reacted with the $NaHCO_3$ that was formed from the Na_2CO_3 and the original $NaHCO_3$ that was present in the mixture.

If half of Na_2CO_3 was neutralised by 10.00 cm^3, then the remaining half that was present in solution will also react with the same volume of HCl.

Volume of HCl that reacted with the original $NaHCO_3$ in the mixture = z cm^3

$= Titre\ value - Volume\ that\ reacted\ with\ half\ of\ NaHCO_3$

$z = y - x$

$= 19.00 - 10.00$

$= 9.00\ cm^3$

The reaction equation is given by:

$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$

The mole ratio for this reaction is

$$\frac{n(NaHCO_3)}{n(HCl)} = \frac{1}{1}$$

$$C_{NaHCO_3} \times V_{NaHCO_3} = C_{HCl} \times V_{HCl}$$

$$C_{NaHCO_3} = \frac{C_{HCl} \times V_{HCl}}{V_{NaHCO_3}}$$

$$C_{NaHCO_3} = \frac{0.1 \times 9.00}{25.00}$$

$$= 0.0360\ mol\,dm^{-3}$$

c. Mass concentration, ρ = Concentration (C) x Molar mass (M)

i. $C_{Na_2CO_3} = 0.04\ mol\,dm^{-3}$

$$M_{Na_2CO_3} = 2(23) + 12 + 3(16)$$

$$= 46 + 12 + 48$$

$$= 106\ g mol^{-}$$

$$\rho_{Na_2CO_3} = 0.04 \times 106$$

$$= 4.24\ g\,dm^{-3}$$

ii. $C_{NaHCO_3} = 0.036\ mol\,dm^{-3}$

$$M_{NaHCO_3} = 23 + 1 + 12 + 3(16)$$

$$= 84\ g\,mol^{-3}$$

$$M_{NaHCO_3} = 0.036 \times 84$$

$$= 3.02\ g\,mol^{-3}$$

d. Percentage composition of $Na_2CO_3 = \frac{mass\ of\ substance}{Total\ mass\ of\ mixture} \times 100$

Mass of Na_2CO_3 = 4.24 g

Total mass of mixture = 4.24 + 3.02

= 7.26 g

$$\%\ Na_2CO_3 = \frac{4.24}{7.26} \times 100$$

$$= 58.4\%$$

i. Discontinuous Method

a. The titration apparatus is set-up where the titrant (HCl) is placed in the burette.

b. 20 or 25 cm^3 of the basic mixture is pipetted into a conical flask and about two to three drops of phenolphthalein indicator added. The conical flask is swirled gently to develop a uniform colour before it is placed under the burette.

c. The titrant is carefully added to the solution in the conical flask until a permanent colour change is developed. The volume of the titrant (HCl) used is the then recorded as $V_1\ cm^3$.

The equation for the reaction at the phenolphthalein indicator is given by:

$Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$

d. Where $V_1\ cm^3$ of HCl was used to convert all the Na_2CO_3 in the mixture to $NaHCO_3$. Therefore $V_1\ cm^3$ is used to calculate the amount of Na_2CO_3 in the mixture.

e. The solution obtained after the phenolphthalein endpoint is discarded.

f. A new 20 or 25 cm^3 of the mixture is pipetted into a different conical flask. Two drops of methyl orange are added and swirled to develop a uniform colour.

g. It is then titrated against the HCl until a colour change develops. The volume of HCl used is then recorded as $V_2\ cm^3$. In this titration all the components of the mixture, together with the intermediate reacts with the HCl in the equation given by:

$Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$

$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$

This implies that $V_2\ cm^3$ of HCl was used to convert all the Na_2CO_3 in the mixture to $NaHCO_3$ and then further converted it together with the original $NaHCO_3$ to NaCl, H_2O and CO_2.

h. The volume of HCl that reacted with the original $NaHCO_3$ that was present in the mixture is given by the relation ($V_3\ cm^3$), where

$V_3\ cm^3 = V_2\ cm^3 - 2V_1\ cm^3$

Therefore $V_3\ cm^3$ is used to calculate the amount of original $NaHCO_3$ that was present in the mixture

Worked Example

A sample of Na_2CO_3 and $NaHCO_3$ were mixed together to form a solution. 25 cm^3 of this solution was titrated against a 0.1 moldm^{-3} HCl solution using phenolphthalein as an indicator.

The results obtained was recorded as:

Burette readings/cm^3	**I**	**II**	**III**
Volume of acid used (V_1)	9.60	10.00	10.00

Another 25 cm^3 portions of the same solution was also titrated against 0.1 $mol\,dm^{-3}$ HCl solution using methyl orange as an indicator.

The results obtained at the methyl orange endpoint was recorded as:

Burette readings/cm^3	**I**	**II**	**III**
Volume of acid used (V_2)	23.90	23.50	23.50

a. Determine the average titre value at the;

- **i.** Phenolphthalein endpoint
- **ii.** Methyl orange endpoint

b. Determine the volume of HCl that reacted with;

- **i.** All of the Na_2CO_3 present
- **ii.** The original quantity of $NaHCO_3$ present

c. Calculate the concentration in $mol\,dm^{-3}$ of;

- **i.** Na_2CO_3
- **ii.** $NaHCO_3$

d. Determine the mass concentration of;

- **i.** Na_2CO_3
- **ii.** $NaHCO_3$

e. Determine the percentage composition of $NaHCO_3$ in the mixture

Solution

a. Calculation of the average titre values at the various stages of the titration

i. At the phenolphthalein endpoint, the average titre value is given by;

$$= \frac{10.00 + 10.00}{2} \text{ (NOTE : } \textit{the } 9.60 \textit{ value is not consistent)}$$

$$= 10.00\ cm^3$$

ii. At the methyl orange endpoint, the average titre value is given by;

$$= \frac{23.50 + 23.50}{2}$$

$= 23.50\ cm^3$

b. To determine the volume of HCl that reacted with the various components of the mixture;

i. The equation for the complete reaction of Na_2CO_3 with HCl goes by the equation:

$Na_2CO_3 + 2HCl \rightarrow NaCl + H_2O + CO_2$

But at the phenolphthalein endpoint, the reaction that took place goes by the equation:

$Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$

This means that 10.00 cm^3 of HCl reacted with only half of the Na_2CO_3

If 0.5 of Na_2CO_3 required 10.00 cm^3 of HCl

Then 1.0 of Na_2CO_3 will require $= \frac{1.0}{0.5} \times 10.00$

$= 20.00\ cm^3$

Therefore 20.00 cm^3 of HCl reacted with all of the Na_2CO_3 that was present in the mixture.

ii. At the methyl orange endpoint, two sets of reaction occurred where all the components of the mixture were neutralised by the HCl. These reactions can be represented by the equations;

$Na_2CO_3 + 2HCl \rightarrow NaCl + H_2O + CO_2$ ………(1)

$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$ ……………(2)

Equation (1) required 20.00 cm^3

But from the methyl orange titre value,

Equation (1) + (2) required a volume of 23.50 cm^3

$\Rightarrow (1) + (2) = 23.50\ cm^3$

$but\ (1) = 20.00\ cm^3$

substituting $(1) = 20.00$ into the equation

$20.00 + (2) = 23.50$

$(2) = 23.50 - 20.00$

$(2) = 3.50\ cm^3$

Therefore 3.50 cm^3 of HCl reacted with the original $NaHCO_3$ that was present in the mixture.

c.

i. The equation for the reaction of Na_2CO_3 at the phenolphthalein endpoint is given by:

$Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$

The mole ratio of this reaction is $\frac{n(Na_2CO_3)}{n(HCl)} = \frac{1}{1}$

But n = CV

$$C_{Na_2CO_3} \times V_{Na_2CO_3} = C_{HCl} \times V_{HCl}$$

$$C_{Na_2CO_3} = \frac{C_{HCl} \times V_{HCl}}{V_{Na_2CO_3}}$$

$$C_{Na_2CO_3} = \frac{0.1 \times 10.00}{25.00}$$

$$= 0.0400\ mol\,dm^{-3}$$

Therefore, the concentration of Na_2CO_3 present in the mixture is 0.0400 mol dm^{-3}

ii. The reaction that took place between the original $NaHCO_3$ in the mixture and HCl is represented by:

$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$

This required 3.50 cm^3 of the HCl.

The mole ratio of this reaction is

$$\frac{n(NaHCO_3)}{n(HCl)} = \frac{1}{1}$$

$$\frac{C_{NaHCO_3} \times V_{NaHCO_3}}{C_{HCl} \times V_{HCl}} = \frac{1}{1}$$

$$C_{NaHCO_3} \times V_{NaHCO_3} = C_{HCl} \times V_{HCl}$$

$$C_{NaHCO_3} = \frac{C_{HCl} \times V_{HCl}}{V_{NaHCO_3}}$$

$$C_{NaHCO_3} = \frac{0.1 \times 3.50}{25.00}$$

$$C_{NaHCO_3} = 0.0140\ mol\,dm^{-3}$$

Therefore, the concentration of the original $NaHCO_3$ in the solution is 0.0140 $mol\,dm^{-3}$

d. Mass concentration, $\rho = Concentration\ (C) \times Molar\ mass\ (M)$

i. $C_{Na_2CO_3} = 0.04\ mol\,dm^{-3}$

$M_{Na_2CO_3} = 106\ g\,mol^{-}$

$$\rho_{Na_2CO_3} = 0.04 \times 106$$

$$= 4.24\ g\,dm^{-3}$$

Therefore, the mass concentration of Na_2CO_3 is 4.24 $g\,dm^{-3}$

ii. $C_{NaHCO_3} = 0.014\ mol\,dm^{-3}$

$M_{NaHCO_3} = 84\ g\,dm^{-3}$

$$\rho_{NaHCO_3} = 0.014 \times 84$$

$$= 1.18\ g\,dm^{-3}$$

Therefore, the mass concentration of $NaHCO_3$ is 1.18 $g\,dm^{-3}$

$$\text{Percentage composition} = \frac{\text{Mass of substance}}{\text{Total mass of mixture}} \times 100$$

$$= \frac{.18}{(4.24 + 1.18)} \times 100$$

$$= 21.8\ \%$$

Double indicator titrations have widespread applications. Examples of such reactions include:

a. HCl reacting with Na_2CO_3

b. HCl reacting with a mixture of Na_2CO_3 and NaOH

c. NaOH reacting with a mixture of H_3PO_4 and HCl.

Learning Tasks

1. Set up a titration apparatus.
2. Fill the burette with the titrant, remove air bubbles and set the level of the solution to the zero mark.
3. Pipette 20 or 25 cm^3 of the analyte into a conical flask, add few drops of methyl orange and record the colour that develops.
4. Perform the titration to identify the endpoint.
5. Record the volume of the titrant used.
6. Draw a titration table.
7. Perform calculation.

PEDAGOGICAL EXEMPLARS

Exploratory learning

Perform titrations and use the data obtained to determine the concentration of an analyte in a solution. The titrations include:

1. Simple acid-base titrations for determining percentage purity, percentage of water of crystallisation in a hydrated salt, relative atomic mass.
2. Back or Indirect titration.
3. Double-indicator titration.
4. Perform titrations whilst monitoring the pH of the analyte in order to produce volume vs pH graphs to determine equivalence points for each type of titration:
   - **a.** strong acid vs strong base
   - **b.** strong acid vs weak base
   - **c.** strong base vs weak acid
   - **d.** weak base vs weak acid

KEY ASSESSMENTS

Level 1: Write a balanced chemical equation to represent the titration of HCl with NaOH.

Level 2

1. Discuss the types and applications of titration in pharmaceutical manufacturing industry in Ghana.
2. If 22.60 cm^3 of a 0.6 $moldm^{-3}$ NaOH was required to neutralise a 25 cm^3 of a strong dibasic acid. Determine the concentration of the acid.
3. Copy and complete the titration table below;

Burette readings/cm^3	1	2	3
Final	21.30		25.40
Initial	0.00	21.30	
Titre value		21.40	4.00

Hint

The recommended mode of assessment for week 13 and 14 is ***individual project work****.*
Use the level 2 question 1 as a sample question.

APPENDIX D: END OF SEMESTER EXAMINATION

Nature

This end of semester examination will comprise of paper 1 and 2. Paper 1 will be made up of compulsory 30 multiple choice questions while Paper 2 will have sections A and B. Section A will consist of 5 short answer questions for learners to attempt all and section B will have 3 essay type questions of similar difficulty levels which learners must answer only 1. The exam should cover lessons learnt from weeks 1-12. Time allowed for this paper is 2hours

Resources

Stationery, Printer, Scannable answer sheet, Calculators, Answer booklets, Pens, Pencils, etc

Sample question

MCQ

A solution with a hydrogen ion concentration of 0.01M will have a P^H of…

- **A.** 1
- **B.** 2
- **C.** 12
- **D.** 14

Short Answer Type

a) Define acid and base according to Lewis

b) Differentiate between acid salt and normal salt

c) State any 2 major ideas in the collision theory?

d) State 3 factors that affect equilibrium position

e) Identify Bronsted-Lowry acids in each reaction:$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Essay Type

Discuss 5 factors that affect the rate of a chemical reaction

How to Administer

1. Prepare examination questions and materials
2. Inform learners of examination details (date, time location, etc)
3. Distribute examination materials (questions, answer sheets, etc)
4. Monitor learners to prevent cheating and maintain a quiet and secure environment

Table of Specification

Weeks	Focal Area(s)	Type of Questions	DoK Levels				Total
			1	2	3	4	
1	Enthalpy Change and Standard Enthalpy Changes	Multiple Choice	1	1	-	-	2
		Essay	-	-	1	-	1
2	Experimental Determination of Enthalpy Change	Multiple Choice	1	2	-	-	3
		Short Answer	-	1	-	-	1
3	Hess's Law of Constant Heat Summation	Multiple Choice	-	2	1	-	3
4	Rate of Reaction	Multiple Choice	1	-	1	-	2
		Essay			1	-	1
5	Collision Theory	Multiple Choice	2	1	-	-	3
		Short Answer	-	1	-	-	1
6	Dynamic Equilibrium	Multiple Choice	1	2	-	-	3
		Short Answer		1	-	-	1
7	Equilibrium Constant Expressions	Multiple Choice	1	2	-	-	3
8	Le Chateliers Principle and Factors That Affect Equilibrium Positions	Multiple Choice	1	1	1	-	3
		Short Answer	1	-	-	-	1
9	Concepts of Acids and Bases	Multiple Choice	1	1	1	-	3
		Short Answer	-	1	-	-	1
10	Physical and Chemical Properties of Acids and Bases	Multiple Choice	-	1	-	-	1
		Essay	-	-	1	-	1
11	Strengths of Acids and Bases	Multiple Choice	-	1	1	-	2
		Short Answer	-	1		-	1
12	Salts	Multiple Choice	-		1	-	1
	Total		**10**	**19**	**9**		**38**

APPENDIX E: INDIVIDUAL PROJECT WORK ASSESSMENT

Presentation of written report (approx. 2-3 pages) demonstrating a comprehensive understanding of the process of titration in pharmaceutical industry in Ghana.

Scoring Rubrics for Project Task

Criteria	Excellent (4)	Good (3)	Fair (2)	Poor (1)
Understanding of Concepts	Comprehensive understanding of at least 4 types of titrations; such as acid-base, Redox, precipitation, potentiometric and complexometric titrations	Good understanding of 3 types of acid-base, Redox, precipitation, potentiometric and complexometric titrations	Good understanding of 2 acid-base, Redox, precipitation, potentiometric and complexometric titrations	Lacks understanding acid-base, Redox, precipitation, potentiometric and complexometric titrations
Relevance to Ghana	Thoroughly explains 4 titration applications as quality control, regulatory compliance, formulation development and patient specific drug dosage adjustment to Ghana's pharmaceutical industry.	Explains 3 titration applications as quality control, regulatory compliance, formulation development and patient specific drug dosage adjustment to Ghana's pharmaceutical industry.	Explains 2 titration applications as quality control, regulatory compliance, formulation development and patient specific drug dosage adjustment to Ghana's pharmaceutical industry.	Explains 1 titration applications as quality control, regulatory compliance, formulation development and patient specific drug dosage adjustment to Ghana's pharmaceutical industry.

Criteria	Excellent (4)	Good (3)	Fair (2)	Poor (1)
Clarity and Structure	Score for all 3 well-organised presentations; ideas flow logically; minimal grammatical errors.	Score for 2 well-organised presentations; ideas flow logically; minimal grammatical errors.	Score for 1 well-organised presentation; ideas flow logically; many grammatical errors.	Unclear presentation; disorganised; many grammatical errors.
Research Depth	Extensive research with 4 examples of the applications such as P^H, active pharmaceutical, halide ion and metal ion concentrations determination with 2 sources cited effectively.	Research with 3 examples of the applications such as P^H, active pharmaceutical, halide ion and metal ion concentrations determination with 2 sources cited effectively.	Research with 2 examples of the applications such as P^H, active pharmaceutical, halide ion and metal ion concentrations determination with at least 1 source cited effectively.	Research with 1 example of the applications such as P^H, active pharmaceutical, halide ion and metal ion concentrations determination with at least 1 source cited effectively.

Total = 16 marks

SECTION 5: TRENDS OF CHEMICAL AND PHYSICAL PROPERTIES OF ELEMENTS AND THEIR COMPOUNDS IN THE PERIODIC TABLE

Strand: Systematic Chemistry of the Elements

Sub-Strand: Periodicity

Learning Outcome: *Predict and describe the trends of chemical and physical properties of elements and their compounds in the periodic table.*

Content Standard: Demonstrate Understanding of the regular patterns of chemical and physical properties of elements and their compounds in the periodic table.

INTRODUCTION AND SECTION SUMMARY

This two-week section explores the interesting world of Period 3 elements and their compounds. Learners will discover the patterns and trends in the physical and chemical properties of these elements and their hydrides, oxides, hydroxides, and chlorides. By looking at the relationships between atomic number, electron configuration, and resulting properties, learners will understand what makes each element and compound unique.

By the end of this section, learners will be able to

1. Describe the patterns in physical and chemical properties of Period 3 elements.
2. Describe the patterns in physical and chemical properties of compounds (hydrides, oxides, hydroxides, and chlorides) of Period 3 elements.
3. Compare and contrast the properties of different Period 3 elements and their compounds.

The weeks covered by the section are

Week 15: *Describe the patterns in physical and chemical properties of the Period 3 elements.*

Week 16: *Describe the patterns in physical and chemical properties of compounds of Period 3 elements.*

SUMMARY OF PEDAGOGICAL EXEMPLARS

Learners will explore the properties of Period 3 elements and their compounds over two weeks. In Week 15, they will discover trends in metallic, physical, and chemical properties through discussions, digital research, and graphing activities. They will analyse data, identify patterns, and develop problem-solving skills. In Week 16, learners will predict and describe the physical properties of Period Three element compounds, including hydrides, oxides, hydroxides, and chlorides. They will investigate the thermal stability of carbonates and nitrates, conduct experiments, and write chemical equations. Through Talk-for-Learning, Think-Pair-Share, Collaborative Learning, and Experiential Learning approaches, learners will engage in a dynamic and interactive journey, deepening their understanding of chemical properties and reactions.

ASSESSMENT SUMMARY

Evaluate learners' understanding of the patterns in physical and chemical properties of hydrides, oxides, hydroxides, and chlorides of Period 3 elements.

1. **Assessment Types**
   - **a.** MCQs: Basic recall of properties and trends. Provide visual aids for reference.
   - **b.** Short Answer Questions: Describe and explain specific properties. Use guiding questions to structure responses.
   - **c.** Practical Experiment: Compare thermal stability of Na_2CO_3, K_2CO_3, and $CaCO_3$. Provide step-by-step instructions and scaffolded support.
   - **d.** Essay Questions: Analyse and evaluate trends and properties. Allow choice of format (written, presentation, poster).
2. **Delivery**
   - **a.** MCQs and Short Answer Questions: Administer in class or as homework.
   - **b.** Practical Experiment: Conduct in the laboratory with hands-on participation.
   - **c.** Essay Questions: Assign as a take-home project with format choice.

WEEK 15

Learning Indicator: *Describe the patterns in physical and chemical properties of the period 3 elements*

FOCAL AREA: PHYSICAL AND CHEMICAL PROPERTIES OF PERIOD 3 ELEMENTS

The Modern Periodic Law states that the physical and chemical properties of elements are periodic functions of their atomic numbers. This means that elements with similar electronic configurations exhibit similar properties, such as chemical properties (e.g., reactivity, valence), physical properties (e.g., melting, boiling points and density) and chemical bonding (e.g., ionic, covalent and metallic).

Understanding the relationships between electrons, atomic numbers and molecular structure helps us predict and explain the trends in chemical properties (e.g., why certain elements are more reactive than others) and physical properties (e.g., why certain elements have higher melting points than others).

Period three (3) elements refer to elements that are found in the third row (Period) of the periodic table. They are called Period three elements because their outer electrons are found in the third energy level. There are a total of eight elements in Period 3, which include: sodium (Na), magnesium (Mg), aluminum (Al), silicon (Si), phosphorus (P), sulphur (S), chlorine (Cl), argon (Ar).

Physical properties of period three elements

1. **Metallic Character:** This is a measure of how easily an atom loses its electrons. Metallic character generally decreases from left to right across the period. This is because across the period electron affinity increases. Elements have a greater tendency to gain electrons as you move across the period leading to a decrease in metallic character. Furthermore, across the same period, ionisation energy increases making it harder for the elements to lose electrons and exhibit metallic character.

   In addition, across the period reactivity decreases making elements less likely to lose electrons as you move from left to right, leading to a decrease in metallic character.

   Sodium (Na) and magnesium (Mg) and aluminium (Al) are typical metals, exhibiting high metallic character. Silicon (Si) is a metalloid, exhibiting some metallic properties, but also showing significant nonmetallic character. Phosphorus (P), sulphur (S), chlorine (Cl), and argon (Ar) are nonmetals, exhibiting low metallic character.

2. **Melting and Boiling Points:** Melting and boiling points of Period 3 elements increases from left to right across the period from sodium (Na) to silicon (Si) and decreases from silicon (Si) to argon (Ar). The elements in this range (Na, Mg, Al) are metals, forming metallic bonds. As you move from left to right, the number of electrons in the outermost energy level increases, allowing for stronger metallic bonds. Silicon (Si) is a metalloid and has a complete outer energy level, forming a strong covalent bond similar to carbon in diamond. Stronger bonds require more energy to break, resulting in higher melting and boiling points from sodium to silicon.

   From phosphorus to argon of the period are nonmetals existing as simple molecules with Van der Waals forces between them. Argon, however, is monatomic, existing as separate individual atoms with Van der Waals forces between them. Van der Waals forces are very weak intermolecular forces of attraction. Van der Waals forces require less energy to break, resulting in lower melting and boiling points. These forces increase with increasing molecular size and decrease with decreasing molecular size. The melting points and boiling points follow this order:

   $S_8 > P_4 > Cl_2 > Ar$

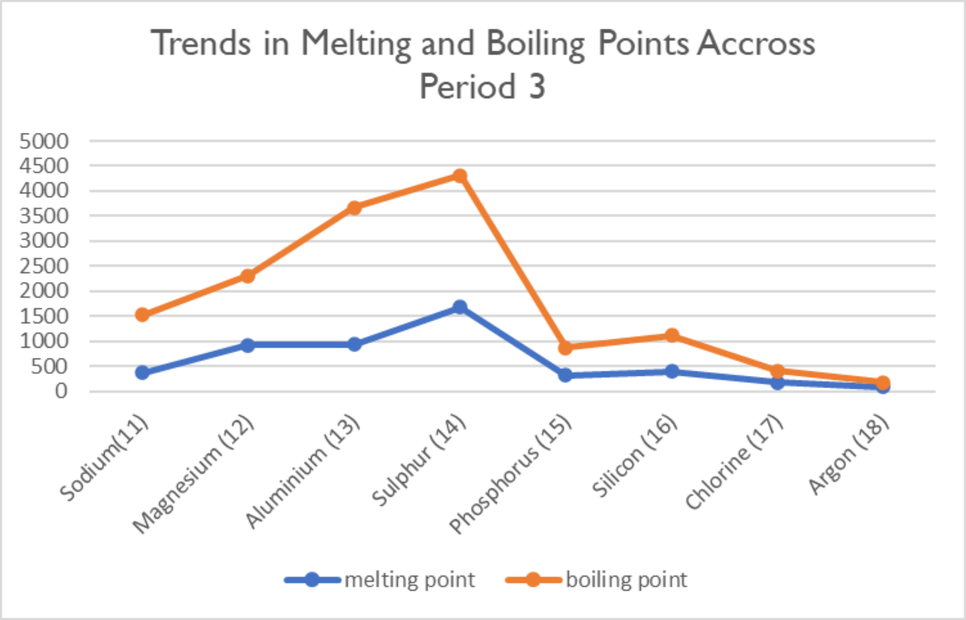


Figure 5.1: *Trends in Boiling and Melting Points Across Period 3*

3. **Density:** Density increases from left to right, with a maximum at aluminium (Al), and then decreases. In other words, the density of Period 3 elements increases from sodium (Na) to aluminium (Al), with aluminium being the highest, and then begins to decrease from aluminium (Al) to argon (Ar). This trend is due to the increase in nuclear charge and the decrease in atomic radius as you move across the period, resulting in a more compact and dense atomic structure. However, the decrease in density after aluminium is due to the increase in atomic radius and the decrease in nuclear charge, leading to a less dense atomic structure.

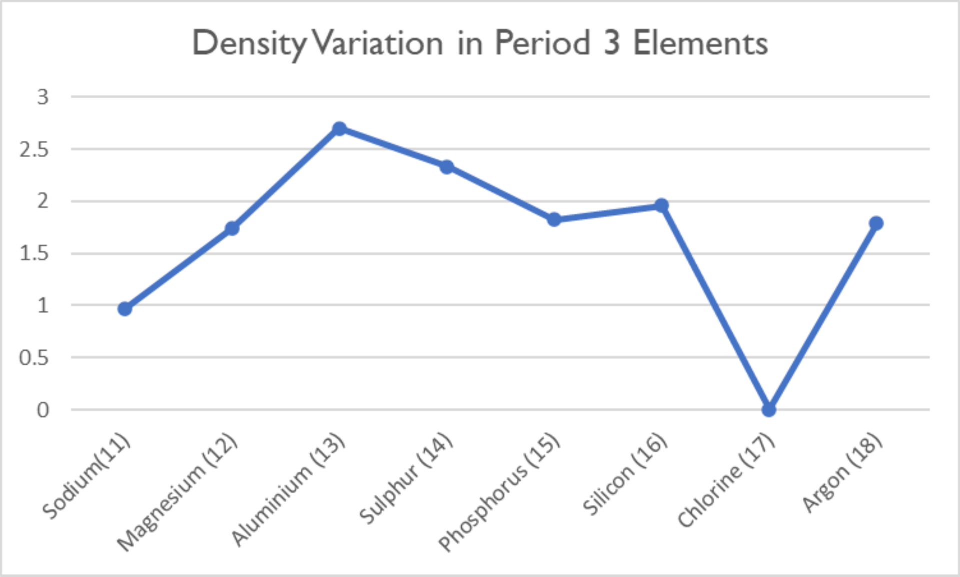


Figure 5.2: *Density Variation in Period 3 Elements*

4. **Electrical Conductivity:** Electrical conductivity is the ability of a substance to conduct electricity. It is measured by the amount of electric current that flows through a material when a voltage is applied.

In general, the electrical conductivity of Period 3 elements decreases from left to right across the period. This trend is related to the elements' electron configurations and the strength of their bonds. As you move from left to right across the period, the elements' valence electrons become more tightly bound, reducing their ability to conduct electricity.

Sodium (Na), magnesium (Mg), and aluminium (Al) are good conductors of heat and electricity. They have a high number of free electrons that can move freely, carrying electrical charge. They are often used in electrical wiring and circuits. Silicon (Si) is a semiconductor. It has a limited number of free electrons, but can be modified to increase conductivity. It is used in electronic devices such as computers and smartphones. Phosphorus (P), sulphur (S), chlorine (Cl), and argon (Ar) are poor conductors of electricity (insulators). They have a low number of free electrons, making it difficult for electricity to flow. They are often used in electrical insulation and protective gear.

Table 5.1: *A Comparison of Physical Properties Period 3 Elements at a glance*

Element	Atomic radius (picometer)	Phase at room temperature	Melting point (K)	Boiling point (K)	Density at STP (gcm^{-3})
$^{23.0}_{11}Na$	186	Solid	371	1156	0.968
$^{24.3}_{12}Mg$	160	Solid	922	1380	1.738

Element	Atomic radius (picometer)	Phase at room temperature	Melting point (K)	Boiling point (K)	Density at STP (gcm^{-3})
$^{27.0}_{13}Al$	143	Solid	933	2740	2.7
$^{28.1}_{14}Si$	137	Solid	1683	2628	2.33
$^{31.0}_{15}P$	128	Solid	317	553	1.823
$^{32.1}_{16}S$	123	Solid	392	718	1.96
$^{35.5}_{17}Cl$	119	Gas	172	238	0.003214
$^{39.9}_{18}Ar$	188	Gas	84	87	1.789

Chemical properties of compound of Period 3 elements

Across the period from left to right, elements move from highly electropositive metals through semi-metal or metalloids with giant structures to simple molecular non-metals. Reactivity of elements generally decreases across the period from left to right as elements becomes less reactive across the period.

Reaction with water

Only sodium reacts with water at room temperature. The remaining elements of the period are generally unreactive towards pure water at room temperature and normal atmospheric pressure.

a. Sodium reacts violently with cold water to form a very strong alkaline compound, sodium chloride and hydrogen gas:

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

b. Magnesium reacts very slowly with cold water to produce sparingly soluble magnesium hydroxide and hydrogen gas but reacts vigorously with water at higher temperatures (steam) to produce magnesium oxide and hydrogen gas:

$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(s) + H_2(g)$

$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$

Reaction with Oxygen: Apart from argon which does not react with oxygen due to stable electron configuration, the rest of the period 3 elements reacts with oxygen at different conditions:

a. $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

b. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

c. $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

d. $Si(s) + O_2(g) \rightarrow SiO_2(s)$

e. $P_4(g) + 5O_2(g) \rightarrow P_4O_{10}(g)$

f. $S_8(g) + O_2(g) \rightarrow 8SO_2(g)$

Reactions of Period 3 Elements with Chlorine: Chlorine (Cl_2) is a highly reactive gas. Reactions with chlorine are important in understanding the properties of Period 3 elements. Reactivity of period 3 elements generally decrease across the period with chlorine being an exception. Sodium is the most reactive with argon being completely inert due to its completely filled outer electron configuration.

a. $2Na + Cl_2 \rightarrow 2NaCl$

b. $Mg + Cl_2 \rightarrow MgCl_2$

c. $2Al + 3Cl_2 \rightarrow 2AlCl_3$

d. $Si + 2Cl_2 \rightarrow SiCl_4$

e. $P + 3Cl_2 \rightarrow PCl_3$

f. $S + Cl_2 \rightarrow SCl_2$

Learning Tasks

1. Create a table with the following columns: Element, atomic number, density, melting point, boiling point, atomic radius and phase at room temperature. Ask learners to fill in the table with the corresponding information for each Period 3 element.
2. Provide learners with a graph showing the trends in physical properties (density, melting point, boiling point) across Period 3 elements. Ask them to analyse the graph and write a short essay (150-200 words) explaining the trends, patterns, and any correlations they observe.
3. Assign learners a real-world scenario or application involving Period 3 elements (e.g., developing a new alloy, designing a chemical reaction). Ask them to research and propose a solution, taking into account the properties and trends of the elements involved.

PEDAGOGICAL EXEMPLARS

1. **Talk-for- learning**
    - **a.** Introduce the topic, ask learners to share their prior knowledge, and write key terms on the board.

- **b.** Explain what metallic property is, how it is measured, and use a chart to show the trend across Period 3 elements. Discuss how the property decreases from left to right.
- **c.** Define physical properties, explain how they are measured, and use a chart to show trends in density, melting and boiling points, and hardness. Discuss the changes across the period.
- **d.** Define chemical properties, explain how they are measured, and briefly describe the reactivity of Period 3 elements with water, acids, alkalis, air/oxygen, and chlorine. Discuss how the reactions change across the period.
- **e.** Divide learners into mixed-ability groups and assign each group a Period 3 element. Ask them to research and discuss its properties using books or the internet.
- **f.** Summarise the trends, ask learners to reflect on their learning, and provide additional resources for further exploration.
- **g.** Provide tailored materials for groups that need extra support, such as visual aids or simplified texts and provide additional challenges or questions for advanced groups, such as researching real-world applications or comparing properties across elements.

2. Problem-based learning

- **a.** Ask learners to use the set of data on physical properties (e.g., atomic radius, melting point, boiling point, density) of Period 3 elements which they have created in **task one** to plot a graph of those data against the atomic numbers of the elements and visualise the trends.
- **b.** Assist learners who may be struggling in getting appropriate scales and labels and make allowance for learners who may want to use digital tools in plotting.
- **c.** Have learners explore the data and graph to identify patterns and relationships.
- **d.** Ask open-ended questions like:
  - i. What do you notice about the graph?
  - ii. How do the physical properties change across the period?
  - iii. What might explain these trends?
- **e.** Provide resources and guidance as needed to help learners understand the physical properties and their units.
- **f.** Identify and explain any outliers or anomalies.
- **g.** Facilitate Analysis and Interpretation
- **h.** Have learners analyse the graph to identify trends and patterns.

i. Ask questions like:
  i. How does the atomic radius change across the period?
  ii. What is the relationship between melting point and atomic number?
  iii. How do the trends relate to the elements' positions in the period?

j. Evaluate learners understanding through their graphs, analysis, and reflections.

3. Digital learning

a. Divide learners into small groups, assign each a Period 3 element, and ask them to research and explore its reactions using smartphones or tablets.

b. Provide specific questions or topics (e.g., sodium with water, magnesium combustion) and tailored support to groups based on their ability level.

c. Suggest credible websites, such as Britannica, Chemistry LibreTexts, or ScienceDirect, to ensure accurate information.

d. Encourage learners to take notes and organise information as they research.

e. After researching, ask each group to create a multimedia presentation or presentation in any format that showcases their findings.

f. Encourage active listening and peer feedback.

g. Facilitate class discussions to assess learners' knowledge and understanding.

KEY ASSESSMENTS

Level 1. What happens to the melting points of the Period 3 elements as you move from left to right?

Level 2: Describe the trend in reactivity of the Period 3 elements. How does it relate to their atomic structure?

Level 3

1. Explain why the physical and chemical properties of Period 3 elements, such as atomic radius, electronegativity, and reactivity, show periodic trends. Use examples from the period to support your answer.
2. What happens when each of the following compounds is heated strongly? Support your answer with appropriate chemical equation(s):
   a) Lithium trioxonitrate (V)- (LiNO3)

Hint

*The recommended mode of assessment for week 15 is **questioning**. Use the level 3 question 2 as a sample question. Refer to the Teacher Assessment Manual and Toolkit page 127 for additional information on this mode of assessment.*

WEEK 16

Learning Indicator: *Describe the patterns in physical and chemical properties of compound (hydrides, oxides, hydroxides and chlorides) of Period 3 elements*

FOCAL AREA 1: PHYSICAL AND CHEMICAL PROPERTIES OF COMPOUNDS OF PERIOD 3 ELEMENTS

Physical and chemical properties of Hydrides of Period 3

Period 3 hydrides are compounds formed between hydrogen and the elements in the third period of the periodic table. Examples include: sodium hydride (NaH), magnesium hydride (MgH_2), aluminium hydride (AlH_3), silicon hydride (SiH_4), phosphorus hydride (PH_3), sulphur hydride (H_2S) and chlorine hydride (HCl).

Physical Properties

1. Bonding and Structure

   NaH and MgH_2: These are ionic hydrides, where hydrogen forms H^- ions. They have high melting and boiling points due to strong ionic bonds.

   AlH_3: This is a covalent hydride with polymeric structure, showing characteristics between ionic and covalent bonding.

   SiH_4, PH_3, H_2S, and HCl: These are covalent hydrides with molecular structures. They have lower melting and boiling points compared to ionic hydrides.

2. Melting and Boiling Points

   Ionic hydrides such as NaH and MgH_2 have high melting and boiling points due to strong ionic bonds.

   Covalent hydrides such as SiH_4, PH_3, H_2S, HCl have lower melting and boiling points compare to the ionic hydrides because of weaker Van der Waals forces between molecules.

3. Solubility in Water

   Ionic hydrides, NaH and MgH_2 react with water to produce hydrogen gas and corresponding hydroxides.

   AlH aluminium $_3$ reacts with water to form aluminum hydroxide and hydrogen gas. SiH_4, PH_3, H_2S, and HCl are generally soluble in water, forming acidic or basic solutions depending on the hydride.

4. Thermal Stability

   NaH and MgH_2 are thermally stable and decompose at high temperatures.

AlH_3: Less thermally stable compared to NaH and MgH_2.

SiH_4, PH_3, H_2S, and HCl decompose at relatively lower temperatures.

Summary of physical properties of Period 3 hydrides

These trends as seen above reflect the changes in bonding and structure as you move across Period 3 from metals to non-metals. The properties shift from those typical of ionic compounds (high melting points, high reactivity with water) to those of covalent compounds (lower melting points, varied reactivity).

Chemical Properties

1. **Reactivity:** Increases from NaH to AlH_3, with AlH_3 being highly reactive and pyrophoric (potential to spontaneously combust upon exposure to oxygen or air).
2. **Acid-Base character:** Becomes more acidic from NaH to AlH_3, with AlH_3 being a strong acid.
3. **Redox reactions:** Hydrides become more reducing from NaH to AlH_3.

Physical and chemical properties of oxides of Period 3

Period 3 oxides are compounds formed between oxygen and the elements in the third period of the periodic table. The oxides of period 3 elements include: Na_2O, MgO, Al_2O_3, SiO_2, P_4O_{10}, SO_3, Cl_2O_7, P_4O_6, SO_2 and Cl_2O.

Physical properties

1. **Structure:** Transition from ionic (e.g., Na_2O, MgO) to giant covalent (e.g., SiO_2) to molecular oxides (e.g., P_4O_{10}, SO_3).
2. **Melting and Boiling Points:** High for ionic and giant covalent oxides due to strong bonds; lower for molecular oxides due to weaker Van der Waals forces.
3. **Electrical Conductivity:** None of these oxides conduct electricity in solid form. However, ionic oxides can conduct electricity when molten due to ion movement.

Types of Oxides of period 3 elements

1. Ionic oxides: Examples: sodium oxide (Na_2O), magnesium oxide (MgO).
2. Bonding: These oxides have strong ionic bonds between metal cations (Na^+, Mg^{2+}) and oxide anions (O^{2-}).
3. Melting and Boiling Points: High, due to the strong electrostatic forces of attraction between the oppositely charged ions. A significant amount of energy is required to break these bonds.
   a. Giant Covalent oxides: Example: silicon dioxide (SiO_2).Bonding: SiO_2 has a giant covalent structure where each silicon atom is covalently bonded to four oxygen atoms in a tetrahedral arrangement, forming a continuous network.

b. Melting and Boiling Points: Very high, because breaking the extensive network of strong covalent bonds requires a lot of energy.

c. Molecular Oxides: Examples: phosphorus pentoxide (P_4O_{10}), sulphur trioxide (SO_3).

d. Bonding: These oxides consist of discrete molecules held together by weaker Van der Waals forces (intermolecular forces).

e. Melting and Boiling Points: Lower compared to ionic and giant covalent oxides because the intermolecular forces are much weaker than ionic or covalent bonds. Less energy is needed to overcome these forces.

  i. Electrical Conductivity: Ionic oxides conduct electricity when molten; covalent and molecular oxides do not conduct electricity.

Chemical properties of Period 3 oxides

Acid-Base Character: Basic (e.g., Na_2O, MgO), amphoteric (e.g., Al_2O_3), and acidic (e.g., SiO_2, P_4O_{10}, SO_3).

Reactivity with Water: Basic oxides form hydroxides (e.g., $Na_2O + H_2O \rightarrow 2NaOH$); acidic xides form acids (e.g., $SO_3 + H_2O \rightarrow H_2SO_4$).

Physical and chemical properties of Period 3 hydroxides

The hydroxides of Period 3 elements are compounds formed between elements of period 3 and hydroxide ions (OH^-). Hydroxides of period 3 include: NaOH, $Mg(OH)_2$, $Al(OH)_3$, $Si(OH)_4$, $P(OH)_3$, H_2SO_4, HClO.

Chemical properties of Period 3 hydroxides

1. **Solubility in Water:** NaOH is highly soluble in water, forming a strong alkaline solution. This high solubility is due to the ionic nature of NaOH, which dissociates completely in water. $Mg(OH)_2$ is sparingly soluble, forming a weakly alkaline solution. Its lower solubility is due to the stronger lattice energy of the ionic compound, which is not completely overcome by hydration energy. $Al(OH)_3$ is insoluble in water but dissolves in both acids and bases due to its amphoteric nature. This means it can form soluble complexes with both H^+ and OH^- ions. $Si(OH)_4$ and $P(OH)_3$ are generally insoluble in water because they form covalent bonds, which do not dissociate easily in water.

2. **Thermal Stability:** The thermal stability of hydroxides generally increases across the period. This is because the lattice energy of the hydroxides increases, making them more stable at higher temperatures. NaOH and $Mg(OH)_2$ decompose at relatively lower temperatures compared to $Al(OH)_3$ and $Si(OH)_4$. This is due to the stronger ionic bonds in NaOH and $Mg(OH)_2$, which require less energy to break compared to the covalent bonds in $Al(OH)_3$ and $Si(OH)_4$.

3. **Melting and Boiling Points:** NaOH has a high melting point due to its ionic nature, which requires a lot of energy to break the ionic bonds. $Mg(OH)_2$ also has a high melting point but lower than NaOH due to the presence of weaker ionic bonds. $Al(OH)_3$ has a lower melting point compared to NaOH and $Mg(OH)_2$ because it has both ionic and covalent character. $Si(OH)_4$ and $P(OH)_3$ have much lower melting points as they are molecular compounds with weaker intermolecular forces.
4. **Electrical Conductivity:** NaOH and $Mg(OH)_2$ conduct electricity in aqueous solution due to the presence of free ions. $Al(OH)_3$, $Si(OH)_4$, and $P(OH)_3$ do not conduct electricity in solid or aqueous forms because they do not have free ions in solution.

Summary of physical properties of hydroxides of Period 3

These trends as seen above reflect the changes in bonding and structure as you move across Period 3 from metals to non-metals. The properties shift from those typical of ionic compounds (high solubility, high melting points, good conductivity) to those of covalent compounds (low solubility, lower melting points, poor conductivity).

Chemical properties of hydroxides of Period 3

1. Acid-Base behaviour

Metals hydroxides: They are basic in nature with $Al(OH)_3$ exhibiting both acidic and basic properties. Example:

a. NaOH is a strong base and dissociates completely in water to form hydroxide ions (OH^-). Example: $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$. It reacts with acids to form salts and water. Example: $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

b. $Mg(OH)_2$ is a weak base and only partially dissociates in water. Example: $Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$. It reacts with acids to form salts and water. Example: $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$

c. $Al(OH)_3$ can act as both an acid and a base. Reacts with acids to form salts and water. Example: $Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$. Reacts with bases to form complex ions. Example: $Al(OH)_3(s) + NaOH(aq) \rightarrow NaAl(OH)_4(aq)$

Non-metal hydroxides: They are generally acidic. Example:

i. H_2SO_4 completely dissociates in water. $H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$. Reacts with bases to form salt and water. Example: $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

ii. HClO is a weak Acid and partially dissociates in water. Example: $HClO(aq) \rightleftharpoons H^+(aq) + ClO^-(aq)$. Reacts with Bases to form salts and water. Example: $HClO(aq) + NaOH(aq) \rightarrow NaClO(aq) + H_2O(l)$

2. **Reactivity:** Reactivity varies widely among the hydroxides. NaOH and $Mg(OH)_2$ are Basic and react with acids to form salts and water. $Al(OH)_3$ is amphoteric, reacting with both acids and bases. Non-metal hydroxides (H_2SO_4 and HClO) are acidic and reacts with bases to form salts and water.

Patterns in physical and chemical properties of Period 3 chlorides

The chlorides of Period 3 elements are compounds formed between these elements and chlorine. They include: NaCl, $MgCl_2$, $AlCl_3$, $SiCl_4$, PCl_5 and S_2Cl_2

Table 5.2: *Chlorides of period 3*

Compound	Name	Type of Chloride
NaCl	Sodium chloride	Ionic chlorid
$MgCl_2$	Magnesium chloride	Ionic chloride
$AlCl_3$	Aluminum chloride	Covalent chlorid
$SiCl_4$	Silicon tetrachloride	Covalent chloride
PCl_5	Phosphorus pentachloride	Covalent chloride
S_2Cl_2	Disulphur dichloride	Covalent chloride

Physical Properties of Period 3 chlorides

1. **Structure:** Ionic chlorides such sodium chloride (NaCl) and magnesium chloride ($MgCl_2$) have giant ionic lattice structures. Covalent chlorides such as aluminium chloride ($AlCl_3$), silicon tetrachloride ($SiCl_4$), phosphorus pentachloride (PCl_5), and sulphur dichloride (S_2Cl_2) all have covalent structures. $AlCl_3$ can exist as a di-mer (Al_2Cl_6) in the solid state.
2. **Melting and Boiling Points:** Ionic chlorides like NaCl and $MgCl_2$ have high melting and boiling points due to strong ionic bonds whist covalent Chlorides like $SiCl_4$, PCl_5, and S_2Cl_2 have lower melting and boiling points due to weaker Van der Waals forces.
3. **Electrical Conductivity:** Ionic chlorides such as NaCl and $MgCl_2$ conduct electricity when molten due to the movement of ions. Covalent chlorides such as $AlCl_3$, $SiCl_4$, PCl_5, and S_2Cl_2 do not conduct electricity in solid or molten states because they lack free ions or mobile electrons.

Chemical Properties Period 3 chlorides

1. **Reactivity with Water**
   a. Ionic chlorides such as NaCl and $MgCl_2$ dissolve in water to form hydrated ions. Example: $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$.

- **b.** Covalent chlorides: React with water, often hydrolysing to form acids and other products.
    - i. Aluminium chloride ($AlCl_3$) hydrolyses to form hydrochloric acid (HCl) and Aluminium hydroxide. Example: $AlCl_3(s) + 3H_2O(l) \rightarrow Al(OH)_3(s) + 3HCl(aq)$.
    - ii. Silicon Tetrachloride ($SiCl_4$): Hydrolyses to form silicon dioxide (SiO_2) and hydrochloric acid. Example: $SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl(g)$
    - iii. Phosphorus pentachloride (PCl_5): Hydrolyses to form phosphoric acid (H_3PO_4) and hydrochloric acid. Example: $PCl_5(s) + 4H_2O(l) \rightarrow H_3PO_4(aq) + 5HCl(g)$

2. **Acid-Base behaviour**
    - **a.** Ionic chlorides: Generally neutral in aqueous solution.
    - **b.** Covalent chlorides: Tend to form acidic solutions upon hydrolysis. Example: $SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl(g)$

Thermal stability of carbonates and nitrates of period 2 and 3 elements

1. **Carbonates (CO_3^{2-}):** Thermal decomposition of carbonates can be represented generally as: $MCO_3(s) \rightarrow MO(s) + CO_2(g)$, where M represents a metal cation.
    - **a.** Period 2 Carbonates:
        - i. Lithium carbonate (Li_2CO_3): Decomposes upon heating to form lithium oxide (Li_2O) and carbon dioxide (CO_2). $Li_2CO_3(s) \rightarrow Li_2O(s) + CO_2(g)$.

          Lithium carbonate is less thermally stable compared to carbonates of heavier alkali metals.
    - **b.** Period 3 Carbonates
        - i. Potassium carbonate (K_2CO_3): Does not decompose easily upon heating.
        - ii. Calcium carbonate ($CaCO_3$): Decomposes upon heating to form calcium oxide (CaO) and carbon dioxide (CO_2). $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$.

Thermal stability increases down the group. Potassium carbonate is more thermally stable than lithium carbonate, the reason can be attributed to polarizability.Smaller cations (like Li^+) have higher charge density and polarise the carbonate ion more, making it less stable and easier to decompose. Larger cations (like K^+) have lower charge density and polarise the carbonate ion less, making it more stable.

2. **Nitrates (NO_3^-):** Thermal Decomposition of nitrates can be represented generally by $2M(NO_3)_2(s) \rightarrow 2MO(s) + 4NO_2(g) + O_2(g)$, where M represents a metal cation.
   - **a.** Period 2 Nitrates
     - i. Lithium nitrate ($LiNO_3$): Decomposes upon heating to form lithium oxide (Li_2O), nitrogen dioxide (NO_2), and oxygen (O_2).

       $4LiNO_3(s) \rightarrow 2Li_2O(s) + 4NO_2(g) + O_2(g)$

       Lithium nitrate is less thermally stable compared to nitrates of heavier alkali metals.
   - **b.** Period 3 nitrates
     - i. Potassium nitrate (KNO_3): Decomposes upon heating to form potassium nitrite (KNO_2) and oxygen (O_2).

       $2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$
     - ii. Calcium nitrate ($Ca(NO_3)_2$): Decomposes upon heating to form calcium oxide (CaO), nitrogen dioxide (NO_2), and oxygen (O_2).

       $2Ca(NO_3)_2(s) \rightarrow 2CaO(s) + 4NO_2(g) + O_2(g)$

       Thermal stability increases down the group. Potassium nitrate is more thermally stable than lithium nitrate.

       Reason: Similar to carbonates, smaller cations (like Li^+) have higher charge density and polarise the nitrate ion more, making it less stable and easier to decompose. Larger cations (like K^+) have lower charge density and polarise the nitrate ion less, making it more stable.

Learning Tasks

1. List the oxides of the Period 3 elements and classify each oxide as acidic, basic, or amphoteric.
2. Write balanced chemical equations for the reactions of sodium oxide (Na_2O) and sulphur dioxide (SO_2) with water.
3. Explain the differences in the chemical behavior of these oxides based on their bonding and structure.

PEDAGOGICAL EXEMPLARS

1. **Talk-for-learning approach**
   - **a.** Briefly explain Period 3 elements and their compounds (hydrides, oxides, hydroxides, chlorides) using visual aids and verbal explanations to cater to different learning styles.

b. Have learners individually consider the properties of each compound type, providing additional resources (handouts, interactive simulations) for those who need more support.

c. Learners discuss their thoughts with a partner, allowing them to express their ideas in different ways (written notes, drawings, verbal explanations).

d. Pairs share their ideas with the class, using guiding questions to help struggling learners and challenge advanced learners with deeper questions.

e. Cover structure and bonding (ionic and covalent), acid-base characteristics (basic and neutral/slightly acidic), and reactions with water and air, providing scaffolded support (equation templates, step-by-step guides) for learners who need it.

f. Discuss their structure, bonding, acid-base characteristics, hydrolytic behavior, and reactions, using a mix of visual aids, verbal explanations, and hands-on activities to cater to different learning preferences.

g. Summarise key points and use quizzes, exit tickets, and quick checks for understanding to gauge learners' progress, adjusting instruction as needed to ensure all learners are supported.

2. Collaborative Learning

a. Briefly explain carbonates and nitrates, and their thermal stability. Use a mix of visual aids (diagrams, videos) and verbal explanations to cater to different learning styles.

b. Divide the class into small groups. Group learners by mixed abilities to encourage peer learning and support.

c. Explain ionic size, charge density, and polarizing power. Provide additional resources (handouts, interactive simulations) for learners who need more support.

d. learners individually think, discuss in pairs, and share with the group. Allow learners to express their thoughts in different ways (written notes, drawings, verbal explanations).

e. Discuss trends in thermal stability and compare Li, K, and Ca compounds. Use guiding questions to help struggling learners and challenge advanced learners with deeper questions.

f. Guide learners to write decomposition reactions for the compounds. Provide scaffolded support (equation templates, step-by-step guides) for students who need it.

g. Each group presents their findings and equations. Allow learners to choose their presentation format (poster, slide show, oral presentation).

- **h.** Clarify concepts and assess understanding through discussion and quizzes. Use a variety of assessment methods (quizzes, exit tickets, group discussions) to cater for different strengths and preferences

3. Experiential learning

- **a.** Explain the purpose of the experiment and the concept of thermal stability using visual aids and verbal explanations to cater to different learning styles.
- **b.** Review safety procedures, ensuring all students understand and follow them.
- **c.** Prepare materials and set up the apparatus, providing additional support for students who need it.
- **d.** Heat Na_2CO_3, K_2CO_3, and $CaCO_3$, observe changes, and test for CO_2 gas, allowing hands-on participation for kinesthetic learners.
- **e.** Guide learners to write balanced equations for the decomposition reactions, offering scaffolded support for those who need it.
- **f.** Analyse results and discuss differences in thermal stability, using guiding questions to help struggling learners and challenge advanced learners.
- **g.** Use diagrams and videos for visual learners, hands-on activities for kinesthetic learners, and additional questions for advanced learners.
- **h.** Summarise key points and assess understanding through quizzes and discussions, adjusting instruction as needed to support all learners.

KEY ASSESSMENTS

Level 1: List the Period 3 elements and their corresponding hydrides. (Provide a periodic table for reference)

Level 2: Compare the thermal stability of Na_2CO_3, K_2CO_3, and $CaCO_3$.(Provide a graphic organiser to help students organise their thoughts).

Level 3

1. Explain why the thermal stability of carbonates increases down the group in Period 3. (Allow students to use diagrams or written explanations based on their preference).
2. Use computational tools to compare and rank the reactivity of Fluorine(F_2), Chlorine (Cl_2) and Bromine (Br_2)

Hint

The recommended mode of assessment for week 16 is ***computational****. Use the level 3 question 2 as a sample question.*

SECTION 6: PHYSICAL AND CHEMICAL PROPERTIES OF THE HALOGENS

Strand: Systematic Chemistry of the Elements

Sub-Strand: Periodicity

Learning Outcome: *Predict and describe the physical and chemical properties of the halogens as well as their uses in everyday life.*

Content Standard: Demonstrate knowledge and understanding of the pattern of the physical and chemical properties of the halogens as well as their uses.

Hint

Remind learners of the Mid-Semester Examination in Week 18. Refer to Appendix F at the end of this section for a Table of Specification to guide you to set the questions for the exams.

INTRODUCTION AND SECTION SUMMARY

In weeks 17 and 18 of the Systematic Chemistry of the Elements strand, students will explore the periodicity of chemical elements, focusing on halogens (Group 17). In week 17, learners will examine the physical and chemical properties of halogens, emphasising their unique characteristics, reactivity, and trends within the group. This foundational knowledge will set the stage for week 18, where students will describe the reactions of halide salts and differentiate the acid strengths of hydrogen halides, underlining the significance of these acids in various applications. Furthermore, learners will investigate the practical uses of halogens in industry and daily life. Throughout these lessons, the integration of 21st-century skills will be emphasised, promoting digital learning and ensuring that perspectives of gender and social inclusion are acknowledged to foster a comprehensive understanding of chemistry's role in society.

The weeks covered by the section are:

Week 17: *Physical and Chemical properties of halogens*

Week 18*: Reaction of halide salts and uses of halogens*

SUMMARY OF PEDAGOGICAL EXEMPLARS

In delivering the lessons on halogens, a variety of pedagogical strategies will be employed to enhance student engagement and understanding. In Week 17, the talk for learning approach will facilitate discussions on the variation in physical properties of halogens and collaborative learning techniques will encourage learners to explore electron configurations and variable oxidation states within groups. To enrich the content, digital learning will be utilised, allowing learners to watch informative videos on halogen reactions with water and alkalis, followed by guided discussions on reagents, conditions, and products. In Week 18, the think-pair-share method will further deepen understanding as students explain the reducing power of halides and illustrate reactions with concentrated tetraoxosulphate (VI) acid. Mixed-ability groups will collaboratively discuss relative bond strengths and Ka values of hydrogen halides, as well as the practical applications of halogens in water purification, bleaching, and disinfection, connecting chemistry to everyday life.

ASSESSMENT SUMMARY

The assessment summary for weeks 17 and 18 focuses on evaluating student understanding of halogens through various formative methods. Assessments will include short answers, oral and written presentations, and lab reports, allowing learners to demonstrate their grasp of the physical and chemical properties of halogens and their reactions. Practical assessments will involve manipulation of equipment, safety procedures, and deductions from experimental data. Homework will extend learning and application of concepts. Group projects will encourage collaboration, targeting the DoK framework levels 1 through 3, while fostering critical thinking, communication, and creativity. Class exercises, including measurements and group tasks, will assess comprehension in real-time, supplemented by discussions on acid strengths and practical uses of halogens in everyday life. Records of all assessments will be documented in transcripts to facilitate continuous assessment and guide future instruction.

WEEK 17

Learning Indicator: *Explain the physical and chemical properties of the halogens (Group 17 elements)*

FOCAL AREA: PHYSICAL AND CHEMICAL PROPERTIES OF THE HALOGENS (GROUP 17 ELEMENTS)

Halogens belong to Group 17 (VII A) of the periodic table. It is the only group that exhibit all the three states of matter at room temperature. Fluorine and chlorine exist as gases, bromine is a liquid and iodine exist as a solid. Astatine is radioactive with a very short half-life so very little is known about it. All the halogens exist as non-metals in a diatomic molecule except the radioactive Astatine. Halogens are highly reactive and so are found in nature combined to other elements and never as free elements. They have the general electron configuration of $ns^2 np^5$ where the last principal quantum number (n) is greater than two. They fall short of one electron to attain the noble gas configuration. They are the most reactive of the non-metallic elements. Anions derived from halogens are called halides. They easily gain an electron to form an ionic halide but share an electron to form a covalent compound.

Physical properties of the halogens

1. There is gradual change in the physical state of halogens down the group from gas (fluorine and chlorine) to liquid (bromine) to solid (iodine and astatine). This reflects the nature of intermolecular forces present which also influences colour at room temperature.
2. Halogens exhibit unique colours that vary significantly among the group. Fluorine has a pale yellow colour, chlorine forms a yellow – green colouration in the gaseous form as well as in solutions. Bromine has a reddish-brown appearance while iodine is a dark purple solid. The colour of halogens deepens down the group of the periodic table.
3. The density of halogens increases down the group. Down the group, the atomic mass increases. This increase in mass results from the addition of more protons and neutrons in the nucleus. Flourine and chlorine are denser than air, bromine is denser than water and iodine is a solid with a high density.
4. The halogens are held together by intermolecular force of attraction in the form of Van der Waals forces. The strength of Van der Waals forces increases with increasing molecular masses. Fluorine has the smallest mass as compared to iodine and so the strength of the intermolecular forces in fluorine is less than that in iodine. The lower the intermolecular forces in fluorine keeps it in the gaseous state at room temperature.

5. They are strong oxidising agents. The standard reduction potentials of halogens which indicates their oxidising strength decreases from F_2 to I_2. This means that Cl_2 will oxidise I^- ions to I_2 in aqueous solution.

$(X_2/2X^-)$	Standard Reduction Potential, E°
$F_2/2F^-$	+2.87
$Cl_2/2Cl^-$	+1.36
$Br_2/2Br^-$	+1.09
$I_2/2I^-$	+0.54

6. Halogens in their pure form have lower bond energies compared to many other non-metal diatomic molecules such as O_2 and N_2. $F - F$ has a bond energy of 154 kJ/mol while $O = O$ and $N \equiv N$ have a bond energies of 495 kJ/mol and 941 kJ/mol respectively. This is partly due to the number of paired electrons shared between the elements. Halogens in their pure form share a pair of electrons (i.e. made up of single bonds), two pairs of electrons are shared between two oxygen atoms in oxygen gas (forms a double bond) while three pairs of electrons are shared between two nitrogen atoms in nitrogen gas (forms a triple bond). The higher the number of bonds shared between molecules the greater the bond energy. The presence of lone pairs causes repulsion between atoms in a diatomic molecule. Halogens have relatively small size so they experience a greater repulsive force from the three lone pairs which weakens the bond between them.

7. Generally, halogens have low melting and boiling point

   The primary forces responsible for holding halogen molecules together are Van der Waals forces. These are relatively weak forces that arise from temporary dipoles created when electron density is distributed unequally in a molecule. Compared to other types of intermolecular forces (like hydrogen bonding or dipole-dipole interactions), Van der Waals forces are weaker, which explains why halogens have comparatively low melting and boiling points. However, the strength of Van der Waals forces increases with the size of the molecule (as more electrons are present), leading to higher melting and boiling points.

Halogen	Melting point/ °C	Boiling point/ °C
F_2	-219.6	-188.1
Cl_2	-101.5	-34.0
Br_2	-7.2	58.8
I_2	113.7	184.3

8. Halogens have a strong and often nasty smell

Halogens are known for their strong, pungent odors, which can often be described as unpleasant or "nasty." This characteristic smell is primarily associated with their molecular properties and the nature of the halogen compounds. Halogens are generally volatile substances hence they readily evaporate into the gas phase at room temperature. Compounds like chlorine (Cl_2) and bromine (Br_2) can easily transition from liquid to gas, allowing the odor to be detectable even in small concentrations. The gaseous state of halogens allows their particles to disperse quickly into the air, contributing to a more intense olfactory perception.

9. They are poisonous elements

Down the group of the halogens, interatomic bond energy generally decreases. Atomic size increases down the group which causes the nuclei of atoms to be farther away from the shared electron cloud when these elements are covalently bonded. This reduces the force of attraction which then forms a weaker bond. The energy required to break the weaker covalent bond therefore reduces. Bond dissociation energy is the heat change that occurs when one mole of a covalent bond is broken down into its constituent gaseous atoms whereas bond energy is the standard enthalpy change that occurs when one mole of a covalent bond is formed from its constituent gaseous atoms. Certain anomalies can be observed in the bond dissociation energy of the halogens which arise from factors such as atomic size, lone pair repulsion and the nature of bonding in the molecular form of the elements. Bond dissociation energy of halogens decreases from Cl_2 to I_2 with F_2 having an abnormally low value. Each of the halogens have three lone pairs of electrons but Fluorine has the smallest size among them. The electrons forming the bond are packed close to each other which contributes to a strong lone pair – lone pair repulsion that reduces the covalent bond attraction between the nuclei and the shared electrons. This weakens the $F-F$ bond so less energy is required to break the bond. This is in contrast to other halogens where such repulsive forces are less significant due to larger atomic sizes.

Halogen (X − X)	Bond energy/ $kJmol^{-}$
F – F	159
Cl – Cl	243
Br – Br	193
I – I	151

Chemical properties of the halogens

1. Halogens react with metals to form metal halides that are mostly ionic compounds. In the reaction, the metal atom(s) lose their valence electron(s) to form a cation while the halogen atom(s) receives the electron(s) to form an anion. The metallic ions and the halide ions are then attracted to each other due to the opposite charges which results in an ionic bond. These compounds are usually characterised by a crystal lattice structure.

   Eg. $M+X_2 \rightarrow MX$

   $Mg+Cl_2 \rightarrow MgCl_2$

2. They form molecular compounds among themselves (such as ICl and BrF_3) and with other non-metals in other groups (such as NF_3 and PCl_5)

3. React with hydrogen to form hydrogen halides

   $H_2(g)+X_2(g) \rightarrow 2\,HX(g)$

   The high reactivity of fluorine makes its reaction with hydrogen gas very violent (explosive) but this becomes less violent when it's substituted with chlorine, bromine or iodine. Chlorine and hydrogen explode in bright sunlight but react slowly in the dark. Bromine reacts slowly with hydrogen on heating with a platinum catalyst. Iodine combines partially and very slowly with hydrogen even on heating. Hydrogen halides (HX) dissolve in water to form hydrohalic acids. HF ionizes partially in solution so it is a weak acid, but HCl, HBr and HI are all strong acids.

4. Halogens take part in displacement reactions with halide salts. The more reactive halogen will displace a less reactive halide from its compounds.

5. Halogens have very high electronegativity as well as high first electron affinity.

   Halogens have very high electron affinity so form ionic compounds with electropositive elements (such as metals) by gaining an electron. In binary ionic compounds ($MgCl_2$), halogens have an oxidation state of -1. In the pure form (X_2), they exhibit oxidation state of zero (0). However, halogens can also exhibit multiple positive oxidation states due to factors such as lone pair electrons, presence of d-orbitals and stability of oxidation states based on electronegativity. Fluorine belongs to the second shell which has only two orbitals (s and p orbitals) and so does not have any low energy d orbital which lone pairs can be promoted to when it becomes excited. As such does not commonly display positive oxidation states. Heavier halogens have empty d-orbitals in which electrons can be accommodated when the atoms become excited. Chlorine can form an oxidation state of +1, +3, +5 and +7. Bromine exhibits similar oxidation states though higher oxidation states are typically less stable compared to chlorine. Iodine also shows oxidation states of +1, +3, +5 and +7 but the stability of these states varies based on the specific compound.

Element	-1	0	+1	+3	+4	+5	+6	+7
F	NaF	F_2						
Cl	$MgCl_2$	Cl_2	HClO	$HClO_2$	ClO_2	$HClO_3$		$HClO_4$
Br	$CaBr_2$	Br_2	BrCl	BrF_3	BrO_2	$HBrO_3$	BrO_3	$HBrO_4$
I	KI	I_2	ICl, HIO	ICl_3		HIO_3		HIO_4

Figure 5.3: *Variable oxidation states of halogens and some examples of their compounds*

Halogens are highly electronegative elements. They can accept electrons in their outer shell to become stable and so they are also described as highly oxidising agents. Down the group of the periodic table, electronegativity decreases which also decreases the ability of the halogens to accept electrons and the overall reactivity of halogens. The measure of the ability to accepts electrons from other substances or oxidising strength of a substance is given by the standard electrode potential (E^{θ}). All halogens have positive standard electrode potential values which indicate that they are good oxidising agents. The stronger oxidising agent has a more positive value.

Fluorine is most powerful oxidising agent among the halogens, followed by chlorine then bromine with iodine being the least powerful. The trend of reactivity among the halogens can be described as $F_2 > Cl_2 > Br_2 > I_2$. A more reactive halogen can displace a less reactive halogen from an aqueous halide solution. This sets up a redox reaction in which electrons are transferred between atoms. Halogens that are lower in the group are oxidised by those that are above them on the periodic. For example, chlorine can oxidise bromide ions to bromine and iodide ions to iodine.

$$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$$

$$Cl_2 + 2KI \rightarrow 2KCl + I_2$$

From the standard electrode potential values, bromine can only oxidise iodide ions to iodine. This can be represented by the equation

$$Br_2 + 2KI \rightarrow 2KBr + I_2$$

Halogens can also undergo precipitation when halide ions react with solution containing silver ions. It helps in the identification of unknown samples that contain halide ions due to the characteristic nature of the precipitate that forms. The formation of silver halides that precipitates out of solution depends on the halide present. The general reaction equation can be represented as:

$X^-(aq) + Ag^+(aq) \rightarrow AgX\ (s)$

Reactions with chloride ions:

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

White precipitate of silver chloride is formed when chloride ions combine with silver ions in an aqueous solution. The precipitation of AgCl occurs instantly upon the mixing of the solutions due to the strong tendency of silver ions to bond with the chloride ions. The white mass that is formed is insoluble in water which makes it easily separable from the reaction mixture. It is however soluble in ammonia and alkaline solutions. This process is a classic example of a precipitation reaction which is often utilised in analytical chemistry to detect the presence of chloride ions in solution. The reaction is also used to determine chloride concentration in solutions by measuring the amount of silver nitrate required to precipitate all the chloride ions as silver chloride.

Reaction with bromide ions

$Br^-(aq) + Ag^+(aq) \rightarrow AgBr\ (s)$

Aqueous solution of silver ions reacts with a solution containing bromide ions to form a yellowish-white precipitate of silver bromide. The solid mass formed is insoluble in water and dilute ammonia but soluble in concentrated ammonia.

Reaction with iodide ions

$I^-(aq) + Ag^+(aq) \rightarrow AgI\ (s)$

The reaction of aqueous solutions containing silver ions and iodine ions leads to the formation of a bright yellow precipitate of silver iodide. The precipitate does not dissolve in dilute or concentrated ammonia.

Learning Tasks

1. Present an analysis of the physical states of halogens.
2. Identify and outline the variations in melting and boiling points of halogens. Construct a graph comparing these values and identify any noticeable trends.
3. Create a visual chart or poster that highlights the physical state, melting and boiling points, and bond energies of each halogen, ensuring visual clarity. (Include a discussion on anomalies in bond energies and provide examples).
4. Work together to create a table that compares the electron configurations of halogens, identifying their commonalities and differences. Based on the differences in their electron configurations, explain why some halogens exhibit variable oxidation states and reactivity.

5. Discuss and identify the similarities in the chemical properties of halogens. Consider aspects such as their tendency to form salts with metals, their reactions with hydrogen to form hydrogen halides, and their high reactivity.
6. Prepare a simple table of the standard electrode potential values of the halogens. Use the values to discuss how these values determine the reactivity of halogens in displacement and precipitation reactions. Create a simple flowchart that illustrates which halogen can displace another. Share examples of both reaction types and explain the outcomes based on reactivity.
7. In mixed-ability groups, watch selected videos that demonstrate the reactions of halogens with water and alkalis. After viewing, analyse the reagents, reaction conditions, and products formed in detail. Prepare a comprehensive report that includes the balanced chemical equations for each reaction and an explanation of the significance of these reactions in real-world applications.

PEDAGOGICAL EXEMPLARS

1. **Talk-for-learning approach**
   - **a.** In mixed-ability groups, discuss the variations in the physical properties of the halogens under the following headings:
     - i. Physical State
     - ii. Melting and Boiling points
     - iii. Bond energies
   - **b.** Discuss any anomalies in the bond energies.
2. **Collaborative learning**
   - **a.** In mixed-ability groups, discuss the electron configurations and the exhibition of variable oxidation states of the halogens.
   - **b.** Discuss the similarities in chemical nature of the halogens.
   - **c.** Deduce the differences in the reactivity of the halogens based on their standard electrode potential values.
   - **d.** Use the following reactions for your discussions:
     - i. Halogens with halide ions (Displacement reactions).
     - ii. Halide ions with silver ions (Precipitation reactions).
3. **Digital learning**
   - **a.** In mixed-ability groups, watch videos and discuss the reagents, conditions and products formed when the halogens react with water and alkalis.
   - **b.** Write balanced chemical equations to show how the reactions occur.

KEY ASSESSMENTS

Level 1: Write a balanced chemical equation between aqueous solutions of sodium chloride and silver nitrate and state the nature of the products formed.

Level 2

1. You are given four substances: a fuming red liquid, a dark metallic-looking solid, a pale-yellow gas, and a yellow-green gas that attacks glass. You are told that these substances are the first four members of Group 7A, the halogens. Name each one.
2. Explain the chemical basis for the uses of halogen reactions in everyday life under the following headings:
    a) Purification of water for drinking

Level 3: Consider the halogens chlorine, bromine, and iodine. The melting point and boiling point of chlorine are –101.5°C and –34.0°C and those of iodine are 113.7°C and 184.3°C, respectively. Thus, chlorine is a gas and iodine is a solid under room conditions. Estimate the melting point and boiling point of bromine. Compare your values with those from the webelements.com website.

Hint

*The recommended mode of assessment for week 17 is **homework**. Use the level 2 question 2 as a sample question.*

WEEK 18

Learning Indicator: *Describe the reaction of halide salts and explain the differences in acid strength of the hydrogen halides as well as the uses of the halogens*

FOCAL AREA: REACTION OF HALIDE SALTS, DIFFERENCES IN ACID STRENGTH OF THE HYDROGEN HALIDES AND THE USES OF THE HALOGENS.

Reaction of halide salts

Halogens are generally good oxidising agents and so they get reduced to their anionic forms (halides). The halide ions have the tendency to lose electrons to form the halogen.

$2F^-(aq) \rightarrow F_2(g) + 2e$

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

$2Br^-(aq) \rightarrow Br_2(g) + 2e$

$2I^-(aq) \rightarrow I_2(g) + 2e$

Iodide ions have the highest reducing power, followed by bromide ions. Fluoride ions have very low reducing properties. This is because as one moves down the group, the number of shells in the ion increases, which increases the shielding effect by the core electrons. So the outer electrons become less strongly held. These electrons are thus lost more easily and the halide ion is more readily oxidised.

Reactions of halides salts with concentrated tetraoxosulphate (VI) acid

Halides, which are compounds formed from halogen elements exhibit a variety of chemical behaviours based on their respective reactivity and the nature of the anions. Among the halides, bromide and iodide ions are particularly noted to act as reducing agents when reacting with concentrated tetraoxosulphate (VI) acid. A reducing agent is a substance that donates electrons in a chemical reaction, leading to a decrease in oxidation state of another substance. The reactivity of halides as reducing agents increases down the group in the periodic table. Down the group, ions become larger and the larger the ion, the farther the outer electrons are from the nucleus. It is easier to lose electrons from larger ions due to increased shielding effect. Thus, iodide is the strongest reducing agent among the halides, followed by bromide, while fluoride and chloride show very little reducing behavior. The reactions typically proceed through two steps. Initially forming hydrohalic acids which can further reduce tetraoxosulphate (VI) to Sulphur (IV) oxide depending on the reduction potential.

Reaction of NaF

Sodium fluoride does not exhibit strong reducing properties. It typically reacts to form hydrofluoric acid and sodium hydrogen tetraoxosulphate (VI). The H-F bond in HF is particularly strong, making it less likely for HF to dissociate into H^+ and F^- ions, which is a crucial step for fluoride to act as a reducing agent.

$NaF + H_2SO_4 \rightarrow HF + NaHSO_4$

Reaction of NaCl

Sodium chloride does not show significant reducing behavior.

Forms hydrochloric acid (HCl) and sodium hydrogen sulfate ($NaHSO_4$).

$NaCl + H_2SO_4 \rightarrow HCl + NaHSO_4$

Oxidation of Cl^- to Cl_2 typically requires stronger oxidising agents or specific conditions to facilitate the reaction. Concentrated sulphuric acid, while being a strong acid, is not sufficient to achieve oxidation of Cl^- because it does not have a strong enough electron affinity to oxidise chloride ions effectively.

Reaction of NaBr

When a halide salt like sodium bromide (NaBr) is treated with concentrated sulphuric acid, they engage in a series of oxidation-reduction reactions. Initially the reaction leads to the formation of HBr which serves as a strong reducing agent.

$NaBr + H_2SO_4 \rightarrow HBr + NaHSO_4$

The bromide ion formed in HBr is relatively stable in its -1 oxidation state; however, unlike chloride ion, it has a lower electronegativity and a larger atomic radius. This makes it easier for bromide ions to lose electrons and be oxidised to bromine (Br_2) during reactions with oxidising agents like concentrated tetraoxosulphate (VI) acid. The Bromide, though a good reducing agent can only reduce the Sulphur from +6 to +4.

$H_2SO_4 + 2H^+ + 2Br^- \rightarrow Br_2 + SO_2 + 2H_2O$

Reaction of NaI

Sodium iodide dissociate in solution to produce iodide ions which possess unique chemical properties. It can act as reducing agents due to its size. When Iodide ions are exposed to concentrated tetraoxosulphate (VI) acid, they easily lose electrons and so gets reduced to Iodine. The reaction initially leads to the formation of hydroiodic acid according to the reaction:

$NaI + H_2SO_4 \rightarrow HI + NaHSO_4$

Hydroiodic acid is a strong acid and a highly effective reducing agent. The presence of the iodide ion (I^-) in hydroiodic acid allows it to lose electrons readily. When HI is treated with a strong oxidising agent like concentrated tetraoxosulphate (VI) acid (H_2SO_4), the I^- ions can be oxidised to molecular iodine (I_2) reducing the Sulphur from +6 to +4 (SO_2), to 0 (S) or to -2 (H_2S).

$$H_2SO_4 + 2H^+ + 2I^- \rightarrow I_2 + SO_2 + 2H_2O$$

$$H_2SO_4 + 6H^+ + 6I^- \rightarrow I_2 + S + 4H_2O$$

$$H_2SO_4 + 8H^+ + 8I^- \rightarrow 4I_2 + H_2S + 4H_2O$$

Acid strengths of hydrogen halides

The acid strength of hydrogen halides varies significantly across the group. The hydrogen halides include hydrofluoric acid (HF), hydrochloric acid (HCl), hydrobromic acid (HBr) and hydroiodic acid (HI). Aqueous HF acts as a weak Bronsted – Lowry acid in water whereas the other hydrohalic acids (HCl, HBr and HI) are all strong acids. The acid strength increases in the order HF << HCl < HBr < HI.

a. Relative bond strengths of Hydrogen halides (HX)

The relative bond strength of the H – X bonds play a crucial role in determining the acid strength of hydrogen halides.

Hydrogen halide (H – X)	Bond strength/ *kJmol*$^-$
H – F	565
H – Cl	432
H – Br	366
H – I	298

For a hydrogen halide to act as an acid, it must dissociate in water to release the hydrogen ion (H^+). From the bond strength of the hydrogen halides, HF has a very strong bond which makes it difficult to dissociate into its ions. HCl, HBr, and HI have relatively weaker bonds, which means they easily ionize in solution, contributing to their stronger acidic nature. The weaker the bond strength, the easier it is for the molecule to dissociate into H^+ and X^- ions and the stronger the acid.

b. K_a values of HX

The acid dissociation constant (K_a) is a measure of the strength of an acid in solution. It quantifies the extent to which an acid dissociates into its constituent ions in water. The dissociation reaction for a hydrogen halide can be represented as follows:

$$HX(aq) \rightarrow H^+(aq) + X^-(aq)$$

Hydrogen halide (HX)	Acid dissociation constant (K_a)
H – F	6.6×10^{-4}
H – Cl	1.3×10^{6}
H – Br	5.0×10^{9}
H – I	5.0×10^{10}

The size of the halogen atom increases down the group from fluorine to iodine, which affects the ability to bond with hydrogen. Iodine is significantly larger than fluorine. This increase in atomic size results in longer and thus weaker H-X bonds for the heavier halogens (Cl, Br, I), contributing to their stronger acidic nature.

Also, the acid dissociation constant value is influenced by the electronegativity of the elements. It determines how tightly the electrons in the H – X bond are held. Fluorine is the most electronegative element and so pulls the bonding electrons closer resulting in a strong bond that resists dissociation. Hence contributing to the lower K_a value. Iodine on the other hand is the least electronegative among the halogens and so forms a weaker bond with hydrogen. Therefore, HI dissociate more easily and thus have the highest K_a value.

c. Thermal stability of the halogen halides

The thermal stability of the halogen halides is the ability of the compound to withstand heat without undergoing decomposition. It determines how easy the compound is broken up into its constituent elements when heated. Halogen halides exhibit varying thermal stabilities. This thermal stability is significantly influenced by the bond energies of the H-X bonds, and understanding these relationships is essential in assessing their reactivity and stability under heat. Bond energies decrease from HF and HI, indicating that the strength of the H – X bond weakens down the group in the periodic table. Hydrogen fluoride (HF) has the strongest H-F bond, requiring the most energy to break. As a result, HF is more thermally stable than the other hydrogen halides due to its high bond energy. HF can withstand higher temperatures without dissociating compared to HCl, HBr, or HI. In contrast, hydrogen iodide (HI), which has the weakest bond among the hydrogen halides, is less thermally stable. The weaker the H-I bond, the easier it is for HI to dissociate into H^+ and I^- ions when heated. This makes HI more likely to decompose or react at elevated temperatures.

Everyday uses of halogens

Halogens and their compounds have widespread usage in modern society. Its application ranges from industry, health care and environmental management. They play a crucial role in providing solutions and improving quality of life. Their unique chemical properties enable innovations across diverse sectors in society through technological advancements.

a. Purification of water for drinking

Chlorine is used extensively in water treatment plants to disinfect drinking water, removing harmful microorganisms. Chlorine dissolves in water as described in the chemical equation:

$Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HClO(aq)$

The key component of this reaction that aids in disinfection is the hypochlorous acid (HClO). Hypochlorous acid is highly effective as a disinfectant due to its strong oxidising properties. It acts on bacteria cell membrane, proteins and DNA through oxidative stress. Bacteria are protected by a cell membrane that acts as a barrier, keeping essential components inside. Hypochlorous acid disrupts the integrity of the cell membrane making it leaky which allows essential components (like proteins, nutrients and DNA) to spill out of the organism. The function of these essential components are based on their structures. Due to the oxidising properties of the hypochlorous acid the essential components of the bacteria are oxidised that changes their structure which effectively inactivates them or stops their reproduction. Hypochlorous acid is effective against a wide range of pathogens (including fungi, viruses, gram-positive bacteria and gram-negative bacteria) due to its non-specific mechanism of action. It stability in aqueous solutions make it a preferred choice for disinfection in various settings including water treatment facilities, swimming pools and healthcare environments.

b. You use bleach to oxidise dyes and other coloured molecules.

c. Disinfecting toilets

Halogens, such as chlorine, bromine, and iodine, are widely used as disinfectants due to their effective antimicrobial properties. The use of halogens in disinfecting toilets is primarily based on their ability to kill bacteria, viruses, and other pathogens, thus promoting hygiene and sanitation. Halogens are strong oxidising agents, meaning they readily accept electrons during chemical reactions. This property allows them to react with a wide variety of organic and inorganic substances, including the cellular components of microorganisms. When halogens come into contact with pathogens present in toilets (such as bacteria and viruses), they can oxidise vital cellular components, leading to the destruction of the pathogen.

Halogens, particularly chlorine, can react with the organic matter commonly found in toilets (e.g., fecal material, urine) and disinfect the surfaces. A common disinfectant is sodium hypochlorite (NaClO), a chlorine compound often used in household bleach. When sodium hypochlorite dissolves in water, it forms hypochlorous acid (HOCl), which is the active disinfecting agent. Hypochlorous acid is effective at penetrating and breaking down organic matter and is particularly potent against various bacteria and viruses.

d. Use of halogenated hydrocarbons as solvents, refrigerants and in aerosols.

Halogenated hydrocarbons, which are organic compounds containing carbon and one or more halogen atoms (such as fluorine, chlorine, bromine, or iodine), have a diverse range of applications, including as solvents, refrigerants, and in aerosol propellants.

Halogenated hydrocarbons often exhibit varying degrees of polarity due to the electronegativity of halogen atoms. Halogenated hydrocarbons, such as dichloromethane (DCM) and chloroform, are generally polar solvents that are excellent at dissolving a wide range of organic compounds, including fats, oils, and waxes. The dipole moment generated by the C-X (carbon-halogen) bond allows these solvents to interact favorably with polar solutes, enhancing their solvent capacity.

Certain halogenated hydrocarbons possess low volatility and thermal stability, which are crucial for applications like refrigeration and aerosol formulation. Compounds such as chlorofluorocarbons (CFCs) and hydrofluorocarbons (HFCs) are effective refrigerants due to their low boiling points, which allows them to absorb and release heat efficiently in vapor-compression cycles. Their thermal stability also means they do not readily decompose under the conditions found in refrigeration systems.

Halogenated hydrocarbons are often used as propellants in aerosol products. Their stability and low reactivity ensure they can safely deliver active ingredients without undergoing unwanted reactions during storage or use. For instance, substances like dichlorodifluoromethane (CFC-12) were used due to their stability and ability to maintain pressure in aerosol containers.

Halogenated alkanes often have higher densities than common organic solvents, making them suitable for specific applications where heavier solvents are advantageous. Their low boiling points facilitate easy vaporisation, which is crucial for both refrigeration cycles and aerosol dispersal. NB it should be noted that the majority of these chemicals are now banned for use as refrigerants and aerosols due to the contribution they make to global warming.

Learning Tasks

1. Explain the reducing power of halides using factors such as electronegativity, bond strength, and ion size.
2. With the help chemical equations explain how halogens react with concentrated tetraoxosulphate (VI) acid.
3. Identify and explain the trend in acid strength of hydrogen halides based on bond strength and Ka values.
4. How does bond energies affect the thermal stability of halogen halides
5. State two uses of halogens and explain the chemical basis behind their usage.

PEDAGOGICAL EXEMPLARS

1. Talk-for-learning

- **a.** Ask learners to think about the halides (chloride, bromide, iodide) and their reducing power. Consider the following questions:
- **a.** What is reducing power?
- **b.** How do halides differ in their reducing power?
- **c.** What factors influence reducing power?
- **b.** Pair learners up to discuss their thoughts and ideas. Ask each pair to write down the differences in reducing power of halides and provide examples. Circulate around the room to facilitate discussion and provide guidance.
- **c.** Ask each pair to share their findings with the class.
- **d.** Write the chemical equations for the reactions of halides salts with concentrated tetraoxosulphate (VI) acid on the board:
  - i. $2Cl^- + H_2SO_4 \rightarrow Cl_2 + SO_4^{2-} + H_2O$
  - ii. $2Br^- + H_2SO_4 \rightarrow Br_2 + SO_4^{2-} + H_2O$
  - iii. $2I^- + H_2SO_4 \rightarrow I_2 + SO_4^{2-} + H_2O$
- **e.** Highlight the differences in reducing power of halides and how it affects the reactions.
- **f.** Lead a class discussion to reinforce understanding and address any misconceptions. Ask questions like:
- **a.** What is the trend in reducing power of halides?
- **b.** How does the concentration of tetraoxosulphate (VI) acid affect the reactions?
- **c.** What are the practical applications of halides' reducing power?

2. **Collaborative learning**
   - **a.** Divide learners into mixed-ability groups of 3-4 and assign each group a specific topic related to halogens (acid strengths, thermal stability, or everyday uses).

     Consider learners' learning styles and abilities when forming groups, and provide extra support or challenges as needed.
   - **b.** Provide guiding questions and relevant information to facilitate conversation.

     Offer additional support for struggling students by providing graphic organisers or visual aids, while challenging advance learners with more complex questions.
   - **c.** Instruct groups to research and collect data on their assigned topic.

     Encourage struggling learners to focus on key points, while challenging advanced learners to explore additional resources or conduct experiments.
   - **d.** Ask groups to create visual aids (tables, graphs, concept maps) to illustrate their findings.
   - **e.** Provide extra support for struggling learners by offering templates or examples, while challenging advanced learners to create more complex or interactive visual aids.
   - **f.** Instruct groups to share their findings with the class.

     Encourage struggling learners to focus on key points, while challenging advanced students to lead the class discussion or provide additional insights.
   - **g.** Ask learners to reflect on their learning and identify areas for further exploration.

     Provide extra support for struggling learners by offering additional resources or review sessions, while challenging advanced learners to design their own assessments or projects.

KEY ASSESSMENTS

Level1: State one industrial use of chlorine and state clearly the chemical basis for use. Support the answer with a chemical equation.

Level 2

1. Consider the reaction of tetraoxosulphate (VI) acid with:
   - **a.** NaCl
   - **b.** NaI
     - i. State the observations for each reaction
     - ii. Write balanced chemical equation for each reaction
     - iii. State the role of tetraoxosulphate (VI) acid in each reaction

2. Arrange HF, HCl, HBr and HI in order of increasing acid strength and give reasons.

Hint

The recommended mode of assessment for week 18 is ***Mid-Semester Examination.*** *Refer to Appendix F for a Table of Specification to guide the setting of the questions.*

Section 6 Review

In this section, learners explored the properties and reactions of halogens, classifying them by physical states and analysing trends in melting and boiling points. They studied bond energies of diatomic halogen molecules, creating graphs to visualize data and identify anomalies. Collaborative tasks included comparing electron configurations, understanding variable oxidation states, and discussing common chemical properties, such as forming salts with metals and producing hydrogen halides. Key reactions, including displacement and precipitation reactions, highlighted halogen reactivity as predicted by standard electrode potential values. Learners also examined the differences in acid strength among hydrogen halides and their thermal stability, connecting theoretical concepts to real-world applications. Discussions on halogens' interactions with water and alkalis reinforced the practical relevance of their chemistry in various industries. Overall, this section emphasised the significance of halogens in both theoretical chemistry and practical contexts.

APPENDIX F: MID-SEMESTER EXAMINATION

Structure

This mid-semester exam will comprise 40 multiple-choice questions (MCQ). Learners should attempt all questions by choosing the correct options. The exam should cover lessons learned from weeks 13-17. Time allowed for this paper is 50minutes.

Resources

Stationery, printer, scannable answer sheet, calculator, pencils, etc

Sample question

MCQ

The halogen with the highest density is…

A. bromine (Br_2).

B. chlorine (Cl_2).

C. fluorine (F_2).

D. iodine (I_2).

How to Administer

Prepare examination questions and materials

Inform learners of examination details (date, time location, etc)

Distribute examination materials (questions, answer sheets, etc)

Monitor learners to prevent cheating and maintain a quiet and secure environment

Table of Specification

Week	Focal Area	Type of Question	DoK Level				Total
			1	2	3	4	
13	Acid - Base Titration	Multiple Choice	3	3	2	-	8
14	Acid - Base Titration	Multiple Choice	2	4	3	-	9
15	Physical and Chemical Properties of Period 3 Elements	Multiple Choice	2	3	1	-	6
16	Physical and Chemical Properties of Compounds of Period 3 Elements	Multiple Choice	3	4	3	-	10

Week	Focal Area	Type of Question	DoK Level				Total
			1	2	3	4	
17	Physical and Chemical Properties of the Halogens (Group 17 Elements)	Multiple Choice	2	3	2	-	7
	Total		**12**	**17**	**11**		**40**

SECTION 7: STRUCTURE, CHEMICAL BONDING AND PROPERTIES OF MOLECULAR COMPOUNDS

Strand: Bonding

Sub-Strand: Systematic Chemistry of the Elements

Learning Outcome: *Describe the structure, chemical bonding and properties of molecular compounds*

Content Standards

1. Demonstrate knowledge and understanding of the role of modelling, evidence and theory in explaining the structure, bonding as well as properties of molecular compounds.
2. Demonstrate knowledge and understanding of multiple bonds and use it to account for the structure of compounds

INTRODUCTION AND SECTION SUMMARY

This two-week section covers the concepts of **hybridisation** and its application to various structures. Learners will explore hybridisation, predict the shape and bond angles for different species, and distinguish between sigma and pi bonds.

By the end of this section, learners will be able to:

1. Predict the shape and bond angles for different species
2. Distinguish between sigma and pi bonds
3. Explain hybridisation and relate it to various structures

The weeks covered by the section are

Week 19: *Hybridisation, structure and bonding in organic compounds*

Week 20: *Hybridisation, structure and bonding in organic compound*

SUMMARY OF PEDAGOGICAL EXEMPLARS

Learners will explore the concepts of hybridisation and chemical bonding in one week. They will use a collaborative learning approach to predict the relative polarity of bonds using electronegativity values and predict the shape and bond angles for species with 4, 3, and 2 negative charge centers using VSEPR theory.

They will also determine molecular polarity based on shape and bond polarities. Through Talk-for-Learning, learners will discuss in small mixed-ability groups how orbital overlap forms molecular orbitals and distinguish between sigma and pi bonds.

Learners will engage in teacher-led discussions on hybridisation, explaining it as the mixing of atomic orbitals to form new orbitals of equivalent energy and shape.

They will work in pairs using molecular models to identify and explain various types of hybridisation (sp, sp^2, sp^3, sp^3d, sp^3d^2, dsp^2). Collaborative learning will continue as learners identify the relationship between Lewis dot structures and hybrid orbitals, use orbital diagrams to explain hybrid orbitals, and predict shapes and bond angles of molecules (e.g., CH_4, $BeCl_2$, BCl_3, NH_3, H_2O, PCl_5, SF_6, $[Ni(CN)_4]^{2-}$).

For their home task, learners will research and compare the structure and bonding in diamond and graphite, and describe the structure and bonding of silicon (IV) oxide.

ASSESSMENT SUMMARY

Assessment types

1. Use MCQs test basic recall of hybridisation types, VSEPR theory, and bond polarity. Include visual aids for reference.
2. Use Short Answer Questions: Explain the formation of sigma and pi bonds, predict molecular shapes and bond angles. Use guiding questions to structure responses.
3. Practical Experiment: Use molecular models to demonstrate hybridisation types (sp, sp^2, sp^3, sp^3d, sp^3d^2). Provide step-by-step instructions and scaffolded support.
4. Essay Questions: Analyse and evaluate the relationship between Lewis dot structures and hybrid orbitals. Allow choice of format (written, presentation, poster).

Delivery

1. MCQs and Short Answer Questions: Administer in class or as homework.
2. Practical Experiment: Conduct in the laboratory with hands-on participation.
3. Essay Questions: Assign as a take-home project with format choice.

WEEK 19

Learning Indicator: *Predict the shape and bond angles for species and distinguish between sigma and pi bonds*

FOCAL AREA: SHAPE AND BOND ANGLES OF MOLECULAR COMPOUNDS AND DIFFERENCES BETWEEN SIGMA AND PI BONDS.

Shape and bond angles for molecular compounds

Electronegativity is a fundamental concept in chemistry that describes the ability of an atom to attract electrons towards itself in a chemical bond. This property is crucial for predicting the nature of bonds between atoms and understanding the behaviour of molecules. The concept of electronegativity was first introduced by Linus Pauling, and it is typically measured on the Pauling scale, which assigns values to elements based on their electron-attracting power.

Electronegativity values range from around 0.7 for elements like cesium and francium to 4.0 for fluorine, the most electronegative element. These values are unitless and provide a relative measure of an atom's ability to attract electrons. Generally, electronegativity increases across a period from left to right and decreases down a group in the periodic table. This trend is due to the increasing nuclear charge and decreasing atomic radius as you move across a period, and the increasing atomic radius and electron shielding effect as you move down a group.

Bond Polarity

The difference in electronegativity values between two bonded atoms can predict the type of bond and its polarity:

1. **Non-polar Covalent Bonds**: When the electronegativity difference between two atoms is less than 0.5, the electrons are shared equally, resulting in a non-polar covalent bond. An example is the bond in a hydrogen molecule (H_2), where both hydrogen atoms have the same electronegativity.
2. **Polar Covalent Bonds**: If the electronegativity difference is between 0.5 and 1.7, the electrons are shared unequally, leading to a polar covalent bond. The more electronegative atom attracts the shared electrons more strongly, creating a partial negative charge (δ-) on that atom and a partial positive charge (δ+) on the less electronegative atom. Water (H_2O) is a classic example, where oxygen is more electronegative than hydrogen, resulting in a polar molecule.
3. **Ionic Bonds**: When the electronegativity difference exceeds 1.7, the more electronegative atom can attract the bonding electrons so strongly that it essentially

takes them away from the less electronegative atom, forming ions. This results in an ionic bond, as seen in sodium chloride (NaCl), where sodium (Na) loses an electron to become Na^+, and chlorine (Cl) gains an electron to become Cl^-.

Dipole Moments

A polar bond creates a dipole moment, which is a measure of the separation of positive and negative charges in the bond. The dipole moment is represented by an arrow pointing from the positive to the negative end, indicating the direction of electron density shift. The magnitude of the dipole moment is typically measured in Debye units (D).

Figure 7.1: *Dipole Moment*

Examples:

1. Hydrogen chloride (HCl): The electronegativity difference between hydrogen (2.1) and chlorine (3.0) is 0.9, indicating a polar covalent bond. Chlorine, being more electronegative, attracts the shared electrons more strongly, resulting in a partial negative charge (δ-) on the chlorine atom and a partial positive charge (δ+) on the hydrogen atom. This creates a dipole moment, which is a measure of the separation of positive and negative charges in the molecule. The arrow points from the positive end (hydrogen) to the negative end (chlorine), indicating the direction of the electron density shift. The presents of the dipole moment in HCl is what makes it polar.

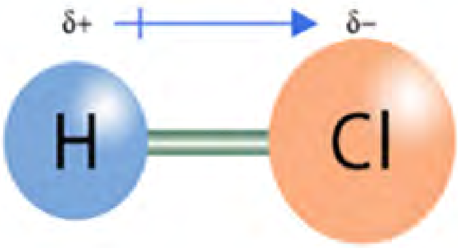


Figure 7.2: *polarity of HCl*

2. Methane (CH_4): The electronegativity difference between carbon (2.5) and hydrogen (2.1) is 0.4, resulting in non-polar covalent bonds. The molecule is symmetrical and due to its symmetrical tetrahedral shape, the dipole moments of the C-H bonds point in different directions and cancel each other out and as a result, methane has no net dipole moment, making it a non-polar molecule.

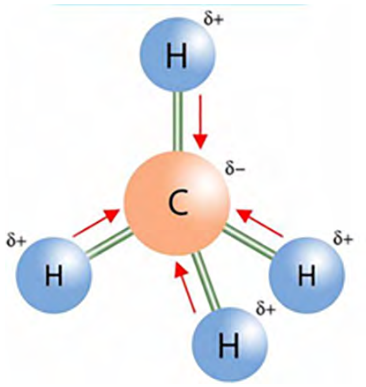


Figure 7.3: *polarity of* CH_4

3. Carbon dioxide (CO_2): The electronegativity difference between carbon (2.5) and oxygen (3.5) is 1.0, indicating the linear shape of CO_2 is highly symmetrical. This symmetry ensures that the dipole moments of the individual C=O bonds point in opposite directions and cancel each other out and as a result, CO_2 has no net dipole moment, making it a non-polar molecule.

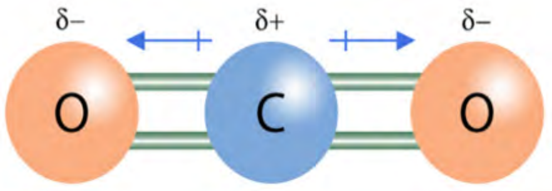


Figure 7.4: *polarity of* CO_2

Predicting Molecular Polarity from Molecular Shape and Bond Polarities Using VSEPR Theory

The valence-shell-electron-pair repulsion (VSEPR) theory considers that molecular shape is determined primarily by the repulsions between pairs of electrons in the molecule or molecular ion, be they bonding pairs or lone pairs. To minimise these repulsions, electron pairs around an inner atom within a molecule will be situated as far apart as possible in the preferred three-dimensional structure.

To determine the shape of a molecule using VSEPR theory, we use the following procedure.

1. Determine and draw the Lewis structure of the molecule.
2. Count the number of sets of bonding pairs and lone pairs of electrons around any inner atom.
3. Modify the geometry, if necessary, to take account of the fact that the magnitudes of repulsions between sets of electron pairs depend on whether the electron pairs involved are bonding pairs (BP) or lone pairs (LP). The repulsions are in the following order. LP–LP > BP–LP > BP–BP.

Therefore, two neighbouring lone pairs on an inner atom will repel each other to a greater extent than will two neighbouring bonding pairs. This is because lone pairs occupy a larger volume than bonding pairs. Note that double bonds occupy a larger volume than single bonds, and some structures may have to be modified for this.

1. Species with 4 Negative Charge Centers

Example: Methane (CH_4)

Methane consists of one carbon atom bonded to four hydrogen atoms. The carbon atom has four valence electrons, and each hydrogen atom has one valence electron, forming four single covalent bonds. The carbon atom has four electron groups (four bonding pairs) around it. According to VSEPR theory, these four electron groups repel each other and arrange themselves in a **tetrahedral shape** to minimise repulsion. The bond angles in a tetrahedral molecule are approximately 109.5°.

In a 3D model, the hydrogen atoms form a tetrahedron around the central carbon atom.

Methane (CH_4) is a non-polar because the bond polarities cancel out due to its symmetrical tetrahedral shape.

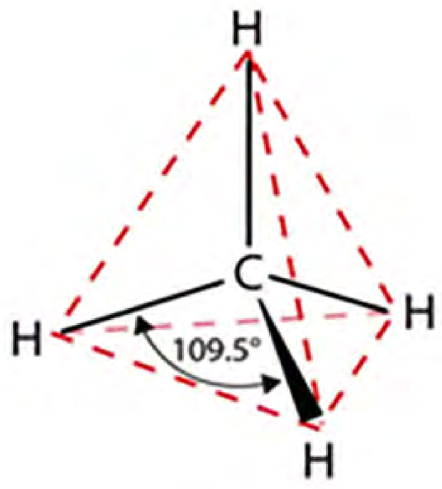


Figure 7.5: *Tetrahedral shape of* CH_4

2. Species with 3 Negative Charge Centers

Example: (Boron trifluoride, BF_3)

Boron trifluoride consists of one boron atom bonded to three fluorine atoms. The boron atom has three valence electrons, and each fluorine atom has seven valence electrons, forming three single covalent bonds.

The boron atom has three electron groups (three bonding pairs) around it and according to VSEPR theory, these three electron groups repel each other and arrange themselves in a **trigonal planar shape** to minimise repulsion.

The bond angles in a trigonal planar molecule are approximately 120°. In a 3D model, the fluorine atoms form an equilateral triangle around the central boron atom. Boron trifluoride (BF_3) is a non-polar molecule because the bond polarities cancel out due to its symmetrical trigonal planar shape.

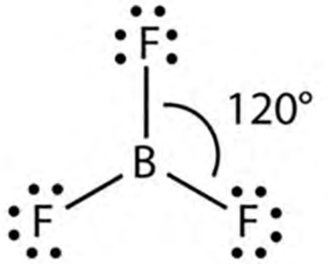


Figure 7.6: *Triagonal planar shape of* BF_3

3. Species with 2 Negative Charge Centers

Example: (Carbon dioxide, CO_2).

Carbon dioxide consists of one carbon atom double-bonded to two oxygen atoms. The carbon atom has four valence electrons, and each oxygen atom has six valence electrons, forming two double bonds. The carbon atom has two electron groups (two double bonds) around it. According to VSEPR theory, these two electron groups repel each other and arrange themselves in a **linear shape** to minimise repulsion. The bond angles in a linear molecule are 180°.

In a 3D model, the oxygen atoms are positioned in a straight line on either side of the central carbon atom.

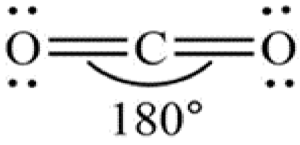


Figure 7.7: *linear shape of* CO_2

Carbon dioxide (CO_2) is a non-polar molecule because the bond polarities cancel out due to its symmetrical linear shape.

Predicting Molecular Polarity from Molecular Shape and Bond Polarities

To determine if a molecule is polar, it is important to consider both the bond polarities and the geometry of the molecule. A molecule with polar bonds can be non-polar if the bond dipoles cancel out due to the molecule's symmetry. Equally, a molecule with an asymmetrical shape and polar bonds will be polar. Example:

1. Water (H_2O): The O-H bonds in water are polar due to the significant difference in electronegativity between oxygen (3.44) and hydrogen (2.20). This creates a partial

negative charge (δ-) on the oxygen atom and a partial positive charge (δ+) on the hydrogen atoms. Water has a bent shape due to the two lone pairs on the oxygen atom. The bond angle is approximately 104.5°. The bond dipoles do not cancel out because of the bent shape, making water a polar molecule with a net dipole moment. This polarity is responsible for many of water's unique properties, such as its high boiling point and its ability to dissolve many substances.

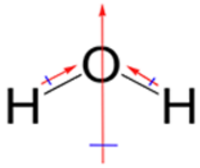
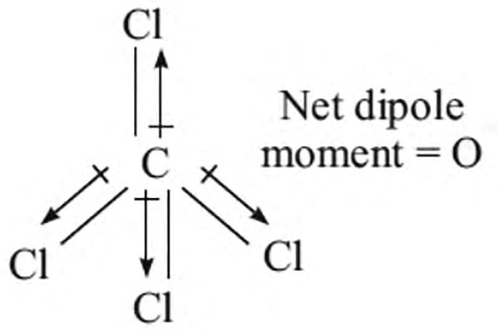


Figure 7.8: *Bent shape of* H_2O

2. Carbon tetrachloride (CCl_4): The C-Cl bonds are polar due to the difference in electronegativity between carbon (2.55) and chlorine (3.16). CCl_4 has a tetrahedral shape, with bond angles of 109.5°. Despite having polar bonds, the tetrahedral symmetry of CCl_4 causes the bond dipoles to cancel out, resulting in a non-polar molecule. This non-polarity makes CCl_4 a good solvent for non-polar substances.



3. Ammonia (NH_3): The N-H bonds are polar due to the difference in electronegativity between nitrogen (3.04) and hydrogen (2.20). NH_3 has a trigonal pyramidal shape due to the lone pair on the nitrogen atom, with bond angles of approximately 107°. The bond dipoles do not cancel out because of the trigonal pyramidal shape, making ammonia a polar molecule with a net dipole moment. This polarity contributes to ammonia's high solubility in water and its ability to form hydrogen bonds.

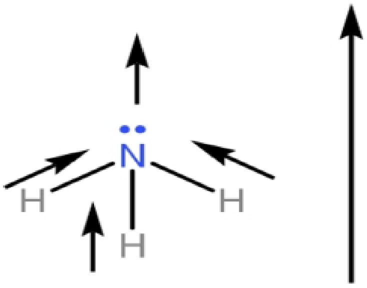
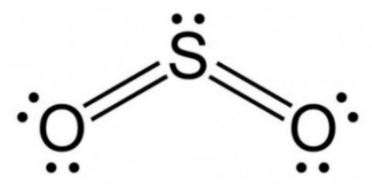


Figure 7.9: *trigonal pyramidal shape of* NH_3

4. Sulphur (IV) oxide (SO_2): The S=O bonds are polar due to the difference in electronegativity between sulphur (2.58) and oxygen (3.44). SO_2 has a bent shape due to the lone pair on the sulphur atom, with bond angles of approximately 119°. The bond dipoles do not cancel out because of the bent shape, making sulphur dioxide a polar molecule with a net dipole moment. This polarity influences SO_2's interactions.

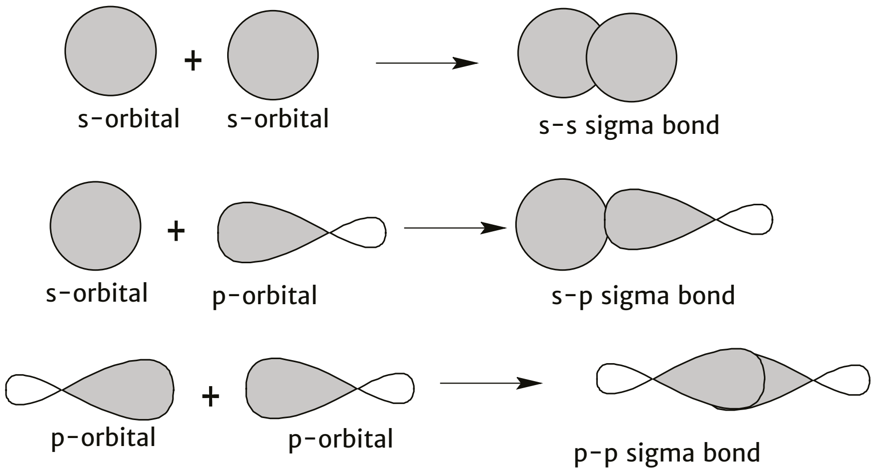


Differences between sigma and pi bonds

In chemistry, understanding the nature of chemical bonds is crucial for comprehending molecular structures and reactions. Sigma (σ) and pi (π) bonds are two fundamental types of covalent bonds that play a significant role in the formation of molecules. This section aims to describe and distinguish between sigma and pi bonds, highlighting their formation, characteristics, and differences.

Sigma Bonds (σ Bonds)

Sigma bonds are formed by the **head-on (axial) overlap** of atomic orbitals. This overlap can occur between two s orbitals, one s and one p orbital, or two p orbitals. Example:



Characteristics of sigma bonds

a. Sigma bonds are the first bonds formed between two atoms and are present in all single bonds.

b. They are generally stronger than pi bonds due to the direct overlap of orbitals, which allows for a greater degree of electron sharing.

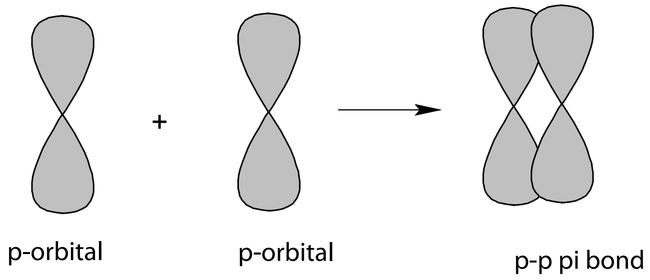
c. Sigma bonds have cylindrical symmetry around the bond axis, meaning the electron density is evenly distributed along the axis connecting the two nuclei.

d. Atoms connected by a sigma bond can freely rotate around the bond axis without breaking the bond. The bond between two hydrogen atoms in an H_2 molecule is a sigma bond formed by the overlap of two 1s orbitals. In methane (CH_4), the carbon-hydrogen bonds are sigma bonds formed by the overlap of sp^3 hybrid orbitals of carbon with the 1s orbitals of hydrogen.

Pi Bonds (π Bonds)

Pi bonds are formed by the **side-by-side (lateral) overlap** of p orbitals. This overlap occurs above and below the plane of the atoms involved in the bond.

Pi bonds can only be formed when there is already a sigma bond between the atoms, as they require the presence of unhybridised p orbitals. Example:



Characteristics of pi bonds

a. Pi bonds are found in double and triple bonds, where they accompany sigma bonds. A double bond consists of one sigma bond and one pi bond, while a triple bond consists of one sigma bond and two pi bonds.

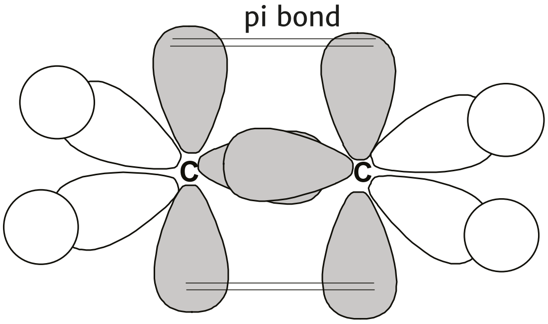
b. Pi bonds are generally weaker than sigma bonds because the sideways overlap is less effective in sharing electron density.

c. Pi bonds do not have cylindrical symmetry. Instead, the electron density is concentrated above and below the plane of the nuclei.

d. The presence of pi bonds restricts the rotation of atoms around the bond axis, leading to rigidity in the molecular structure.

Examples of pi-bonds

In ethene (C_2H_4), the carbon-carbon double bond consists of one sigma bond and one pi bond formed by the lateral overlap of unhybridised p orbitals.



Differences Between Sigma and Pi Bonds

1. **Formation:** Sigma bonds are formed by the head-on overlap, while pi bonds are formed by side-by-side overlap of atomic orbitals.
2. **Strength:** Sigma bonds are stronger due to the direct overlap of orbitals, whereas pi bonds are weaker because of the less effective sideways overlap.
3. **Symmetry:** Sigma bonds have cylindrical symmetry around the bond axis, while pi bonds have electron density above and below the plane of the nuclei.
4. **Rotation:** Sigma bonds allow free rotation of bonded atoms, whereas pi bonds restrict rotation, leading to rigidity in the molecular structure.
5. **Occurrence:** Sigma bonds are present in all single bonds and are the first bonds formed between atoms. Pi bonds are found in multiple bonds (double and triple bonds) and always accompany sigma bonds.

Learning Tasks

1. Calculate the difference in electronegativity for the following pairs of elements: H-Cl, O-H, N-H, C-H, and F-H, and classify each bond as nonpolar covalent, polar covalent, or ionic.
2. Determine the electronegativity difference for each bond in the following molecules: H_2O, CO_2, NH_3, CH_4, and HCl, and predict the overall polarity based on bond polarities.
3. Use VSEPR theory to predict the shape and bond angles of the following simple molecules: CH_4, NH_3, H_2O, and $BeCl_2$, and create 3D models.
4. Predict the shape and bond angles for the following molecules with 4, 3, or 2 negative charge centers on the central atom: SF_4, BF_3, CO_2, and PCl_5 using VSEPR theory.

PEDAGOGICAL EXEMPLARS

1. **Collaborative learning approach**
   - **a.** Introduction: Begin with a brief explanation of electronegativity, VSEPR theory, and molecular polarity to ensure all students have a foundational understanding.
   - **b.** Group Formation: learners into mixed-ability groups to encourage peer learning and support.
   - **c.** Electronegativity Task: Provide each group with a table of electronegativity values and a list of element pairs. Have them calculate the electronegativity differences and classify the bonds as nonpolar covalent, polar covalent, or ionic.
   - **d.** Discussion: Groups discuss their findings and share with the class, allowing for clarification and deeper understanding.
   - **e.** VSEPR Theory Task: Assign each group different molecules with 4, 3, or 2 negative charge centers. Students use VSEPR theory to predict the shapes and bond angles, creating 3D models if possible.
   - **f.** Presentation: Groups present their molecular shapes and bond angles, explaining their reasoning and addressing any questions from peers. Make room for groups to choose their made of presentation.
   - **g.** Polarity Prediction Task: Provide groups with molecules and their Lewis structures. Students predict the molecular shape, determine bond polarities, and predict overall molecular polarity.
   - **h.** Analysis: Groups analyse how molecular shape and bond polarities contribute to overall polarity, discussing their predictions with the class.
   - **i.** Differentiation: Offer additional support and resources to groups or individuals who need it, and provide extension tasks for advanced students to explore more complex molecules.
   - **j.** Reflection: Conclude with a class discussion reflecting on the learning process, key concepts, and any remaining questions.

2. **Talk-for-Learning approach**
   - **a.** Start with a brief overview of molecular orbitals and the importance of orbital overlap in bond formation to ensure all learners have a basic understanding.
   - **b.** Divide learners into small, mixed-ability groups to promote diverse perspectives and peer support.
   - **c.** Provide each group with diagrams of s and p orbitals. Have them discuss how the overlap of these orbitals leads to the formation of molecular orbitals, focusing on sigma and pi bonds.

d. Offer guiding questions to facilitate discussion, such as “What happens when two s orbitals overlap?” and “How does the overlap of p orbitals differ from s orbitals?”

e. Each group presents their discussion points to the class, explaining how orbital overlap results in molecular orbitals and the formation of sigma and pi bonds.

f. Assign each group the task of describing and distinguishing between sigma and pi bonds, using examples and diagrams to illustrate their points.

g. Groups create a Venn diagram comparing sigma and pi bonds, highlighting similarities and differences.

h. Facilitate a class-wide discussion where groups share their Venn diagrams and discuss the key characteristics of sigma and pi bonds.

i. Provide additional resources and support for groups or individuals who need it, and offer extension activities for advanced learners to explore more complex bonding scenarios.

j. Conclude with a reflection session where learners discuss what they learned, any challenges they faced, and remaining questions.

KEY ASSESSMENT

Level 1: Identify the shape of CH_4 molecule using VSEPR theory

Level 2

1. Explain why carbon dioxide (CO_2) is a nonpolar molecule despite having polar bonds.
2. Compare the molecular shapes of boron trifluoride (BF_3) and ammonia (NH_3) and explain the differences in their bond angles.
3. How do the unique properties of diamond and graphite make them suitable for specific industrial applications?

Hint

The recommended mode of assessment for Week 19 is ***research****. Use the level 2 question 3 as a sample question.*

WEEK 20

Learning Indicators

1. *Explain hybridisation and relate it to various structures*
2. *Apply the concept of hybridisation to explain sigma and pi bond of organic compounds and explain their structure and bonding*

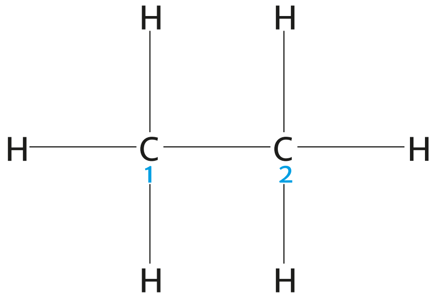
FOCAL AREA: HYBRIDISATION, STRUCTURE AND BONDING IN ORGANIC COMPOUNDS

Hybridisation is a fundamental concept in organic chemistry that explains the bonding and geometry of carbon-containing compounds. Organic compounds are primarily composed of carbon atoms, which have the ability to form stable covalent bonds with other carbon atoms and various elements, leading to an immense variety of molecular structures. The concept of hybridisation helps us understand how carbon atoms can form these diverse structures by mixing their atomic orbitals to create new hybrid orbitals.

In organic compounds, carbon atoms typically undergo three main types of hybridisation: sp^3, sp^2, and sp. Each type of hybridisation corresponds to a different molecular geometry and bonding pattern, which are important for the stability and reactivity of organic molecules. Hybridisation thus helps us to understand the bonding in organic molecules.

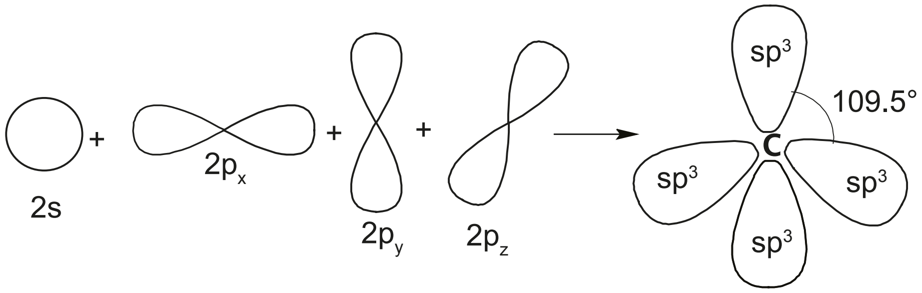
Types of Hybridisation and Bonding

1. **sp^3 hybridisation**: This type of hybridisation is seen in alkanes, where each carbon atom forms four sigma (σ) bonds with other atoms, resulting in a tetrahedral geometry with bond angles of approximately 109.5°. Example ethane (C_2H_6).

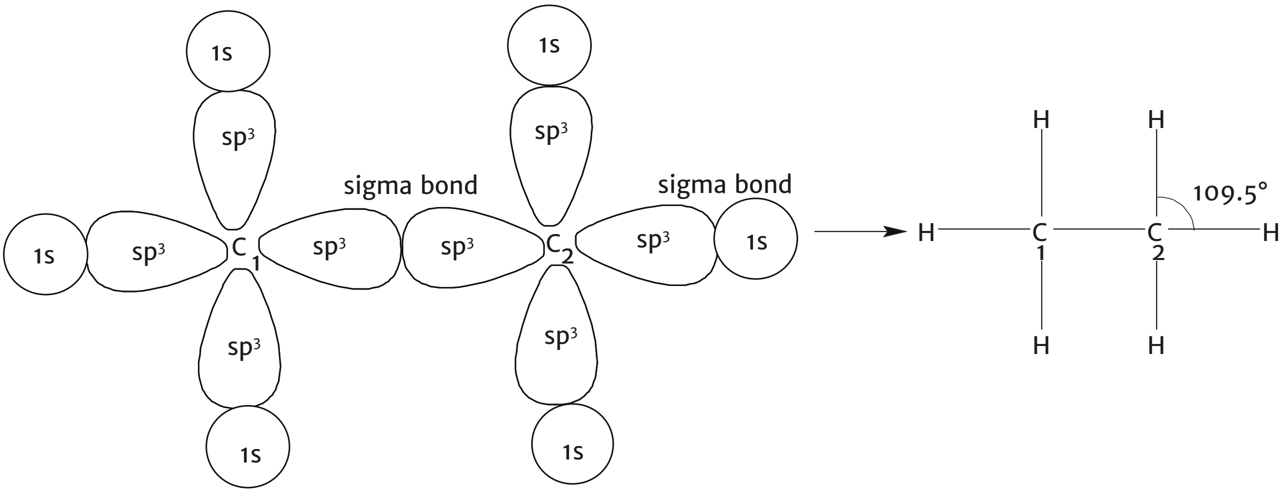


Ethane has two central atoms and each carbon atom in has the ground state electron configuration of $1s^2\ 2s^2\ 2p^2$. In this state, each carbon atom has two unpaired electrons in the 2p orbitals, which is insufficient to form four bonds. One electron from the 2s orbital of each carbon atom is promoted to the empty $2p_z$ orbital and this results in the configuration $1s^2\ 2s^1\ 2p^3$ for each carbon atom, giving four unpaired electrons.

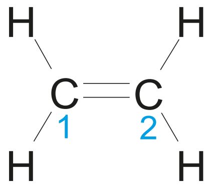
The 2s orbital and the three 2p orbitals (2px, 2py, 2pz) of each carbon atom mix to form four equivalent sp^3 hybrid orbitals. Each sp^3 hybrid orbital consists of 75% p character and 25% s character.



The four sp^3 hybrid orbitals on each carbon atom arrange themselves in a tetrahedral geometry to minimise electron repulsion. This arrangement results in bond angles of approximately 109.5°. In ethane (C_2H_6), each carbon atom uses one of its sp^3 hybrid orbitals to overlap with an sp^3 hybrid orbital of the other carbon atom, forming a sigma (σ) bond. The remaining three sp^3 hybrid orbitals on each carbon atom overlap with the 1s orbitals of hydrogen atoms, forming six sigma (σ) bonds (three per carbon atom).



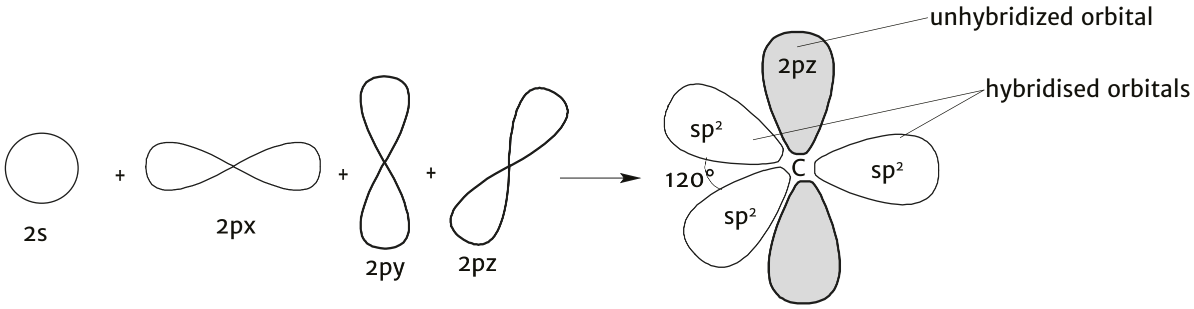
2. **sp^2 hybridisation:** This hybridisation is typical in alkenes, such as ethene (C_2H_4), where each carbon atom forms three σ bonds and one pi (π) bond. The sp^2 hybrid orbitals lie in a plane with bond angles of about 120°, while the unhybridised p orbital forms the π bond.



Each carbon atom in ethene has the ground state electron configuration of $1s^2\ 2s^2\ 2p^2$.

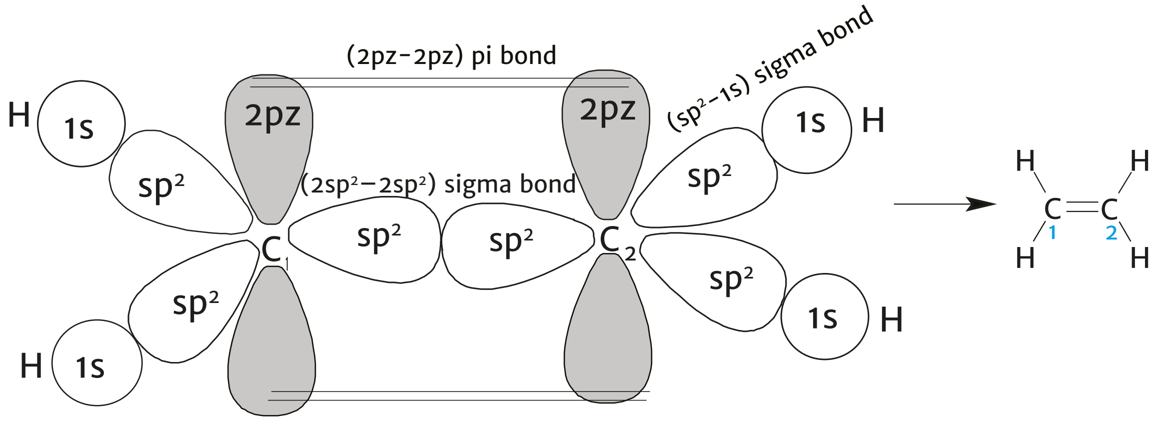
In this state, each carbon atom has two unpaired electrons in the 2p orbitals, which is insufficient to form four bonds. One electron from the 2s orbital of each carbon atom is promoted to the empty 2p orbital and this results in the configuration $1s^2\ 2s^1\ 2px^1\ 2py^1\ 2pz^1$ for each carbon atom, giving four unpaired electrons.

The 2s orbital and two of the 2p orbitals (2px and 2py) of each carbon atom mix to form three equivalent sp^2 hybrid orbitals. There is one leftover 2pz orbital remaining on the carbon atom unhybridised (pure orbital). Each sp^2 hybrid orbital consists of 67% p character and 33% s character.

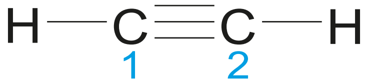


The three sp^2 hybrid orbitals on each carbon atom arrange themselves in a trigonal planar geometry to minimise electron repulsion. This arrangement results in bond angles of approximately 120°.

Each carbon atom uses its three sp^2 hybrid orbitals to form sigma (σ) bonds, two with hydrogen atoms and one with the other carbon atom. The unhybridised 2pz orbitals on each carbon atom overlap side-by-side to form a pi (π) bond. This results in a double bond between the carbon atoms, consisting of one sigma bond and one pi bond.

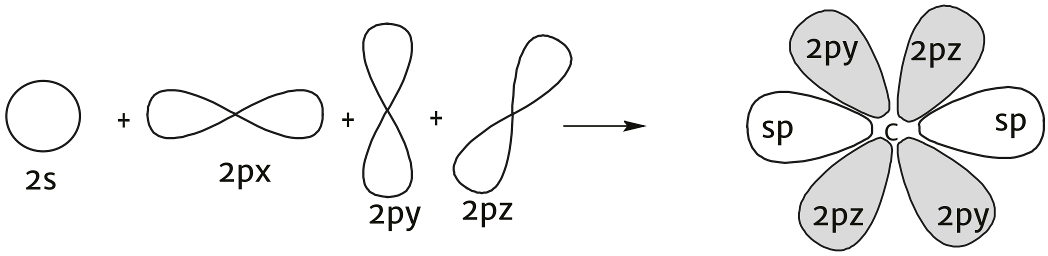


3. **sp hybridisation**: This type of hybridisation is found in alkynes, such as ethyne (C_2H_2), where each carbon atom forms two σ bonds and two π bonds. The sp hybrid orbitals are oriented linearly with bond angles of 180°.

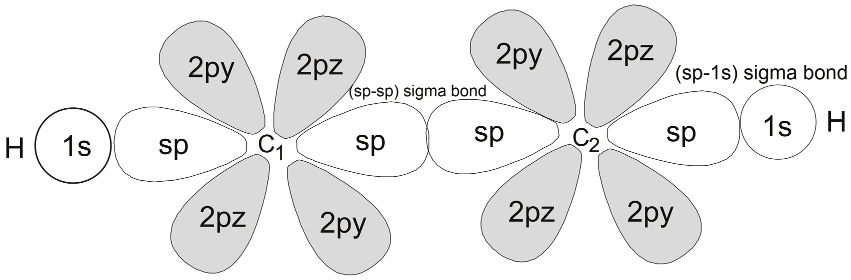


$$H-\underset{1}{C}\equiv\underset{2}{C}-H$$

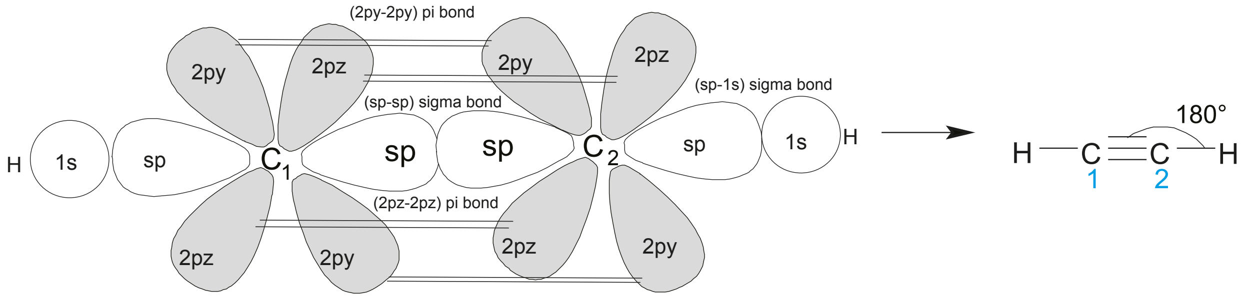
Ethyne has two carbons and each atom has the ground state electron configuration of $1s^2$ $2s^2$ $2p^1$ $2py^1$. In this state, each carbon atom has two unpaired electrons in the 2p orbitals, which is insufficient to form four bonds. One electron from the 2s orbital of each carbon atom is promoted to the empty 2p orbital and this results in the configuration $1s^2$ $2s^1$ $2px^1$ $2py^1$ pz^1 for each carbon atom, giving four unpaired electrons. The 2s orbital and one of the 2p orbitals (2px) of each carbon atom mix to form two equivalent sp hybrid orbitals. Each sp hybrid orbital consists of 50% s character and 50% p character. The remaining two 2p orbitals (2py and 2pz) remain unhybridised.



The two sp hybrid orbitals on each carbon atom arrange themselves in a linear geometry to minimise electron repulsion and this arrangement results in bond angles of 180°. Each carbon atom uses one of its sp hybrid orbitals to overlap with an sp hybrid orbital of the other carbon atom, forming a sigma (σ) bond. The remaining sp hybrid orbital on each carbon atom overlaps with the 1s orbital of a hydrogen atom, forming two sigma (σ) bonds (one per carbon atom).



The unhybridised 2py and 2pz orbitals on each carbon atom then overlap side-by-side to form two pi (π) bonds. This results in a triple bond between the carbon atoms, consisting of one sigma bond and two pi bonds.



Learning Tasks

1. Using molecular model kits, have learners build models of ethene (C_2H_4) and ethyne (C_2H_2). Ensure they correctly represent the sigma and pi bonds.

   Offer step-by-step guidance for learners who need more support, while allowing proficient learners to work independently and explore additional molecules.
2. Ask learners to compare the bonding and geometry of ethene and ethyne. They should describe how the sigma and pi bonds are formed and explain the differences in molecular geometry.

   Provide guiding questions to help learners structure their analysis. Proficient learners can be encouraged to include additional examples and explore the implications of these differences on the molecules' properties.
3. Using molecular model kits, have learners build models of ethene (C_2H_4) and ethyne (C_2H_2). Ensure they correctly represent the sigma and pi bonds.

   Offer step-by-step guidance for learners who need more support, while allowing proficient learners to work independently and explore additional molecules.
4. Ask learners to compare the bonding and geometry of ethene and ethyne. They should describe how the sigma and pi bonds are formed and explain the differences in molecular geometry.

   Provide guiding questions to help learners structure their analysis. Proficient learners can be encouraged to include additional examples and explore the implications of these differences on the molecules' properties.

PEDAGOGICAL EXEMPLARS

1. Exploratory Learning

a. Divide the class into small mixed-ability groups, ensuring each group has a balanced mix of skills and knowledge levels to promote peer learning and collaboration.

b. Assign specific roles within each group, such as leader, recorder, model builder, and presenter, to ensure active participation from all members and cater to different strengths.

c. Briefly explain the purpose of using molecular models, emphasising how they help visualize and understand the structure of carbon-carbon double and triple bonds.

d. Distribute molecular model kits to each group, including carbon and hydrogen atoms, and connectors representing single, double, and triple bonds, to provide hands-on learning experiences.

e. Instruct learners to identify the components needed to construct a carbon-carbon double bond, guiding them to use two carbon atoms and four hydrogen atoms for accurate model building.

f. Have learners connect two carbon atoms using two connectors to represent the double bond, and attach two hydrogen atoms to each carbon atom using single connectors, ensuring they observe the planar geometry with bond angles of approximately 120°.

g. Next, instruct learners to identify the components needed to construct a carbon-carbon triple bond, guiding them to use two carbon atoms and two hydrogen atoms for accurate model building.

h. Have learners connect two carbon atoms using three connectors to represent the triple bond, and attach one hydrogen atom to each carbon atom using single connectors, ensuring they observe the linear geometry with bond angles of 180°.

i. Explain that a carbon-carbon double bond consists of one sigma (σ) bond and one pi (π) bond, describing how the sigma bond is formed by the head-on overlap of sp^2 hybrid orbitals, while the pi bond is formed by the side-by-side overlap of unhybridised p orbitals.

j. Explain that a carbon-carbon triple bond consists of one sigma (σ) bond and two pi (π) bonds, describing how the sigma bond is formed by the head-on overlap of sp hybrid orbitals, while the two pi bonds are formed by the side-by-side overlap of two sets of unhybridised p orbitals.

2. Inquiry-based learning

a. Divide the class into small mixed-ability groups, ensuring diverse perspectives and collaborative learning.

b. Provide relevant videos, charts, or pictures to each group, illustrating the formation of pi and sigma bonds in unsaturated hydrocarbons.

c. Explain the purpose of the visual aids, emphasising how they will help in understanding the bonding in unsaturated hydrocarbons.

d. Have each group watch the provided videos that demonstrate the formation of pi and sigma bonds, catering to visual learners.

e. Encourage learners to study the charts and pictures, focusing on the orbital overlaps that form pi and sigma bonds, to support different learning styles.

f. Guide each group to discuss their observations from the visual aids, promoting inquiry and deeper understanding.

g. Ask learners to identify sigma bonds (formed by linear/co-axial overlap) and pi bonds (formed by lateral/sideway overlap) to reinforce their learning.

h. Provide molecular model kits to each group and instruct them to construct models of unsaturated hydrocarbons, such as ethene (C_2H_4) and acetylene (C_2H_2).

i. Have learners use the models to visualize and demonstrate the linear/co-axial overlap for sigma bonds and the lateral/sideway overlap for pi bonds.

j. Ask each group to present their models and explain the formation of pi and sigma bonds, using the visual aids as references, to consolidate their understanding and communication skills

KEY ASSESSMENTS

Level 1: Identify the type of bond (sigma or pi) in a carbon-carbon single bond.

Level 2: Describe the geometry of a molecule with a carbon-carbon double bond.

Level 3: Analyse the impact of pi bonds on the reactivity of unsaturated hydrocarbons.

Level 4

1. Synthesise information from various sources to explain how the presence of pi bonds influences the physical properties of unsaturated hydrocarbons.
2. Use molecular kits to construct carbon- carbon double and triple bonds, and compare their bond lengths and bond angles

Hint

*The recommended mode of assessment for Week 20 is **practical**. Use the level 4 question 2 as a sample question. Refer to the Teacher Assessment Manual and Toolkit page 128 for further information about how to use this assessment mode.*

Section 7 Review

The two-week exploration of molecular shape, bond angles, and hybridisation was a resounding success, fostering active participation and deep understanding among learners. Collaborative activities, such as predicting bond polarity and determining molecular shapes using VSEPR theory, were particularly effective. Insightful discussions on orbital overlap and molecular orbital formation helped learners distinguish between sigma and pi bonds. Teacher-led explanations of hybridisation clarified the concept, while hands-on activities with molecular models reinforced theoretical knowledge. The home task, which involved researching real-world examples like diamond, graphite, and silicon (IV) oxide, further solidified connections between classroom learning and practical applications. The combination of collaborative, discussion-based, and hands-on learning approaches created a dynamic environment, significantly enhancing learners' grasp of hybridisation and chemical bonding concepts.

SECTION 8: ORGANIC COMPOUNDS

Strand: Chemistry of Carbon Compounds

Sub-Strand: Organic Functional Groups

Learning Outcome: *Explain the source, structure, nomenclature, properties, uses of the various classes of organic compounds and the test for their functional groups.*

Content Standard: Demonstrate knowledge and understanding of source, structure, nomenclature, properties and uses of organic compounds as well as the characteristic test for their functional groups.

Hint

*Remind learners of the end of semester examination in Week 24. Refer to **Appendix G** at the end of this section for a Table of specification to guide the setting of the items.*

INTRODUCTION AND SECTION SUMMARY

This four-week section digs into the captivating world of organic chemistry, exploring the structure, bonding, properties, and everyday uses of alkynes, alkanes, alkenes, benzene, alkanols, and alkanoic acids. Learners will gain a comprehensive understanding of these fundamental concepts, enabling them to analyse and evaluate the relationships between structure, bonding, and properties.

By the end of this section, learners will be able to

1. Explain the structure and bonding of alkanes, alkenes, alkynes, benzene, alkanols, and alkanoic acids.
2. Identify the properties and everyday uses of these organic compounds.
3. Analyse and evaluate the relationships between structure, bonding, and properties.

The weeks covered by the section are

Week 21: *Alkanes*

Week 22: *Alkenes and Alkynes*

Week 23: *Benzene and alkanols*

Week 24: *Alkanoic acids*

SUMMARY OF PEDAGOGICAL EXEMPLARS

Learners will engage in a collaborative learning approach, exploring the concepts of organic chemistry through:

1. Research and presentations on the structure, bonding, properties, and everyday uses of organic compounds.
2. Talk-for-Learning discussions on the relationships between structure, bonding, and properties.
3. Pair work to identify and explain various types of isomerism (structural, stereoisomerism).
4. Teacher-led discussions on the nomenclature of organic compounds.

ASSESSMENT SUMMARY

1. **Assessment types**
    - **a.** MCQs test basic recall of organic compound structures, properties, and uses
    - **b.** Short Answer Questions: Explain the structure, bonding, and properties of organic compounds.
    - **c.** Practical Experiment: Use molecular models to demonstrate isomerism types
    - **d.** Essay Questions: Analyse and evaluate the relationships between structure, bonding, and properties.
2. **Delivery**
    - **a.** MCQs and Short Answer Questions: Administer in class or as homework.
    - **b.** Practical Experiment: Conduct in the laboratory with hands-on participation.
    - **c.** Essay Questions: Assign as a take-home project with format choice.

WEEK 21

Learning Indicator: *Explain the nomenclature, structure, properties and everyday uses of alkanes*

FOCAL AREA: ALKANES

Alkanes are the type of hydrocarbon molecule in which the carbon atoms are connected only by single bonds. In effect, alkanes are saturated hydrocarbons with the general formula (C_nH_{2n+2}). One of the many uses of alkanes is being used as fuel.

Naming of a Straight Chain hydrocarbon

a. Identify the longest continuous carbon chain in the molecule. The number of carbon atoms in the longest straight chain determines the parent name of the alkane.

b. Identify and name any substituent that is present. Number the chain in such a way to give the carbon that bears the substituent the lowest possible number. This becomes the position of the substituent which is written as a number followed by a hyphen and the parent name of the alkane. Methane is an alkane with only one carbon atom. Ethane, Propane and Butane contain two, three and four long-chain carbon atoms respectively. Alkanes with five or more carbon atoms in a straight chain are named using the Greek prefixes: pent, hex, hept, oct, non and dec to represent five, six, seven, eight, nine and ten respectively.

c. When two or more of the same substituents are present in an alkane, precede the name of the substituent with the positions they are located and a numerical prefix. Two, three, and four are assigned the numerical prefixes di, tri and tetra respectively.

d. When two or more different substituents are present, indicate the position followed by the name of each substituent. The positions of the substituent must bear the lowest number possible on the longest carbon chain. The names of the substituents are written in alphabetical order regardless of their position.

e. For the same alkane where it is possible to obtain the same number of longest carbon chain in different directions, the one with the most substituent is used.

Number of C atoms (n)	Molecular Formula (C_nH_{2n+2})	Prefix	Parent Name
1	CH_4	Meth -	Methane
2	C_2H_6	Eth -	Ethane
3	C_3H_8	Prop -	Propane
4	C_4H_{10}	But -	Butane

Number of C atoms (n)	Molecular Formula (C_nH_{2n+2})	Prefix	Parent Name
5	C_5H_{12}	Pent -	Pentane
6	C_6H_{14}	Hex -	Hexane
7	C_7H_{16}	Hept -	Heptane
8	C_8H_{18}	Oct -	Octane
9	C_9H_{20}	Non -	Nonane
10	$C_{10}H_{22}$	Dec -	Decane

Names of Substituents

Formula	Name
CH_3	Methyl
CH_3CH_2	Ethyl
$CH_3CH_2CH_2$	Propyl
F	Fluoro
Cl	Chloro
Br	Bromo
I	Iodo
C_6H_5	Phenyl
NH_2	Amino

Drawing the Structure of alkanes

1. Identify the number of carbon atoms present in the alkane using the parent name
2. Write the number of carbon atoms and draw lines to join them forming the carbon chain
3. Attach hydrogen atoms to the carbon atoms using single bonds. This represents the expanded structural formula.
4. The expanded structural formula can be converted to the condensed structural formula by combining the H atoms with their corresponding C atoms to which they are bonded.

5. The condensed structural formula can also be converted to the line-angle formula as a zig-zag line in which the ends and corners represent C atoms. H atoms are not indicated in the line-angle formula.

Types of organic reactions

The reactivity of alkanes is very low. They do not react with acids, bases, oxidants and reductants under mild conditions. Bonds in alkanes a fully saturated which does not give room for additional bonds. Also, the primary bonds in alkanes are sigma bonds which are strong non-polar bonds, making it difficult for them to be broken. The low polarity of alkanes which is due to their non-polar nature is also another factor. There is no significant difference in electronegativity between carbon and hydrogen atoms making it difficult to attract polar substances to react with them. The absence of lone pairs or empty orbitals which are often involved in initiating reactions also contributes to the low reactivity of alkanes. However, under certain conditions, alkanes can undergo combustion, halogenation and cracking.

Halogenation

Halogenation is a type of substitution reaction where a halogen atom (F, Cl, Br or I) replaces a hydrogen atom in an alkane. Alkanes require special conditions to react with halogens. They do not react with halogens in the cold and dark. This is because the low temperatures reduce the average kinetic energy of the molecules making it difficult to overcome the high activation energy barrier. To initiate the halogenation process, light energy is needed to excite the halogen molecule for it to break into halogen radicals. The conditions necessary for an alkane to undergo halogenation are light, heat and the presence of a catalyst. The general reaction of halogenation is given by:

$R-H+X_2 \rightarrow R-X+HX$

Chlorination

If chlorine, Cl_2, and methane gas, CH_4, are mixed together at room temperature in the absence of strong light, no reaction occurs. However, if the two gases are heated above 100°C, or irradiated with an intense light source, a reaction begins and heat is liberated. Two new substances can be detected: chloromethane and hydrogen chloride. If the reaction is allowed to progress further with more chlorine gas being added, chloromethane itself begins to react and a mixture of dichloromethane, CH_2Cl_2, trichloromethane, $CHCl_3$ (also known as chloroform), and tetrachloromethane, CCl_4 (also known as carbon tetrachloride), is generated.

The initial step of the halogenation reactions involves the breaking of the covalent bond present in Cl_2. This leads to the formation of two chlorine atoms which contains an unpaired electron each (Free radical).

$$Cl - Cl\ h\xrightarrow{eat\ or\ light} Cl\bullet + \bullet Cl$$

The chlorine radical formed is very reactive and can remove hydrogen atoms from the molecules to gain its stable filled state. The reaction between the chlorine radical and methane leads to the formation of HCl and a methyl radical $(\bullet CH_3)$.

$$\overset{\curvearrowright}{Cl\bullet} + \overset{\curvearrowleft\,\curvearrowright}{H - CH_3} \rightarrow H - Cl + \dot{C}H_3$$

The methyl radical is also reactive enough to abstract a chlorine atom from the excess Cl_2 if present.

$$H_3\overset{\curvearrowright}{C\bullet} + \overset{\curvearrowleft\,\curvearrowright}{Cl - Cl} \rightarrow H_3C - Cl + \dot{Cl}$$

The reaction continues until two radical encounter each other to produce a non-radical species.

Bromination of alkanes is also position whereby bromine radicals substitute hydrogen atoms in an alkane molecule. Under controlled conditions, bromine can react with ethane to generate bromoethane but can lead to the formation of dibromoethane and tribromoethane if not controlled.

$$\underset{Ethane}{CH_3CH_3} + \underset{Bromine}{Br_2} \xrightarrow{Heat\ or\ light} \underset{Bromoethane}{CH_3CH_2Br} + HBr$$

Fluorine, F_2, reacts with alkanes under these conditions, but the reactions are highly exothermic (reflecting the high bond strength of the C—F) and so are very difficult to control. Iodine, $I_{2,}$ on the other hand is rarely used because the reaction is endothermic and too much energy is required to initiate the reaction, making it difficult to control.

Combustion

Combustion is a chemical reaction between an alkane and oxygen, resulting in the release of energy. The amount of oxygen available determines the nature of the combustion process. In the presence of excess oxygen only carbon (IV) oxide and water are produced which indicates a complete combustion.

$$C_nH_{2n+2} + \frac{(3n+1)}{2}O_2 \underset{\Delta}{\rightarrow} nCO_2 + (n+1)H_2O$$

Eg. $CH_4 + 2O_2 \underset{\Delta}{\rightarrow} CO_2 + 2H_2O$

$$C_2H_6 + \frac{7}{2}O_2 \underset{\Delta}{\rightarrow} 2CO_2 + 3H_2O$$

Under conditions of oxygen deficiency, incomplete combustion may also occur with the formation of Carbon(II) oxide as a by-product.

$$C_nH_{2n+2} + \frac{(2n+1)}{2}O_2 \underset{\Delta}{\rightarrow} nCO + (n+1)H_2O$$

Eg. $CH_4 + \frac{3}{2}O_2 \rightarrow /\Delta\ CO + 2H_2O$

$$C_2H_6 + 5O_2 \rightarrow \Delta\ 2CO + 3H_2O$$

Carbon (II) oxide is a poisonous gas because it has a high affinity for haemoglobin and gets transported to tissues. Among the organic compounds, alkanes are one of the best chemical sources of energy due to the large amount of energy they release on combustion.

Cracking

Thermal cracking of alkanes involves breaking down large hydrocarbon molecules into smaller ones by exposing them to high temperatures (around 450–750°C) and pressures (70 atm or higher). This process produces smaller alkanes and a significant amount of alkenes, which are valuable for producing polymers like plastics. For example, thermal cracking of decane ($C_{10}H_{22}$) can produce octane (C_8H_{18}) and ethene (C_2H_4):

$C_{10}H_{22}. \rightarrow C_8H_{18} + C_2H_4$

Catalytic cracking, on the other hand, uses a catalyst (often zeolites) to lower the required temperature (about 450–500°C) and pressure compared to thermal cracking. The presence of a catalyst facilitates the breaking of carbon-carbon bonds, leading to a higher yield of branched alkanes and aromatic hydrocarbons, which are more suitable for high-octane fuels. For instance, catalytic cracking of hexadecane ($C_{16}H_{34}$) can produce octane (C_8H_{18}) and butene (C_4H_8):

$C_{16}H_{34} \rightarrow C_8H_{18} + 2C_4H_8$

Reformation

Reformation, also known as catalytic reforming, is a chemical process that converts straight-chain alkanes into branched-chain alkanes, cycloalkanes, and aromatic hydrocarbons, enhancing the octane rating of gasoline. This process uses a catalyst, typically platinum or a combination of platinum and rhenium, under high temperatures (around 500°C) and moderate pressures. For example, when hexane (C_6H_{14}) undergoes reformation, it can be converted into benzene (C_6H_6) and hydrogen gas (H_2):

$C_6H_{14} \rightarrow C_6H_6 + 4H_2$

Specific terminology

An **electrophile** is a species with a positive or partial positive charge which are attracted to regions of high electron density. Electrophiles act as Lewis acids. An electrophile may be a cation (Al^{3+}, Fe^{2+}, H^+, CH_3^+, $CH_3CH_2^+$), a positive portion of a polar molecule ($H^{\sigma+} - Cl^{\sigma-}$, $CH_3CH_2^{\sigma+} - Cl^{\sigma-}$) or molecule with empty orbitals on the central atom (BeO, BCl_3). The simplest electrophile is the hydrogen ion (proton), due to its vacant orbital in the lowest orbital, 1s orbital.

A **nucleophile** is a species with a negative or partial negative charge which are attracted to regions deficient in electrons. Nucleophiles act as Lewis bases. A nucleophile may be an anion (Cl^-, OH^-, HSO_4^-, CH_3COO^-), the negative portion of a polar molecule ($H - Cl$, $CH_3CH_2CH_2 - Br$) or a neutral molecule with at least one lone pair of electron on the central atom (H_2O, NH_3, CH_3NH_2). A lone pair of electrons is typically high in energy because they do not benefit from the stabilization bonding electrons get from being shared between two nuclei. The lone pairs of electrons (one for N, two of equal energy for O) occupy sp^3 orbitals.

Free radicals are chemical species that contain unpaired electrons. They are produced from the homolytic cleavage of covalent bonds where each atom retains one of the shared electrons which results in the formation of radicals. They are highly reactive due to the presence of unpaired electrons. Examples of free radicals include hydroxyl radical ($\bullet OH$), methyl radical ($\bullet CH_3$) and propyl radical ($CH_3\dot{C}HCH_3$). Primary radicals (ie carbon atom with an unpaired electron which is bonded to only one carbon) are more reactive than secondary radicals (ie carbon atom with unpaired electron that is bonded to two other carbon atoms) which are also more reactive than tertiary radicals (ie carbon atom with an unpaired electron which is bonded to three other carbon atoms).

Isomerism is the phenomenon whereby two or more compounds have the molecular formula but different structural formula which causes differences in physical and chemical properties. A set of compounds that exhibit isomerism are called isomers. Isomerism in alkanes is caused by branching, i.e., isomers of alkanes are straight and branched chain compounds.

Isomerism of alkane

The number of structural isomers is determined by the following:

1. Identify the number of carbon atoms present in the molecular formula of the compound.
2. Write a straight-chain carbon skeleton for the given compound.
3. Determine the number of possible chains with a side group of one carbon. Remember that the molecule is named such that the branch occurs at the lowest carbon number. Hence (for example) 2-methyl butane is possible but 3-methyl butane is not (as this is just 2-methyl butane going backwards).
4. Determine the number of possible chains with two side groups each containing one carbon atom. Remember to apply the same conventional rule as above.
5. Determine the number of possible chains with three or more side groups each containing one carbon atom (If possible). Remember to apply the same conventional rule as above.
6. Determine the number of possible chains with one or more side groups each containing two carbon atoms. Remember to apply the same conventional rule as above.
7. Ensure that no carbon atom has more than four bonds.

8. Each of the possible chains is a new structural isomer.

Physical properties of alkanes

Physical properties of alkanes are important because they determine the conditions under which the compound can be handled and used. The form in which a drug is manufactured and dispensed is affected by its physical properties. Alkanes are non-polar compounds that usually exist as gases, volatile liquids or soft solids at room temperature. They are held together by Van der Waals forces. The nature in which a particular alkane is found is due to the strength of intermolecular bonds that hold the molecules together. The strength of Van der Waals forces increases with the length of the carbon chains present in the molecule. Alkanes with a low number of carbons (C1 to C4) exist as gases at room temperature due to the weak Van der Waals forces that hold them together while those between (C5 to C17) are liquids and those beyond C18 are present as solids. Melting and boiling points are greatly affected by the size of compounds as well as the number of carbon atoms present in compounds.

Melting and boiling points are influenced by the length of carbon chains and their branching. An increase in the number of carbon chains of an alkane increases the size of the compound. This leads to an increase in Van der Waals forces and a higher surface area where the molecule has more electrons in its structure, leading to stronger temporary dipoles and, consequently, stronger intermolecular attractions. For example, butane has a boiling point of -1.0°C while pentane has a boiling point of 36.1°C. Pentane has a longer carbon chain compared to butane, which leads to an increase in surface area and subsequently a stronger Van der Waals force due to more significant interactions between molecules. As a result, it requires more energy (heat) to break these interactions, which is reflected in the higher boiling point. However, branches reduce the surface area that is exposed to the intermolecular forces. Branching reduces the boiling point of alkane compounds. For example, pentane has a boiling point of 36.1°C while 2-methylbutane has a boiling point of 27.8°C. Although 2-methylbutane has five carbon atoms, just like pentane, its structure is branched, which reduces the surface area available for intermolecular contact. The rounder shape of branched alkanes means fewer points of contact between adjacent molecules, which leads to weaker Van der Waals forces and consequently a lower boiling point than would be expected from the number of carbons alone.

Alkanes are non-polar, which makes them insoluble in water. However, they are soluble in non-polar solvents such as other alkanes. The bonds between carbon and hydrogen are considered non-polar because the difference in electronegativity between carbon and hydrogen is minimal (i.e. 2.5 and 2.1 respectively). This means neither atom pulls on the shared electrons significantly more than the other, resulting in a balanced electrical charge distribution throughout the molecule. Because water molecules attract each other strongly through intermolecular forces (hydrogen bonds), they do not interact well with non-polar substances like alkanes. When an alkane is introduced to water, the water molecules will remain clustered together and exclude the alkane molecules, leading to a lack of solubility.

Alkanes have densities from 0.62 g/mL to about 0.79 g/mL, which is less than the density of water (1.0 g/mL). The molecular structure of alkanes consists of long chains or branched arrangements which are made up of only carbon and hydrogen. Hydrogen is the lightest atom (atomic mass of about 1.008 g/mol). The predominance of these lighter atoms in alkanes contributes to their overall low molecular mass. The tetrahedral arrangement around carbon atoms in alkanes means that the molecules occupy a larger volume for a given mass, contributing to their lower density.

Reactivity of alkanes in terms of bond enthalpies and bond polarity

Alkanes are saturated hydrocarbons which are made up of single covalent bonds (C – C and C – H). The bond enthalpy of alkanes is high particularly in the C – H bonds which are strong and stable. This requires a significant amount of energy or a catalyst to break the bonds.

The electronegativity of atoms in alkanes are quite similar and so they are characterised by relatively non-polar bonds. This implies that alkanes do not engage readily in polar reactions, such as those with strong electrophiles or nucleophiles. Reactivity often depends on the presence of polar bonds, which facilitate interactions with other reactive species. Since alkanes do not exhibit significant polarity, they have limited ability to participate in such reactions, further contributing to their inertness, especially under standard conditions.

Effect of combustion on the Environment

The combustion of alkanes is an exothermic process which leads to the formation carbon (IV) oxide or carbon (II) oxide depending on the availability of oxygen. Increased CO_2 contributes to the greenhouse effect and global warming. When sunlight reaches the Earth, some of it is absorbed by the surface, warming the planet. The Earth then emits this energy back as infrared radiation (heat). CO_2 and other greenhouse gases absorb a portion of this heat and re-radiate it back toward the surface. This process keeps the Earth warmer than it would be without these gases. Factors that increase the global warming are tied to the conditions that increase the amount of CO_2 in the environment such as burning coal or natural gases for energy, deforestation which reduces the number of plants that can absorb CO_2 from the atmosphere and industrial processes that release CO_2 directly or indirectly.

Uses of alkanes

The first four alkanes (methane, ethane, propane, and butane) are gases at room temperature and are widely used as heating fuels. Methane is the primary component of natural gas, which is extracted from underground reservoirs. It can also be found in biogas, a renewable source produced from the decomposition of organic matter (e.g., in landfills or anaerobic digesters). It is widely used in household kitchens for cooking, industrially as a feedstock for producing chemicals and also in commercial heating systems. Liquefied Petroleum Gas (LPG) primarily consists of two alkanes: propane (C_3H_8) and butane (C_4H_{10}). These hydrocarbons are predominantly gaseous at room temperature and pressure but can be easily liquefied under pressure, which allows for convenient storage and transport. Both propane and butane provide a high energy yield when combusted, making LPG an efficient fuel source.

Alkanes having five to eight carbon atoms (pentane, hexane, heptane, and octane) are liquids at room temperature. They are highly volatile, which makes them useful in fuels such as gasoline.

Liquid alkanes with 9 to 17 carbon atoms have higher boiling points. The longer saturated carbon chains provide higher energy density, meaning more energy is released per unit of fuel burned. This is vital for engines that need to run efficiently over long distances without needing frequent refueling. They are found in kerosene, diesel, and jet fuels. Diesel fuel has a higher boiling point compared to gasoline, which means it can withstand higher compression in the engine. This leads to better fuel efficiency and torque, making it suitable for heavy-duty applications. The properties of jet fuel ensure that it remains liquid at low temperatures experienced at high altitudes, while also burning efficiently to provide the thrust necessary for flight.

Motor oil and mineral oil are both important products derived from hydrocarbons, and they serve distinct purposes based on their chemical properties. Motor oil is a mixture of high-molecular-weight liquid hydrocarbons, used to lubricate and clean the internal components of engines. Mineral oil is a mixture of liquid hydrocarbons and is used as a laxative and a lubricant. It works best as a laxative because it is insoluble in water and can coat the digestive tract, which helps in the easy passage of stool.

Alkanes with 18 or more carbon atoms are waxy solids at room temperature. The long hydrocarbon chains make them hydrophobic meaning they do not have polar functional groups that can attract water molecules. Known as paraffins, they are used in waxy coatings added to fruits and vegetables to retain moisture, inhibit mold, and enhance appearance. It creates a thin, protective layer that acts as a physical barrier. This barrier limits the evaporation of moisture from the surface of the produce.

Vaseline, is a semisolid mixture of hydrocarbons with more than 25 carbon atoms used in ointments, cosmetics and as a lubricant. The longer carbon chains makes vaseline thicker, so it does not flow like liquids. Due to this it retains a consistent, spreadable form, which is ideal for lotions and ointments. Its thick and greasy nature makes it an effective lubricant. This is useful not just in beauty products but also in mechanical applications where a thick lubricant is needed.

Learning Tasks

1. State the general formula of alkanes and use it to develop the first 10 alkanes
2. Explain why alkanes are described as saturated hydrocarbons
3. Differentiate between a nucleophile and an electrophile using specific examples.
4. Draw the isomers of butane and pentane
5. Give the IUPAC name of $CH_3CH_2CHBrCH_2CH(CH_3)CH_3$
6. Arrange the following alkanes in increasing order of boiling point and explain the order

   [2,3-dimethylbutane, hexane, 2-methylpentane, methane]
7. Write a balanced equation for the combustion of butane.
8. Describe the formation of acid rain
9. State and explain two uses of alkanes in the food industry.

PEDAGOGICAL EXEMPLARS

1. **Activity-Based Learning**
   - **a.** Begin the lesson by introducing the concept of alkanes, inquire about the general formula of alkanes from the learners.
   - **b.** Guide learners to use the general formula of alkanes to develop the first ten alkanes.
   - **c.** Guide learners through the steps of IUPAC naming of alkanes in naming straight chain alkanes, branched chain alkanes and alkanes with substitutes.
2. **Talk-for-Learning:** In a whole class discussion,
   - **a.** Explain saturated and unsaturated compounds and guide learners to identify alkanes as a saturated hydrocarbon.
   - **b.** Use models, charts or drawings to explain and guide learners to identify electrophiles, nucleophiles and free radicals.
   - **c.** Explain the types of organic reactions alkanes can undergo.
   - **d.** State and explain the factors that affect reactions of alkanes.
   - **e.** Explain the types of bond cleavage in alkanes (homolytic and heterolytic fission) and the nature of the products that are formed in these reactions.

f. Explain the effect of functional groups on alkanes

g. Use models or charts to explain isomerism (chain, position, functional group, geometrical)

3. **Activity-based Learning**
   - **a.** Guide learners to model the isomers of butane and pentane
   - **b.** Present the IUPAC rules of naming alkanes
   - **c.** Guide learners to name straight-chain alkanes (C1 – C10)
   - **d.** Assist learners in drawing the structural formula of a named alkane as well as writing its molecular formula
   - **e.** Guide learners to name branched-chain alkanes (up to C6)
   - **f.** Assist learners in drawing the structural formula of branched alkane as well as writing its molecular formula
4. **Inquiry-based learning**
   - **a.** Investigate the physical properties of alkanes such as melting points, boiling points, density, volatility and solubility.
   - **b.** Explain the low reactivity of alkanes in terms of bond enthalpies and bond polarity.
   - **c.** Describe using equations, complete and incomplete combustion reactions of alkanes
   - **d.** Discuss the effect of their products on the environment (greenhouse effect and global warming),
   - **e.** The release of CO and the use of catalytic converters to minimise pollution.
5. **Collaborative Learning**
   - **a.** Discuss the effect of other products of burning fuel (formation of acid rain)
   - **b.** Describe using equations, the reactions of methane with chlorine. Discuss the effect of the product (CFCs) on the ozone layer.
   - **c.** Discuss free radical mechanism of the reaction of methane and chlorine
   - **d.** Discuss cracking and reformation of alkanes
   - **e.** Discuss the uses of alkanes. (i.e., Uses of the various petroleum fractions obtained from fractional distillation of crude oil).

KEY ASSESSMENTS

Level 1: Define isomerism

Level 2: Crude oil is made up of alkanes that can be separated into fractions.

a. Give the names of four petroleum fractions and state the uses of each.

b. Name the process by which these fractions are obtained from crude oil.

c. Naphtha fractions are cracked to produce useful products. Write an equation to illustrate thermal cracking, using $C_{12}H_{26}$ as an example.

Level 3

1. Natural gas is a fossil fuel with methane as its main component and sulphur – containing impurities.

  a. Write balanced equation for the combination of methane in both a limited and an excess supply of oxygen.

  b. Identify two pollutants formed in the combustion of natural gas and state why each is considered as a pollutant.

  c. Describe the effect of the products of complete combustion of methane on the environment.

2. The economic benefits of petroleum production (which contains alkanes) outweigh the environmental impacts of alkane combustion and extraction processes. Write an essay arguing "for" or "against" the motion

Hint

*The recommended mode of assessment for week 21 is **debate**. Use the level 3 question 2 as a sample question. Refer to the Teacher Assessment Manual and Toolkit page 130 for more information on how to administer this assessment mode.*

WEEK 22

Learning Indicator: *Explain the nomenclature, structure, properties of alkenes and alkynes*

FOCAL AREA 1: ALKENES

Alkenes are hydrocarbon compounds containing carbon-carbon double bonds. Alkenes are found in some plant extracts like citrus fruits (limonene), steroids (cheolesterol) and insect pheromones (muscalure). The industrial production of alkenes is mainly by cracking of other hydrocarbon compounds.

Physical properties of Alkenes

The boiling points of alkenes are low, they increase with the length of the carbon-carbon chain and decrease with branching.

Compared to alkanes, alkenes have a relatively lower boiling point and densities. This is because the presence of the double bond introduces a degree of rigidity in the molecular structure, which affects how tightly the molecules can pack together. Density is affected by both the mass of the molecules and how closely they pack together in a given volume. Since alkenes have a less compact arrangement due to their structural configuration, they tend to have lower densities than alkanes of similar molecular weight. Also, alkenes have a lower molecular weight and lower surface area exposed for Van der Waals forces.

Alkenes are less dense than water and are non-polar organic compounds.

Alkenes are insoluble in water but soluble in alkene solvents and in other organic solvents.

Alkenes with fewer carbons (less than five carbon chains) are colourless gases whereas those with higher carbon chains are volatile liquids.

Alkenes have the general formula (C_nH_{2n}) where n is the number of carbon atoms present in the molecule. They form homologous series with straight and branched chain.

Isomerism in Alkenes

Alkenes show both structural and geometric isomerism

Structural isomerism refers to the phenomenon where compounds with the same molecular formula have different arrangements of atoms within the molecule. In alkenes, this can occur due to the position of the carbon-carbon double bond or the variations in the carbon chain length. There are two types of structural isomerism in alkenes. These are position isomerism and chain isomerism.

1. ***Position isomerism***: This occurs when the position of the carbon-carbon double bond changes in the carbon chain. For example, in 1-butene, the double bond is

between the first and second carbon, while in 2-butene, it is between the second and third carbon.

1-Butene (CH_2=CH-CH_2-CH_3) vs. 2-Butene (CH_3-CH=CH-CH_3)

2. ***Chain isomerism*** in alkenes refers to the phenomenon where alkenes with the same molecular formula have different arrangements of their carbon atoms, leading to distinct structures. This type of isomerism arises because the carbon chain can be either straight or branched, even though the number of carbon and hydrogen atoms remains the same.

   For example, consider the molecular formula C_4H_8, which can represent different chain isomers:

   But-1-ene: A straight-chain alkene where the double bond is between the first and second carbon atoms:

   CH_2=CH-CH_2-CH_3

   2-methylpropene: A branched-chain isomer with the double bond in a different carbon framework:

   CH_2=C(CH_3)-CH_3

   Each of these compounds has the same molecular formula but differs in the arrangement of the carbon chain, demonstrating chain isomerism.

IUPAC nomenclature of alkenes up to C6

Alkene nomenclature follows the rules for naming hydrocarbon compounds as discussed in the previous weeks.

a. Identify the longest carbon chain containing the carbon double bond and replace the suffix "-ane" of alkanes with "-ene." When more than one double bond is present, di, tri, tetra etc precedes the suffix.

b. Number the carbon-carbon chain from the end nearest the multiple bonds so that the carbon atoms in that bond have the lowest possible numbers.

c. If the double bond is equidistant from both ends of the chain, number the chain from the end nearest the first branch point.

d. Indicate the position of the multiple bond using the lower numbered carbon atom of that bond.

Synthesis of Alkenes

The synthesis of alkenes is important in organic chemistry due to their reactivity and utility as intermediates in various chemical reactions. They can be synthesised from alcohols, alkyl halides and alkynes.

Preparation from alcohols: Alcohols are dehydrated in the presence of catalysts to obtain alkenes. The reaction takes place under special conditions such as excess concentrated H_2SO_4 and heat.

$H_2SO_4 \rightarrow HSO_4^- + H^+$

$CH_3CH_2 - \ddot{O}H + H^+ \rightarrow CH_3CH_2 - OH_2^+ \rightarrow CH_3CH_2^+ + H_2O\,CH_3 - CH_2^+ + HSO_4^- \rightarrow CH_2$
$= CH_2 + H_2SO_4$

Synthesis from alkyl halides

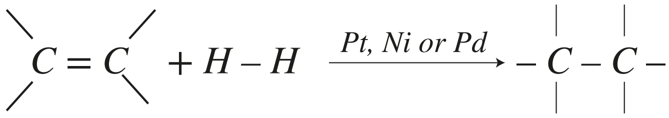
This process typically involves elimination reactions, where a hydrogen halide (HX) is removed from the alkyl halide, resulting in the formation of a carbon-carbon double-bond. In eliminating the HX from the alkyl halide the hydrogen atom is preferentially eliminated from the carbon with the least number of H atoms.

$$CH_3CH_3 \xrightarrow{OH^-} CH_2 = CH_3 + H - OH + Br^-$$

Reactions of symmetrical and unsymmetrical alkenes with:

a. Hydrogen

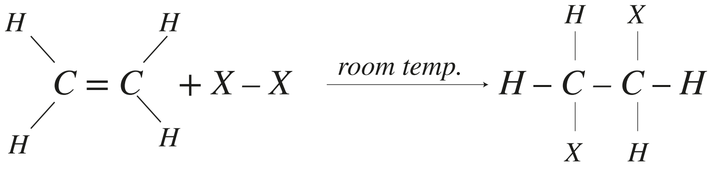
Hydrogen gas is an example of a symmetrical reagent. Under ordinary conditions, hydrogen and alkenes do not react. Special conditions such as temperature, pressure and catalyst are required for this reaction. The suitable conditions include heat (above 130 degrees C and nickel, platinum or palladium catalyst. When an alkene and hydrogen gas are mixed in the presence of catalysts such as palladium, two hydrogen atoms are added to the double bond to give a saturated alkane. This reaction is called a hydrogenation reaction and it is very exothermic. The reaction can be represented as;



This reaction is very useful in margarine production from vegetable oils which are soft and have a high proportion of cis-alkenes.

b. Halogens

Halogens are also categorized as symmetrical reagents. Conditions required for this reaction are room temperature, dark and polar environment. The halogen is added across the double bond to form a saturated dihaloalkane. The equation for the reaction between a halogen and the simplest alkene (ethene) can be represented as;



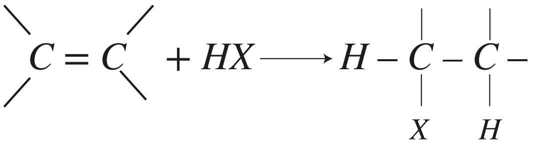
c. Steam

The reaction is referred to as an acid-catalyzed hydration of alkenes. An acid is used to catalyze the reaction because the water molecules a too weak to dissociate into its ions.

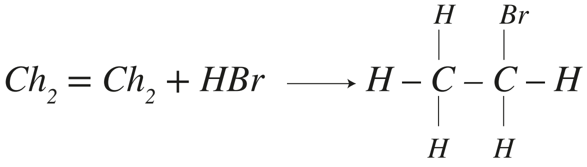
$$CH_2 = CH_2 + H_2SO_4 \xrightarrow{COLD} CH_3CH_2 - OSO_4 \xrightarrow{STEAM} CH_3CH_2OH + H_2SO_4$$

e. Hydrogen halides

This is an example of an unsymmetrical reagent. When it reacts with an alkene the double bond becomes saturated with hydrogen halide to form haloalkanes. The reaction is given by;

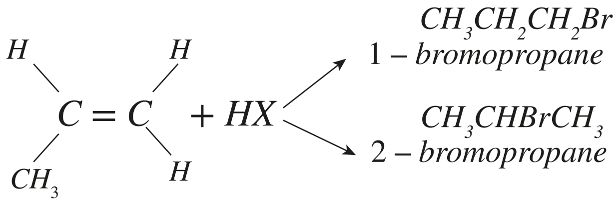


Where HX is HCl, HBr or HI. The reactivity of the hydrogen halides increases down the group of the halogens with HI having the highest reactivity.



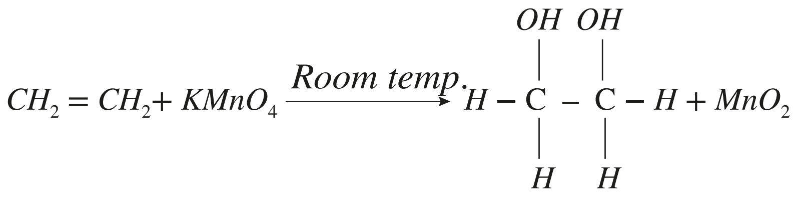
$$Ch_2 = Ch_2 + HBr \longrightarrow H - \overset{\displaystyle H}{\underset{\displaystyle H}{C}} - \overset{\displaystyle Br}{\underset{\displaystyle H}{C}} - H$$

When an unsymmetrical reagent reacts with an unsymmetrical alkene, more than one product is formed. A major product and the others which can be described as minor or trace products. The major product is formed from the more stable secondary (2°) carbonium ion which forms faster and easily.



f. $KMnO_4$

When neutral aqueous $KMnO_4$ reacts with excess ethane at room temperature, it leads to the production of 1,2-ethanediol $HOCH_2CH_2OH$. The oxidation number of manganese is reduced from +7 to +4. This reaction is characterised by a colour change from purple to brown.



$$CH_2 = CH_2 + KMnO_4 \xrightarrow{Room\ temp.} H - \overset{\displaystyle OH}{\underset{\displaystyle H}{C}} - \overset{\displaystyle OH}{\underset{\displaystyle H}{C}} - H + MnO_2$$

Distinguishing between alkanes and alkenes using bromine water or bromine in tetrachloromethane

Saturated hydrocarbons with single C-C bonds have very low reactivity. They are generally unreactive with bromine under normal conditions due to the lack of a double bond. Alkenes on the other hand react with bromine, resulting in the addition of bromine across the double bond. When chlorine water which is reddish-brown solution is added to an alkene a dibrominated alkane is formed which decolourizes the bromine water.

Importance of the reactions of alkenes (production of margarine and alcohol)

Alkenes serve as essential raw materials for the synthesis of various chemicals such as alcohols, polymers, and other organic compounds.

They are also used in the synthesis of polymers.

1. Alkenes like butene and propene are used in gasoline formulations to improve combustion and performance.
2. Additionally, alkenes can be converted into more complex hydrocarbon.
3. They are used in the production of detergents.

They are also used in the production of ethylene glycol (antifreeze chemical) used in automobile radiators.

Learning Tasks

1. Write the general formula for alkenes and use it to develop the molecular formula of the first six alkenes.
2. Model the isomers of a given set of alkenes
3. Use chemical equations or diagrams to explain how alkenes are formed
4. Use chemical equations or charts to explain how alkenes react with a given set of reagents.
5. Explain how alkenes can be distinguished from an alkane using relevant chemical equations and stating the observations that will be made.
6. Outline three uses of alkanes in the food industry.

PEDAGOGICAL EXEMPLARS

1. **Collaborative learning approach:** Using think-pair-share,
    a. Guide learners to recall the general formula of alkenes
    b. In whole class discussion guide learners to develop the molecular formula of ethene after which each group will be assigned the task of developing the

molecular formulae of a given set of alkenes with a specific number of carbons. Ask learners to share their molecular they have developed with other groups.

- **c.** Lead a class discussion on how to model or draw the structural formula of ethene. Encourage all learners to contribute in the modeling. Task learners to draw or model the structure of the alkenes whose molecular formula have been developed by other groups.

2. Activity-based Learning

- **a.** Explain isomerism in alkenes.
- **b.** Model each isomer of butene.
- **c.** Task learners to draw the structural formula the isomers of butene that were modeled.
- **d.** Ask learners to model and draw the isomers of pentene.
- **e.** Guide learners through the IUPAC rules of naming alkenes. Use the rules to name a specific alkene. Encourage all learners to participate in the naming.
- **f.** Task learners to apply the IUPAC rules to name alkenes up to C6 (straight chain and branched chain).
- **g.** Design a mind map to show the sources, the laboratory preparation and physical properties of alkenes.
- **h.** Write chemical equations to illustrate the reactions of symmetrical and unsymmetrical alkenes with:
  - i. Hydrogen,
  - ii. Halogens,
  - iii. Steam,
  - iv. Hydrogen halides
  - v. KMnO4

Note

Use Markornikov's rule for the unsymmetrical alkenes.

Experiential learning: Distinguish between alkanes and alkenes using bromine water or bromine in tetrachloromethane.

Initiating Talk-for-Learning

1. Outline the economic importance of the reactions of alkenes (production of margarine and alcohol)
2. Watch video to reinforce the concepts taught

KEY ASSESSMENTS

Level 1: Give two (2) importance of ethane.

Level 2: Distinguish between $CH_3CH = CH_2$ and $CH_3CH_2CH_3$ chemically

Level 3: Butene (C_4H_8) has three isomers.

a. Draw the structures of each isomer and give its IUPAC name.

b. Which of the isomers undergo geometrical isomerism? Draw structures to show the cis and trans isomers.

FOCAL AREA 2: ALKYNES

Introduction to Alkynes

Alkynes are a unique and important class of hydrocarbons characterised by the presence of at least one carbon-carbon triple bond. This triple bond is a defining feature that sets alkynes apart from other hydrocarbons like alkanes (single bonds) and alkenes (double bonds). The general formula for alkynes is $C_nH_{2n-2,}$ where $n \geq 2$, indicating that they are unsaturated hydrocarbons with fewer hydrogen atoms compared to their alkane counterparts.

when $n = 2$, C_2H_2 (ethyne)

when $n = 3$, C_3H_4 (propyne)

when $n = 4$, C_4H_6(butyne), etc.

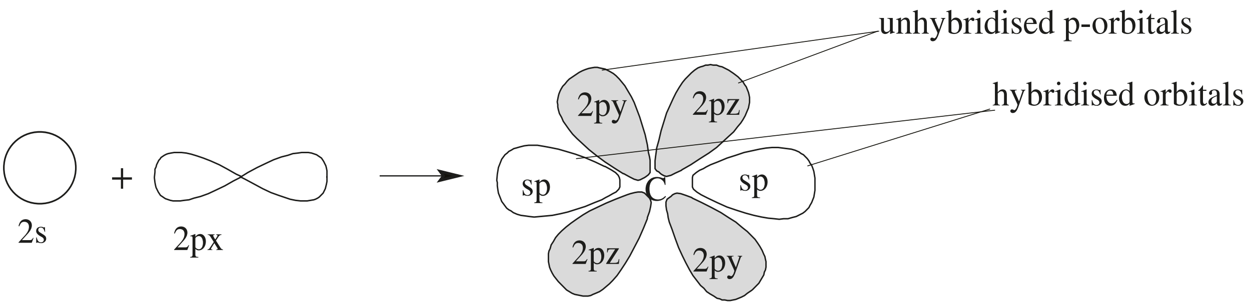
Alkynes are less common than alkenes, both in the laboratory and in living organisms.

The triple bond in alkynes consists of one sigma bond and two pi bonds, which makes it shorter and stronger than single or double bonds. This bond structure also contributes to the linear geometry of alkynes, resulting in a bond angle of 180^0.

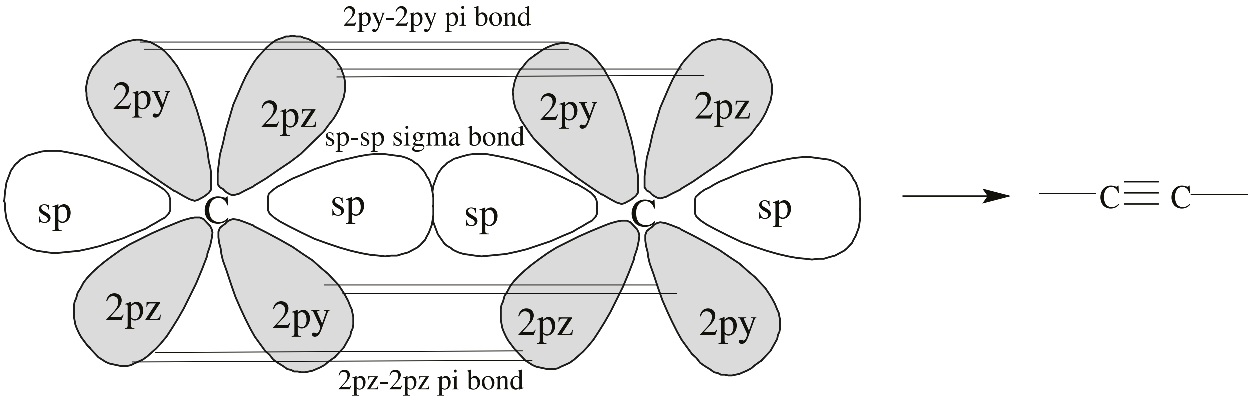
Alkynes are known for their reactivity, which is higher than that of alkanes and alkenes. This increased reactivity is due to the high electron density in the triple bond, making alkynes more susceptible to addition reactions. These reactions are crucial in organic synthesis, allowing for the creation of a wide range of complex molecules.

Structure of alkynes

In alkynes, the carbon atoms involved in the triple bond are sp-hybridised. This hybridisation involves the mixing of one s orbital and one p orbital from each carbon atom to form two sp hybrid orbitals. These sp orbitals form a linear structure with a bond angle of 180 degrees, resulting in a straight-line geometry.



The triple bond in alkynes consists of one sigma bond and two pi bonds. The sigma bond is formed by the head-on overlap of sp orbitals, while the pi bonds are formed by the side-by-side overlap of unhybridised p orbitals.

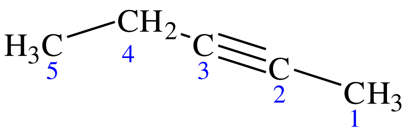


This arrangement makes the triple bond shorter and stronger than single or double bonds, contributing to the unique reactivity of alkynes.

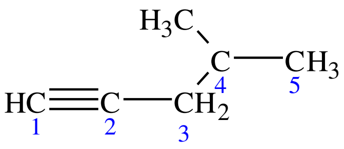
The linear structure of alkynes also affects their physical properties. For instance, they are generally non-polar and have low solubility in water but dissolve well in organic solvents. The presence of the triple bond also makes alkynes more reactive than alkanes and alkenes, allowing them to participate in a variety of chemical reactions, such as hydrogenation and halogenation.

Nomenclature of alkynes

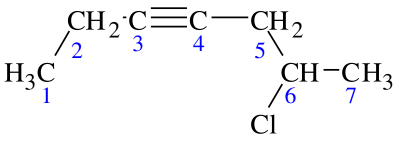
Alkyne nomenclature adheres to the general rules for hydrocarbons discussed in weeks 20 and 21. The suffix "-yne" is used to denote the presence of a triple bond. The position of the triple bond is indicated by the number of the first carbon involved in the bond. Numbering the main chain starts from the end closest to the triple bond, ensuring that the triple bond receives the lowest possible number. Examples:



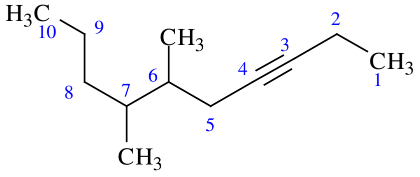
pent-2-yne



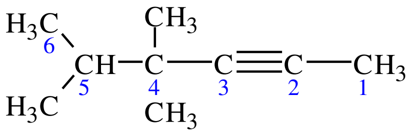
4-methylpent-1-yne



6-chlorohept-3-yne



6,7-dimethyldec-3-yne



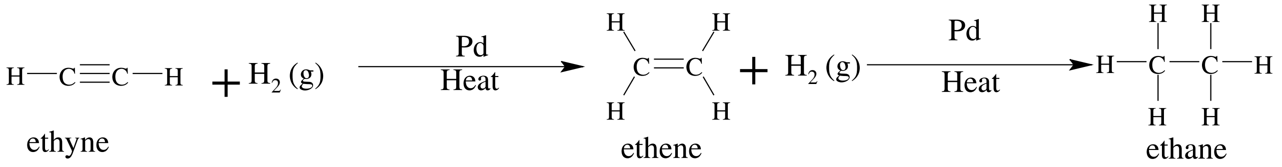
4,4,5-trimethylhex-2-yne

Properties of alkynes

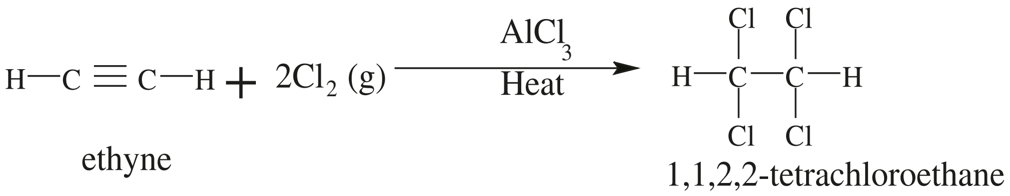
1. State and Appearance: The first three members of the alkyne series (ethyne, propyne, and butyne) are gases at room temperature. Higher order alkynes are liquids or solids. All alkynes are colourless, and ethyne has a distinctive garlic-like odour.
2. Polarity and Solubility: Alkynes are generally non-polar molecules. They are slightly soluble in polar solvents like water but dissolve well in organic solvents such as benzene, acetone, and chloroform. This solubility pattern is similar to that of alkanes and alkenes.
3. Boiling and Melting Points: The boiling and melting points of alkynes increase with molecular weight. Alkynes have slightly higher boiling points than alkanes and alkenes of similar molecular weight due to the increased electron density in the triple bond, which enhances London dispersion forces.
4. Density: Alkynes are less dense than water. Their densities increase with molecular weight but remain lower than those of corresponding alkanes and alkenes.

Chemical properties of alkynes

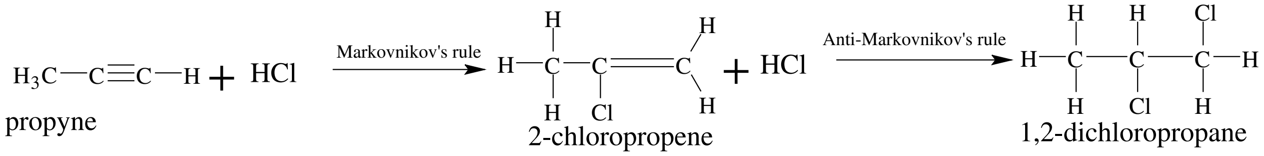
1. **Hydrogenation**: Alkynes can undergo hydrogenation, where hydrogen molecules add across the triple bond. This reaction typically requires a metal catalyst such as palladium, platinum, or nickel. The addition of one equivalent of hydrogen converts the alkyne to an alkene, while the addition of two equivalents results in an alkane. For example, the hydrogenation of ethyne (acetylene) can produce ethene (ethylene) and eventually ethane.



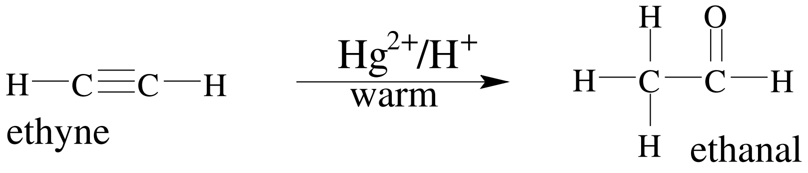
2. **Halogenation:** Alkynes react with halogens (e.g., chlorine, bromine) to form dihaloalkenes and tetrahaloalkanes. The reaction with one equivalent of halogen yields a dihaloalkene, while excess halogen leads to a tetrahaloalkane. This reaction is useful for synthesizing halogenated compounds. Example:

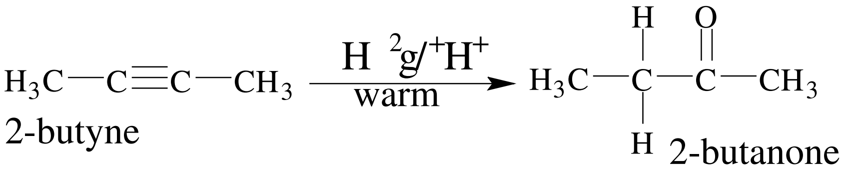


3. **Hydrohalogenation:** Alkynes react with hydrogen halides (e.g., HCl, HBr) to form haloalkenes and dihaloalkanes. The addition follows Markovnikov's rule, where the hydrogen atom attaches to the carbon with the greater number of hydrogen atoms, and the halogen attaches to the carbon with fewer hydrogen atoms. Example:



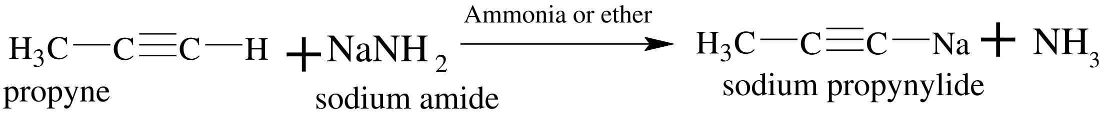
4. Alkynes can undergo **hydration** in the presence of a catalyst, such as mercuric sulphate in aqueous sulphuric acid, to form ketones or aldehydes. This reaction also follows Markovnikov's rule. For example, the hydration of ethyne yields ethanal, while the hydration of other alkynes typically produces ketones.





5. Alkynes can be **oxidised** by strong oxidising agents like potassium permanganate ($KMnO_4$) or ozone (O_3). These reactions cleave the triple bond, resulting in the formation of carboxylic acids or ketones. For instance, the oxidation of ethyne with potassium permanganate produces ethanedioc acid (oxalic acid).

6. Terminal alkynes (those with a triple bond at the end of the carbon chain) exhibit slight acidity due to the sp-hybridised carbon atom, which is more electronegative than sp^2 or sp^3 hybridised carbons. This acidity allows terminal alkynes to **react with strong bases** like sodium amide ($NaNH_2$) to form acetylide ions. These ions are valuable intermediates in organic synthesis, enabling the formation of new carbon-carbon bonds.

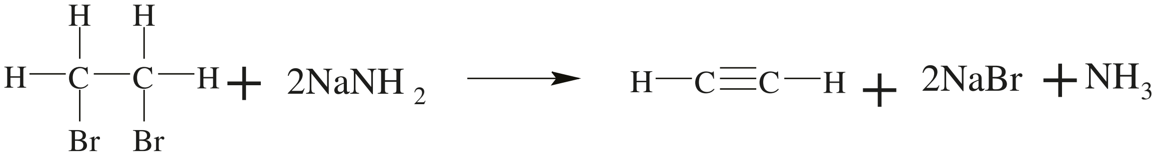


7. **Polymerisation**: Under specific conditions, alkynes can undergo Polymerisation to form complex polymers. This property is utilised in the production of synthetic materials and plastics. For example, the Polymerisation of acetylene can produce polyacetylene, a conductive polymer.

Laboratory preparation of alkynes

1. **Elimination Reactions of Dihalides:** Alkynes can be synthesised from dihalides through a double elimination reaction. This process involves the removal of two molecules of hydrogen halide (HX) from a dihalide compound. The dihalides used can be either vicinal (halogens on adjacent carbons) or geminal (halogens on the same carbon).

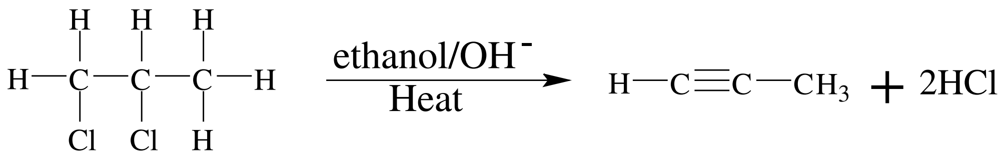
   Mechanism: A strong base, such as sodium amide ($NaNH_2$) in liquid ammonia (NH_3), is used to abstract a proton adjacent to a halogen atom, leading to the formation of a double bond (alkene). The base then abstracts another proton from the alkene, resulting in the formation of a triple bond (alkyne). Example:



2. **De-hydrohalogenation of alkyl dihalides**

   This method involves the removal of hydrogen halides from vicinal or geminal dihalides to form alkynes. The reaction typically requires a strong base and high temperatures to proceed efficiently.

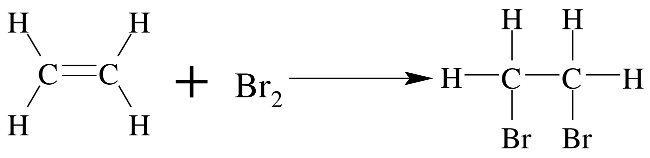
Mechanism: The base abstracts a proton from the carbon adjacent to the halogen, forming a double bond (alkene). A second equivalent of the base abstracts another proton, leading to the formation of a triple bond (alkyne).

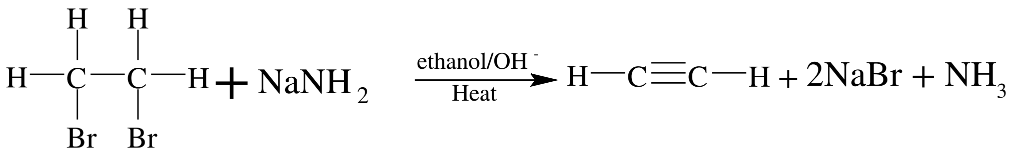


$$\mathrm{CH_2Cl{-}CHCl{-}CH_3 \xrightarrow[\text{Heat}]{\text{ethanol/OH}^-} H{-}C{\equiv}C{-}CH_3 + 2HCl}$$

3. **Preparation from Alkenes:** Alkynes can also be prepared from alkenes through halogenation followed by dehydrohalogenation. This method involves the addition of halogens (e.g., Br_2) to an alkene to form a dihaloalkane, which is then subjected to dehydrohalogenation to yield the alkyne.

   Mechanism: Halogenation of the alkene to form a dihaloalkane followed by dehydrohalogenation of the dihaloalkane using a strong base to form the alkyne.





$$\mathrm{CH_2Br{-}CH_2Br + NaNH_2 \xrightarrow[\text{Heat}]{\text{ethanol/OH}^-} H{-}C{\equiv}C{-}H + 2NaBr + NH_3}$$

4. The preparation of alkynes from calcium carbide is a well-established **industrial method,** particularly for producing acetylene (ethyne). This process involves a series of chemical reactions starting from limestone (calcium carbonate).

   a. Step1: limestone ($CaCO_3$) is heated to produce quicklime (CaO) and carbon dioxide (CO_2).

   $CaCO_3 \rightarrow CaO + CO_2$

   b. Step 2: Quicklime is then heated with coke (carbon) in an electric furnace to produce calcium carbide (CaC_2) and carbon monoxide (CO).

   $CaO + 3C \rightarrow CaC_2 + CO$

   c. Step 3: Calcium carbide is then made to react with water to produce ethyne (C_2H_2) and calcium hydroxide ($Ca(OH)_2$).

   $CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$

Everyday uses of alkynes

Alkynes are vital in many everyday applications, from industrial processes to consumer products. Their usefulness and reactivity make them valuable in numerous fields, contributing to advancements in technology, agriculture, and healthcare, notably among them are:

1. **Fuel and Energy**: Acetylene, the simplest alkyne, is widely used as a fuel in welding and cutting metals due to its high flame temperature. It is also used in portable lighting, such as miner's lamps.
2. **Chemical Synthesis**: Alkynes are crucial intermediates in organic synthesis. They are used to produce a variety of organic compounds, including pharmaceuticals, agrochemicals, and dyes. Their reactivity allows for the creation of complex molecules through various chemical reactions.
3. **Polymer Production**: Alkynes play a significant role in the production of polymers. For instance, acetylene is used to produce polyvinyl chloride (PVC), a common plastic used in pipes, cables, and clothing. Other alkynes are involved in the synthesis of polyethylene and other plastics.
4. **Artificial Ripening**: Ethyne (acetylene) is used in the agricultural industry to artificially ripen fruits. This process helps in controlling the ripening time, ensuring that fruits reach the market in optimal condition.
5. **Solvents and Additives**: Alkynes are used in the production of solvents and as additives in fuels, paints, and coatings. These applications take advantage of their chemical properties to improve the performance and stability of various products.
6. **Pharmaceuticals**: In the pharmaceutical industry, alkynes are used to synthesise active pharmaceutical ingredients (APIs). Their unique chemical structure allows for the development of drugs with specific therapeutic effects, etc.

Learning Tasks

1. List the physical properties of alkynes.
2. Create a Venn diagram comparing and contrasting alkynes, alkanes, and alkenes.
3. Compare the bond angles and lengths in alkynes with those in alkanes and alkenes and explain the differences.
4. Provide the IUPAC names for a series of given alkyne structures.
5. Identify and list five everyday products that contain alkynes or are produced using alkynes.
6. Research and write an essay on the role of alkynes in modern technology and medicine.

PEDAGOGICAL EXEMPLARS

Collaborative learning approach

1. Divide learners into small groups and ask them to brainstorm what they know about alkynes. Encourage wild ideas and provide guidance for struggling learners.
2. Assign each group a topic related to alkynes (e.g. structure, acidity, reactions). Ask them to research and gather information, providing scaffolding for struggling learners.
3. Have each group member become an expert on a specific aspect of alkynes. Ask them to teach their group members, providing support for struggling learners.
4. Provide groups with problems related to alkynes (e.g. write chemical equations, identify isomers). Ask them to work together to solve the problems, encouraging advanced learners to take on more challenging tasks.
5. Ask each group to present their findings to the class. Encourage peer feedback and provide guidance for struggling learners.
6. Facilitate a class discussion on the topics researched by each group. Encourage advanced learners to lead the discussion and provide additional support for struggling learners.
7. Ask learners to reflect on what they learned and what they would like to learn more about. Provide prompts for struggling learners and encourage advanced learners to think critically.
8. Show a video to reinforce concepts taught. Provide additional support for struggling learners by offering a graphic organiser.
9. Create a quiz with questions related to alkynes. Ask learners to work in groups to complete the quiz, providing scaffolding for struggling learners.
10. Summarise key concepts and ask learners to reflect on what they learned. Provide additional support for struggling learners and encourage advanced learners to set goals for further learning.

KEY ASSESSMENTS

Level 1: What is the general formula for alkynes?

Level 2

1. Explain the process of naming an alkyne using the IUPAC system.
2. Draw the following compounds:
   - **a.** 4,4-dimethyl-2-pentyne
   - **b.** 3-octyne
   - **c.** 3-methyl-1-hexyne

Level 3

1. Compare and contrast the physical and chemical properties of alkanes, alkenes, and alkynes.
2. Design an experiment to demonstrate the combustion of an alkyne and explain the expected results.
3. The mystery of the unknown hydrocarbon. Analyse the following clues:
   - **a)** I am a hydrocarbon with 6 carbon atoms
   - **b)** I have a single double bond
   - **c)** I am a key player in plastic production.

   Use your knowledge of Alkenes to figure out my identity. Who am I?

Hint

The recommended mode of assessment for week 22 is **gamification.** Use the level 3 question 3 as a sample question. Refer to the *Teacher Assessment Manual and Toolkit page 148 for more information on how to go about this assessment mode.*

WEEK 23

Learning Indicators

1. *Describe the structure, bonding, properties of benzene and its uses in everyday life*
2. *Explain the structure, nomenclature, preparation, properties and uses of alkanols as well as their characteristics test*

FOCAL AREA 1: BENZENE

Introduction to benzene

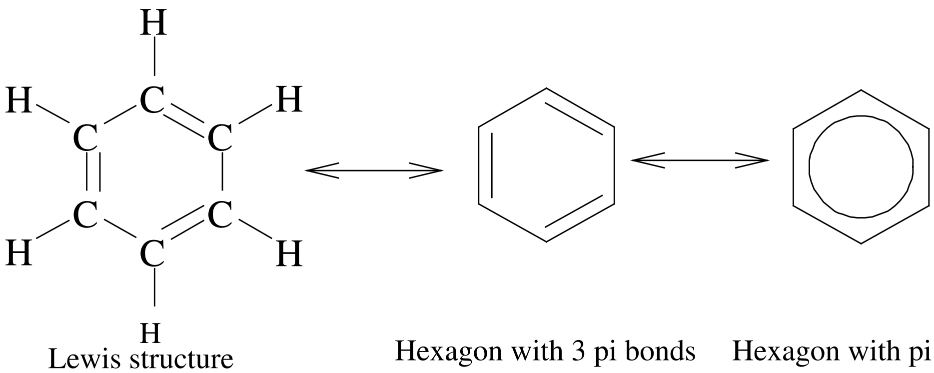
Benzene is one of the most fundamental and widely studied organic compounds in chemistry. Its unique structure and properties make it a cornerstone in the field of organic chemistry. Benzene is a naturally occurring substance produced by volcanoes and forest fires and present in many plants and animals, but benzene is also a major industrial chemical made from coal and oil. Benzene is found in many useful compounds, for example in pharmaceuticals, pesticides, polymers and dyes. The common painkillers aspirin, paracetamol, ibuprofen and morphine all contain benzene rings.

Discovery of Benzene

Benzene was first discovered by the English scientist Michael Faraday in 1825 while studying illuminating gas. Later, in 1834, the German chemist Eilhardt Mitscherlich produced benzene by heating benzoic acid with lime. Another significant contribution came from the German chemist A.W. von Hofmann, who isolated benzene from coal tar in 1845.

Structure of benzene

Benzene has the molecular formula **C_6H_6**, indicating it consists of six carbon atoms and six hydrogen atoms.



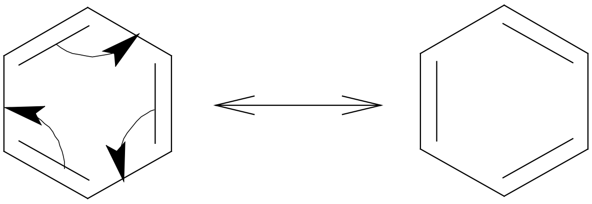
Lewis structure

Hexagon with 3 pi bonds

Hexagon with pi electron delocalisation

The structure of benzene has been a subject of interest since its discovery. In 1865, the German chemist August Kekulé proposed that benzene consists of a ring of six carbon atoms with alternating single and double bonds. This suggests that benzene should react in the same way that an unsaturated alkene does, however, this is not the case.

This structure as proposed by August Kekulé was later refined to show that the double bonds are not fixed but resonate between different positions, giving benzene a unique stability. This concept is known as **resonance**.

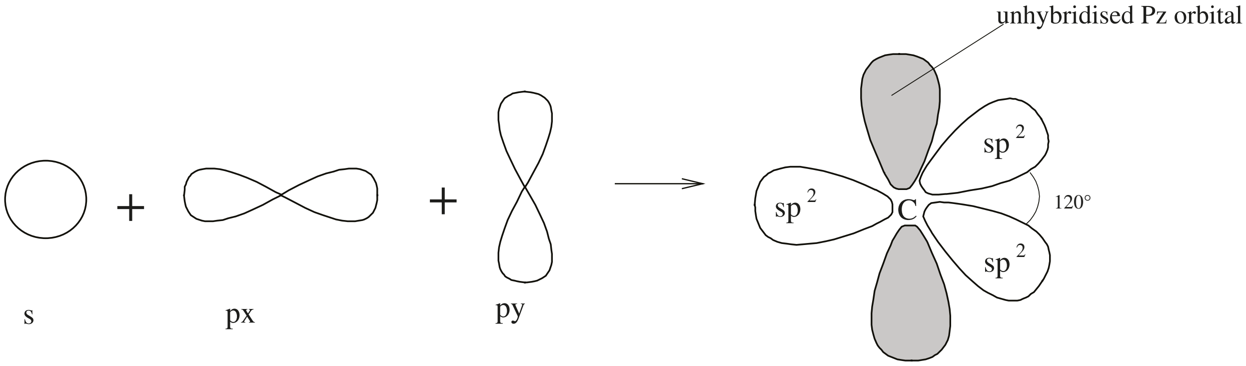


Kekule structures of benzene

Bonding in Benzene

Modern bonding theories, such as valence-bond and molecular orbital theories, explain benzene's structure in terms of delocalization of electrons.

Each carbon atom in benzene undergoes sp^2 hybridisation. This hybridisation involves the mixing of one s orbital and two p orbitals to form three sp^2 hybrid orbitals. These sp^2 hybrid orbitals lie in a plane, 120° apart, forming a trigonal planar structure around each carbon atom.



Each carbon atom uses two of its sp^2 hybrid orbitals to form σ bonds with two adjacent carbon atoms, creating a continuous ring of carbon atoms. The third sp^2 hybrid orbital on each carbon atom forms a σ bond with a hydrogen atom. This results in a hexagonal ring structure with alternating single and double bonds when viewed in a simplified Kekulé structure.

Each carbon atom in benzene has one unhybridised p orbital that is perpendicular to the plane of the sp^2 hybrid orbitals. These p orbitals overlap sideways with the p orbitals of adjacent carbon atoms. The sideways overlap of p orbitals creates a π bond above and below the plane of the carbon atoms.

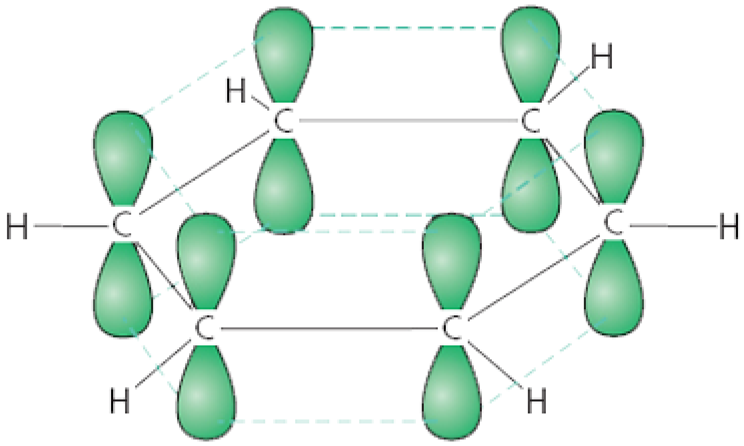


Figure 8.1: *sideways overlap of unhybridised p orbitals to form pi bonds*

This overlap is not localized between specific carbon atoms but extends over the entire ring, forming a continuous π system. The electrons in these π bonds are delocalized, meaning they are shared equally among all six carbon atoms.

The delocalization of π electrons across the benzene ring provides extra stability to the molecule, a phenomenon known as aromatic stabilization. This delocalization is often represented by a circle inside the hexagonal ring in structural diagrams of benzene, indicating the equal distribution of π electrons.

Benzene can be represented by two resonance structures, where the positions of the double bonds alternate. The true structure of benzene is a resonance hybrid of these two structures, with the π electrons delocalized over the entire ring.

The delocalization of electrons results in all carbon-carbon bonds in benzene having the same length.

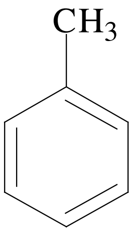
Implications of Benzene's Bonding

1. **Chemical Reactivity**: Benzene's delocalized π system makes it less reactive in addition reactions, which would disrupt the aromatic ring. Instead, benzene typically undergoes electrophilic aromatic substitution reactions, where one hydrogen atom is replaced by another atom or group without disturbing the delocalized π system.
2. **Aromaticity**: Benzene is the simplest example of an aromatic compound, characterised by its cyclic, planar structure and delocalized π electrons.

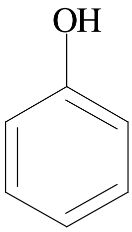
Nomenclature of benzene

The nomenclature of benzene derivatives can be a bit complex due to the variety of substituents that can attach to the benzene ring. Here are some key points to help you understand how benzene compounds are named:

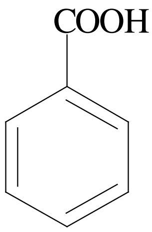
1. **Common Names**: Some benzene derivatives have common names that are widely used. For example: Toluene (methylbenzene), Phenol (hydroxybenzene), Aniline (aminobenzene), Benzoic acid (carboxybenzene), etc.



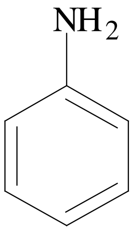
toluene



phenol

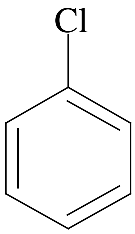


benzoic acid



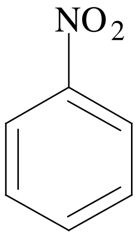
aniline

2. **Simple Benzene Derivatives**: When a single substituent is attached to the benzene ring, the compound is named by prefixing the name of the substituent to "benzene." For example:



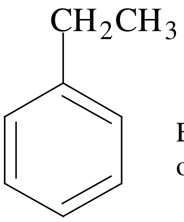
Benzene with a chlorine atom

Chlorobenzene



Benzene with a nitro group (-NO_2)

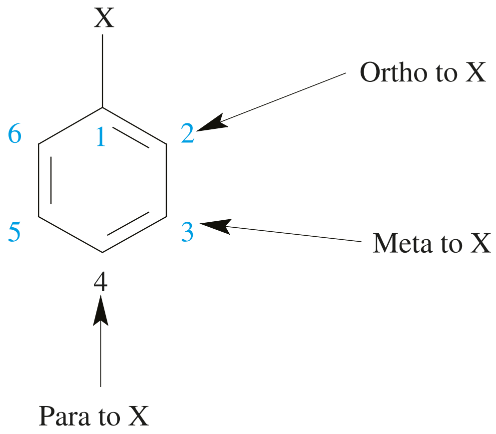
Nitrobenzene



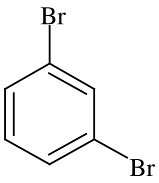
Ethyl group on benzene

Ethylbenzene

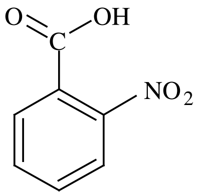
3. **Disubstituted Benzenes**: When two substituents are attached to the benzene ring, their relative positions are indicated using the terms ortho (o-), meta (m-), and para (p-):



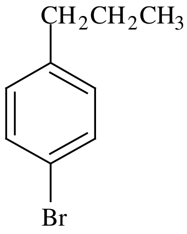
Examples include:



meta-Dibromobenzene

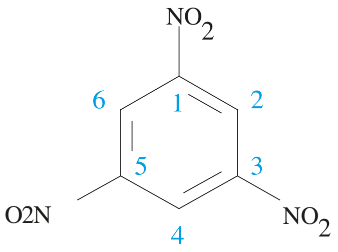


ortho-Nitrobenzoic acid



para-Bromopropylbenzene

4. **Polysubstituted Benzenes**: When more than two substituents are present, their positions are indicated by numbering the carbon atoms in the benzene ring. The substituent that gives the compound its base name is given the lowest possible number. For example:1,3,5-trinitrobenzene (three nitro groups at positions 1, 3, and 5).



1,3,5-trinitrobenzene

5. **Complex Substituents**: When the benzene ring is a substituent on a larger molecule, it is referred to as a "phenyl" group. For example: Phenylalanine (an amino acid with a phenyl group). Example:

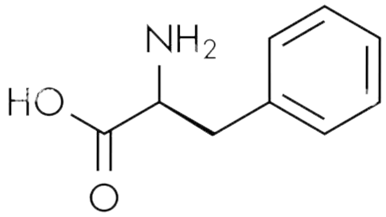


Figure 8.1: *Structure of phenylalanine*

Properties of benzene

Physical Properties of Benzene

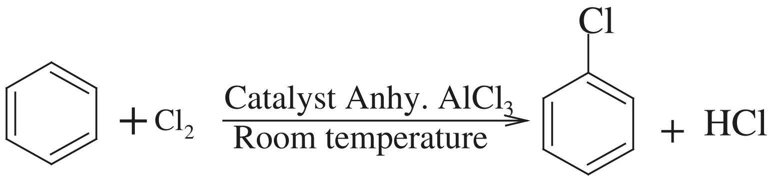
1. Benzene is a clear, colourless liquid. It has a distinctive sweet odour, which is quite noticeable even at low concentrations.
2. The density of benzene is around 0.8765 g/cm³ at 20°C. This means it is less dense than water, which is why benzene floats on water.
3. Benzene has a melting point of 5.53°C, meaning it solidifies just above the freezing point of water. Its boiling point is 80.1°C, which is relatively low compared to many other organic compounds, making it easy to vapourise.
4. Benzene is not soluble in water due to its non-polar nature. However, it is highly soluble in organic solvents like alcohol, ether, chloroform, and acetone. This makes it useful in various chemical reactions and industrial processes.
5. Benzene is highly flammable and burns with a sooty flame due to incomplete combustion. This property requires careful handling and storage to prevent fire hazards.

6. The refractive index of benzene is 1.5011 at 20°C. This property is important in optical applications and helps in identifying the purity of the substance.

Chemical properties of Benzene

Benzene primarily undergoes electrophilic substitution reactions rather than addition reactions. This is because addition reactions would disrupt its stable aromatic ring structure. Common electrophilic substitution reactions include halogenation, nitration, sulfonation, and Friedel-Crafts alkylation and acylation.

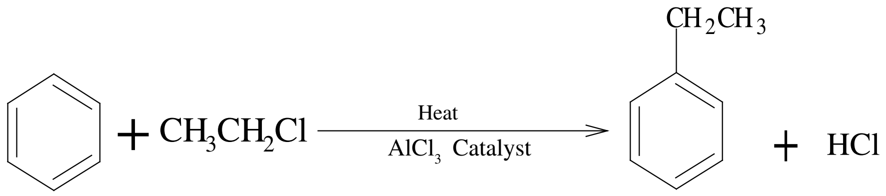
1. Halogenation reaction: Benzene reacts with halogens in the presence of a metal halide carrier iron(III) bromide ($FeBr_3$) or aluminium chloride ($AlCl_3$). The metal halide carrier acts as catalyst and creates the electrophile, X^+, where X is a halogen atom.



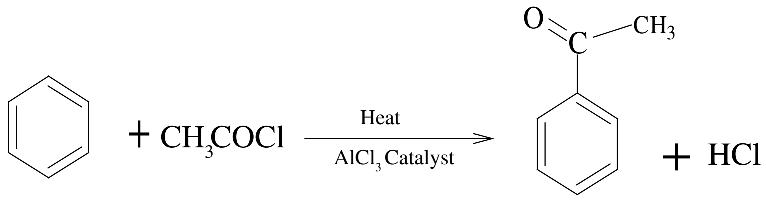
2. Friedel-Crafts reactions: Friedel-Crafts reactions are used to substitute a hydrogen atom in the benzene ring for an alkyl group (Friedel-Crafts alkylation) or an acyl group (Friedel-Crafts acylation).

   $AlCl_3 + Cl_2 \rightarrow AlCl_4^- + Cl^+$ and $FeBr_3 + Br_2 \rightarrow FeBr_4^- + Br^+$

   a. Friedel-Crafts alkylation

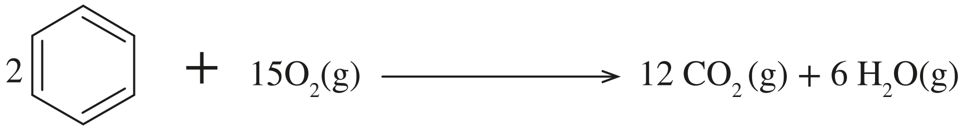


   b. Friedel-Crafts acylation

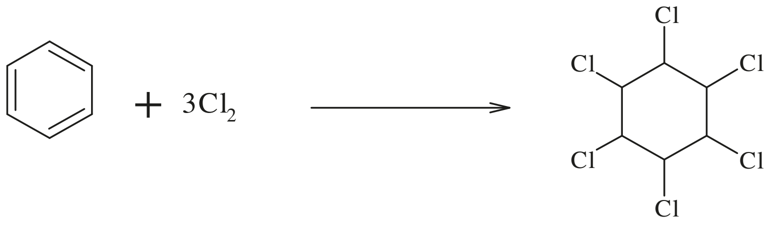


3. **Nitration:** Benzene reacts with a mixture of concentrated nitric acid (HNO_3) and concentrated sulphuric acid (H_2SO_4) at a temperature between 25 and 60 °C.

4. **Reaction with Oxygen:** Hydrocarbons burn in air or oxygen to produce carbon dioxide and water provided sufficient oxygen is available. Benzene also follows this pattern but burns with very smoky flame due to the relatively high carbon content.

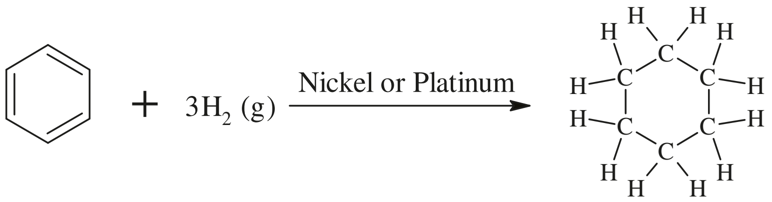


5. **Halogenation reaction:** bromine reacts with chlorine in the presence of sunlight to form benzene hexachloride.



Benzene hexachloride

6. **Hydrogenation reaction**



1. Solvent
   a. Industrial Applications: Benzene is used as a solvent in the production of rubber, lubricants, dyes, detergents, drugs, and pesticides. Its ability to dissolve other substances makes it valuable in these processes.
   b. Laboratory Use: In research and development, benzene is often used as a solvent for chemical reactions and extractions.
2. Chemical Production
   a. Plastics and Resins: Benzene is a key starting material in the production of styrene, which is then polymerized to make polystyrene, a common plastic. It's also used to produce phenol, which is further processed into epoxy resins.
   b. Synthetic Fibers: Benzene is used to produce nylon and other synthetic fibers, which are essential in the textile industry.
3. Printing Industry: Benzene is found in some printing inks and cleaning agents used to maintain printing equipment. Its solvent properties help in achieving the desired consistency and quality of the ink.

4. Fuel Additive: Benzene is added to gasoline to increase its octane rating, which helps in improving engine performance and reducing knocking. However, due to its toxicity, its use in fuels is regulated and often limited.
5. Pharmaceuticals: Benzene is used as an intermediate in the synthesis of various drugs. Its chemical structure allows for the creation of complex molecules needed in pharmaceuticals.
6. Other Uses
    a. Benzene is used in the manufacture of certain explosives, where its chemical properties are harnessed for controlled reactions.
    b. Some pesticides are synthesised using benzene as a starting material, leveraging its reactivity to create effective pest control agents.

Learning Tasks

1. Draw the chemical structure of benzene and label its components.
2. Write balanced chemical equations for the nitration, sulfonation, and halogenation of benzene.
3. Create a detailed report on the various industrial uses of benzene, including its role in the production of plastics and synthetic fibers.

PEDAGOGICAL EXEMPLARS

1. Collaborative learning approach

- **a.** Begin with a brief lecture or video on benzene's discovery and its importance in organic chemistry. Provide visual aids and simplified explanations for learners who may need additional support.
- **b.** Divide the class into small groups, ensuring a mix of abilities in each group. Assign roles within groups (e.g., leader, note-taker, presenter) to ensure active participation from all learners.
- **c.** Each group researches and discusses the Kekulé structure, focusing on alternating single and double bonds. Provide additional resources or guided questions for groups that may need more help understanding the concept.
- **d.** Groups create diagrams showing the resonance structures of benzene and discuss how these structures contribute to its stability. Use colour-coded diagrams and step-by-step guides for learners who need more visual support.
- **e.** Groups draw the molecular orbital diagram for benzene and explain the concept of delocalized electrons.
- **f.** Provide molecular model kits or interactive simulations for hands-on learning.

g. Groups present their findings to the class, highlighting the strengths and weaknesses of each theory. Allow learners to present in various formats (e.g., posters, slideshows, verbal presentations) based on their strengths.

h. Provide problem sets or case studies related to benzene's structure and stability for groups to solve. Offer tiered problem sets with varying levels of difficulty to cater to different learning needs.

i. Conduct a class discussion or have learners write reflections on what they learned and how they worked as a group. Use reflective prompts or sentence starters to help learners articulate their thoughts.

2. Digital Learning Approach

a. Begin with an introductory video explaining substitution reactions and their significance in organic chemistry. Provide subtitles or transcripts for students who may need additional support.

b. Use a virtual laboratory platform where students can simulate reactions of benzene with halogens, concentrated HNO_3, concentrated H_2SO_4, alkyl halides, and acyl halides. Offer guided tutorials within the virtual lab for students who need step-by-step instructions.

c. Have learners watch videos or animations showing different substitution reactions and note the patterns they observe. Provide graphic organisers or templates to help learners organise their observations.

d. In small groups, learners discuss their findings from the virtual lab and videos, focusing on the mechanisms and outcomes of each reaction.

e. Facilitate discussions with guiding questions and provide additional resources for deeper understanding.

f. Use interactive quizzes or online assessments to test learners' knowledge of the reactions and their mechanisms. Offer quizzes with varying levels of difficulty.

g. Have learners write reflections or participate in a class discussion about what they learned through the digital tools and how it helped their understanding. Use reflective prompts or sentence starters to assist learners in articulating their thoughts.

3. Inquiry-Based Learning Integration

a. Have learners watch videos or read articles about the addition reactions of benzene with hydrogen and halogens. For example, benzene reacts with hydrogen in the presence of a nickel or platinum catalyst to form cyclohexane. Benzene also reacts with halogens like chlorine or bromine in the presence of sunlight to form hexahalocyclohexanes.

b. If possible, use a virtual lab to simulate these reactions. Provide guided questions and additional resources for learners who need more support.

c. learners create concept maps to visually organise and summarise the reactions of benzene, including both addition and substitution reactions. They can use online tools like chemsketch, GoConqr or draw them by hand.

d. Encourage learners to work in groups to discuss and refine their concept maps. Provide templates or examples of concept maps for learners who need more guidance.

e. learners research the uses of benzene in everyday life and present their findings. Benzene is used in the production of plastics, synthetic fibers, rubber, dyes, detergents, and more. Allow learners to present their findings in various formats (e.g., slideshows, posters, written reports) based on their strengths.

KEY ASSESSMENTS

Level 1: What type of reactions does benzene typically undergo?

Level 2: Describe the conditions required for the halogenation of benzene.

Level 3

1. Compare and contrast the reactions of benzene with those of alkenes.

2. Predict the product of the Friedel-Crafts alkylation of benzene with ethyl chloride.

Level 4: Investigate the historical development of the benzene structure and its impact on modern chemistry.

FOCAL AREA 2: ALKANOLS

Introduction and structure of Alkanols

Alkanols are characterised by a hydroxyl (-OH) group attached to an sp^3-hybridised carbon atom. They are viewed structurally in two ways, for example, alkanols are viewed as alkane molecules in which one of the hydrogen atoms has been replaced with a hydroxyl group or as water molecule in which one hydrogen atom has been replaced by an ethyl group. They form a homologous series of compounds with the general formula $C_nH_{2n+1}OH$ for where $n \geq 1$.

when $n = 1$, CH_3OH (methanol)

when $n = 2$, C_2H_5OH (ethanol), etc.

Nomenclature of Alkanols

The systematic naming of alkanols follows the International Union of Pure and Applied Chemistry (IUPAC) rules, which ensure consistency and clarity in chemical nomenclature.

Basic Principles of IUPAC Nomenclature for Alkanols

1. Identify the Longest Carbon Chain: The first step in naming an alkanol is to identify the longest continuous carbon chain that contains the hydroxyl group. This chain is considered the parent hydrocarbon. The name of the parent hydrocarbon is derived from the corresponding alkane by replacing the suffix '-ane' with '-anol'.
2. Number the Carbon Chain: The carbon atoms in the parent chain are numbered starting from the end nearest to the hydroxyl group. This ensures that the hydroxyl group receives the lowest possible number. The position of the hydroxyl group is indicated by this number.
3. Name and Number Substituents: Any substituents (alkyl groups, halogens, etc.) attached to the carbon chain are named and numbered according to their positions on the chain. The substituents are listed in alphabetical order in the final name.
4. Combine the Elements: The final name of the alkanol is constructed by combining the position number of the hydroxyl group, the name of the parent hydrocarbon, and the names and positions of any substituents. The suffix '-anol' is used to indicate the presence of the hydroxyl group.

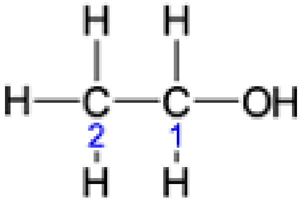
Examples

Give the IUPAC name for each compound

1. CH_3CH_2OH
2. $CH_3CH(OH)CH_3$
3. $CH_3C(OH)(CH_3)CH_3$
4. $(CH_3)_2C(OH)CH_2CH_2CH(Br)CH_3$

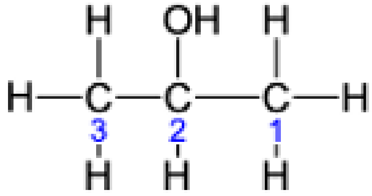
Solution

1. CH_3CH_2OH



   a. Parent hydrocarbon: Ethane
   b. Replace '-ane' with '-anol': Ethanol
   c. The hydroxyl group is on the first carbon, so no number is needed.

2. $CH_3CH(OH)CH_3$

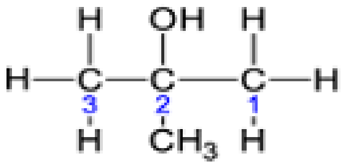


a. Parent hydrocarbon: Propane

b. Number the chain: The hydroxyl group is on the second carbon.

c. Replace '-ane' with '-anol' and indicate the position: Propan-2-ol.

3. $CH_3C(OH)(CH_3)CH_3$



a. Parent hydrocarbon: Propane

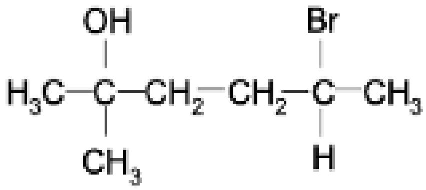
b. Identify and number the longest chain: The hydroxyl group is on the second carbon.

c. Name the substituent: A methyl group is attached to the second carbon.

d. Replace '-ane' with '-anol' and indicate the position: Propan-2-ol.

e. Combine the name of the substituent and the parent name as one name: 2-Methylpropan-2-ol.

4. $(CH_3)_2C(OH)CH_2CH_2CH(Br)CH_3$



a. Parent hydrocarbon: Hexane

b. Identify and number the longest chain: The hydroxyl group is on the second carbon.

c. Name the substituents: A methyl group is attached to the second carbon and a bromine on the fifth carbon.

d. Replace '-ane' with '-anol' and indicate the position: Hexan-2-ol.

f. Combine the names of the substituent in alphabetical order and the parent name as one name: 5-bromo-2-methylhexan-2-ol

Naming of Polyhydroxy Alkanols

When an alkanol contains more than one hydroxyl group, the suffixes '-diol,' '-triol,' etc., are used to indicate the number of hydroxyl groups. The positions of the hydroxyl groups are indicated by numbers.

Example:

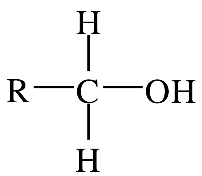
1. 1,2-ethanediol $HOCH_2CH_2OH$
   - **a.** Parent hydrocarbon: Ethane
   - **b.** Two hydroxyl groups: '-diol'
   - **c.** Indicate positions: 1,2-ethanediol.
2. Propan-1,2,3-triol $CH_2(OH)CH(OH)CH_2(OH)$
   - **a.** Parent hydrocarbon: Propane
   - **b.** Three hydroxyl groups: '-triol'
   - **c.** Indicate positions: 1,2,3-propantriol.

Classification of Alkanols

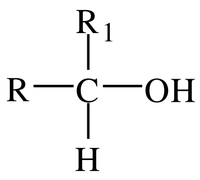
Alkanols can be classified based on the degree of substitution of the carbon atom to which the hydroxyl group is directly attached. This classification helps in understanding their chemical behaviour and reactivity. The three main types of alkanols are primary, secondary, and tertiary alkanols.

1. Primary (1°) Alkanols: In primary alkanols, the carbon atom bearing the hydroxyl group is attached to only one other carbon atom or none (in the case of methanol). This means the hydroxyl group is at the end of the carbon chain. Primary alkanols are generally more reactive in oxidation reactions compared to secondary and tertiary alkanols. Examples:
   - **a.** Methanol (CH_3OH): The simplest primary alkanol with no other carbon atoms attached to the carbon bearing the hydroxyl group.
   - **b.** Ethanol (CH_3CH_2OH): The hydroxyl group is attached to a carbon atom bonded to one other carbon atom.
2. Secondary (2°) Alkanols: In secondary alkanols, the carbon atom with the hydroxyl group is attached to two other carbon atoms. This positioning makes secondary alkanols less reactive in oxidation reactions compared to primary alkanols but more reactive than tertiary alkanols. Examples:
   - **a.** Propan-2-ol ($CH_3CH(OH)CH_3$): The hydroxyl group is attached to the second carbon in a three-carbon chain, which is bonded to two other carbon atoms.
   - **b.** Butan-2-ol ($CH_3CH_2CH(OH)CH_3$): The hydroxyl group is on the second carbon of a four-carbon chain.

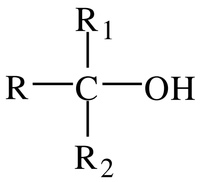
3. Tertiary (3°) Alkanols: In tertiary alkanols, the carbon atom bearing the hydroxyl group is attached to three other carbon atoms. Tertiary alkanols are the least reactive in oxidation reactions due to the absence of hydrogen atoms attached to the carbon bearing the hydroxyl group. Examples:
    - **a.** Tert-butanol ($(CH_3)_3COH$): The hydroxyl group is attached to a carbon atom bonded to three other carbon atoms.
    - **b.** 2-Methylpropan-2-ol ($(CH_3)_2C(OH)CH_3$): The hydroxyl group is on a carbon atom bonded to three other carbon atoms, including a methyl group.



primary (1^0)



secondary (2^0)



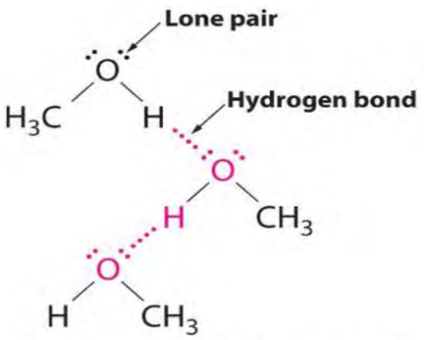
tertiary (3^0)

where R = alkyl group

Properties of Alkanols

1. Physical Properties: Alkanols, exhibit a range of physical properties that are significantly influenced by the presence of the hydroxyl (-OH) group. These properties include boiling points, solubility, density, viscosity, and odor. Understanding these properties is important for understanding the behaviour of alkanols in various chemical and industrial processes.
    - **a.** Boiling Points: The boiling points of alkanols are generally higher than those of alkanes and ethers of similar molecular masses. This is primarily due to the presence of hydrogen bonding between the hydroxyl groups of alkanol molecules. Hydrogen bonds are relatively strong intermolecular forces that require more energy to break, resulting in higher boiling points.

      The hydroxyl group in alkanols can form hydrogen bonds with other alkanol molecules. For example, in methanol (CH_3OH), the hydrogen atom of one molecule can form a hydrogen bond with the oxygen atom of another molecule.



Hydrogen bonding in methanol

Figure 8.2: *hydrogen bonding in methanol*

Methanol (CH_3OH) has a boiling point of 65°C, whereas methane (CH_4), with a similar molecular mass, has a boiling point of -164°C. This significant difference highlights the impact of hydrogen bonding.

b. Solubility: The solubility of alkanols in water is another important physical property. Alkanols with shorter carbon chains are highly soluble in water due to their ability to form hydrogen bonds with water molecules. As the carbon chain length increases in alkanols, their solubility in water decreases. This trend can be attributed to the increasing hydrophobic character of the alkyl chain, which reduces the overall ability of the molecule to interact with water molecules through hydrogen bonding.

Short-chain alkanols example methanol, ethanol (CH_3CH_2OH), and propanol ($CH_3CH_2CH_2OH$) are completely miscible with water whereas long-chain alkanols with longer carbon chains, such as octanol ($C_8H_{17}OH$), have limited solubility in water due to the increasing hydrophobic character of the alkyl group.

c. Viscosity: Viscosity is a measure of a fluid's resistance to flow. Viscosity of alkanols is influenced by the molecular structure, particularly the length of the carbon chain and the presence of the hydroxyl (-OH) group. As we move down the homologous series of alkanols, the viscosity generally increases. This trend can be attributed to several factors, including molecular weight and intermolecular forces.

i. Molecular Weight: As the molecular weight of alkanols increases, the size and mass of the molecules also increase. Larger molecules have more surface area, leading to greater intermolecular interactions, which in turn increases viscosity.

ii. Intermolecular Forces: Alkanols exhibit hydrogen bonding due to the presence of the hydroxyl group. Hydrogen bonds are strong intermolecular forces that significantly contribute to the viscosity of alkanols. As the carbon chain lengthens, the number of hydrogen bonds and van der Waals forces increases, leading to higher viscosity.

d. Odour: Alkanols exhibit distinctive odours that vary depending on their molecular structure. Methanol has a sweet, pungent odour, ethanol is characterised by its familiar alcoholic smell, often associated with beverages, isopropanol (isopropyl alcohol) has a sharp, medicinal odour. Higher alkanols (those with 4 to 10 carbon atoms) tend to have heavier, fruity odors and are somewhat viscous or oily. These odours are due to the presence of the hydroxyl group (-OH) attached to the carbon chain, which influences the volatility and sensory perception of these compounds. The variation in odour among different alkanols can be attributed to the differences in their molecular size and structure.

2. **Chemical Properties of Alkanols:** Alkanols exhibit a variety of chemical properties due to the presence of the hydroxyl (-OH) group. This functional group significantly influences their reactivity and interactions with other substances. Some of the chemical properties of alkanols include their reactions with oxidising agents, acids, bases, and other compounds.

Acidity of Alkanols

Alkanols, or alcohols, exhibit weak acidic properties due to the presence of the hydroxyl group (-OH). The acidity of an alkanol is determined by its ability to donate a proton (H^+) from the hydroxyl group, forming an alkoxide ion (RO^-).

Factors Influencing Acidity

a. Polarity of the O-H Bond: The O-H bond in alkanols is polar, with the oxygen atom being more electronegative than hydrogen. This polarity facilitates the release of a proton, making the alkanol weakly acidic.

b. Inductive Effects: Electron-withdrawing groups (e.g., halogens) attached to the carbon chain increase the acidity of alkanols by stabilising the negative charge on the alkoxide ion through inductive effects. Conversely, electron-donating groups (e.g., alkyl groups) decrease acidity by increasing electron density on the oxygen atom, making it less likely to release a proton.

c. Resonance Stabilisation: Alkanols that are in conjugation with a pi bond or aromatic ring (e.g., phenol) are more acidic because the conjugate base is resonance-stabilised. This stabilisation makes it easier for the alkanol to lose a proton. Phenol, an aromatic alcohol, has a pKa of about 10, making it significantly more acidic than aliphatic alkanols due to resonance stabilization.

The order of acidity for alkanols generally follows:

$$1^\circ \text{ alkanols} < 2^\circ \text{ alkanols} < 3^\circ \text{ alkanols}$$

$$\xrightarrow[\text{Increasing acidity}]{\hspace{6cm}}$$

This is because primary alkanols have fewer electron-donating alkyl groups, making them more acidic than secondary and tertiary alkanols. Also, As the carbon chain length increases, the electron-donating effect of the alkyl groups becomes more pronounced, reducing the acidity of the alkanol. Therefore, methanol (CH_3OH) is more acidic than ethanol (CH_3CH_2OH), which is more acidic than propanol ($CH_3CH_2CH_2OH$), and so on.

Amphoteric Nature of Alkanols

1. Reaction with Bases

   Alkanoic acids react with bases (e.g., sodium hydroxide) to form salts (alkanoates) and water. This is a typical neutralisation reaction. Example: $R\text{-}COOH+NaOH \rightarrow R\text{-}COONa+H_2$

   The carboxyl group donates a proton (H^+) to the hydroxide ion (OH^-), forming water and the corresponding alkanoate salt.

2. Reaction with Carbonates and Bicarbonates

   Alkanoic acids react with carbonates (e.g., sodium carbonate) and bicarbonates (e.g., sodium bicarbonate) to produce carbon dioxide, water, and a salt. The acid reacts with the carbonate or bicarbonate, releasing carbon dioxide gas, which is observed as effervescence.

   Example: $2R\text{-}COOH + Na_2CO_3 \rightarrow 2R\text{-}COONa + H_2O + CO_2$

3. **Reaction with sodium Metal**

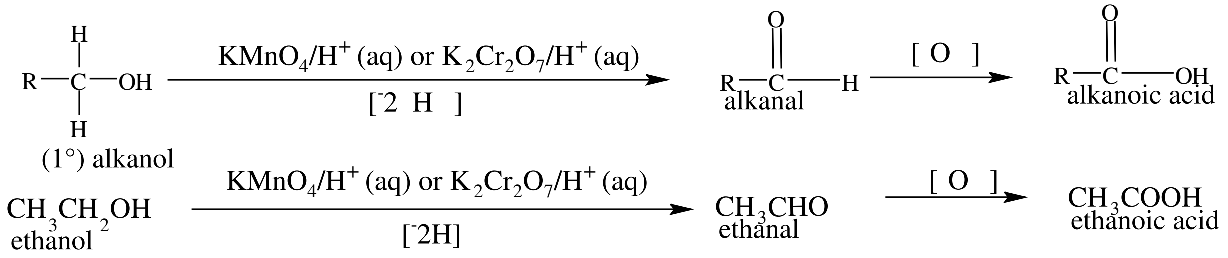
   Alkanols react with sodium metal to produce alkoxide ions and hydrogen gas. This reaction is often used to demonstrate the acidity of alkanols in a laboratory setting. The hydroxyl group donates a proton to the sodium metal, forming an alkoxide ion and releasing hydrogen gas.

   Example: $2ROH + 2Na \rightarrow 2RONa + H_2$

4. Oxidation reactions

The oxidation of alkanols is one of their most important chemical properties. The products of oxidation depend on the type of alkanol (primary, secondary, or tertiary). The oxidation process is typically carried out using oxidising agents such as potassium dichromate ($K_2Cr_2O_7$) or potassium permanganate ($KMnO_4$), often in the presence of an acid like sulphuric acid (H_2SO_4).

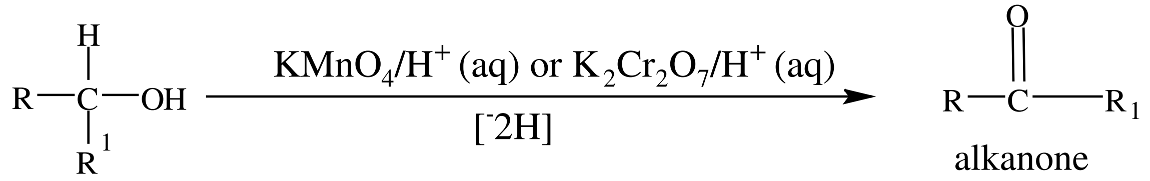
a. Primary alkanols are oxidised to alkanals and further to carboxylic acids. For example, ethanol (a primary alkanol) can be oxidised to ethanal (acetaldehyde) and then to ethanoic acid (acetic acid).



$$R\text{-}CH_2\text{-}OH \xrightarrow[\text{[2 H]}]{KMnO_4/H^+ (aq) \text{ or } K_2Cr_2O_7/H^+ (aq)} R\text{-}C(=O)\text{-}H \xrightarrow{[O]} R\text{-}C(=O)\text{-}OH$$

$$\underset{\text{ethanol}}{CH_3CH_2OH} \xrightarrow[\text{[-2H]}]{KMnO_4/H^+ (aq) \text{ or } K_2Cr_2O_7/H^+ (aq)} \underset{\text{ethanal}}{CH_3CHO} \xrightarrow{[O]} \underset{\text{ethanoic acid}}{CH_3COOH}$$

b. Secondary alkanols are oxidised to alkanones. For instance, propan-2-ol can be oxidised to propanone (acetone).



$$R-\overset{\overset{H}{|}}{\underset{\underset{R^1}{|}}{C}}-OH \xrightarrow[\text{[-2H]}]{KMnO_4/H^+(aq) \text{ or } K_2Cr_2O_7/H^+(aq)} R-\overset{\overset{O}{\|}}{C}-R_1$$

(2°) alkanol alkanone

$$\underset{\text{propan-2-ol}}{CH_3CH(OH)CH_3} \xrightarrow[\text{[-2H]}]{KMnO_4/H^+(aq) \text{ or } K_2Cr_2O_7/H^+(aq)} \underset{\text{propan-2-one}}{CH_3COCH_3}$$

c. Tertiary alkanols do not oxidise easily because there is no hydrogen atom attached to the carbon bearing the hydroxyl group. They generally resist oxidation under mild conditions.

5. Combustion reactions

The combustion of alkanols (alcohols) is a highly exothermic reaction, meaning it releases a significant amount of heat. This property makes alkanols useful as fuels. During combustion, alkanols react with oxygen to produce carbon dioxide and water. The general equation for the complete combustion of an alkanol is:

$$C_nH_{2n+1}OH + \left(n + \tfrac{1}{2}\right)O_2 \rightarrow nCO_2 + (n+1)H_2O$$

In complete combustion, there is sufficient oxygen for the alkanol to fully oxidise, producing carbon dioxide and water. Example, lower alkanols like ethanol burn with a nearly invisible blue flame, which is characteristic of complete combustion. The invisible nature of the flame is what makes alcohol fires particularly dangerous. Its combustion releases a significant amount of heat and this makes it an effective fuel.

Incomplete combustion occurs when there is insufficient oxygen. This results in the formation of carbon monoxide (CO) or even carbon (soot) along with water. Incomplete combustion is less efficient and produces less energy compared to complete combustion.

6. Reaction with alkanoic acids

Alkanols react with alkanoic acids to form akylalkanoates (esters) and water in a process known as esterification. This reaction is an important method for synthesizing esters, which are widely used in fragrances, flavourings, and as solvents. The general reaction can be represented as:

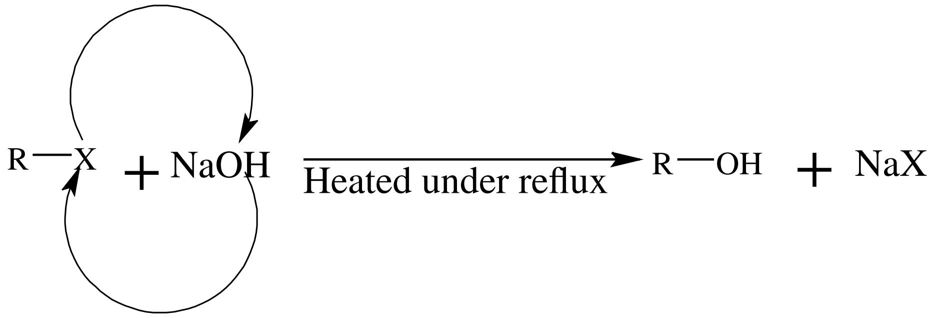
$$\underset{\text{alkanol}}{R-OH} + \underset{\text{alkanoic acid}}{R1-COOH} \xrightarrow[D]{H^+} \underset{\text{alkylalkanoate (ester)}}{R1-COOR} + H_2O$$

$$\underset{\text{ethanol}}{CH_3CH_2OH} + \underset{\text{propanoic acid}}{CH_3CH_2COOH} \xrightarrow[D]{H^+} \underset{\text{propylethanoate (ester)}}{CH_3CH_2COOCH_2CH_3} + H_2O$$

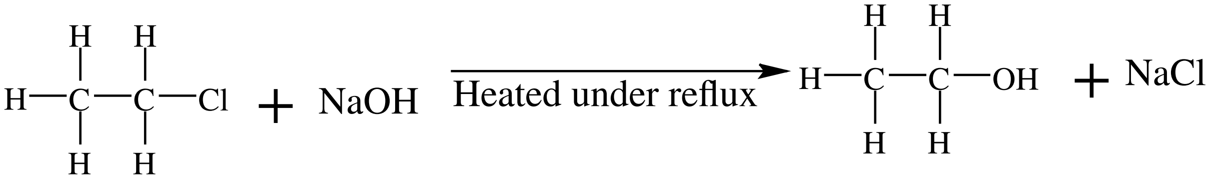
where R and R1 are alkyl groups. The reaction typically requires an acid catalyst, such as concentrated sulphuric acid (H_2SO_4), to proceed efficiently. The acid catalyst not only speeds up the reaction but also helps in removing the water produced, driving the equilibrium towards the formation of the ester.

Preparation of Alkanols in the Laboratory

1. Hydrolysis of Halogenoalkanes: Hydrolysis of halogenoalkanes is a common method for preparing alkanols. Halogenoalkanes, also known as alkyl halides, are compounds where one or more hydrogen atoms in an alkane are replaced by halogen atoms (e.g., chlorine, bromine). The halogenoalkane is heated under reflux with the aqueous alkali (Aqueous sodium hydroxide (NaOH) or potassium hydroxide (KOH)). This is a nucleophilic substitution reaction where the hydroxide ion (OH^-) replaces the halogen atom in the halogenoalkane. The reaction mixture is heated under reflux to ensure complete reaction. The product, an alkanol, is then distilled off. Example:



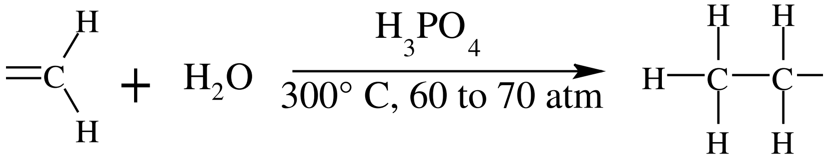
$$R{-}X + NaOH \xrightarrow{\text{Heated under reflux}} R{-}OH + NaX$$



Chloroethane

Ethanol

2. Hydration of Alkenes: Hydration of alkenes is another important method for preparing alkanols. This involves the addition of water to an alkene in the presence of an acid catalyst. The alkene is made to reacted with water in the presence of the acid catalyst (e.g., phosphoric acid, H_3PO_4). The reaction proceeds via the formation of a carbocation intermediate, followed by the addition of a water molecule. Example:



3. Fermentation of Sugars: Fermentation is a biological process used to produce ethanol from sugars. This method is widely used in the production of alcoholic beverages and biofuels. The sugars are converted to ethanol and carbon dioxide by the action

of yeast. The enzyme (zymase) in the yeast catalyze the breakdown of sugars to ethanol and carbon dioxide according to the reaction:

$$C_6H_{12}O_6 \xrightarrow{\text{zymase}} 2\ CH_3CH_2OH\ (aq) + 2\ CO_2(g) + \text{Heat}$$

After fermentation, the mixture contains ethanol, water, and other by-products. The ethanol is separated from this mixture through distillation. The distillation process involves heating the mixture to vapourise the ethanol, which is then condensed back into liquid form.

Test for alkanols

Testing for alkanols (alcohols) involves several chemical reactions that can identify the presence and type of alcohol. Below are some common tests:

1. **Lucas Test for Alkanols:** The Lucas test is a qualitative test used to differentiate primary, secondary, and tertiary alcohols based on their reactivity with Lucas reagent, which is a mixture of anhydrous zinc chloride ($ZnCl_2$) in concentrated hydrochloric acid (HCl). The test is based on the formation of alkyl chlorides, which are insoluble in the aqueous solution, leading to turbidity.

   Procedure:

   a. Add 2-3 mL of Lucas reagent to a test tube.

   b. Add 1-2 drops of the alcohol to be tested.

   c. Shake the test tube and observe the time taken for the solution to turn turbid.

   Observations:

   i. Primary Alcohols: $R\text{-}CH_2OH + HCl \rightarrow R\text{-}CH_2Cl + H_2O$. Primary alcohols react very slowly or not at all at room temperature, showing no immediate turbidity.

   ii. Secondary Alcohols: $R_2CH\text{-}OH + HCl \rightarrow R_2CH\text{-}Cl + H_2O$. Secondary alcohols react within 5-10 minutes, forming turbidity.

   iii. Tertiary Alcohols: $R_3C\text{-}OH + HCl \rightarrow R_3C\text{-}Cl + H_2O$. Tertiary alcohols react immediately, forming turbidity almost instantly.

   Lucas test is particularly useful for distinguishing tertiary alcohols from primary and secondary alcohols due to the rapid formation of turbidity.

2. **Iodoform Test for Alkanols:** The Iodoform test is used to identify alcohols with the structure $R\text{-}CH(OH)\text{-}CH_3$, which includes ethanol and secondary alcohols with a methyl group adjacent to the hydroxyl group. The test involves the reaction of the alcohol with iodine (I_2) and sodium hydroxide (NaOH), leading to the formation of a yellow precipitate of iodoform (CHI_3).

Procedure

- **a.** Add a few drops of the alcohol to be tested into a test tube.
- **b.** Add a few drops of iodine solution.
- **c.** Add sodium hydroxide solution dropwise until the iodine colour is discharged.
- **d.** Warm the mixture gently if necessary.

Observations

$\text{R-CH(OH)-CH}_3 + 3\text{I}_2/4\text{NaOH} \rightarrow \text{CHI}_3 + \text{R-COONa} + 3\text{NaI} + 3\text{H}_2\text{O}$.

A positive result is indicated by the formation of a yellow precipitate of iodoform (CHI_3), which has a characteristic antiseptic smell. The Iodoform test is specific for alcohols that have a methyl group adjacent to the hydroxyl group, making it a valuable test for identifying ethanol and certain secondary alcohols.

3. **Ester Test**
   - **a.** Take 1ml of the organic liquid to be tested in a clean dry test tube.
   - **b.** Add 1ml of glacial acetic acid and 2-3 drops of conc. sulphuric acid.
   - **c.** Heat the mixture in a water bath for 10 minutes.
   - **d.** The hot mixture is poured into a beaker containing cold water.
   - **e.** Smell the water in the beaker.
   - **f.** A fruity smell confirms the presence of alcoholic group.

Everyday Uses of Alkanols

1. Household Uses
   - **a.** Cleaning Agents: Alkanols like ethanol and isopropanol are used in household cleaning products due to their ability to dissolve grease and stains. They are effective in disinfecting surfaces and are commonly found in window cleaners and multi-surface cleaners.
   - **b.** Personal Care Products: Ethanol is a key ingredient in hand sanitizers, perfumes, and hairsprays. Its antimicrobial properties make it effective in killing bacteria and viruses, ensuring hygiene and freshness.
2. Medical and Pharmaceutical Uses
   - **a.** Antiseptics and Disinfectants: Ethanol and isopropanol are widely used in medical settings as antiseptics and disinfectants. They are used to clean wounds, disinfect skin before injections, and sterilize medical instruments.
   - **b.** Pharmaceutical Solvents: Alkanols are used as solvents in the pharmaceutical industry to dissolve and extract active ingredients from plant materials and other

sources. This makes it easier to formulate medications in various forms, such as pills, tinctures, and extracts.

c. Preservatives: Ethanol acts as a preservative in pharmaceutical products, ensuring their stability and longevity.

3. Industrial Uses

a. Solvents: Alkanols are used as solvents in the manufacture of paints, varnishes, and coatings. They help dissolve other substances and ensure a smooth application.

b. Fuel Additives: Ethanol is used as a fuel additive in gasoline to improve combustion efficiency and reduce emissions. It is also used as a renewable fuel source in the form of bioethanol.

c. Chemical Synthesis: Alkanols are used as intermediates in the synthesis of various chemicals, including esters, ethers, and aldehydes.

4. Food and Beverage Industry

a. Alcoholic Beverages: Ethanol is the primary alcohol found in alcoholic beverages such as beer, wine, and spirits. It is produced through the fermentation of sugars by yeast.

b. Flavouring Extracts: Alkanols are used to extract flavors from natural sources, which are then used in food products and beverages.

5. Laboratory and Research Uses

a. Solvents for Reactions: Alkanols are commonly used as solvents in chemical reactions and processes in research laboratories. They are particularly useful for dissolving organic compounds.

b. Cleaning and Sterilization: Ethanol is used to clean and sterilize laboratory equipment due to its effectiveness in killing microorganisms and its quick evaporation rate.

c. Sample Preparation: Alkanols are used in the preparation of samples for analysis, as they can dissolve many organic compounds that are insoluble in water.

d. Precipitation of Biomolecules: Ethanol is used to precipitate DNA and RNA, making it easier to isolate and study these molecules in biochemical research.

6. Automotive Uses

a. Antifreeze: Methanol and ethanol are used as antifreeze agents in automobile radiators to lower the freezing point of the coolant and prevent it from freezing in cold temperatures.

- **b.** Fuel: Ethanol is used as a fuel in racing cars and rockets due to its high energy content and clean-burning properties.

Learning Tasks

1. List at least three common uses of ethanol in everyday life. Include pictures or diagrams to illustrate each use.
2. Practice naming simple alkanols (up to 4 carbon atoms) using IUPAC rules. Provide both the structural formula and the name.
3. Write a short essay on the differences between primary, secondary, and tertiary alkanols, including examples and their structural differences.
4. Write a report on the oxidation reactions of primary, secondary, and tertiary alkanols. Include balanced chemical equations and explain the products formed.
5. Design an experiment to demonstrate the esterification reaction of an alkanol with a carboxylic acid. Conduct the experiment, record your observations, and present your findings to the class.

PEDAGOGICAL EXEMPLARS

1. Collaborative Learning Approach

- **a.** Have learners individually think about and write down the general formula for alkanols and the specific formulae for methanol (CH_3OH), ethanol (C_2H_5OH), propanol (C_3H_7OH), butanol (C_4H_9OH), pentanol ($C_5H_{11}OH$), and hexanol ($C_6H_{13}OH$).
- **b.** Pair learners to discuss and compare their formulae.
- **c.** Pairs share their results with the class, ensuring everyone has the correct formulae.
- **d.** Have learners individually think about the classification of alkanols into primary, secondary, and tertiary based on the carbon atom bonded to the hydroxyl group.
- **e.** In pairs, let learners discuss examples of each class and draw structural isomers for alkanols up to C6.
- **f.** Pairs present their findings, highlighting the differences between primary, secondary, and tertiary alkanols, and the various isomers.
- **g.** Ask learners to individually review the IUPAC rules for naming alkanols, such as identifying the longest carbon chain and numbering the chain to give the hydroxyl group the lowest possible number.
- **h.** In pairs, have learners practice naming various alkanols and their isomers, such as 1-butanol, 2-butanol, 2-methyl-1-propanol, etc.

   - **i.** Pairs share their names with the class, and the teacher provides feedback and corrections as needed.

2. **Experiential Learning**
   - **a.** Introduce the preparation of alkanols, including fermentation and hydration methods, with visual aids and simplified explanations. For beginners focus on basic concepts, intermediates take detailed notes, advanced learners analyse processes.
   - **b.** Divide the class into mixed-ability groups, assigning roles (leader, note-taker, presenter) to ensure active participation.
   - **c.** Groups research and discuss ethanol preparation by fermentation and hydration. For beginners Summarise steps, intermediates compare processes, advanced learners analyse efficiency and impact.
   - **d.** Groups compare and contrast fermentation and hydration methods.
   - **e.** Groups discuss physical properties of alkanols (boiling points, solubility, density).
   - **f.** Groups present findings, using molecular model kits or simulations for hands-on learning. For beginners present summaries, intermediates provide explanations, advanced learners lead discussions and analyses.

3. **Activity-based Learning**
   - **a.** Introduce the preparation of alkanols, including fermentation and hydration methods, with visual aids and simplified explanations.
   - **b.** Divide the class into mixed-ability groups, assigning roles (leader, note-taker, presenter) to ensure active participation.
   - **c.** Groups research and discuss ethanol preparation by fermentation and hydration.
   - **d.** Groups compare and contrast fermentation and hydration methods.
   - **e.** Groups discuss physical properties of alkanols (boiling points, solubility, density).
   - **f.** Groups present findings, using molecular model kits or simulations for hands-on learning.

4. **Project-based learning**
   - **a.** Begin with a brief lecture or video on alcohol breath analysers and the role of potassium dichromate in detecting alcohol.
   - **b.** Divide the class into small mixed-ability groups, assigning roles (leader, note-taker, presenter) to ensure active participation.

c. Groups research how alcohol breath analysers work, focusing on the redox reaction involving potassium dichromate.

d. Groups plan and design an alcohol breath analyser using potassium dichromate.

e. Groups create a prototype of their alcohol breath analyser.

f. Groups research the advantages and disadvantages of biofuels, focusing on energy efficiency

g. Groups evaluate the energy efficiency of biofuels and discuss their findings.

h. Groups present their alcohol breath analyser designs and biofuel evaluations to the class.

i. Conduct a class discussion or have learners write reflections on their projects and group work.

KEY ASSESSMENTS

Level 1: What is the general formula for alkanols? Provide the formula and an example.

Level 2

1. Compare and contrast the physical properties of methanol (CH_3OH) and ethanol (C_2H_5OH). Discuss their boiling points and solubility in water.
2. Analyse the mechanism of dehydration of a secondary alcohol under acidic conditions. What is the primary factor that determines whether an elimination or substitution reaction occurs?

   A. The temperature of the reaction.

   B. The strength of the acid used.

   C. The stability of the carbocation intermediate formed during the reaction.

   D. The concentration of the alcohol.

Level 3: Explain the process of esterification involving ethanol and acetic acid. Write the balanced chemical equation for the reaction and describe the conditions required for the reaction to occur.

Level 4: Design an experiment to demonstrate the oxidation of a primary alkanol (e.g., ethanol) to an aldehyde and then to a carboxylic acid. Include the materials, procedure, and safety precautions. Analyse the results and explain the chemical changes that occur at each step.

Hint

The recommended mode of assessment for week 22 is **multiple choice questions.** Use the level 2 question 2 as a sample question.

WEEK 24

Learning Indicator: *Explain the structure, nomenclature, preparation, properties and uses of alkanoic acids as well as their characteristic test*

FOCAL AREA: ALKANOIC ACIDS

Introduction to Alkanoic Acids

Alkanoic acids, also known as carboxylic acids, are a class of organic compounds characterised by the presence of a carboxyl group (R-COOH), where R is an alkyl group or hydrogen or a phenol group. This functional group consists of a carbonyl group (C=O) and a hydroxyl group (-OH) attached to the same carbon atom. Alkanoic acids are widely found in nature and are known for their acidic properties. They play a crucial role in various biochemical processes and are used in numerous industrial applications.

They form a homologous series of compounds with the general molecular formular $C_n H_{2n+1}COOH$ *for where* $n \geq 0$.

when $n = 0$, HCOOH (methanoic acid)

when $n = 1$, CH_3COOH (ethanoic acid)

when $n = 2$, C_2H_5COOH (propanoic acid)

when $n = 3$ C_3H_7COOH (butanoic acid), etc.

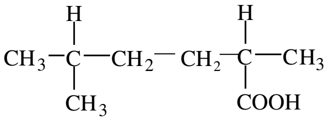
Alkanoic acids are weak acids, meaning they do not completely dissociate in water. This partial dissociation is responsible for their characteristic properties, such as their ability to conduct electricity in solution and their reactivity with bases to form salts and water. The acidity of alkanoic acids is influenced by the presence of electron-withdrawing groups, which stabilise the negative charge on the conjugate base, making the acid stronger.

Nomenclature of Alkanoic Acids

The nomenclature of alkanoic acids follows the International Union of Pure and Applied Chemistry (IUPAC) guidelines. The name of an alkanoic acid is derived from the corresponding alkane by replacing the suffix "-e" with "-oic acid." For example, methane becomes methanoic acid, and ethane becomes ethanoic acid. The carboxyl group is always assigned the highest priority in numbering the carbon chain, and it is always located at the end of the chain, making it carbon number one, example:

$CH_3CH(Cl)CH_2COOH$
3-Chlorobutanoic acid

$CH_3CH=CHCH_2COOH$
3-pentenoic acid

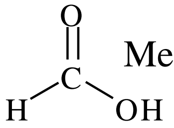


$$CH_3-\overset{H}{\underset{CH_3}{C}}-CH_2-CH_2-\overset{H}{\underset{COOH}{C}}-CH_3$$

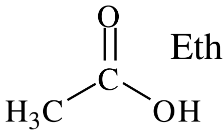
2,5-dimethylhexanoic acid

$Cl-COOH$
Chloromethanoic acid

Common names of alkanoic acids are often derived from their natural sources. For instance, formic acid gets its name from the Latin word “formica,” meaning ant, as it is found in ant venom. Acetic acid derives its name from “acetum,” the Latin word for vinegar. These common names are widely used in industry and everyday language.

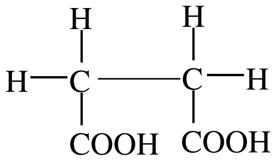


Methanoic acid (formic acid)



Ethanoic acid (acetic acid)

When naming alkanoic acids with multiple carboxyl groups, the suffix “-dioic acid” is used. For example, a compound with two carboxyl groups on an ethane chain is called ethanedioic acid.

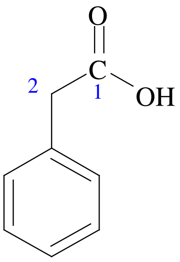


Ethanedioic acid

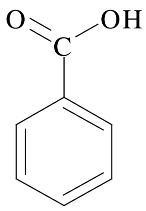
$HOOC-CH_2-COOH$

1,3-propanedioic acid

Additionally, if the alkanoic acid group is attached to a ring structure, the suffix “-alkanoic acid” is used, and the carbon bearing the alkanoic group is designated as carbon number one.



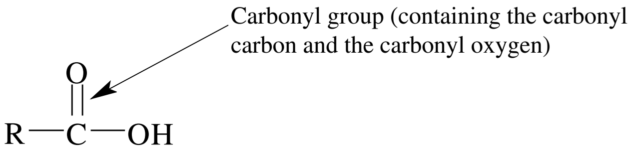
2-phenylethanoic acid



Phenylmethanoic acid (benzoic acid)

Structure of alkanoic acids

The structure of alkanoic acids is defined by the presence of the carboxyl group (-COOH). This group consists of a carbon atom double-bonded to an oxygen atom (C=O) and single-bonded to a hydroxyl group (-OH).



The general formula for alkanoic acids is R-COOH, where R represents an alkyl group, a hydrogen or a phenyl group.

The carboxyl group is planar and exhibits resonance, which stabilises the molecule. The resonance structure involves the delocalization of electrons between the carbonyl and hydroxyl groups, resulting in partial double-bond character between the carbon and oxygen atoms. This resonance contributes to the acidity of alkanoic acids, as it stabilises the conjugate base formed after the loss of a proton.

In terms of molecular geometry, the carbon atom in the carboxyl group is sp^2 hybridised, leading to a trigonal planar arrangement. The bond angles around the carbonyl carbon are approximately 120^0. The hydroxyl group is capable of forming hydrogen bonds, which significantly influence the physical properties of alkanoic acids, such as their boiling points and solubility in water.

Properties of Alkanoic Acids

They exhibit a range of physical properties that are influenced by the length of their carbon chain and the presence of the carboxyl group.

1. **Boiling and Melting Points**: Alkanoic acids generally have higher boiling and melting points compared to other organic compounds of similar molecular weight. This is due to the strong hydrogen bonding between the carboxyl groups of adjacent molecules.

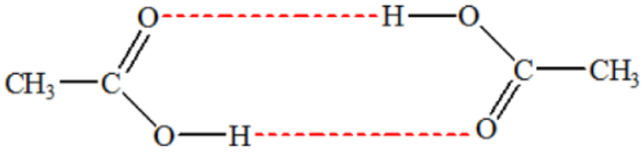


Figure 8.3: *Hydrogen bonding ethanoic acid*

The boiling and melting points increase with the length of the carbon chain due to the increased van der Waals forces.

Alkanoic acids exhibit stronger hydrogen bonding compares to their counterparts, alkanol. Alkanoic acids form dimers through two hydrogen bonds between the hydrogen atom of the hydroxyl group and the oxygen atom of the carbonyl group (C=O) of another molecule. This dual hydrogen bonding results in higher boiling and melting points for alkanoic acids compared to alkanols of similar molar mass. For example, ethanoic acid (acetic acid) has a boiling point of 118°C, which is significantly higher than that of ethanol (78°C) despite having a similar molecular weight.

2. **Solubility**: The solubility of alkanoic acids in water is another notable physical property. Lower members of the series, such as methanoic acid (formic acid) and ethanoic acid, are highly soluble in water due to their ability to form hydrogen bonds with water molecules. However, as the carbon chain length increases, the hydrophobic nature of the alkyl group reduces solubility. For example, butanoic acid (butyric acid) is less soluble in water compared to ethanoic acid.

3. **Density**: Alkanoic acids are generally denser than water. The density of these acids decreases with the length of the carbon chain. For instance, the density of ethanoic acid is about 1.049 g/cm³, while that of hexanoic acid is around 0.927 g/cm³.
4. **Physical State**: At room temperature, the physical state of alkanoic acids varies with the length of the carbon chain. Lower members, such as methanoic acid and ethanoic acid, are liquids, while higher members, such as octanoic acid and decanoic acid, are waxy solids. This change in physical state is due to the increasing van der Waals forces as the carbon chain increases.
5. **Odour**: Alkanoic acids typically have strong, pungent odours. For example, ethanoic acid has a characteristic vinegar-like smell, while butanoic acid has a rancid butter odour. The odour becomes less intense as the molecular weight increases.

Chemical properties of Alkanoic Acids

This functional group is highly reactive and participates in a variety of chemical reactions. The reactivity of alkanoic acids is largely influenced by the carboxyl group's ability to donate a proton (acidic behaviour) and its electrophilic carbonyl carbon, which can undergo nucleophilic attack.

1. **Acid-Base Reactions:** One of the most fundamental reactions of alkanoic acids is their reaction with bases. Alkanoic acids are weak acids, meaning they partially dissociate in water to form hydronium ions (H_3O^+) and carboxylate ions ($R\text{-}COO^-$). When they react with strong bases like sodium hydroxide (NaOH), they form salts and water in a neutralisation reaction. For example, acetic acid reacts with sodium hydroxide to form sodium acetate and water:

   $$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$

   This reaction is important in various industrial processes and in the formation of soaps, where long-chain alkanoic acids react with bases to form soap molecules.

2. **Esterification reactions:** Alkanoic acids react with alcohols in the presence of an acid catalyst to form esters and water. This reaction, known as esterification, is widely used in the production of fragrances, flavorings, and plastics. For example, ethanoic acid reacts with ethanol to form ethyl ethanoate and water:

   $$\underset{\text{ethanoic acid}}{CH_3COOH} + \underset{\text{ethanol}}{CH_3CH_2OH} \xrightarrow[\text{Warm}]{H^+} \underset{\substack{\text{ethylethanoate} \\ \text{(ester)}}}{CH_3COOCH_2CH_3} + \underset{\text{water}}{H_2O}$$

   Esterification is an equilibrium reaction, and the yield of ester can be increased by removing water from the reaction mixture or by using an excess of one of the reactants.

3. **Decarboxylation reactions:** Decarboxylation is the process by which alkanoic acids lose a carbon dioxide molecule (CO_2) to form hydrocarbons. This reaction typically requires heating and sometimes a catalyst. For example, the decarboxylation of ethanoic acid to methane can be represented as:

   $CH_3COOH \rightarrow CH_4+CO_2$

   Decarboxylation is an important reaction in biochemistry, particularly in metabolic pathways where carboxylic acids are converted to smaller molecules.

4. **Formation of Amides:** Alkanoic acids react with amines to form amides in a reaction known as amidation. This reaction typically requires heating and sometimes a dehydrating agent. For example, ethanoic acid reacts with ammonia to form acetamide:

   $CH_3COOH + NH_3 \rightarrow CH_3CONH_2 + H2O$

   Amides are important in the synthesis of pharmaceuticals and polymers.

Test for Alkanoic Acids

1. **Litmus Test**

   Procedure: Dip a piece of blue litmus paper into the sample solution.

   Observation: If the blue litmus paper turns red, it indicates the presence of an acid.

2. **Reaction with Sodium Bicarbonate**

   Procedure: Add a small amount of sodium bicarbonate ($NaHCO_3$) to the sample solution.

   Observation: Effervescence (bubbling) due to the release of carbon dioxide gas (CO_2) indicates the presence of a carboxylic acid.

   Equation of reaction: $RCOOH + NaHCO_3 \rightarrow RCOONa + H_2O + CO_2$

3. **Esterification Test**

   Procedure: Mix the sample with a small amount of ethanol and a few drops of concentrated sulphuric acid (H_2SO_4). Heat the mixture gently.

   Observation: A sweet, fruity smell indicates the formation of an ester, confirming the presence of a carboxylic acid.

   Equation of reaction: $RCOOH + C_2H_5OH \rightarrow RCOOC_2H_5 + H_2O$

4. **Reaction with Magnesium**

   Procedure: Add a small piece of magnesium ribbon to the sample solution.

   Observation: Effervescence due to the release of hydrogen gas (H_2) indicates the presence of a carboxylic acid.

Equation of reaction: $2RCOOH + Mg \rightarrow (RCOO)_2Mg + H_2$

5. **Formation of a Salt**

   Procedure: Add a few drops of the sample to a solution of sodium hydroxide (NaOH).

   Observation: The formation of a salt and water indicates the presence of a carboxylic acid.

   Equation of reaction: $RCOOH + NaOH \rightarrow RCOONa + H_2O$

Preparations of Alkanoic Acids

1. *Oxidation of Primary Alcohols*: Primary alcohols can be oxidised to carboxylic acids using strong oxidising agents like potassium permanganate ($KMnO_4$) or potassium dichromate ($K_2Cr_2O_7$) in the presence of an acid.

   $R\text{-}CH_2OH + 2[O] \rightarrow R\text{-}COOH + H_2O$

   Example: $CH_3CH_2OH + 2[O] \rightarrow CH_3COOH + H_2O$

   Conditions: Acidic medium (dilute sulphuric acid) and heating under reflux

2. *Oxidation of Aldehydes*: Alkanals or aldehydes can be further oxidised to alkanoic acid acids using similar oxidising agents.

   $R\text{-}CHO+[O]\rightarrow R\text{-}COOH$

   Example: $CH_3CHO+[O]\rightarrow CH_3COOH$

   Conditions: Acidic medium and heating under reflux

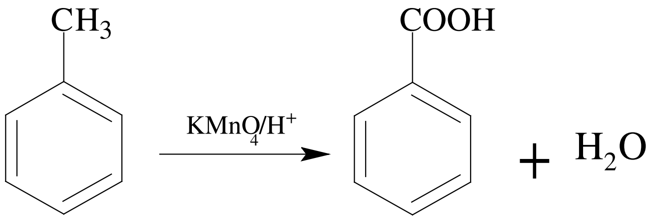
3. *Hydrolysis of Nitriles*: Nitriles can be hydrolyzed to carboxylic acids in the presence of an acid or base.

   Reaction: $R\text{-}CN + 2H_2O \rightarrow R\text{-}COOH + NH_3$

   Example: $CH_3CN + 2H_2O \rightarrow CH_3COOH + NH_3$

   Conditions: Acidic or basic medium and heating

4. *Oxidation of Alkylbenzenes*: Alkylbenzenes can be oxidised to aromatic alkanoic acids using strong oxidising agents like potassium permanganate. Example:



   Conditions: Acidic or basic medium and heating.

5. *Hydrolysis of Esters*: Esters can be hydrolyzed to alkanoic acids in the presence of an acid or base.

General reaction: $R\text{-}COOR^1 + H_2O \rightarrow R\text{-}COOH + R^1\text{-}OH$

Example: $CH_3COOCH_3 + H_2O \rightarrow CH_3COOH + CH_3OH$

Conditions: Acidic or basic medium and heating

Uses of Alkanoic Acids

1. Preservatives: Benzoic acid is used as a preservative in pharmaceutical products to prevent microbial growth.
2. Flavoring Agents: Carboxylic acids like citric acid and lactic acid are used to add tartness and enhance flavors in various foods and beverages.
3. Soaps and Detergents: Fatty acids, such as stearic acid, are used in the production of soaps and detergents.
4. Medicinal Compounds: Salicylic acid is used in the production of aspirin and other medications for its anti-inflammatory properties.
5. Fragrances: Carboxylic acids are used in the synthesis of esters, which are key components in many fragrances and perfumes.
6. Animal Feed: Propionic acid is used as a preservative in animal feed to prevent mold growth.

Learning Tasks

1. What is the general formula of alkanoic acids?
2. Perform a laboratory experiment to test for the presence of alkanoic acids using sodium bicarbonate and litmus paper and write a lab report detailing the procedure, observations, and conclusions.
3. Create a list of 10 alkanoic acids with their common names and IUPAC names.
4. Research and present a report on the various uses of alkanoic acids in food, pharmaceuticals, and industry.

PEDAGOGICAL EXEMPLARS

1. **Collaborative learning approach**
   - a. Divide learners into small groups of 3-4.
   - b. Assign each group a specific task related to alkanoic acids, such as:
     - i. Researching and presenting on a specific alkanoic acid
     - ii. Creating a concept map illustrating the production of alkanoic acids
     - iii. Developing a graphic organiser to compare physical properties

c. Encourage groups to work together to complete the task, promoting active discussion and problem-solving.

d. Circulate around the room to facilitate group work, providing guidance and support as needed.

e. Encourage groups to share their findings with the class, promoting peer-to-peer learning.

f. Use think-pair-share, jigsaw, or number-headed-together strategies to facilitate collaboration.

g. Emphasise teamwork, communication, and mutual respect.

h. Assess group work and provide feedback, highlighting strengths and areas for improvement.

Note

Jigsaw: Assign each group member a specific aspect of alkanoic acids (e.g., production, physical properties, biological importance). After researching, have each member teach their aspect to the group.

Concept Map Challenge: Divide groups into pairs and assign each pair a concept map to create. After completion, have pairs share and compare their maps.

2. Exploratory Learning

Objective: To design and conduct experiments to investigate the reactions of alkanoic acids with various substances.

Materials needed

- Test tubes
- Alkanoic acids (e.g., ethanoic acid, propanoic acid)
- Substances for reaction (, NaOH, $NaHCO_3$, Na_2CO_3, NH_3, ROH, $LiAlH_4$)
- Virtual laboratory or video resources (optional)

1. Divide students into small groups and assign each group a substance to investigate.

2. Ask each group to:

   a. Research the expected reaction between the alkanoic acid and their assigned substance.

   b. Design a simple test tube experiment to test their hypothesis.

   c. Conduct the experiment and record observations.

   d. Analyse results and draw conclusions.

3. Alternatively, use virtual laboratory or video resources to simulate experiments.
4. Have each group present their findings to the class.
5. Assess their ability to think critically and draw conclusions from their results.

Note

Provide additional guidance on experiment design and data analysis.

3. **Experiential Learning**
   - **a.** Conduct an experiment to test for alkanoic acids through the reaction with alkanols to produce esters. Provide a template for recording observations.
   - **b.** Talk about other ways of testing for the presence of alkanoic acids
   - **c.** Assess student understanding through participation, discussion engagement, and analysis, providing additional support and feedback for struggling students and encouraging advanced students to self-assess and peer review.
4. **Initiating Talk for learning**
   - **a.** Introduce the topic and ask learners to share prior knowledge, providing additional support for struggling learners by offering a graphic organiser and encouraging advanced learners to share more detailed information.
   - **b.** Describe synthetic pathways of alkanoic acids, using visual aids and simple language for struggling learners and providing more complex details and examples for advanced learners.
   - **c.** Explain uses of alkanoic acids, using real-life examples and multimedia resources to engage struggling learners and encouraging advanced learners to research and present on additional uses.
   - **d.** Facilitate a group discussion on the uses of alkanoic acids, providing guiding questions and prompts for struggling learners and encouraging advanced learners to lead the discussion and explore more complex topics.
   - **e.** Ask learners to reflect on what they have learned, providing sentence starters and encouraging drawing or writing for struggling learners and encouraging advanced learners to think critically and make more abstract connections.
   - **f.** Assess learners understanding through participation, discussion engagement, and a brief written reflection, providing additional support for struggling learners by offering a template and encouraging advanced learners to self-assess and peer review.

g. Provide feedback and additional support as needed, offering constructive feedback and suggestions for improvement for struggling learners and encouraging advanced learners to set goals for further learning.

KEY ASSESSMENTS

1. **Level 1:** What is the general formula for alkanoic acids?
2. **Level 2:** Name the following alkanoic acid using IUPAC nomenclature:
    a. $(CH_3)_3C\text{-}COOH$
    b. $CH_3\text{-}CH_2\text{-}C(Br)_2\text{-}COOH$
    c. $C_6H_5\text{-}CH_2\text{-}CH(Cl)\text{-}COOH$
    d. $HOOC\text{-}CH_2\text{-}CH_2\text{-}COOH$
3. **Level 3:** Compare the boiling points of acetic acid (CH_3COOH) and ethanol (CH_3CH_2OH). Explain why acetic acid has a higher boiling point.
4. **Level 4:** Design an experiment to test for the presence of an alkanoic acid in a given sample. Include the procedure, expected results, and safety precautions.
5. **Level 4:** Evaluate the impact of alkanoic acids in everyday life by discussing their uses in at least three different industries. Provide specific examples and explain their significance.

Hint

The recommended mode of assessment for Week 24 is ***End of second semester examination****. Refer to* ***Appendix G*** *at the end of this section for a Table of specification to guide the setting of the items.*

Section 8 Review

In this section, learners delved into the fascinating world of organic chemistry, exploring the structure, bonding, properties, and everyday uses of alkynes, alkanes, alkenes, benzene, alkanols, and alkanoic acids. Through collaborative research and presentations, learners classified and analysed the properties of these organic compounds, identifying trends and relationships between structure and bonding. They engaged in discussions on isomerism, nomenclature, and the practical applications of these compounds in various industries. Key concepts, such as hybridisation and molecular orbitals, were reinforced through interactive activities and experiments. Learners also examined the differences in properties and uses among these organic compounds, connecting theoretical concepts to real-world applications. Overall, this section emphasised the significance of organic chemistry in both theoretical understanding and practical contexts, preparing learners for further exploration of the subject. By the end of this section, learners gained a comprehensive understanding of the fundamental concepts of organic chemistry.

APPENDIX G: END OF SEMESTER-EXAMINATION

Structure

This end of semester examination will comprise of paper 1 and 2. Paper 1 will be made up of compulsory 30 multiple choice questions while Paper 2 will have sections A and B. Section A will consist of 5 short answer questions for learners to attempt all and section B will have 3 essay type questions of similar difficulty levels which learners must answer only 1. The exam should cover lessons learnt from weeks 13-24. Time allowed for this paper is 2hours

Resources

Stationery, Printer, Scannable answer sheets, Answer booklets, calculators, pens, pencils, etc

Sample question

The hybridization of the carbon atom in methane is…

A. sp

B. sp^2

C. sp^3

D. sp^2d

Short Answer

a. Explain why iodine is solid but chlorine is gas at room temperature

b. What happens to the melting points of metallic Period 3 elements as you move from left to right?

c. Write a balanced chemical equation between aqueous solutions of Sodium chloride and silver nitrate and state the nature of the products formed.

d. Differentiate between saturated and unsaturated hydrocarbons

e. Give 3 general uses of alkanes

Essay type

Discuss how the structure and bonding in diamond affects its uses

How to Administer

1. Prepare examination questions and materials
2. Inform learners of examination details (date, time location, etc)

3. Distribute examination materials (questions, answer sheets, etc)

4. Monitor learners to prevent cheating and maintain a quiet and secure environment

Table of Specification

Weeks	Focal Area(s)	Type of Questions	DoK Levels				Total
			1	2	3	4	
13	Acid - Base Titration	Multiple Choice	1	1	1	-	3
		Essay		-	1	-	1
14	Acid - Base Titration	Multiple Choice	1	2	-	-	3
		Short Answer		1	-	-	1
15	Physical and Chemical Properties of Period 3 Elements	Multiple Choice	1	1	1	-	3
16	Physical and Chemical Properties of Compounds of Period 3 Elements	Multiple Choice	1	-	1	-	2
		Short Answer		1	-	-	1
17	Physical and Chemical Properties of The Halogens (Group 17 Elements)	Multiple Choice	1	1	-	-	2
18	Reaction of Halide Salts, Differences in Acids Strength of the Hydrogen Halides and the Uses of the Halogens	Multiple Choice	1	1	-	-	2
		Essay	-	-	1	-	1
19	Angles of Molecular Compounds and Differences Between Shapes and Bond Sigma and Pi Bonds	Multiple Choice	1	2	-	-	3
		Short Answer	-	1	-	-	1
20	Hybridization, Structure and Bonding in Organic Compounds	Multiple Choice	1	1	-	-	2

Weeks	Focal Area(s)	Type of Questions	DoK Levels				Total
			1	2	3	4	
21	Alkanes	Multiple Choice	1	1	1	-	3
		Short Answer	-	1	-	-	1
22	Alkenes and Alkynes	Multiple Choice	-	2	-	-	2
		Short Answer		1	-	-	1
23	Benzene and Alkanols	Multiple Choice	1	1	1	-	3
		Essay	-		1	-	1
24	Alkanoic Acids	Multiple Choice		2			2
	Total		**10**	**20**	**8**	**-**	**38**
	Eassy	**3**					
	MCQ	**30**					
	Short Answers	**5**					

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