MINISTRY OF EDUCATION

CHEMISTRY

For Senior High Schools

TEACHER MANUAL

YEAR 1 - BOOK 2

NATIONAL COUNCIL FOR
CURRICULUM & ASSESSMENT
OF MINISTRY OF EDUCATION

MINISTRY OF EDUCATION

REPUBLIC OF GHANA

Chemistry

For Senior High Schools

Teacher Manual

Year One - Book Two

NaCCA
NATIONAL COUNCIL FOR
CURRICULUM & ASSESSMENT
OF MINISTRY OF EDUCATION

CHEMISTRY TEACHER MANUAL

Enquiries and comments on this manual should be addressed to:
The Director-General
National Council for Curriculum and Assessment (NaCCA)
Ministry of Education
P.O. Box CT PMB 77
Cantonments Accra

Telephone: 0302909071, 0302909862
Email: info@nacca.gov.gh
website: www.nacca.gov.gh

CONTENTS

INTRODUCTION

The National Council for Curriculum and Assessment (NaCCA) has developed a new Senior High School (SHS), Senior High Technical School (SHTS) and Science, Technology, Engineering and Mathematics (STEM) Curriculum. It 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 Book Two of the Teacher Manual for Chemistry covers all aspects of the content, pedagogy, teaching and learning resources and assessment required to effectively teach Year One of the new curriculum. It contains information for the second 12 weeks of Year One. Teachers are therefore to use this Teacher Manual to develop their weekly Learning Plans as required by Ghana Education Service.

Some of the key features of the new curriculum are set out below.

Learner-Centred Curriculum

The SHS, SHTS, and STEM curriculum places the learner at the center of teaching and learning by building on their existing life experiences, knowledge and understanding. Learners are actively involved in the knowledge-creation process, with the teacher acting as a facilitator. This involves using interactive and practical teaching and learning methods, as well as the learner's environment to make learning exciting and relatable. As an example, the new curriculum focuses on Ghanaian culture, Ghanaian history, and Ghanaian geography so that learners first understand their home and surroundings before extending their knowledge globally.

Promoting Ghanaian Values

Shared Ghanaian values have been integrated into the curriculum to ensure that all young people understand what it means to be a responsible Ghanaian citizen. These values include truth, integrity, diversity, equity, self-directed learning, self-confidence, adaptability and resourcefulness, leadership and responsible citizenship.

Integrating 21st Century Skills and Competencies

The SHS, SHTS, and STEM curriculum integrates 21st Century skills and competencies. These are:

- **Foundational Knowledge:** Literacy, Numeracy, Scientific Literacy, Information Communication and Digital Literacy, Financial Literacy and Entrepreneurship, Cultural Identity, Civic Literacy and Global Citizenship
- **Competencies:** Critical Thinking and Problem Solving, Innovation and Creativity, Collaboration and Communication
- **Character Qualities:** Discipline and Integrity, Self-Directed Learning, Self-Confidence, Adaptability and Resourcefulness, Leadership and Responsible Citizenship

Balanced Approach to Assessment - not just Final External Examinations

The SHS, SHTS, and STEM curriculum promotes a balanced approach to assessment. It encourages varied and differentiated assessments such as project work, practical demonstration, performance assessment, skills-based assessment, class exercises, portfolios as well as end-of-term examinations and final external assessment examinations. Two levels of assessment are used. These are:

- Internal Assessment (30%) – Comprises formative (portfolios, performance and project work) and summative (end-of-term examinations) which will be recorded in a school-based transcript.
- External Assessment (70%) – Comprehensive summative assessment will be conducted by the West African Examinations Council (WAEC) through the WASSCE. The questions posed by WAEC will test critical thinking, communication and problem solving as well as knowledge, understanding and factual recall.

The split of external and internal assessment will remain at 70/30 as is currently the case. However, there will be far greater transparency and quality assurance of the 30% of marks which are school-based. This will be achieved through the introduction of a school-based transcript, setting out all marks which learners achieve from SHS 1 to SHS 3. This transcript will be presented to universities alongside the WASSCE certificate for tertiary admissions.

An Inclusive and Responsive Curriculum

The SHS, SHTS, and STEM curriculum ensures no learner is left behind, and this is achieved through the following:

- Addressing the needs of all learners, including those requiring additional support or with special needs. The SHS, SHTS, and STEM curriculum includes learners with disabilities by adapting teaching and learning materials into accessible formats through technology and other measures to meet the needs of learners with disabilities.
- Incorporating strategies and measures, such as differentiation and adaptative pedagogies ensuring equitable access to resources and opportunities for all learners.
- Challenging traditional gender, cultural, or social stereotypes and encouraging all learners to achieve their true potential.
- Making provision for the needs of gifted and talented learners in schools.

Social and Emotional Learning

Social and emotional learning skills have also been integrated into the curriculum to help learners to develop and acquire skills, attitudes, and knowledge essential for understanding and managing their emotions, building healthy relationships and making responsible decisions.

Philosophy and vision for each subject

Each subject now has its own philosophy and vision, which sets out why the subject is being taught and how it will contribute to national development. The Philosophy and Vision for Chemistry is:

Philosophy: All learners can engage in an exciting and fascinating learning experience in Chemistry, with inquiry and experimental skills and competencies for transition to further studies, lifelong learning or the world of work.

Vision: Learners who exhibit competencies in the critical evaluation of scientific and technological development, capable of developing products, processes in chemistry and related fields as well as further studies.

SUMMARY SCOPE AND SEQUENCE

S/N	STRAND	SUB-STRAND	YEAR 1			YEAR 2			YEAR 3		
			CS	LO	LI	CS	LO	LI	CS	LO	LI
1.	Physical Chemistry	Matter and its Properties	3	4	22	1	2	8	-	-	-
		Equilibria	1	1	3	2	2	9	4	4	10
2..	Systematic Chemistry of the Elements	Periodicity	1	1	2	2	2	4	1	1	2
		Bonding	2	2	5	1	1	2	-	-	-
3.	Chemistry of Carbon Compounds	Characterization of Organic Compounds	1	1	2	1	1	1	-	-	-
		Organic Functional groups	1	1	2	1	1	5	2	2	4
Total			**9**	**10**	**36**	**8**	**9**	**29**	**7**	**7**	**16**

Overall Totals (SHS 1 – 3)

Content Standards	**24**
Learning Outcomes	**26**
Learning Indicators	**81**

SECTION 4: KINETIC THEORY AND STATES OF MATTER

Strand: **Physical Chemistry**

Sub-Strand: Matter and its properties

Content Standard: Demonstrate understanding of the use of the kinetic theory of matter to explain the behaviour of solids, liquids and gases under different conditions and describe the laboratory preparation of gases as well as their uses in everyday life.

Learning Outcome: *Use the kinetic theory of matter to explain the behaviour of solids, liquids and gases under different conditions and describe the laboratory preparation of gases as well as their uses in everyday life.*

INTRODUCTION AND SECTION SUMMARY

This section is a continuation of the previous section and covers weeks 13 and 14.

The content standard focuses on demonstrating understanding and practical application of kinetic theory, including the laboratory preparation of gases and their everyday uses. The learning outcome is for learners to proficiently utilise the kinetic theory to explain the properties of different states of matter and comprehend the practical implications of gas preparation and usage. Throughout this section, tailored tasks will be employed to cater to diverse learning needs, ensuring comprehensive coverage of the content.

SUMMARY OF PEDAGOGICAL EXEMPLARS

Various pedagogical approaches will be utilised, ranging from hands-on experiments to visual aids and discussions, to engage learners effectively. This will help in the acquisition of concepts to be learned. Varied teaching strategies and resources are employed to cater for the diverse needs of learners. Conscious effort should be made to use mixed-ability groupings and help learners to respect group dynamics.

ASSESSMENT SUMMARY

Key assessment strategies will be implemented to measure learners' depth of knowledge, incorporating tasks of varying complexity to assess their understanding and application of the kinetic theory and gas-related concepts. Prominent among them are quizzes, assignments and projects/practical/laboratory works. Support any learner who is struggling to understand the concept.

The weeks covered by the section are:

Week 13: *The ideal gas equation, Non-ideal gas behaviour and Van der Waals equation.*

Week 14: *Preparation and test for hydrogen gas, Preparation and test for Carbon dioxide gas and Preparation and test for Ammonia gas*

Week 13

Learning Indicators:

1. *Write the ideal gas equation and apply it in simple calculations using the different numerical values of R and units of pressure and volume.*
2. *Explain why gases show deviation from ideal behaviour and suggest how the ideal gas equation could be modified to describe gas behaviour more accurately.*

Theme or Focal Area: The ideal gas equation

The ideal gas equation, also known as the ideal gas law, is a fundamental equation in chemistry that describes the behaviour of gases under certain conditions. It relates the pressure (P), volume (V), temperature (T) and amount of gas in moles (n) of an ideal gas sample. It combines Boyle's law, Charles's Law and Avogadro's law.

$$PV = nRT$$

Where;

P = Pressure of gas

V = Volume of gas

n = Moles of gas

T = Temperature of gas

R = Ideal gas constant = 8.314 $Jmol^{-1}K^{-1}$(S.I. unit of R), Other units of R = 0.082057L atm $mol^{-1}K^{-1}$, 62.364L Torr $mol^{-1}K^{-1}$ and 8.3145 m^3 Pa $mol^{-1}K^{-1}$

NB: Because of the different units of R, it is important to always match the units of pressure, volume, number of moles and temperature given with the units of R.

If the value of R is given as 0.082057L atm $mol^{-1}K^{-1}$**,** the unit for pressure must be atm, for volume must be litre and for temperature must be Kelvin.

If the value of R is given as 62.364L Torr $mol^{-1}K^{-1}$**,** the unit for pressure must be Torr, for volume must be litre, and for temperature must be Kelvin.

The ideal gas equation is essential for understanding and predicting the behaviour of gases in various chemical reactions and processes.

Example: What is the volume of 10 g of nitrogen gas at 25 ^{0}C and 101 kPa? [N =14, R = 8.314 $Jmol^{-1}K^{-1}$]

Solution (Use problem-solving strategy)

a. Analyse the question

Known

N =14, R = 8.314 $Jmol^{-1}K^{-1}$

Temperature, T = 25 + 273 = 298 K

Mass, m = 10 g

Pressure, P = 101 kPa = 101 000 Pa

R = 8.314 $Jmol^{-1}K^{-1}$

Unknown

Volume, $V = ?$

b. **Apply the problem-solving strategy**

$$\begin{aligned} M(N_2) &= 2(14) \\ &= 28\text{g} \\ n &= \frac{m}{M} \\ &= \frac{10}{28} \\ &= 0.357 \text{ mol} \\ PV &= nRT \\ V &= \frac{nRT}{P} \\ &= \frac{0.357 \times 8.314 \times 298}{101000} \\ &= 0.0088 \text{ m}^3 \end{aligned}$$

$\therefore$ volume $= 0.0088 \text{ m}^3$

Learning Tasks

1. What is the ideal gas equation and how is it derived?
2. A gas occupies a volume of 10 litres at a pressure of 2 atm and a temperature of 50°C. How many moles of gas are present?
3. A 150 dm^3 cylinder of oxygen gas at a hospital exerts a pressure of 800 kPa at 25°C. Calculate the mass of oxygen contained in the cylinder. [$R = 8.314 \text{ Jmol}^{-1}\text{K}^{-1}$, O = 16]

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

a. **Collaborative Learning:**

  i. Begin the lesson by providing a brief overview of Boyle's Law, Charles' Law and Avogadro's Law, explaining each law in simple terms and using relatable examples

  ii. Facilitate group discussions where learners can share their understanding of each gas law and how they relate to one another. Encourage learners to ask questions and challenge each other's thinking.

  iii. Divide the class into small mixed-ability groups and assign specific roles within each group, such as researcher, presenter, recorder and facilitator, to ensure everyone participates and contributes to the learning process.

  iv. Allow learners to work in groups and show how the laws above can be combined to give the ideal gas law, explaining each of the variables in the ideal gas Law.

b. **Problem-solving approach:**

  1. While maintaining the mixed ability groupings, present learners with real-life scenarios or problems that require the application of the ideal gas equation ($PV = nRT$) and encourage groups to discuss and solve the problems together.

  2. Incorporate interactive quizzes during the lesson to gauge learners' understanding of key concepts related to the ideal gas equation in real-time. Use multiple-choice questions or short-answer prompts to assess learners' grasp of the concept and provide immediate feedback to guide their learning.

Key Assessments (DoK)

1. **Level 2:** State the units of pressure, volume and temperature used in the ideal gas equation.
2. **Level 2:** A gas sample has a volume of 3.00 L at a pressure of 2.00 atm and a temperature of 300 K. Calculate the number of moles of gas using the ideal gas equation.
3. **Level 3:** Analyse the implications of using the ideal gas equation to predict the behaviour of a gas at extremely high pressures or low temperatures, considering factors such as phase changes and molecular interactions.

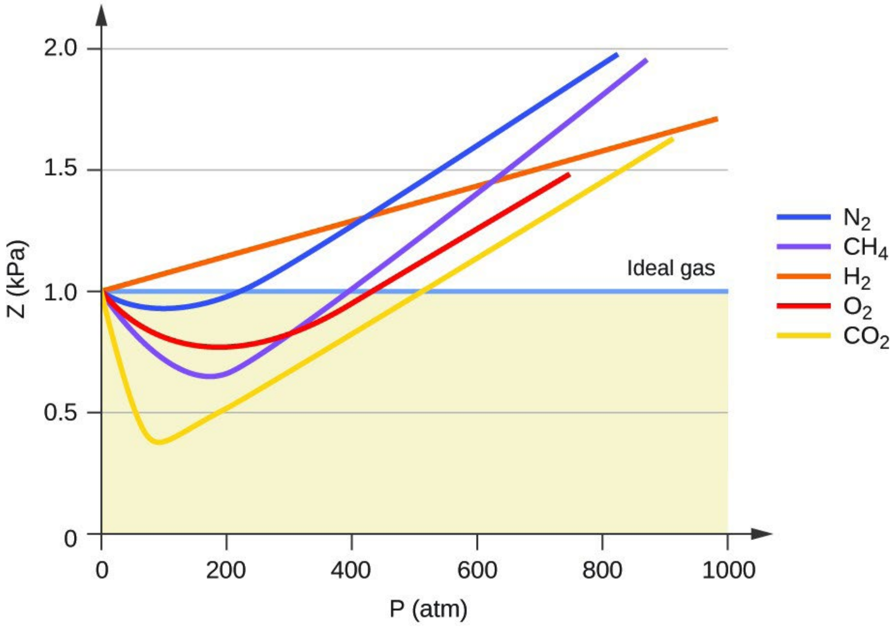
Theme or Focal Area: Non-ideal gas behaviour

A real gas behaves differently from what is expected in ideal conditions at:

- **a.** High-pressure
- **b.** Low temperature

This is due to the following:

- **a.** Real gases have an actual volume of molecules, which is significant at very high pressure. At very high pressure, the value of PV becomes greater than the ideal volume. The deviation increases when the relative molecular mass (Mr) increases.
- **b.** Intermolecular forces always exist in real gases. At lower temperatures, the kinetic energy of the molecules is at their lower value, intermolecular forces increase, which reduces the pressure P, making the PV in value less than the ideal value. Gases with greater intermolecular forces are more polar.



Example: Arrange these gases in their order of deviation from ideal behaviour and explain the order: O_2, Ne, NH_3

Solution

$$\underrightarrow{Ne < O_2 < NH_3}$$

Increasing order of deviation

Reasons: Neon, a noble gas, exhibits the least deviation from ideal behaviour. This is because noble gases have very weak intermolecular forces and their molecules are far apart, resulting in minimal interactions between them. Oxygen, while also a non-polar molecule like neon, has slightly stronger intermolecular forces due to its higher molar mass. However, these forces are still relatively weak, compared to other gases, resulting in a moderate deviation from ideal behaviour. Ammonia, a polar molecule with hydrogen bonding, experiences the highest deviation from ideal behaviour. Hydrogen bonding leads to stronger intermolecular attraction between ammonia molecules, compared to neon and oxygen, causing greater deviation from the ideal gas law.

Learning Tasks

1. Compile a list of examples in daily life where gases behave differently than predicted by the ideal gas law.
2. Explain in simple terms why real gases deviate from ideal behaviour, using concepts such as molecular size, intermolecular forces and volume occupied by gas particles.
3. Use online simulations to explore how gas behaviour changes under different conditions such as high pressure or low temperature. Record your observations and discuss them with your classmates.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

Project-based Learning:

1. Start by introducing the ideal gas law equation ($PV = nRT$) and its assumptions (particles have zero volume and experience no intermolecular forces).
2. Explain that while the ideal gas law is useful for many situations, real gases often deviate from ideal behaviour under certain conditions.
3. Engage learners in discussions about common scenarios where real gases deviate from ideal behaviour, such as high pressure, low temperature or when gas molecules are large or have strong intermolecular forces.
4. Use examples and demonstrations to illustrate these deviations. For instance, show how gases liquefy at low temperatures and high pressures, contrary to the behaviour predicted by the ideal gas law.
5. Through various sources such as the internet, videos or books, assign learners to research various factors that contribute to deviations from ideal gas behaviour, such as molecular size, intermolecular forces or gas mixtures. Have them do presentations on their findings.
6. Integrate quizzes during the lesson to gauge learners' understanding of key concepts related to non-ideal gas behaviour in real time and provide immediate feedback to guide their learning

Key Assessment (DoK)

1. **Level 2:** What is meant by non-ideal gas behaviour?
2. **Level 2:** Describe the impact of molecular size on the deviation of gases from ideal behaviour
3. **Level 3:** Evaluate the significance of intermolecular forces in causing deviations from ideal behaviour in real gases. Discuss how these forces affect gas behaviour under different conditions.

Theme or Focal Area: Van der Waals equation

Van der Waals equation is an equation of state that extends the ideal gas law to include the non-zero size of gas molecules and the interactions between them.

For one mole of gas, the equation is $(\boldsymbol{P} + \frac{a}{V^2})(V - b) = \boldsymbol{RT}$ and for n moles of gas, the Van der Waals equation is $[\boldsymbol{P} + (an^2/v^2)]\ (\boldsymbol{V} - \boldsymbol{nb}) = \boldsymbol{nRT}$

Where;

a = a measure of the strength of the intermolecular forces.

b = The excluded molar volume

Learning Tasks

1. State Van der Waals equation for one mole of an ideal gas
2. Describe what the Van der Waals equation is used for and why it was developed.
3. Compare the Van der Waals equation with the ideal gas law. Discuss situations where the ideal gas law is appropriate and where the Van der Waals equation is more accurate. Provide examples to support your explanation.

Pedagogical Exemplars (*with the cross-cutting themes integrated*)

Talk for Learning:

1. Start the lesson by introducing the topic on the Van der Waals equation and its purpose in describing the behaviour of real gases.
2. Encourage learners to share their prior knowledge or experiences related to gas behaviour and ideal gas laws.
3. Use open-ended questions to spark discussion and promote critical thinking about the limitations of the ideal gas law and the need for corrections provided by the Van der Waals equation.
4. Divide the class into small mixed-ability groups and assign each group a specific aspect of the Van der Waals equation to discuss, e.g., the significance of the a and b constants or the conditions under which the equation is most applicable.
5. At the end of the lesson or discussion, provide opportunities for learners to reflect on what they have learned about the Van der Waals equation.
6. Ask students to summarise the key concepts from the discussion, either individually or as a group, to reinforce their understanding.

Key Assessment (DoK)

1. **Level 1:** State one difference between the ideal gas law and the Van der Waals equation.
2. **Level 2:** Use the Van der Waals equation to calculate the pressure of 1 mole of carbon dioxide gas in a 2-litre container at 273 K. Given: $\boldsymbol{a} = 3.592$, $\boldsymbol{b} = 0.0427$ and $\mathbf{R} = 8.314\ \text{Jmol}^{-1}\text{K}^{-1}$
3. **Level 3:** Discuss the significance of the $\boldsymbol{a}$ and $\boldsymbol{b}$ constants in the Van der Waals equation. How do these constants account for the deviations of real gases from ideal behaviour?

References

1. SHS Chemistry Curriculum
2. Gilbert, T. R., Davies, G., & Davies, N. (2009). Chemistry (2nd ed.). W. W. Norton & Company, Inc.
3. Addison-Wesley. (2000). Chemistry (5th ed.). Prentice Hall.
4. Ryan, L., & Norris, R. (2014). Cambridge International AS and A-level Chemistry Coursebook. University Printing House.
5. Chemistry. (2015). Science Resource Centre Handbook, Itec Global.
6. Doc Brown. (n.d.). Gas Laws. Retrieved from https://docbrown.info/page03/3_52gaslaws.htm
7. Khan Academy. (n.d.). Ideal Gas Law. Retrieved from https://www.khanacademy.org/science/ap-chemistry-beta/x2eef969c74e0d802:intermolecular-forces-and-properties/x2eef969c74e0d802:ideal-gas-law/e/ideal-gas-law
8. Chang, R. (2008). General Chemistry: The Essential Concepts (5th ed.). McGraw Hill.
9. https://courses.lumenlearning.com/suny-mcc-chemistryformajors-1/chapter/non-ideal-gas-behavior/

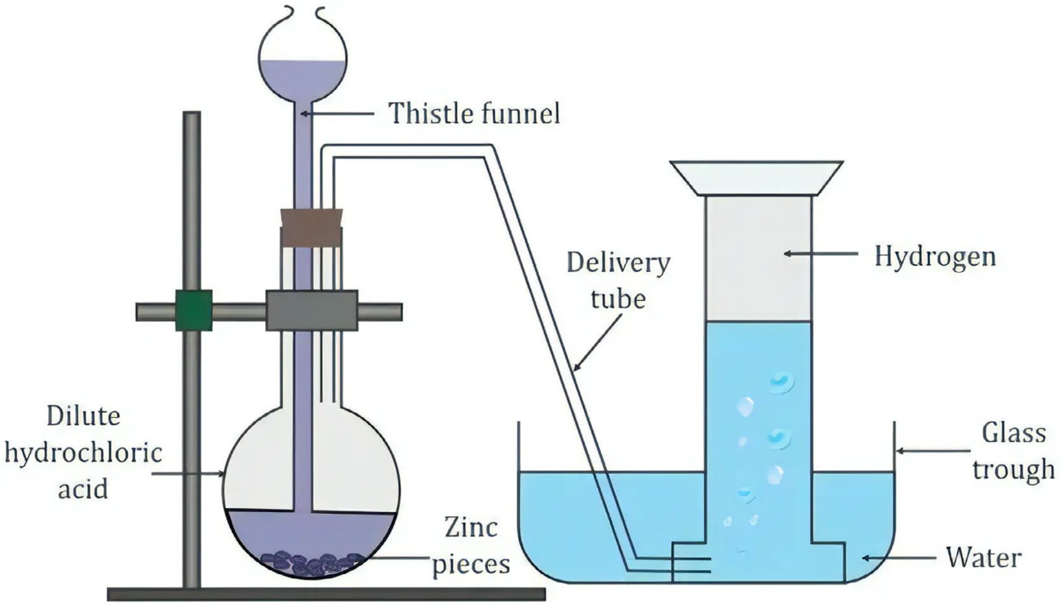
Week 14

Learning Indicator: *Design and perform experiments to prepare and test for gases (hydrogen, ammonia and carbon dioxide gases).*

Theme or Focal Area: Preparation and test for hydrogen gas

Equation for preparation:

$Mg + 2HCl \rightarrow MgCl_2 + H_2$



Preparation of Hydrogen gas

Test for hydrogen gas: Put a burning/lighted splint in the gas; a "pop" sound indicates the presence of hydrogen gas.

How to dry hydrogen gas: Pass the gas produced through anhydrous $CaCl_2$.

Physical properties of hydrogen gas

i. It is the lightest gas and has the lowest density.
ii. It is colourless and odourless.
iii. It is insoluble in water.

Chemical properties of hydrogen gas

i. It is neutral to litmus.
ii. It is unreactive under normal conditions.
iii. It does not support combustion.
iv. It burns in oxygen (air) with the pop sound to produce water.
   $2H_2 + O_2 \rightarrow 2H_2O$
v. It reacts with halogens
   $H_2 + F_2 \rightarrow 2HF$

Uses of hydrogen gas in everyday life

- **i.** It is used to produce ammonia gas in the Haber process.
- **ii.** It is used to manufacture margarine by hydrogenation of unsaturated fats.
- **iii.** It is used in oxy-hydrogen flame for cutting and welding of metals.
- **iv.** It is used in fuel cells.

Learning Tasks

1. State at least two metals that can be used to prepare hydrogen gas.
2. Use a balanced chemical equation to explain the 'pop' sound heard when hydrogen gas is tested with a lighted splint.
3. Explain why in the preparation of hydrogen gas in the laboratory, the end of the thistle funnel in the flask must be dipped in the acid.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

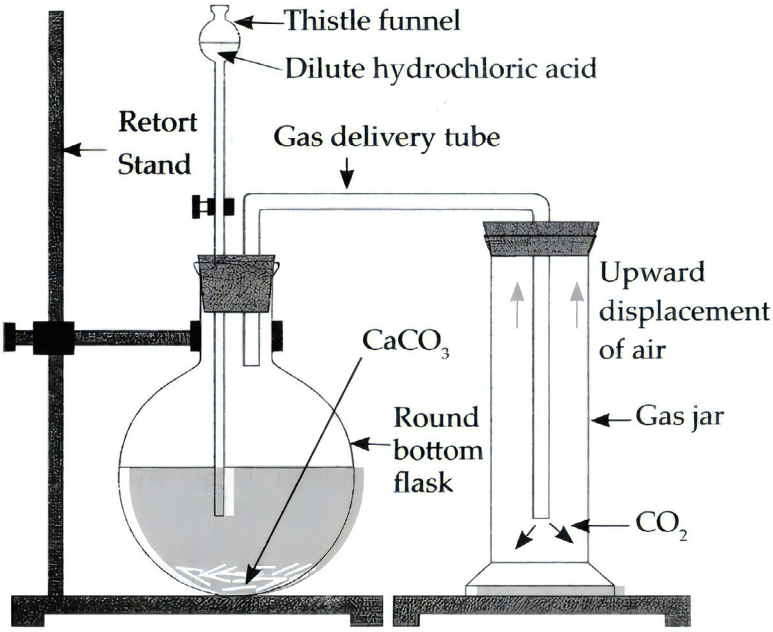
1. **Collaborative learning:**
    a. Divide learners into mixed-ability groups and guide them to discuss the properties of hydrogen gas. Encourage slow learners to participate in the discussion by asking them questions.
    b. Guide learners to state the uses of hydrogen gas in everyday life.
2. **Experiential learning:**
    a. In mixed-ability groups, guide learners to perform experiments to prepare hydrogen gas.
    b. In mixed-ability groups, guide learners to create charts to illustrate the experimental set-up for the preparation of hydrogen gas.
    c. Guide learners to make deductions from the experiment to prepare hydrogen gas.

Key Assessments (DoK)

1. **Level 1:** State two properties of hydrogen that enable it to be used to fill balloons.
2. **Level 1:** State three practical uses of hydrogen in everyday life.
3. **Level 2:** State the method of collecting hydrogen gas and explain why it is collected by this method.
4. **Level 4:** You are provided with the thistle funnel, delivery tube, split cork, conical flask, gas jar, beehive stand, water trough, Magnesium ribbon and dilute HCl. Design an experiment to prepare and test for hydrogen gas.

Theme or Focal Area: Preparation and test for Carbon dioxide gas

1. Equation for preparation: $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$



Preparation of Carbon dioxide

2. **Test for carbon dioxide gas**: Pass the gas through lime water (saturated solution of calcium hydroxide), which turns milky.
3. **How to dry carbon dioxide gas**: Pass the gas produced through concentrated H_2SO_4.
4. **Physical properties of carbon dioxide gas**
   - i. It is colourless and odourless.
   - ii. It is denser than air.
   - iii. It is soluble in water.
   - iv. It condenses into a white solid called dry ice.
5. **Chemical properties of carbon dioxide gas**
   - i. It turns moist blue litmus red.
   - ii. It turns lime water ($Ca(OH)_2$ solution) milky.
     $CO_2 + Ca(OH)_2 \rightarrow CaCO_3 + H_2O$
   - iii. Excess passage of CO_2 causes milkiness to disappear.
     $CaCO_3 + H_2O + CO_2 \rightarrow Ca(HCO_3)_2$
   - iv. It reacts with water in the presence of sunlight and chlorophyll to produce glucose and oxygen (photosynthesis). $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$
   - v. It does not support combustion.
6. Uses of Carbon dioxide in everyday life
   - i. It is used in photosynthesis to produce glucose and oxygen.
   - ii. It is used to produce fire extinguishers.
   - iii. It is used to manufacture fizzy drinks.
   - iv. It is used to manufacture refrigerants.

Learning Tasks

1. Name the chemicals required for the preparation of carbon dioxide in the laboratory.
2. Why is sulphuric acid not used for the preparation of carbon dioxide in the laboratory?
3. Explain what happens when carbon dioxide gas is passed through a saturated solution of calcium hydroxide briefly and in excess.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

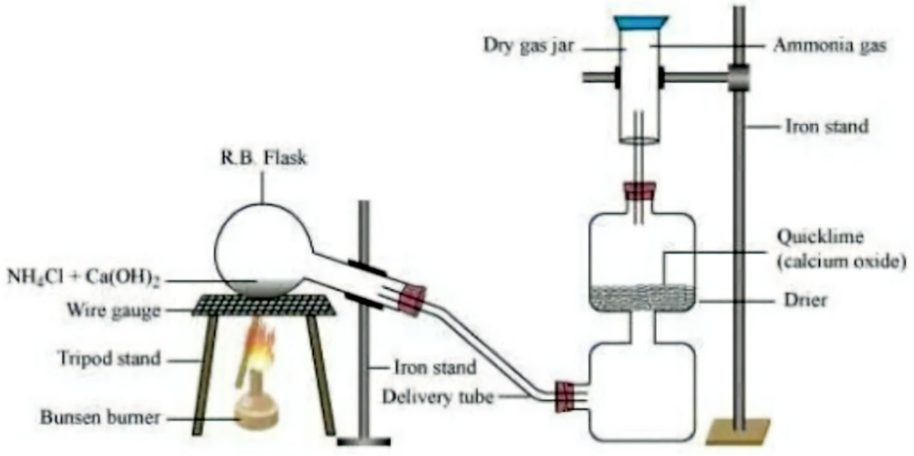
1. **Collaborative learning:**
   a. Divide learners into mixed-ability groups and guide them to discuss the properties of carbon dioxide gas. Encourage slow learners to participate in the discussion by asking them questions.
   b. Guide learners to state the uses of carbon dioxide gas in everyday life.
2. **Experiential learning:**
   a. In mixed-ability groups, guide learners to perform experiments to prepare carbon dioxide gas.
   b. In mixed-ability groups, guide learners to create charts to illustrate the experimental set-up for the preparation of carbon dioxide gas.
   c. Guide learners to make deductions from the experiment to prepare carbon dioxide gas.

Key Assessments (DoK)

1. **Level 1:** State the properties of CO_2 that allow it to be used in the manufacture of:
   a) fizzy drinks
   b) as refrigerants.
2. **Level 2:** State the method collecting carbon dioxide gas and explain why it is collected by this method.
3. **Level 3:** With the help of a neatly labelled diagram, explain the laboratory preparation of carbon dioxide gas.

Theme or Focal Area: Preparation and test for Ammonia gas

1. **Equation for preparation:** $Ca(OH)_2 + 2NH_4Cl \rightarrow CaCl_2 + 2H_2O + 2NH_3$



2. **Test for gas:**
   i. It turns moist red litmus blue.
   ii. It forms white fumes with concentrated HCl vapour to form NH_4Cl.
3. **How to dry the gas:** Pass the gas produced through calcium oxide (CaO).
4. **Physical properties:**
   i. It is a colourless gas with a pungent, choking smell.
   ii. It is less dense than air.
   iii. It is soluble in water.
5. **Chemical properties:**
   i. It turns moist red litmus blue.
   ii. It forms dense white fumes with HCl ($NH_3 + HCl \rightarrow NH_4Cl$)
   iii. It burns in oxygen with a pale yellow flame.
6. **Uses of Ammonia gas in everyday life.**
   i. It is used to manufacture fertilisers.
   ii. It is used to manufacture explosives.
   iii. It is used to manufacture nylon.
   iv. It is used to manufacture plastics.
   v. It is used to manufacture pigment.

Learning Tasks

1. How is ammonia gas collected in the laboratory?
2. Concentrated sulphuric acid is used as a drying agent. Explain why it is not advisable to use it as a drying agent in the test for ammonia gas. Support your answer with an equation.
3. Explain why ammonium nitrate is not used in the preparation of ammonia and why ammonia gas is not collected over water.

Pedagogical Exemplars (*with the cross-cutting themes integrated*)

1. **Collaborative learning:**
   a) Divide learners into mixed-ability groups and guide them to discuss the properties of ammonia gas. Encourage all learners to participate in the discussion by asking them questions.
   b) Guide learners to state the uses of ammonia gas in everyday life.
2. **Experiential learning:**
   a) In mixed-ability groups, guide learners to perform experiments to prepare ammonia gas.
   b) In mixed-ability groups, guide learners to create charts to illustrate the experimental set-up for the preparation of ammonia gas.
   c) Guide learners to make deductions from the experiment to prepare ammonia gas.

Key Assessments (DoK):

1. **Level 1:** State the effect of ammonia gas on
   a) moist blue litmus paper;
   b) moist red litmus paper.
2. **Level 2:** How is ammonia gas prepared in the laboratory, starting from NH_4Cl? State the conditions and balanced equation for preparation
3. **Level 3:** In the laboratory preparation of ammonia gas, the round bottom flask containing the reagents is slanted downwards whilst being heated. Explain why these actions are necessary in the preparation of the gas.

Reference

1. SHS Chemistry Curriculum
2. https://docbrown.info/page13/ChemicalTests/GasPreparation.htm
3. Experimental setup for hydrogen gas preparation –‘’ Retrieved [April 16, 2024] from https://www.bing.com/images/search?q=experimental+setup+for+hydrogen+gas+preparation&form=HDRSC3&first=1’’
4. Laboratory preparation of carbon dioxide gas-‘’ Retrieved [April 16, 2024] from https://www.sarthaks.com/1664312 ‘’
5. preparation of ammonia gas-‘’Retrieved [April 16, 2024] https://hi-static.z-dn.net/files/d66/498677059191d6e494623ea7d566c44e.jpg’’

SECTION 5: SOLUBILITY AND ITS APPLICATIONS IN QUALITATIVE ANALYSIS

Strand: **Physical Chemistry**

Sub-Strand: Equilibria

Content Standard: Demonstrate knowledge and application of solubility and solubility rules.

Learning Outcome: *Apply the solubility rules to analyse and predict the behaviour of common ionic compounds in qualitative analysis*

INTRODUCTION AND SECTION SUMMARY

During Weeks 15 and 16, learners will delve into the content standard of demonstrating knowledge and application of solubility and solubility rules. The learning outcome for this section is to effectively apply the solubility rules to analyse and predict the behaviour of common ionic compounds in qualitative analysis. To achieve this, a variety of tailored tasks will be utilised, focusing on indicators such as explaining solubility, describing factors influencing solubility, determining the solubility of different substances and conducting tests to identify ions based on solubility rules.

SUMMARY OF PEDAGOGICAL EXEMPLARS

Various pedagogical approaches including talk for learning, demonstrations and inquiry-based learning will be employed. By engaging with these methods, learners will develop a comprehensive understanding of solubility concepts and their practical applications in qualitative analysis. Varied teaching strategies and resources were employed to cater for the diverse needs of learners. Conscious effort should be made to use mixed-ability groupings and help learners to respect group dynamics.

ASSESSMENT SUMMARY

Different assessment strategies will be implemented to measure learners' depth of knowledge, incorporating tasks of varying complexity to assess their understanding and application of the kinetic theory and gas-related concepts. Prominent among them are quizzes, assignments and projects/practical/laboratory works. Support any learner who is struggling to understand the concept. Support learners who could not perform to expectation.

The weeks covered by the section are:

Week 15: *Solubility and Solubility Rules*

Week 16: *Qualitative Chemical Analysis*

Week 15

Learning Indicator: *Explain the term solubility and describe the factors that affect the solubility of substances.*

Theme or Focal Area: Solubility

Review the previous knowledge of the learner on solutions, solutes and solvents from the CCP (JHS) curriculum.

When a solution is formed from substances in different states of matter, the solvent is the substance that retains its state of matter. This implies that the solute changes its state. The constituent of the mixture that is present in a smaller amount is called the solute and the one present in a larger quantity is called the solvent. Demonstrate using table salt (NaCl) and water to prepare a salt solution.

Types of Solutions based on states of constituent

Solutions can be prepared in any of the three states of matter (solid, liquid, and gas). Even though the most common type of solution that is seen in our immediate environment is where the solute is solid and the solvent is in the liquid state (for example a salt or sugar solution), there are different types of solutions in nature. Any of the three states of matter can behave as a solute or a solvent depending on its quantity in the solution.

Type of Solution	State of solute	State of solvent	Example
Solid-liquid	solid	liquid	Salt solution
Liquid-liquid	liquid	liquid	Alcohol in water
Gas-liquid	Gas	Liquid	dew
Gas - gas	Gas	Gas	Air
Gas - solid	Gas	Solid	H_2 with Platinum
Solid - solid	Solid	Solid	Alloys

Based on the extent of solute dissolution in a solvent at a given temperature, a solution can be categorised as unsaturated, saturated or supersaturated. An unsaturated solution is a solution that can dissolve more solute at a given temperature. A saturated solution is a solution that cannot dissolve any more solute at a given temperature. A supersaturated solution contains more solute in a solution than the maximum amount it can dissolve at a given temperature in the presence of undissolved solutes.

Demonstrate unsaturated and saturated solutions using a named salt

1. Pour about 100cm^3 of water into a clean dry beaker.
2. Use a spatula to transfer the salt into the water in the beaker.
3. Stir with a rod until all dissolves.
4. Describe the type of solution formed.
5. Transfer more solute and stir to dissolve.
6. As more solute is added, observe and describe what happens.
7. Filter any solute that may be left and collect the filtrate in a beaker.
8. Describe the type of solution formed.
9. Transfer more solute into the filtrate and stir.

10. As more solute is added, observe and describe what happens.

The solubility of a solute at a given temperature is the maximum quantity of the solute in moles or grammes that dissolves in one (1) cubic decimetre of a solvent to form a saturated solution.

Factors that affect Solubility

1. Nature of the solvent concerning the solute
2. Temperature
3. Pressure

1. **Nature of the solvent concerning the solute:** Polar solvents have molecules with permanent dipoles, making them interact with polar solutes while non-polar solutes lack permanent dipoles and interact with non-polar solutes. When two substances are similar in terms of polarity, they are more likely to dissolve in each other. Polar solutes dissolve in polar solvents while non-polar solutes dissolve in non-polar solvents. Ionic compounds are more soluble in polar solvents than non-polar solvents. Inorganic acids, bases and ionic salts tend to be much more soluble in water.
2. **Temperature:** Generally, solubility increases with the rise in temperature and decreases with the fall of temperature, although this is not always the case. Heat is needed to break the bonds holding the molecules or ionic lattices in the solid together. At the same time, the formation of new solute-solvent bonds gives off an amount of heat.
3. **Pressure:** Pressure can affect the solubility of a gas (solute) in a liquid solvent but has very little effect when the solute is a liquid or solid sample. Henry's Law states that the amount of gas that is dissolved in a liquid is directly proportional to the partial pressure of that gas.

Learning Tasks

1. Define solubility.
2. Explain the concept of solute and solvent.
3. State the units of solubility and use one of them to explain unsaturated, saturated and supersaturated solutions.
4. Conduct an experiment to investigate how temperature affects the solubility of a specific substance and analyse the results.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. **Talk-for-learning:**
   a) Put learners in groups and encourage them to share what they remember about solutions and their properties. With the aid of sample solutions, engage learners in a brainstorming session to explain the terms solute, solvent and solution.
   b) With the aid of sample solutions such as sugar solution, a mixture of kerosene and water or a mixture of coke and carbon dioxide, describe the various types of solutions based on the state of solute and solvent and give specific examples (solid-liquid, solid-gas, gas-liquid, solid-solid, liquid-liquid, gas-gas solutions).
2. **Demonstration:** Watch a video or prepare samples of the various types of solutions (unsaturated, saturated and supersaturated). Connect the concept of solubility to what students already know about solutions. Involve learners in the preparation process by allowing them to measure and add the solute to the solvent. Encourage active participation and observation to enhance their understanding of the concept.

3. **Inquiry-based learning:** In groups, experiment to investigate the factors that affect solubility: temperature, pressure, nature of solute and solvent (molecular size and polarity)

   NB: Stirring or shaking increases the rate of solubility.

   Relate the factors to everyday life such as the dissolution of salt in soup, sugar in hot and cold water and carbon dioxide in fizzy drinks.

Key Assessments (DoK):

1. **Level 1:** Define solubility.
2. **Level 1:** Define a saturated solution.
3. **Level 2:** Explain each of the following terms;
   a) Unsaturated solution
   b) Saturated solution
   c) Supersaturated solution
4. **Level 2:** Explain how temperature affects the solubility of gases in liquids.

Theme or Focal Area: Solubility Rules

Soluble Salts

1. The Na^+, K^+, and NH_4^+ ions form soluble salts. Thus, NaCl, KNO_3, $(NH_4)_2SO_4$, Na_2S, and $(NH_4)_2CO_3$ are soluble.
2. The nitrate (NO_3^-) ion forms soluble salts. Thus, $Cu(NO_3)_2$ and $Fe(NO_3)_3$ are soluble.
3. The chloride (Cl^-), bromide (Br^-), and iodide (I^-) ions generally form soluble salts. Exceptions to this rule include salts of the Pb^{2+}, Hg^{2+}, Ag^+, and Cu^+ ions. $ZnCl_2$ is soluble, but CuBr is not.
4. The sulphate (SO_4^{2-}) ion generally forms soluble salts. Exceptions include $BaSO_4$, $SrSO_4$ and $PbSO_4$, which are insoluble, and Ag_2SO_4, $CaSO_4$ and Hg_2SO_4, which are slightly soluble

Insoluble Salts

1. Sulphides (S^{2-}) are usually insoluble. Exceptions include Na_2S, K_2S, $(NH_4)_2S$, MgS, CaS, SrS, and BaS.
2. Oxides (O^{2-}) are usually insoluble. Exceptions include Na_2O, K_2O, SrO and BaO, which are soluble, and CaO, which is slightly soluble.
3. Hydroxides (OH^-) are usually insoluble. Exceptions include NaOH, KOH, $Sr(OH)_2$, and $Ba(OH)_2$, which are soluble, and $Ca(OH)_2$, which is slightly soluble.
4. Chromates (CrO_4^{2-}) are usually insoluble. Exceptions include Na_2CrO_4, K_2CrO_4, $(NH_4)_2CrO_4$, and $MgCrO_4$.
5. Phosphates (PO_4^{3-}) and carbonates (CO_3^{2-}) are usually insoluble. Exceptions include salts of the Na^+, K^+, and NH_4^+ ions

Design an experiment to determine the solubility of soluble substances.

Determination of solubility of solutes

1. Put distilled water into a clean beaker and heat gently above room temperature.
2. Add the named salt a little at a time while stirring until a saturated solution is formed.
3. Cool to room temperature

4. A known volume of the saturated solution is transferred into a previously weighed clean and dry evaporating dish.
5. Weigh the evaporating dish and the solution.
6. Evaporate to dryness using a water bath.
7. Allow to cool and then weigh.
8. Repeat the drying and cooling process to ensure a constant mass of the salt is obtained.
9. Determine the number of moles of the salt in the volume of solution that was evaporated.

Solubility Curve:

It is a graph that shows how temperature affects the solubility of salts. It is plotted with the solubility as the y-axis and temperature as the x-axis. Each point on the graph represents the maximum amount of solute that can dissolve in a solvent at a particular temperature. A change in temperature corresponds with a change in the maximum amount of the solute that can dissolve a solvent. The various points on the graph are then joined to form a curve better known as the solubility curve. The following deductions can then be made from the curve:

1. Any point below the curve represents an unsaturated solution at that particular temperature.
2. At any particular temperature, the solubility of the salt can be traced vertically to the curve.
3. Any point above the curve represents a supersaturated solution at that particular temperature.
4. To determine how much extra salt has been dissolved or crystallised out, compute the difference in the corresponding y-axis values at the respective temperatures.
5. Determine the point at which crystallisation starts and be able to calculate the mass of solute that will be crystallised out when a solution is cooled from one temperature to the other by simple subtraction on the y-axis axis.

Learning Tasks

1. Determine the insoluble compound that is formed when the calcium chloride solution is mixed with a zinc carbonate solution.
2. Design a poster to depict the steps involved in determining the solubility of a given salt.
3. Provide learners with a sample solubility curve and ask them to identify the maximum amount of solute that can be dissolved at specific temperatures.
4. Use information from a solubility curve to determine the amount of solute that will be deposited when a solution is cooled from a particular temperature to another.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. **Activity-based learning:**
    a) Introduce the concept of solubility rules and outline its importance in predicting the nature of a solution that will be formed when a salt is mixed with water.
    b) Prepare a number of cards with different ions written on each.
    c) Based on the understanding that cations combine with anions to form salts, let learners work in groups to sort a combination of the cards into soluble and insoluble salts using the solubility rules.
    d) Design an experiment to determine the solubility of a named salt at different temperatures.

2. **Collaborative Learning:** In groups, allow learners to analyse data on a given salt and determine its solubility experimentally at a particular temperature.
3. **Exploratory Learning**
   a) Give learners a set of data on the solubility of a known salt at different temperatures for them to analyse.
   b) Guide learners to plot a solubility curve using data from an experiment.
   c) Analyse and deduce information from the curve.

Key Assessments (DoK)

1. **Level 1:** What does the point above a solubility curve represent?
2. **Level 3:** Use the information below to answer the questions that follow;

Temperature/°C	Solubility of NaCl (g/100g water)
0	33.0
20	33.5
40	34.0
60	34.5
80	35.0

   i. Plot the solubility curve for NaCl
   ii. What is the solubility of NaCl at 30°C?

3. **Level 2:** The solubility values of Potassium chlorate at different temperatures are given below:

Temperature/°C	Solubility/(g/dm^3)
0	5.3
10	5.9
20	6.5
30	7.0
40	7.5
50	8.0
60	8.5
70	9.0

   i. Plot the solubility curve for $KClO_3$ from the result above.
   ii. From the graph, determine the solubility of $KClO_3$ at 35°C

References:

1. SHS Curriculum
2. Chang, Raymond (2008). General Chemistry - Essential Concepts. 5th Edition, The McGraw Hill Companies

Week 16

Learning Indicator: *Perform tests on water-soluble compounds to identify ions based on the solubility rules.*

Theme or Focal Area: Qualitative Chemical Analysis

The branch of chemistry that deals with the identification of elements or groups of elements that are present in a sample is referred to as qualitative chemical analysis. The complexity of the techniques employed in qualitative analysis depends on the nature of the sample and its interaction with other substances. It does not provide numerical information about the components in a sample, e.g. concentrations or other quantities.

Preliminary Examination

The appearance of the substance and its reaction to litmus may indicate the possible presence of many ions. For the solubility of the sample falls under the preliminary examination. When a sample is insoluble, the colour and nature of the precipitate formed are noted, as well as its solubility in excess of the precipitate above

Examination of cations

For systematic examination, groups of cations are precipitated successively from solution by reagents. The individual cations are then identified by further examination of the group precipitates. The usual reagents that are used for the identification of cations are aqueous solutions of sodium hydroxide and ammonia. The table below describes the effect of an aqueous solution of sodium hydroxide on some cations.

Cation	Drops of NaOH	Excess NaOH
Ca^{2+}	White chalky precipitate	precipitate insoluble
Pb^{2+}	White chalky precipitate	precipitate soluble
Al^{3+}	White gelatinous precipitate	precipitate soluble
Zn^{2+}	White gelatinous precipitate	precipitate soluble
Cu^{2+}	Blue gelatinous precipitate	precipitate insoluble
Fe^{2+}	Dark green gelatinous precipitate	precipitate insoluble
Fe^{3+}	Reddish brown precipitate	precipitate insoluble
NH_4^+	No visible reaction	No visible reaction

To identify cations, the test with an aqueous solution of NaOH should be repeated with an aqueous ammonia solution (also known as ammonium hydroxide, NH_4OH). The observations with ammonia solution are usually similar, but not always the same and the differences can be as important as identifying metal ions. The table below describes the effect of NH_3 on cations in drops and then in excess.

Cations	Drops of NH_3	Excess NH_3
Ca^{2+}	No visible reaction	No precipitate formed
Pb^{2+}	White chalky precipitate	precipitate insoluble
Al^{3+}	White gelatinous precipitate	precipitate insoluble
Zn^{2+}	White gelatinous precipitate	precipitate soluble
Cu^{2+}	Blue gelatinous precipitate	precipitate dissolves
Fe^{2+}	Dark green gelatinous precipitate	precipitate insoluble
Fe^{3+}	Reddish brown ppt	precipitate insoluble

Examination of anions

The procedure for detecting anions in aqueous solutions is called anion analysis. Some preliminary tests are done before anion analysis. The preliminary tests that are done include physical examination (i.e. colour and smell), effect on litmus paper and dry heating test (for the evolution of gas).

The other preliminary tests are based on the fact that CO_3^{2-}, S^{2-}, NO_3^- and SO_3^{2-} react with dilute H_2SO_4 to produce CO_2, H_2S, NO_2 and SO_2 gases respectively. These gases on identification indicate the nature of the anion present in the salt.

Cl^-, Br^-, I^- and NO_3^- react with concentrated H_2SO_4 but not with diluted H_2SO_4 to produce Cl_2, Br_2, I_2 and NO_2 gases respectively. SO_4^{2-} does not react with diluted H_2SO_4 or concentrated H_2SO_4.

The table below describes the smell, colour and effect of the various gases on litmus

Gas	Smell	Colour	Litmus
SO_2	Irritating, Pungent	Colourless	Turns red
NO_2	Irritating, Pungent	Reddish-brown	Turns red
H_2S	Rotten egg	Colourless	Faintly red
CO_2	Odourless	Colourless	Faintly red
HCl	Choking	Colourless	Turns red
HBr	Choking	Colourless	Turns red
HI	Choking	Colourless	Turns red

Anions that give white precipitates

Bench solutions that produce a white precipitate with anions include $BaCl_2$, $AgNO_3$, $Ca(OH)_2$ and $Pb(NO_3)_2$

Actions of $BaCl_2$ on some anions

This is a precipitation reaction caused by barium ions and sulphate, sulphite and carbonate ions clumping together to give a white solid mass (precipitate). $Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \longrightarrow BaSO_{4(s)}$

$BaSO_4$ is insoluble in dilute HCl.

$Ba^{2+}_{(aq)} + SO_3^{2-}{}_{(aq)} \longrightarrow BaSO_{3(s)}$

$BaSO_3$ dissolves when dilute HCl is added to produce a soluble $BaCl_2$, with the evolution of a colourless gas that has a choking smell. The gas, SO_2, produced turns the purple colour of acidified $KMnO_4$ to colourless.

$Ba^{2+}_{(aq)} + CO_3{}^{2-}{}_{(aq)} \longrightarrow BaCO_{3(s)}$

$BaCO_3$ dissolves when dilute HCl is added to produce a soluble $BaCl_2$ with the evolution of a colourless odourless gas. The gas CO_2 produced turns lime water milky.

$BaSO_4$, $BaSO_3$, and $BaCO_3$ all produce a white precipitate upon reacting with $BaCl_2$. However, each of the resulting products has a peculiar reaction with dilute HCl.

Actions of $AgNO_3$ on some anions

This is a precipitation reaction caused by silver ions with chloride, sulphite and carbonate ions clumping together to give a white solid mass (precipitate).

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

$AgCl_{(s)}$ is insoluble in dilute HNO_3 but becomes soluble when excess NH_3 is added to the resulting insoluble precipitate.

$Ag^+_{(aq)} + SO_3{}^{2-}{}_{(aq)} \longrightarrow Ag_2SO_{3(s)}$

$Ag_2SO_{3(s)}$ dissolves when dilute HNO_3 is added to produce a soluble $AgNO_3$, with the evolution of a colourless gas that has a choking smell.

The gas, SO_2, produced turns the purple colour of acidified $KMnO_4$ to colourless.

$Ag^+_{(aq)} + CO_3{}^{2-}{}_{(aq)} \longrightarrow Ag_2CO_{3(s)}$

$Ag_2CO_{3(s)}$ dissolves when dilute HNO_3 is added to produce a soluble $AgNO_3$, with the evolution of a colourless odourless gas. The gas CO_2 produced turns lime water milky.

Actions of $Ca(OH)_2$ on some anions

$Ca(OH)_2$ reacts with sulphites and carbonates to give calcium $CaSO_3$ and $CaCO_3$, which are seen as a white precipitate. Further tests or reactions with dilute hydrochloric acid help to evolve characteristic gases.

$CaSO_3$ dissolves when dilute HCl is added to produce a soluble $CaCl_2$, with the evolution of a colourless gas that has a choking smell. The gas, SO_2, produced turns the purple colour of acidified $KMnO_4$ to colourless.

$CaCO_3$ dissolves when dilute HCl is added to produce a soluble $CaCl_2$, with the evolution of a colourless odourless gas, CO_2, which turns lime water milky.

Actions of $Pb(NO_3)_2$ on some anions

Lead(II) nitrate, $Pb(NO_3)_2$, can have various effects on different anions, depending on their chemical properties. It is often used in qualitative analysis to test for the presence of specific anions. Here are some common effects of lead (II) nitrate on certain anions such as halogens, sulphate sulphites and carbonate. Regarding the halogens, different colours of precipitates are formed with respect to the type of the halogen being tested.

$Pb(NO_3)_2 + 2Cl^-_{(aq)} \longrightarrow PbCl_{2(s)} + 2NO_3{}^-$

Lead(II) nitrate reacts with chloride ions to form a white precipitate of lead chloride ($PbCl_2$)

$Pb(NO_3)_2 + 2I^-_{(aq)} \longrightarrow PbI_{2(s)} + 2NO_3{}^-$

Lead(II) nitrate reacts with iodide ions to form a yellow precipitate of lead iodide (PbI_2)

$Pb(NO_3)_2 + 2Br^-_{(aq)} \longrightarrow PbBr_{2(s)} + 2NO_3^-$

Lead(II) nitrate reacts with bromide ions to form a creamy-white precipitate of lead bromide ($PbBr_2$).

$SO_4{}^{2-}{}_{(aq)} + Pb(NO_3)_{2(aq)} \longrightarrow PbSO_{4(s)} + 2NO_3^-$ Lead(II) nitrate reacts with sulphate ions to form a white precipitate of lead sulphate ($PbSO_4$)

$Pb(NO_3)_{2(aq)} + CO_3{}^{2-}{}_{(aq)} \longrightarrow PbCO_{3(s)} + 2NO_3^-$

Lead(II) nitrate reacts with carbonate ions to form a white precipitate of lead(II) carbonate ($PbCO_3$), which is slightly soluble in water.

It is essential to remember that the formation of these precipitates indicates the presence of the corresponding anions in the solution. Confirmatory tests or additional analyses may be necessary to verify the presence of specific anions, especially if multiple anions produce similar precipitates. Additionally, some anions might require specific pre-treatments or adjustments in pH before performing the test to ensure accurate results.

Learning Tasks

1. Provide learners with a set of unknown solutions containing common ions such as Cu^{2+}, $Fe^{2+}$and Al^{3+} and ask them to perform simple precipitation reactions to identify the ions present.
2. Provide learners with a mixture of two salts which contain at least one of the following ions: (Cu^{2+}, Fe^{2+}, Pb^{2+}, Al^{3+}, Zn^{2+}, Fe^{3+}) and anions (Cl^-, $SO_4{}^{2-}$, $CO_3{}^{2-}$), and ask them to perform qualitative analysis on an unknown solution containing a mixture of cations and write a lab report in a tabular form.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. **Talk-for-learning:**
   a) Start with a brief overview of solubility rules and why they are important in identifying ions in water-soluble compounds. Use real-world examples to illustrate the relevance of understanding solubility rules.

   b) Conduct a demonstration where you dissolve various known compounds in water and observe their solubility. Encourage learners to predict the solubility of each compound based on the rules they have learned.

   c) Divide the class into mixed-ability small groups and provide them with different water-soluble compounds. Task each group to perform tests to identify the ions present in their compounds based on solubility rules. Encourage collaboration and discussion within groups.

   d) Assist groups that are struggling.

2. **Inquiry-based learning:**
   a) In mixed-gender grouping where applicable, design an experiment for learners to test for the presence of the following cations (Al^{3+}, Ca^{2+}, Cu^{2+}, Fe^{2+}, Fe^{3+}, Zn^{2+}, Pb^{2+}, $NH_4{}^+$) using NaOH(aq) and NH_3(aq) as precipitating reagents. Support and guide as students apply their knowledge.

   b) While maintaining the groupings, provide more opportunities for hands-on practice by tasking learners to perform tests to identify and describe the behaviour of the following anions ($CO_3{}^{2-}$, Cl^-, Br^-, I^-, NO_3^-, $SO_3{}^{2-}$, $SO_4{}^{2-}$, S^{2-})

   c) Design an experiment for learners to determine the presence of salt using the grid method

d) Assess learners' understanding through quizzes, group presentations or individual assignments

Key Assessments (DoK)

1. **Level 1:** What are the general solubility rules for common ions in aqueous solutions?
2. **Level 2:** Imagine you have a solution containing an unknown compound. Describe a step-by-step procedure to identify the ions present in the compound using solubility rules. Include the tests you would perform and how you would interpret the results.
3. **Level 3:** A student has conducted a series of tests on a water-soluble compound and determined that it contains either chloride ions (Cl^-) or carbonate ions (CO_3^{2-}). Design an experimental procedure to identify which ion is present in the compound. Explain your rationale for each step of the procedure and how it will lead to a conclusive result.

References

1. SHS Chemistry Curriculum
2. Chang, R. (2008). General Chemistry: The Essential Concepts (5th ed.). McGraw Hill.
3. Nagwa. (n.d.). Identifying Ions in Aqueous Solutions Using Solubility Rules. Retrieved from https://www.nagwa.com/en/lessons/804197359707/

SECTION 6: PERIODIC PROPERTIES

Strand: **Systematic Chemistry of the Elements**

Sub-Strand: Periodicity

Content Standard: Demonstrate knowledge and understanding of how periodic properties change with atomic number and principal quantum number.

Learning Outcome: *Describe and explain the trends of periodic properties on the periodic table.*

INTRODUCTION AND SECTION SUMMARY

In this section, learners will explore the content standard of demonstrating knowledge and understanding of how periodic properties evolve with atomic numbers and quantum numbers. The learning outcome aims for learners to proficiently describe and explain the trends of periodic properties found on the periodic table. Through a series of tailored tasks, learners will engage with indicators such as utilising electron configurations to determine element placement on the periodic table and articulating the changes in periodic properties in relation to atomic numbers and quantum numbers.

SUMMARY OF PEDAGOGICAL EXEMPLARS

A range of pedagogical approaches including activity-based learning, collaborative learning, and talk for learning will be implemented to cater to varied learning preferences. Varied teaching strategies and resources were employed to cater for the diverse needs of learners. Conscious effort should be made to group learners according to mixed ability and help learners to respect group dynamics.

ASSESSMENT SUMMARY

Assessment strategies will gauge learners' grasp of the material at different levels of depth, ensuring a comprehensive understanding of periodic trends. By involving themselves in these activities and assessments, learners will master the periodic properties and their variations across the periodic table effectively. Support any learner who is struggling to understand the concept. Support learners who could not perform to expectation.

The weeks covered by the section are:

Week 17: *Periodicity and Periodic Table and Periodic law*

Week 17

Learning Indicators:

1. *Use the electron configuration of elements to determine their position on the periodic table.*
2. *Explain how periodic properties change with atomic number and principal quantum number.*

Theme or Focal Area: Periodicity and Periodic Table

Periodicity refers to the periodic or repeating trends in the chemical and physical properties of elements as they are arranged in increasing atomic numbers in the periodic table.

The periodic table is a tabular arrangement of elements, organised based on their atomic structure and chemical properties. The periodic table consists of a row of element, called periods, and a column of elements, called groups. Each element in the same group shares similar chemical properties while atoms generally decrease in size as we move across a period (due to the increasing nuclear charge attracting the electrons).

There are several types of elements in the periodic table, including metals, non-metals and metalloids (semi-metals), each with their characteristic properties and behaviour. The periodic table is an important tool and is widely used in the fields of chemistry, physics and material science.

Classification of elements according to the blocks (s, p, d)

Elements can be classified into four main blocks in the periodic table according to their electron configurations. These are the s-block, p-block, d-block and f-block. We will only focus on the s, p and d-blocks. The classification of elements into s, p and d-blocks is based on the configuration of their valence electrons and their position within the periodic table.

1. **s-block elements:** Elements in groups 1 and 2 are called s-block elements because their valence electrons are located in the s-subshell. These elements include metals like lithium, sodium, potassium and calcium, that have low electronegativity and are highly reactive. Helium is also a s-block element.
2. **p-block elements:** Elements in groups 13-18 (except for helium) are called p-block elements because their valence electrons are located in the p-subshell. These elements include non-metals like carbon, nitrogen and oxygen, as well as metalloids like boron and silicon. P-block elements generally exhibit a wide range of properties.
3. **d-block elements:** Elements in groups 3-12 are called d-block elements because their valence electrons are located in the d-subshell. These elements include iron, copper and gold, and are generally harder than s-block elements. D-block elements also exhibit a wide range of properties.

Classification of elements according to groups: (IUPAC system and Roman numeral system)

The classification of elements according to groups is based on their electron configurations. The most commonly used systems for classifying elements are the IUPAC system and the Roman numeral system.

1. **IUPAC system:** This system is used to classify elements based on the number of valence electrons they have in their outermost electron shell. The groups are numbered from 1 to 18, where group 1 contains elements with one valence electron, group 2 contains elements with two valence electrons, and so on until group 18, which contains elements with eight electrons. Here is a summary of how elements are classified in the IUPAC system:

(a) Group 1: 1 valence electron.
(b) Group 2: 2 valence electrons.
(c) Group 13: 3 valence electrons.
(d) Group 14: 4 valence electrons
(e) Group 15: 5 valence electrons.
(f) Group 16: 6 valence electrons.
(g) Group 17: 7 valence electrons.
(h) Group 18: 8 valence electrons (except helium, which has 2 valence electrons).

2. **Roman Numeral System:** This system is also used to classify elements based on their electron configuration. However, it assigns group numbers using Roman numerals, which are based on the number of valence electrons in the outermost shell and the subshell that contains the last electron. The numerals I to VIII are used to denote the number of outer shell electrons, and the letters A and B indicate the subshells in which the last electron is present. Here is a summary of how elements are classified using the Roman numerals:
(a) Group I: 1 outer shell electron.
(b) Group II: 2 outer shell electrons.
(c) Group III A: 3 outer electrons in subshell A.
(d) Group IV A: 4 outer electrons in subshell A.
(e) Group V A: 5 outer electrons in subshell A.
(f) Group VI A: 6 outer electrons in subshell A.
(g) Group VII A: 7 outer electrons in subshell A.
(h) Group VIII: 8 outer electrons (except helium, which has 2 outer electrons).

In both systems, elements within the same group exhibit similar chemical and valence electrons in the shell.

3. **Period of elements:** The periodic table is arranged in periods which are horizontal rows containing elements arranged in increasing atomic number from left to right. The first period (period 1) contains only two elements, hydrogen and helium. The second period (period 2) contains the elements with atomic numbers 3 to 10, and the third period (period 3) contains elements with atomic numbers 11 through 18.

Each period on the periodic table corresponds to a shell of electrons in an atom. For example, the first-period elements (hydrogen and helium) have only one shell of electrons while the second-period elements have two shells of electrons. The number of shells increases with each subsequent period. Some of the properties of elements within a period change gradually as the atomic number increases, e.g. atomic size. This is due to the increasing nuclear charge, gradual filling of electron shells and the changes in the valence electrons.

4. **Classification of elements according to metals, semi-metals, and non-metals:** Elements can be classified into three broad categories based on their properties. These are metals, non-metals and semi-metals (also called metalloids). The classification of elements into metals, non-metals and semi-metals is based on their physical and chemical properties and their position on the periodic table.
a) **Metals:** Metals are typically shiny, solid at room temperature (except for mercury), malleable (ability to be hammered into different shapes), ductile (ability to be drawn into wires) and good conductors of heat and electricity. They generally have high melting and

boiling points and tend to lose electrons to form positively charged ions (cations). Examples of metals include sodium, aluminium, copper, iron, lead, silver and gold.

b) **Non-metals:** Non-metals are typically dull, brittle (if solid) and poor conductors of heat and electricity. They generally have low melting and boiling points and tend to gain electrons to form negatively charged ions (anions). Examples of non-metals include oxygen, nitrogen, sulphur and carbon.

c) **Semi-metals:** Semi-metals have properties that are intermediate between metals and non-metals. They are usually solids at room temperature, but their properties vary widely. For example, they can be shiny or dull, conductors or insulators and brittle or ductile. Examples of semi-metals include boron, silicon, germanium, arsenic and antimony.

Physical properties (hardness, density, melting point, boiling point and physical state) of groups 1, 2 and noble gases.

Group 1 elements (alkali metals):

1. Hardness: They are relatively soft and can be easily cut with a knife.
2. Density: They have low densities compared to most metals.
3. Melting point: They have relatively low melting points.
4. Boiling point: They have relatively low boiling points.
5. Physical state: At room temperature, all alkali metals are solid.

Group 2 elements (alkaline earth metals):

1. Hardness: They are harder than alkali metals but softer than most metals.
2. Density: They have higher densities than the alkali metals.
3. Melting point: They have higher melting points than alkali metals but lower than most other metals.
4. Boiling point: They have higher boiling points than the alkali metals but lower than most other metals.
5. Physical state: All alkaline earth metals are solid at room temperature.

Noble gases:

1. Density: They have very low densities, particularly helium.
2. Melting point: They have very low melting points.
3. Boiling point: They have very low boiling points.
4. Physical state: All noble gases are gases at room temperature except for radon, which is a radioactive solid.

Chemical properties of groups 1, 2 and noble gases

The chemical properties of groups 1, 2 and noble gases are described below:

1. **Group 1 elements (alkali metals):**
   a) Reactivity: Group 1 elements have only one electron in their outermost shell, making it easy to lose that electron in a chemical reaction. Hence, they are highly reactive and readily give up their outermost electron to form cations with a charge of +1.
   b) Electronegativity: Alkali metals have low electronegativity and are highly electropositive. This makes them excellent reducing agents in chemical reactions.
   c) Reactivity with water: Alkali metals are so reactive that they can only be stored in oil or inert gases. They react vigorously with water to form hydroxides and hydrogen gas. When dropped into water, they float on the surface and release hydrogen gas with a hissing sound.

2. **Group 2 elements (Alkaline earth metals):**
   a) Reactivity: Alkaline earth metals are reactive, but less so than alkali metals. They have two electrons in their outermost shell and readily give up those two electrons to form cations with a charge of +2.
   b) Electronegativity: Alkaline earth metals have low electronegativity and are highly electropositive. This makes them good reducing agents in chemical reactions.
   c) Reactivity with water: Alkaline earth metals react with water, but less vigorously than alkali metals. They form hydroxides with a release of hydrogen gas.
3. **Noble gases:** Noble gases are usually chemically inactive because they have a complete outer shell of electrons. This makes them stable and non-reactive with other elements.

Learning Tasks

1. How does the electron configuration of an atom relate to the block in which the element is found on the periodic table?
2. How will the electron configuration be used to predict the group of an element on the periodic table?
3. State the properties of sodium based on the group it belongs to on the periodic table

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. **Activity-based learning:** Help learners to arrange atomic orbitals in order of increasing energy levels. Based on the number of electrons present in an atom, encourage learners to independently write the electron configuration of the first thirty elements based on the orbital notation.
2. **Collaborative Learning**
   a) In pairs, classify the elements according to the following categories:
      i. The blocks (s, p, d)
      ii. Groups (IUPAC system and the Roman numeral system).
      iii. the period in which the element belongs.
      iv. metals, semi-metals and non-metals.
   b) Describe the physical properties (hardness, density, melting point, boiling point and physical state) and chemical properties of some representative elements (groups 1, 2, 7 and the noble gases).
   c) Have the learners work in pairs to design the periodic table.

Key Assessments (DoK)

1. **Level 1:** Element X has electron configuration $1s^2\ 2s^2$
   a) State the block that X belongs to.
   b) Which group does X belong to?
2. **Level 2:** Consider the following elements: X and Y with atomic numbers, 17 and 11 respectively. Write the electron configuration of each element and use it to determine the:
   a) block to which each belongs.
   b) group to which each belongs.
   c) period to which each belongs.

3. **Level 2:** Construct a periodic table using the first 20 elements.

Theme or Focal Area: Periodic law

The periodic law is a statement that helps explain the repeating patterns of chemical and physical properties of elements as they are arranged in the periodic table. The periodic law states that “the properties of elements are periodic functions of their atomic numbers”. In other words, elements display periodic trends in some of their properties when arranged in order of increasing atomic number.

This pattern arises because the properties of elements are largely determined by the number and arrangement of electrons in their atoms, and the atomic number reflects both the number of protons and electrons in an atom. The periodic table organises elements in ways that allow us to see how electrons are arranged in the outermost or valence electron shell.

Periodic properties

A periodic property is a physical or chemical property that displays a repeating trend as one moves across the periodic table. These trends arise from the periodic variation in the electron configurations of elements.

Some examples of periodic properties are listed below:

1. **Atomic size:** Atomic size, also known as atomic radius, refers to the size of the atom. It is difficult to measure due to the uncertainty in the location of the outermost electron. It is sometimes defined as half the distance between the nuclei of two covalently bonded atoms (covalent radius).
2. **Ionic size:** Ionic radius is the distance from the nucleus of an ion up to which it influences its electron cloud.
3. **Ionisation energy:** It is the energy required to remove an electron from the outermost shell of each atom in one mole of gaseous atoms to form a cation.
4. **Electron affinity:** It is the energy change that occurs when each atom in one mole of atom in its gaseous state gains an electron to form an anion.
5. **Electronegativity:** Electronegativity is the ability of an atom to attract electrons towards itself in a chemical bond.

Factors that affect periodic properties

Factors that affect periodic properties include the nuclear charge, distance from the nucleus, shielding effects, electron configuration and nuclear size.

1. **Nuclear charge (atomic number):** The number of protons in the nucleus defines the nuclear charge of an atom. The greater the nuclear charge, the greater the pull on the electrons and the smaller the atomic radii as you go across a period. Consequently, ionisation energy and electronegativity both increase as one moves across a period on the periodic table.
2. **Distance from the nucleus (energy level or shell):** The distance that a valence electron is located from the nucleus is a significant factor that affects periodic properties. As the number of energy levels or shells an electron occupies increases, the electron is farther from the nucleus, decreasing the nuclear attraction to the outermost electrons.
3. **Shielding effects**: Shielding occurs when negatively charged valence electrons are shielded from attraction to the nucleus by inner electrons. The higher the number of inner shells of electrons, the weaker the attraction between the nucleus and the valence electron. Therefore, as the number of inner shells increases, the shielding effect increases. This effect makes it easier for outermost electrons to be removed and shields bonding electrons from the nuclear charge, causing ionisation energy and electronegativity to decrease as one moves down a group.

4. **Electron configuration:** The electron configuration of an atom, including the number and orientation of the electrons in the valence shell, affects periodic properties. A filled valence shell or half-filled subshells contribute to stability, making these configurations highly desirable. As a result, atoms will gain or lose electrons to achieve these stable configurations.

Variation of the periodic property in the periodic table

Different periodic properties show variations across the periodic table due to a variety of factors, such as atomic structure, chemical bonding, electronegativity and electron configurations.

Here are a few examples of periodic properties and their variations:

1. **Atomic radius:** Atomic radius refers to the distance between the nucleus and the outermost electrons. Generally, the atomic radius decreases from left to right across a period due to the increasing nuclear charge from left to right across a period, which attracts the electrons more strongly, making them more compact. This reduces the size of the atom. Conversely, atomic size increases down the group. This is because as one moves down a group, the number of energy levels (shells) increases, causing an increase in the distance between the valence electrons and the nucleus. As the number of core shells increases down the group, the shielding/screening effect increases. This decreases the nuclear attraction between the protons and the valence electrons that make the atom larger.
2. **Ionic radius:** When atoms gain or lose electrons to form ions, their sizes change. Cations (positively charged ions) are smaller than their parent atoms because they have lost electrons, which means less electron-electron repulsion and a smaller electron cloud. Anions (negatively charged ions) are larger than their parent atoms because they have gained electrons, resulting in more electron-electron repulsion and a larger electron cloud.
3. **Ionisation energy:** Ionisation energy tends to increase from left to right across a period because as you move from left to right across a period, there is an increase in nuclear charge which attracts the electrons more tightly, making it more difficult to be removed. Ionisation energy decreases down a group. This is because, down a group, there is an increase in atomic size and a decrease in nuclear charge which makes the valence electrons less bound to the nucleus, thereby making it easier to be removed.
4. **Electronegativity:** Electronegativity typically increases as you move from left to right across a period due to an increase in the effective nuclear charge. However, there are exceptions to this trend such as boron, which has a lower electronegativity than expected due to its partially filled p-orbital. Conversely, electronegativity decreases down a group. This is because down a group, the atomic radius increases, reducing the attraction of the nucleus and bonding electrons.
5. **Electron affinity:** Electron affinity tends to increase from left to right across a period and decreases from top to bottom within a group. This is because the closer the electron is to the electrostatic attraction of the nucleus, the more energy will be released when it is added. Meanwhile, within a group, the increasing atomic size reduces the electrostatic attraction of the nucleus, making it harder to add an electron.

Discrepancies in the periodic properties with respect to beryllium, boron, nitrogen and oxygen

Several factors can account for discrepancies in periodic properties with respect to beryllium, boron, oxygen nitrogen and oxygen.

1. **Size of the atoms and ions:** The size of an atom or ion affects many periodic properties. In general, atoms tend to be smaller as you go across a period due to an increase in the effective nuclear charge. However, there are some anomalies in this trend. For example, beryllium has a smaller size than boron, despite being to the left of it in the periodic table. This can be attributed to the fact that beryllium has a fully filled 2s orbital, which makes it more compact than boron, which has a partially filled 2p orbital with greater electron-electron repulsion.

2. **Electron configuration:** The electron configuration of an element can also affect its periodic properties. For example, nitrogen has higher ionisation energy than oxygen because nitrogen has a half-filled 2p orbital, which makes it more stable than the partially filled 2p orbital of oxygen. Also, the extra electron is shielded by the half-filled 2p orbital electrons. Hence, more energy is required to remove electrons from nitrogen than is required in oxygen.
3. **Nuclear charge and shielding:** Boron has lower ionisation energy than beryllium because the 2p electron in boron is easily removed, as it exhibits increased shielding from the nucleus due to the filled 2s orbital. The 2s electrons in beryllium are strongly attracted by the nucleus, hence increasing the nuclear charge. Thus, more energy is needed to remove the 2s electrons of beryllium.

Learning Tasks

1. Define and explain the periodic law in your own words.
2. Research how atomic size is defined and explain the factors that influence the size of an atom.
3. Describe the trends in the periodic properties of elements
4. Define electronegativity and explain its role in determining the nature of chemical bonds that are formed between atoms.
5. Identify the periodic properties that show similar trends across the period

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. **Talk for learning**
   a) Begin by introducing the concept of the periodic table and its significance in chemistry. Explain that the periodic table organises elements based on their atomic number, electron configuration and chemical properties.
   b) Engage students in a discussion about the structure of the periodic table. Encourage them to identify patterns and trends in the arrangement of elements.
   c) Discuss the significance of periods (rows) and groups (columns) in the periodic table.
   d) Explain the periodic law, stating that the properties of elements are periodic functions of their atomic numbers.
2. **Collaborative learning:**
   a) Use visual aids such as diagrams, charts and interactive simulations to help learners visualise the periodic table and its patterns. Highlight specific trends, such as atomic radius, ionisation energy and electronegativity, across periods and down groups.
   b) Divide learners into small groups and assign each group a specific periodic trend to investigate, such as atomic radius, ionization energy or electronegativity.
   c) Ask each group to prepare a short presentation or mini-lesson to teach their assigned periodic trend to the rest of the class and account for discrepancies in the periodic properties with respect to beryllium, boron, oxygen and nitrogen.
   d) Encourage learners to explain their findings, discuss the concepts and provide real-life examples to illustrate the trend.
   e) Conclude with a reflection of the session where learners discuss what they have learned and how their understanding has deepened.

Key Assessments (DoK)

1. **Level 1:** State the periodic law
2. **Level 2:** Distinguish between electronegativity and electron affinity.
3. **Level 2:** Explain why the first ionisation energy of oxygen is less than that of nitrogen.
4. **Level 2:** Arrange the following elements in increasing order of electronegativity and explain your order: C, F, Mg, Br, S, O, Na.
5. **Level 3:** Arrange the following species in order of decreasing size K^+, Cl^- and Ca^{2+}.
6. **Level 3:** Explain what is meant by ionisation energy and state how it varies across a period in the periodic table.

References

1. SHS Chemistry Curriculum
2. Zumdall, S.S., & DeCoste (2017). Chemistry. Cengage Learning.
3. Chang, Raymond (2008). General Chemistry - Essential Concepts. 5th Edition, The McGraw Hill Companies
4. Royal Society of Chemistry. (n.d). The Periodic Table. https://www.rsc.org/periodic-table KhanAcademy. (n.d). Periodic Trends.
5. https://www.khanacademy.org/science/chemistry/periodic-table/periodic-table-trends-bonding/e/periodic-trends

SECTION 7: INTERATOMIC BONDING

Strand: **Systematic Chemistry of the Elements**

Sub-Strand: Bonding

Content Standard: Demonstrate knowledge and understanding of the formation and properties of interatomic bonding.

Learning Outcome: *Predict and explain ionic, covalent and metallic bonding, as well as their characteristic properties.*

INTRODUCTION AND SECTION SUMMARY

This section focuses on demonstrating knowledge and understanding of the formation and properties of interatomic bonding. The content standard is centred on predicting and explaining ionic, covalent and metallic bonding along with their characteristic properties. Various tailored tasks will be utilised to cover the content effectively.

SUMMARY OF PEDAGOGICAL EXEMPLARS

The section will incorporate pedagogical approaches like activity-based learning, exploratory learning, digital learning and talk for learning. These diverse methods aim to engage all learners and enhance their comprehension accordingly. Learners can expect an interactive and dynamic learning experience.

ASSESSMENT SUMMARY

Varied assessment strategies will be employed to cater for the diverse needs of learners. Also, key assessment strategies with varying depths of knowledge will be implemented to find out if the learning indicators and learning outcomes have been achieved. Support learners who have not performed to meet expectations.

The weeks covered by the section are:

Week 18: *Ionic bonding*

Week 19: *Covalent bonding and Metallic Bonding*

Week 18

Learning Indicator: *Explain ionic bonding and its formation, and state the properties of ionic compounds.*

Theme or Focal Area: Ionic bonding

Chemical bonds are formed whenever two or more atoms are held strongly together. They do so in a particular way, which tends to give the constituent atoms specific whole number ratios. Gilbert Lewis, an American scientist, explained phenomenally that atoms combine to achieve a more stable electron configuration. Maximum stability is attained when an atom loses, gains or shares electrons to achieve a similar electron configuration to that of a noble gas (inert gas).

In chemical bond formation, only the outermost shells of the constituent atoms come into contact. So, only the valence electrons are involved in bond formation. The Lewis dot symbol, which consists of the symbol of an element and a dot for one valence electron in an atom, helps to explain how atoms interact to form chemical bonds. Elements in the same group on the periodic table have the same number of valence electrons, hence similar Lewis' dot symbols, except for their atomic symbols. The constituent atoms in a molecule can either donate, accept or share a certain number of electrons to form a specific type of chemical bond. The type of bond created depends on the electronegativity difference between the atoms involved.

Atoms connect with each other through chemical bonds. Electronegativity strongly influences how atoms interact with each other and how they bond. In fact, the electronegativity difference between two bonded atoms determines the nature of the chemical bond that forms between them. If the electronegativity difference is large (greater than approx. 1.7), the bond that forms between the atoms will be an ionic bond, and if it is small (less than approx. 1.7), a covalent bond will generally form.

1. **Ionic bond**

   It is an electrostatic force of attraction formed between a positive ion and a negative ion. Usually, the less electronegative element completely transfers its valence electron(s) to the more electronegative element.

   This exchange results in a more stable, noble gas electron configuration for both atoms involved. An ionic bond is based on attractive electrostatic forces between two ions of opposite charge.

   Ionic bonds can also be formed between species that are already ions rather than neutral atoms forming ions e.g. precipitation reactions.

2. **Cations and Anions**

   Ionic bonds involve a cation and an anion. The bond is formed when an atom, typically a metal, loses an electron or electrons and becomes a positive ion, or cation. One example of an ionic bond is the formation of sodium fluoride, NaF, from a sodium atom and fluorine. In this reaction, the sodium atom loses its single valence electron to the fluorine atom, which has just enough space to accept it.

   $Na \rightarrow Na^+ + e^-$ ***and*** $F + e^- \rightarrow F^-$

   The ions produced are oppositely charged and are attracted to one another due to electrostatic forces. $Na^+ + F^- \rightarrow NaF$

3. **Factors that affect ionic bond formation**
   i. Low ionisation energy of the metallic element which forms the cation.
   ii. Large electron affinity of the non-metallic element which forms the anion.
   iii. Large lattice energy i.e., the smaller size and higher charge of the ions.

The electronegativity difference of the elements affects the bonding of atoms. Elements with high electronegativity tend to form ionic bonds with elements of low electronegativity.

4. **Properties of ionic compounds**
    i. Ionic compounds have high melting and boiling points.
    ii. Some dissolve in polar solvents like water.
    iii. They conduct electricity in the aqueous form or molten state (as the ions are free to move).
    iv. In their solid form, they serve as good insulators (as the ions cannot move).
    v. They are hard and brittle in nature.

Learning Tasks

1. How are ionic compounds held together?
2. Use the Lewis dot structures to describe how $MgCl_2$ is formed.
3. Explain why $NaCl_{(s)}$ does not conduct electricity even though it is an ionic compound.
4. Explain why ionic compounds have high boiling or melting points.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. **Talk for learning:**
    a) Guide learners to review the meaning of chemical bonding and the types of interatomic bonding (ionic, covalent and metallic) from the JHS curriculum through questions and answers.

       Encourage slow learners to participate in the review by asking them questions.
    b) Divide learners into mixed-ability groups and guide them to discuss the factors that affect ionic bond formation (ionisation energy, electron affinity, electronegativity).

       Encourage slow learners to participate in the discussion by asking them questions.
    c) In mixed-ability groups, guide learners to use illustrations to explain the formation of cations and anions and relate the charge on simple ions to the group number of the element on the periodic table. Encourage slow learners to participate actively by asking them questions.
2. **Activity-based learning**:
    a) In mixed-ability groups, guide learners to use atomic models or simulations or electron dots to describe and explain the formation of ionic bonds between metals and non-metals. Encourage slow learners to participate actively by paying more attention to them.
    b) In mixed-ability groups, guide learners to use models to illustrate and describe the formation of sodium chloride crystals. Encourage all learners to participate actively by paying more attention to them.
    c) In mixed-ability groups, guide learners to discuss the properties of ionic compounds.

Key Assessment

1. **Level 1:** State two properties of ionic compounds.
2. **Level 2:** Sodium oxide, Na_2O, is an ionic compound formed when sodium reacts with oxygen. Describe, in terms of electrons, what happens when sodium oxide is formed in this reaction.
3. **Level 2:** Which of the following pairs of compounds is more ionic?

    a) $BeCl_2$ and $AlCl_3$

b) Li_2O and CaO. Give reasons for your answer.

4. **Level 2:** Describe how calcium and oxide ions are formed, and using the Lewis dot structures, describe how calcium oxide is formed.

References

1. SHS Curriculum
2. Chang, Raymond (2008). General Chemistry - Essential Concepts. 5th Edition, The McGraw Hill Companies

Week 19

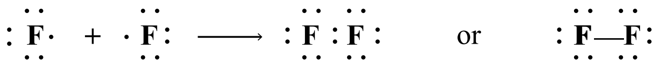
Learning Indicator: *Explain covalent bonding and its formation, and state the properties of covalent compounds.*

Theme or Focal Area: Covalent bonding

1. **Covalent bond:** A covalent bond is a chemical bond that is formed when two atoms mutually share a pair of electrons. Covalent bonds are usually found between non-metal atoms. By doing so, the atoms attain a stable duplet or octet electronic configuration. In covalent bonding, overlapping of the atomic orbitals having one electron from each of the two atoms takes place, resulting in the sharing of the pair of electrons. Generally, the orbitals of the electrons in the valence shell of the atoms are used for electron sharing. For example, in a hydrogen molecule (H_2), a covalent bond is formed by the overlap of the two s-orbitals each containing one electron from each of the two H atoms of the molecule. Each electron in a shared pair of electrons is attracted to the nuclei of both atoms. Each H atom attains a $1s^2$ configuration. $\mathbf{H\cdot + \cdot H \rightarrow H : H}$

   Covalent bonding between many compounds which comprise atoms beyond the first shell involves only the valence electrons. In the covalent bond formation by two fluorine atoms in F_2, there are seven (7) valence electrons, which signifies that only one (1) unpaired electron exists on a fluorine atom

   Only two (2) valence electrons participate in the formation of the covalent bond in F_2, leaving six (6) valence electrons in each atom not involved in forming the covalent bond. The pairs of electrons that are not used in bonding are called lone pairs. Therefore, F_2 has six lone pairs of electrons (3 lone pairs on each F atom).



Source: (*Raymond Chang and Jason Overby, 2009)*

2. **Pure covalent bond:** When the atoms that form a covalent bond are identical, as in H_2, Cl_2 and other diatomic elements, then the electrons in the bond are shared equally. We refer to this as a pure covalent bond**.**

   Electrons shared in pure covalent bonds have an equal probability of being near each nucleus.

3. **Dative bond or coordinate covalent bonding:** This is a type of covalent bond in which both electrons in the bond are donated by one atom. That is, the pair of electrons that is shared between the two atoms comes from only one of the bonding atoms This type of bonding is different from typical covalent bonds, where each atom contributes one electron to form a shared pair.

4. **Formation of dative bonds or coordinate covalent bonding:**
   a) Formation of Ammonium ion (NH_4^+). The reaction between an ammonia molecule (NH_3) and a proton (H^+) to form an ammonium ion (NH_4^+). $NH_3 + H^+ \rightarrow NH_4^+$. In ammonia, three hydrogen atoms combine directly with the nitrogen atom by normal covalent bonding. The nitrogen atom has a lone pair of electrons in its outermost shell. It is acting as a proton acceptor, i.e. can combine with the hydrogen ion. It shares the lone pair of electrons with a proton from an acid to produce the ammonium ion, NH_4^+. The proton carries over its positive charge to give the ammonium ion, NH_4^+.
   b) Formation of a hydroxonium ion (H_3O^+). $H_2O + H^+ \rightarrow H_3O^+$. In a water molecule, two hydrogen atoms share two pairs of electrons with an oxygen atom by normal covalent

bonding. The oxygen atom in a water molecule has two lone pairs of electrons in its outermost shell. It shares this with a proton (hydrogen ion, H^+) from an acid to produce the hydroxonium ion H_3O^+. The positive charge on the hydrogen ion is carried over to give the positively charged hydroxonium ion, H_3O^+.

5. **Polar covalent bond:** Polar covalent bond is a type of chemical bond formed between two atoms, where the electrons are shared unequally. In this type of bond, the bonding electrons are more attracted to one atom than the other, giving rise to a shift of electron density toward that atom. The atom that attracts the electrons more strongly acquires a partial negative charge whilst the other acquires a partial positive charge. For example, the electrons in the H–Cl bond of a hydrogen chloride molecule spend more time near the chlorine atom than near the hydrogen atom.

   Whether a bond is nonpolar or polar covalent is determined by a property of the bonding atoms called electronegativity. Electronegativity is a measure of the tendency of an atom to attract electrons (or electron density) towards itself. It determines how the shared electrons are distributed between the two atoms in a bond. The more strongly an atom attracts the electrons in its bonds, the larger its electronegativity. Electrons in a polar covalent bond are shifted toward the more electronegative atom; thus, the more electronegative atom is the one with the partial negative charge. The greater the difference in electronegativity, the more polarised the electron distribution and the larger the partial charges of the atoms.

6. **Factors that affect the formation of covalent bonds**
   a) Number of valence electrons
   b) High ionisation energy
   c) Comparable electron affinity
   d) Comparable electronegativities
   e) Atomic size
   f) High nuclear charge and small internuclear distance.

7. **Properties of covalent compounds:** Compounds that contain covalent bonds exhibit different physical properties than ionic compounds, such as:
   a) Covalent compounds exist in all states of matter namely solid, liquid and gas.
   b) Covalent compounds do not conduct electricity in solid, molten or aqueous state.

8. **Polarisation of ion:** During the formation of an ionic compound or ionic molecules, two oppositely charged ions (cations and anions) must come closer to each other. During this process, the cation attracts the electron charge cloud of the outermost shell of the anion toward itself. Therefore, the symmetrical shape of the anion gets distorted, deformed or polarised.

   The phenomenon of the distortion of the symmetrical shape of the electron cloud of anion in the nearby cation is called polarisation of anion.

   There is also the chance of the polarisation of a cation by an anion. However, due to the smaller size of the cation, its electron cloud is strongly held to the nucleus, and the shape of the cloud is not distorted to an appreciable extent. Hence, the polarisation of a cation by an anion is not generally considered.

9. **Polarising power:** The ability of a cation to polarise (distort) an anion is called its polarising power or polarising ability.

10. **Factors affecting polarising power:** The factors affecting the magnitude of a cation's polarising power are listed below:
    a) Magnitude of positive charge on the cation

b) Size of cation

i. **Magnitude of positive charge on the cation:** The greater the charge on the cation, the more strongly it attracts the outermost shell electron cloud of an anion toward itself and polarises the given anion easily. Therefore, the polarising power of a cation is directly proportional to the magnitude of the positive charge on it. E.g. $Na^+ < Mg^{2+} < Al^{3+}$. If the same element has a different positive charge, the higher positive charge has greater power of polarisation. E.g., $Sn^{4+} > Sn^{2+}$

ii. **Size of cation:** The smaller the size of the cation, the more strongly it attracts the outermost shell electron cloud of an anion towards itself, hence, the greater its polarising ability. In other words, with the decreasing size of the cation, the polarising power of the cation increases. Thus, the polarising power of the cation is inversely proportional to the size of the cation. Example: Li+> Na+> K+> Rb+

11. **Polarisability of anions:** The tendency of an anion to get polarised by a cation is called its polarisability. The factors affecting the polarisability of an anion are magnitude of the negative charges on the anion and the size of the anion
    a) **Magnitude of negative charges on anion:** The higher the negative charge on the anion, the more easily its outermost electron cloud is attracted by cations, hence an anion is directly proportional to the magnitude of the negative charges on it. E.g., $C^{4-} > N^{3-} > O^{2-} > F^-$
    b) **Size of the anion**: The larger the size of the anion, the more easily its outermost shell electron cloud is attracted by the cation towards itself, hence, the greater the polarisability of an anion. Thus, the polarisability of an anion is directly proportional to the size of the anion.

Learning Tasks

1. How are covalent compounds formed?
2. How are valence electrons represented in Lewis dot structures?
3. Use Lewis dot structures to explain the nature of chemical bonding found in oxygen gas (O_2) and identify the number of bonds formed in the molecule
4. What differentiates a polar covalent bond from a non-polar covalent bond?

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

2. **Talk for learning:**
    a) Through whole class discussion, guide the learners to review the previous knowledge on covalent bonding from the JHS curriculum. Encourage all learners to actively take part in the discussion by asking them questions.
    b) In mixed-ability groups, guide learners to identify and distinguish between the types of covalent bonds (simple covalent, dative or co-ordinate and polar covalent bonds)
3. **Exploratory learning:**
    a) In mixed-ability groups, guide learners to use atomic models to explain the formation of covalent bonds between different non-metals. Special attention must be given to visually impaired learners.
    b) In mixed-ability groups, guide learners to use electron dot structures or models to illustrate the formation of simple (H_2), dative ($NH4^+$) and polar (HF) covalent bonds. Special attention must be given to visually impaired learners.

4. **Initiate talk for learning:**
    a) In mixed-ability groups, guide learners to discuss and explain the following terms: polarisation, polarizability and polarising power, and state the factors that introduce ionic character into covalent bonds.
    b) In mixed-ability groups, guide learners to:
        i. use electronegativity differences between atoms to predict bond type.
        ii. discuss the properties of covalent compounds.
        Encourage slow learners to participate in the discussion by asking them questions.

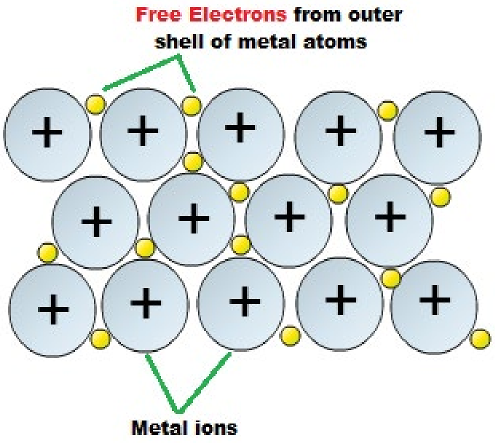
Key Assessments (DoK)

1. **Level 1:** Define a covalent bond.
2. **Level 1:** State the type(s) of bonds in NH_4^+.
3. **Level 2:** Describe how covalent bond is formed in the following molecules: H_2, N_2, Cl_2 and O_2
4. **Level 2:** Distinguish between polar covalent bonds and dative bonds.

Theme or Focal Area: Metallic bonding

1. **Metallic bond:** Metallic bond is a type of chemical bond that occurs in metals. It is an electrostatic force of attraction between the fixed positive metal ions and the delocalised electrons around the cations. The lattice structure of the atoms is held together by a sea of delocalised valence electrons. The delocalised electrons are referred to as a sea of electrons or mobile electrons.

    Metallic bonds are formed by the process of metal atoms transferring their outermost electrons to an electron sea, which is the collection of shared electrons that surround the positively charged atomic cores of the metal atoms. When heated, the metal atoms release their outermost electrons, which then move freely throughout the lattice. Since metal atoms have very low electronegativity, they tend to lose their valence electrons, readily forming positive ions in the process. These positive ions are then surrounded by a cloud of delocalised electrons, which form the metallic bond that holds the ions together in a regular lattice structure.



    The strength of the metallic bond depends on factors such as the number of valence electrons in the metal atom, the size of the atoms and the proximity of the atoms in the lattice. Metallic bond strength increases as the number of valence electrons increases. Thus, aluminium is harder than magnesium, which is in turn harder than sodium because 3, 2, and 1 valence electrons are attracted by the fixed positive lattice points respectively. As the atomic size of these atoms reduces with increased nuclear attraction, their melting points also increase accordingly.

2. **Properties of metals**
   a) Malleable
   b) Ductile
   c) Good conductors of heat and electricity

Learning Tasks

1. Arrange the following metals in order of increasing metallic bond: Na, K, Mg and Al.
2. Give reasons for your order in question 1 above.
3. Draw a diagram to show the metallic bonding in magnesium ions.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. **Use of Digital Learning:** Guide learners to watch a video or listen to a presentation (with the aid of relevant charts) on metallic bonding. Special attention must be given to visually impaired learners and those with hearing challenges.
2. **Activity-Based learning:**
   a) Based on the presentation on metallic bonding, divide learners into mixed-ability groups and guide them to design a mind map on the explanation of metallic bonding, factors that affect its formation, as well as its properties. Special attention must be given to visually impaired learners.
   b) Guide learners to individually design models for metallic bonding (lattice of positive ions in a pool of electrons).

   Special attention must be given to visually impaired learners.
3. **Talk for Learning:** In mixed-ability groups, guide learners to:
   a) explain each of the factors that affect metallic bond formation.
   b) discuss the properties of metallic solids and link the properties to metallic bonding. Encourage all learners to actively partake in the discussion by asking them questions.

Key Assessments (DoK)

1. **Level 1:** What is a metallic bond?
2. **Level 2:** State and explain two factors that affect the strength of metallic bonds.
3. **Level 2:** Describe the metallic bonding in sodium metal and use it to explain how it conducts electricity.

References:

1. SHS Curriculum
2. Polarization of ion-polarizing power-and polarizability,’’Retrieved [April 16, 2024] from https://chemistnotes.com/inorganic/polarization-of-ion-polarizing-power-and-polarizability/#’’
3. Raymond Chang and Jason Overby. (2009). *General Chemistry, the essential concepts*. New York: McGraw-Hill.

SECTION 8: INTERMOLECULAR BONDING

Strand: **Systematic Chemistry of the Elements**

Sub-Strand: Bonding

Content Standard: Demonstrate knowledge and understanding that the type of chemical bond in a compound determines the physical and chemical properties of that compound.

Learning Outcome: *Predict and describe the type of intermolecular bonds that will be formed between a group of compounds.*

INTRODUCTION AND SECTION SUMMARY

This section looks at the types of intermolecular forces, how they arise from molecular structures and how these forces influence the physical properties of compounds.

In Week 20, the focus is on describing the types of intermolecular forces and compounds that exhibit these forces. Week 21 shifts towards explaining how these forces affect the properties of compounds. The learning outcome is for students to predict and describe intermolecular bonds.

SUMMARY OF PEDAGOGICAL EXEMPLARS

Tailored tasks will guide learners through the content. Incorporating pedagogical methods, such as talk for learning, experiential learning and collaborative learning, to foster engagement and understanding for all learners. Varied resources will be employed to cater for the diverse needs of learners. Conscious effort should be made to group learners according to mixed ability and help learners to respect group dynamics.

ASSESSMENT SUMMARY

Assessment strategies will gauge learners' grasp of the material at different levels of depth, ensuring a comprehensive understanding of periodic trends. By involving themselves in these activities and assessments, learners will master the concept of intermolecular forces effectively. Support any learner who is struggling to understand the concept. Support learners who could not perform to expectation.

The weeks covered by the section are:

Week 20: *Intermolecular Bonding*

Week 21: *Effects of intermolecular forces on physical properties of compounds*

Week 20

Learning Indicator: *Describe the types of intermolecular forces and explain how they arise from the structural features of molecules*

Theme or Focal Area: Intermolecular Bonding

Intermolecular bonds are attractive forces that exist between molecules. This is different from intramolecular forces that hold atoms together in a molecule, e.g. covalent bonds. Covalent molecules are usually joined together by intermolecular bonds. These bonds arise because of the presence of electric charges, dipoles and hydrogen bonds. The most common types of intermolecular bonds include dipole-dipole forces, hydrogen bonds and induced dipole–induced dipole forces. These intermolecular forces are types of Van der Waals attractions. The strength of a covalent bond is, however, larger than an intermolecular bond that joins the molecules together. Generally, the covalent bond strength of molecules ranges from 50-200 kJ/mol while that of intermolecular forces ranges from 1-12 kJ/mol. Intramolecular forces stabilise individual molecules while intermolecular bonds determine a range of properties of substances, including melting and boiling points, solubility and viscosity. For a substance to boil, it needs enough energy to overcome the attractive forces that hold the molecules together before they can enter the vapour phase. Therefore, when given the boiling points of molecular substances, it is reflective of the strength of intermolecular forces that hold molecules of the compound.

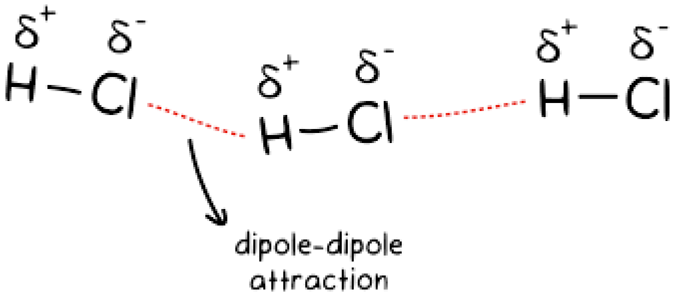
Van der Waal Forces

Covalent molecules are generally held together by Van der Waals forces of attraction. The attractive forces are brought about by unlike charges that are permanently created or induced based on the atoms that make up the compound. There are three forms of Van der Waals forces based on the charge distribution:

1. Dipole-dipole interactions
2. Hydrogen bonding
3. Induced dipole–induced dipole (London dispersion) forces

Dipole - dipole interaction

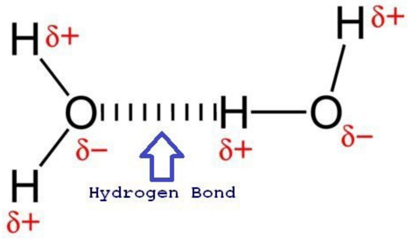
Dipole-dipole interaction, also known as dipolar interaction, is a type of intermolecular force that occurs between two polar molecules. Polar molecules have a permanent dipole moment, which is a separation of positive and negative charge due to differences in electronegativity between the atoms in the molecule. Since each atom has a different affinity for electrons, the 'push and pull' of their shared electrons results in one atom maintaining most of the electron density and a partial negative charge, leaving the other atom with a partial positive charge. In dipole-dipole interactions, the positive end of one molecule is attracted to the negative end of another, which can result in an attractive force between them.



Dipole-dipole interactions are generally weaker than covalent bonds but stronger than London dispersion forces, which are another type of intermolecular force (which will be discussed later). Dipole-dipole interactions are responsible for many physical and chemical properties of polar molecules. They also play a role in determining the solubility of polar molecules in polar solvents, as the dipole-dipole interactions between the molecule and solvent can make the molecule more soluble. The presence of dipole-dipole interactions can also affect the geometry and stability of molecules in solid and liquid states.

Hydrogen Bond

It is a special type of dipole-dipole interaction where a hydrogen atom bonded to a small but highly electronegative atom, such as oxygen, nitrogen, or fluorine, is attracted to another electronegative atom in a nearby molecule. The electronegative atom creates a partial negative charge while the hydrogen atom creates a partial positive charge. The partially positive hydrogen atom is attracted to the partially negative electronegative atom in another molecule, resulting in a hydrogen bond.



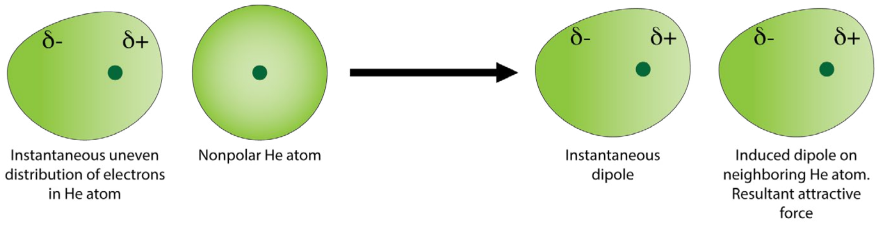
Hydrogen bonds are stronger than London dispersion forces and dipole-dipole forces but weaker than covalent bonds. They play a significant role in determining the properties of water, ice and many biomolecules, such as DNA and proteins. The unique properties of water, including its high boiling point, surface tension and density, are due to the presence of hydrogen bonding between water molecules. In DNA, hydrogen bonding between complementary base pairs facilitates the double helix structure of the molecule. Some examples of compounds that contain hydrogen atoms bonded to highly electronegative atoms such as oxygen, nitrogen and fluorine include:

1. Water (H_2O)
2. Ammonia (NH_3)
3. Hydrogen fluoride (HF)
4. Methanol (CH_3OH)
5. Ethanol (C_2H_5OH)
6. Acetic acid (CH_3COOH)
7. Phenol (C_6H_5OH)
8. Sugars (such as glucose and fructose and their macromolecules)
9. Proteins (such as collagen and keratin)
10. RNA

These compounds have unique properties that make them important in many different fields, such as medicine, chemistry and biology.

Induce dipole–induce dipole (London dispersion) forces

It is a kind of intermolecular bond that exists between non-polar covalent molecules due to fluctuations in electron density, which gives rise to uneven charge distribution. The uneven charge distribution causes temporary dipoles to be created. This then induces a dipole in a nearby molecule, which results in their attraction to each other.



Examples of non-polar molecules which have weakly induced dipole-induced dipole intermolecular forces are diatomic molecules (O_2, H_2, l_2, Br_2), noble gases (He, Ar, Ne), CO_2, CH_4, CCl_4, polythene and rubber.

Learning Tasks

1. Distinguish between intramolecular and intermolecular forces and give examples.
2. Explain the formation of hydrogen bonds, using the structural formula of two identical molecules.
3. Identify molecules that are capable of forming hydrogen bonds and explain the conditions under which hydrogen bonding occurs.
4. Define dipole–dipole interactions and compare them to other intermolecular forces.
5. The molecular compound of propanol is $CH_3CH_2CH_2OH$. Explain the nature of intermolecular forces that will exist in a 5L sample of propanol.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. **Talk for Learning:**
   a) With the aid of charts, models or other resources, differentiate between intramolecular and intermolecular forces. Explain intermolecular bonding and use concept maps to guide learners to identify the types of intermolecular forces found in compounds (covalent compounds). Note: Ionic compounds are linked by inter-ionic or electrostatic forces
   b) Watch a video or use a chart to reinforce the types of intermolecular forces in molecules.
2. **Collaborative learning:**
   a) Using think-pair-share, describe the three main types of intermolecular forces: Hydrogen bonding, dipole-dipole interactions and London dispersion forces.
   b) Define each type of interaction and provide examples to illustrate their significance.
   c) Using charts of molecular mass (molecular size, number of electrons per molecule) of the halogens and their boiling points, deduce and explain the factors that affect the strength of London dispersion forces of attraction.
   d) In pairs, and with the aid of charts or any other data, visualise and deduce the factors that affect the strength of hydrogen bonds.
      i. Electronegativity and size of elements directly bonded to the hydrogen.
      ii. Number of hydrogen bonds per molecule.
      iii. Orientation of the hydrogen bond.

Key Assessments (DoK)

1. **Level 2:** Deduce the type of intermolecular force present in each of the following compounds:
   a) H_2S
   b) NH_3
   c) CH_4
   d) SF_6
   e) HF
2. **Level 2:** In terms of intermolecular forces, explain why the solubility of ethanol in water is greater than the solubility of ether in water.
3. **Level 3:** Consider the following compounds: CH_4, CO_2, NH_4Cl. Select the compound capable of undergoing dipole-dipole interactions and explain your answer.

References

1. SHS Curriculum
2. Chang, Raymond (2008). General Chemistry - Essential Concepts. 5th Edition, The McGraw Hill Companies
3. https://phet.colorado.edu/en/simulations/molecule-polarity
4. https://studiousguy.com/wp-content/uploads/2021/02/Dipole-Dipole-Interaction.png retrieve on 16-04-2024
5. https://www.bing.com/images/blob?bcid=qNckYDUvc-oGqNT-eFsvrTYgA2jh.....48 retrieve on 16-04-2024
6. https://chem.libretexts.org/Bookshelves/Introductory_Chemistry/Chemistry_for_Allied_Health_(Soult)/05%3A_Properties_of_Compounds/5.03%3A_Polarity_and_Intermolecular_Forces

Week 21

Learning Indicator: *Explain how intermolecular forces affect the physical properties of compounds.*

Theme or Focal Area: Effects of intermolecular forces on physical properties of compounds

The understanding of intermolecular forces helps us to understand and explain the physical properties of substances since it is intermolecular forces that account for physical properties such as boiling points, melting points, viscosities, etc.

Intermolecular forces give rise to a number of structural features and properties of covalent compounds.

1. **Boiling and Melting point:** The strength of intermolecular forces directly influences the amount of energy required to change the state of a substance, leading to variations in melting and boiling points. Substances with stronger intermolecular forces tend to have higher melting and boiling points. This is because more energy is required to break these strong attractions between molecules. Conversely, substances with weaker intermolecular forces typically have lower melting and boiling points. It takes less energy to overcome these weaker attractions.

   Generally, different types of intermolecular forces have varying strengths:

   i. Induced dipole-induced dipole (London dispersion) forces are the weakest.

   ii. Dipole-dipole interactions are stronger than induced dipole-induced dipole interactions.

   iii. Hydrogen bonding, which is a special type of dipole-dipole interaction, is the strongest.

   E.g. Substances like water, with strong hydrogen bonding, have higher melting and boiling points compared to substances like methane, which only have induced dipole-induced dipole interaction. Comparing molecules of similar size and shape, those with hydrogen bonding will generally have higher melting and boiling points than those with only dipole-dipole or induced dipole-induced dipole interaction.

2. **Solubility:** The ability of a substance, known as the solute, to dissolve in a solvent to form a homogeneous mixture called a solution. Solubility depends on several factors which include:

   i. **Like Dissolves Like Principle:** The solubility of a substance in a solvent is largely determined by the similarity of their intermolecular forces. This principle states that "like dissolves like", which means that substances with similar types of intermolecular forces tend to dissolve in each other.

   ii. **Polarity**: Polar solvents, such as water, have molecules with uneven charge distribution, leading to dipole-dipole interactions. Substances with polar molecules, like alcohols, tend to be soluble in polar solvents due to the attraction between polar molecules.

   iii. **Hydrogen Bonding:** Hydrogen bonding is a special dipole-dipole interaction between molecules containing hydrogen bonded to highly electronegative atoms like oxygen, nitrogen or fluorine. Substances capable of hydrogen bonding, such as alcohols, amines and carboxylic acids, tend to be more soluble in solvents that can also participate in hydrogen bonding.

   iv. **Non-polar Substances:** Non-polar substances, like hydrocarbons (e.g., oils, fats), have weak-induced dipole-induced dipole forces between molecules. They are generally soluble in non-polar solvents such as hexane or benzene, due to the similar nature of their intermolecular forces.

   v. **Temperature:** In general, increasing temperature can increase solubility by providing more energy to overcome intermolecular forces. However, this trend may not always hold true for all substances, as changes in temperature can also affect the solubility of substances differently, depending on their specific intermolecular forces.

3. **Surface tension:** Surface tension is a property of liquids that arises due to the cohesive forces between molecules at the surface of the liquid. Intermolecular forces, such as hydrogen bonding, dipole-dipole interactions and induced dipole-induced dipole forces, play a crucial role in determining the magnitude of surface tension.
   i. Liquids with strong intermolecular forces, such as water (which exhibits hydrogen bonding), have high surface tension. This is because the cohesive forces between molecules at the surface are strong, causing the surface to resist being stretched or broken.

   ii. Liquids with weaker intermolecular forces, such as nonpolar molecules like hexane or pentane, have lower surface tension. In these liquids, the cohesive forces between molecules at the surface are weaker, allowing the surface to be more easily stretched or broken.

   iii. Increasing the temperature of a liquid generally decreases its surface tension. This is because higher temperatures increase the kinetic energy of molecules, causing them to move more rapidly and reducing the strength of the intermolecular forces holding them together at the surface.

4. **Enthalpy of Vaporisation:** Stronger intermolecular forces lead to higher enthalpies of vaporisation. This is because substances with stronger intermolecular forces require more energy to overcome these forces and transition from the liquid phase to the gas phase. For e.g., substances with hydrogen bonding, such as water, have high enthalpies of vaporisation due to the strong attraction between molecules in the liquid phase. In contrast, substances with weaker intermolecular forces, such as noble gases, have low enthalpies of vaporisation because they require less energy to transition to the gas phase.

5. **Viscosity:** Intermolecular forces also affect the viscosity of a liquid, which is its resistance to flow. Substances with stronger intermolecular forces generally have higher viscosities because the molecules are more strongly attracted to each other and resist movement. E.g., liquids with extensive hydrogen bonding, such as honey or syrup, have high viscosities and flow slowly. In contrast, liquids with weaker intermolecular forces, such as water or alcohol, have lower viscosities and flow more easily.

6. **Volatility:** Stronger intermolecular forces result in lower volatility, as molecules are more tightly held together and require more energy to overcome these forces and enter the vapour phase. Conversely, weaker intermolecular forces lead to higher volatility, as molecules can more easily break free from each other and transition into the vapour phase.

Learning Tasks

1. Provide learners with a list of substances (e.g., water, ethanol, methane etc.) and ask them to identify the physical properties of each substance, such as boiling point, melting point and state at room temperature (solid, liquid, gas).
2. Present learners with two substances known to have different intermolecular forces, such as water (hydrogen bonding) and methane (London dispersion forces or induced dipole-induced dipole forces), and ask them to compare and contrast the physical properties of these substances, focusing on factors like boiling point, melting point and viscosity.
3. Provide learners with a series of molecules and their structures and ask them to predict which molecule would have the highest boiling point and justify their prediction using their understanding of intermolecular forces. Guide students to consider factors such as molecular size, polarity and hydrogen bonding.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

Talk for learning:

1. Start by engaging learners in a whole-class discussion about the concept of intermolecular forces and their importance in determining the physical properties of substances. Encourage learners to share their prior knowledge and experiences related to the topic.
2. Divide learners into pairs of small mixed-ability groups and provide them with a list of physical properties affected by intermolecular forces (solubility, density, viscosity, etc.). Ask each group to discuss how they think intermolecular forces influence each property. Then, have them share their ideas with the class.
3. Provide real-world examples or case studies where the effects of intermolecular forces on physical properties are evident. For instance, discuss how the intermolecular forces in water contribute to its high surface tension and boiling point, making it an effective solvent.
4. Allow learners to take on the role of the teacher by having them explain the relationship between intermolecular forces and physical properties to their peers. Encourage them to use examples and analogies to make the concepts more accessible.
5. Incorporate interactive quizzes during the lesson to gauge learners' understanding of key concepts related to intermolecular forces and physical properties in real-time. Use multiple-choice questions, short-answer prompts or concept-matching activities to assess learners' grasp of the concept and provide immediate feedback to guide their learning.

Key Assessment (DoK)

1. **Level 1:** Which type of intermolecular force is responsible for the high boiling point of water?
2. **Level 2:** Compare and contrast the physical properties of ethanol (C_2H_5OH) and methane (CH_4) in terms of their intermolecular forces. Explain why ethanol has a higher boiling point than methane.
3. **Level 3:** Design an experiment to investigate the relationship between the strength of intermolecular forces and the surface tension of different liquids. Include a hypothesis, materials list, procedure and expected results.
4. **Level 4:** Analyse the impact of intermolecular forces on the solubility of various substances in water. Consider how factors such as molecular structure, polarity and hydrogen bonding influence solubility. Provide examples to support your analysis.

References

1. SHS Curriculum
2. Chang, R. (2008). General Chemistry: The Essential Concepts (5th ed.). McGraw Hill.
3. PhET Interactive Simulations. (n.d.). Retrieved from https://phet.colorado.edu/
4. Ryan, L., & Norris, R. (2014). Cambridge International AS and A-level Chemistry Coursebook. University Printing House.

SECTION 9: QUALITATIVE AND QUANTITATIVE ANALYSIS OF ORGANIC COMPOUNDS

Strand: **Chemistry of Carbon Compounds**

Sub-Strand: Characterisation of Organic Compounds

Content Standard: Demonstrate knowledge and understanding of the general processes involved in qualitative and quantitative elemental analysis of organic compounds.

Learning Outcome: *Apply the knowledge and understanding in science to describe qualitative and quantitative elemental analysis of organic compounds.*

INTRODUCTION AND SECTION SUMMARY

This session looks at qualitative and quantitative analysis of organic compounds. The content standard emphasises demonstrating an understanding of the processes involved in this analysis. By the end of the section, learners are expected to apply their scientific knowledge to describe both qualitative and quantitative analysis of organic compounds. Week 22 will specifically focus on outlining these analyses, while Week 23 will challenge learners to design and carry out experiments to detect and measure the presence of carbon, hydrogen, sulphur, nitrogen and halogens in organic compounds. Tailored tasks will be utilised throughout the section to facilitate learning.

SUMMARY OF PEDAGOGICAL EXEMPLARS

Various pedagogical approaches such as digital learning, experiential learning and inquiry-based learning will be used to engage learners. Where possible, engage learners in practical activities to maximise learning. Provide support systems; peer-to-peer support and teacher-to-learner support system. Encourage learners to respect each other’s views.

ASSESSMENT SUMMARY

Varied assessment tools will be used to support teaching and learning. Prominent among them are quizzes, assignments and projects/practical/laboratory works. The assessment strategies will vary in complexity to ensure that all learners participate in the learning process to meet the learning outcome. Support any learner who is struggling to understand the concepts.

The weeks covered by the section are:

Week 22: *Methods of separation and purification of organic compounds*

Week 23: *Test for carbon and hydrogen in an organic compound and Quantification of carbon, hydrogen and halogens in organic compounds*

Week 22

Learning Indicator: *Describe qualitative and quantitative analysis of organic compounds*

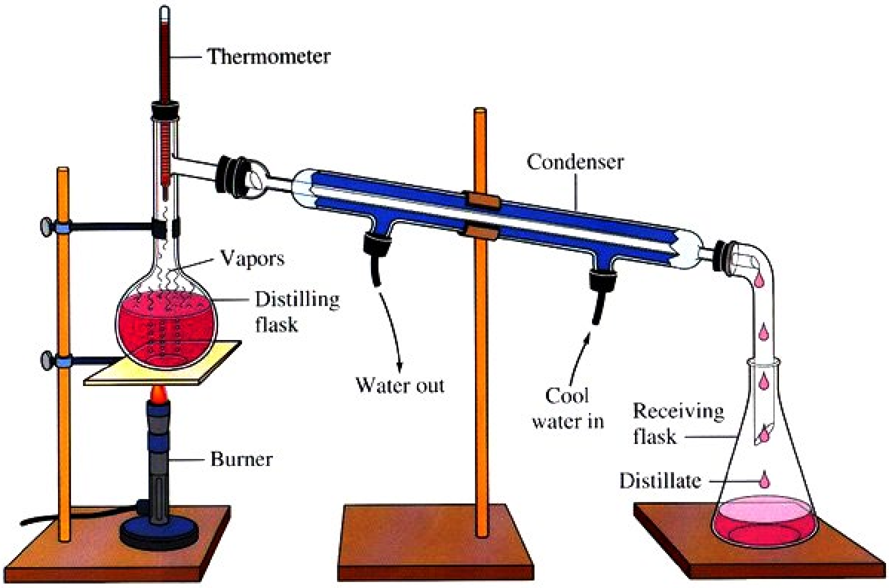
Theme or Focal Area: Methods of separation and purification of organic compounds

Most substances naturally occur in the form of mixtures. These mixtures can be in different forms based on the physical and chemical properties of the components that make up the mixture. Isolating a specific component from a mixture is possible when that component possesses a property markedly different from the other components. The greater the difference in the property, the more effective the separation process. Properties such as boiling point, melting point, solubility, surface adsorption and chemical reactivity can serve as the basis for separating mixtures. Below are some of the methods used for the separation and purification of organic compounds.

1. Distillation

It is used to separate volatile liquids from a mixture of non–volatile liquids or solids. In this method of separation, the components of the mixture must have wide differences in their boiling points. This method of separation involves two physical processes, namely:

- **a)** The liquid in the mixture (solution) boils and evaporates when heated (liquid → gas).
- **b)** Vapour passes through a condenser, where it is condensed back to liquid (distillate) and collected in a flask (gas → liquid).
- **c)** Any dissolved solid (residue) is left in the solution in the flask due to the high boiling point of the solid.
- **d)** The distillation apparatus usually consists of a distillation flask, a condenser and a receiving flask. A hot plate or a Bunsen burner is used to heat the distillation flask which contains the mixture to be separated.



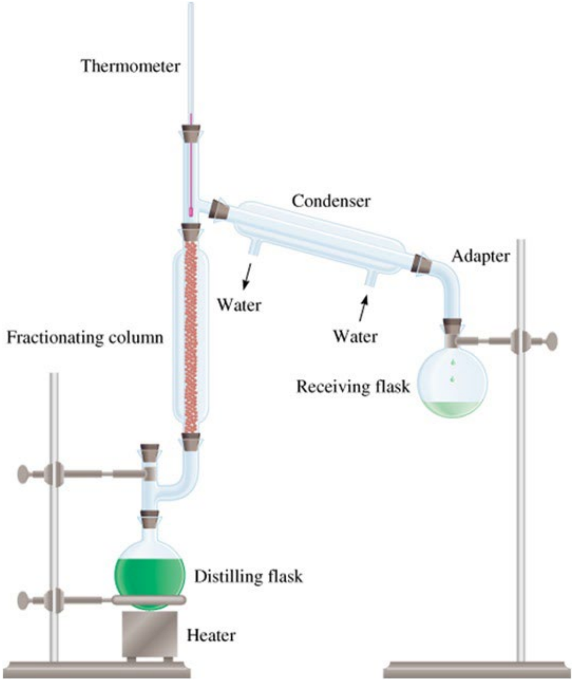
Practical Application

- **a.** It is used to recover or purify seawater.
- **b.** It is used in the separation of ethanol from a mixture of ethanol and water.
- **c.** Purification of essential oils from plants.
- **d.** Production of distilled water for laboratory use

2. Fractional distillation

This method is used to separate two or more miscible liquids, having similar (close) boiling points. The process involves two main physical processes.

- **a)** Solution mixture is boiled to vaporise the most volatile compound in the mixture (liquid → gas).
- **b)** The vapour mixture is passed through a fractionating column, which separates the mixture.
- **c)** The vapour is cooled by cold water in a condenser, which condenses it back to liquid (gas → liquid).
- **d)** In the laboratory, it is carried out in a fractionating column which is attached to the top of a distillation flask and connected to a condenser. The fractionating column is made up of a glass tube packed with glass or porcelain beads. This provides a greater surface area upon which repeated vaporisation and condensation can occur, leading to greater separation efficiency.



Practical Application

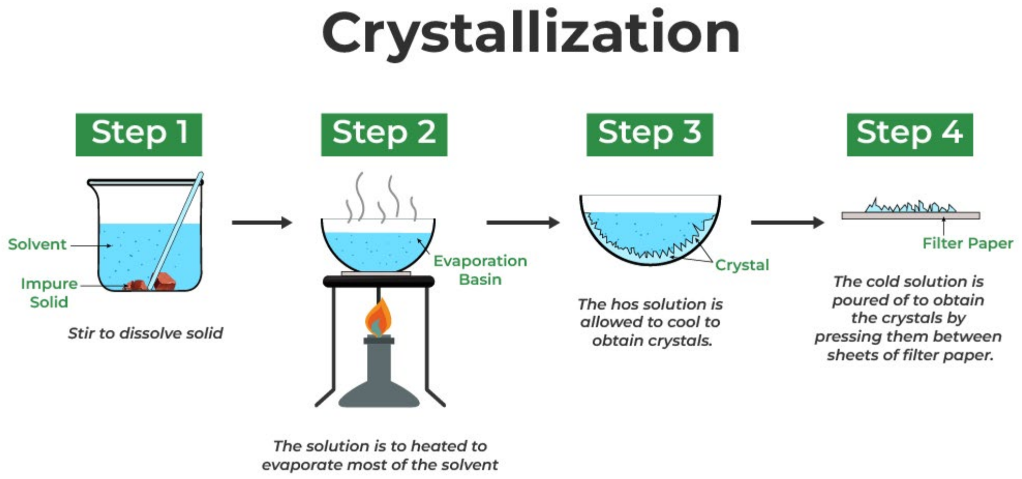
- **a.** Distillation of methanol-ethanol mixtures.
- **b.** Distillation of crude oil into useful fractions.
- **c.** Isolating oxygen gas from liquid air

3. Crystallisation

It is an essential purification technique under which an atom or molecule arranges itself in a well-defined three-dimensional lattice that minimises the overall energy of the system.

- **a)** It is a purification method used to remove soluble impurities from a solid compound.
- **b)** It involves dissolving the solid compound in a solvent (in which the impurities must also be soluble) and evaporating the liquid from the solution to a point beyond the solubility limit of the dissolved solid.
- **c)** Solid crystals grow out of the solution due to the solution being too concentrated for the solid to remain dissolved at that temperature.
- **d)** On cooling, the hot concentrated pure crystals form as the solubility gets less and the solubility limit is exceeded (as the impurities are at a lower concentration, they remain in the solution).

e) Crystals are filtered and dried in a desiccator.



Practical Application

Substances that can be isolated by crystallisation include:

- **a.** Benzoic acid
- **b.** Glucose
- **c.** Paracetamol

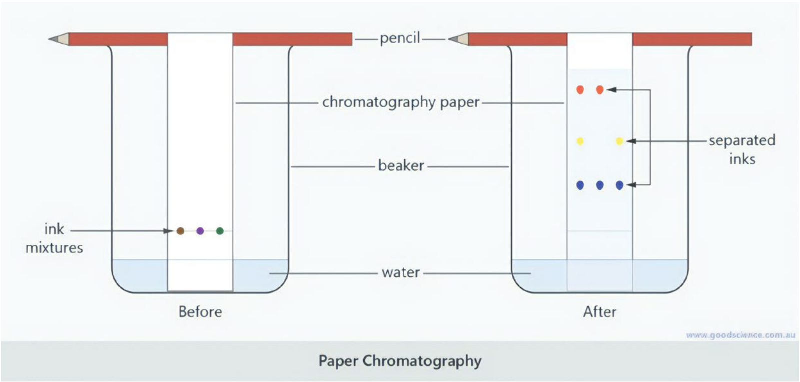
4. Paper chromatography

It is a method separating a mixture based on the different partition coefficients of the components in a mixture corresponding to the relative solubilities the mobile phase moves over the stationary phase (adsorbent paper).

- **a)** The substances in the mixture have different affinities for the solvent (mobile phase) and stationary phase (water trapped in the paper's structure), giving rise to different rates of movement over the paper.
- **b)** The highest point reached by the solvent is called the solvent front.
- **c)** The final result of the separation of the components is called a chromatogram.
- **d)** The Rf value of each component is the ratio of how far the spot travels relative to the distance moved by the solvent front.

$$\text{Rf} = \frac{\text{Distance moved by dissolved substance spot}}{\text{Distance moved by solvent front}}$$

The conditions of same temperature and same solvent used must be quoted in reference to the Rf data table to identify the sample.



Paper Chromatography

Applications in everyday life:

Practical application of the technique is used:

- **a)** in food dye analysis
- **b)** to identify substances
- **c)** in forensic analysis
- **d)** to analyse and identify natural products

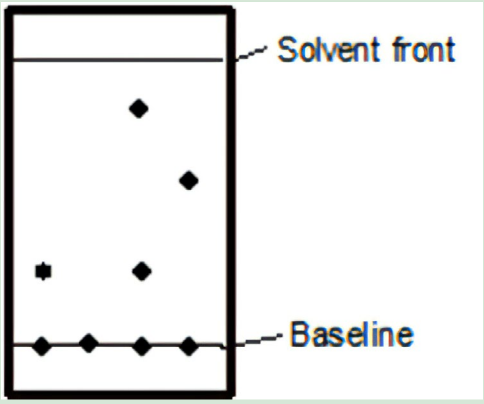
Learning Task

1. Define distillation and explain its purpose in the separation of mixtures.
2. Sketch the set-up of a simple distillation apparatus and explain the role of each component in the distillation process.
3. How will you recover and recycle a solvent containing dissolved solid impurities for an industrial process?
4. Consider the boiling points of the following alkanols

Alkanol	Boiling point (°C)
Methanol	65
Ethanol	79
1-propanol	97
1-butanol	117

   - **a)** Suggest the appropriate method of separating the alkanols
   - **b)** Which alkanol will be collected first? Explain your answer.

5. Explain crystallisation and outline its importance
6. What are the steps involved in crystallisation?
7. A student used paper chromatography to separate colour dye. The following is the chromatogram for the dyes A, B, C and D.



   - **a)** Deduce one conclusion that can be drawn about dye C.
   - **b)** Suggest why B remained on the baseline.
   - **c)** Calculate the R_f values of dye D and A

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. Using digital learning strategies, let learners watch video clips or observe teacher's demonstration of:
   a) Distillation
   b) fractional distillation
   c) crystallisation.

   Engage learners in mixed-ability groups to discuss the key steps in using crystallisation, drying and distillation to purify a given impure organic compound.

2. Let learners explore the use of melting and boiling points to determine the purity of a given organic compound using an inquiry-based learning approach.

3. Experiential Learning: Demonstrate the determination of melting points of some organic solids (benzoic acid, oxalic acid, ethanamide)

   Describe paper chromatography as an analytical technique that separates components in a mixture of organic compounds and state its uses in everyday life, such as in forensics, natural products, environmental analysis, etc.

   Analyse and interpret simple paper chromatograms, including the use of Rf-values.

Key Assessments (DoK)

Level 1: What is meant by the term crystallisation?

Level 1:

(a). What is paper chromatography?
(b). How can you identify substances in a mixture using paper chromatography?

Level 2: Explain Rf-values and their significance in paper chromatography. How can you calculate the Rf-value of a particular component?

Level 3: The green pigment chlorophyll in leaves of plants can be obtained using the following procedures:

Step 1: The neem leaf is ground with alcohol until the solution is saturated

Step 2: The green pigment obtained is separated from the mixture

Step 3: The colours in the pigment are then separated.

a. Name the apparatus used in step 1.
b. Suggest why the leaves are ground with alcohol instead of water in step 1
c. Name the type of separation method used in step 2.
d. Describe how step 3 is carried out

Level 4: A label on a can of green drink contains no artificial colours. Plan an investigation to show that the green colour of the drink was not a mixture of these two artificial colours. Blue E125
Yellow E 110

You are provided with the green colour from the drink, samples of E 125 and E 110 and some common laboratory apparatus

References

1. SHS Chemistry Curriculum
2. Chemistry. (2015) Science Resource Centre Handbook, Itec Global
3. Chang, Raymond (2008). General Chemistry - Essential Concepts. 5th Edition, The McGraw Hill Companies
4. RoseMarie Gallagher & Paul Ingram. (2014). Complete Chemistry for Cambridge IGCSE. 3rd ed. Oxford University Press, Great Clarendon Street, Oxford OX2 6DP
5. https://www.docbrown.info/page01/ElCpdMix/EleCmdMix2.htm
6. Crystallization. “Retrieved [April 18, 2024] from https://www.bing.com/images/blob?bcid=qFaFwhxnReoGqNT-eFsvrTYgA2jh.....5M
7. Chromatography. “Retrieved [April 18, 2024] from https://eduspots.org/wp-content/uploads/2020/06/paper-chromatography-diagram-1024x486.jpg
8. Distillation. “Retrieved [April 18, 2024] from https://www.bing.com/images/blob?bcid=qGoe6cjfXeoGqNT-eFsvrTYgA2jh.....3w

Week 23

Learning Indicator: *Design and experiment a test for the presence and mass composition of carbon, hydrogen, sulphur, nitrogen and halogens in organic compounds*

Theme or Focal Area: Test for carbon and hydrogen in an organic compound

Procedure

a. Organic sample is heated in the presence of the oxidising agent copper (II) oxide.

b. Carbon present is converted into carbon dioxide.

c. Hydrogen presence is converted into water.

d. Presence of carbon dioxide is detected by passing it through lime water, which turns milky.

e. Presence of water is detected by passing it through white anhydrous copper (II) sulfate which turns blue.

Test for Nitrogen, Sulphur and halogens in organic compounds

Procedure

Use Lassaigne sodium fusion test to convert N, S and halogens into inorganic ions.

Test for Sulphur: Add a few drops of freshly prepared sodium nitroprusside to the filtrate from the sodium fusion test. A purple or violet colour formation indicates the presence of sulphur in organic compounds.

Test for nitrogen:

a. Add a few drops of dilute NaOH to the filtrate of the sodium fusion test.

b. Heat and cool solution.

c. Add iron sulfate solution followed by drops of iron (III) chloride.

d. Formation of Prussian blue precipitate indicates the presence of nitrogen.

Test for halogens:

a. Add excess concentrated nitric acid to the filtrate from the sodium fusion test and heat.

b. Cool the solution and aqueous $AgNO_3$ solution

c. Formation of a white precipitate shows chlorine is present.

d. Formation of cream precipitate shows bromine is present.

e. Formation of yellow precipitate shows iodine is present.

LearningTask

1. How will you test for the presence of nitrogen and sulphur in an organic compound?
2. Describe how you will test for chlorine in an organic compound.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

Experiential learning: In small mixed-ability groups, design and conduct experiments to test for the presence of carbon, hydrogen, sulphur, nitrogen and halogens in organic compounds.

Key Assessments (DoK):

Level 1: Identify the colour change when testing for chlorine in an organic compound by adding excess concentrated nitric acid to the filtrate from a sodium fusion test after heating and then $AgNO_3$ is added.

Level 2: List the steps involved in the determination of nitrogen in an organic compound.

Level 3: How can N, S and halogens be converted into inorganic ions?

Level 4: Design an experiment to determine the presence of sulphur in an organic compound.

Theme or Focal Area: Quantification of carbon, hydrogen and halogens in organic compounds

Procedure:

- **a.** Convert known mass of dry sample to carbon dioxide and water by heating it over an oxidising agent.
- **b.** Water produced is passed through a pre-weighed U-tube containing magnesium perchlorate and the carbon dioxide is passed through a pre-weighed U-tube containing sodium hydroxide.
- **c.** The mass of carbon and hydrogen present are calculated as follows:

$$\text{Mass of carbon} = \frac{12}{44} \times \text{mass of } CO_2 \text{ produced}$$
$$\text{Mass of hydrogen} = \frac{2}{18} \times \text{mass of } H_2O \text{ produced}$$

- **d.** Calculate the percentage composition by mass of C and H.
- **e.** Determine the empirical formula and the molecular formula.

Estimation of halogen

- **a.** Use Carius method
- **b.** Use the formula:

$$\text{Mass of X} = \frac{\text{Relative atomic mass of X}}{\text{Molecular mass of AgX}} \times \text{mass of AgX produced}$$

Example:

0.95g of an organic compound on combustion in pure oxygen gave 2.85g of CO_2 and 1.50 g of water.

Calculate the percentage of carbon and hydrogen present. [C = 12, H = 1]

Solution: Use the problem-solving strategy

(a) Analyse the question

Known

Mass of organic compound = 0.95 g

Mass of CO_2 produced = 2.85 g

Mass of H_2O produced = 1.50 g

Ar. of C = 12, H = 1

Unknown

% C and % H =?

(b). Solve: Apply the problem-solving strategy

Mass of carbon $= \frac{12}{44} \times$ mass of CO_2 produced

Mass of carbon in the sample $= \frac{12}{44} \times 2.85$

$= 0.778$ g of carbon

$$\text{Percentage of carbon} = \frac{\text{mass of carbon}}{\text{mass of organic compound}} \times 100\%$$

$$\%\ of\ C = \frac{0.778}{0.95} \times 100$$

$$= 81.89\%$$

Mass of hydrogen in the sample $= \frac{2}{18} \times 1.50$

$= 0.1665$ g of hydrogen

$$\text{Percentage of hydrogen} = \frac{\text{mass of hydrogen}}{\text{mass of organic compound}} \times 100$$

$$\%\ of\ H = \frac{0.1665}{0.95} \times 100$$

$$= 17.52\%$$

Learning Tasks

1. Describe how to test for carbon and hydrogen in a hydrocarbon compound
2. Outline the procedure to test for nitrogen, sulphur and halogens in an organic compound
3. Explain the process that is used for the estimation of carbon, hydrogen and halogen in organic compounds, etc.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

Inquiry-based learning:

a. Describe how the mass of elements (C, H, X) are obtained.

b. Perform calculations involving percentage composition using secondary data and review the calculation of empirical and molecular formulae.

Key Assessments (DoK)

Level 1:

a. Name the reagents used to test for the presence of carbon in an organic compound.

b. Define the empirical formula of a compound.

Level 2: If the molar mass of a compound, X, is given, explain how to determine the molecular formula from the empirical formula

Level 3: An organic compound, Q, contains C, H, O and N. A mass of 0.132 g of Q was burnt completely in oxygen to produce 0.072 g of water, 0.176 g of carbon (IV) oxide and 24.0 cm3 of nitrogen.

- **i.** Calculate the empirical formula of Q
- **ii.** If the molar mass of Q is 132 gmol-1. Deduce the molecular formula of Q

  [O = 16, N = 14, C 12, H = 1, Vm = 22400 cm3 mol-1]

DoK level 4: Sample Q is a crystal obtained from the bark of a tree. Plan an investigation to show that carbon, hydrogen and nitrogen are contained in the sample. You are provided with common laboratory apparatus.

References

1. SHS Chemistry Curriculum
2. Chang, Raymond (2008). General Chemistry - Essential Concepts. 5th Edition, The McGraw Hill Companies

SECTION 10: CLASSIFICATIONS OF ORGANIC COMPOUNDS

Strand: **Chemistry of Carbon Compounds**

Sub-Strand: Organic Functional groups

Content Standard: Demonstrate knowledge and understanding of organic chemistry to classify organic compounds.

Learning Outcome: *Predict and classify organic compound*

INTRODUCTION AND SECTION SUMMARY

This session focuses on the chemistry of carbon compounds, with specific emphasis on differences between organic and inorganic compounds, classes of organic compounds, homologous series and properties of homologous series. The session will last for only one week. Throughout the week, learners will be engaged in activities designed to help them demonstrate their knowledge and understanding of organic chemistry in order to effectively classify organic compounds. The learning outcome of the session is for learners to accurately predict and classify a variety of organic compounds.

SUMMARY OF PEDAGOGICAL EXEMPLARS

To facilitate this learning, pedagogical exemplars such as inquiry-based learning, collaborative learning, activity-based learning and talk for learning will be utilised. These methods will encourage active participation and deeper understanding among the students. Where possible, engage learners in practical activities to maximise learning. Provide support systems; peer-to-peer support and teacher-to-learner support system. Encourage learners to respect each other's views.

ASSESSMENT SUMMARY

Assessment strategies of varying complexities will be employed to evaluate learners' grasp of the material and ensure that the learning outcome is successfully achieved. Prominent among them are quizzes, assignments and projects/practical/laboratory works. The assessment strategies will vary in complexity to ensure that all learners participate in the learning process to meet the learning outcome. Support any learner who is struggling to understand the concept.

The week covered by the section is:

Week 24: *Organic chemistry and Homologous series*

Week 24

Learning Indicators:

1. *Distinguish between organic and inorganic compounds and classify organic compounds.*
2. *Explain homologous series and state their properties.*

Theme or Focal Area: Organic chemistry

Organic chemistry is a branch of chemistry which studies the structure and properties of carbon compounds, except oxides of carbon, carbonates, carbides and cyanides.

An organic compound is a compound containing carbon atoms covalently bonded to other atoms.

Properties of carbon that make it possible to form many stable compounds

1. Carbon has a valency of four (tetravalent) and can form four covalent bonds.
2. Carbon atoms can join together to form straight chains, branched chains and ring structures.
3. Carbon can form single, double or triple bonds with itself and other elements
4. Carbon can form isomers.

Differences between organic and inorganic compounds

Inorganic compounds	Organic compounds
Contain any element except organic carbon.	Must contain a carbon atom.
Are usually ionic compounds and some covalent compounds.	Are usually covalent compounds.
Usually have relatively high melting and boiling points.	Usually have relatively low melting and boiling points.
Often soluble in polar solvents.	Usually soluble in nonpolar solvents.
Usually occur as solids at room temperature.	Often exist as liquids and gases.

Classes of organic compounds

Organic compounds can be classified into:

1. Aliphatic hydrocarbons
2. Alicyclic hydrocarbons
3. Aromatic hydrocarbons
4. Heterocyclic compounds

Hydrocarbons

Hydrocarbons are organic compounds containing only carbon and hydrogen atoms. Hydrocarbons can be saturated or unsaturated.

Saturated hydrocarbons have single bonds between the carbon and hydrogen atoms. Examples are alkanes.

Unsaturated hydrocarbons have double or triple bonds between two carbon atoms. Examples are alkenes and alkynes.

Aliphatic hydrocarbons:

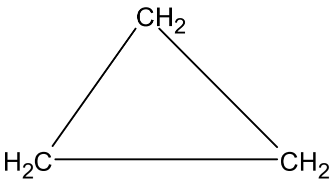
These are organic compounds containing carbon and hydrogen atoms that are usually linked together in chains via single, double or triple bonds. They can be open straight chains or branched chains. Aliphatic hydrocarbons may be saturated or unsaturated.

Examples are alkanes (which have $C - C$ bonds), alkenes (which have $C = C$ bonds) and alkynes (which have $C \equiv C$ bonds).

Many of the aliphatic compounds are flammable and so they are used as fuels such as butane in LPG and ethylene in welding.

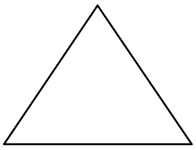
Alicyclic hydrocarbons: These are organic compounds that have closed rings of carbon atoms.

Examples are



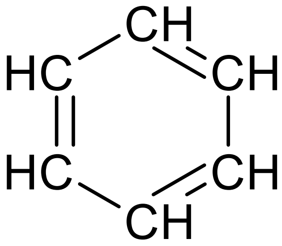
cyclopropane

or



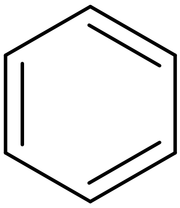
cyclopropane

Aromatic hydrocarbons: These are organic compounds that have one or more benzene rings in their structure. The different structures of benzene are shown below.



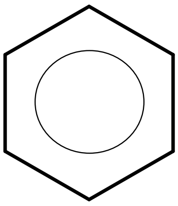
1

benzene



2

benzene

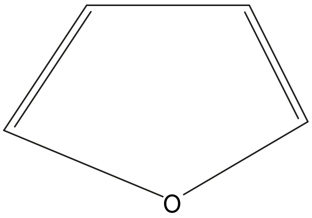


3

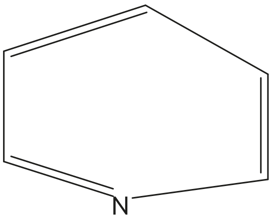
benzene

Heterocyclic compounds: These are organic compounds which are not hydrocarbons. They contain rings of atoms and carbon and hydrogen atoms, including other atoms such as oxygen, nitrogen and sulphur.

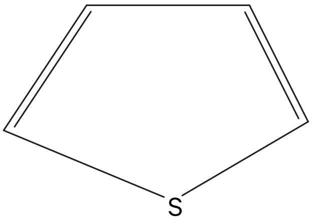
Examples are



furan



pyridine



thiophene

Learning Tasks

1. What are organic compounds?
2. State four reasons why carbon forms many stable compounds.
3. Distinguish between alicyclic hydrocarbons and aromatic hydrocarbons.

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. **Inquiry-Based learning:**
   a. Divide learners into mixed-ability groups and guide them to discuss the meaning of organic chemistry. Encourage slow learners to participate in the discussion by asking them questions.

   b. Guide learners to explain why carbon forms many compounds.

2. **Collaborative learning:**

   Through a whole class activity, guide learners to differentiate between organic compounds and inorganic compounds, and give examples.

   Encourage slow learners to participate in the discussion by asking them questions.

3. **Activity-based learning:**
   a. In mixed-ability groups, guide learners to use cut-outs on the types and examples of organic compounds to classify different organic compounds into aliphatic hydrocarbons, alicyclic hydrocarbons, aromatic hydrocarbons and heterocyclics.

   Special attention must be given to visually impaired learners.

   b. In small mixed-ability groups, guide learners to model the various classes of organic compounds and draw their structures.

Key Assessments (DoK)

1. **Level 1:** What is meant by the term aromatic hydrocarbon?
2. **Level 2:** In a tabular form, state the differences between organic compounds and inorganic compounds.
3. **Level 2:** Describe the general differences in chemical bonding between organic and inorganic compounds.
4. **Level 3:** Compare and contrast the physical and chemical properties of organic and inorganic compounds.

Theme or Focal Area 2: Homologous series

It is a group of compounds having the same general molecular formula and similar chemical properties, and in which each member differs from the neighbour in composition by CH_2.

The tables below show examples of homologous series for alkanes and alkenes.

1. **Alkanes:**

Number of carbon atoms (n)	Molecular formula, (C_nH_{2n+2})
1	CH_4
2	C_2H_6
3	C_3H_8
4	C_4H_{10}
5	C_5H_{12}

2. **Alkenes:**

Number of carbon atoms (n)	Molecular formula, (C_nH_{2n})
2	C_2H_4
3	C_3H_6
4	C_4H_8
5	C_5H_{10}
6	C_6H_{12}

Properties of a homologous series

1. Each member of the homologous series differs from the neighbour in composition by a CH_2 group.
2. They have a general molecular formula.
3. They have a general method of preparation.
4. They exhibit similar chemical properties.
5. They have the same functional group.
6. They exhibit a gradual change in physical properties along the series.

How to represent organic compounds

Organic compounds are represented by using their molecular formula, condensed formula or structural formula.

The tables below show examples of alkanes and alkenes

1. **Alkanes:**

Number of carbon atoms (n)	Molecular formula, (C_nH_{2n+2})	Structural formula
1	CH_4	CH_4
2	C_2H_6	CH_3CH_3
3	C_3H_8	$CH_3CH_2CH_3$
4	C_4H_{10}	$CH_3CH_2CH_2CH_3$
5	C_5H_{12}	$CH_3CH_2CH_2CH_2CH_3$

2. **Alkenes:**

Number of carbon atoms (n)	Molecular formula, (C_nH_{2n})	Structural formula
2	C_2H_4	$CH_2{=}CH_2$
3	C_3H_6	$CH_3CH{=}CH_2$
4	C_4H_8	$CH_3CH_2CH{=}CH_2$
5	C_5H_{10}	$CH_3CH_2CH_2CH{=}CH_2$
6	C_6H_{12}	$CH_3CH_2CH_2CH_2CH{=}CH_2$

Learning Tasks

1. What is a homologous series?
2. Write the molecular formulae of the first four members of the organic compound with the general molecular formula, C_nH_{2n-2}.
3. Draw the structures for the following organic compounds:
   a. $CH_3CH_2CH_3$
   b. $CH_3CH = CH_2$

Pedagogical Exemplars *(with the cross-cutting themes integrated)*

1. **Talk for learning:**

   Divide the learners into mixed-ability groups and use the inductive approach to guide them to deduce the meaning of homologous series.

   Encourage slow learners to participate in the discussion by asking them questions.

2. **Inquiry-based learning:**
   a. In mixed-ability groups, guide learners to write down the homologous series for alkanes / alkenes.
   b. Guide learners to discuss the properties of homologous series based on the examples.
   c. In mixed-ability groups, guide learners to demonstrate ways of representing organic compounds (molecular formula, condensed formula and structural formula) with the aid of charts, models, etc

Key Assessments (DoK)

1. **Level 1:** Explain the term homologous series.
2. **Level 1:** State three properties of a homologous series.
3. **Level 3:** The following compound is a member of the alkane homologous series: CH_3CH_3
   Draw the structures of the next three members of the series that follow the compound.

References

1. SHS Chemistry Curriculum
2. Addison – Wesley. (2000). *Chemistry*. 5th ed. Prentice Hall, New York
3. Morrison & Boyd. (1983). *Organic Chemistry*. 4 ed. Allyn & Bacon Inc. 7 Wells Avenue, Newton Massachusetts 02159.
4. Raymond Chang and Jason Overby. (2009). *General Chemistry, the essential concepts*. New York: McGraw-Hill.

ACKNOWLEDGEMENTS

Special thanks to Professor Edward Appiah, Director-General of the National Council for Curriculum and Assessment (NaCCA) and all who contributed to the successful writing of the Teacher Manuals for the new Senior High School (SHS), Senior High Technical School (SHTS) and Science Technology, Engineering and Mathematics (STEM) curriculum.

The writing team was made up of the following members:

NaCCA Team	
Name of Staff	**Designation**
Matthew Owusu	Deputy Director-General, Technical Services
Reginald Quartey	Ag. Director, Curriculum Development Directorate
Anita Cordei Collison	Ag. Director, Standards, Assessment and Quality Assurance Directorate
Rebecca Abu Gariba	Ag. Director, Corporate Affairs
Anthony Sarpong	Director, Standards, Assessment and Quality Assurance Directorate
Uriah Kofi Otoo	Senior Curriculum Development Officer (Art and Design Foundation & Studio)
Nii Boye Tagoe	Senior Curriculum Development Officer (History)
Juliet Owusu-Ansah	Senior Curriculum Development Officer (Social Studies)
Eric Amoah	Senior Curriculum Development Officer (General Science)
Ayuuba Sullivan Akudago	Senior Curriculum Development Officer (Physical Education & Health)
Godfred Asiedu Mireku	Senior Curriculum Development Officer (Mathematics)
Samuel Owusu Ansah	Senior Curriculum Development Officer (Mathematics)
Thomas Kumah Osei	Senior Curriculum Development Officer (English)
Godwin Mawunyo Kofi Senanu	Assistant Curriculum Development Officer (Economics)
Joachim Kwame Honu	Principal Standards, Assessment and Quality Assurance Officer
Jephtar Adu Mensah	Senior Standards, Assessment and Quality Assurance Officer
Richard Teye	Senior Standards, Assessment and Quality Assurance Officer
Nancy Asieduwaa Gyapong	Assistant Standards, Assessment and Quality Assurance Officer
Francis Agbalenyo	Senior Research, Planning, Monitoring and Evaluation Officer

NaCCA Team	
Name of Staff	**Designation**
Abigail Birago Owusu	Senior Research, Planning, Monitoring and Evaluation Officer
Ebenezer Nkuah Ankamah	Senior Research, Planning, Monitoring and Evaluation Officer
Joseph Barwuah	Senior Instructional Resource Officer
Sharon Antwi-Baah	Assistant Instructional Resource Officer
Dennis Adjasi	Instructional Resource Officer
Samuel Amankwa Ogyampo	Corporate Affairs Officer
Seth Nii Nartey	Corporate Affairs Officer
Alice Abbew Donkor	National Service Person

Subject	Writer	Designation/Institution
Home Economics	Grace Annagmeng Mwini	Tumu College of Education
	Imoro Miftaw	Gambaga Girls' SHS
	Jusinta Kwakyewaa (Rev. Sr.)	St. Francis SHTS
Religious Studies	Dr. Richardson Addai-Mununkum	University of Education Winneba
	Dr. Francis Opoku	Valley View University College
	Aransa Bawa Abdul Razak	Uthmaniya SHS
	Godfred Bonsu	Prempeh College
RME	Anthony Mensah	Abetifi College of Education
	Joseph Bless Darkwa	Volo Community SHS
	Clement Nsorwineh Atigah	Tamale SHS
Arabic	Dr. Murtada Mahmoud Muaz	AAMUSTED
	Dr. Abas Umar Mohammed	University of Ghana
	Mahey Ibrahim Mohammed	Tijjaniya Senior High School
French	Osmanu Ibrahim	Mount Mary College of Education
	Mawufemor Kwame Agorgli	Akim Asafo SHS
Performing Arts	Dr. Latipher Osei Appiah-Agyei	University of Education Winneba
	Desmond Ali Gasanga	Ghana Education Service
	Chris Ampomah Mensah	Bolgatanga SHS, Winkogo

Subject	Writer	Designation/Institution
Art and Design Studio and Foundation	Dr. Ebenezer Acquah	University for Education Winneba
	Seyram Kojo Adipah	Ghana Education Service
	Dr. Jectey Nyarko Mantey	Kwame Nkrumah University of Science and Technology
	Yaw Boateng Ampadu	Prempeh College
	Kwame Opoku Bonsu	Kwame Nkrumah University of Science and Technology
	Dzorka Etonam Justice	Kpando Senior High Sschool
Applied Technology	Dr. Sherry Kwabla Amedorme	AAMUSTED
	Dr. Prosper Mensah	AAMUSTED
	Esther Pokuah	Mampong Technical College of Education
	Wisdom Dzidzienyo Adzraku	AAMUSTED
	Kunkyuuri Philip	Kumasi SHTS
	Antwi Samuel	Kibi Senior High School
	Josiah Bawagigah Kandwe	Walewale Technical Institute
	Emmanuel Korletey	Benso Senior High Technical School
	Isaac Buckman	Armed Forces Senior High Technical School
	Tetteh Moses	Dagbon State Senior High School
	Awane Adongo Martin	Dabokpa Technical Institute
Design and Communication Technology	Gabriel Boafo	Kwabeng Anglican SHTS
	Henry Agmor Mensah	KASS
	Joseph Asomani	AAMUSTED
	Kwame Opoku Bonsu	Kwame Nkrumah University of Science and Technology
	Dr. Jectey Nyarko Mantey	Kwame Nkrumah University of Science and Technology
	Dr. Ebenezer Acquah	University for Education Winneba
Business Studies	Emmanuel Kodwo Arthur	ICAG
	Dr. Emmanuel Caesar Ayamba	Bolgatanga Technical University
	Ansbert Baba Avole	Bolgatanga Senior High School, Winkogo
	Faustina Graham	Ghana Education Service, HQ
	Nimako Victoria	SDA Senior High School, Akyem Sekyere
Agriculture	Dr. Esther Fobi Donkoh	University of Energy and Natural Resources
	Prof. Frederick Adzitey	University for Development Studies
	Eric Morgan Asante	St. Peter's Senior High School

Subject	**Writer**	**Designation/Institution**
Agricultural Science	David Esela Zigah	Achimota School
	Prof. J.V.K. Afun	Kwame Nkrumah University of Science and Technology
	Mrs. Benedicta Carbiliba Foli	Retired, Koforidua Senior High Technical School
Government	Josephine Akosua Gbagbo	Ngleshie Amanfro SHS
	Augustine Arko Blay	University of Education Winneba
	Samuel Kofi Adu	Fettehman Senior High School
Economics	Dr. Peter Anti Partey	University of Cape Coast
	Charlotte Kpogli	Ho Technical University
	Benjamin Agyekum	Mangoase Senior High School
Geography	Raymond Nsiah Asare	Methodist Girls' High School
	Prof. Ebenezer Owusu Sekyere	University for Development Studies
	Samuel Sakyi Addo	Achimota School
History	Kofi Adjei Akrasi	Opoku Ware School
	Dr. Anitha Oforiwah Adu-Boahen	University of Education Winneba
	Prince Essiaw	Enchi College of Education
Ghanaian Language	David Sarpei Nunoo	University of Education Winneba, Ajumako
	Catherine Ekua Mensah	University of Cape Coast
	Ebenezer Agyemang	Opoku Ware School
Physical Education and Health	Paul Dadzie	Accra Academy
	Sekor Gaveh	Kwabeng Anglican Senior High Technical School
	Anthonia Afosah Kwaaso	Junkwa Senior High School
	Mary Aku Ogum	University of Cape Coast
Social Studies	Mohammed Adam	University of Education Winneba
	Simon Tengan	Wa Senior High Technical School
	Jemima Ayensu	Holy Child School
Computing and Information Communication Technology (ICT)	Victor King Anyanful	OLA College of Education
	Raphael Dordoe Senyo	Ziavi Senior High Technical School
	Kwasi Abankwa Anokye	Ghana Education Service, SEU
	Millicent Heduvor	STEM Senior High School, Awaso
	Dr. Ephriam Kwaa Aidoo	University for Education Winneba
	Dr. Gaddafi Abdul-Salaam	Kwame Nkrumah University of Science and Technology

Subject	Writer	Designation/Institution
English Language	Esther O. Armah	Mangoase Senior High School
	Kukua Andoh Robertson	Achimota School
	Alfred Quaittoo	Kaneshie Senior High Technical School
	Benjamin Orrison Akrono	Islamic Girls' Senior High School
	Fuseini Hamza	Tamale Girls' Senior High School
Intervention English	Roberta Emma Amos-Abanyie	Ingit Education Consult
	Perfect Quarshie	Mawuko Girls Senior High School
	Sampson Dedey Baidoo	Benso Senior High Technical School
Literature-in-English	Blessington Dzah	Ziavi Senior High Technical School
	Angela Aninakwah	West African Senior High School
	Juliana Akomea	Mangoase Senior High School
General Science	Dr. Comfort Korkor Sam	University for Development Studies
	Saddik Mohammed	Ghana Education Service
	Robert Arhin	SDA SHS, Akyem Sekyere
Chemistry	Ambrose Ayikue	St. Francis College of Education
	Awumbire Patrick Nsobila	Bolgatanga SHS, Winkogo
	Bismark Tunu	Opoku Ware School
	Gbeddy Nereus Anthony	Ghanata Senior High School
Physics	Dr. Linus Labik	Kwame Nkrumah University of Science and Technology
	Henry Benyah	Wesley Girls High School
	Sylvester Affram	Kwabeng Anglican SHS
Biology	Paul Beeton Damoah	Prempeh College
	Maxwell Bunu	Ada College of Education
	Ebenezer Delali Kpelly	Wesley Girls' SHS
	Doris Osei-Antwi	Ghana National College
Mathematics	Edward Dadson Mills	University of Education Winneba
	Zacharia Abubakari Sadiq	Tamale College of Education
	Collins Kofi Annan	Mando SHS
Additional Mathematics	Dr. Nana Akosua Owusu-Ansah	University of Education Winneba
	Gershon Mantey	University of Education Winneba
	Innocent Duncan	KNUST SHS
Intervention Mathematics	Florence Yeboah	Assin Manso SHS
	Mawufemor Adukpo	Ghanata SHS
	Jemima Saah	Winneba SHS

Subject	**Writer**	**Designation/Institution**
Robotics	Dr. Eliel Keelson	Kwame Nkrumah University of Science and Technology
	Dr. Nii Longdon Sowah	University of Ghana
	Isaac Nzoley	Wesley Girls High School
Engineering	Daniel K. Agbogbo	Kwabeng Anglican SHTS
	Prof. Abdul-Rahman Ahmed	Kwame Nkrumah University of Science and Technology
	Valentina Osei-Himah	Atebubu College of Education
Aviation and Aerospace Engineering	Opoku Joel Mintah	Altair Unmanned Technologies
	Sam Ferdinand	Afua Kobi Ampem Girls' SHS
Biomedical Science	Dr. Dorothy Yakoba Agyapong	Kwame Nkrumah University of Science and Technology
	Jennifer Fafa Adzraku	Université Libre de Bruxelles
	Dr. Eric Worlawoe Gaba	Br. Tarcisius Prosthetics and Orthotics Training College
Manufacturing Engineering	Benjamin Atribawuni Asaaga	Kwame Nkrumah University of Science and Technology
	Dr. Samuel Boahene	Kwame Nkrumah University of Science and Technology
	Prof Charles Oppon	Cape Coast Technical University
Spanish	Setor Donne Novieto	University of Ghana
	Franklina Kabio Danlebo	University of Ghana
	Mishael Annoh Acheampong	University of Media, Art and Communication
Assessment	Benjamin Sundeme	St. Ambrose College of Education
	Dr. Isaac Amoako	Atebubu College of Education
Curriculum Writing Guide Technical Team	Paul Michael Cudjoe	Prempeh College
	Evans Odei	Achimota School