MINISTRY OF EDUCATION

Chemistry

TEACHER MANUAL

YEAR 1 - BOOK 1

NATIONAL COUNCIL FOR CURRICULUM & ASSESSMENT OF MINISTRY OF EDUCATION

MINISTRY OF EDUCATION

REPUBLIC OF GHANA

Chemistry

Teacher Manual

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NATIONAL COUNCIL FOR
CURRICULUM & ASSESSMENT
OF MINISTRY OF EDUCATION

CHEMISTRY TEACHER MANUAL

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INTRODUCTION

The National Council for Curriculum and Assessment (NaCCA) has developed a new Senior High School (SHS), 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 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 this information for the first 12 weeks of Year One, with the remaining 12 weeks contained within Book Two. 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.

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SCOPE AND SEQUENCE

Chemistry Summary

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 1: INTRODUCTION TO CHEMISTRY, SCIENTIFIC METHOD AND ATOMS

Strand: Physical Chemistry

Sub-Strand: Matter and its properties

Content Standard: Demonstrate understanding of the scientific practices in chemistry using relevant acquired skills to solve problems as well as explaining the structure of the atom and its stability

Learning Outcome: *Use the knowledge and understanding of the scientific practices in chemistry to explain the structure of the atom as well as the stability of its nucleus*

INTRODUCTION AND SECTION SUMMARY

This section introduces learners to the world of Chemistry, its branches, and its relevance to various aspects of life. Learners will engage with the scientific method of inquiry and explore significant milestones in atomic theory, including Bohr's model, Dalton's atomic theory, J.J. Thomson's cathode ray tube experiment, and the Rutherford model of the atom. Additionally, they will examine electron configuration and explore the study of radioactivity and the properties of radioactive radiations.

PEDAGOGICAL EXEMPLARS SUMMARY

To maximize learning, the section employs diverse pedagogical strategies such as digital learning, talk-for-learning, exploratory learning, inquiry-based learning, group learning, and project-based learning. These approaches are designed to promote the acquisition of skills like collaboration, communication, and digital literacy, preparing learners for the challenges of the 21st century.

ASSESSMENT SUMMARY

Different types of assessments will be used to measure how well learners understand and apply what they have learned. These assessments will focus on areas like chemical storage, lab safety, atomic theories, electron arrangements, and radioactivity. This approach ensures that learner's knowledge and abilities are evaluated, covering a wide range of concepts taught in class. By tailoring the assessments to these key areas, learners' understanding and application of chemistry concepts can be accurately evaluated.

Week 1

Learning Indicator(s):

1. *Describe chemical processes around us and their applications in everyday life.*
2. *Discuss and explain safety rules and hazard symbols in the laboratory.*

Theme or Focal Area: Meaning of Chemistry

Chemistry is a scientific discipline that focuses on the study of matter, its composition, structure, and properties as well as the principles governing its behaviour. It intersects with fields like physics, biology, environmental science, and engineering and is crucial in understanding and explaining the natural world. Chemistry plays a vital role in developing new technologies, materials, and drugs for various applications.

Branches of Chemistry

1. Pure chemistry is the study of basic principles and theories of chemistry without considering practical use or application. It involves exploring the properties, structure, and behaviour of matter at a molecular and atomic level, analysing the interactions and transformation of substances, understanding the behaviour of atoms and molecules, discovering new compounds, and improving technologies.

Branches of Pure Chemistry

The main branches of pure chemistry are as follows.

a. Physical Chemistry

b. Organic Chemistry

c. Inorganic Chemistry

   a. **Physical Chemistry** is a branch of chemistry that combines principles from physics and chemistry to study the relationship between the physical properties of matter and its chemical composition and behaviour.
   b. **Organic Chemistry** is the branch of chemistry that studies carbon-based molecules and their properties, composition, and reactions.
   c. **Inorganic chemistry** is the branch of chemistry that studies non-carbon-based compounds and their properties, composition, and reactions. It includes the study of the properties of elements, their compounds and their behaviours in different conditions.

2. Applied chemistry is the branch of chemistry that studies the practical applications of chemical knowledge in various fields. It focuses on using chemistry to solve real-world problems using scientific methods and principles. It has diverse applications including food science, medicine, pharmaceuticals, material sciences, agriculture and environmental science.

The centrality of chemistry as a science discipline

Chemistry is a central scientific discipline that plays a critical role in various aspects of our daily lives, from health and well-being to the environment around us. It is often called the "central – science" because it is connected to other disciplines such as physics and biology. It is critical in the development of new materials in various industries such as electronics, textiles, and construction.

Chemistry is a crucial discipline that provides a fundamental understanding of our world. Its applications are vast and include technology, medicine, industry, and environmental management, making it central to scientific progress and human development.

What is the relationship between chemistry and other subjects?

Chemistry has close relationships with various other subjects, including physics, biology, and environmental science, due to the fact that it overlaps with them in terms of content and techniques.

Effect of chemistry on daily lives

Chemistry has an enormous impact on daily life, as it is essential for various aspects of modern life. Here are some ways in which chemistry affects daily life.

1. **Food and nutrition:** Chemistry has a significant impact on food and nutrition by improving food quality, safety, and preservation (e.g. Treatment of water at Kpong and Weija.; Standardisation of products at Ghana Standards Authority). It helps us understand the composition of different foods, develops various food processing techniques, and uses chemicals as food additives to improve taste and prevent spoilage. It also provides tools and techniques for analysing food components, contaminants, and nutrients, contributing to research aimed at improving health and disease prevention.
2. **Agriculture:** Chemistry is crucial in agriculture to maximise crop yield and quality while minimising costs and environmental impact. It impacts agriculture through the development of animal feed (KoudijsGh Ltd.), fertilisers (Glofert- fertiliser company in Tema, Ghana), chemical pesticides to control pests, understanding soil chemistry, genetic modifications (Ghana Atomic Energy Commission), and water management with chemicals. Chemistry has revolutionised agriculture, providing valuable insights, technologies, and solutions to enhance crop yields, control pests and diseases and improve soil and water quality.
3. **Medicine:** Chemistry has a significant impact on medicine as it contributes to the development of drugs (Tobinco Pharmaceutical Ltd.) and medical devices (Intravenous infusions PLC, Koforidua, Ghana), their production, and analysis. Chemistry plays a role in discovering new compounds and synthesising them to optimise their therapeutic use.
4. **Transportation:** Chemistry greatly affects transportation in various ways, which include fuel production (Tema Oil Refinery, TOR, Ghana), vehicle material designs, lubricants and additives, emissions control, and battery technologies. These chemical advancements enhance fuel efficiency, decrease emissions, and improve the transition to eco-friendly transportation methods.
5. **Energy:** Chemistry affects energy through its involvement in the production of traditional and renewable energy, energy storage solutions, the development of energy-efficient technologies, and technologies that reduce emissions from energy production. Through chemical principles, researchers can identify solutions that promote more sustainable and environmentally friendly energy production and consumption. Chemistry plays a crucial role in the production of traditional energy sources such as coal, oil (TOR), and natural gas through processes such as extraction, refining and combustion. Chemistry is also involved in the production of renewable energy sources such as solar panels (Global Engineering and Drilling Ghana Ltd., East Lagon) and wind turbines through the development of new materials and processes. Battery technology relies on electrochemistry.

Careers in chemistry and chemistry-related fields

There are many career opportunities in the field of chemistry and chemistry-related fields. Below are just a few examples

1. Pharmacist
2. Medical doctor
3. Biochemist
4. Chemical engineer
5. Chemistry teacher

6. Nurse
7. Laboratory technician

Education and training required for careers in chemistry

The education and training required for careers in chemistry and related fields vary depending on the specific job and employer.

Some jobs in chemistry or chemistry-related fields require a minimum of a bachelor's degree, while some specialised positions may require an advanced degree. Employers often require laboratory or research experience, relevant work experience and problem-solving skills. Some routine laboratory jobs may not require a degree but would need school qualifications in chemistry.

Learning Tasks

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

Administer the following learning tasks to help learners to reinforce and acquire new knowledge and skills.

Learners:

1. a. watch video or slides/ pictures on a variety of natural and artificial phenomena that can be explained by chemistry.

   b. deduce and discuss the meaning of chemistry from their observations.

   c. discuss and distinguish between the traditional branches of chemistry.

   d. discuss the centrality of chemistry as a science discipline which is related to other science subjects.

   e. draw flowcharts and pictures to summarise ways in which chemistry affects daily lives under the following headings: (Food and nutrition, Agriculture, Medicine, Transportation, Energy, etc.)

   f. discuss careers in chemistry and chemistry-related fields.

   g. discuss the education and training required for the careers in chemistry chemistry-related fields.

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

Digital learning:

a. Divide learners into mixed ability groups and show them video or slide/pictures on variety of natural and artificial phenomena that can be explained by chemistry. Visually impaired learners should be considered and alternative plans made for them.

b. Guide learners to deduce and discuss the meaning of chemistry from their observations. Encourage slow learners to contribute to the discussion by directing questions to them.

c. Guide learners to discuss and distinguish among the traditional branches of chemistry. Ask questions and encourage all learners to talk in the group.

d. In mixed ability groups, guide learners to discuss the centrality of chemistry as a science discipline and its relation to other science subjects.

e. Using mixed ability groups, guide learners to draw flowcharts and pictures to summarise ways in which chemistry affects daily lives under the following headings: (Food and nutrition, Agriculture, Medicine, Transportation, Energy, etc.)

   Visually impaired learners should be considered and alternative plans made for them.

f. In mixed-ability groups, guide learners to discuss careers in chemistry and chemistry-related fields. Encourage shy learners to contribute to the discussion by directing questions to them.

g. Guide leaners to discuss the education and training required for careers in chemistry and chemistry-related fields. Encourage slow learners to contribute to the discussion by directing questions to them.

Key Assessment

DoK level 1: Recall/Reproduction

State three careers that are related to chemistry.

DoK level 2: Skill/Concept

State and explain at least three benefits of chemistry to the Ghanaian society.

DoK level 3: Strategic Reasoning/Thinking

Undertake research on the importance of chemistry to the Ghanaian society in the field of agriculture and present your findings using PowerPoint or flipcharts.

Theme or Focal Area: Rules and regulations in the chemistry laboratory

The school chemistry laboratory must be a safe place for effective learning. Given this, the following rules and regulations must be explained and observed.

1. Do not eat or drink anything in the laboratory.
2. Never taste chemicals in the laboratory.
3. Any water spilled on the floor must be wiped off immediately.
4. Do not add water to acid but rather acid to water.
5. Never walk barefoot in the laboratory.
6. Keep the laboratory clean and organised.
7. Report all accidents and spills immediately.
8. Follow proper handling and disposal procedures for chemicals.
9. Follow instructions for conducting experiments and using equipment.
10. Wear appropriate protective equipment such as a lab coat, safety goggles and gloves.
11. Know the location and use of emergency equipment such as fire extinguishers, eye wash stations, and safety showers.
12. Be aware of the potential hazards in the laboratory and take precautions to prevent accidents.

Chemical Hazards

Chemical hazards are potential hazards and dangers can be posed by solids, liquids, gases, and solutions.

These hazards can have various effects on human health and the environment depending on the chemical, the dose, the exposure route, and the duration of exposure.

Some common chemical hazards include:

1. **Explosives**

   These are chemicals that can rapidly release energy in the form of heat, light, gas and sound, causing physical damage and injury.

   Examples are dynamite, nitro-glycerine, ammonium nitrate and nitrocellulose.

2. **Flammable liquids and gases**

   These are chemicals that can ignite (catch fire) or explode when exposed to heat, sparks, or flames, causing burns, fires, and explosions.

   Examples are gasoline, propane, butane, ethanol, diesel, acetone, paint thinners, aerosol sprays, lubricating oils, cooking oils, and fats.

3. **Corrosive substances**

   These are chemicals that can destroy or damage materials such as metals, plastics, or human tissue, causing severe burns and tissue damage.

   Examples are hydrochloric acid, sulfuric acid, nitric acid, sodium hydroxide, potassium hydroxide, bleach and ammonia solution.

4. **Toxic substances**

   These are chemicals that can harm or kill living organisms such as humans, animals, or plants by interfering with biological functions or disrupting vital organ systems.

   Examples are arsenic, lead, mercury, carbon monoxide, pesticides, cyanide, benzene, chlorine gas and ammonia.

5. **Oxidising substances**

   These are chemicals that can accelerate and promote combustion (burning) in other materials by providing oxygen or other oxidising agents. They cause severe burns, respiratory damage, and explosions.

   Examples are hydrogen peroxide, potassium permanganate, oxygen gas, chlorine gas, bleach, nitric acid, and potassium nitrate.

6. **Radioactive substances**

   These materials emit radiations spontaneously as a result of decay of their atomic nuclei. It is essential to handle and dispose of them properly as they can pose significant hazards due to their potential for radiation exposure and contamination.

   Examples of radioactive substances are uranium, radon, iodine-131, and cobalt-60.

7. **Irritant substances**

   These are materials that can cause irritation or inflammation when they come into contact with the skin, eyes, respiratory system, or other organs.

   Irritant substances can have a range of adverse effects on humans, such as itching, pain, redness, swelling, and blistering of the skin.

   Examples are ammonia, bleach, hydrochloric acid, detergents, insecticides, sodium hydroxide and gasoline.

8. **Harmful substances**

   These are materials that can pose a risk to the health and safety of humans or the environment. They can cause acute or chronic effects on exposure, depending on the dose, duration and mode of exposure.

   Examples are lead, asbestos, pesticides, carbon monoxide, tobacco smoke, mercury, and arsenic.

9. **Biohazard substances**

   These are materials that can pose a threat to the health and safety of living organisms, including humans, plants, and animals. These substances may contain living and non-living biological agents that can cause harm such as bacteria, viruses, toxins and biological waste.

   Examples of biohazard substances include blood, bodily fluids, tissues, organs, and microorganisms.

Prohibition Signs

Prohibition signs are warning signs that indicate that certain activities, actions, or objects are not allowed (prohibited) in a particular area.

Explanation	Symbol
1. **No naked flame:** It is a prohibition sign that indicates that open flames and any activity involving unprotected flames are restricted or prohibited in a certain area. The sign is intended to improve safety, guard against fire hazards, and ensure adherence to applicable legislation. Common locations include industrial settings, laboratories, fuel storage areas, construction sites, and places with flammable materials. A 'no naked flame' sign as shown.	 **Figure 1.0:** *No Naked Flame Symbol*
2. **Danger:** 'Danger' is a term used to describe a specific situation, activity, or condition that poses a significant risk of harm, injury or damage to individuals, property or the environment. Any situation that has the potential to cause harm, injury or damage can be considered as dangerous, and it is important to respond to such situations promptly to prevent harm. A 'danger' sign as shown.	 **Figure 1.1:** *Danger Symbol.*
3. **No smoking:** 'No smoking' is a common sign that is seen in public places, workplaces, and other areas where smoking is prohibited. A 'no smoking' sign as shown.	 **Figure 1.2:** *No smoking Symbol*

Hazard symbols and their meanings

Hazard symbols

Hazard symbols are visual signs or markings used to indicate the potential danger or risks associated with a particular substance or product. These pictograms are usually displayed on containers or packing to help users identify and handle hazardous materials safely.

Explanation	Symbols
1. **Harmful symbol:** The harmful symbol is used to indicate that a substance is harmful if ingested, inhaled, or absorbed through the skin. The harmful symbol as shown.	 **Figure 1.3:** *Harmful Symbol.*
2. **Irritant symbol:** The irritant symbol is used to indicate that a substance may irritate the skin, eyes or respiratory system. The irritant symbol as shown.	 **Figure 1.4:** *Irritant Symbol*
3. **Corrosive symbol:** The corrosive symbol is used to indicate that a substance is capable of causing irreversible damage to living tissues or corroding materials including metals, plastics, and other substances. The corrosive symbol as shown.	 **Figure 1.5:** *Corrosive Symbol.*
4. **Toxic symbol:** The toxic symbol is used to indicate that a substance is highly poisonous and can cause serious harm to human health or the environment. The toxic symbol shown.	 **Figure 1.6:** *Toxic Symbol*
5. **Oxidising symbol:** The oxidising symbol is used to indicate that a substance is capable of promoting the combustion (burning) or ignition of other materials. The oxidising symbol as shown.	 **Figure 1.7:** *Oxidising Symbol*

Explanation	Symbols
6. **Flammable symbol:** The flammable symbol is used to indicate that a substance is combustible and can catch fire easily. The symbol for flammable as shown	 **Figure 1.8:** *Flammable Symbol.*
7. **Explosive symbol :** The explosive symbol is a sign that warns people about the presence of explosives or other hazardous substances. The explosive symbol as shown.	 **Figure 1.8:** *Explosive Symbol.*
8. **Radioactive symbol:** The radioactive symbol, also known as radiation hazard symbol, is a warning symbol that is used to indicate the presence of radioactive materials or areas that emit radiation. The radiation hazard symbol is shown below:	 **Figure 1.9:** *Radioactive Symbol*
9. **Biohazard symbol:** The biohazard symbol is used to indicate the presence of biological hazards such as infectious agents, toxins, and other biohazardous materials that can cause harm to human health or the environment. The biohazard symbol is shown below:	 **Figure 1.10:** *biohazard Symbol*

First aid Signs

First aid signs serve as visual markers intended to designate the whereabouts of first aid facilities, equipment, or stations within a given area. Their essential function lies in enhancing workplace safety and furnishing explicit directions to individuals during emergency situations. These signs commonly employ universally acknowledged symbols, colors, and text to communicate details regarding the accessibility and position of first aid resources.

1. **First aid:** First aid is the immediate assistance provided to a person who has been injured or has suddenly taken ill. It involves a series of simple, life-saving techniques and procedures that can be performed by anyone with basic training. The primary objective of first aid is to preserve life, prevent the condition from worsening, and promote recovery while waiting for professional medical attention. A first aid sign as shown.	 **Figure 1.11:** *First aid Symbol*
2. **Safety Shower:** This sign is a visual cue that shows where first aid supplies and safety showers are located in a building. This sign is usually located in areas where there is a possibility of exposure to hazardous compounds that may require quick decontamination or first aid, is essential to emergency response. This sign is placed strategically in places where there is a greater chance of coming into contact with potentially harmful products, such as industrial facilities, laboratories, chemical storage areas, and other workspaces with potentially harmful materials. A safety shower sign as shown.	 **Figure 1.12:** *1. Safety Shower Symbol*
3. **Eye Wash:** This sign is a visual indicator designed to identify the location of emergency eye wash stations. These stations are crucial in environments where there is a risk of exposure to hazardous substances that can cause eye injuries or irritation. The sign helps individuals quickly locate the nearest eye wash station, promoting prompt action in case of an emergency. It is strategically placed in areas where there is a risk of eye exposure to chemicals, dust, or other hazardous materials. Common locations include laboratories, manufacturing facilities, chemical storage areas, and places where workers handle potentially harmful substances. An eye wash sign as shown.	 **Figure 1.13:** *Eye Wash Symbol.*

Personal Protective Equipment (PPE)

Personal protective equipment (PPE) is any equipment or clothing worn by individuals to protect against the specific hazards in the workplace or other environments. PPE is designed to protect the wearer from potential hazards that could cause injury, illness, or death. Some of the common types of PPE include:

1. **Respirators/Gas masks:** These are devices designed to reduce the inhalation of hazardous substances such as dust, fumes, and gases.
2. **Hand gloves:** These are specialised hard-worn protective coverings that protect the skin against harmful substances or injuries.

3. **Eye protectors:** These devices include safety glasses, chemical goggles, or face shields that protect the eyes from flying particles, dust, or splashes of hazardous substances.
4. **Protective clothing :** This includes specialised clothing such as lab coats, aprons, and full-body suits that protect against chemicals, heat, and other hazardous materials.

Safety Equipment

Safety equipment refers to devices or clothing that are designed to protect individuals from injuries or hazards while performing activities or tasks. This includes:

1. **Eye shower station:** An eye shower station is a safety device found in workplaces where hazardous exposure is possible. It provides immediate treatment to individuals who have contact with hazardous materials or chemicals in the eye. The device consists of a basin attached to a water supply, and an injured person is instructed to flush the eyes with water directed at the eyes and not the face.
2. **Fume chamber:** A fume chamber is an enclosed space or hood designed to contain and capture hazardous fumes, dust, or vapours that may be produced during laboratory experiments or industrial processes. It is often used to protect workers from dangerous substances that may be emitted during experiments, testing, or production and to prevent contamination of the external environment.

Learning Tasks

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

Learners

1. a. (i) watch videos or see pictures/drawings of wrong practices in the chemistry laboratory.
   (ii) discuss the ‘dos’ and ‘don’ts’ in the chemistry laboratory.
   (iii) discuss the rules and regulations that should be followed in the chemistry laboratory.
   b. (i) examine the assay of chemical containers or reagents, electrical gadgets, and other materials and identify the hazard symbols on them.
   (ii) discuss chemical hazards under the following headings: (corrosive, toxic, oxidising, flammable, explosive, radioactive, irritant/harmful, biohazard)
   (iii) discuss prohibition signs under the following headings: (First aid, danger, No smoking, High voltage, etc.)
   c. sketch the hazard symbols and explain what they mean.
   d. exchange ideas on how to handle hazardous chemicals using personal protective equipment and safety equipment.
2. discuss laboratory rules and hazard symbols in the chemistry laboratory.

Pedagogical Exemplars (with the cross-cutting themes integrated)

Talk-for-Learning:

a. i. Divide learners into mixed groups and show them videos or pictures of practices in the chemistry laboratory. Virtually impaired learners should be considered, and alternative plans made for them.

ii. Guide learners to discuss the 'dos' and 'don'ts' in the chemistry laboratory. Encourage shy learners to contribute to the discussion by directing questions to them.

iii. Guide learners to discuss the rules and regulations that should be followed in the chemistry laboratory. Encourage slow learners to contribute to the discussion by directing questions to them.

b. i. In mixed ability groups guide learners to examine the assay of chemical containers or reagents, electrical gadgets, and other materials and identify the hazard symbols on them. Virtually impaired learners should be considered, and alternative plans made for them.

ii In mixed ability groups, guide learners to discuss chemical hazards under the following headings: corrosive, toxic, oxidising, flammable, explosive, radioactive, irritant/harmful, biohazard. Encourage slow learners to contribute to the discussion by directing questions to them.

iii. In mixed ability groups, guide learners to discuss Prohibition signs under the following headings: First aid, danger, No smoking, High voltage, etc. Encourage slow learners to contribute to the discussion by directing questions to them.

c. In ability groups, guide learners to sketch at least two hazard symbols and explain what they mean. Virtually impaired learners should be considered, and alternative plans made for them.

d. Using think-pair-share, guide learners to exchange ideas on how to handle hazardous chemicals safely using personal protective equipment such as chemical goggles, hand gloves, apron/laboratory coat, respirator/gas mask and safety equipment such as eye shower station and fume hood. Encourage slow learners to talk by directing questions to them.

Collaboration and Communication:

In ability groups, guide learners to discuss at least five laboratory rules and at least five hazard symbols. Ask each group to select a leader who will do presentation on behalf of the group. Encourage slow learners to talk by directing questions to them.

Key Assessment

DoK level 1: Recall/Reproduction

State at least two examples of personal protective equipment used in the chemistry laboratory**.**

DoK level 1:Recall/Reproduction

State the precautions to take when handling chemicals with the following hazard labels.

a. Corrosive substance

b. Toxic substance

DoK level 3:Strategic Reasoning/Thinking

In a certain school, the authorities are concerned about the increasing rate of accidents in the chemistry laboratory during practical sessions.

a. Suggest three possible causes of these accidents

b. Write and explain three (3) ways of curbing the problem identified

c. Assuming one of the identified causes of the accident is improper handling and usage of reagents, provide two remedies to addressing this issue.

Reference

1. SHS Chemistry Curriculum : 1.x. 1. CS.1,1.x.1.1.LI.1-2.x.1.1 AS.1-3 (p. 14-15)
2. Zumdal, S,S.,& Decoste(2017). Chemistry. Cengage learning
3. Chang, R(2017). Chemistry, McGraw-Hill Education.
4. Atkins, P.W.,& de Paula J(2017). Atkins, Physical Chemistry, Oxford University Press
5. Khan Academy-Introduction to chemistry: https://www.khanacademy.org/science/chemistry/atomic-structure-and-properties/introduction-to-the-atom/v/introduction-to-chemistry
6. American chemical society.(n.d). Careers in chemistry. https://www.acs.org/careers/chemical-sciences.html
7. OSHA. Personal Protective Equipment. Retrieved from https://www.osha.gov/personal-protective-equipment.
8. Laboratory Safety Institute-https:// www.labsafetyinstitute.org
9. Occupational Safety and Health Administration (OSHA)-Hazard Communication Standard. https://www.osha.gov/hazcom

WEEK 2

Learning Indicator(s):

1. *Explain why chemicals should be stored by compatibility and not alphabetically in the laboratory.*
2. *Investigate the scientific method of inquiry.*

Theme or Focal Area: Storage of Chemicals

Chemicals in the laboratory should be stored in a safe and organised manner to prevent accidents and ensure the safety of the laboratory workers. It is better to store these chemicals by compatibility and not alphabetically.

Here are some reasons why chemicals should be stored by compatibility and not alphabetically:

1. Chemicals should be stored by compatibility in the laboratory because some chemicals can react explosively or dangerously when they come into contact with other chemicals.
2. Storing chemicals alphabetically can result in incompatible substances being stored next to each other. This can cause chemical reactions that can result in fires, explosions, toxic fumes, and other hazardous situations.
3. Storing chemicals by compatibility reduces the risk of accidents and ensures the safety of laboratory workers.
4. Storing chemicals by compatibility also helps in organising the storage of chemicals, making it easier to locate specific chemicals when needed.

Therefore, it is important to follow the guidelines for the storage of chemicals by compatibility to prevent harmful incidents or accidents in the laboratory.

The following guidelines can be used to store chemicals in the laboratory.

Table 2.0: *Chemical Compatibility Chart.*

Chemical Incompatibility Chart

	Acids Inorganic	Acids Oxidizing	Acids Organic	Alkalis Bases	Oxidizers	Poisons Inorganic	Poisons Organic	Water Reactives	Organic Solvents
Acids Inorganic			X	X		X	X	X	X
Acids Oxidizing			X	X		X	X	X	X
Acids Organic	X	X		X	X	X	X	X	
Alkalis Bases	X	X	X				X	X	X
Oxidizers			X				X	X	X
Poisons Inorganic	X	X	X				X	X	X
Poisons Organic	X	X	X	X	X	X			
Water Reactives	X	X	X	X	X	X			
Organic Solvents	X	X		X	X	X			

How to put out a Small Fire Using Fire Blanket and Fire Extinguisher.

If a small fire breaks out in the laboratory, it is important to act quickly and appropriately to minimise any potential damage or injuries.

Here is how to use a fire blanket and fire extinguisher to put out a small fire in laboratory setting:

1. **Using Fire Blanket**
   (a) If there is a small fire on or near a person, the person should stop, drop, and roll to smother the flames. If the fire is caused by a flammable liquid or in a pan, turn off the heat source.
   (b) Pull the blanket out of its bag or storage container.
   (c) Hold the corner of the blanket and if possible, cover the fire starting from the base. If the fire is on a person's clothing wrap the blanket around him or her to smother the flames.
   (d) Make sure the edges of the blanket create a seal with the surface around the fire to prevent the spread of flames.
   (e) Leave the blanket in place until the fire has completely stopped or until help arrives.

2. **Using Fire Extinguisher**
   (a) Before using the fire extinguisher, pull the fire alarm and make sure everyone in the area is aware of the fire.
   (b) Identify the type of fire extinguisher you are using to ensure it is appropriate for the fire you are dealing with.
   (c) Hold the fire extinguisher in an upright position and aim it at the base of the flames.
   (d) Squeeze the handle or trigger to release the extinguishing agent.
   (e) Sweep the nozzle from side to side while aiming at the base of the flames until the fire is completely extinguished.
   (f) Stay alert for any reignition of the flames and keep the fire extinguisher aimed at the base of the flames until it is safe to put away.

Understanding the type of fire extinguisher, you have and the type of fire you are dealing with is crucial to ensure the right type of action.

Learning Tasks

1. List at least two key rules to follow when storing chemicals in the laboratory.
2. State at least two reasons why chemicals should be stored by compatibility method and not in alphabetical order.
3. Explain how to extinguish small fires in the laboratory using:
   a. Fire blanket
   b. Fire extinguisher

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

Exploratory Learning / Collaborative Learning

a. Using field trip as a strategy, support learners to visit the school chemical store or chemistry laboratory to observe how chemicals are stored.

b. In mixed ability groups ask learners to explain why chemicals should be stored by compatibility and not alphabetically in the laboratory.

c. Show a video to learners on a violent reaction between chemicals as a result of them being stored right next to each other *(e.g., hydrogen peroxide and hydrazine, or, oxidising materials and* *flammable materials, acetic acid and nitric acid or reaction between potassium permanganate and glycerol).* Lead a whole class discussion related to the key concepts demonstrated in the video.

- **d.** Use think-pair-share approach to lead discuss on how to put out small fire using fire blanket and fire extinguisher.
- **e.** Guide learners in pairs to demonstrate how to put out small fire using fire blanket and fire extinguisher.

NB: Take all the necessary precautions to prevent accidents and burns. Make provision for first aid. Also, be mindful of learners who may be allergic to smoke.

Key Assessment

DoK Level 1 Reproduction/Recall

State at least one (1) type of fire that can be extinguished using a carbon dioxide fire extinguisher.

DoK Level 2 Skills of conceptual understanding

Explain why potassium permanganate should not be stored near a bottle containing ethanol.

DoK Level 3 Strategic reasoning

Design a chemical storage compatibility chart using the following chemicals in the laboratory: HCl, HNO_3, CH_3COOH, NaOH, H_2O_2, CH_3CH_2OH.

Theme or Focal Area(s): The scientific Method of Inquiry

The Scientific method of inquiry is a systematic approach used by scientists to investigate and learn about the natural world around us. The scientific method involves a series of steps that are followed to ensure that scientific investigations are made in a logical, objective, and repeatable manner.

Here are the steps involved in the scientific method:

1. **Make Observations:** Scientists begin by carefully observing and recording information about a phenomenon or problem they wish to investigate, or gather prior knowledge about a certain topic or concept
2. **Formulate a Question:** Based on their observations and prior knowledge, scientists create a question or issue they want to investigate.
3. **Develop a Hypothesis:** A hypothesis is an educated guess about the answer to the question or issue that has been formulated.
4. **Conduct Experiments:** In this step, scientists design and conduct experiments to test the hypothesis.
5. **Collect and Analyse Data:** Scientists record their observations and collect data from their experiments.
6. **Draw Conclusions:** Based on the results of their experiments and observations, scientists use logic to draw conclusions about their hypothesis.
7. **Communicate Results:** Finally, scientists communicate their findings through scientific papers, presentations or other means.

The scientific method allows scientists to avoid bias and to ensure that their results are valid, reliable and replicable. Through this methodical approach, scientists can discover new knowledge, solve problems and explore and understand the natural world.

Learning Task

1. List at least five steps involved in the scientific method of enquiry.
2. Discuss at least three steps involved in the scientific method of enquiry.
3. a. Describe how to apply the steps in the scientific methods of enquiry to solve a problem in the school environment or nearby community.
   b. Design a poster outlining the scientific methods of enquiry to be used in solving the problem and share with the class for discussion.

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

Inquiry-Based Learning:

a) In mixed - ability groupings learners discuss at least five steps involved in the scientific method of enquiry.

b) In ability groupings guide learners to apply the steps in the scientific methods of enquiry to solve a problem in the school environment or nearby community.

c) In ability groupings support learners to design a poster outlining the method used and share with the class for discussion.

Key Assessment

DoK Level 1 Reproduction/Recall

Outline at least five steps involved in the scientific method of enquiry.

DoK Level 2 Skills of Conceptual Understanding

Identify at least a problem in the school environment that can be solved using the scientific method of enquiry.

DoK Level 3 Strategic Reasoning

Formulate a hypothesis to drive the investigation for at least one of the problems identified.

DoK Level 4 Extended Thinking

Design an experiment that can be used to solve the problem(s) identified.

Reference

1. Occupational Safety and Health Administration (OSHA)-Hazard Communication. Retrieved on December 18, 2023, from https://www.osha.gov/hazcom.
2. National Academy of Science. (1998). Teaching about the Scientific method: *A guide for schools*. National Academy Press.
3. Harrison, A.G., &Treagust, D.F., (2016). Teaching and Learning the Nature of Science and Scientific Inquiry: *An International Review of Science Education*. Springer.

Week 3

Learning Indicator(s):

1. *Identify the main postulates of Dalton's atomic theory and explain the weaknesses of the theory.*
2. *Describe the cathode ray experiment and alpha particle scattering experiment and identify the weaknesses of J. J. Thompson and Rutherford's models of the atom.*

Theme or Focal Area: Dalton's Atomic Theory

The idea that elements are made up of atoms is called the **atomic theory.**

The postulates of the Dalton's atomic theory are as follows:

1. All elements are made up of small indivisible particles called atoms.
2. Atoms cannot be created or destroyed.
3. Atoms of the same element are identical, that is, they have the same mass and size, but atoms of different elements have different masses and sizes.
4. Atoms of different elements combine in simple whole number ratios to form compounds.

The theory in its broad outline is still valid, however, some of the postulates have been modified in the light of subsequent discoveries.

Modification of the Dalton's Atomic Theory

1. **All elements are made up of small indivisible particles called atoms**:
   Atoms are not indivisible. This is because it was later discovered that atoms are not the smallest particles and further broken down into subatomic particles such as protons, neutrons and electrons.
2. **Atoms cannot be created or destroyed:**
   This postulate is still acceptable for ordinary chemical reactions. In nuclear reactions however, atoms of the same element are destroyed and new ones are created.
3. **Atoms of the same element are identical, that is, they have the same mass and size, but atoms of different elements have different mass and size:**
   The discovery of isotopes (atoms of the same element having the same number of protons but different number of neutrons) contradicts this postulate.
4. **Atoms of different elements combine in simple whole number ratios to form compounds:**
   This postulate is still acceptable for inorganic compounds which usually contain few atoms per molecule. Carbon, however, forms very large organic compounds such as polymers, proteins, and starch which can contain thousands of atoms.

   Silicon, which is inorganic, also forms a very complex silicates involving a large number of atoms.

Learning Task:

1. State at least one postulate of Dalton's atomic theory.
2. State at least two postulates of Dalton's atomic theory that were modified.

3. State at least two postulates of Dalton's atomic theory that were modified and give the reasons for their modification.

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

1. Talk-for- learning: In a whole class discussion, guide learners to review the description of the atom and its sub–atomic particles from JHS syllabus.
2. Group work: Learners in mixed-ability groups,
   a. Discuss the postulates of Dalton's atomic theory. Moderate the discussion in such a way that every learner can state at least one of the main postulates of Daltons atomic theory.
   b. Evaluate and criticise each of the postulates of Dalton's atomic theory with the aid of relevant charts or other resources. Encourage learners in their groups to explain the basis upon which some postulates of Dalton's atomic theory were modified.
3. Project–based learning:
   a. Construct a model to represent the atom as a simple sphere with no internal structure.
   b. Draw a diagram of the atom modelled.
   c. Display the model and diagram for class discussion.

Key Assessment

DoK Level 1 (Reproduction/Recall)

In your own words, state the main postulates of Dalton's atomic theory.

DoK level 4 (Extended critical thinking and reasoning)

State at least one strength and one weakness of Dalton's atomic theory.

DoK Level 2 (Skills of conceptual understanding)

Explain at least one strength and one weakness of Dalton's atomic theory.

DoK Level 3 (Strategic reasoning)

How relevant is Dalton's atomic theory to the evolution of modern chemistry?

Theme or Focal Area(s): J.J. Thomson's Cathode ray experiment

J.J. Thomson's cathode ray experiment was a series of experiments that laid the foundation for the discovery of the electron.

Thomson passed an electric current through a vacuum tube and observed a stream of negatively charged particles that travelled from the negatively charged electrode, known as the cathode, to the positively charged electrode, known as the anode. These particles were called cathode rays. The cathode ray tube used by Thomson is shown below.

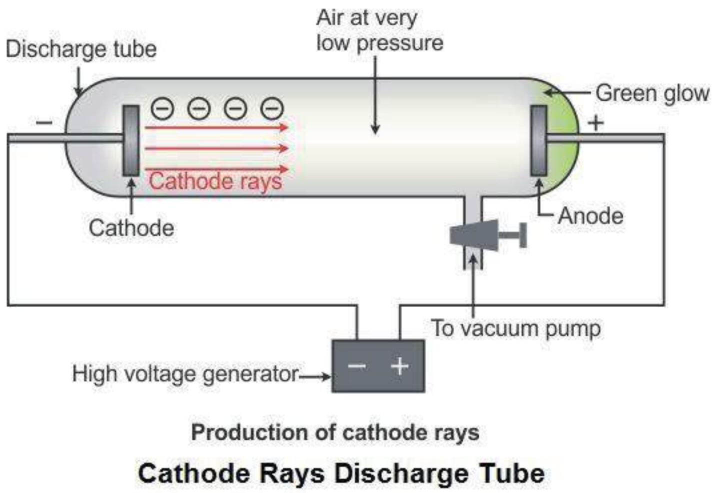


Figure 3.0: *A Cathode Ray Tube*

To further study the cathode rays, Thomson conducted experiments using electrical and magnetic fields. He found that the rays were deflected by both fields, indicating that the particles that made up the ray had a negative charge. He also measured the charge-to-mass ratio of the particles and found that it was much smaller than any known atom, leading him to conclude that the cathode rays were made of particles smaller than the atoms. These particles were named as electrons.

Thomson's discovery of the electron was a major breakthrough in the understanding of the structure of matter.

J.J. Thomson's model of the atom

J.J. Thomson's model of the atom, also known as the 'plum pudding' model was developed on the basis of his discovery of the electron. According to this model, the atom is composed of a positively charged sphere, like a pudding, in which negatively charged electrons are embedded, like plums.

The diagram below shows Thomson's model of the atom.

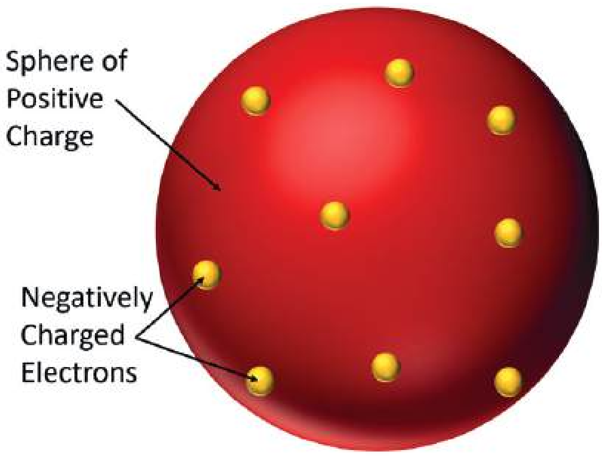


Figure 3.1: *Plum Pudding Model of the Atom.*

While this model was a significant step in the understanding of atomic structure, it had some weaknesses.

Some of the main weaknesses of J.J. Thomson's model of the atom are:

1. The model assumed that the positive and negative charges were spread evenly throughout the atom, which would produce a neutral electric charge. However, the model failed to explain the presence of a nucleus in atoms.
2. The model also did not provide any information about the number or arrangement of the electrons in an atom. Thomson's model just suggested that the electrons were dispersed throughout the atom but not in a pattern or orbit.
3. The model did not explain the overall mass of the atom. The electrons are much lighter than the protons and neutrons that made up the nucleus, and it was not clear how the overall mass of the atom was distributed.
4. The discovery of the atomic nucleus by Rutherford disproved Thomson's model by showing that it was not accurate in describing the atomic structure.

Despite the limitations, Thomson's model was significant in showing that atoms are not indivisible but could be further broken down into constituent particles. It also led to the development of further models for the structure of the atom, including Rutherford's model, which corrected some of the shortcomings of Thomson's model.

Rutherford' alpha scattering experiment

Rutherford's alpha scattering experiment was a landmark experiment aimed to investigate the structure of the atom and the nature of its constituent particles.

The experiment involved firing positively charged alpha particles at a thin gold foil and observing their trajectory as they passed through the foil. The expectation was that the alpha particles would pass straight through the foil or be slightly deflected by the atomic structure of the atoms within the foil. However, some alpha particles were scattered at very large angles, and some even scattered backwards.

The diagram below shows Rutherford's alpha scattering experiment.

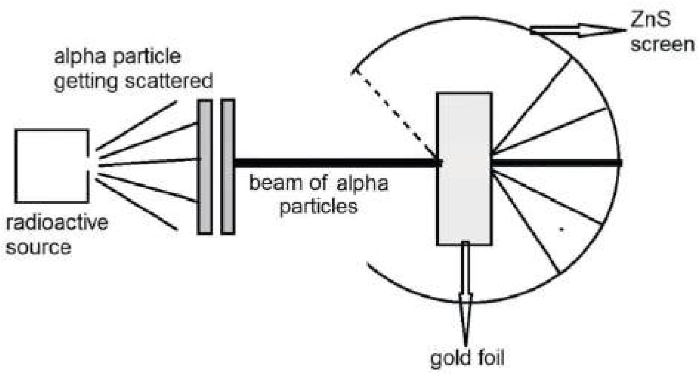


Figure 3.2: *Rutherford's alpha scattering experiment.*

Rutherford analysed the results of the experiment and proposed a new atomic model that showed that the atom has a small, dense, positively charged nucleus at its centre, surrounded by negatively charged electrons. He concluded that the deflected alpha particles were deflected by a strongly charged nucleus, while others passed straight through the atom's empty space.

This experiment provided evidence for the existence of the atomic nucleus and paved the way for further discoveries about the structure and behaviour of atoms.

Rutherford's model of the atom

Rutherford's atomic model, also known as the nuclear model, was proposed based on his famous alpha particle scattering experiment. The model introduced the concept of a small, positively charged nucleus in the centre of the atom, surrounded by negatively charged electrons. The electrons would orbit the nucleus in specific energy levels and paths, similar to the planets orbiting the sun.

Rutherford's model of the atom is shown below.

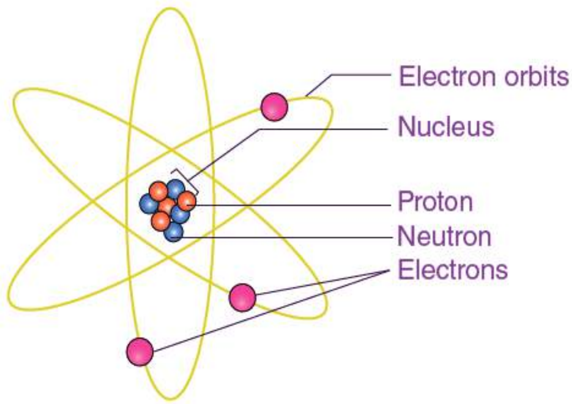


Figure 3.3: *Rutherford Model of the Atom.*

Despite its contribution to the understanding of the structure of the atom, Rutherford's atomic model had several weaknesses which are as follows:

1. The model is unable to explain the stability of atoms. In the model, negatively charged electrons move around the positively charged nucleus and should eventually lose energy and spiral into the nucleus, causing the atom to collapse. However, this does not happen in reality, and the model failed to explain why.
2. The model cannot account for the high energy emission spectra of atoms. According to the model, electrons must travel in specific paths and can only transition between certain energy levels, which would result in a limited spectrum of radiation. But experimental observations showed that the atoms emitted a much larger range of radiation than predicted by the model.
3. Rutherford's atomic model could not explain the existence of isotopes, atoms of elements with the same atomic number but different mass numbers. The model proposes that the number of electrons in an element be equal to its atomic number, which would also determine the number of protons in the element. However, isotopes of the same element have a different number of neutrons even though they have the same number of protons.

Despite these limitations, Rutherford's atomic model was a fundamental stepping stone towards the modern understanding of atomic structure and formed the basis of further models, including Bohr's atomic model, which built upon Rutherford's concept of a nucleus and attempted to address some of the model's weaknesses.

The structure of the atom

Based on the results of J.J. Thomson's cathode ray and Rutherford's alpha scattering experiments, the following structure of the atom was proposed:

The atom consists of a positively charged nucleus, which contains most of the mass of the atom. The protons are positively charged. The electrons which have a negative charge, are located outside the nucleus.

The number of electrons in an atom is equal to the number of protons, giving an atom a neutral overall charge. The electrons are held in the shells by the electrostatic attraction to the positively charged nucleus.

The diagram below shows the location, charges and relative masses (amu = atomic mass units) of the subatomic particles of the atom.

Subatomic Particles			
Particle	**Charge**	**Relative Mass (AMU)**	**Location**
Proton	Positive	1	Nucleus
Neutron	Neutral	1	Nucleus
Electron	Negative	1/1840	Outside Nucleus

Learning Task:

1. Ask learners to copy and complete the table below:

Particle	**Location**	**Relative mass**	**Charge**
electron	Outside nucleus		
proton			+1
neutron		1amu	

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

Activity-Based Learning: In mixed-ability groups,

- **a.** Use simulation videos or charts to investigate the properties of cathode rays under the following headings:
  - i. Effect in a magnetic field,
  - ii. Effect in electric field,
  - iii. Effect on photographic plate.
- **b.** Use simulation to investigate or chart to illustrate or video to show and describe J. J. Thomson's cathode ray experiment and Rutherford's alpha scattering experiment.
- **c.** Describe the structure of the atom based on analysis of the evidence gathered from both experiments.

Project – based learning: Individually or in groups, guide learners to use the evidence gathered from the experiments of JJ Thomson and Rutherford to;

- **a.** Construct a model to represent the atom.
- **b.** Draw a diagram of the atom modelled
- **c.** Display the model and diagram for class discussion.

Key Assessment

DoK level 1(Reproduction/Recall)

Which subatomic particles did Thomson include in the plum-pudding model of the atom?

DoK level 2(Skills of conceptual understanding)

How did the results of the Rutherford's gold foil experiment differ from his expectations?

References

1. SHS Chemistry Curriculum
2. Zumdall, S.S., & DeCoste (2017). Chemistry. Cengage Learning
3. Chang, R. (2017). Chemistry. McGraw-Hill Education
4. Tro, N.J., (2017) Chemistry: A Molecular Approach. Pearson Education.
5. Atkins,P.W., & de Paula, J. (2017) Atkins' Physical Chemistry. Oxford University Press.
6. Encyclopaedia Britannica. https://www.britannica.com/science/atomic-theory
7. ThoughtCo. H: https: https://www.thoughtco.com/john-daltons-atomic-model-607777
8. Encyclopaedia Britannica https://www.britannica.com/science/Rutherford-atomic-model
9. Cathode ray tube experiment, "Retrieved [February 27, 2024] from Cathode Rays – Sir William Crookes (wordpress.com)"
10. Plum pudding model of the atom, "Retrieved [February 27, 2024] from https://keystagewiki.com/index.php/Plum_Pudding_Model#/media/File%3APlumPuddingDiagram.png"
11. Rutherford alpha scattering experiment, "Retrieved [February 27, 2024] from rutherford alpha scattering experiment - Search Images (bing.com) "
12. Rutherford atomic model, "Retrieved [February 28, 2024] fromhttps://cdn1.byjus.com/wp-content/uploads/2018/07/Rutherford-Atomic-Model-and-Limitations-.png"
13. Subatomic particles displayed in a table, "Retrieved [February 27, 2024] from subatomic particles displayed in a table - Search Images (bing.com)"

Week 4

Learning Indicator(s):

1. *State the main postulates of Bohr's planetary theory and explain the importance of the quantum numbers to the electron structure of the atom.*
2. *Apply Aufbau's principle, Pauli's exclusion principle and Hund's rule of maximum multiplicity to write electron configuration of the first thirty elements of the periodic table.*

Theme or Focal Area (s): Bohr's planetary theory

Bohr's planetary theory explained the structure of an atom and the behaviour of electrons.

According to this theory, electrons orbit the nucleus in a fixed, circular orbits at specific energy levels, like the planets in the solar system. These energy levels were quantised, meaning electrons could only occupy specific energy states and could only transition between those levels by either absorbing or emitting a discrete amount of energy in the form of a photon. This theory also introduced the concept of ground state, where the electron orbits the nucleus at its lowest energy level, excited states, where the electron absorbs energy and jumps to higher energy levels.

Bohr's theory provided a foundation for modern atomic theory and led to further developments in the understanding of quantum mechanics.

Main postulates of Bohr's planetary theory

Bohr's planetary theory of the atom was based on the following postulates:

1. Electrons move around the nucleus of an atom in fixed, circular orbits.
2. The electrons can exist only in certain allowed orbits, which correspond to specific energy levels.
3. While an electron is in a particular energy level, it does not radiate energy. This energy is only emitted when the electron jumps from one energy level to another.
4. The energy of the emitted radiation corresponds to the difference in energy between the initial and the final energy levels.
5. The size of the orbit and the energy of the electron are related. Electrons in larger orbits have more energy than those in smaller orbits.
6. Electrons can only make transitions between energy levels that correspond to a specific amount of energy, known as a quantum. These transitions produce or absorb photons that have a frequency proportional to difference in energy.

These postulates provided a framework for understanding the behaviour of electrons within atoms.

The diagrams below show Bohr's model of the atom.

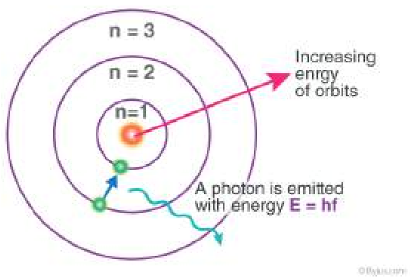


Figure 4.0: *Bohr's model of the atom*

Continuous and line spectra

a. Continuous spectrum

A continuous spectrum is a spectrum of electromagnetic radiation containing photons of all energy levels within a specific range. Unlike atomic spectra, which consists of only specific energy levels, a continuous spectrum contains a radiation of all energies resulting in a smooth display of colours or wavelengths.

Examples of sources of continuous spectra include a hot solid object, such as a light bulb filament, a glowing gas or plasma.

Continuous spectra are important in both astronomy and laboratory experiments, as they can be used to help identify the composition and temperature of objects emitting radiation.

b. Line spectrum

A line spectrum is a spectrum produced by an excited atom or molecule that contains only discrete wavelengths or colours of electromagnetic radiation. These wavelengths correspond to specific energy level transitions within the atom or molecule.

The line spectrum appears as a series of coloured lines or bands rather than a continuous spectrum, which contains radiation at all wavelengths. For example, the line spectrum of hydrogen consists of several discrete lines of colour in the visible spectrum. Each line corresponds to a specific transition between energy levels in the hydrogen atom. Similarly, each chemical element has a unique line spectrum, which can be used to identify the element based on the wavelengths of the lines observed.

Line spectra are important in many areas of science, including astronomy, chemistry and physics. Chemists use line spectra to help identify unknown substances or verify the purity of a sample.

Physicists use line spectra to study the behaviour of electrons within atoms and molecules, providing insights to the nature of matter and energy.

The diagrams below show examples of continuous and line spectra.

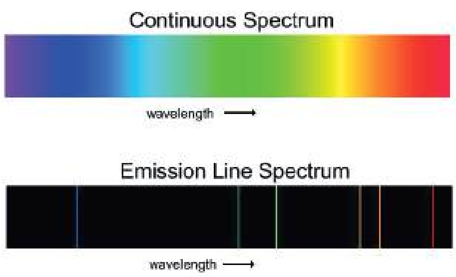


Figure 4.1: *Diagram showing continuous spectrum and emission spectrum*

Differences between a continuous spectrum and a line spectrum

There are several differences between a continuous spectrum and a line spectrum. These are:

Definition: A continuous spectrum contains radiation of all energies within a certain range, whereas an emission line spectrum contains only specific wavelengths of radiation.

Source: A (nearly) continuous spectrum is emitted by a hot, dense object such as a light bulb filament or a star, whereas an emission line spectrum is produced when an excited atom or molecule emits light.

Appearance: A continuous spectrum appears as a smooth display of colours or wavelengths while an emission line spectrum appears as a series of discrete lines or bands.

Composition: A continuous spectrum contains radiation of all energies, while an emission line spectrum contains only specific energies that correspond to the energy level transitions of the emitting atom or molecule.

Usage: Continuous spectra are used to identify the temperature and composition of a source emitting light while emission line spectra are used to identify the chemical composition of a source emitting light.

Examples: Examples of sources for continuous spectra include the sun (almost continuous), light bulbs, and blackbody radiators, while emission line spectra are typically emitted by excited atoms or molecules, such as hydrogen, helium or neon.

Relationship of the lines in the emission spectrum of hydrogen to electron energy levels

The lines in the emission spectrum of hydrogen are directly related to the electron energy levels of the hydrogen atom. When an electron in a hydrogen atom is excited, it moves from the ground state (lowest energy level) to a higher energy level. This higher energy state is not stable, and the electron will eventually return to its ground state by releasing energy in the form of a photon of light.

The energy of this photon is directly proportional to the difference between the higher and lower energy levels of the electron. Since each energy level of the electron in a hydrogen atom is fixed, the energy of the released photon is also fixed and corresponds to specific wavelength or colour of light. Therefore, each emission line in the hydrogen spectrum corresponds to a specific energy level transition for the electron in the hydrogen atom.

The lines in the spectrum represent wavelengths of the photons emitted as the electron transitions back to a lower energy level.

The diagram below shows the relationship between the line spectrum of hydrogen and energy levels.

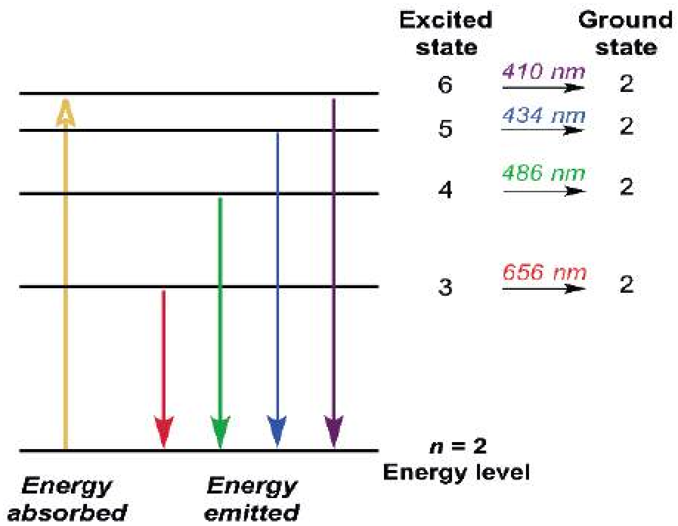


Figure 4.2: *Line Spectrum of Hydrogen and Energy Levels*

Contribution of Quantum theory towards the development of the atomic structure.

Quantum theory has made significant contributions toward the development of the atomic structure. Some of these contributions are:

Wave-particle duality: Quantum theory introduced the concept of wave-particle duality, which suggests that particles could exhibit both wave-like and particle-like behaviours. This theory helped scientists to understand the behaviour of electrons in atoms, as electrons exhibit wave-like behaviour in their movement around the nucleus.

Discrete energy levels: Quantum theory introduced the concept of energy levels in atoms. This means that electrons can only exist at certain energy levels around the nucleus and cannot exist anywhere in between.

Uncertainty principle: Quantum theory introduced the uncertainty principle, which states that it is impossible to know both the position and the momentum of an electron at the same time.

Quantum numbers: Quantum theory introduced the concept of quantum numbers which describe the energy levels and positions of electrons in atoms. These quantum numbers help predict the properties of atoms and their behaviour during chemical reactions.

Electron spin: Quantum theory also introduced the concept of electron spin, which explains why two electrons in the same electron orbital have opposite spins. This concept helps scientists understand the behaviour of electrons in atoms and their contribution to the magnetic properties of materials. These concepts introduced by quantum theory have helped scientists understand the fundamental behaviour of electrons and predict their properties and behaviour during chemical reactions.

Quantum numbers

Quantum numbers are integers or half-integers that describe the properties of electrons in an atom. There are four types of momentum/azimuthal quantum number, magnetic quantum number and spin quantum number.

Here is a brief explanation of each quantum number.

a. **Principal quantum number(n):** This quantum number determines the energy level of an electron and describes the size of the electron cloud. It can have any positive value starting from 1.

b. **Angular Momentum/Azimuthal quantum number(l):** This quantum number indicates the shape of the electron cloud or the subshell in which an electron is present. It can have values from 0 to (n-1).

c. **Magnetic quantum number(m):** This quantum number specifies the orientation of the electron cloud in space. It can have values from (-l) to (+l).

d. **Spin quantum number(s):** This quantum number describes the spin of an electron, which is a fundamental property of all particles. It can be either (+½) or (-½).

These quantum numbers play a crucial role in determining the electron configuration of an atom and understanding its behaviour. They also help us in predicting the position of electrons in an atom by providing a framework for describing energy states and orbitals.

The importance of quantum numbers to the electron structure of the atom

Quantum numbers play an important role in understanding the electron structure of an atom. The electron structure refers to the arrangement of electrons in an atom’s different energy levels and subshells.

Here are some reasons why quantum numbers are important in this regard:

a. **Describing the electron energy levels:** The principal quantum number (n) allows us to determine the energy levels available to electrons in an atom. Each energy level corresponds to an electron shell, and the value of n determines the number of subshells and electrons that can reside in each shell.

b. **Specifying the subshells:** The angular momentum/azimuthal quantum number (l) provides information about the subshells within each shell. It determines the shape of the subshell. This helps us to predict the electron distribution more accurately.

c. **Determining the electron orientation:** The magnetic quantum number (m), indicates the orientation of the electron cloud in space. It helps us to understand spatial arrangement of electrons within subshells.

d. **Predicting electron spin:** The spin quantum number (s), describes the spin of each electron, which is important for understanding the electron configuration of an atom. Two electrons with opposite spins can occupy the same orbital, which has important consequences for the chemical and physical properties of different elements.

Orbitals

An orbital is defined by a set of quantum numbers (n, l, m). Here are a few examples of orbitals:

a. The **s** orbital is the lowest orbital, with a spherical shape and can hold up to two electrons. The **s** orbital is shown below.

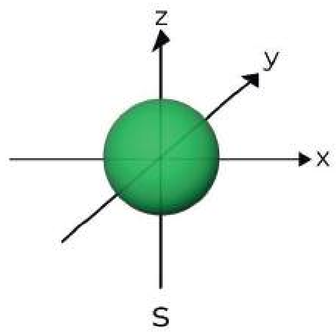


Figure 4.3: *Shape of the S-orbital*

b. The **p** orbitals (p_x, p_y and p z), have a dumb-bell shape and can hold up to six electrons (two in each orbital). These orbitals are found in the second and higher energy levels. The **p** orbitals are shown below;

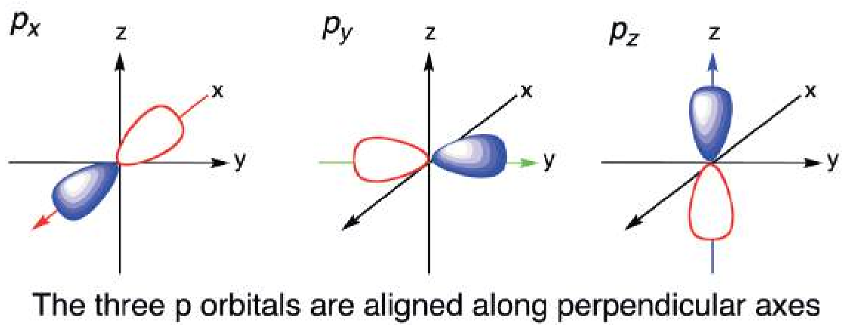


Figure 4.4: *The three P orbitals aligned along perpendicular axes*

c. The **d** orbitals have more complex shapes and can hold up to ten electrons in total (5 orbitals with 2 electrons in each orbital). These orbitals are found in the third energy level and higher.

Learning Task:

1. State Bohr's planetary model of the atom.
2. Investigate how Bohr's model explain the stability of the atom.
3. Create a comparative analysis detailing the differences and similarities between continuous and line spectra while delving into the fundamental principles that govern each of them

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

Collaborative Learning: In small mixed ability or mixed-gender (where applicable) groups,

a) Provide learners with diagrams or models illustrating the planetary structure of the atom, highlighting the fixed orbits and energy levels as propose by Bohr.

b) Provide test books, or have learners' surf the internet for articles on Bohr's theory and the stability of the atom, followed by structured writing tasks such as summarising each postulate or composing explanations in their own words.

c) Facilitate a discussion on continuous and line spectra. Assign roles within the group such as group leader, note-taker and presenter and have them create a table of differences and similarities between continuous and line spectra and have them do group presentation.

d) Provide additional readings or extension of tasks for advanced learners to explore deeper aspects of spectral analysis and offer additional support to struggling learners.

Demonstrative Learning:

a) Begin the lesson with a hands-on demonstration using physical models to illustrate the concept of quantum numbers. Use tangible objects to represent electrons and their energy levels within the atom.

b) Utilise visual aids such as diagrams to visually represent the relationship between quantum numbers and electron orbitals. Provide worked examples on assigning quantum numbers to electrons in different energy levels and provide opportunities for learners to practice.

c) Scaffold the calculations processes of the quantum numbers by providing step by step guidance. Start with simple examples and gradually increase the complexity as learners becomes more comfortable with the calculation.

d) Conclude the concept by stressing the importance of quantum numbers to the electron structure of the atom.

Project-based learning:

a) In a whole class discussion, explain the characteristics and properties of s and p orbital verbally proving real- life examples.

b) In small a mix-ability groups have learners build orbital models using clay, balloons or materials in their environment. This will help them visualise and understand the spatial arrangement of s and p orbitals.

c) Offer extension activities for advanced learners and scaffold tasks for struggling learners.

Key Assessment

DoK level 1 (Reproduction/Recall):

What was the name of Bohr's model of the atom

DoK Level 2 (Skills of conceptual understanding):

Briefly describe how Bohr's model of the atom contributed to our understanding of the stability of the atom

DoK Level 3 (Strategic reasoning):

Evaluate the possible values of the magnetic quantum number (ml) for an electron in an orbital with $l = 2$

DoK Level 4 (Extended critical thinking and reasoning):

Compare and contrast the properties of *s* and *p* in terms of shape, orientation energy level. Justify your reasoning with relevant examples

Theme or Focal Area(s): Aufbau's principle

Aufbau's principle, also known as the building-up principle, is a fundamental principle in chemistry stating that the atomic orbitals are filled with electrons in order of increasing energy. Specifically, electrons will fill lower energy atomic orbitals before moving on to higher energy levels. This principle is used to determine the electron configuration of atoms and the order in which the orbitals are filled. Electrons fill subshells in the following order: 1s, 2s, 2p, 3s, 3p, 4s, 3d.

Pauli's exclusion principle:

Pauli's exclusion principle is a fundamental principle in quantum mechanics that states that no two electrons in an atom can have the same set of quantum numbers. In other words, if two electrons are in the same orbital, they must have opposite spins. The principle is essential in determining the structure of atoms, as it limits the number of electrons that can occupy each orbital and organises the electron configuration of atoms in the periodic table.

Hund's rule of maximum multiplicity

Hund's rule of maximum multiplicity is a principle in quantum mechanics that states that within a subshell, electrons will occupy orbitals singly, with their spins parallel' before they pair up with opposite spins.

In other words, if two or more orbitals having the same amount of energy are unoccupied, then electrons will start occupying them singly, before they fill them in pairs.

This means that electron pairing in **p** and **d** orbitals cannot occur until each orbital of a given subshell contains one electron or is singly occupied.

This rule is based on the fact that, electrons in orbitals with parallel spins repel each other less than electrons with opposite spins, leading to a lower potential energy for the system. Therefore, when electrons occupy a subshell, they will make the least energetic configuration by occupying orbitals singly before pairing up.

This rule is important in determining the electronic configuration of atoms.

How to express electron configuration using s, p, d notation

To express the electronic configuration of an atom using s, p, d notation, you first need to identify the principal quantum number of the highest energy level occupied by electrons, which is equal to the period number of the element in the periodic table. Then you assign the electrons to each subshell in the following order:

1. First, the 1s orbital is filled before any other orbital.
2. Next, the 2s orbital is filled before the 2p orbitals start to fill.
3. Then, the 3s orbital is filled before the 3p orbitals start to fill.
4. After that, the 4s orbital is filled before the 3d orbitals begin to fill.

That is, 1*s* 2s 2*p* 3*s* 3*p* 4*s* 3*d*

So, for example, the electron configuration of Carbon(C), which has 6 electrons, can be expressed in s ,p and d notation as follows:
$1s^2\ 2s^2\ 2p^2$.

This indicates that the first energy level (n = 1) is filled with two electrons in the 1s orbital, while the second energy level (n = 2) is filled with four electrons in the 2*s* and 2*p* orbitals, (with two electrons in the 2*s* orbital and two electrons in the 2*p* orbitals).

How to express electron configuration using 'electrons-in-boxes' method

The electron configuration of an atom can also be expressed using the 'electrons-in-boxes' method.

In this method, each orbital is represented as a box, and the electrons are represented by arrows, with the direction indicating their spin.

For example, the electron configuration of carbon (C) with 6 electrons can be represented as follows:

1s

↑↓

2s

↑↓

2p

↑	↑	

In this representation, the first energy level (n= 1) has only one orbital, the 1s orbital, with electrons represented by a pair of arrows pointing up and down. The second energy level (n=2) contains four orbitals: the 2s orbital, which has two electrons represented by arrows pointing up and down, and the three 2p orbitals, which has two electrons represented by two arrows pointing up.

Differences in stability between fully filled, half-filled and partially filled orbitals

The stability of an atom depends on the electron configuration of its orbitals. A fully filled or half-filled subshell is more stable than a partially filled subshell.

Learning Tasks

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

1) Provide learners with a list of elements with their atomic numbers and instruct them to write the electron configuration for each element using the basic rules ($_3$Li, etc.).

2) Provide learners with elements that have unusual electron configuration or that require exceptions to the basic rules and task them to write their electron configuration ($_{24}$Cr, etc.)

3) Provide learners with scenarios involving elements and ask them to predict how the electron configuration of the atoms involved will affect the properties of their compound (Noble gases and group one elements)

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

Group Learning:

a) Begin by assessing learners' prior knowledge on electron configuration and their understanding of the first twenty elements using a short quiz or discussion.

b) Put learners in mixed ability groupings and let them research from the internet, textbooks and any other sources on the rules governing the writing of electron configuration and their applications in writing electron configuration. Have learners summarise the rules. Offer support to learners who may require additional assistance on the rules governing the writing of electron configuration.

c) In ability groupings, offer worksheets or practice problems on the application of the rules in writing electron configuration with varying level of difficulty and provide additional support or challenge as needed by the ability groupings

Activity-based Learning:

a) In ability groupings, assign learners with practice questions that requires writing of electron configuration in different notation (i.e. s, p and d notion and the electron in box method)

b) Have groups come out to present their solutions. Task other groups to identify violations if any in the presentations.

c) In a whole inclusive class, discuss the stability associated fully filled, half-filled and partially filled orbitals in subshells.

Digital learning: Have learners watch videos or observe demonstrations of the process of filling orbitals using the s, p and d notations and electron in box method

Key Assessment

DoK Level 2: (Skills of conceptual understanding):

Write the electron configuration of oxygen using s, p and d notation

DoK Level 2: (Skills of conceptual understanding):

Explain the Aufbau principle and how it applies to writing of electron configurations

DoK Level 3: (Strategic reasoning):

Explain why the electron configuration of chromium (Cr) is different from what is expected based on the Aufbau principle

DoK Level 4: (Extended critical thinking and reasoning):

Analyse the relationship between electron configuration and the chemical properties of elements within a period.

References

1. SHS Chemistry Curriculum.
2. Zumdall, S.S., & DeCoste (2017). Chemistry. Cengage Learning.
3. Chang, R. (2017). Chemistry. McGraw-Hill Education.
4. Tro, N. J. (2017). Chemistry: A Molecular Approach. Pearson Education.
5. Encyclopaedia Britannica. https://britannica.com/science/Bohr-atomic-model
6. https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)

7. https://www.khanacademy.org/science/physics/quantum-physics
8. https://www.britannica.com/science/Pauli-exclusion-principle
9. s and p-orbital diagram. Retrieve on [February 27, 2024] from www.google.com
10. Bohr model of atom. Retrieve on [February 27, 2024] from bing.com
11. Continuous spectrum and emission spectrum. Retrieve on [February 27, 2024] from Google Search
12. Line spectrum showing H and energy level. Retrieve on [February 27, 2024] from Google Search

Week 5

Learning Indicator(s): *Describe radioactivity, and the properties of radiations and compare isotopes based on their stability as well as their applications in everyday life.*

Theme or Focal Area: Relative atomic mass

Relative atomic mass is the average mass of an atom of an element, considering all its naturally occurring isotopes, relative to the mass of an atom of carbon-12, which has been assigned a mass of exactly 12.

This value is expressed in atomic mass units (amu). The relative atomic mass is determined by the abundance of each isotope of an element.

Relative molecular mass

Relative molecular mass is the sum of the relative atomic masses of all the atoms in a molecule, relative to the mass of carbon-12, which has been assigned a mass of exactly 12.

The relative molecular mass is calculated by adding up the atomic masses of all the atoms in a molecule, considering the number of atoms in each element present in a molecule.

Mass spectrometer

A mass spectrometer is an instrument used for measuring the mass-to-charge ratio (m/z) of ions in a sample. It operates by generating ions from a sample and then separating the ions based on their mass-to-charge ratio (m/z) using electric and magnetic fields. The separated ions then hit a detector where they create a signal that is proportional to their abundance. This signal is then analysed to determine the mass-to-charge ratio (m/z) of the ions and their relative abundance, providing information about the composition of the sample.

The mass spectrometer is shown below.

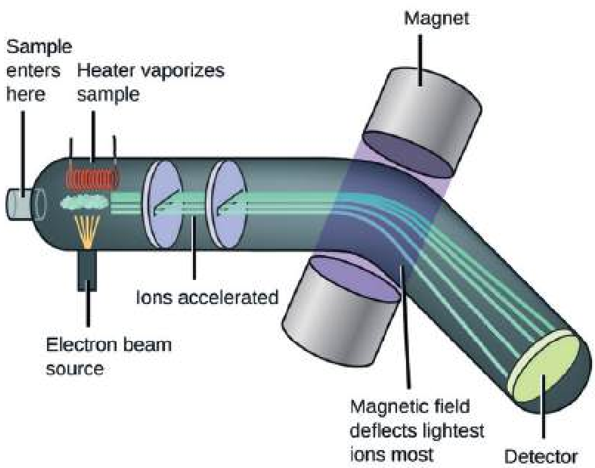


Figure 5.0: *Mass spectrometer.*

Mass spectrometers are used in a wide variety of fields, including chemistry, biochemistry, physics, geology and environmental science, for applications such as identifying the presence of specific compounds in a sample, analysing the composition of proteins and other biomolecules, studying the isotopic composition of elements, and monitoring air and water pollution.

Parts of the mass spectrometer and how they work

The five main processes which take place in the mass spectrometer are:

1. A sample of the substance is vaporised in the vaporisation chamber.
2. The vapour is ionised into positive ions in the ionisation chamber.
3. The positive ions are attracted and accelerated by an electric field.
4. The accelerated positive ions are deflected in the magnetic field and focused onto a detector according to their mass-to-charge ratio (m/z).
5. The ions generate a current or signal that is proportional to their abundances. The detector then draws a mass spectrum of the different ions.

The diagrams below show how the mass spectrometer is used.

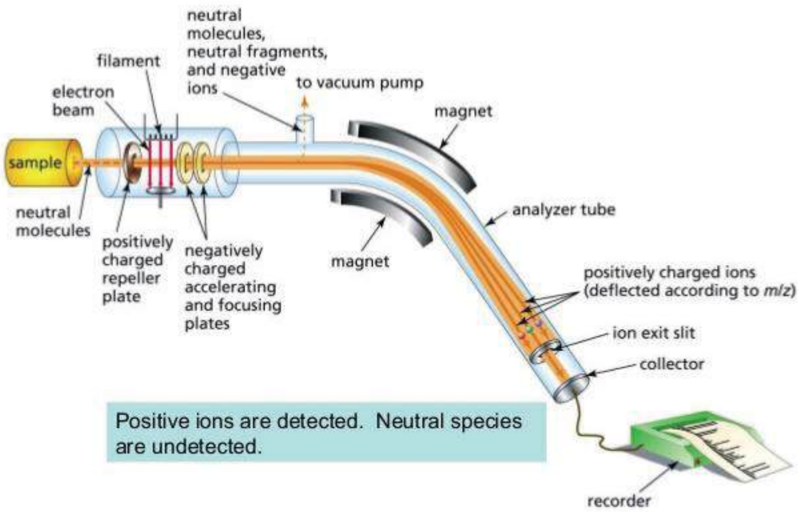


Figure 5.1: *Diagram of a Mass Spectrometer Processes that takes place in it*

Source: *K. Bhavyasri et al., IJSRR 2019, 8(2), 3161-3176*

Mass spectrum

A mass spectrum is a graphical representation of the ion intensities (relative abundances) as a function of their mass-to-charge ratios (m/z). Simply put, a mass spectrum is a graph of the percentage abundance versus the relative atomic masses of the ions present in the sample.

Generally, the mass spectrum shows a series of peaks, each representing an ion with a specific mass-to-charge ratio (m/z). The height of the peak is proportional to the abundance of the ion in the sample.

The mass spectrum of chlorine is shown below.

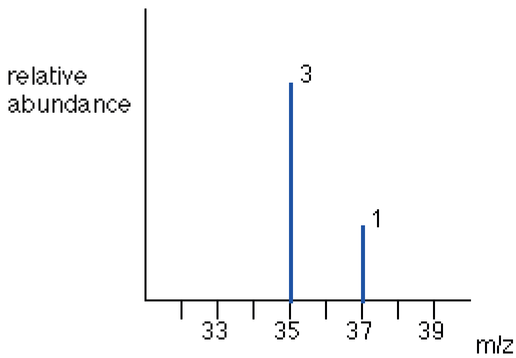


Figure 5.2: *Diagram of a Mass Spectra of Chlorine Element.*

How to calculate the relative atomic mass of different elements

The relative atomic mass of an element can be calculated by using

1. The mass spectrum
2. Percentage abundance data.

Generally, the relative atomic mass is calculated by using the following steps:

1. Identify the isotopes of the element and their percentage abundances. The percentage abundance refers to the amount of each isotope present in a sample, expressed as a percentage of the total number of atoms of the elements.
2. Multiply the mass of each isotope by its percentage abundance.
3. Add up the products obtained in step 2.
4. Divide the sum obtained in step 3 by 100 to obtain the relative atomic mass of the element.

For example, let us consider the element boron. Boron has two isotopes, boron-10 and boron-11 with natural abundances of approximately 20% and 80% respectively. To calculate the relative atomic mass of boron, we can use the steps as follows:

1. Boron-10 has a percentage abundance of 20% and boron-11 has a percentage abundance of 80%.
2. The relative atomic mass (A_r) of boron-10 is 10 amu and the relative atomic mass (A_r) of boron-11 is 11 amu.

   Therefore, 10 x 20 =200 and 11x 80 = 880
3. 200 + 880 = 1080
4. $\frac{1080}{100} = 10.80$

   Therefore, the relative atomic mass of boron is 10.80 amu

Example:

Chlorine naturally exists as two isotopes, chlorine-35 and chlorine-37. The abundance of chlorine-35 is 75% and the abundance of chlorine-37 is 25%.

Calculate the relative atomic mass of chlorine.

Solution:

1. % of chlorine-35 = 75% and % of chlorine-37 = 25%
2. The A_r of chlorine-35 is 35 amu and the A_r of chlorine-37 is 37 amu.

   Therefore 75 x 35 = 2625 and 25 x 37 = 925
3. 2625 + 925 = 3550
4. $\frac{3550}{100} = 35.50$

Therefore, the relative atomic mass of chlorine is 35.50 amu.

Learning Task:

1. Explain relative atomic mass and relative molecular mass.
2. Describe how the mass spectrometer operates.
3. a). Design a model to show how to calculate the relative atomic mass of different elements
   b). Use the information provided in table1 to calculate the relative atomic mass of naturally occurring copper isotopes. Leave your answer in 1 decimal place.

Table 1:

Mass number	% Abundance
63	69
65	31

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

1. a) Through a whole class discussion, guide learners to explain relative atomic mass and relative molecular mass.
   b) In ability groupings describe the principal parts of a mass spectrometer and explain how it works, using charts or pictures, or models.
2. **Problem – solving approach**
   a) With ability groupings, guide learners to identify peaks on a simple mass spectrum and use them to calculate the relative abundance and masses of isotopes. *NB: Support those who are struggling to do the calculations and increase the difficulty levels of the calculations for those who are finding it very easy.*
   b) Individually, let learners design a model as a guide to calculate the relative atomic mass of different elements from:
      i). mass spectrum
      ii). percentage abundance data

Key Assessment

DoK Level 1 Reproduction/Recall

1. Define at least one of the following:
   a). Relative atomic mass
   b). Relative molecular mass.

DoK Level 2 Skills of conceptual understanding

Describe how the mass spectrometer operates.

DoK Level 3 Strategic reasoning

An instrument called a mass spectrometer can be used to determine the masses and relative abundances of different isotopes of an element. Find out how a mass spectrometer works and design a poster to illustrate your findings.

DoK Level 3 Strategic reasoning

1. Design a model to demonstrate how to calculate the relative atomic mass of different elements.
2. Copper has two isotopes, copper-63 and copper-65. The percentage abundance of copper-63 is 69% and that of copper-65 is 31%.

Sketch the mass spectrum of copper.

3. In a sample containing 100 atoms of boron, 80 were boron-11 atoms while 20 were the boron-10 isotope. Determine the relative atomic mass of boron.

Theme or Focal Area: Radioactivity

Radioactivity refers to the process in which unstable nuclei decay and emit energetic particles or electromagnetic radiation. These emissions are referred to as ionising particles and include alpha particles, beta particles, and gamma rays.

Radioactivity is a natural phenomenon that occurs spontaneously in certain isotopes of elements. These isotopes are called radioactive isotopes or radionuclides. Radioactive decay proceeds at a fixed and predictable rate, characterised by a half-life. The half-life is the time required for half of the radioactive atoms in a sample to decay into stable atoms.

Radioactivity has many practical applications in various fields including medicine (e.g. nuclear medicine), industry (radiography), and energy production (nuclear power). However, it also poses certain health risks and safety concerns, which require careful management and regulation.

Differences between nuclear reactions and chemical reactions

Nuclear reactions and chemical reactions are two different reactions that involve changes in the composition of matter.

The main differences between these two types of reactions are as follows:

1. **Nature of particles involved:** In a chemical reaction, the atomsof the reacting substances do not change their identities. In contrast, in a nuclear reaction, the atomic nuclei of the reacting substances undergo changes, resulting in the formation of different nuclei and subatomic particles.
2. **Energy changes:** Nuclear reactions involve much larger energy changes than chemical reactions. This is because nuclear reactions involve changes in the binding energies of atomic nuclei, which are much larger than the energies involved in the breaking and forming of chemical bonds.
3. **Rate of reaction:** Chemical reactions generally occur at a faster rate than nuclear reactions. This is because chemical reactions involve the interaction of electrons, which are much lighter and move much faster than atomic nuclei.
4. **Triggering factors:** Chemical reactions are triggered by factors such as temperature, pressure, and concentration, while nuclear reactions are triggered by factors such as particle bombardment and radioactive decay.

   In other words, chemical reactions are affected by environmental conditions such as temperature and pressure but nuclear reactions are not.
5. **Product stability:** The products of a chemical reaction are typically more stable than the products of a nuclear reaction. This means that the products of a nuclear reaction may undergo further decay or transformation into products over time.
6. **Emission of radiation:** Nuclear reactions emit radiation such as alpha particles, beta particles, and gamma rays while chemical reactions do not.

Properties of alpha, beta, and gamma radiation

Alpha, beta, and gamma radiation are types of ionising radiation, which have different properties based on their composition and energy.

Here are the properties of each type of radiation:

Alpha radiation:

1. Consists of a helium nucleus of two protons and two neutrons.
2. Has a charge of +2 and a mass number of 4.
3. Travels only a few centimetres in the air and are blocked by a sheet of paper or skin.
4. Has a high ionisation potential and can cause significant damage to living tissue, if inhaled, ingested, or absorbed.
5. Emitted by heavy nuclei like uranium and radium as they decay into lighter elements.
6. Symbol $^{4}_{2}He$

Beta radiation:

1. Consists of a negatively charged particle identical to a high-energy electron.
2. It has a charge of -1 and has a negligible mass.
3. Are more penetrating than alpha particles and can penetrate several millimetres of aluminium or plywood.
4. Have lower ionisation potential than alpha particles and may cause damage to living tissue if they penetrate the skin or are inhaled.
5. Are emitted by elements with an excess or deficiency of neutrons, like carbon-14.
6. Symbol $^{0}_{-1}e$

Gamma radiation:

1. Consists of high-energy photons of electromagnetic radiation.
2. Has no charge or mass.
3. Is highly penetrating and can pass through several centimetres of lead or concrete.
4. Has a lower ionisation energy potential than alpha and beta radiation, but can still cause damage to living tissues by ionising atoms along their path.
5. Emitted by the nucleus of some radioactive atoms, such as Cobalt-60.
6. Symbol γ

Factors that affect the stability of nuclei

The stability of atomic nuclei is determined by the balance between the strong nuclear force, which holds the nucleus together and the electrostatic repulsion between protons in the nucleus.

The neutron-to-proton ratio and the binding energy per nucleon *(one of the subatomic particles of the nucleus)* are the two most important factors that determine the stability of an atomic nucleus.

1. **Neutron-to-proton ratio:** The neutron-to-proton ratio is the ratio of the number of neutrons to the number of protons in the nucleus. This ratio affects the stability of the nucleus because it determines the balance between the strong nuclear force and the electrostatic repulsion between protons.

When the ratio is low, meaning there are more protons than neutrons, the electrostatic forces between the protons are stronger than the strong nuclear forces holding the nucleus together, making the nucleus unstable. On the other hand, nuclei with high neutron-to-proton ratios are also unstable as there are more neutrons than protons and the strong nuclear force becomes weaker leading to decay of the nucleus.

For stability, the ratio must be close to one (1).

2. **Binding energy per nucleon:** The energy needed to keep the protons, neutrons, and electrons together in an atom is generally called binding energy.

   The binding energy per nucleon refers to the amount of energy required to remove a nucleon from the nucleus of an atom. It is calculated by dividing the total binding energy of a nucleus by the total number of nucleons (protons and neutrons) in the nucleus. The binding energy per nucleon reflects the strength of the strong nuclear force that holds the nucleus together.

   When the binding energy per nucleon is high, the nucleus is more stable, meaning the strong nuclear force is strong enough to overcome the electrostatic repulsion between the protons. Nuclei with low binding energy per nucleon are less stable and tend to undergo radioactive decay to increase their binding energy per nucleon.

Half-life of a nuclide

The half-life of a nuclide is the time it takes for half of the atoms in a sample of the nuclide to decay. This means that after one half-life has passed, half of the original nuclide atoms will have decayed, and the remaining half will remain.

The half-life is a characteristic property of each radioactive nuclide and is constant for that nuclide, which means that regardless of the size of the sample, each atom has the same probability of decaying during a certain period. The half-life of a nuclide is dependent on the type of decay process it undergoes, such as alpha, beta, or gamma decay.

Some nuclides have very short half-lives, which means they decay rapidly and are highly radioactive for only a short time, while others have very long half-lives, which means they decay slowly and can remain radioactive for thousands or millions of years.

Half-life is an important concept in the study of nuclear chemistry and is used to calculate the amount of time needed for a given amount of radioactive material to decay to a desired amount, as well as how much of a substance will remain after a certain period of time. It is also used in radioactive dating methods to estimate the age of geological samples or archaeological artefacts.

How to calculate the half-life using experimental data and calculation

To calculate the half-life of a radioactive nuclide using experimental data, you need to measure the decay rate or the amount of radioactive material remaining at specific time intervals.

Here is how to calculate half-life using experimental data and calculation:

1. Measure the initial activity (or quantity) of the radioactive sample i.e. No.
2. Measure the activity (or quantity) of the radioactive sample at a specific time interval, Δt i.e. Nt
3. Calculate the fraction of the original material that has decayed during the time interval, Δt Using the formula: Fraction decayed $= \frac{\text{activity at } \Delta t}{\text{initial activity}}$

$$= \frac{Nt}{No}$$

4. **4.** Calculate the natural logarithm of the fraction decayed:

   i.e. $\ln\left(\frac{Nt}{No}\right) = \lambda \Delta t$ (λ is the decay constant of the nuclide)

5. **5.** Solve for the decay constant:

   $\lambda = \frac{-In(Nt\_No)}{\Delta t}$

6. **6.** Calculate the half-life, $t_{1/2}$:

   i.e. $t_{\frac{1}{2}} = \frac{In2}{\lambda}$, where (In 2) is the natural logarithm of 2, which is approximately 0.693.

   i.e. $t_{\frac{1}{2}} = \frac{0.693}{\lambda}$.

Note that you can repeat this process with different time intervals, and the resulting half-lives should be consistent if the sample is homogeneous and the decay process is steady.

NB:

The expressions $In\left(\frac{No}{Nt}\right) = \lambda \Delta t$ and $\log\left(\frac{No}{Nt}\right) = \frac{\lambda t}{2.303}$ can also be used.

Examples:

1. **1.** A sample of pure iodine-131 has a decay rate of 600 counts per second. 16 days later, the decay rate has dropped to 150 counts per second. Calculate the half-life of the sample of iodine.

Solution:

No = 600 counts per second, Nt = 150 counts per second and

Δt = 16 days.

$In\left(\frac{No}{Nt}\right) = \lambda \Delta t$

$In\left(\frac{600}{150}\right) = 16\,\lambda$

$In\ 4 = 16\,\lambda$

$\lambda = \frac{In\ 4}{16} = \frac{1.386}{16}$

$= 0.0866\ day^{-1}$

Half-life, $t_{\frac{1}{2}} = \frac{0.693}{\lambda}$

$= \frac{0.693}{0.0866} = 8$ days.

2. **2.** How long does it take a 100.0 g sample of As-81, with a half-life of 33 seconds to decay to 6.25g?

Solution:

No = 100 g, Nt = 6.25 g, and $t_{\frac{1}{2}} = 33$ s.

$t_{\frac{1}{2}} = \frac{0.693}{\lambda}$

$\lambda = \frac{0.693}{t_{1/2}}$

$= \frac{0.693}{33}$

$= 0.021 s^{-1}$

$In\left(\frac{No}{Nt}\right) = \lambda \Delta t$

$In\left(\frac{100}{6.25}\right) = 0.021 \Delta t$

$In\ 16 = 0.021 \Delta t$

$$\Delta t = \frac{\text{In}16}{0.021} = \frac{2.773}{0.021}$$

$\Delta t = 132.0$ s

Uses of radioisotopes and the principle behind each use

Radioisotopes are used in a wide range of applications, from medical diagnosis and treatment to industry and scientific research.

Here are some common uses of radioisotopes and the principle behind each use:

1. **Medical diagnosis:** Radioisotopes like technetium-99, iodine-131, and gallium-67 are commonly used in medical imaging techniques such as positron emission tomography (PET), single photon emission computed tomography (SPECT,) and computed tomography (CT) scans.

   These isotopes emit gamma rays that can be detected by special cameras to create images of internal body structures, allowing doctors to diagnose and monitor diseases such as cancer, heart disease, and neurological disorders.

2. **Medical treatment:** Radioisotopes can be used to treat certain types of cancer, such as thyroid cancer, by selectively targeting and destroying cancerous cells.

   Iodine-131 is used to treat thyroid cancer, while strontium-89 is used to relieve pain in bone cancer patients.

3. **Industrial applications:** Radioisotopes such as cobalt-60 and iridium-192 are used to sterilise medical equipment and food products, as well as measure the thickness of materials in manufacturing processes.
4. **Agricultural applications:** Radioisotopes can be used to measure plant and soil properties, such as moisture content and nutrient uptake, to improve crop yield and quality. Radioisotopes are also used for insect and pest control.
5. **Scientific research:** Radioisotopes can be used to study various biological and chemical processes in cells, tissues, and organisms. Carbon-14 is used in dating archaeological artefacts and geological samples, while tritium (hydrogen-3) is used to label molecules and track their movement within biological systems.
6. **Source of heat energy:** Controlled nuclear fission is used to generate electricity and heat energy.

   The principle behind each use of radioisotopes is based on their unique properties, including their half-life, decay mode, and energy spectrum.

   For example, in medical imaging, radioisotopes that emit gamma rays are used because they can penetrate the body and be detected by specialised cameras. In cancer treatment, radioisotopes that emit beta particles or alpha particles are used because they can travel short distances and deliver a high dose of radiation to cancerous cells while sparing healthy tissue.

   In industrial and agricultural applications, radioisotopes are used to measure various properties based on their ability to emit radiation or interact with other materials.

   In a nutshell, the uses of radioisotopes are based on their ability to provide useful information, diagnose and treat diseases, improve quality control in manufacturing, and advance scientific research.

How to complete and balance simple nuclear reactions

Balancing a nuclear reaction involves ensuring that the total number of protons (atomic number, or proton number) and the total mass numbers (nucleon number) are the same on both sides of the equation. Additionally, the sum of the charges on each side must be equal. Example of a balanced nuclear reaction is shown below:

$$^{238}_{92}\mathrm{U} \longrightarrow {}^{234}_{90}\mathrm{Th} + {}^{4}_{2}\mathrm{He}$$

Assist learners to balance the following nuclear reactions:

1. $^{232}_{690}Th \rightarrow \; ^{288}_{88}Ra + \quad X$
2. $^{14}_{6}C \rightarrow ^{14}_{7}N + Y$
3. $^{9}_{4}Be + Z \rightarrow ^{12}_{6}\mathrm{c} + ^{1}_{0}n \; (^{1}_{0}n = neutron)$

Note that nuclear reactions involve the emission or capture of particles, such as alpha particles (helium nuclei) or beta particles (electrons).

These particles should be accounted for in the equation, with the appropriate symbol.

Risks associated with radioactivity

Radioactivity can pose several health and environmental risks. Here are some of the risks associated with radioactivity:

1. **Radiation exposure:** Exposure to ionising radiation emitted by radioactive materials can damage the DNA in cells, leading to cellular mutations and cancer. Long-term exposure to low levels of radiation is associated with an increased risk of cancer, while high levels of radiation exposure can cause sickness and death.
2. **Health risk to workers:** Individuals who work with or around sources of radiation, such as nuclear power plant workers and medical professionals, are at a higher risk of radiation exposure and associated health risks.
3. **Environmental risks:** Radioactive materials can contaminate soil, water, and air, leading to environmental contamination and increased radiation exposure for humans and other living organisms.
4. **Nuclear accidents:** Accidents at nuclear power plants can release large amounts of radioactive materials into the environment, leading to increased radiation exposure and environmental contamination.
5. **Nuclear weapons:** The use of nuclear weapons can release large amounts of ionising radiation, causing both immediate and long-term health effects for humans and the environment.

   To mitigate the risks associated with radioactivity, it is important to follow proper safety protocols when working with radioactive materials. This includes using protective equipment, maintaining proper monitoring and dosimetry practices, and ensuring that radioactive waste is properly disposed of and secured.

   In the event of a nuclear accident or radiation release, evacuation procedures and clean-up exercise should be carried out to minimise exposure and contamination.

Learning Task:

1. List at least three uses of radioisotopes
2. Explain at least two uses of radioisotopes
3. Determine how to complete and balance simple nuclear reactions

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

Collaborative Learning

a) Guide learners in pairs to discuss radioactivity and distinguish between nuclear reactions and chemical reactions. Learners present their responses in a power point or flow charts or any appropriate format of their choice.

b) In mixed-ability groupings, support learners to use the internet to investigate and present the descriptions of the properties of alpha, beta and gamma radiations to the whole class.

c) In ability groupings, guide learners to complete and balance simple nuclear reaction equations.

d) Guide learners in pairs to discuss and present their findings on why certain nuclei are unstable in terms of neutron–to–proton ratio and binding energy per nucleon to the whole class.

e) Using think-pair-share strategy, let leaners discuss the uses of radioisotopes and explain the principle behind each use as well as the risks associated with radioactivity.

Problem – solving approach,

a) Guide learners individually to determine the half-life of a nuclide from experimental data through calculations.

b) Lead an activity on how to use half-life information to determine the number of radioisotopes remaining at a given time.

c) Learners individually determine how to complete and balance simple nuclear reactions and present their findings to the whole class using any appropriate mode of their choice

Key Assessment

DoK Level 1: Reproduction/Recall

List at least three (3) uses of radioisotopes.

DoK Level 2: Skills of conceptual understanding

Explain at least two uses of radioisotopes.

DoK Level 3: Strategic reasoning

1. Explain how to complete and balance simple nuclear reactions.
2. Complete and balance the equation below:

$${}^{27}_{13}Al + {}^{4}_{2}He \rightarrow \quad {}^{30}_{15}P + X$$

Reference

1. SHS Chemistry Curriculum
2. Gross, J.H. (2015). Mass Spectrometry: A textbook. Springer.
3. McLafferty, F.W.,& Turecek, F. (1993). Interpretation of mass spectra. University Science Books
4. Skoog, D.A., Holler, F,J.,& Crouch, S. R. (2017). Principles of Instrumental Analysis. Cengage Learning
5. Knoll, G.F., (2010). Radiation detection and measurement. John Wiley & Sons.
6. International Atomic Energy Agency. (2019). Application of Nuclear Technique. https://www.iaea.org/topics/nuclear-technology-and-applications
7. https://www.epa.gov/radiation/radionuclides

8. Keller, C.,&Nornberg, P. (2018). Radioactivity. In Encyclopedia of Geochemistry (pp 1-3). Springer. https://doi.org/10.1007/978-3-1-319-39312-4_348-1

9. Pomme, S., et al (2014). Half-life. In Fundamentals of Nuclear Science and Engineering (3rd ed.,pp. 129-155) Springer-Verlag. https://doi.org/10.0007/978-3-642-40719-6_5

SECTION 2: CONCEPT OF THE MOLES

Strand: **Physical Chemistry**

Sub-Strand: Matter and its properties

Content Standard: Demonstrate an understanding of the mole concept and its significance to the quantitative analysis of chemical reaction.

Learning Outcome: *Use the mole concept to determine the amount and quantity of various substances involved in chemical reactions.*

INTRODUCTION AND SECTION SUMMARY

This section introduces learners to fundamental concepts such as relative atomic mass, relative molecular mass, and the mole as a unit of amount of substance. They will learn to perform calculations based on the amount of substance and understand the importance of the mole concept in preparing standard solutions.

SUMMARY OF PEDAGOGICAL EXEMPLARS

Various pedagogical approaches, including activity-based learning, talk-for-learning, problem-based approaches, think-pair-share activities, and experiential learning, will be use to facilitate understanding. These methods aim to engage learners actively in the learning process, promoting critical thinking, problem-solving, and practical application skills.

ASSESSMENT SUMMARY

Assessment of learners will include different levels of depth of knowledge (DoK) to evaluate the effectiveness of the pedagogical strategies employed. This multifaceted approach ensures that learners not only grasp the concepts but also develop critical thinking and problem-solving abilities. By undergoing diverse assessments, learners gain a comprehensive understanding of the subject matter and its everyday life applications.

Week 6

Learning Indicator(s):

1. *Explain relative atomic mass and relative molecular mass.*
2. *Describe the atomic mass unit as an average mass.*
3. *Describe the mole as a unit of the amount of substance.*

Theme or Focal Area: Relative Atomic mass (Ar)

It is defined as the average mass of one atom of the element compared to 1/12th of the mass of one atom of carbon – 12.

Mathematically,

$$\text{Ar} = \frac{\text{Average mass of one atom of the element}}{\frac{1}{12}\text{th of the mass of one atom of carbon} - 12}$$

It is denoted by Ar and has no unit.

Relative Molecular Mass (Mr)

It is defined as the average mass of one molecule of a substance compared with 1/12th of the mass of one atom of carbon-12.

$$Ar = \frac{\textit{Average mass of one molecule of the substance}}{\frac{1}{12}\text{th } \textit{of the mass of one atom of carbon} - 12}$$

It also has no unit. For ionic compounds, the relative molecular mass is called its Relative formula mass (as ionic substances do not exist as molecules).

Relative molecular mass in the sum of the masses of the elements that make up the molecule.

Example:

Determine the relative molecular mass of ammonia gas, (NH_3). [N = 14, H = 1].

Solution

$$\begin{aligned} Mr\ (NH_3) &= Ar\ (N) + 3Ar(H) \\ &= 14 + 3(1) \\ &= 17 \end{aligned}$$

Learning tasks:

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

Learners

1. Find out why carbon-12 isotope is use as a reference scale for measurement.
2. Use a beam balance to demonstrate how to determine the mass of an element or compound.

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

Prior to the lesson: Guide learners to find out why carbon-12 isotope is used as a reference scale for measurement.

Activity-based learning

In mixed ability groups, guide learners on how to use a beam balance to determine the mass of an element or compound by placing the standard (Carbon-12) in one pan and the mass to be determined in the other pan. Offer adequate support to slow learners and virtually impaired learners.

Key Assessment:

DoK level 1: Reproduction/ Recall

Define relative atomic mass.

DoK level 2: Skill of conceptual understanding

1. State at least two differences between atomic mass and relative atomic mass.
2. Determine the relative molecular mass of methane, (CH_4). [C = 12, H = 1].

DoK level 3: Strategic Reasoning/Thinking

Explain why the carbon-12 isotope is chosen as a reference atom for the measurement of the masses of elements.

Theme or Focal Area(s): The Atomic mass unit

The relative atomic mass scale is based on an isotope of carbon - 12. Carbon-12 contains 6 protons and 6 neutrons and mass of 12 atomic mass units. The carbon - 12 scale is therefore defined as an atomic mass reference scale in which one atom of carbon - 12 isotope has 12 units.

Therefore,

Mass of one carbon - 12 = 12 amu.

$$1 \text{ amu} = \frac{\text{Mass of one carbon} - 12}{12}$$

Recall that, most naturally occurring elements have different isotopes with different natural abundance and masses. Therefore, relative atomic mass is an average mass.

Application in everyday life

- Relative atomic mass and relative molecular mass are used in calculating the concentration of stock solution from chemical stores.
- The idea of relative molecular mass or formula mass and the law of conservation of mass are used to do quantitative calculations in chemistry.
- The idea of relative atomic mass is used to determine the empirical formula of a substance.

Example:

One atomic mass unit of Carbon – 12 is 1.6603×10^{-24} g. If the average mass of an atom of oxygen is 2.65659×10^{-23} g. Determine its relative atomic mass.

Solution

$$Ar\ (O) = \frac{Average\ mass\ of\ one\ atom\ of\ the\ element}{\text{1/12th of the mass of one atom of carbon} - 12}$$

$$Ar\ (O) = \frac{2.65659 \times 10^{-23} g}{1.6603 \times 10^{-24} g}$$

$$= 16.0$$

***NB**: One atomic mass unit of Carbon – 12 is the same as* 1/12thof the mass of one atom of carbon − 12.

Learning tasks:

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

Learners

1. Explain how the atomic mass unit (amu) of an individual particle (atom or molecule) is obtained.

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

Talk-for-learning:

In mixed ability groups, guide learners to explain how the atomic mass unit, (amu) of an individual particle (atom or molecule) is obtained by using a video or power point presentation. Encourage slow learners to talk by directing questions to them.

Key Assessment:

DoK level 1: Reproduction/ Recall

a. Explain atomic mass unit, amu.

b. State the expression for calculating the relative atomic mass of elements.

DoK level 2: Skills of conceptual understanding

a. The atomic mass unit of Carbon –12 is 1.6603 x 10^{-24} g. If the average mass of an atom X is 6.63310 x 10^{-23} g. Determine its relative atomic mass.

DoK level 3: Strategic Reasoning/Thinking

Explain how the relative atomic mass can be used to calculate the empirical formula of a substance.

Theme or Focal Area: The mole as a unit of amount of substance

In everyday life, units such as pair and dozen are used to represent a specific number of items. Scientists use the term **mole** to represent specific number of elementary entities (atoms, ions or molecules)

One mole of a substance is defined as the amount of substance that contains as many elementary entities as there are atoms in 12g of the Carbon-12 isotope.

1 mole of every substance contains 6.02×10^{23} elementary entities.

For example, 1 mole of magnesium metal contains 6.02×10^{23} atoms inside it.

How do you determine the number of particles (N) of a substance contained in a given number of moles (n)?

Number of formula units in 1 mole of any substance = $1 \times 6.02 \times 10^{23}$

Number of formula units in 2 moles of any substance =$2 \times 6.02 \times 10^{23}$

Number of formula units in n mole of any substance = $n \times 6.02 \times 10^{23}$

$$N = n \times 6.02 \times 10^{23}$$

but 6.02×10^{23} is termed as **Avogadro's number** or **constant** and it is denoted by N_A or L

Mathematically, $n = \frac{N}{N_A}$ or $n = \frac{N}{L}$

where,

n = number of moles (amount of substance) measured in mol. The mole is the base unit of the fundamental quantity called the amount of substance.

N = number of entities

L = Avogadro's number expressed as defined particles mol^{-1}. L is a molar quantity, that is, a quantity expressed per mole.

Example:
Calculate the number of moles contained in 9.5×10^{23} molecules of oxygen. [$L = 6.02 \times 10^{23}$]

Solution: Use the problem-solving approach.

(a) Analyse the question

Known

Number of molecules
Avogadro's Constant
Formula to use: $n = \frac{N}{L}$

Unknown

Amount of substance

Solve: Apply the formula

$N = 9.5 \times 10^{23}$

$L = 6.02 \times 10^{23}$

n = ?

$n = \frac{N}{L}$

Substituting the values,

$$n = \frac{9.5 \times 10^{23}\ molecules}{6.02 \times 10^{23}\ molecules\ mol^{-1}}$$

n = 1.58 mol

Learning tasks:

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

Learners

1. Explain the mole in relation to various elementary entities (atoms, ions, molecules, electrons, protons, neutrons).
2. Discuss the relationship between the Mole and Avogadro's Constant.

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

Talk-for-learning

Through a whole class activity:

1. In mixed ability groups, guide learners to explain the mole in relation to various elementary entities (atoms, ions, molecules, electrons, protons, neutrons). Encourage slow learners to talk by directing questions to them.
2. In ability groups, guide learners to discuss the relationship between the Mole and Avogadro's Constant. Encourage slow learners to talk by directing questions to them.

Key Assessment

DoK level 2: Skills of conceptual understanding

1. The number of molecules of ammonia gas is 12.04 x 10^{23}. Calculate the number of moles of ammonia gas. [L = 6.02 x 10^{23}]
2. Calculate the number of oxygen molecules in 0.5 mol of oxygen gas. [L = 6.02 x 10^{23}]
3. Calculate the number of atoms in 16 g of copper, Cu. [Cu = 63.5, L = 6.02 x 10^{23}]

Reference

1. SHS Chemistry Curriculum. LI1.1.2.1.1
2. RoseMarie Gallagher & Paul Ingram. (2014). Complete Chemistry for Cambridge IGCSE. 3rd ed. Oxford University press, Great Clarendon Street, Oxford OX2 6DP
3. Addison – Wesley. (2000). Chemistry. 5th ed. Prentice Hall, New York
4. Lawrie Ryan &Rogger Norris. (2014). Cambridge International AS and A level Chemistry Coursebook. University Printing House, Cambridge CB2 8B5
5. Christopher T. et al. (2010). Chemistry for the IB Diploma. Hodder Education Hachette UK company, 338 Euston Road London NW1 3BH
6. Catrin Brown & Mike Ford. (2020) Chemistry for IB Diploma. Edinburg Gate, Harlow, Essex, CM 20 2JE

7. Chang Reymond. (2017). General Chemistry – Essential Concepts. The McGraw Hill Companies

8. https://docbrown.info/page04/4_73calcs01ram.htm

9. https://docbrown.info/page04/4_73calcs09mvg.htm

Week 7

Learning Indicator(s):

1. *Calculate different physical quantities (number of entities, mass and volume) based on the amount of substance.*
2. *Explain the mole concept and its relevance in preparation of standard solutions.*

Theme or Focal Area (s): Calculating the Number of Entities

Recall that,

$$\text{Amount of substance} = \frac{\text{Number of entities}}{\text{Avogadro's constant}}$$

$$= \frac{N}{N_A} \text{ or } n = \frac{N}{L}$$

To calculate for the number of entities, multiply both sides of the equation by L

$$N = n \times L$$

Number of Entities = number of moles of substance x Avogadro's Constant

Example:

Calculate the number of atoms contained in 0.25 mol of sodium. [$L = 6.02 \times 10^{23}$]

Solution (Use problem solving strategy)

Analyse the question

Known

$n = 0.25$ mol

$L = 6.02 \times 10^{23}$

Unknown

$N = ?$

Use the formula: $N = n \times L$

Solve: Apply the formula

By definition,

$N = n \times L$

$N = 0.25 \text{ mol} \times 6.02 \times 10^{23}$

$N = 1.51 \times 10^{23}$ atoms

Evaluate: Check to see if the answer makes sense and if the correct unit is stated.

Moles in Mass of Atoms or Molecules

The Molar mass (M) of a substance, is the relative atomic mass (A_r) or relative molecular mass (M_r) expressed in grams per mole e.g. the M(H_2) is 2 $gmol^{-1}$, or the M($CaCO_3$) is 100$gmol^{-1}$

How do you determine the mass of a given number of moles of a substance?

Mass of 1 mole of atom X = 1 × M(X)

Mass of 1 mole of O = 1 × 16 = 16g

Mass of 2 moles of O = 2 × 16 = 32g

Mass of n moles of O = **n** × **M** = **m** g

m = n × M

$$Amount\ of\ substance\ (n) = \frac{mass\ of\ substance\ (m)}{Molar\ mass\ (M)}$$

Example:

Calculate the number of moles contained in 20g of Aluminium atoms [Al = 27]

Solution (Use problem-solving strategy)

Analyse the question

Known

Mass (m) = 20g

Relative atomic mass (A_r) = 27

= Molar mass = 27 $gmol^{-1}$

Formula to use:

$$Amount\ of\ substance\ (n) = \frac{mass\ of\ substance\ (m)}{Molar\ mass\ (M)}$$

Unknown

Number of moles (n)

Solve:

Apply the formula

By definition,

Substituting the values,

$$Amount\ of\ substance\ (n) = \frac{20\ g}{27\ gmol^{-1}}$$

$$= 0.74\ mol$$

Therefore, the number of moles of aluminium atoms is 0.74 mol

Evaluate: Check to see if it makes sense

Calculating for the mass of a given amount of substance

Recall that, $Amount\ of\ substance\ (n) = \frac{mass\ of\ substance\ (m)}{Molar\ mass\ (M)}$

Making m the subject yields, $m = n \times M$

Example:

Calculate the mass of 0.5 mol of water, H_2O [H = 1, O = 16]

Solution: Use the problem-solving approach
Analyse the question

Known
Number of moles (n) = 0.5 mol
Relative Atomic masses [H = 1, O = 16]
Formula to use: $m = n \times M$

Unknown
Mass (m) =?

Solve: Apply the strategy
Calculate the M
M (H_2O) = 2(1) + 16 = 18gmol^{-1}

Calculate the mass
By definition, $m = n \times M$

$m = 0.5 \text{ x } 18$

$m = 9\text{g}$

Therefore, the mass of water is 9g.

Evaluate: Check the answer to see if the correct unit is stated and if the answer makes sense.

Quantity of substance and molar volume of gases

The volume occupied by a gas depends on:

- **i.** Quantity of substance
- **ii.** Temperature
- **iii.** Pressure of the gas

At standard temperature of 273 K and pressure of 101.3 kPa (known as standard temperature and pressure, or s.t.p.), the volume occupied by one mole of any gas is called molar volume, denoted by Vm. Vm is a constant and has a value of 22.4 dm^{3}mol^{-1} at s.t.p.

The Molar Volume **Vm**, the number of moles of substance **n** and volume of gas are related by the formula;

$$\textit{Amount of substance } (n) = \frac{\textit{Volume of substance in } dm^3 (V)}{\text{Molar volume } (V_m)}$$

Multiplying both sides of the equation by Vm gives

$$V = n \times V_m$$

NB:

- **i.** This equation is used to calculate the volume of a gas at s.t.p., given the quantity of substance or number of moles.
- **ii.** If the gas volume is measured in cm^3, convert to dm^3.

iii. You can calculate the volume of a named gas, given the formula and relative atomic masses of the elements.

Example:

Calculate the volume occupied by 0.75 mol of ammonia gas (NH_3) at s.t.p.
[$Vm = 22.4\ dm^3mol^{-1}$]

Solution: Use the problem-solving approach

Analyse the question
Known
Mole (n) = 0.75 mol
$Vm = 22.4\ dm^3mol^{-1}$
Formula to use: V = n x Vm

Unknown
Volume of gas, V =?

Solve: Apply the problem-solving approach

By definition,

$V = n \times Vm$

Substituting the values,

$V = 0.75\ mol \times 22.4\ dm^3mol^{-1}$

$V = 16.8\ dm^3$

Evaluate: Check the answer to see if it makes sense.

The concept map of the relationship between mole and other variables are:

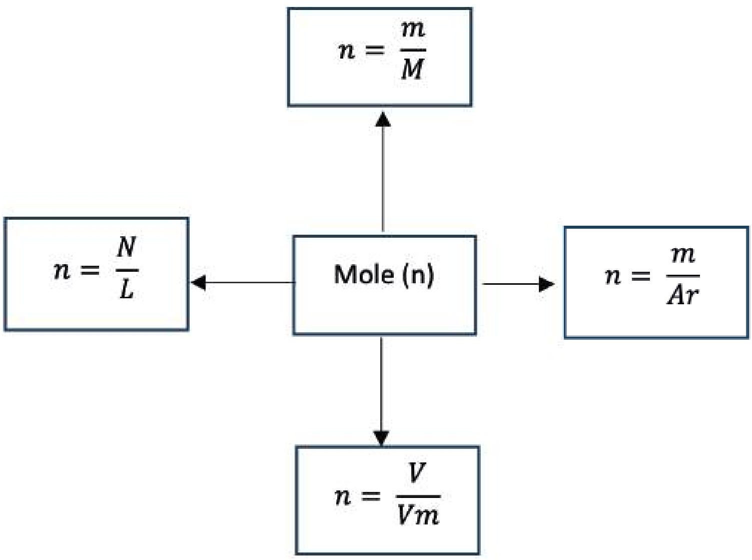


Figure 7.0: *A Concept Map Showing the Relationship between Mole and other Variables*

Learning Task

1. Calculate the number of oxide ions contained in 0.5 mol of Al_2O_3 [$N_A = 6.02 \times 10^{23}$]
2. Calculate the number of moles contained in 17.5g of P_4O_{10}? [P=31, O=16]
3. Calculate the number of moles present in 2.5 g of Na_2SO_4.
4. Calculate the mass of 0.25mol of Na_2SO_4
5. Calculate the number of sodium ions present in 2.5 g of Na_2SO_4.
   [Na = 23, S = 32, O=16, L = 6.02×10^{23} mol^{-1}]
6. Calculate the number of moles contained in 250 cm^3 of carbon dioxide gas at s.t.p.
7. Calculate the volume occupied by 0.25 moles of carbon dioxide gas at s.t.p.
8. Calculate the number of carbon dioxide molecules present in 500 cm^3 of the gas at s.t.p.
   [L = 6.02×10^{23} mol^{-1}, Vm = 22.4dm^3mol^{-1}]

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

Problem-based approach:

1. Describe how the amount of substance (n) can be used to determine the number of entities (atoms, molecules or ions), mass(m) of a substance, volume (V) of a gas using mathematical equations that represent their interconversions.

a. Provide clear definitions and explanations of key terms, such as Avogadro's number and molar mass.

b. Use gradual instructions to guide students through simple mole calculations

c. Support learners who are struggling with scaffolded worksheet.

2. Practice calculations involving amount of substance, number of entities and molar quantities.

a. Give more opportunities for learners to practice and review to help reinforce the basic concepts.

Key Assessment

DoK level 2 (Skills of conceptual understanding)

Calculate the number of sodium ions in 16g of Na_2CO_3. [Na = 23, O = 16, C = 12, $N_A = 6.02 \times 10^{23}$ particles/mol]

DoK level 2 (Skills of conceptual understanding)

The molar mass of CO_2 is 44 g/mol. How many moles of CO_2 are present in 124 g sample of CO_2.

DoK level 2 (Skills of conceptual understanding)

What is the mass of 5.0×10^{23} molecules of NO_2?

Theme or Focal Area(s): The mole concepts and their relevance in preparation of standard solution

A solution is a uniform mixture of a solute and a solvent. The quantity of solute per unit volume of solution is termed as concentration. To be able to compare the concentration of solutions, we use standard units.

Types of concentration

1. **Quantity of substance concentration (Molarity):** It is defined as the quantity of solute (number of moles of solute) dissolved in one cubic decimetre of the solution.

   Mathematically, it is expressed as:

   $$\text{Concentration in mol } dm^{-3}(C) = \frac{\text{amount ofs ubstance in moles } (n)}{\text{Volume of solution in } dm^3 \ (V)}$$

   The number of moles of solute n can be made the subject as

   $$n = C \times V$$

NB:

i. This equation can be used to calculate the number of moles of solute required for a given volume of a specified concentration.
ii. The volume of solution required can be calculated given the number of moles and a specified concentration.
iii. If the volume of solution is given in cm^3, it must be converted to dm^3 by dividing it by 1000.

Example:
Calculate the number of moles in 250 cm^3 of 0.500 mol dm^{-3} sulfuric acid solution

Solution: Use the problem-solving approach

Analyse the question

Known
Volume, V = 250 cm^3
Concentration, C = 0.500 mol dm^{-3}
Formula to use: n = C x V

Unknown
Number of moles, n =?

Solve: Apply this formula
Convert the volume to dm^3 by dividing it by 1000.
V = 250/1000 = 0.250 dm^3

By definition,
n = C × V

Substituting the values,
n = 0.500 mol dm^{-3} x 0.250 dm^3
n = 0.125 mol

Evaluate: Check the answer to see if it makes sense.

2. **Mass Concentration (Concentration in g dm^{-3}):** It is defined as the mass of solute dissolved in one cubic decimetre of the solution. It is denoted by ρ(not to be confused with density).

   Mathematically, it is defined as:

   $$\text{Concentration in g } dm^{-3}(\rho) = \frac{\textit{mass of solute in grams } (m)}{\text{Volume of solution in } dm^3 \ (V)}$$

   The mass of the solute m can be made the subject as follows:

$$m = \rho \times V$$

NB: This equation can be used to calculate:

i. The mass of solute in a given volume of a solution of a specified concentration.
ii. The volume of solution needed when the mass of solute and specified concentration is given.

Mass of solute required to prepare a given volume of a standard solution

Recall that,

$C = \frac{n}{V}$ ....(1)

but $n = \frac{m}{M}$ .... (2)

Substituting (2) into (1) gives $C = \frac{m}{M \times V}$

$$\text{Hence } m = C \times M \times V$$

This equation is used to calculate the mass of solute required to prepare a given volume of a standard solution.

Example:
Calculate the mass of NaOH required to prepare 250cm^3 of 0.50mol dm^{-3} sodium hydroxide solution. [Na = 23, H = 1, O = 16]

Solution: Use the problem-solving approach

Analyse the question

Known
Volume of solution, V = 250 cm^3

Concentration, C = 0.50 mol dm^{-3}

Relative atomic masses: Na = 23, H = 1, O = 16

Formula to use: $m = C \times M \times V$

Unknown
Mass, m =?

Solve: Apply the formula

i. Calculate the M (NaOH) = 23 + 16 + 1 = 40 g mol^{-1}

ii. Convert the volume to dm^3 by dividing it by 1000.

V = 250/1000 = 0.250 dm^3

iii. Substituting the values into the formula,

$m = 0.50 \times 40 \times 0.25 = 5\text{g}$

Therefore, the mass of NaOH required is 5g.
(c) Evaluate: Check the answer to see if it makes sense.

Relationship that exists between molar concentration and mass concentration
Consider the equation:
$\text{C} = \frac{m}{M \times V}$

Also recall that, $\rho = \frac{m}{V}$. Combining the two equations yields, $C = \frac{\rho}{M}$

Where,

c = Concentration in mol dm^{-3}

ρ = mass concentration in g dm^{-3}

M = Molar mass in $gmol^{-1}$

Preparation of Standard Solution

A standard solution is a solution whose concentration is accurately known.

Primary Standard

A primary standard is a substance that is usually available in pure form or a state of known purity, which is used in preparing a standard solution. Examples are sodium carbonate and potassium iodate

Properties of a primary standard

It should be available in pure form or easily purified.

It must be stable, that is, it must not lose weight or take up water during weighing.

It must have a reasonably high relative formula mass.

It must react speedily without side reactions with the substance being standardised.

It should have high solubility

How to prepare a standard solution from a solid solute

1. Determine the mass of the solute required to make the appropriate concentration and volume of desired solution.
2. Weigh accurately the solute in a beaker.
3. Add distilled water to the beaker and its contents and swirl to dissolve the solid.

   NB: The beaker must have a lower volume than the standard volumetric flask being used.
4. Transfer the solution to the required standard volumetric flask through a funnel.
5. Rinse the stirrer and the beaker used into the flask, then add more distilled water until the meniscus lies on the calibration mark.
6. Invert the stoppered flask a few times to mix.
7. Label the solution.

Preparation of standard solution from concentrated solution

1. Use the dilution formula ($C_1V_1=C_2V_2$) to calculate the volume of the concentrated solution required.
2. Pour some distilled water into the required standard volumetric flask.
3. Measure the stock or concentrated solution and transfer it into the distilled water in the volumetric flask.
4. Swirl the flask and its content and top the solution to the calibration mark with distilled water.
5. Label the solution.

Determination of the concentration of a stock solution

The commercial stock solution usually contains chemical assay (that is the label on their container, specifying the purity, density, molecular mass, and other relevant information).

Calculate the mass of the substance in 1 dm^3

Calculate the mass of the pure substance in 1dm^3 by multiplying by the percentage purity.

Divide this mass by the molar mass to get the concentration.

Mathematically use the formula:

$$\frac{Concentration\ (C) = Density\ \rho \times 1000 \times percentage\ purity\ (\%)}{Molar\ mass\ (M) \times 100}$$

Learning Tasks:

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

1. Calculate the number of moles in 2g of Sodium hydroxide (NaOH).
2. Calculate the concentration of sodium hydroxide (NaOH) solution prepared by dissolving 2g in 250 cm^3 of solution.
3. Calculate the volume of distilled water required to prepare 0.05 mol dm^{-3} of solution from 2.5g of Na_2CO_3. [Na = 23, O=16, C = 12, H = 1]
4. Calculate the mass concentration of a solution prepared by dissolving 2.5g of NaOH in 250 cm^3 of distilled water.
5. Calculate the mass of substance contained in 150 cm^3 of 0.5 mol dm^{-3} sodium hydroxide solution.
6. Prepare a 250cm^3 solution of NaOH of concentration 2.0 $moldm^{-3}$
7. The label on a stock solution of HCl sold in a chemical store contains the following information:

   Density = 1.19 g cm^{-3}

   Percentage purity = 37%

   Molar mass of HCl = 36.5 g mol^{-1}

   Using the information given, prepare a 250 cm^3 solution of HCl of concentration 2.0 $moldm^{-3}$

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

Think-Pair Share

a. Recap on how to use mass and relative molecular mass to determine the number of moles of a given substance.

b. Particular attention should be given to learners who are struggling to get the concept.

Initiating Talk-for-Learning: Guide learners to

a. Brainstorm on the need to know the concentration of solutions.

b. Encourage all learners to share their thoughts on the importance of knowing the concentration of solutions in real life.

Experiential Learning:

a. Identify the apparatus for preparing standard solutions in various concentrations.

b. In small groups, apply the mole concept to prepare standard solutions in $moldm^{-3}$ and gdm^{-3}.

c. Provide learners with a named sample of solid and liquid solutes and guide them to use the various apparatus to prepare a standard solution.

d. Emphasis should be laid on the importance of following the steps in preparing standard solutions.

Key Assessment

DoK level 1:(Reproduction/Recall)

List five (5) apparatus used in preparing a standard solution from solid solutes

DoK level 4 (Extended critical thinking and reasoning)

Work in small groups to prepare 250 cm^3 of 2.0 mol dm^{-3} solution of sodium hydroxide in the laboratory.

[H=1.0, O=16.0, Na=23.0]

Reference

1. SHS Chemistry Curriculum LI1.1.2.1.2, LI1.1.2.1.3 and LI1.1.2.1.4
2. Addison – Wesley. (2000). Chemistry. 5th ed. Prentice Hall, New York
3. Lawrie Ryan & Roger Norris. (2014). Cambridge International AS and AA-level Chemistry Coursebook. University Printing House, Cambridge CB2 8B5
4. Christopher T. et al. (2010). Chemistry for the IB Diploma. Hodder Education Hachette UK company, 338 Euston Road London NW1 3BH
5. Catrin Brown & Mike Ford. (2020) Chemistry for IB Diploma. Edinburg Gate, Harlow, Essex, CM 20 2JE
6. https://docbrown.info/page04/4_73calcs09mvg.htm
7. https://www.khanacademy.org/science/ap-chemistry-beta/x2eef969c74e0d802:atomic-structure-and-properties
8. Chang Reymond. (2017). General Chemistry – Essential Concepts. The McGraw Hill Companies

SECTION 3: MOLE RATIOS, CHEMICAL FORMULAE AND CHEMICAL EQUATIONS

Strand: **Physical Chemistry**

Sub-Strand: Matter and its properties

Content Standard: Demonstrate an understanding of the mole concept and its significance to the quantitative analysis of chemical reactions.

Learning Outcome: *Write mole ratios for chemical equations and apply them in quantitative analysis.*

INTRODUCTION AND SECTION SUMMARY

This section emphasizes mastering essential skills including IUPAC nomenclature for inorganic compounds, writing compound formulas based on chemical laws, and balancing equations. Learners will also engage in stoichiometric calculations to deepen their understanding of chemical reactions.

SUMMARY OF PEDAGOGICAL EXEMPLARS

To enhance engagement and understanding, a variety of pedagogical approaches such as group presentations, problem-based learning, collaborative learning, talk-for-learning, inquiry-based learning, and experiential learning will be utilized. These methods aim to foster critical thinking, teamwork, and hands-on learning experiences.

ASSESSMENT SUMMARY

Tailored tasks using various resources like calculators, chemical balances, and worksheets will be implemented to ensure learning outcomes are achieved. Assessment tasks with different levels of depth of knowledge will be employed at the end of each focal area to evaluate pedagogical effectiveness. This comprehensive approach aims to provide learners with a strong foundation in inorganic chemistry principles and practical problem-solving skills.

WEEK 8

Learning Indicator(s): *Use IUPAC nomenclature to name inorganic compounds, write the formulae of compounds based on the laws of chemical combination and write balanced chemical equations.*

Theme or Focal Area(s): Use IUPAC nomenclature to name inorganic compounds, write the formulae of the compounds based on the laws of chemical combination and write a balanced chemical equation

The nomenclature of inorganic compounds is based on the oxidation number system.

Oxidation number of an atom is the number of electrons gained or lost by an atom when forming a compound.

Rules for assigning oxidation number

1. The oxidation number of elements in their elemental state is zero.
2. The oxidation state of oxygen in most compounds is -2 except in a peroxide (-1) and superoxide (-½).
3. The oxidation state of hydrogen is (+1). when it is bonded to a non-metal and (-1) when bound to a metal.
4. The oxidation state of an ion is equal to the charge on the ion.
5. For a neutral molecule or polyatomic ion, the sum of the oxidation numbers of all the atoms must be equal to the total charge on it.

Naming Binary Ionic compounds

Rule

1. Name the cation first followed by the name of the anion.
2. The name of the cation is the name of the metal.
3. For metals with atomic numbers above 20, indicate the oxidation state in Roman numerals and brackets.
4. The anions are named by replacing the suffix with 'ide'

Example: $NaCl$ – Sodium chloride
$FeCl_3$ – Iron(III) chloride

Naming simple acids

Rule:

1. Use the prefix 'hydro', then the root name of the central atom
2. Add the suffix 'ic' to the root name
3. Add the word 'acid'

Example: HCl – Hydrochloric acid
HI – Hydroiodic acid

Naming oxoacids

Rule:

1. Use the prefixes oxo, dioxo, trioxo and tetraoxo to indicate the number of oxygen atoms present.
2. Add the root name of the central atom.

3. Add the suffix 'ate' followed by the oxidation state of the central atom in Roman numerals and brackets.
4. Add the word acid.

Example: H_2CO_3 – trioxocarbonate (IV) acid
H_2SO_4 – tetraoxosulphate (VI) acid

Naming Acid Salts

Rule:

Name the metal cation first followed by the name of the oxosalt.

1. If the cation has a relative atomic number above 20, its oxidation state should be indicated in Roman numerals and brackets.
2. Add the word 'hydrogen'.
3. Name the oxoanion as usual without the word 'ion'.

Example: $NaHSO_4$ – Sodium hydrogen tetraoxosulphate (VI)

Naming Simple non-ionic compounds

Rule:

1. Name the electropositive element first.
2. Add the root name of the anion.
3. Add the suffix 'ide'.

Example: HCl – hydrogen chloride
SiC – Silicon carbide

Molecular compounds where a pair of elements form different compounds with different number of oxygen atoms

Rule:

1. Name the electropositive element first.
2. Indicate its oxidation state in Roman numerals and in brackets.
3. Name the electronegative element as usual.

Example: CO_2 – Carbon (IV) oxide
N_2O_3 – Nitrogen (III) oxide

Naming Hydrated salt

Rule:

1. Name the anhydrous part first.
2. Use the prefixes mono, di, tri, tetra, penta, hexa, hepta, octa, etc. to indicate the moles of water of crystallisation.
3. Add the suffix hydrate to it.

Example: $CuSO_4.5H_2O$ – Copper (II) tetraoxosulphate(VI) pentahydrate

Learning Task

1. What is the IUPAC name of the following compounds?
   - a. $AlBr_3$
   - b. $CrCl_3$
   - c. HBr
   - d. H_3PO_3
   - e. $NaHCO_3$
   - f. Al_4C_3
   - g. Na_3N
   - h. SO_2
   - i. $MgSO_4.7H_2O$

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

Group Presentations: In small groups, research and give class presentations on the IUPAC rules for naming various groups of inorganic compounds.

- **a.** Support learners to touch on the core principles of IUPAC nomenclature, emphasising the importance of systematic naming in chemistry.
- **b.** Assist learners to make visually appealing diagrams or infographics that highlight the main ideas of the IUPAC naming guidelines.

Apply the IUPAC rules to name binary and ternary compounds, oxoacids, salts and hydrated salts

- **a.** Distribute compounds formula cards to learners in their groups and ask them to use the rules to name the various compounds.
- **b.** Design an interactive activity or games that simulate the process of naming binary and ternary compounds.
- **c.** Move round the various groups to inspect progress of work and give assistance when necessary.

Design a chart on the IUPAC rules for naming inorganic compounds to be posted in the classroom

Key Assessment

DoK level 2:(Skills of conceptual understanding)

Write the chemical formulae of the following compounds

1. Potassium chloride
2. Iron (II) bromide
3. Copper (II) tetraoxosulphate(VI)
4. Sodium carbonate decahydrate

Theme or Focal Area(s): Chemical Formulae

Chemical Formulae: It is an expression which shows the chemical composition of a compound in terms of the symbols of the atoms involved.

Types of Chemical Formulae

There are three (3) types of chemical formulae, these are;

1. Empirical formula
2. Molecular formula
3. Structural formula

Empirical formula: The empirical formula of a compound is the simplest whole number ratio of atoms present in a compound.

Determination of the empirical formula

1. State the mass or percentage mass of each element.
2. Divide each mass by their relative atomic mass.
3. Determine the simplest whole number ratio of each element by dividing the results by the least ratio number or scale up by X factor to get the simplest whole number ratio.
4. The simplest integer ratio is then used to write the empirical formula as a right-hand subscript.

Molecular formula: It is the formula that shows the actual number of atoms of each element in the simplest unit of a substance.

NB: Some compounds have their empirical formula being the same as the molecular formula.

Molecular formula is derived from the empirical formula using the following relationship;

$$Molecular\ mass = (Empirical\ mass)_n$$

The integer **n** obtained is used to multiply each element of the empirical formula to get the molecular formula.

Therefore

$$Molecular\ formula = (Empirical\ formula)_n$$

Example:

In an experiment to determine the empirical formula of lead sulphide, the following results were obtained:

Mass of lead = 207 g

Mass of sulphur = 32 g

Calculate the empirical formula of lead sulphide [Pb = 207, S = 32]

Solution:	Pb	S
Mass of element:	207 g	32 g
Mole of element, $\left(\frac{m}{Ar}\right)$	$\frac{207}{207}$	$\frac{32}{32}$
Simplest ratio	1	1

Empirical formula: PbS

Structural Formula: It is the formula that shows how the atoms in the molecule or compound are bonded to each other.

Example: Ethane, with molecular formula, C_2H_6 has a structure formula:

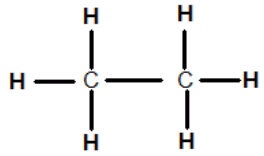


Fig. 3.1

Percentage composition of elements in a compound

The method of calculating the percentage by mass composition of a compound in terms of its constituent element is as follows:

1. Calculate the molecular mass of the compound.
2. Calculate the mass of the specified elements in the compound considering the number of atoms of each element in the formula.

$$\textit{Percentage of element in a compound} = \frac{\textit{Relative atomic mass} \times \textit{Number of atoms of elements} \times 100}{\textit{Relative molecular mass}}$$

Example: Calculate the percentage by mass of nitrogen and hydrogen in NH_3.

Solution:

By definition,

$$\textit{Percentage of element in a compound} = \frac{\textit{Relative atomic mass} \times \textit{Number of atoms of elements} \times 100}{\textit{Relative molecular mass}}$$

$$\%\ \textit{of nitrogen} = \frac{\textit{mass of nitrogen}}{\textit{Molar mass of ammonia}} \times 100$$

$$\%\ \textit{of nitrogen} = \frac{14}{17} \times 100$$

$$= 82.35\ \%$$

$$\%\ \text{of hydrogen} = \frac{\textit{mass of hydrogen}}{\textit{Molar mass of ammonia}} \times 100$$

$$\%\ \text{of hydrogen} = \frac{3}{17} \times 100$$

$$= 17.65\ \%$$

Learning Task

1. A chloride of copper has the following percentage mass 47.4% copper and 52.6% chlorine. Determine the empirical formula. [Cu=64, Cl=35.5]
2. Determine the percentage by mass of carbon and hydrogen in C_2H_6. [C=12, H=1]

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

Problem-based learning: In a whole class activity,

1. Discuss the laws that are applied in writing or determining chemical formulae of various compounds.

   a. Divide the steps of writing chemical formula into manageable steps and guide learners using scaffolding.
   b. Start with simple examples and increase the complexity as learners gain confidence in the skills of writing chemical formula.
   c. Encourage peer tutoring, where learners who are more confident in writing chemical formulae help their peers who may be struggling.

2. Calculate the empirical and molecular formulae of various inorganic compounds.

   a. Initially guide learners through the steps in calculating empirical and molecular formulae of compounds.
   b. Provide a guided worksheet with the various steps and allow learners to compute the empirical and molecular formulae of sample questions.
   c. As learners gain mastery over the process, encourage them to work without the visual aid of the steps.

3. Write the chemical formulae for named binary and ternary compounds, oxoacids, oxosalts and hydrated salts.

   Determine the percentage composition of elements in various compounds based on their formulae. E.g., MgO, Cu_2O and CuO.

Key Assessment

DoK Level 3: (Strategic reasoning)

An organic compound of relative molecular mass 46 on analysis was found to contain 52.0% carbon, 13.3% hydrogen and the remaining being oxygen. Determine its

i. Empirical formula

ii. Molecular formula

[H=1.0, C=12.0, O=16.0]

Theme or Focal Area: Laws of chemical combination

1. The law of conservation of mass

It states that mass is not created or destroyed in a chemical reaction, i.e. the total mass of the products made is equal to the total mass of the reactants.

2. The law of definite proportion

It says that the proportion by amounts of each element in a pure compound is always the same, no matter how the compound is prepared.

3. The law of multiple proportion

When two elements combine to form more than one compound, the different masses of one element that combine with a fixed mass of the other element are in a ratio of small whole numbers.

It is defined as an expression which uses chemical symbols in formula to represent the elements and compounds which occurred in a chemical reaction.

Types of chemical equations

a. Combustion

It is a chemical reaction in which a substance reacts with excess or limited oxygen to give oxides of the components of the substance.

Example:

$CH_4 + 2O_2 \rightarrow CO_2 + H_2O$

b. Synthesis

It is a type of reaction in which two or more simple substances combine to form a more complex compound. That is, the reactants combine to form a single product.

It could be illustrated as:

$A + B \rightarrow AB$

Example:

$2H_2 + O_2 \rightarrow 2H_2O$

Application in everyday life

It is widely used in chemistry for the formation of salts, organic compounds, biomolecules, medicines, pesticides and polymers.

c. Displacement reaction

It is a reaction in which one atom or ion in a reactant is replaced by another atom or ion of another element.

Example:

$Mg + CuSO_4 \rightarrow MgSO_4 + Cu$

Application in everyday life

- Displacement reactions are essential in various chemical processes
- They have practical application in metallurgy, electrochemistry, extraction of metals such as gold from their ores.
- They help to predict reactivity.

d. Decomposition

It is a type of chemical reaction in which a compound breaks down into two or more simpler substances under certain conditions. That is, a simple reactant undergoes a chemical change to produce multiple products.

It is illustrated as:

$AB \rightarrow A + B$

Where AB is the initial reactant and A and B are the products.

The condition under which the reaction occurs could be heat, light and use of catalyst.

Example:

$CaCO_3 \underset{\Delta}{\rightarrow} CaO + CO_2$

Application in everyday

- It is important in natural and industrial processes.
- They are essential in the fields of chemistry, biology and environmental science.

e. Ionic equation

It is a chemical equation involving at least one ionic species as a reactant or product, that is, species of dissolved ionic compounds in terms of their free ions; ions that exist in a chemical equation, but are not involved in the overall equation, are **spectator ions**.

Example:

Write the net ionic equation of the reaction,

$$2KI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow PbI_{2(s)} + 2KNO_{3(aq)}$$

Solution:

Write the ionic equation

$$2K^+_{(aq)} + 2I^-_{(aq)} + Pb^{2+}_{(aq)} + 2NO_{3(aq)} \rightarrow PbI_{2(s)} + 2K^+_{(aq)} + 2NO^-_{3(aq)}$$

Cancel the spectator ions to yield the net ionic equation

$$\cancel{2K^+_{(aq)}} + 2I^-_{(aq)} + Pb^{2+}_{(aq)} + \cancel{2NO_{3(aq)}} \rightarrow PbI_{2(s)} + \cancel{2K^+_{(aq)}} + \cancel{2NO^-_{3(aq)}}$$

Write the net ionic equation,

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

Learning Task:

1. Guide learners to mention the laws of chemical combination.
2. Assist learners to state the laws of chemical combination.
3. Using the reaction between zinc granules and hydrochloric acid, explain the law of conservation of mass.
4. Balance the following chemical equations
   a. $N_2 + O_2 \rightarrow N_2O_5$
   b. $MgCl_2 + K_3PO_4 \rightarrow Mg_3(PO_4)_2 + KCl$
   c. $C_2H_6O + O_2 \rightarrow CO_2 + H_2O$

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

Initiating talk-for- learning:

Using relevant examples, discuss the laws of chemical combination namely:

a. Law of conservation of matter
b. Law of constant proportion
c. Law of multiple proportion

Guide learners using simulations or charts to assist in their discussions.

Collaborative learning:

With the aid of a chart, discuss the rules to be followed in balancing chemical equations.

In pairs or groups, write and balance chemical equations for the following:

a. combustion
b. synthesis

c. displacement or replacement

d. decomposition

e. ionic equation

Inquiry-based learning: In pairs or groups, perform a simple experiment to show that mass is conserved in a chemical reaction.

Use the reaction between molar solutions of Na_2CO_3 and $CaCl_2$.

Key Assessment

DoK Level 2: (Skills of conceptual understanding)

Balance the following chemical equation and use it to answer the questions that follow:

$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$

1. How many moles of water would be obtained from 4 moles of C_2H_6?
2. How many moles of C_2H_6 would be needed to produce 3 moles of water?

DoK Level 2: (Skills of conceptual understanding)

State and explain the three laws of chemical combination.

DoK level 4: (Extended critical thinking and reasoning)

Design an experiment to show that mass is conserved in a chemical reaction.

Reference

1. SHS Chemistry Curriculum LI1.1.2.1.2B
2. Chang Reymond. (2017). General Chemistry – Essential Concepts. The McGraw Hill Companies
3. https://docbrown.info/page04/4_73calcs05emp.htm

Week 9

Learning Indicator: *Perform calculations involving stoichiometric relationships*

Theme or Focal Area (s): Stoichiometry is the relationship between quantities of reactants and products in a chemical reaction.

Mole ratio

The relative quantity of any two substances that take part in a chemical reaction is termed as **mole ratio**. The stoichiometric coefficients in the balanced equation are considered as the number of moles of each reactant or product.

Example: In the reaction,

$$2C \quad + \quad O_2 \quad \rightarrow \quad 2CO$$

$$2 \text{ mol} \quad + \quad 1 \text{ mol} \qquad 2 \text{ mol}$$

The mole ratio between carbon and oxygen is written as;

$$\frac{n(C)}{n(O_2)} = \frac{2}{1}$$

$$n = \frac{m}{M}$$

$$n = \frac{N}{L}$$

$$n = \frac{V}{V_m}$$

Using stoichiometric quantities to calculate numbers of entities involved in a chemical reaction

1. Write the mole ratio using the stoichiometric coefficient of the known substance and substance being calculated.
2. Calculate the number of moles of the substance being calculated.
3. Use the calculated moles and the relation N = n x L to calculate the number of entities.

Example:
Consider the equation,

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Calculate the molecules of ammonia gas produced if 3.01 × 10 23 molecules of hydrogen react with nitrogen gas

Solution:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

$N = 3.01 \times 10^{23}$, $n = ?$

Determine the number of moles of hydrogen

$$n = \frac{N}{L}$$

$$n = \frac{3.01 \times 10^{23}}{6.02 \times 10^{23}}$$

$$= 0.5 \text{ mol}$$

Write mole ratio between ammonia and hydrogen

$$\frac{n(NH_3)}{n(H_2)} = \frac{2}{3}$$

Calculate the number of moles of ammonia produced:

$$n(NH_3) = \frac{2}{3} \times n(H_2)$$

$$n(NH_3) = \frac{2}{3} \times 0.5$$

$$= 0.33 \text{ mol}$$

Solve for the number of molecules of product

$N = n \times L$

$N = 0.33 \times 6.02 \times 10^{23}$

$= 1.99 \times 10^{23}$ molecules

Calculating for the mass of substance

Procedure

1. Write the correct balanced equation.
2. Convert the quantity of the known substance into a number of moles using the correct mole formula

   $n = \frac{m}{M}$

   $n = \frac{N}{L}$

   $n = \frac{V}{V_m}$

3. Write the mole ratio using the stoichiometric coefficient of the known substance and substance being sought for.
4. Calculate the number of moles of the substance being sought for.
5. Use the calculated moles and the relation m = n x M to calculate the mass of substance.

Example:

Consider the reaction,

$N_2 + 3H_2 \rightarrow 2NH_3$

Calculate the mass of ammonia produced if 7 g of nitrogen reacts with excess hydrogen gas. [N = 14, H = 1]

Solution

Given the equation,

$N_2 + 3H_2 \rightarrow 2NH_3$

$m(N_2) = 7g$, $m(NH_3) = ?$

Determine moles of nitrogen

$n = \frac{m}{M}$

$n = \frac{7}{2 \times 14}$

$= 0.25$ mol

Determine the mole ratio

$$\frac{n(NH_3)}{n(N_2)} = \frac{2}{1}$$

$$n(NH_3) = 2 \times n(N_2)$$

$$n(NH_3) = 2 \times 0.25$$

$$= 0.5 \text{ mol}$$

$$M(NH_3) = 14 + 3(1)$$

$$= 17$$

$$m = 0.5 \text{ x } 17$$

$$= 8.5 \text{ g of ammonia}$$

Calculate the concentration of substance (analyte)

Procedure

1. Write the correct balanced equation.
2. Convert the quantity of the known substance into mole using the correct mole formula

   $n = \frac{m}{M}$

   $n = \frac{N}{L}$

   $n = \frac{V}{V_m}$
3. Write the mole ratio using the stoichiometric coefficient of the known substance and substance being sought for.
4. Calculate the number of moles of the substance being sought for.

   $C = \frac{n}{V}$

5) Use the calculated moles and the relation to calculate the concentration of the analyte. Example:

   Consider the reaction,

   $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

Example:

Given that 20 cm^3 of H_2SO_4 reacts completely with 25 cm^3 of 0.5 mol dm^{-3} NaOH, calculate the concentration of H_2SO_4.

Solution

$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$

$$C(NaOH) = 0.5 \text{ mol dm}^{-3}$$

$$V(NaOH) = 25 \text{ cm}^3$$

$$= 0.025 \text{ dm}^3$$

$$C(H_2SO_4) = ?$$

$$V(H_2SO_4) = 20 \text{cm}^3$$

$$= 0.020 \text{dm}^3$$

Determine the number of moles of sodium hydroxide solution using the formula, $n = c \times V$

$n = 0.025 \times 0.5$

$= 0.0125$ mol

Write mole ratio

$$\frac{n(H_2SO_4)}{n(NaOH)} = \frac{1}{2}$$

$$n(H_2SO_4) = \frac{1}{2} \times n(NaOH)$$

$$n(H_2SO_4) = \frac{1}{2} \times 0.0125$$

$$= 0.00625 \text{ mol}$$

Determine the concentration using the formula, $C = \frac{n}{V}$

$$C = \frac{0.00625}{0.020}$$

$$= 0.313 \text{ moldm}^{-3}$$

Calculating for the volume of substance

Procedure

1. Write the correct balanced equation.
2. Convert the quantity of the known substance into a number of moles using the correct mole formula

   $n = \frac{m}{M}$

   $n = \frac{N}{L}$

   $n = \frac{V}{V_m}$
3. Write the mole ratio using the stoichiometric coefficients of the known substance and substance being calculated.
4. Calculate the number of moles of the substance being sought for.
5. Use the calculated number of moles and the relationship $V = n \times V_m$ to calculate the volume of the gas.

Example:

10.5 g of methane reacts with excess oxygen to produce carbon dioxide and water. Calculate the volume of carbon dioxide gas produced at s.t.p.[C = 12, H = 1, O = 16Vm = 22.4 dm^3/mol].

Solution

Use the problem-solving strategy

1. Write the correct balanced equation.

   $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
2. Convert the quantity of the known substance into a number of moles using the correct mole formula

   $m(CH_4) = 10.5g,$

$$M(CH_4) = 12 + (1 \times 4)$$
$$= 16\ gmol^{-1}$$
$$n = \frac{m}{M}$$
$$= \frac{10.5}{16}$$
$$= 0.66\ mol$$

3. Write the mole ratio using the stoichiometric coefficients of the known substance and substance being calculated.

$$\frac{n(CO_2)}{n(CH_4)} = \frac{1}{1}$$

$$n(CO_2) = n\left(CH_4\right)$$

4. Calculate the number of moles of the substance being sought for.

$$n(CO_2) = n(CH_4) = 0.66\ mol$$
$$\therefore n(CO_2) = 0.66\ mol$$

5. Use the calculated number of moles and the relationship $V = n \times Vm$ to calculate the volume of the gas.

$$n(CO_2) = \frac{V}{V_m}$$
$$V = n(CO_2) \times V_m$$
$$= 0.66 \times 22.4$$
$$= 14.78\ dm3$$

Learning tasks

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

1) Balance the chemical equation $H_2 + O_2 \rightarrow H_2O$ and determine the mole ratio of hydrogen to oxygen in the reaction.

2) Balance the chemical equation $Al_{(s)} + Fe_2O_{3(aq)} \rightarrow Al_2O_{3(aq)} + Fe_{(s)}$ and determine the mole ratio of Fe_2O_3 to Fe in the reaction.

3) Consider the reaction, $HCl + NaOH \rightarrow NaCl + H_2O$. If $25cm^3$ of 0.25 mol dm^{-3} HCl reacts completely with excess sodium hydroxide, calculate the mass of sodium chloride produced. [Na = 23, Cl = 35.5]

4) Consider the reaction, $2KOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$. Calculate the volume of KOH of concentration 0.1 mol dm^{-3} required to completely neutralise 20 cm^3 of a 0.25 mol dm^{-3} H_2SO_4 solution.

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

Experiential Learning

a) Start the lesson by assessing learners' prior knowledge in basic chemistry concepts such as balancing of chemical equations from the JHS syllabus.

b) Put learners into ability groupings and offer various instructional materials such written instructions, reagents, hands-on demonstrations to the groups.

c) Guide learners in their ability groupings to carry out simple reactions base on the materials provided and have them follow strictly the written instructions.

d) Have the groups write down and balanced the equations they have carried out and present.

e) While maintaining the groups, let learners write out the mole ratio of the species in each of the balanced reactions and use them to calculate the following quantities. Provide support to groups that are struggling and allow for more independence for advance groups.

i. Number of entities
ii. Amount of substance
iii. Mass of substance
iv. Concentrations in g dm^{-3}, mol $dm^{-3}$and ppm
v. Volume of substance

Key Assessment

DoK Level 2: (Skills of conceptual understanding):

Write a balance chemical equation for the reaction between Zn and HCl

DoK Level 3: (Strategic reasoning):

Zinc metal reacts with hydrogen chloride according to the reaction $Zn + HCl \rightarrow ZnCl + H_2$. Calculate the mass of zinc require to produce 2 moles of hydrogen gas

DoK Level 3: (Strategic reasoning):

Calculate the mass of CO_2 produce when 2 moles of propane (C_3H_8) reacts with excess oxygen. $(C = 12, H = 1, O = 16)$

Theme or Focal Area(s): Limiting and excess reagents

In a reacting system involving two reactants with initial quantities given, or having information to determine their initial quantities, the reactant that is completely used up is called the **limiting reagent**. The reagent which is not completely used up, is called the **excess reagent**.

The maximum quantity of the products formed is determined by the limiting reagent.

Procedure

1. Calculate the initial quantity of each reactant in moles.

2. If the stoichiometric ratio of the reactants is 1:1 the reagent with lower value is the limiting reagent and the other is excess reagent.

3. If the stoichiometric ratio of the reactant is not 1:1 then write a mole ratio between the two reactants and solve for one of them.

4. Compare the calculated number of moles with the initial quantity of moles. If the calculated number of moles is greater than the initial amount, then it is the limiting reagent and if it is less, it is the excess reactant.

Example: Consider the reaction,

$N_2 + 3H_2 \rightarrow 2NH_3$

If 12.0 g of nitrogen and 8.0 g of hydrogen react in the formation of ammonia,

a. Determine the limiting reagent

b. Calculate the mass of ammonia (NH_3) produced.

[N = 14, H = 1]

Solution

$N_2 + 3H_2 \rightarrow 2NH_3$

$m(N_2)= 12g$, $m(H_2) = 8g$

a) Determine the initial moles of both reactants

$$n(N_2) = \frac{m}{M}$$
$$= \frac{12}{28}$$
$$= 0.43 \text{ mol}$$
$$n(H_2) = \frac{m}{M}$$
$$= \frac{8}{2}$$
$$= 4 \text{ mol}$$

b) Determine the limiting reagent

$$\frac{n(N_2)}{n(H_2)} = \frac{1}{3}$$
$$n(N_2) = \frac{1}{3} \times n(H_2)$$
$$n(N_2) = \frac{1}{3} \times 4$$
$$= 1.33 \text{ mol}$$

1.33 moles of N_2 are required to react with 4 moles of H_2, but we only have 0.43 moles, so N_2 is the limiting reactant.

c) Write the mole ratio between the limiting reagent and the product (NH_3)

$$\frac{n(NH_3)}{n(N_2)} = \frac{2}{1}$$
$$n(NH_3) = \frac{2}{1} \times n(N_2)$$
$$n(NH_3) = 2 \times 0.43$$
$$= 0.86 \text{ mol}$$

Mass of NH_3, $m = n \times M$

$$= 0.86 \text{ x } 17$$
$$= 14.62 \text{ g}$$

Learning Tasks

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

1. Given the reaction, $2H_2 + O_2 \rightarrow 2H_2O$, if 3 moles of hydrogen and 2 moles of oxygen reacted, determine the limiting reagent
2. For the reaction $4Fe + 3O_2 \rightarrow 2Fe_2O_3$, if 10 mole of iron and 8 moles of oxygen are available, determine the limiting reagent
3. Calculate the volume of chlorine required to react completely with 50 cm^3 of 1.0 $moldm^3$ sodium bromide (NaBr) solution. Identify the excess reagent in the reaction.

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

Experiential learning

a) Start the lesson by discussing the concept of chemical reactions and the importance of stoichiometry in predicting reactant consumption and product formation

b) Introduce the terms limiting and excess reagents and explain their significance in reactions.

c) Provide real-world applications to illustrate the concept of limiting and excess reagents

d) Put learners into ability groupings and provide worksheets with differentiated tasks and ensure that the tasks are appropriate for each group

e) To one group let learners' complete simple calculations to determine the limiting reagent and calculate mass of product formed. Provide support if learners are struggling.

f) To another group let learners solve problems involving stoichiometry to determine the limiting reagent, calculate the mass of product and identify the excess reactant.

g) Reconvene as a whole class and discuss the findings from the group tasks

Key Assessment

DoK Level 2: (Skills of conceptual understanding):

In a reaction, 20 grams of hydrogen gas (H_2) reacts with 10 grams of oxygen gas (O_2) to produce water (H_2O).

Determine which reactant is the limiting reagent and which is the excess reagent.

DoK Level 3: (Strategic reasoning):

In the combustion of propane (C_3H_8), 40 grams of propane reacts with 100 grams of oxygen gas (O_2).

1. Determine which reactant is the limiting reagent and which is the excess reagent.
2. Calculate the mass of carbon dioxide (CO_2) produced.
3. Calculate the mass of the excess reagent remaining after the reaction is complete.

DoK Level 3: (Strategic reasoning):

In the synthesis of ammonia (NH_3), 50 grams of nitrogen gas (N_2) reacts with 20 grams of hydrogen gas (H_2).

1. Determine which reactant is the limiting reagent and which is the excess reagent.
2. Calculate the mass of ammonia produced.
3. Calculate the volume of nitrogen gas consumed at STP (standard temperature and pressure), assuming the reaction goes to completion.

Theme or Focal Area: Percentage yield of the product of a reaction

In a chemical reaction, the calculated amount of a product is usually not obtained due to the following:

a. The reaction may be reversible.

b. Some reactants may undergo side reactions.

c. Some products cannot be separated or recovered from the mixture.

Actual yield

It is the amount of a product actually obtained from a chemical reaction in practice.

The percentage yield

The percentage yield of a reaction is the percentage of the product obtained compared to their theoretical maximum yield calculated from the balanced equation.

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

Example:

Magnesium metal reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas.

a) If 12 g of magnesium reacts with excess HCl, calculate the maximum theoretical mass of magnesium chloride formed.

b) If 42.0 g of purified anhydrous magnesium chloride was obtained, calculate the percentage yield. [Mg = 24, Cl = 35.5]

Solution

Use problem - solving approach

a) Analyse the question

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

$m(Mg) = 12$ g

$M(Mg) = 24$

$m(MgCl_2) = ?$

b) Determine the number of moles of Mg

$$n(Mg) = \frac{m}{M} = \frac{12}{24} = 0.5 \text{ mol}$$

c) Write mole ratio between the $MgCl_2$ and Mg

$$\frac{n(MgCl_2)}{n(Mg)} = \frac{1}{1}$$

$$n(MgCl_2) = n(Mg) = 0.5 \text{ mol}$$

d) Determine the theoretical mass of $MgCl_2$ formed

$m = n \times M.$

But, M (MgCl 2) = 24 + (2 x 35.5) = 95 gmol −1

m = 0.5 x 95

= 47.5 g

e) Determine the percentage yield.

$$\text{Percentage yield} = \frac{\text{Actual amount obtained}}{\text{maximum theoretical yield}} \times 100$$

$$= \frac{42}{47.5} \times 100$$

= 88.42 %

Learning Tasks

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

1. In a laboratory experiment, students aimed to produce 25 grams of copper sulphate ($CuSO_4$) according to the reaction: $Cu + H_2SO_4 \rightarrow CuSO_4 + H_2$.

   If they actually obtained 20 grams of copper sulphate, calculate the percentage yield of the reaction.

2. In a chemical synthesis, 30 grams of calcium carbonate ($CaCO_3$) reacts with excess hydrochloric acid (HCl) to produce calcium chloride ($CaCl_2$), carbon dioxide (CO_2), and water. If 22 grams of calcium chloride are obtained in the reaction, calculate the percentage yield.

3. In the preparation of aspirin ($C_9H_8O_4$), a student reacted 20 grams of salicylic acid ($C_7H_6O_3$) with excess acetic anhydride ($C_4H_6O_3$). The theoretical yield of aspirin is 24 grams. If the student obtained 18 grams of aspirin, calculate the percentage yield.

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

Problem solving approach:

a) In a whole inclusive class, provide learners with a step-by-step visual guide in the form of a flowchart. Break down the process into clear, sequential steps with concise explanations and illustrative diagrams.

b) Put learners into mix-ability groupings and offer structured worksheets with guided practice problems of increasing complexity. Begin with straightforward examples where the theoretical yield is provided, and learners only need to calculate the percentage yield. Gradually introduce more challenging problems where learners must determine both the actual and theoretical yields before calculating the percentage yield. Include hints and scaffolding for learners who may be struggling.

c) Engage learners in real-world scenarios or case studies where they must apply their knowledge of percentage yield calculations in practical contexts. Present them with authentic examples such as industrial chemical processes, pharmaceutical synthesis, or environmental chemistry.

Key Assessment

DoK Level 2:(Skills of conceptual understanding):

In a laboratory experiment, 25 grams of sodium chloride (NaCl) are reacted with excess silver nitrate ($AgNO_3$) to produce silver chloride (AgCl). If the theoretical yield of AgCl is 30 grams, and the actual yield obtained is 20 grams, calculate the percentage yield of the reaction.

DoK Level 2: (Skills of conceptual understanding):

In the synthesis of aspirin, a student reacts 20 grams of salicylic acid ($C_7H_6O_3$) with 25 grams of acetic anhydride ($C_4H_6O_3$). The theoretical yield of aspirin is 24 grams. If the student obtains 18 grams of aspirin,

1. State the actual yield of the reaction.
2. Based on the actual yield obtained in part (1), calculate the percentage yield of the reaction.

DoK Level 4:(Extended critical thinking and reasoning):

A chemical reaction between hydrogen gas (H_2) and nitrogen gas (N_2) produces ammonia (NH_3). If 15 grams of hydrogen gas reacts with excess nitrogen gas to produce 25 grams of ammonia,

1. State the actual yield of the reaction.
2. The theoretical yield of ammonia in the reaction described in part (1) is 30 grams. Calculate the percentage yield of the reaction based on the actual yield obtained.

References

1. SHS Chemistry Curriculum LI1.1.2.1.2B, LI1.1.2.1.2C, LI1.1.2.1.2D and LI1.1.2.1.2F
2. Gibert T. R. et al. (2009). Chemistry. 2nd ed. W. W. Norton & Company, Inc 500 Fifth Avenue, New York. 10110
3. Addison – Wesley. (2000). Chemistry. 5th ed. Prentice Hall, New York
4. Lawrie Ryan &Rogger Norris. (2014). Cambridge International AS and A level Chemistry Coursebook. University Printing House, Cambridge CB2 8B5
5. Christopher T. et al. (2010). Chemistry for the IB Diploma. Hodder Education Hachette UK company, 338 Euston Road London NW1 3BH
6. Catrin Brown & Mike Ford. (2020) Chemistry for IB Diploma. Edinburg Gate, Harlow, Essex, CM 20 2JE
7. RoseMarie Gallagher & Paul Ingram. (2014). Complete Chemistry for Cambridge IGCSE. 3rd ed. Oxford University press, Great Clarendon Street, Oxford OX2 6DP
8. https://docbrown.info/page04/4_73calcs06rmc.htm
9. https://www.khanacademy.org/science/ap-chemistry-beta/x2eef969c74e0d802:chemical-reactions/x2eef969c74e0d802:stoichiometry/a/stoichiometry
10. Chang Reymond. (2017). General Chemistry – Essential Concepts. The McGraw Hill Companies

SECTION 4: KINETIC THEORY AND THE 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 spans three weeks and covers foundational chemistry concepts. Learners will delve into the kinetic theory of matter, distinguishing properties of solids, liquids, and gases. They will also study Gas Laws, analyse related graphs, and apply Graham's law of diffusion/effusion and Dalton's law of partial pressure in calculations.

SUMMARY OF PEDAGOGICAL EXEMPLARS

To enhance learning, various strategies such as interactive discussions, problem-solving tasks, hands-on experiments, simulations, and collaborative group work will be employed. These approaches aim to actively engage learners and foster effective understanding of the material.

ASSESSMENT SUMMARY

Assessment will feature challenging tasks at different Depth of Knowledge levels to evaluate teaching effectiveness. By engaging with diverse instructional methods and rigorous assessments, learners will not only grasp chemistry fundamentals but also develop critical thinking, analytical skills, and a deep understanding of matter and gas behaviours.

Week 10

Learning Indicator: *Explain the kinetic theory of matter and apply it to distinguish between the properties of solids, liquids and gases*

Theme or Focal Area(s): Kinetic theory of matter

The kinetic theory of matter states that;

1) Matter is made of tiny particles which are in constant random motion.
2) Matter possesses kinetic energy due to the motion of the particles.
3) The difference between the different states of matter is due to the nature and extent of motion and the separation between the particles.

Solid State

1) Solids have a fixed shape and volume at a given temperature.
2) Particles of solids are closely packed in an orderly manner.
3) Solids have the greatest forces of attraction between their particles
4) Particles of solids undergo vibration about their mean position.
5) Increasing the temperature of solids causes faster vibration of particles.

Using the kinetic model to explain the properties of solids

1) Solids have the greatest density because the particles are closest together.
2) Solids have a fixed shape and volume because of the strong force of attraction between the particles.
3) Solids are difficult to compress because of lack of empty space between the particles.

The Liquid State

1) Liquids have fixed volumes that take the shape of the container at a given temperature.
2) Particles of liquids are close together and arranged randomly
3) The particles of liquids move rapidly in all directions.
4) The forces of attraction between the particles are stronger than that of gases, but lower than that of solids.
5) Increasing the temperature of liquids makes their particles move faster due to a gain in kinetic energy.

Using the kinetic model to explain the properties of liquids

1) Liquids have greater densities than gases because the particles are closer due to the attractive forces.
2) Liquids have fixed volume but take the shape of its container because of the increased particle attraction.
3) Liquids are not easily compressed because there is so little empty space between the particles.
4) Liquids expand on heating, but not like gases, but greater than solid, because the greater forces of attraction restricting expansion is greater in liquids than gases, but lesser in solids.

Learning Tasks

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

1. Provide learners with different materials such as water and ice. Ask them to observe and describe the properties of each state of matter (solids and liquids) based on their shape, volume, density. Guide learners to understand how the kinetic theory of matter explains the behaviour of particles in each state. For example, they can discuss how particles in a solid are closely packed and vibrate in fixed positions, while particles in a liquid move freely and rapidly.
2. Provide learners with ice cubes, a heat source, and a thermometer. Ask them to measure the temperature of the ice, observe what happens as it melts into water, and record the temperature changes. Guide learners to analyse their observations in terms of kinetic theory, explaining how heat energy causes particles to gain kinetic energy and transition from a solid to a liquid state.

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

Collaborative learning:

a. In mix-ability groupings, use visual aids such as diagrams, illustrations, and animations to introduce the kinetic theory of matter. Use simple representations to depict the arrangement and motion of particles in solids and liquids. For example, show how particles in a solid are tightly packed and vibrate in fixed positions, while particles in a liquid are more loosely arranged and can slide past each other. Use colourful graphics to make the concepts clearer to learners.

b. Still maintaining the groups, engage them with activities to investigate the properties of solids and liquids in more depth and present their findings. Encourage critical thinking by asking students to relate their observations to the principles of kinetic theory and draw conclusions about the behaviour of particles in solids and liquids.

Experiential learning: In ability groups, conduct hands-on demonstrations to illustrate the behaviour of particles in solids and liquids. Provide learners with materials such as marbles, beads, and water, and ask them to observe and manipulate the substances to simulate the behaviour of particles. For solids, demonstrate how the particles maintain their shape and volume but can be rearranged by external forces. For liquids, show how the particles can flow and take the shape of their container. Encourage learners to ask questions and make observations to deepen their understanding.

Key Assessment

DoK Level 2: (Skills of conceptual understanding):

Draw a simple diagram illustrating the arrangement and motion of particles in a solid and a liquid.

DoK Level 3:(Strategic reasoning):

Compare and contrast the behaviour of particles in solids and liquids according to the kinetic theory of matter.

DoK Level 4:(Extended critical thinking and reasoning):

In a laboratory experiment, students investigate the effect of temperature on the viscosity of a liquid. Describe how the kinetic theory of matter can be applied to explain the observed changes in viscosity as temperature increases.

Theme or Focal Area(s): The Gaseous State

1) Gases have no fixed shape or volume but fill a container.
2) Forces of attraction between the particles of gases are negligible.
3) Particles are so small that the actual volume of individual particles is negligible compared to the volume of the container.
4) Particles are widely spaced and scattered and undergo random and rapid motion.
5) The average kinetic energy of the gas particles is directly proportional to the absolute temperature of the particles.
6) The collision of the gas particles with the surface of the container causes the gas pressure.

Using the kinetic model to explain the properties of gases

1) Gases have very low densities because the particles will space out in the container.
2) Gases have no fixed volume and shape because of the negligible force of attraction.
3) Gases are easily compressed because of the empty space between the particles
4) Order of ease of compression: gas > liquid > solid
5) Gases exert pressure because of the collision of the particles with the walls of the container.

Learning tasks:

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

1. Provide students with a simple demonstration using a balloon filled with air. Ask them to observe and describe the behaviour of the gas particles inside the balloon. Then, provide them with coloured markers and ask them to draw and label diagrams illustrating gas particle movement in the balloon.
2. Assign learners to research and present on real-world applications of gases and provide a list of industries or technologies where gases play a crucial role.

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

Experiential learning:

a. Use computer simulations and modelling software (e.g. PhEt) to provide experiential learning opportunities for learners. Allow them to interact with virtual simulations of gas behaviour under different conditions and observe how changes in temperature, pressure, and volume affect gas particles. Encourage learners to make predictions, test hypotheses, and analyse simulation results to deepen their understanding of gas behaviour based on the kinetic theory.

b. Where simulations and modelling software are not available, learners should be shown videos of these experiments.

Exploratory learning: Organise field trips to facilities where gases are utilised in various industries, such as chemical plants, power plants, or manufacturing facilities. Invite guest speakers, such as engineers or researchers, to share their experiences and insights into the practical applications of gas behaviour and the kinetic theory of matter. Allow learners to engage with professionals and ask questions to gain real-world perspectives on the topic.

Key Assessment:

DoK Level 2:(Skills of conceptual understanding):

Describe how the kinetic theory of matter explains the behaviour of gases. Provide one example to illustrate how gas particles move and interact with each other.

DoK Level 2:(Skills of conceptual understanding):

A sealed container of gas is heated, causing an increase in temperature. Explain how the kinetic theory of matter can be used to predict the changes in gas pressure and volume inside the container.

DoK Level 4: (Extended critical thinking and reasoning):

A gas cylinder is compressed to half its original volume while maintaining a constant temperature. Using the principles of the kinetic theory of matter, explain how the gas pressure changes as a result of compression

Theme or Focal Area: Explanation of change of state processes

Melting: When solids are heated, particles gain kinetic energy and vibrate more strongly. Attractive forces weaken, particles break and become free to move around.

Freezing: When liquids are cooled, particles lose kinetic energy. Attractive forces strengthen, causing particles to be restricted to a fixed position and vibrate.

Evaporation and boiling:

- **a.** Evaporation occurs when the particles of a liquid escape to form vapour.
- **b.** This is due to the collision of particles with a variety of kinetic energy and speed.
- **c.** On heating, surface particles gain more kinetic energy, move faster, and break away from intermolecular forces.
- **d.** Boiling is rapid vaporisation anywhere in the bulk liquid at a fixed temperature.

Condensation: It is the change in the physical state of matter from the gaseous form into the liquid form. This occurs when the temperature of the gas is lowered to the point where its particles lose enough kinetic energy to form bonds and transition into a liquid. Condensation is the reverse process of vaporisation.

Melting Point Determination

- **a.** Heat one end of the capillary tube in a Bunsen flame.
- **b.** Fill about one-quarter of the capillary tube with the substance and tie it to a thermometer.
- **c.** Insert the tube and thermometer into an oil bath.
- **d.** Heat with constant stirring and record the temperature at which the first crystal melts and the temperature at which the last crystal melts

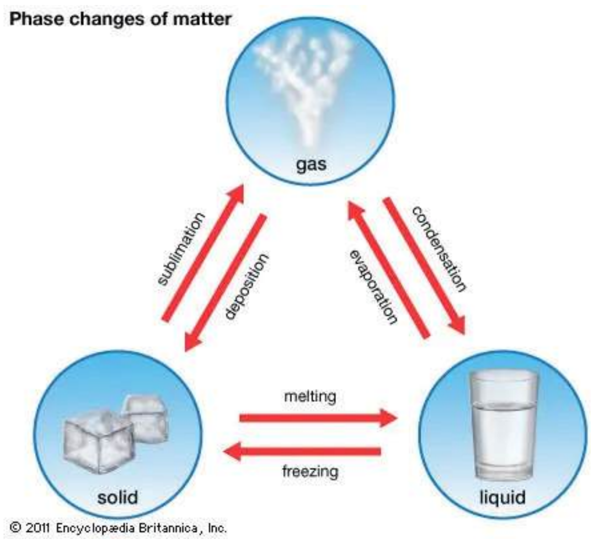


Figure 10.0: *A Diagram Showing the phase exchange of matter.*

Learning Tasks:

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

1. Provide learners with various materials such as ice cubes, water, and a heating source. Ask them to observe and record the changes that occur as the ice melts into water and then boils into steam. Guide students to describe the physical changes associated with each state of matter and explain the process of melting and boiling using simple language.

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

Experiential learning:

In mixed ability groups, provide learners with ice cubes and a heat source to demonstrate melting and boiling.

Use visual aids such as videos or simulations to show change of state processes in real-world contexts, such as the water cycle or industrial processes like distillation.

Key assessments

DoK Level 2: (Skills of conceptual understanding):

Explain the difference between condensation and sublimation, providing examples of substances that undergo each of these changes of state.

DoK Level 3:(Strategic reasoning):

Describe how a solid can change directly into a gas without passing through the liquid state, using examples from everyday life.

DoK Level 4:(Extended critical thinking and reasoning):

Analyse the impact of changes in atmospheric pressure on the boiling point of a liquid. Explain how altitude affects the boiling point of water and its implications for cooking and food preparation.

References

1. SHS Chemistry Curriculum LI1.1.3.1.1
2. Gibert T. R. et al. (2009). Chemistry. 2nd ed. W. W. Norton & Company, Inc. 500 Fifth Avenue, New York. 10110
3. Addison – Wesley. (2000). Chemistry. 5th ed. Prentice Hall, New York
4. Lawrie Ryan & Roger Norris. (2014). Cambridge International AS and AA-level Chemistry Coursebook. University Printing House, Cambridge CB2 8B5
5. Chemistry. (2015) Science Resource Centre Handbook, Itec Global
6. Rosemarie Gallagher & Paul Ingram. (2014). Complete Chemistry for Cambridge IGCSE. 3rd ed. Oxford University Press, Great Clarendon Street, Oxford OX2 6DP
7. Chang Reymond. (2017). General Chemistry – Essential Concepts. The McGraw Hill Companies

Week 11

Learning Indicator: *State and perform calculations involving various Gas Laws and analyse graphs based on the laws.*

Theme or Focal Area: Gas laws

1. **Boyle's Law:** It states that the volume of a fixed mass of gas at constant temperature is inversely proportional to the pressure.

   Mathematically Boyle's law is expressed as,

   $$P_1 V_1 = P_2 V_2$$

   Where P_1 and P_2 are initial and final pressures respectively

   And V_1 and V_2 are initial and final volumes respectively.

   Graphically Boyle's law is represented as follows;

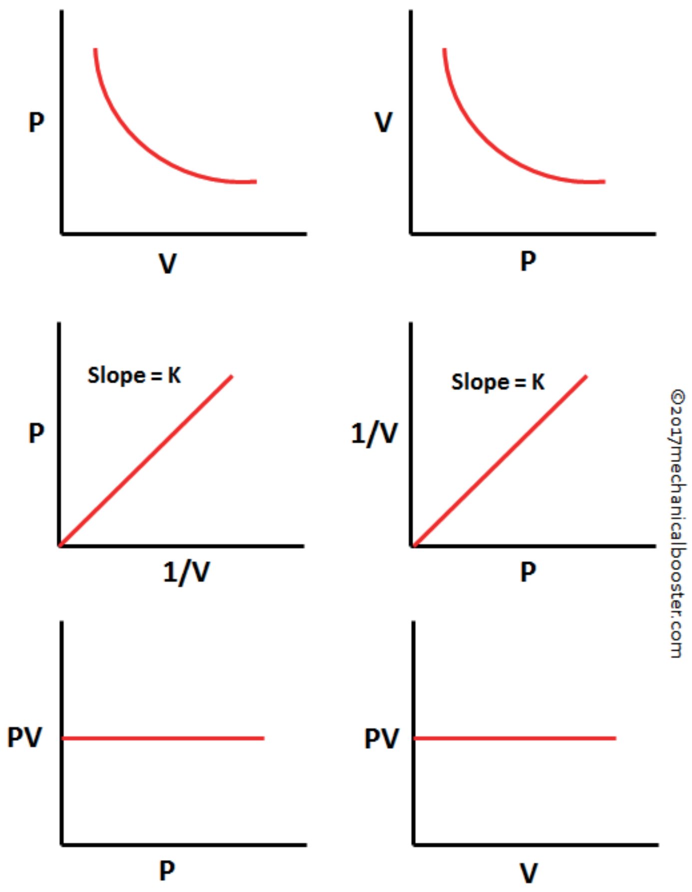


Figure 11.0: *Graphs Showing Different Representation of Boyle's Law.*

Example:

10 m^3 volume of a gas at a pressure of 101 300 Pa was compressed to a volume of 6 m^3 at constant temperature, calculate the final pressure.

Solution: (Use problem solving strategy)

a. Analyse the question

Known

Initial volume, $V_1 = 10\ m^3$

Initial pressure, $P_1 = 101\ 300$ Pa

Final volume, $V_2 = 6\ m^3$

Unknown

Final pressure, P_2 =?

b. Apply the problem-solving strategy

$$P_1V_1 = P_2V_2$$

$$P_2 = \frac{P_1V_1}{V_2}$$

$$= \frac{101300 \times 10}{6}$$

$$= 168833\, P_a$$

2. **Charles' Law:** It states that the volume of a fixed mass of a gas at constant pressure is directly proportional to the absolute temperature.

Mathematically Charles law is represented as,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Where V_1 and V_2 are initial and final volumes respectively

And T_1 and T_2 are initial and final temperatures respectively.

NB: The temperature must always be converted to Kelvin.

Graphically Charle's law is represented as follows;

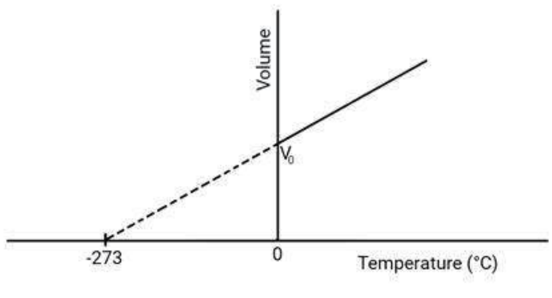


Figure 11.1: *Graphical Representation of Charles Law.*

Example:

10 m^3 volume of a gas in a cylinder is heated from 250 K to 300 K at constant pressure, calculate the final volume of the gas in the cylinder.

Solution: (Use problem solving strategy)

a. Analyse the question

Known

Initial volume, $V_1 = 10\ m^3$

Initial temperature, $T_1 = 250$ K

Final temperature, $T_2 = 300$ K

Unknown

Final volume, $V_2 = ?$

Apply the problem-solving strategy

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = \frac{V_1 T_2}{T_1}$$

$$= \frac{10 \times 300}{250}$$

$$= 12\ m^3$$

3. **Gay-Lussac's law:** It states that for a fixed mass of gas at constant volume, the pressure of a gas is directly proportional to the absolute temperature (K).

Mathematically Gay-Lussac's law is represented as,

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Example:

A fuel and air mixture in a car engine cylinder of volume 1000 cm^3 increases from 20 ^{0}C to 2000 ^{0} C upon combustion. If the normal atmospheric pressure is 100 kPa, calculate the final pressure.

Solution: (Use problem-solving strategy)

Analyse the question

Known

Initial temperature, $T_1 = 20 + 273 = 293$ K

Final temperature, $T_2 = 2000 + 273 = 2{,}373$ K

Initial pressure, $P_1 = 100$ Pa

Unknown

Final pressure, $P_2 =?$

Apply the problem-solving strategy

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$P_2 = \frac{100 \times 2273}{293}$$

$$= 775.8\ kPa$$

Evaluate: Check the answer to see if it makes sense.

4. **Combined gas law:** It is the combination of the Boyle's law, Charle's law and the Gay-Lussac's law.

Mathematically it is expressed as

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Example:

20 cm^3 of a gas at 1 atm and 25 ^{0}Cwas compressed to 16 cm^3 at 40 ^{0}C, calculate the final pressure of the gas.

Solution: (Use problem-solving strategy)

Analyse the question

Known

Initial pressure, P_1 = 1 atm

Initial volume, V_1 = 20 cm^3

Final volume, V_2 = 16 cm^3

Initial temperature, T_1 = 25 + 273 = 298 K

Final temperature, T_2 = 40 + 273 = 313 K

Unknown

Final pressure, P_2 =?

Apply the problem-solving strategy

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$P_2 = \frac{P_1V_1T_2}{T_1V_2}$$

$$= \frac{1 \times 20 \times 313}{298} \times 16$$

= 1.31 atm

5. **Avogadro's Law:** It states that equal volumes of gases at the same temperature and pressure contain the same number of molecules or moles of gas.

Mathematically Avogadro's law is expressed as,

$$\frac{V}{n} = K$$

Where V = Volume occupied by the gas

n = number of moles of the gas

K = constant of proportionality

Example:

Consider the reaction,

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

If the rate of production of CO in anindustrial plant is at 20 dm^3 per minute, what is the rate of production of oxygen to produce CO_2?

Solution

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

$$\frac{V(CO)}{V(O_2)} = \frac{2}{1}$$

$$V(O_2) = \frac{V(CO)}{2}$$

$$= \frac{20}{2} = 10\ dm^3\,min^{-1}$$

Learning Tasks

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

1. State at least one (1) of the following gas laws:
   a. Boyles' law
   b. Charles' law
   c. Gay-Lussac's law
   d. Combined gas law
2. Express any two (2) of the gas laws mathematically taking into consideration the conditions associated with each law.
   i. Identify the variable that are related to each law.
   ii. Illustrate the relationships among the various variables of the law.
3. a. Explain how Boyles' law can be applied in real life situation.
   b. Derive the combined gas law from Boyle's, Charles' and Gay–Lussac's laws.
4. a. 20 m^3 of gas at a pressure of 100000 Pa was compressed to a pressure of 400000 Pa at constant temperature. Calculate the final volume of the gas.
   b. 20 dm^3 of gas in a cylinder is heated from 150 K to a certain temperature. If its volume expands to 35 dm^3, Calculate the final temperature of the gas.
   c. 35 cm^3 of a gas at 1.5 atm at 40 ^{0}C was compressed to 20 cm^3at 50 ^{0}C, calculate the final pressure of the gas.
   d. A balloon was filled to a volume of 2.0 dm^3 with 0.082 moles of helium gas. Suppose 0.015moles of helium is added to the balloon with constant temperature and pressure, what will be the new volume of the balloon?
   e. The temperature of a gas contained in a cylinder undergoes combustion in constant volume, increasing from 30 ^{0}C to 1500 ^{0}C. If the normal atmospheric pressure is 101 kPa, calculate the peak pressure reached after combustion.

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

Exploratory Learning:

1. In small mixed-ability groups, research from the internet, library, books, and other resources about the various gas laws (Boyle's law, Charles' law, Avogadro's law and Gay–Lussac's law) in terms of:
   a. The formulae
   b. Variables
   c. Conditions required
2. Use simulation (PhEt) or charts to illustrate the relationships among the various variables.

3. Sketch graphs of the relationship and deduce the mathematical relationship between the variables.
4. Based on Boyle's, Charles' and Gay–Lussac's laws, determine the expression for the combined gas law.
5. Guide learners to carryout calculations based on the gas laws making references to the mathematical expressions of the laws.

NB: Support those learners who are struggling to express the laws mathematically and also finding it difficult to apply the expressions in solving questions.

Key Assessment

DoK Level 1: Reproduction/Recall

State Boyle's law.

DoK Level 2: Skills of conceptual understanding

A 250 cm^3 sample of a gas has a pressure of 20 Pa. What will be its volume if the pressure is raised to 45 Pa at the same temperature?

DoK Level 3: Strategic reasoning

A balloon can hold 800 cm^3 of air before bursting. If the balloon contains 500 cm^3 of air at 10 °C, determine whether or not the balloon would burst when it is taken into a room of temperature of 20 °C, assuming that the pressure of the gas in the balloon is kept constant.

References

1. SHS Chemistry Curriculum LI1.1.3.1.2
2. Gibert T. R. et al. (2009). Chemistry. 2nd ed. W. W. Norton & Company, Inc. 500 Fifth Avenue, New York. 10110
3. Addison – Wesley. (2000). Chemistry. 5th ed. Prentice Hall, New York
4. Lawrie Ryan & Roger Norris. (2014). Cambridge International AS and AA-level Chemistry Coursebook. University Printing House, Cambridge CB2 8B5
5. Chemistry. (2015) Science Resource Centre Handbook, Itec Global
6. https://docbrown.info/page03/3_52gaslaws.htm
7. https://www.khanacademy.org/science/in-in-class11th-physics/in-in-phy-kinetic-theory/in-in-phy-ideal-gas-laws/v/boyles-law
8. Chang Reymond. (2017). General Chemistry – Essential Concepts. The McGraw Hill Companies

Week 12

Learning Indicator(s): *State Graham's Law of diffusion/effusion and Dalton's Law of partial pressure and apply them to perform calculations.*

Theme or Focal Area: Diffusion

It is the random motion of molecules by which there is a net flow of matter from a region of high concentration to a region of low concentration. A familiar example is the perfume of a flower that quickly permeates the still air of a room.

Graham's law of diffusion/effusion

It states that the rate of diffusion, or effusion, is inversely proportional to the square root of its density at constant temperature and pressure.

Mathematically, $R \propto \frac{1}{\sqrt{D}}$ or $R \propto \frac{1}{\sqrt{M}}$

R = *Rate of diffusion or effusion*

D = *Density of the gas*

M = *Molecular mass of the gas*

Diffusion involves the movement of molecules from an area of high concentration to an area of low concentration due to random motion of molecules. It occurs in gases, liquids and solids and does not require a boundary. On the other hand, effusion specifically refers to the escape of gaseous molecules through a tiny hole into a vacuum or region of lower pressure. It is a type of diffusion but it is specifically about gaseous molecules moving through a tiny opening.

For 2 gases diffusing at the same time

$$\frac{R_1}{R_2} = \sqrt{\frac{D_2}{D_1}} \quad \text{and} \quad \frac{R_1}{R_2} = \sqrt{\frac{M_2}{M_1}}$$

Where R_1 and R_2 are rates of diffusion of gases 1 and 2.

Where D_1 and D_2 are densities of gases 1 and 2.

Where M_1 and M_2 are molecular masses of gases 1 and 2

Graham's law of diffusion can also be expressed in terms of time.

For two gases,

$$\frac{t_1}{t_2} = \sqrt{\frac{M_2}{M_1}}$$

Where t_1 and t_2 are times for the diffusion of gases 1 and 2

Graham's law of diffusion can also be expressed in terms of volume. For two gases,

$$\frac{V_1}{V_2} = \sqrt{\frac{M_2}{M_1}}$$

Where V_1 and V_2 are volumes for gases 1 and 2

Experiment to Demonstrate Graham's law of Diffusion

Figure 12.0: *A diagram Showing Graham's Law of Diffusion.*

Procedure

1. Soak a piece of cotton wool in concentrated NH_3 solution and soak another piece of cotton wool in concentrated HCl.
2. Put at opposite ends of a dry glass tube.
3. Ensure that the glass tube is horizontally mounted to ensure that the diffusion of the gas is not under the influence of gravity.
4. After a few minutes, a white cloud of ammonium chloride appears. This shows the position at which the two gases react.
5. The white cloud of NH_4Cl forms at a position closer to the cotton wool soaked in HCl because particles of NH_3 are lighter than that of HCl and so move faster and where they meet react to give that white fume.

Example:

If 100 cm^3 of methane (CH_4) gas diffuses through a membrane in 40 s, what time will it take 150 cm^3 of ammonia (NH_3) gas to diffuse through the same membrane? [C = 12, H = 1, N = 14]

Solution: (Use problem-solving strategy)

a. **Analyse the question**

Known

[C = 12, H = 1, N = 14]

·Time, t (CH_4) = 40 s

·Volume, V(CH_4) = 100 cm^3

·Volume, V(NH_3) = 150 cm^3

Unknown

·Time, t (NH_3) =?

b. Apply the problem-solving strategy

$M_r(NH_3) = 14 + 3(1) = 17$

$M_r(CH_4) = 12 + 4(1) = 16$

$$R_{(NH_4)} = \frac{V}{t}$$

$$= \frac{100}{40}$$

$$= 2.5\ cm^3 s^{-1}$$

$$\frac{R_{(NH_3)}}{R_{(CH_4)}} = \sqrt{\frac{M_{(CH_4)}}{M_{(NH_3)}}}$$

$$\frac{R_{(NH_3)}}{2.5} = \sqrt{\frac{16}{17}}$$

$$R_{(NH_3)} = 0.9696 \times 2.5$$

$$= 2.4\ cm^3 s^{-1}$$

$$t_{(NH_3)} = \frac{150}{2.4}$$

$$= 62.5\ s$$

Learning Tasks

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

Learners

1. **a.** Discuss Graham's law of diffusion/effusion.
   b. use Graham's law of diffusion/effusion to describe and explain the effect of relative molecular mass and density on the rate of diffusion/effusion of gases.
2. **a.** Design an experiment to investigate the rate of diffusion of ammonia gas and hydrogen chloride gas.
   b. make deductions from the results of the experiment.
3. Perform calculations involving Graham's law of diffusion/effusion.

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

Talk for learning

a. In small mixed-ability groups, guide learners to discuss the Graham's law of diffusion/ effusion. Encourage each learner to play an active role in the discussions by asking questions.

b. In ability groups guide learners to use the Graham's law of diffusion/effusion to describe and to explain the effect of relative molecular mass and density on the rate of diffusion/effusion of gases. Encourage each learner to play an active role in the discussions by asking questions.

Experiential Learning

a. In mixed ability groups, guide learners to design an experiment to investigate the rate of diffusion of ammonia gas and hydrogen chloride gas. Offer enough support to learners who may be experiencing some difficulty.

b. Guide learners to make deductions from the result of the experiments.

Problem solving approach

Guide learners to perform calculations on Graham's law of diffusion/effusion by solving samples questions with them to understand.

In ability groups, give the learners questions on the Graham's law of diffusion/effusion to solve. Offer enough support to leaners who may be slow in solving these questions.

Key Assessment

DoK level 1: Reproduction/ Recall

State Graham's law of diffusion or effusion.

DoK level 2: Skill of conceptual understanding

Given that 100 cm^3 of ethane (C_2H_6) diffuses through a membrane in 40 s, what time will it take 80 cm^3 of propane (C_3H_8) to diffuse through the same membrane at the same temperature and pressure? [C = 12, H = 1].

DoK level 4: Extended critical thinking and reasoning.

Given a cylindrical glass tube, stoppers, two retort stands, cotton wool, concentrated NH_3, concentrated HCl, and tweezers, design an experiment to determine the rate of diffusion of a gas.

Theme or Focal Area(s): Dalton's Law of partial pressures

It states that, in a mixture of gases which do not react, the total pressure exerted is equal to the sum of the partial pressures of the individual gases at constant temperature.

· For a mixture of gases 1 and 2,

$P_T = P_1 + P_2$

where P_T = total pressure

P_1 and P_2 are the partial pressures of gases 1 and 2

$P_1 = X_1 P_T$

$P_2 = X_2 P_T$

where X_1 and X_2 are the mole fractions of gases 1 and 2

$$X_1 = \frac{n_1}{n_1 + n_2}$$

$X_2 = \frac{n_2}{n_1 + n_2}$, where n_1 and n_2 are the number of moles of gases 1 and 2

Example:

Consider the reaction,

$N_2 + 3H_2 \rightarrow 2NH_3$

If the total pressure is 150 atm, calculate the partial pressure of nitrogen gas.

Solution: Use problem – solving approach

Determine the mole ratio of the gaseous species involved

Mole fraction of N_2,

$$X_{N_2} = \frac{n_{N_2}}{n_{N_2} + n_{H_2} + n_{NH_3}}$$

$$= \frac{1}{1 + 3 + 2}$$

$$= \frac{1}{6}$$

$$P_{N_2} = X_{N_2} \times P_T$$

$$= \frac{1}{6} \times 150$$

$$= 25.05\ atm$$

Learning tasks

Carry out a check for relevant/required prior knowledge from the JHS Science curriculum. This can be carried out via questioning, a quiz or think-pair-share. Reteach any material if necessary.

Learners

1. a. explain Dalton's law of partial pressure.

   b. explain how to apply Dalton's law of partial pressure to determine the total pressure of a mixture of gases or the partial pressure exerted by each component in a mixture of gases.
2. perform calculations involving Dalton's law of partial pressure.

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

Talk for learning

a. In small mixed-ability groups, guide learners to explain Dalton's law of partial pressure. Encourage slow learners to talk by directing questions to them.

b. In ability groups, guide learners to explain how to apply the Dalton's law of partial pressure in determining the total pressure of a mixture of gases or the partial pressure exerted by each component in a mixture of gases. Encourage each learner to play an active role by asking questions.

Problem solving approach

Guide learners to perform calculations on Dalton's law of partial pressures by solving sample questions for them to understand.

In ability groups, give learners questions on Dalton's law of partial pressures to solve. Offer enough support to learners who may be slow in solving the questions.

Key Assessment

DoK level 1:Reproduction/ Recall

State Dalton's law of partial pressures.

DoK level 2: Skill of conceptual understanding

1. A mixture of gases contains 4.76 mole of Ne, 0.74 mole of Ar and 2.5 mole of Xe. Calculate the partial pressure of the gases if the total pressure is 2 atm, at a fixed temperature.
2. A mixture of 40.0 g of oxygen and 40.0 g of helium has a total pressure of 0.900 atm. What is the partial pressure of each gas? N = 14, O = 16]

DoK level 3: Strategic Reasoning/Thinking

1. Three gases (8 g of CH_4, 18 g of C_2H_6, and unknown amount of C_3H_8) were added to the same 10.0 dm^3 container. At 23°C , the total pressure in the container was measured to be 4.43 atm. Calculate the partial pressure of each gas in the container. [C = 12, H = 1]

2. A gaseous mixture of O_2 and N_2 contains 32.8% nitrogen by mass. What is the partial pressure of oxygen in the mixture, if the total pressure is 785.0 mmHg? [N = 14, O = 16]

References

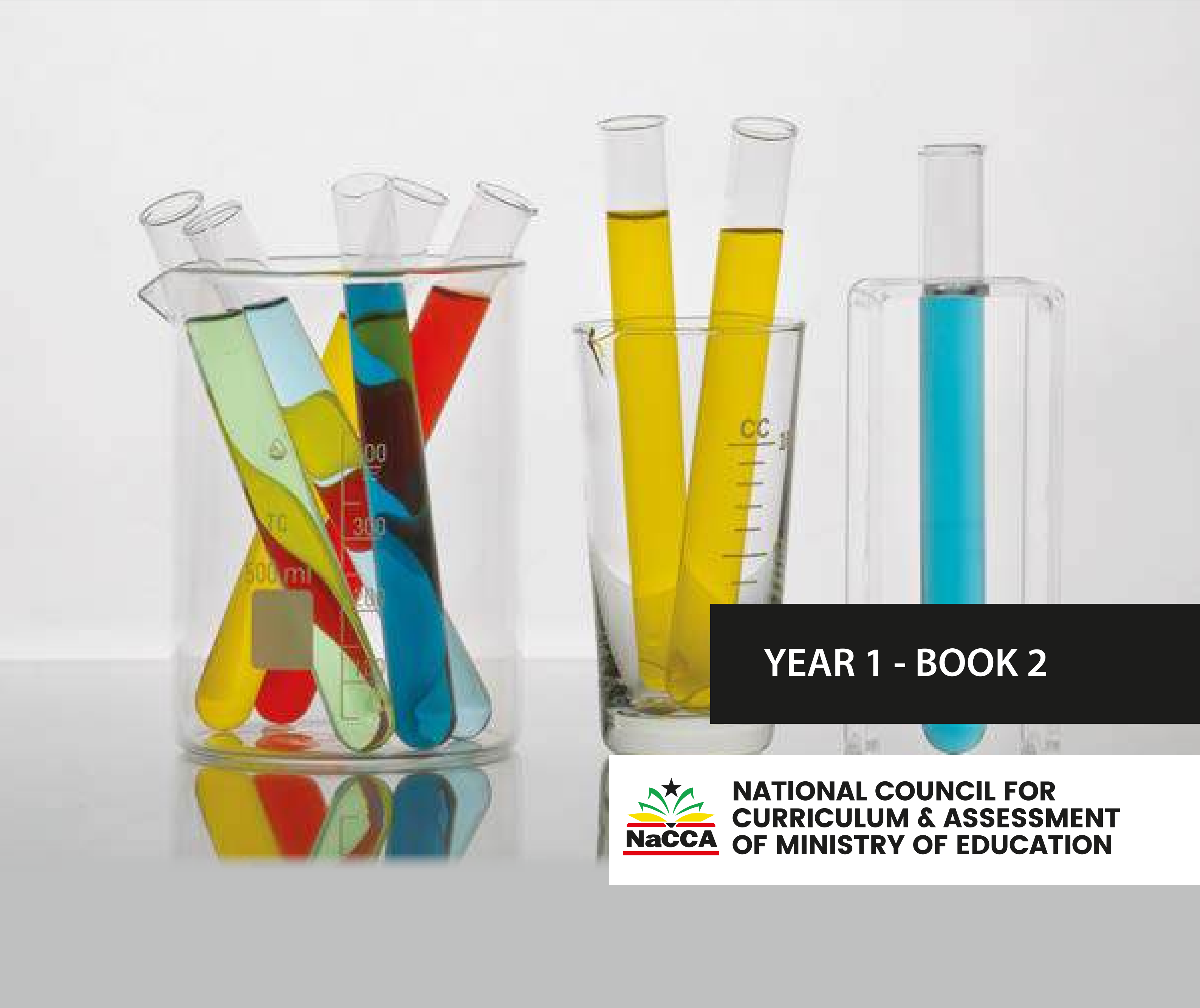
1. SHS Chemistry Curriculum LI1.1.3.1.3
2. Gibert T. R. et al. (2009). Chemistry. 2nd ed. W. W. Norton & Company, Inc. 500 Fifth Avenue, New York. 10110
3. Addison – Wesley. (2000). Chemistry. 5th ed. Prentice Hall, New York
4. Lawrie Ryan & Roger Norris. (2014). Cambridge International AS and AA-level Chemistry Coursebook. University Printing House, Cambridge CB2 8B5
5. Chemistry. (2015) Science Resource Centre Handbook, Itec Global
6. https://docbrown.info/page03/3_52gaslaws2.htm
7. Chang Reymond. (2017). General Chemistry – Essential Concepts. The McGraw Hill Companies

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

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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**

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

$\therefore$ volume = 0.0088 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 $Jmol^{-1}K^{-1}$, 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:

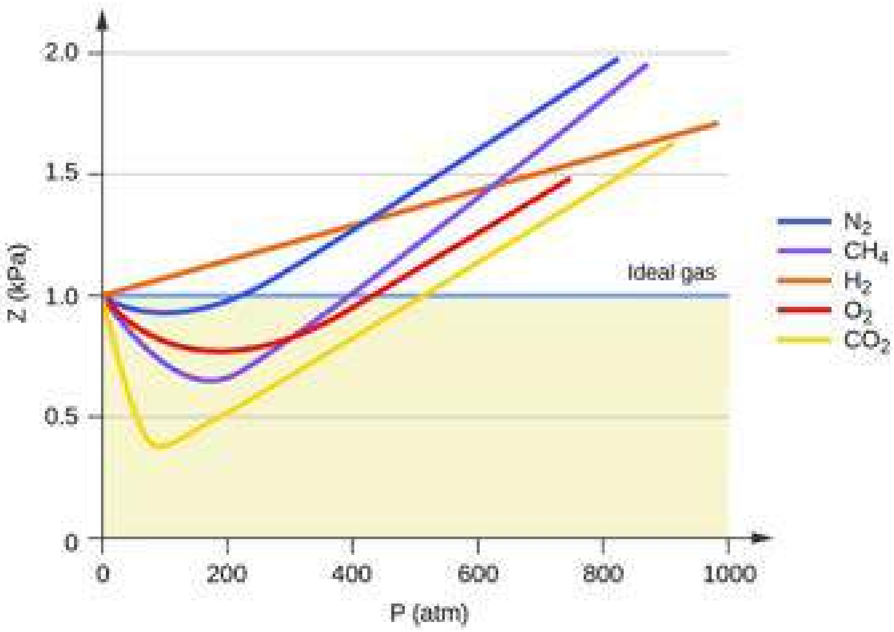
a. High-pressure

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

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

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/

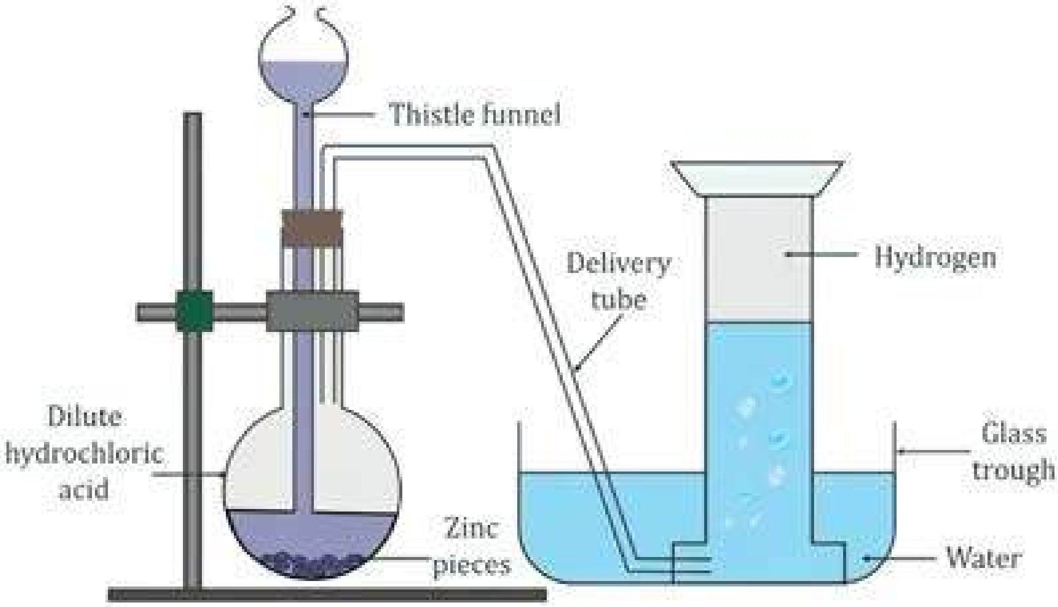
Week 14

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)*

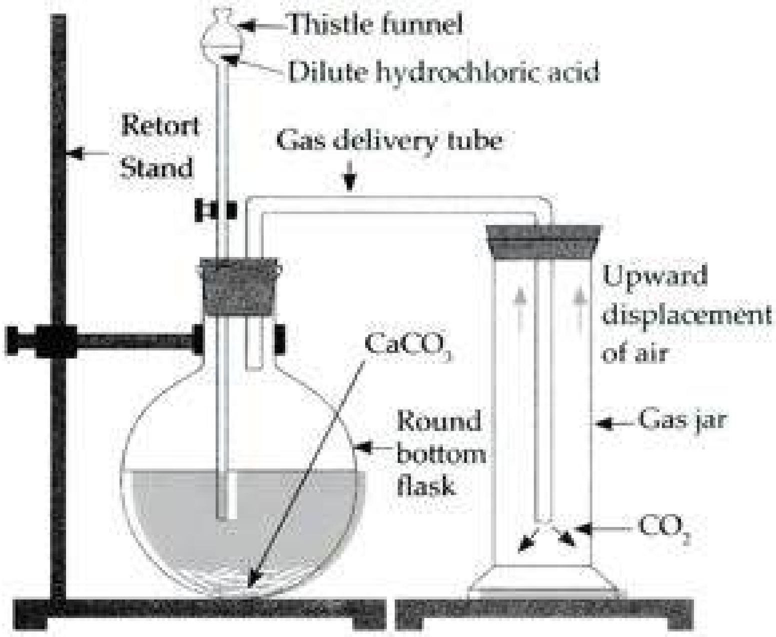
1. **Collaborative learning:**
   a. Divide learners into mixed-ability groups and guide them to discuss the properties of hydrogen gas. Encourage slow learners to participate in the discussion by asking them questions.
   b. Guide learners to state the uses of hydrogen gas in everyday life.
2. **Experiential 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)*

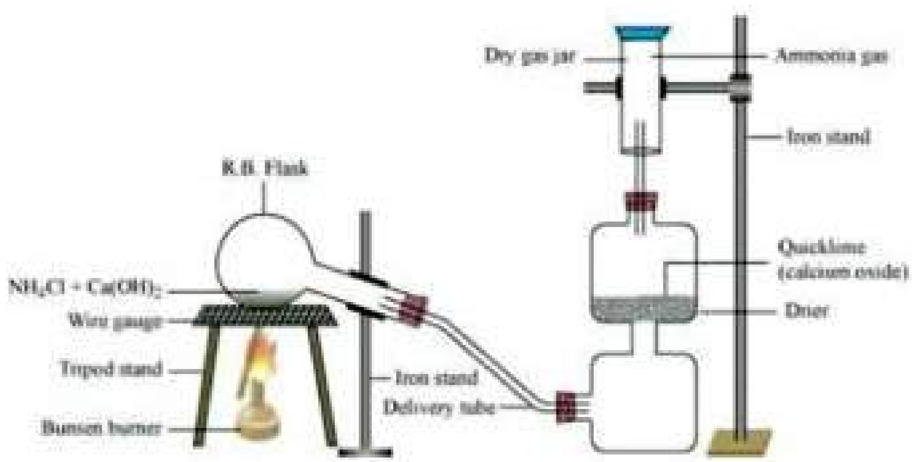
1. **Collaborative learning:**
   a. Divide learners into mixed-ability groups and guide them to discuss the properties of carbon dioxide gas. Encourage slow learners to participate in the discussion by asking them questions.
   b. Guide learners to state the uses of carbon dioxide gas in everyday life.
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

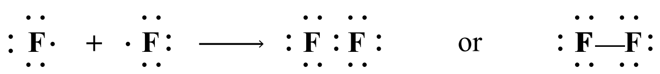
Learning Indicator: *Explain covalent bonding and its formation, and state the properties of covalent compounds.*

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. **H· + ·H → 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.

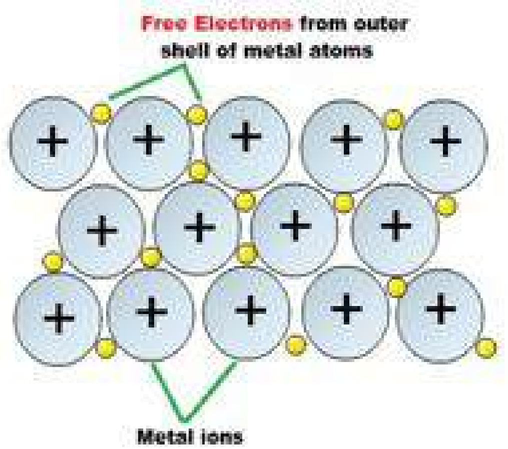
Key Assessments (DoK)

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.

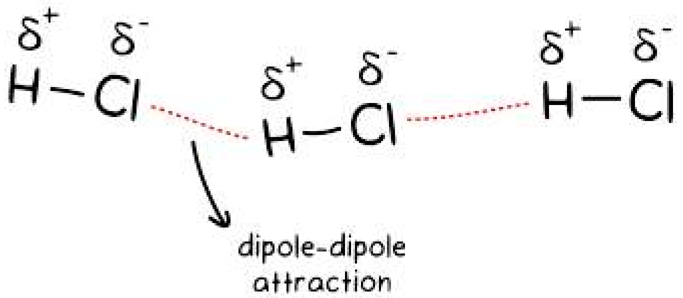
Van der Waal Forces

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

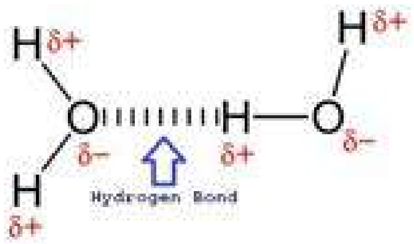
Dipole-dipole interaction, also known as dipolar interaction, is a type of intermolecular force that occurs between two polar molecules. Polar molecules have a permanent dipole moment, which is a separation of positive and negative charge due to differences in electronegativity between the atoms in the 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.



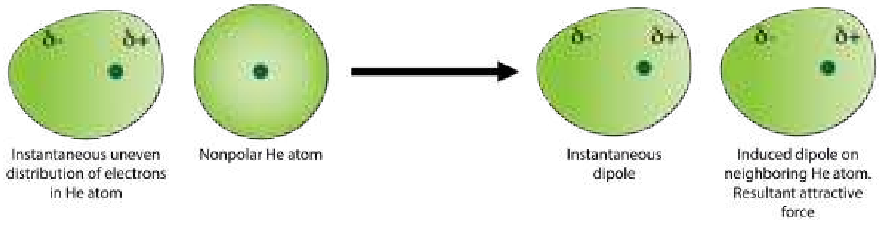
Hydrogen bonds are stronger than London dispersion forces and dipole-dipole forces but weaker than covalent bonds. They play a significant role in determining the properties of water, ice and many biomolecules, such as DNA and proteins. The unique properties of water, including its high boiling point, 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, I_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

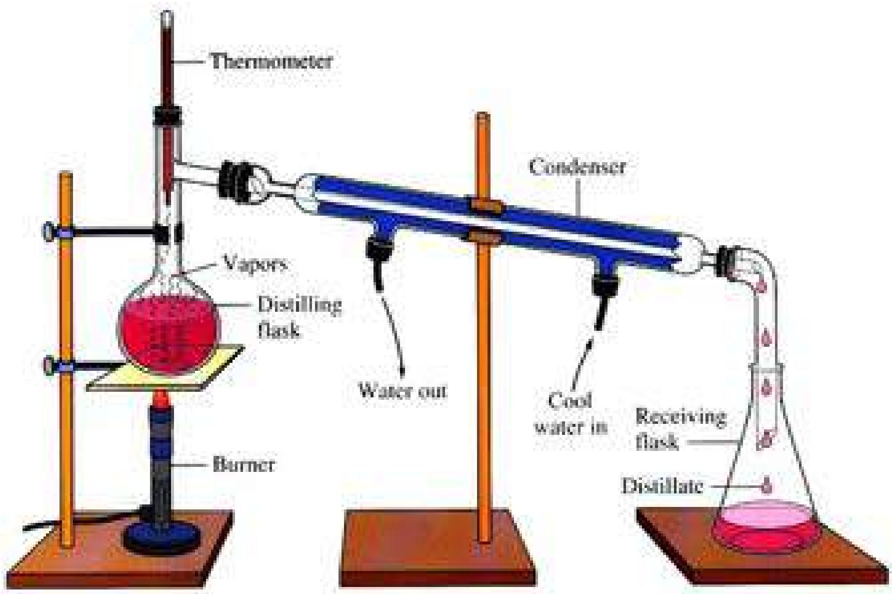
It is used to separate volatile liquids from a mixture of non–volatile liquids or solids. In this method of separation, the components of the mixture must have wide differences in their boiling points. This method of separation involves two physical processes, namely:

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

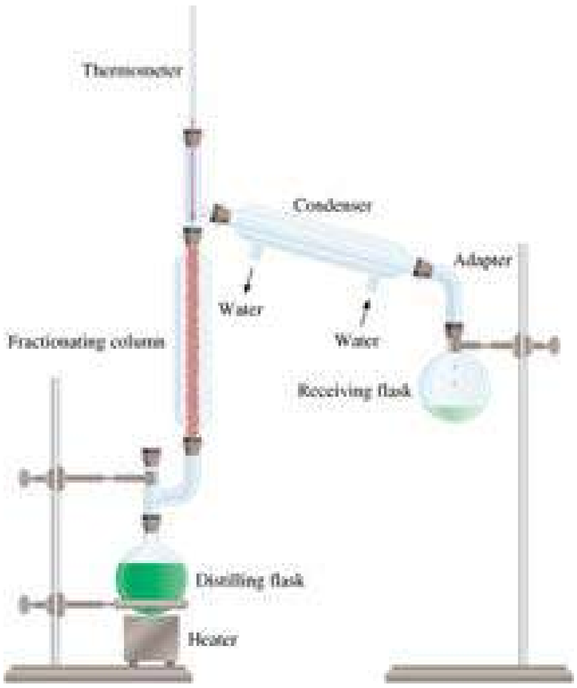
This method is used to separate two or more miscible liquids, having similar (close) boiling points. The process involves two main physical processes.

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.

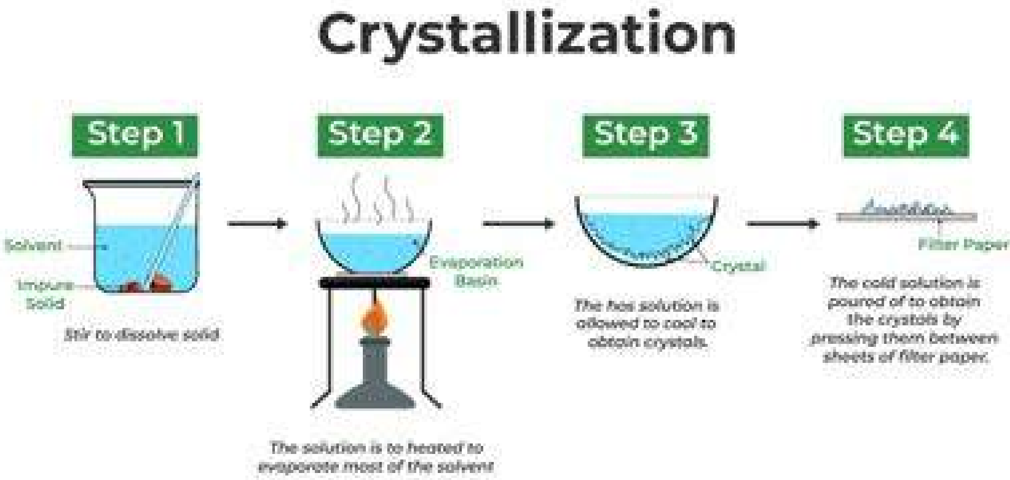
a) It is a purification method used to remove soluble impurities from a solid compound.

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.

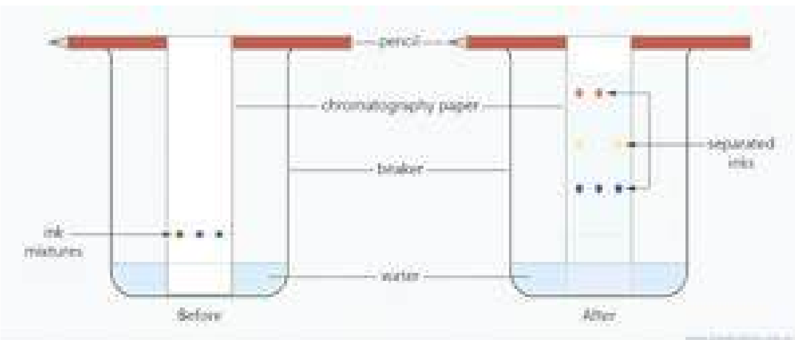
b) The highest point reached by the solvent is called the solvent front.

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.

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

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

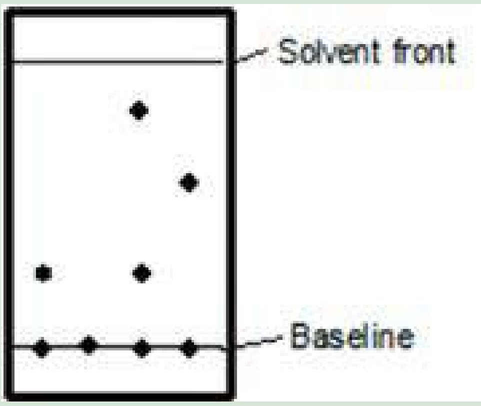
1. Define distillation and explain its purpose in the separation of mixtures.
2. Sketch the set-up of a simple distillation apparatus and explain the role of each component in the distillation process.
3. How will you recover and recycle a solvent containing dissolved solid impurities for an industrial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

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

  Mass of hydrogen $= \frac{2}{18} \times$ mass of H_2O 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:

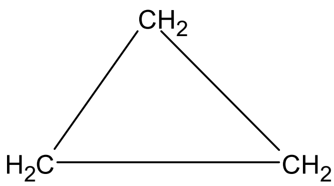
These are organic compounds containing carbon and hydrogen atoms that are usually linked together in chains via single, double or triple bonds. They can be open straight chains or branched chains. Aliphatic hydrocarbons may be saturated or unsaturated.

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.

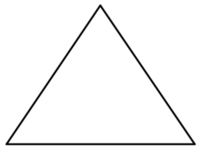
Alicyclic hydrocarbons: These are organic compounds that have closed rings of carbon atoms.

Examples are



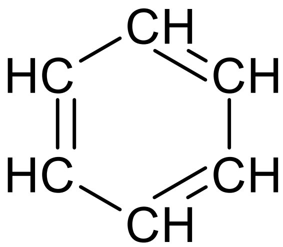
cyclopropane

or



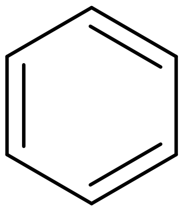
cyclopropane

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



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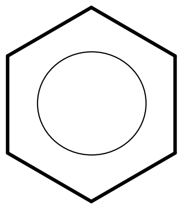
benzene

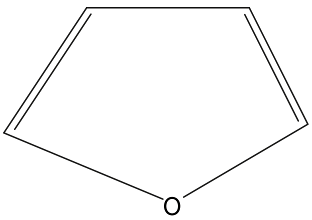


3

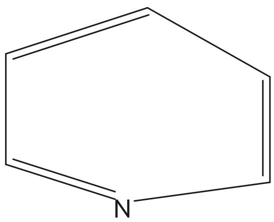
benzene

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.

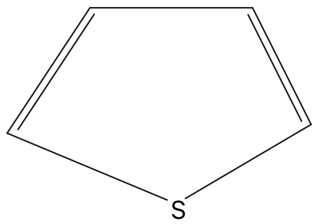
Examples are



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.

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