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

### Year One - Book One

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 21$^{st}$ 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 21$^{st}$ Century Skills and Competencies

The SHS, SHTS, and STEM curriculum integrates 21$^{st}$ 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.

# ACKNOWLEDGEMENTS

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

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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. | <br>**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. | <br>**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. | <br>**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 | <br>**Figure 1.8:**<br>*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. | <br>**Figure 1.8:**<br>*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: | <br>**Figure 1.9:**<br>*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: | <br>**Figure 1.10:**<br>*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. | <br><br>**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. | <br><br>**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. | <br><br>**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.

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# 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.

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# 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.

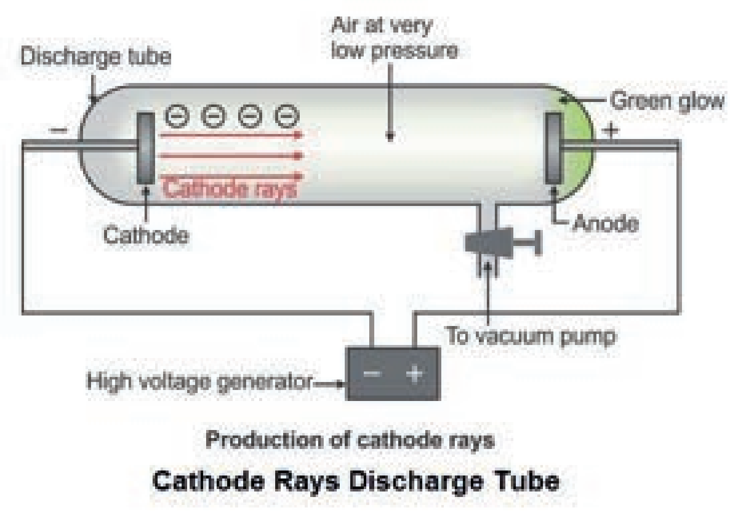


**Figure 3.0:** *A Cathode Ray Tube*

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.

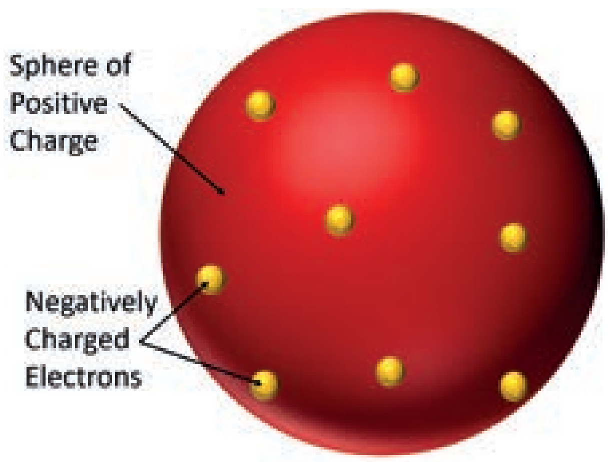


**Figure 3.1:** *Plum Pudding Model of the Atom.*

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.

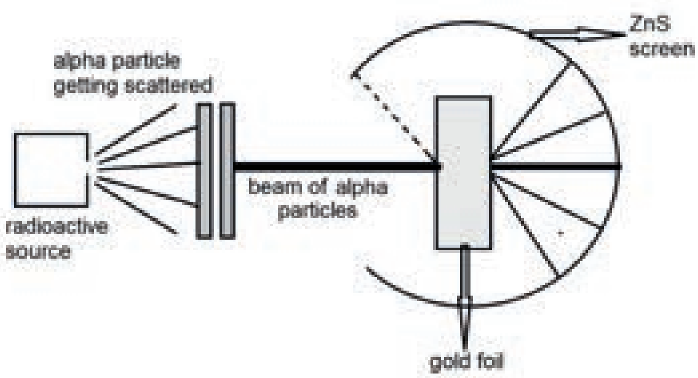


**Figure 3.2:** *Rutherford's alpha scattering experiment.*

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.

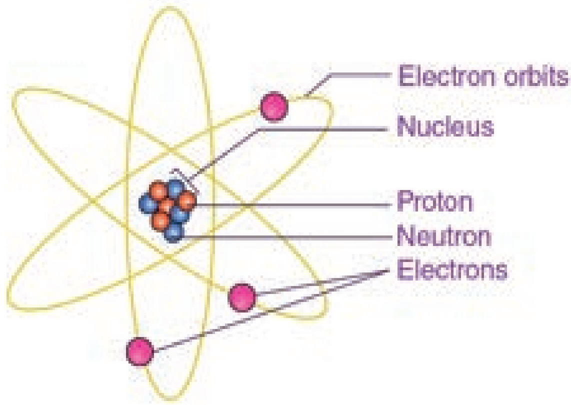


**Figure 3.3:** *Rutherford Model of the Atom.*

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?

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# 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.

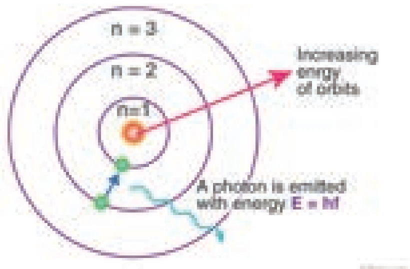


**Figure 4.0:** *Bohr's model of the atom*

## 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.

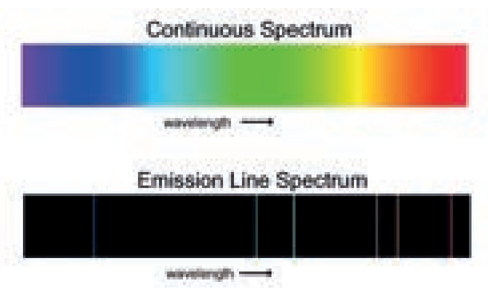


**Figure 4.1:** *Diagram showing continuous spectrum and emission spectrum*

### 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.

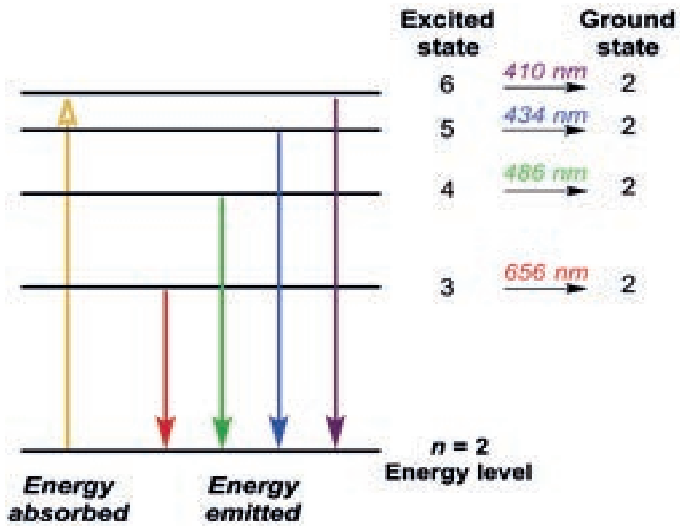


**Figure 4.2:** *Line Spectrum of Hydrogen and Energy Levels*

## 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.

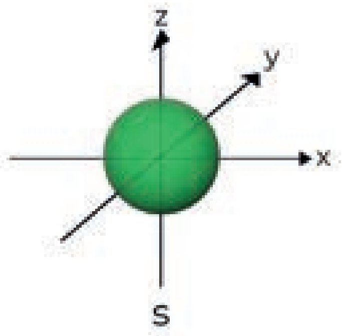


**Figure 4.3:** *Shape of the S-orbital*

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;

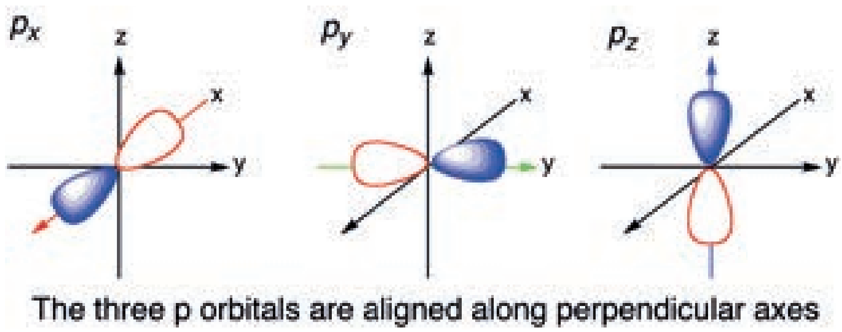


**Figure 4.4:** *The three P orbitals aligned along perpendicular axes*

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.

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# 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.

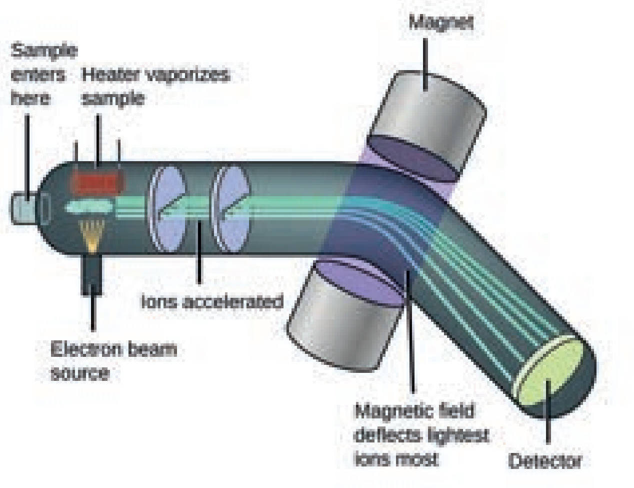


**Figure 5.0:** *Mass spectrometer.*

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.

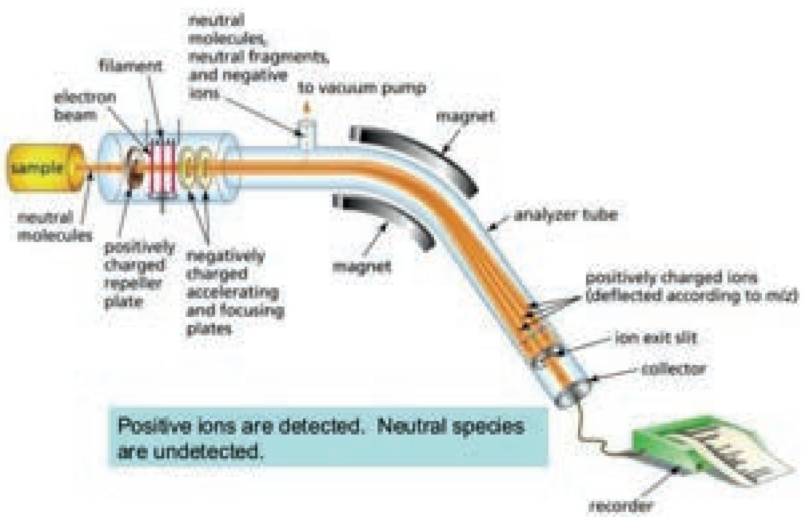


**Figure 5.1:** *Diagram of a Mass Spectrometer Processes that takes place in it*

**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.

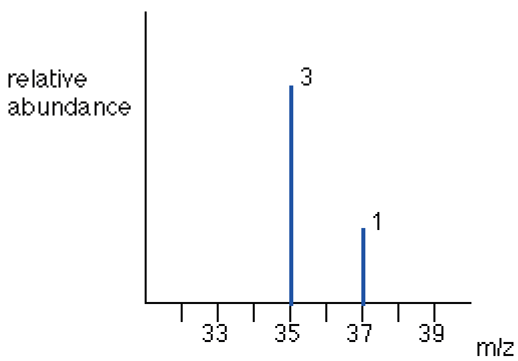


**Figure 5.2:** *Diagram of a Mass Spectra of Chlorine Element.*

**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 $\gamma$

## 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, $\Delta t$ i.e. Nt
3. Calculate the fraction of the original material that has decayed during the time interval, $\Delta t$ Using the formula: Fraction decayed $= \frac{activity\ at\ \Delta t}{initial\ activity}$

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

4. Calculate the natural logarithm of the fraction decayed:

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

5. Solve for the decay constant:

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

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

**Examples:**

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

$\Delta$t = 16 days.

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

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

$In\ 4 = 16\ \lambda$

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

$= 0.0866\ day^{-1}$

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

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

2. 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.

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

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

$= \frac{0.693}{33}$

$= 0.021 s^{-1}$

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

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

$In\ 16 = 0.021 \Delta t$

$\Delta t = \frac{In16}{0.021} = \frac{2.773}{0.021}$

$\Delta t = 132.0$ 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:

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

Assist learners to balance the following nuclear reactions:

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

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:

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

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# 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 $1/12^{th}$ 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 $1/12^{th}$ 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**

$Mr\ (NH_3) = Ar\ (N) + 3Ar(H)$

$= 14 + 3(1)$

$= 17$

**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 $1.6603 \times 10^{-24}$ g. If the average mass of an atom of oxygen is $2.65659 \times 10^{-23}$ 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* 1/12th of 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 $6.02 \times 10^{23}$ elementary entities.

For example, 1 mole of magnesium metal contains $6.02 \times 10^{23}$ 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 $6.02 \times 10^{23}$ 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 $9.5 \times 10^{23}$ 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 $10^{23}$]

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# 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 x 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, $m = n \times M$

$m = 0.5 \times 18$

$m = 9$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;

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

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:

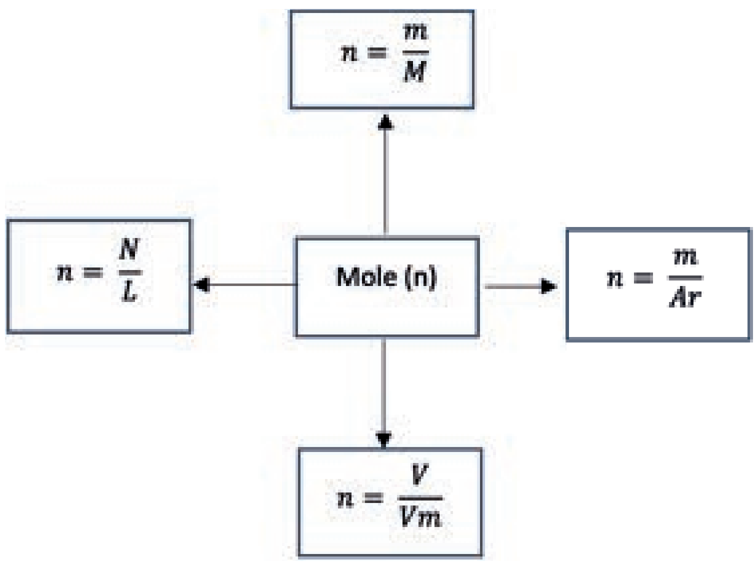


**Figure 7.0:** *A Concept Map Showing the Relationship between Mole and other Variables*

**Learning Task**

1. Calculate the number of oxide ions contained in 0.5 mol of $Al_2O_3$ [$N_A$ = 6.02 x $10^{23}$]
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 x $10^{23}$ $mol^{-1}$]
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 x $10^{23}$ $mol^{-1}$, Vm = 22.4$dm^3mol^{-1}$]

## 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 x $10^{23}$ 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 x $10^{23}$ 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:
$C = \frac{m}{M \times V}$

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 1$dm^3$ by multiplying by the percentage purity.

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

Mathematically use the formula:

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

**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 250$cm^3$ 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]

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# 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:

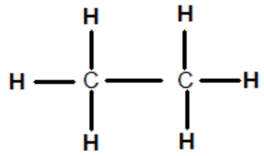

**Fig. 3.1**

**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.

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

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

**Solution**:

By definition,

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

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

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

$$= 82.35\ \%$$

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

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

$$= 17.65\ \%$$

**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
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# 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,

$2C + O_2 \rightarrow 2CO$

2 mol + 1 mol 2 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 $m = n \times M$ 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

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

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

$= 0.25 \text{ mol}$

Determine the mole ratio

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

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

$n(NH_3) = 2 \times 0.25$

$= 0.5$ mol

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

$= 17$

$m = 0.5 \times 17$

$= 8.5$ 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 x Vm 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

   $m(CH_4) = 10.5g,$

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

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

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

$$n(CO_2) = n(CH_4)$$

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

$$n(CO_2) = n(CH_4) = 0.66 \text{ mol}$$
$$\therefore n(CO_2) = 0.66 \text{ mol}$$

**5.** Use the calculated number of moles and the relationship V = n × Vm to calculate the volume of the gas.

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

**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

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

$m(Mg) = 12\text{ g}$

$M(Mg) = 24$

$m(MgCl_2) = ?$

**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.

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# 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.

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# 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,

$$P_1V_1 = P_2V_2$$

   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;

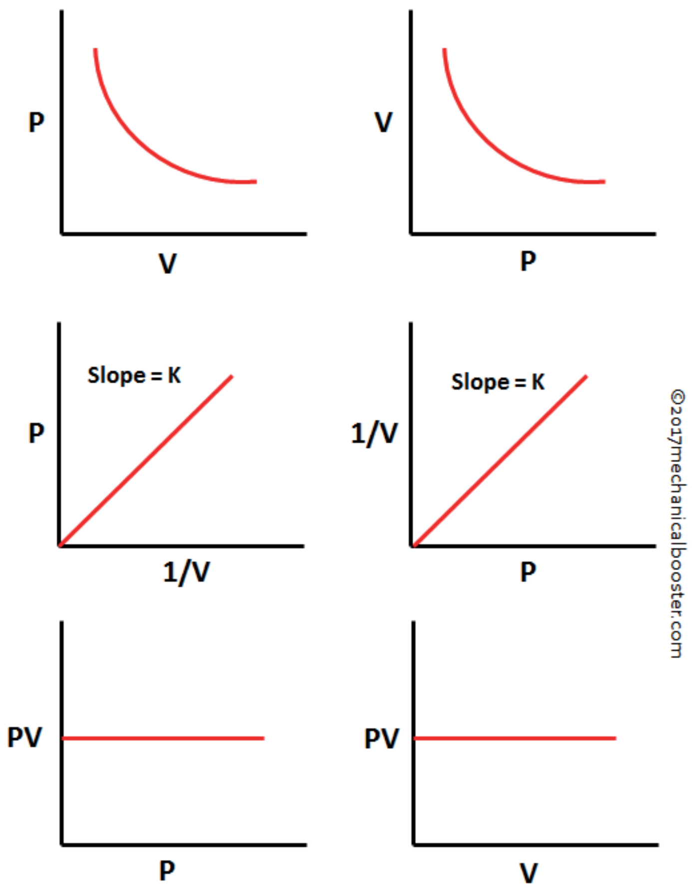


**Figure 11.0:** *Graphs Showing Different Representation of Boyle's Law.*

**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;

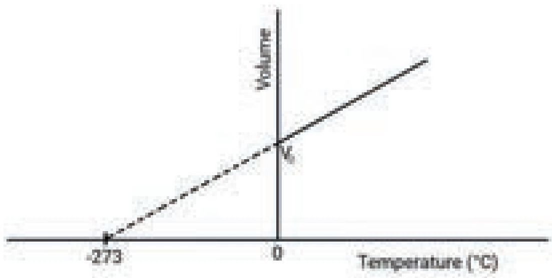


**Figure 11.1:** *Graphical Representation of Charles Law.*

**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 $^0C$ 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 $cm^3$at 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.

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# 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


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