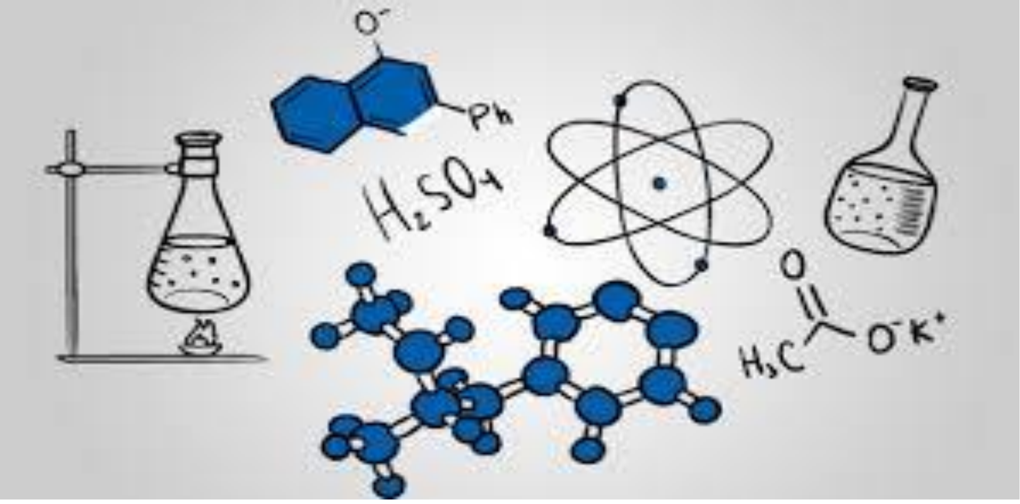
People's Democratic Republic of Algeria
Ministry of Higher Education and Scientific Research

University of Belhadj Bouchaib -Ain Temouchent-
Faculty of technological sciences

Chemistry 1
Structure of matter

Course and Corrected Exercises



Field: Science and Technology
Level: 1st year engineering

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Year 2024/2025

Preface

The study of matter and its structure forms the foundation of many disciplines within science, particularly chemistry. This course material, titled *"The Structure of Matter"*, has been carefully designed to introduce first-year engineering university students to the fundamental principles that govern the behavior and properties of matter.

Understanding the structure of matter is crucial not only for mastering chemistry but also for developing analytical and problem-solving skills that are applicable across various scientific domains. The topics covered in this polycopy include the fundamental concepts of atoms, molecules, chemical bonding, and the physical properties of matter, presented in a progressive and accessible manner.

Our goal with this material is to provide students with a solid grounding in theoretical and practical aspects of chemistry, enabling them to build upon these basics as they advance in their academic and professional journeys.

This document is the result of meticulous research and thoughtful organization, intended to align with the needs and capabilities of first-year university students. It is complemented by diagrams, examples, and exercises to ensure a comprehensive understanding of the topics discussed.

I would like to express my gratitude to my colleagues and students whose insights and feedback have contributed significantly to the creation of this material. I look forward to its role in supporting the academic growth of future engineers.

I hope this polycopy will serve as a reliable guide and resource for students as they embark on their journey in the fascinating world of chemistry.

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Introduction:

Chemistry is often referred to as the "central science" because it bridges the gap between physical and life sciences. Its principles are fundamental to understanding the structure and behavior of matter, which is essential for advancements in engineering, technology, and scientific research.

This course, *"The Structure of Matter"*, is designed for first-year engineering students to provide a comprehensive introduction to the core concepts of chemistry. The aim is to build a strong foundation, enabling students to grasp more advanced topics and their practical applications in engineering disciplines.

The subject is structured into four main chapters, each exploring a specific aspect of the structure of matter:

1. **Fundamental Concepts (3 weeks):**

This chapter lays the groundwork by introducing key principles such as the properties of matter, measurement techniques, and basic chemical laws. It establishes the vocabulary and theoretical framework needed for the subsequent chapters.

2. **Chapter 2: Structure of the atom (3 weeks) :**

Understanding the structure of the atom is a cornerstone of modern chemistry and essential for engineering applications. This chapter explores the fundamental components of the atom and their properties, focusing on key experiments that led to their discovery and characterization.

3. **Radioactivity (3 weeks):**

An exploration of the phenomenon of radioactivity, including its historical discovery, types of radioactive decay, and its applications. This chapter also highlights the role of radioactivity in modern engineering and scientific innovations.

4. **Electronic Structure of the Atom (4 weeks):**

This chapter delves into the quantum mechanical model of the atom, discussing orbitals, energy levels, and the behavior of electrons. A clear understanding of this topic is vital for comprehending chemical bonding and material properties.

5. **Periodic Classification of Elements (2 weeks):**

   The final chapter provides an in-depth look at the periodic table, examining trends and properties of elements. This knowledge is critical for predicting chemical reactivity and understanding the behavior of elements in various contexts.

This course has been meticulously structured to align with the learning pace and needs of first-year engineering students. By the end of this program, students will have a robust understanding of the fundamental concepts of chemistry, preparing them for more specialized topics in their academic and professional journey.

Teaching Units	Intitulés des matières	Code	Credits	Coefficients	Hourly volume Weekly			Half-yearly Hourly Volume (15 weeks)	Assessment mode	
					Course	TUTORIALS	PW		Continuous monitoring	Final exam
UE Fondamentale	Elements of Chemistry (structure of matter)	IST 1.3	7	4	1h30	3h00	1h30	90h00	40% (20% T + 20% PW)	60%

Teaching Objectives

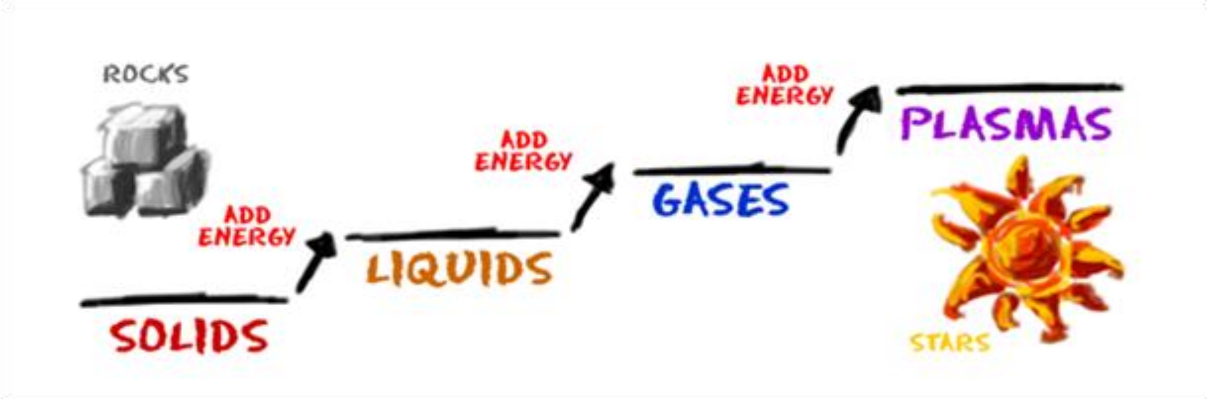
The primary objectives of this course, *"The Structure of Matter"*, are to provide first-year engineering students with a solid foundation in chemistry and to establish connections between chemical principles and engineering applications. Upon completing this course, students should be able to:

1. **Understand Fundamental Concepts:**
    - Grasp the basic properties and behavior of matter.
    - Comprehend essential chemical laws and measurement techniques.
    - Describe chemical processes around us and their applications in everyday life.
2. **Analyze Radioactive Phenomena:**
    - Describe the types and mechanisms of radioactive decay.
    - Recognize the applications and implications of radioactivity in engineering and scientific fields.
    - Describe radioactivity, and the properties of radiations and compare isotopes based on their stability as well as their applications in everyday life.
3. **Explore the Electronic Structure of Atoms:**
    - Explain how electron configurations influence chemical bonding and material properties.
    - 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

    - State the main postulates of Bohr’s planetary theory and explain the importance of the quantum numbers to the electron structure of the atom.
    - Understand the quantum mechanical model of atoms, including orbitals and energy levels.
4. **Utilize the Periodic Table Effectively:**
    - Predict trends in elemental properties using periodic classification.
    - Relate periodic properties to chemical reactivity and material design.
5. **Develop Problem-Solving Skills:**
    - Apply theoretical knowledge to solve practical problems in chemistry and engineering.
    - Analyze and interpret experimental data to draw meaningful conclusions.
6. **Foster Interdisciplinary Thinking:**
    - Connect principles of chemistry to broader engineering challenges.
    - Appreciate the role of chemistry in technological advancements and innovation.

By achieving these objectives, students will be well-prepared to tackle more advanced courses in chemistry and other scientific disciplines, ensuring a successful academic and professional journey in engineering.

Chapter 1
Fundamental concepts



I. Definition of Matter

Matter is defined as anything that has mass and occupies space. It is made up of small particles such as atoms and molecules, which are the building blocks of all substances.
Matter exists in different **states**, or **phases**, which are primarily solid, liquid, gas, and plasma.

- ***Solid:*** In this state, matter has a definite shape and volume. The particles are tightly packed and only vibrate in place (carbon, iron, copper ….).
- ***Liquid:*** Liquids have a definite volume but no fixed shape, taking the shape of their container. Particles are loosely packed and move more freely than in solids (water, milk …).
- ***Gas:*** Gases have neither a definite shape nor a definite volume. The particles are far apart and move freely (H_2, CO_2…).
- ***Plasma***: Plasma is an ionized gas with free-moving electrons and ions. It is found in stars and lightning (lightning, a fireball of a nuclear explosion, solar wind...)

II. Changes in the state of Matter

Matter can undergo two types of changes:

1- Physical Changes of Matter

A **physical change** involves a change in a substance's physical properties without altering its chemical composition. A change in the form or appearance of matter without forming a new substance. These changes are usually reversible, and the identity of the matter remains the same. **Physical changes** are generally **reversible** because no new substances are formed.

a)- Key Characteristics of Physical Changes:

1. **No new substances** are formed.
2. The **chemical composition** of the substance remains unchanged.
3. These changes are often **reversible**.
4. They involve changes in **physical properties** such as shape, size, phase (state), or texture.
5. Usually, there is **no energy change** that alters the substance chemically.

b)- Common Types of Physical Changes

1. Change in State (Phase Change)

Matter can change from one state (solid, liquid, gas) to another without altering its chemical structure. These changes are physical and include:

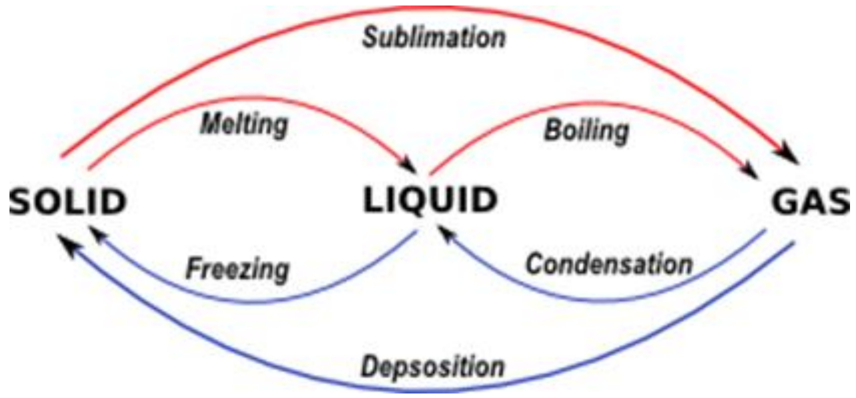


Figure I-1 : Change of state of a matter

- **Example:** Melting ice to form water. The chemical formula H_2O remains the same whether in solid or liquid form.

2. Change in Shape or Size

When matter is physically altered by cutting, bending, stretching, or crushing, it undergoes a physical change. The material's structure remains the same.

- **Example**: Breaking a glass bottle into pieces changes the shape and size of the glass but does not change the substance, which remains glass (SiO_2).

3. Dissolution (Dissolving)

When a substance dissolves in a solvent, it undergoes a physical change. The substance does not chemically react but simply disperses into the solvent.

- **Example**: Dissolving **salt (NaCl)** in water. The salt remains the same compound even though it's dispersed in water, and it can be recovered by evaporating the water.

4. Mixing/Substitution

When two or more substances are mixed without a chemical reaction, the change is physical. The individual components retain their identities and properties.

- **Example**: Mixing **sand and water**. The sand and water can still be separated, and no new substance is formed.

2- Chemical Changes of Matter

A **chemical change** involves a process in which one or more substances are transformed into entirely new substances with different chemical properties. These changes result in a chemical reaction, where the original substances (reactants) change into new products. Chemical changes often involve the breaking and forming of chemical bonds and are typically irreversible under normal conditions. Chemical changes are typically **irreversible** because the substances created are fundamentally different from the reactants.

a)- Key Characteristics of Chemical Changes:

1. **New substances** are formed.
2. The chemical composition of the matter **changes**.
3. These changes are generally **irreversible**.
4. Energy is often involved in the form of **heat**, **light**, or **sound**.
5. Chemical changes are accompanied by observable signs such as a **color change**, **temperature change**, the formation of a **gas**, or the appearance of a **precipitate**.

b)- Common Types of Chemical Changes

1. Combustion (Burning)

Combustion is a chemical reaction where a substance reacts with oxygen, releasing energy in the form of heat and light. This is often accompanied by the production of gases like carbon dioxide and water vapor.

- **Example**: Burning wood or fuel. When wood burns, it combines with oxygen to produce ash, carbon dioxide, and water, which are entirely different substances.

2. Oxidation and Reduction (Redox Reactions)

Oxidation involves the loss of electrons by a substance, while reduction involves the gain of electrons. Many reactions, including rusting and burning, are redox reactions.

- **Example**: **Rusting of iron**. Iron (Fe) reacts with oxygen (O_2) in the presence of water to form iron oxide (rust), changing the metal into a new substance.

3. Decomposition

In decomposition reactions, a single compound breaks down into two or more simpler substances. This often occurs when heat, light, or electricity is applied.

- **Example**: **Electrolysis of water**. When an electric current is passed through water (H_2O), it decomposes into hydrogen (H_2) and oxygen (O_2) gases.

4. Synthesis (Combination Reactions)

In a synthesis reaction, two or more simple substances combine to form a more complex substance.

- **Example**: The combination of hydrogen (H_2) and oxygen (O_2) to form water (H_2O) is a chemical change where the elements bond to create a new compound.

5. Precipitation

In a precipitation reaction, two solutions react to form an insoluble solid called a precipitate. The solid forms because the product is not soluble in the liquid mixture.

- **Example**: Mixing **silver nitrate ($AgNO_3$)** and **sodium chloride (NaCl)** results in the formation of **silver chloride (AgCl)** as a white precipitate.

6. Acid-Base Reactions (Neutralization)

In this type of reaction, an acid reacts with a base to form water and a salt. These reactions are important in chemistry and biological systems.

- **Example**: When **hydrochloric acid (HCl)** reacts with **sodium hydroxide (NaOH)**, they neutralize each other to form **water (H_2O)** and **sodium chloride (NaCl)** (table salt)

c) - Signs of a Chemical Change

There are several observable clues that a chemical change has occurred:

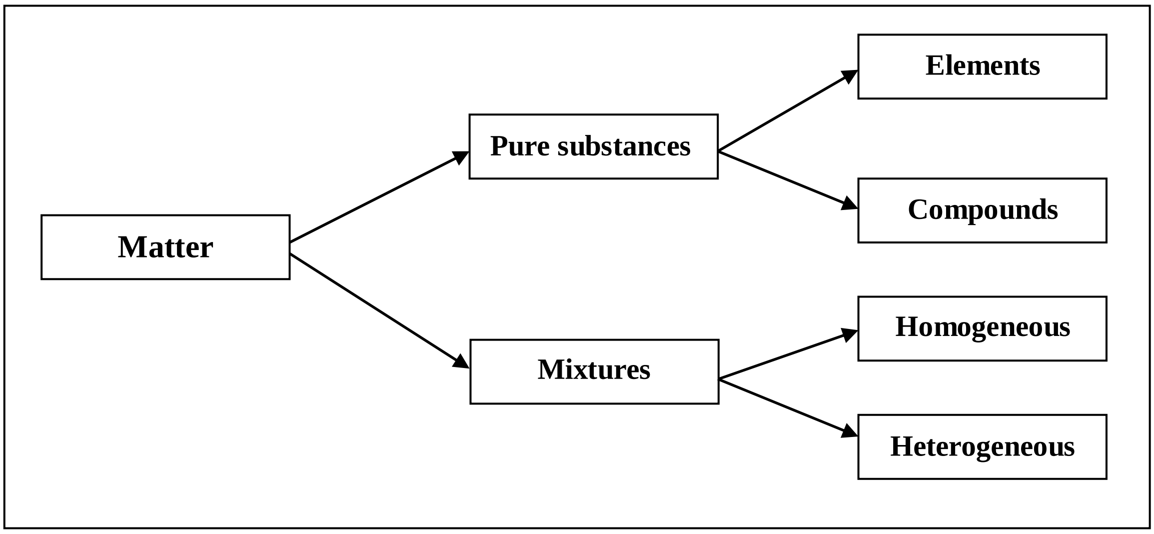
1. **Color change**: A substance changes color, indicating a chemical reaction.
    - **Example:** The tarnishing of silver turns it black as it reacts with sulfur in the air.
2. **Temperature change**: A chemical reaction can release or absorb heat.

    - **Example:** The burning of wood releases heat, indicating an exothermic reaction.

3. **Formation of a gas**: Bubbles or the release of gas (without boiling) suggests a chemical change.
    - **Example:** The reaction between vinegar (acetic acid) and baking soda (sodium bicarbonate) produces carbon dioxide gas.
4. **Formation of a precipitate**: A solid that forms from a liquid mixture is a sign of a chemical change.
    - **Example:** Mixing silver nitrate and sodium chloride forms a solid silver chloride precipitate.
5. **Production of light or sound**: Some chemical reactions release energy in the form of light or sound.
    - **Example:** Fireworks produce both light and sound due to the rapid chemical reactions inside them.

III. Classification of Matter

Matter can be classified based on its physical and chemical properties. The two primary categories are **pure substances** and **mixtures**. Each category has subcategories that help us better understand the nature of matter.



a)- Pure Substances:

A **pure substance** consists of only one type of particle (either an element or a compound) and has a uniform composition throughout. Pure substances have consistent physical and chemical properties and can be further divided into:

1. Elements: Elements are pure substances that consist of only one type of atom. They cannot be broken down into simpler substances by chemical means.

- **Example**: Oxygen (O), Hydrogen (H), Gold (Au), and Carbon (C).

2. Compounds: Compounds are pure substances composed of two or more elements chemically combined in a fixed ratio. Compounds can be broken down into simpler substances by chemical reactions.

- **Example**: Water (H_2O), Carbon Dioxide (CO_2), Sodium Chloride (NaCl).

b)- Mixtures :

A **mixture** consists of two or more substances (elements, compounds, or both) that are physically combined, but not chemically bonded. Mixtures can vary in their composition and properties. They can be separated into their individual components by physical means. Mixtures are divided into:

1. Homogeneous Mixtures (Solutions) : A homogeneous mixture has a uniform composition throughout. The different components are mixed at the molecular level, making it impossible to distinguish between them visually.

- **Example**: Salt water , Vinegar.

2. Heterogeneous Mixtures : A heterogeneous mixture has a non-uniform composition, and the different components can be distinguished visually or with a microscope.

- **Example**: Sand and Iron filings, Oil and Water.

IV- Concept of Atom, Molecules, Mole, and Avogadro's Number

These are fundamental concepts in chemistry that help us understand the structure of matter and quantify the amount of substances in reactions and compounds.

1. Atom :

- **Definition**: **Atoms** are the fundamental units of matter. An **atom** is the smallest unit of matter that retains the properties of an element. It is the basic building block of all substances.
- **Structure of the Atom**:
  - Atoms consist of a **nucleus** (containing **protons** and **neutrons**) and **electrons** that orbit the nucleus in energy levels.
    1. **Protons**: Positively charged particles located in the nucleus (center) of the atom.
    2. **Neutrons**: Neutral particles also found in the nucleus.
    3. **Electrons**: Negatively charged particles that orbit the nucleus in electron clouds or shells.

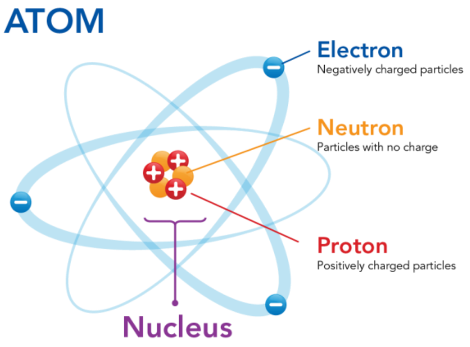


Figure I-2 : Atomic model of a lithium atom which includes the protons, neutrons and electrons

  - **Lithium atom (Li)**: Simplest atom with 3 protons and 3 electrons.

2. Molecule

- **Definition**: A **molecule** is a group of two or more atoms bonded together. It is the smallest unit of a compound that retains the chemical properties of that compound.
- **Types of Molecules**:

    - **Molecules of elements**: Consist of two or more atoms of the same element bonded together.
        - Example: Oxygen molecule (O_2), Nitrogen molecule (N_2).
    - **Molecules of compounds**: Consist of atoms of different elements bonded together.
        - Example: Water molecule (H_2O), Carbon dioxide molecule (CO_2).
- **Bonding**:
    - Atoms in a molecule are held together by chemical bonds, which can be **covalent** (sharing electrons) or **ionic** (transfer of electrons).

3. Mole

- **Definition**: The **mole** is a fundamental unit in chemistry used to measure the amount of a substance. One mole of any substance contains the same number of entities (atoms, molecules, ions, etc.) as there are in 12 grams of carbon-12.
- **Mole Formula**:
- Number of moles n

$$n = \frac{mass\ of\ substance(g)}{molar\ mass\ (g/mol)}$$

4. Avogadro's Number

- **Definition**: **Avogadro's number (N_A)** is the number of atoms, molecules, or other particles in one mole of a substance. It is a fixed constant with a value of approximately: $N_A=6.022\times10^{23}$
- **Meaning**: One mole of any substance contains 6.022×10^{23} atoms, molecules, ions, or other particles.
- **Examples**:
    - 1 mole of **carbon atoms** contains 6.022×10^{23} carbon atoms.
    - 1 mole of **water molecules** contains 6.022×10^{23} water molecules.

Example Problem: Calculating the Number of Molecules in a Sample

If you have 18 grams of water (H_2O), how many water molecules are in the sample?

1. **Step 1**: Calculate the molar mass of water.
    - $H_2O = (2 \times 1) + 16 = 18$ g/mol.
2. **Step 2**: Calculate the number of moles of water.
    - $n = \text{mass} / \text{molar mass} = 18\text{ g} / 18\text{ g/mol} = 1\text{ mol}$
3. **Step 3**: Multiply the number of moles by Avogadro's number to find the number of molecules.
    - $1\text{ mol} \times 6.022 \times 10^{23} = 6.022 \times 10^{23}$ molecules of water.

So, 18 grams of water contains 6.022×10^{23} water molecules.

5- Atomic mass units (amu)

Atomic mass units (amu) are a standard unit of mass used to express atomic and molecular weights. One atomic mass unit is defined as one twelfth the mass of a carbon-12 atom, which is approximately equal to:

$1\text{ amu} = 1.66053906660 \times 10^{-27}\text{ kg}$

V- The Law of Conservation of Matter

This fundamental principle states that matter cannot be created or destroyed in a closed system. During any physical or chemical change, the total mass of the matter remains constant. This principle is vital in understanding chemical reactions.

Key Points of the Law of Conservation of Matter

1. **Matter is Conserved**: During any physical or chemical change, the amount of matter remains constant. The mass of the substances involved in a reaction does not change, even though their forms and properties may.
2. **Closed Systems**: This law applies strictly to closed systems where no matter can enter or leave. In open systems, matter can move in and out, so conservation is considered in terms of what remains within the system.

3. **Chemical Reactions**: In a chemical reaction, atoms are rearranged to form new substances, but the total number of each type of atom remains the same. This means the total mass of the reactants before the reaction is equal to the total mass of the products after the reaction.

Mathematical Representation

In a chemical equation:
Total Mass of Reactants= Total Mass of Products
For example, in the reaction where hydrogen gas reacts with oxygen gas to form water:

$$2H_2+O_2\rightarrow 2H_2O$$

- Reactants: 2 molecules of hydrogen (H_2) and 1 molecule of oxygen (O_2)
- Products: 2 molecules of water (H_2O)

The number of atoms of hydrogen and oxygen remains the same on both sides of the equation:

- Hydrogen: 4 atoms on both sides (2 × 2 in H_2)
- Oxygen: 2 atoms on both sides (2 in O_2 and 2 in H_2O)

VI- Qualitative and Quantitative Aspects of Matter

Understanding matter involves examining both its qualitative and quantitative aspects. These aspects help describe and measure the properties of matter, allowing scientists and researchers to identify substances and predict their behaviors in different conditions.

1- Qualitative Aspects of Matter

Qualitative aspects of matter refer to the observable characteristics and properties of a substance that do not involve measurements or numerical values. These properties describe the qualities of a substance and can be perceived through the senses (sight, Odor, touch, Color, Texture, Solubility, etc.).

- **Color**: Like blue green and yellow.
- **Texture**: soft, rough.
- **Temperature** is another property of the matter like cold or hot.
- **Weight**: A material can be heavy, or it can be light.
- **Hardness**: Hard or soft.

- **Shape**: Materials can have different shapes: For example it can be round, square or rectangular.
- **Size**: It can be big or small
- **Strenght** is another of the physical properties of matter. A material can be strong or weak.
- **Flexibility**: Rigid versus Flexible.
- Qualitative properties are often used for the preliminary identification of substances.

2. Quantitative Aspects of Matter

Quantitative aspects of matter involve measurable characteristics that can be expressed with numbers and units. These properties describe matter in terms of its size, amount, or behavior under specific conditions. Quantitative properties allow for precise and objective measurements, essential for scientific experiments and calculations. They are crucial for predicting the behavior of substances, calculating reaction yields, and determining concentrations in solutions

Examples of Quantitative Properties:

1. **Mass**: The amount of matter in an object, usually measured in grams (g) or kilograms (kg).
2. **Volume**: The amount of space occupied by an object, measured in liters (L), cubic centimeters (cm^3), etc.
3. **Density**: The mass per unit volume of a substance, typically expressed in grams per cubic centimeter (g/cm^3).
4. **Melting Point**: The temperature at which a substance changes from a solid to a liquid (°C).
5. **Boiling Point**: The temperature at which a substance changes from a liquid to a gas (°C).
6. **Pressure**: The force exerted by a substance per unit area, measured in pascals (Pa) or atmospheres (atm).
7. **Molecular Mass** (molecular weight): The mass of a single molecule of a substance, measured in atomic mass units (amu). When molecular mass is expressed in grams per mole (g/mol), it is known as the molar mass and is used to relate the mass of a substance to the number of moles.

8. **Molarity (M): Molarity** is a measure of the concentration of a solute in a solution. It is defined as the number of moles of a solute dissolved in one liter of solution. Molarity is a commonly used unit in chemistry for expressing concentrations in a straightforward and precise manner.
    - The unit for molarity is **moles per liter (mol/L)**, often represented simply as **M**. (exemple : "1 M" solution contains 1 mole of solute in 1 liter of solution.

The formula to calculate molarity (M) is:

$$\mathrm{M} = \frac{n}{V} \qquad \textit{Equation 1}$$

Where:

- M = Molarity (moles per liter, mol/L)
- n = Number of moles of the solute (mol)
- V = Volume of the solution in liters (L)

Example: Calculating the Molarity of a Sodium Chloride (NaCl) Solution

- **Given**:
    - Mass of NaCl = 5.85 grams
    - Molar mass of NaCl = 58.44 g/mol
    - Volume of the solution = 0.5 liters
- **Step 1**: Calculate the number of moles of NaCl (n):

$$\mathrm{n} = \frac{mass}{molar\ mass} = \frac{5.85g}{58.44g/l} = 0.1\mathrm{mol}$$

- **Step 2**: Substitute into the molarity formula:

$$\mathrm{M} = \frac{n}{V} \qquad \mathrm{M= 0.1\ mol/\ 0.5\ L=0.2\ mol/L}$$

Molarity of the NaCl solution = 0.2 M

9. Molality (m)

Molality is another measure of the concentration of a solution, but unlike molarity, it is based on the mass of the solvent rather than the volume of the solution. Molality is defined as the number of moles of a solute per kilogram of solvent. It is particularly useful when dealing with changes in temperature and pressure, as it does not change with fluctuations in temperature since mass remains constant.

- The unit for molality is **moles per kilogram (mol/kg)**, often represented as **m** (example: "1 m" solution contains 1 mole of solute in 1 kilogram of solvent).

The formula to calculate molality (mmm) is:

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

Where:

- m = Molality (moles per kilogram, mol/kg)
- n = Number of moles of the solute (mol)
- msolvent= Mass of the solvent in kilograms (kg)

Example 1: Calculating the Molality of a Sodium Chloride (NaCl) Solution

- **Given**:
    - Mass of NaCl = 5.85 grams
    - Molar mass of NaCl = 58.44 g/mol
    - Mass of the solvent (water) = 0.5 kilograms
- **Step 1**: Calculate the number of moles of NaCl (nnn):

$$\text{n} = \frac{mass}{molar\ mass} = \frac{5.85g}{58.44g/mol} = 0.1\text{mol}$$

- **Step 2**: Substitute into the molality formula:

$$\text{m} = \frac{n}{msolvent} = \frac{0.1mol}{0.5\ kg} = 0.2\ \text{mol/Kgm}$$

Molality of the NaCl solution = 0.2 m

10. Normality (N)

Normality is a measure of concentration that is particularly useful in acid-base chemistry, redox reactions, and precipitation reactions. It is defined as the number of equivalents of a solute per liter of solution. The concept of equivalents is specific to the reaction being considered, so normality can vary depending on the context.

- **Equivalent**: An equivalent is the amount of a substance that can donate or accept one mole of reactive species (like hydrogen ions in acid-base reactions or electrons in redox reactions).
- **Normality and Chemical Reactions**:
  - **Acid-Base Reactions**: In acid-base reactions, normality measures the concentration of hydrogen ions (H^+) or hydroxide ions (OH^-) a solution can provide.
  - **Redox Reactions**: For redox reactions, it measures the number of electrons an oxidizing or reducing agent can accept or donate.
  - **Precipitation Reactions**: In these reactions, it relates to the charge of the ions involved.

The formula for normality (N) is:

$$\mathrm{N} = \frac{neq}{V}$$

Where:

- N = Normality (equivalents per liter, eq/L)
- neq = Number of equivalents of the solute (ex : number of H^+)
- V = Volume of the solution in liters (L)

Calculating the Number of Equivalents

The number of equivalents (neq) depends on the substance and the type of reaction:

- For **acids**: Number of equivalents = moles of acid ×number of replaceable H^+ ions per molecule.
- For **bases**: Number of equivalents = moles of base ×number of replaceable OH^- ions per molecule.

- For **salts**: Number of equivalents = moles of salt ×total positive or negative charge per formula unit.
- For **oxidizing/reducing agents**: Number of equivalents = moles of the agent ×number of electrons transferred in the reaction.

Relationship between Normality and Molarity

Normality can be related to molarity using the following formula:

$$\mathrm{N} = \mathrm{Mx\ neq}$$

Where:

- M = ***Molarity*** of the solution
- neq= Number of equivalents per mole of solute (ex : number of H^+)

Examples

Example 1: Normality of Sulfuric Acid (H_2SO_4)

- **Given**: A 1 M solution of H_2SO_4.
- **Sulfuric acid (H_2SO_4)** can donate 2 hydrogen ions (H^+ per molecule, so it has 2 equivalents per mole.
  - **Normality**:

$$\mathrm{N = M\ x\ neq = 1Mx\ 2 = 2N}$$

Normality of the H_2SO_4 solution = 2 N.

Example 2: Normality of Hydrochloric Acid (HCl)

- **Given**: A 0.5 M solution of HCl.
- **Hydrochloric acid (HCl)** can donate 1 hydrogen ion (H^+) per molecule, so it has 1 equivalent per mole.
  - **Normality**:

$$\mathrm{N = M\ x\ neq = 0.5\ Mx\ 1 = 0.5\ N}$$

- **Normality of the HCl solution** = 0.5 N.

11- Mole Fraction (χ)

The **mole fraction** is a way of expressing the concentration of a component in a mixture. It is defined as the ratio of the number of moles of a particular component to the total number of moles of all components in the mixture. Unlike molarity or molality, the mole fraction is a dimensionless quantity, meaning it has no units.

For a component A in a mixture, the mole fraction (χ_A) is given by:

$$\chi A = n_A / n_{Total}$$

Where:

- χ_A = Mole fraction of component A
- n_A = Number of moles of component A
- n_{total}= Total number of moles of all components in the mixture

If there are multiple components (e.g., A, B, and C) in the mixture, the total number of moles is:

$$\text{ntotal} = \text{nA} + \text{nB} + \text{nC} + \cdots$$

Example Calculation

Example 1: Mole Fraction in a Binary Mixture

- **Given**:
  - Number of moles of ethanol ($n_{ethanol}$ = 2 moles
  - Number of moles of water (n_{water}) = 3 moles
- **Step 1**: Calculate the total number of moles:

$$\text{n total} = \text{n ethanol} + \text{n water} = 2 + 3 = 5 \text{ moles}$$

- **Step 2**: Calculate the mole fraction of ethanol ($\chi_{ethanol}$):

$$\chi\text{ethanol} = \frac{\text{nethanol}}{\text{ntotal}} = \frac{2}{5} = 0.4$$

- **Step 3**: Calculate the mole fraction of water (χ_{water}):

$$\chi\text{water} = \frac{\text{nwater}}{\text{ntotal}} = \frac{3}{5} = 0.6$$

Sum of Mole Fractions:

$$\chi\text{ethanol} + \chi\text{water} = 0.4 + 0.6 = 1$$

13- Mass Concentration (C_m)

Mass concentration is a measure of how much mass of a solute is present in a given volume of solution. It is an important concept in chemistry, particularly in fields such as analytical chemistry, environmental science, and biochemistry, where knowing the exact mass of a substance within a specific volume is crucial. The unit of mass concentration is **grams per liter (g/L)** (It can also be expressed in other units such as (mg/mL) or (kg/m^3)).

The formula to calculate mass concentration ρ (rho) is:

$$C_m = \frac{m}{V}$$

Where:

- C_m = Mass concentration (typically expressed in units like grams per liter, g/L)
- m = Mass of the solute (in grams)
- V= Volume of the solution (in liters)

Example 1: Mass Concentration of Sodium Chloride (NaCl) Solution

- **Given**:
  - Mass of NaCl (solute) = 5 grams
  - Volume of the solution = 0.5 liters
- **Step 1**: Use the mass concentration formula:

$$C_m = \frac{m}{V} = \frac{5\ g}{0.5\ L} = 10 g/L$$

- **Mass Concentration** of the NaCl solution = 10 g/L.

12- Mass Percentage Formula

Mass percentage (also known as **weight percentage** or **% w/w**) is a way of expressing the concentration of a component in a mixture. It is defined as the ratio of the mass of the solute to the total mass of the solution, multiplied by 100 to convert it to a percentage.

The formula to calculate mass percentage (% mass) is:

$$\%\,\text{mass} = \frac{\text{msolute}}{\text{msolution}}\ x\ 100$$

Where:

- m_{solute} = Mass of the solute (in grams)
- $m_{solution}$ = Total mass of the solution (mass of solute + mass of solvent) (in grams)

Steps to Calculate Mass Percentage

1. **Determine the mass of the solute** (msolute).
2. **Determine the total mass of the solution** (msolution), which is the sum of the mass of the solute and the mass of the solvent.
3. **Apply the formula** to find the mass percentage.

Example Calculation

Example 1: Mass Percentage of Sodium Chloride (NaCl) in a Solution

- **Given**:
  - Mass of NaCl (solute) = 5 grams
  - Mass of water (solvent) = 95 grams
- **Step 1**: Calculate the total mass of the solution:

$$\text{mass solution} = \text{msolute} + \text{msolvant} = 5\text{g} + 95\text{g} = 100\text{g}$$

- **Step 2**: Apply the mass percentage formula:

$$\%\,\text{mass} = \frac{\text{msolute}}{\text{msolution}}\ \text{x}\ 100 = \left(\frac{5\text{g}}{100\text{g}}\right)\text{x}\ 100 = 5\%$$

- **Mass Percentage** of NaCl in the solution = 5%.

Exercises of chapter 1

Exercise 1:

The masses of the proton, neutron and electron are respectively: mp = 1.6726 10^{-24} g; mn = 1.6749 10^{-24} g and me = 9.109 10^{-28} g.

1- Define the atomic mass unit and give its value in kilograms.
2- Express the masses of these particles in atomic mass units (u).

Exercise 2:

1- How many atoms and molecules are there in: 6g of Fe, 6g of C and 6g of Ag at room temperature?
2- Calculate the mass (g) of: 1.52 mol of Cu, 1.52 mol of Na , 1.52 mol of Au.
3- How many moles and atoms of Fe and S are there in a 0.5 kg sample of pyrite "FeS_2".
4- Which of the following samples contains the most iron:

a) 0.2 moles of $Fe_2(SO_4)_3$. **b)** 20g of iron. **c)** 2.5x10^{23} atoms of iron

Data: Molar mass (g/mol): C = 12; O = 16; Na = 23; Cu = 63.5; S = 32; Fe = 56; Ag = 108; Au = 197.

Exercise 3 :

a) - A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% chlor. Its molar mass is 98.96 g.

1- What is its molecular formula?

M(C): 12g/mole; M(Cl): 35.5g/mole; M(H): 1g/mole

b) - Analysis of a 12.04 g sample of a liquid compound composed of carbon, hydrogen, and nitrogen showed it to contain 7.34 g C, 1.85 g H, and 2.85 g N.

1- What is the percent composition of this compound?

Exercise 4 :

1- A sample of 0.892 g of potassium chloride (KCl) is dissolved in 54.6 g of water. What is the percent by mass of KCl in solution?
2- Calculate molality of 2.5 g of ethanoic acid (CH_3COOH) in 75 g of benzene.
3- Calculate the molarity of NaOH in the solution prepared by dissolving its 5 g in enough water to form 450 mL of the solution.
4- The density of a 2.45 M aqueous solution of methanol (CH_3OH) is 0.976 g/ml. what is the molality of the solution? The molar mass of methanol is 32.04 g/mol.

M(Na): 23g/mole , M(H): 1g/mole, M(O): 16g/mole

Exercise 5 :

The commercial solution of sulphuric acid contains 80% by weight H_2SO_4. Specific gravity of this solution is 1.71.

1- Calculate the normality of this commercial solution.
2- What volume of this solution should be used to prepare 500 ml of 1N of H_2SO_4 solution?

(M.M H_2SO_4 = 98.076 g/mol)

Exercise 6:

Determine degree of acidity of a vinegar (acetic acid solution) or degree of purity of a vinegar indicated on the label:

Consider a solution of volume 1 L and a concentration of acetic acid C_A = 1.4 mol.L^{-1}

Density of vinegar: d = 1.06 , Molar mass of acetic acid: M = 60 g.mol^{-1}.

Corrected exercises of chapter 1

Exercise 1:

1- Define the atomic mass unit (amu).

The atomic mass unit (amu) is a standard unit of mass that quantifies mass on an atomic or molecular scale. It is defined as one twelfth of the mass of a carbon-12 atom.

$$1\mathrm{u} = \frac{1}{12} \text{ x } \mathrm{m}\ ({}^{12}C) \qquad \text{with:} \qquad \mathrm{m}\ ({}^{12}C) = \frac{12}{\mathrm{NA}} = \frac{12}{6.0230.10^{23}}$$

$$1\mathrm{u} = \frac{1}{12} \text{ x } \mathrm{m}\ ({}^{12}C) = \frac{1}{12} \text{ x } \frac{12}{\mathrm{NA}} = \frac{1}{12} \text{ x } \frac{12}{6.0230.10^{23}} = 1.6605 \times 10^{-24} \text{ g}$$

$$1\mathrm{u} = 1.6605 \times 10^{-27}\ Kg$$

2- Convert the atomic mass unit to kilograms.

Express the masses of the proton, neutron, and electron in atomic mass units (u).

To convert the masses from grams to atomic mass units, we will use the conversion factor:

1 amu = 1.6605×10^{-24} g= 1.6605×10^{-27} Kg

- $\mathrm{mP} = \frac{1.6726 \times 10^{-24}\ \mathrm{g}}{1.6605 \times 10^{-24}\ \mathrm{g}} = 1.007276\ \mathrm{u}$

- $\mathrm{mn} = \frac{1.6749 \times 10^{-24}\ \mathrm{g}}{1.6605 \times 10^{-24}\ \mathrm{g}} \approx 1.008665\ \mathrm{u}$

- $\mathrm{me} = \frac{9.109 \times 10^{-24}\ \mathrm{g}}{1.6605 \times 10^{-24}\ \mathrm{g}} \approx 5,4858.10^{-4}\ \mathrm{u}$

Exercise 2 :

1- How many atoms and molecules are there in: 6g of Fe, 6g of C and 6g of Ag at room temperature?

Calculate the number of moles for each element using the formula:

$$\mathrm{n\ (moles)} = \frac{\mathrm{mass\ (g)}}{\mathrm{molar\ mass\ (g/mol)}}$$

- **6g of Fe**:

n of Fe=6 /56 =0.1071 moles , N=n. N_A

N atoms of Fe=0.1071 mol×6.022×10^{23} atoms/mol=6.44×10^{22} atoms

- **6g of** C:

n of C=6 /12 =0.5 moles , N=n. N_A

$\text{N atoms of C} = 0.5\ \text{mol} \times 6.022 \times 10^{23}\ \text{atoms/mol} = 3.011 \times 10^{23}\ \text{atoms}$

- **6g of Ag**:

$\text{n Ag} = 6\ \text{g}/108\ = 0.0556\ \text{moles}$
$\text{Natoms of Ag} = 0.0556\ \text{mol} \times 6.022 \times 10^{23}\ \text{atoms/mol} = 3.34 \times 10^{22}\ \text{atoms}$

2- **Calculate the mass (g) of: 1.52 mol of Cu, 1.52 mol of Na, 1.52 mol of Au.**

$\text{mass (g)} = \text{moles} \times \text{molar mass (g/mol)}$

- ***<u>For Cu (Molar mass = 63.5 g/mol):</u>***

$\text{mass of Cu} = 1.52\ \text{mol} \times 63.5\ \text{g/mol} = 96.52\ \text{g}$

- ***<u>For Na (Molar mass = 23 g/mol):</u>***

$\text{mass of Na} = 1.52\ \text{mol} \times 23\ \text{g/mol} = 34.96\ \text{g}$

- ***<u>For Au (Molar mass = 197 g/mol):</u>***

$\text{mass of Au} = 1.52\ \text{mol} \times 197\ \text{g/mol} = 299.44\ \text{g}$

3- **How many moles and atoms of Fe and S are there in a 0.5 kg sample of pyrite “FeS2”?**

Convert 0.5 kg to grams:

$0.5\ \text{kg} = 500\ \text{g} 0.5\ \text{kg} = 500\ \text{g}$

- Calculate the molar mass of pyrite (FeS2):

$\text{Molar mass of } FeS_2 = 56\ \text{g/mol (Fe)} + 2 \times 32\ \text{g/mol (S)} = 56 + 64 = 120\ \text{g/mol}$

- Calculate the number of moles of FeS2:

$\text{n moles of } FeS_2 = 500\ \text{g}/120\ \text{g/mol} = 4.167\ \text{moles of FeS2}$

- Calculate moles of Fe and S in the sample:

$$FeS_2 \longrightarrow Fe + 2S$$

- Moles of Fe:
  - $\text{n moles of Fe} = 4.167\ \text{mol } (FeS_2)\text{moles}$,
  - $\text{atoms of Fe} = 4.167\ \text{mol} \times 6.022 \times 10^{23}\ \text{atoms} = 2.51 \times 10^{24}\ \text{atoms}$
- Moles of S:
  - $\text{n moles of S} = 2 \times n\ (FeS_2) = 2 \times 4.167\ \text{mol} = 8.334\ \text{moles}$
  - $\text{atoms of S} = 8.334\ \text{mol} \times 6.022 \times 10^{23}\ \text{atoms} = 5.02 \times 10^{24}\ \text{atoms}$

4- Samples contains the most iron

a) **0.2 moles of** Fe2(SO4)3 :

0.2 moles of Fe2(SO4)3 contains:

0.2 moles×2 moles Fe=0.4 moles of Fe.

$N= n. N_A= 0.4 . 6,023 .10^{23} = 2.40 .10^{23}$ atoms of Fe

b) 20 g of iron :

$N= n. N_A = (m/M) . N_A= (20/56) .6,0.23\ 10^{23} = 2.15\ 10^{23}$.

c) 2.5 10^{23} atoms of iron

The largest mass of iron is from sample : c) 2.5x10^{-23} atoms of iron

Exercise 3 :

a)- Molecular formula:

- % C : (n x 12) g of C $\longrightarrow$ 98.96 g/mol
  % C : 24.27 $\longrightarrow$ 100 g

  (n x 12) x 100 = 24.27 x 98.96 $\Rightarrow$ n = 2

- % H : (n x 1) g of H $\longrightarrow$ 98.96 g
  % H : 4.07 $\longrightarrow$ 100 g

  (n x 1) x 100 = 4.07 x 98.96 $\Rightarrow$ n = 4

- % Cl : (n x 35.5)= 96 g of Cl $\longrightarrow$ 98.96 g
  % Cl ; 71.65 $\longrightarrow$ 100 g

  (n x 35.5) x 100 = 71.65 x 98.96 $\Rightarrow$ n = 2

Molecular formula: $C_2H_4Cl_2$

b)- Analysis of a 12.04 g sample of a liquid compound composed of carbon, hydrogen, and nitrogen showed it to contain 7.34 g C, 1.85 g H, and 2.85 g N.

1- What is the percent composition of this compound?

- ➢ Calculate the percentage of carbon in the compound.
  Percentage of carbon = (mass of carbon / mass of compound) x 100%
  Percentage of carbon = (7.34 g / 12.04 g) x 100% = 61.0%
- ➢ Calculate the percentage of hydrogen in the compound.
  Percentage of hydrogen = (mass of hydrogen / mass of compound) x 100%
  Percentage of hydrogen = (1.85 g / 12.04 g) x 100% = 15.4%
- ➢ Calculate the percentage of nitrogen in the compound.
  Percentage of nitrogen = (mass of nitrogen / mass of compound) x 100%
  Percentage of nitrogen = (2.85 g / 12.04 g) x 100% = 23.7%

Exercise 4:

5- A sample of 0.892 g of potassium chloride (KCl) is dissolved in 54.6 g of water. What is the percent by mass of KCl in solution?.

1- Percent by mass of KCl :

$$\text{Percent by mass} = \left(\frac{mass\ of KCl}{total\ mass\ of\ solution}\right) x\ 100$$

Calculate the total mass of the solution

$$Total\ mass\ of\ solution = mass\ of KCl + mass\ of\ water$$

$$Total\ mass\ of\ solution = 0.892 + 54.6 = 55.492$$

$$\text{Percent by mass} = \left(\frac{0.892}{55.492}\right) x\ 100$$

$$\text{Percent by mass} = 1.606\ \%$$

2- Molality of CH_3COOH

Calculate <u>molality</u> of 2.5 g of ethanoic acid (CH_3COOH) in 75 g of benzene

$$\text{Molality m} = \left(\frac{Moles\ of\ \text{solute}}{mass\ of\ solvent\ (Kg)}\right)$$

Molar mass of $C_2H_4O_2$: 12 x2 + 1 x 4 + 16 x 2 = 60g/mol

$$\text{Moles of C2H4O2} = \left(\frac{2.5}{60}\right) = 0.0417\ mol$$

Mass of benzene in Kg = 75 g/1000 g.Kg^{-1} = 75 .10^{-3} Kg

$$\text{Molality of C2H4O2} = \left(\frac{Moles\ of\ \text{C2H4O2}}{m\ \ of\ bezene}\right) = \frac{0.0417}{75.10^{-3}}$$

$$\text{Molality of C2H4O2} = 0.556\ mol/kg$$

3- Molarity of NaOH :

$$\text{Moles NaOH} = \left(\frac{4g}{40\ g/mol}\right) = 0.125\ mol$$

$$\text{Molarity of NaOH} = \frac{0.125}{450.10^{-3}} = 0.278\ \ M$$

4- Density of CH_3OH :

The density of a 2.45 M ***<u>aqueous</u>*** solution of methanol (CH_3OH) is 0.976 g/ml. what is the molality of the solution? The molar mass of methanol is 32.04 g/mol.

$$\text{Molality m} = \left(\frac{Moles\ of\ \text{solute}}{mass\ of\ solvent\ (Kg)}\right)$$

We need to find : 1) – moles of solute (CH_3OH)

2) – mass of solvent in Kg (water)

1) 2.45 M means the molarity of solution = 2.45 mol/l, **moles of solute (**CH_3OH) = 2.45 mol

2) – mass of solvent in Kg (water) :

$$n \text{ CH3OH} = \left(\frac{m}{M}\right)$$

$$m \text{ CH3OH} = n \text{ CH3OH x M} = 2.45 \text{ x } 32.04 = 78.54\ g$$

The density = 0.976 g/ml

Mass of solution = 0.979 x 1000 = 979 g = mass of solute + mass of water

Mass of water (solvent) = mass of solution – mass of solute

= 979 – 78.54 = 897.46 g = 0.89746 Kg

$$\text{Molality m} = \left(\frac{Moles\ of \text{ solute (moles)}}{mass\ of\ solvent\ (Kg)}\right) = \left(\frac{2.45}{0.89746}\right) = 2.73$$

Molality = 2.73 m

Exercise 5 :

1- ***Calculate the normality of the solution.***

- 80% by weight means the mass of H_2SO_4 in 100 g of the solution is 80 g
- The specific gravity of the solution is 1.71, which means that the density of the solution is 1.71 g/mL

Method 1:

Density = mass/ Volume, so , Mass= Density x Volume = 1.71 x 1000

= 1710 g in 1 L

1 L of commercial solution contains 1710 g of solution

- Calculate de mass of H_2SO_4 in this commercial solution :

80 g of H_2SO_4 $\longrightarrow$ 100 g of the solution

m H_2SO_4 $\longrightarrow$ 1710 g of solution

$$\text{m H2SO4} = \frac{1710 \text{ x } 80}{100} = 1368\ g$$

Calculate the molarity of the solution.

The molarity of the solution is the number of moles of H_2SO_4 per liter of solution.

$$\text{n H2SO4} = \frac{\text{m H2SO4}}{\text{M H2SO4}} = \frac{1368}{98.076} = 13.94\ \text{M}$$

Method 2: General formula :

$$\text{C H2SO4} = \frac{10 \text{ x d x } \%}{\text{M}}$$

$$\text{C H2SO4} = \frac{10 \text{ x } 1.71 \text{ x } 80}{98.076}$$

The molar concentration of H_2SO_4 = 13.94 M

5- *Calculate the normality of the solution.*

The normality of the solution is the number of equivalents of H_2SO_4 per liter of solution.

Since H_2SO_4 has two acidic protons, its normality is twice its molarity.

$$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$$

Normality of the solution = 2 x Molarity = 2 x 13.93 N = 27.86 N

2- *Calculate the volume of the commercial solution needed to prepare 500 mL of 1 N H_2SO_4 solution*.

We can use the following formula: $N_1 \text{ x } V_1 = N_2 \text{ x } V_2$

where:

N_1 , V_1= Normality and Volume of the commercial solution

N_2 , V_2= Normality and Volume of the desired solution

Substituting the values, we get:

27.86 N x V_1 = 1 N x 500 mL

V_1 = (1 N x 500 mL) / 27.86 N = 17.94 mL

V_2= 17.94 mL of the commercial solution should be used to prepare 500 mL of 1 N

Exercise 6 :

- Degree of acidity of a vinegar or degree of purity of a vinegar

This is the mass purity of the solution, i.e. the mass of solute (acetic acid) in 100 g of solution (vinegar).

1- Calculate the mass of 1 L of vinegar using the density:

Mass of vinegar = d x V = 1.06 g/mL x 1000 mL = 1060 g

1 L weight 1060 g

2- Calculate the mass of acetic acid in 1 L of vinegar using the concentration and molar mass:

$$\text{n acetic acid} = \frac{\text{m acetic acid}}{\text{M acetic acid}}$$

in 1 L of solution $C_A = 1.4\ mol.L^{-1}$

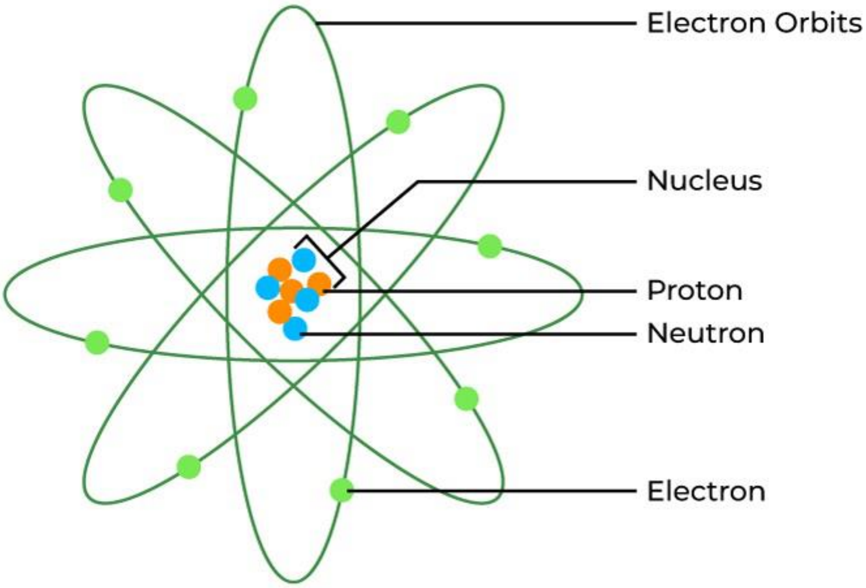
$$\text{m acetic acid} = \text{n acetic acid x M acetic acid} = 1.4 \text{ x } 60 = 84 \text{ g}$$

Calculate the mass purity of acetic acid in the vinegar:

Mass purity = (Mass of acetic acid / Mass of vinegar) x 100% = (84 g / 1060 g) x 100% = 7.92.

Chapter 2

Structure of the atom



I. Electron: J.J. Thomson's experiment, Properties of cathode rays:

JJ Thomson was interested in knowing about other subatomic particles other than the protons. Thomson's experiment, commonly known as the **cathode ray experiment**, was a groundbreaking series of experiments conducted by **J.J. Thomson** in 1897 century. It led to the discovery of the **electron**, one of the most important findings in physics.

Various experiments demonstrated proprieties of Cathode ray:

1. JJ Thomson created a sealed cathode ray tube with minimal air inside.

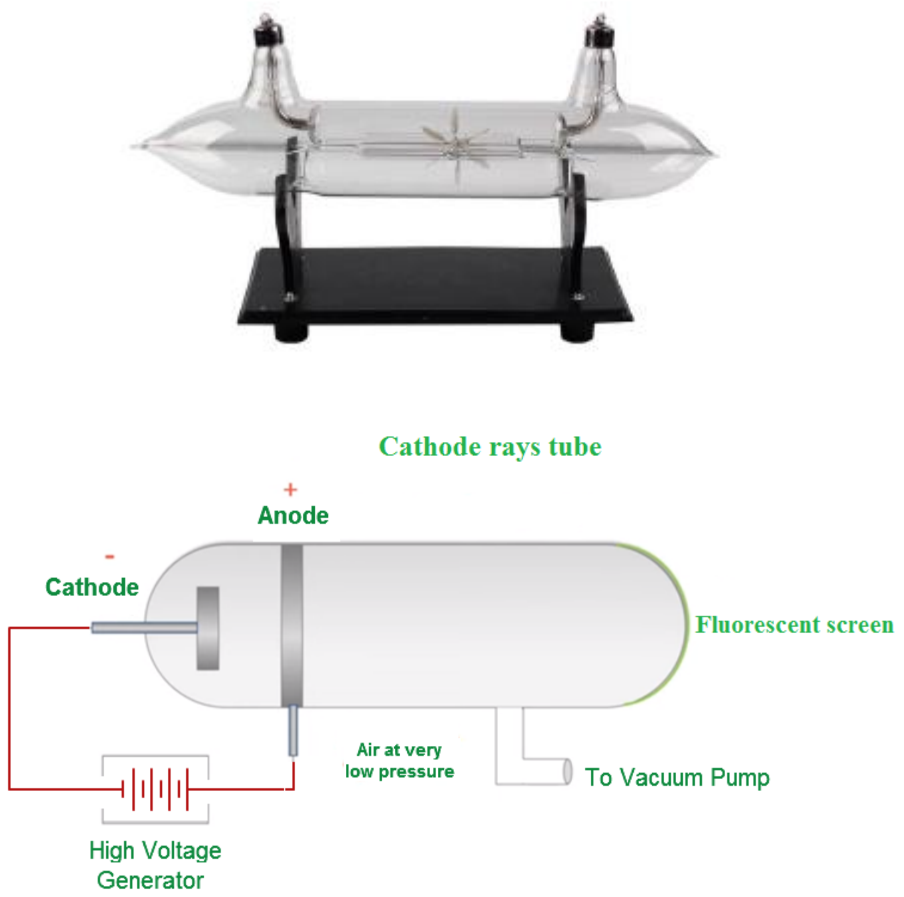
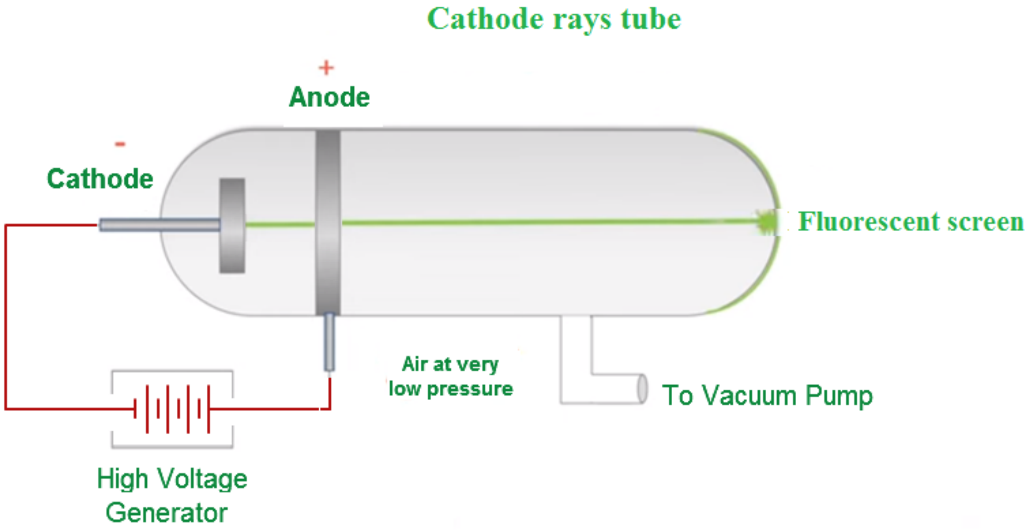


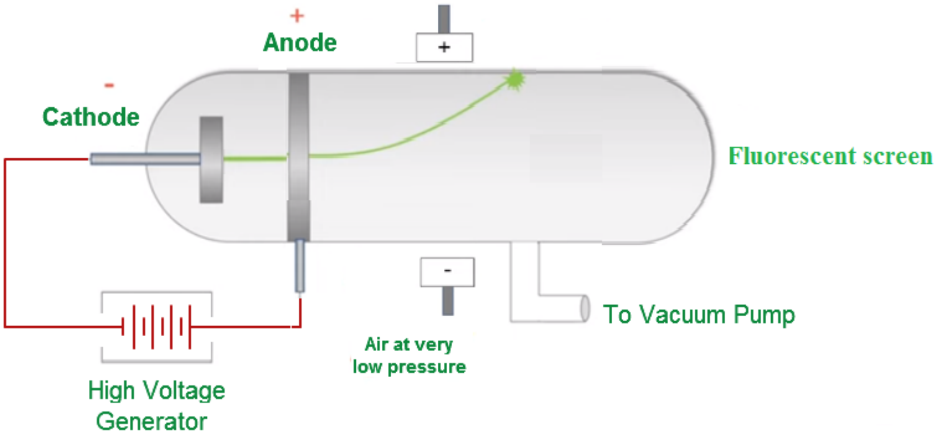
***Figure II- 1 :** The apparatus to determine the charge to the mass ratio of electron*

- A high voltage generator is connected to the anode and cathode. When the Generator is switched on, Rays are formed from the cathode. So, the rays are called cathode rays.
- A vacuum pump is connected to the tube because when the air pressure is low we can easily create cathode rays.

2. Connected the tube to a power source, causing electrons (cathode rays) to shoot out.
   - Observed electrons moving in straight lines inside the vacuum of the tube.

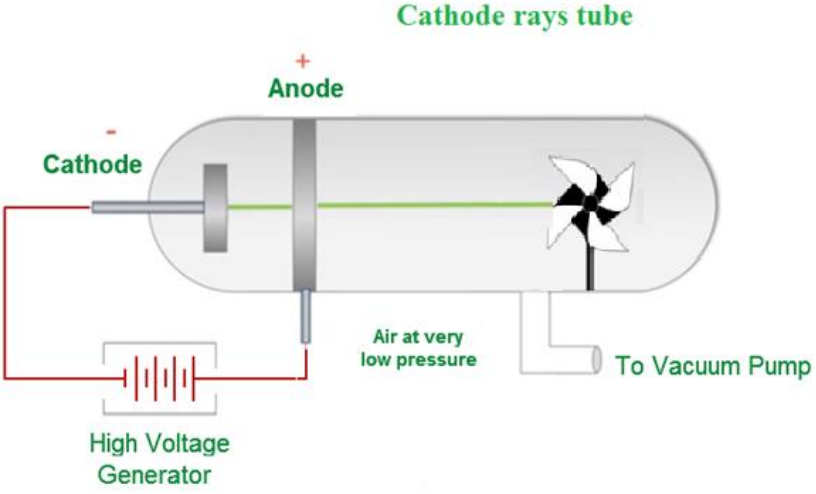


4- JJ Thomson was interested in knowing about these rays. So, he put 2 oppositely charged plates next to each other in the tube.



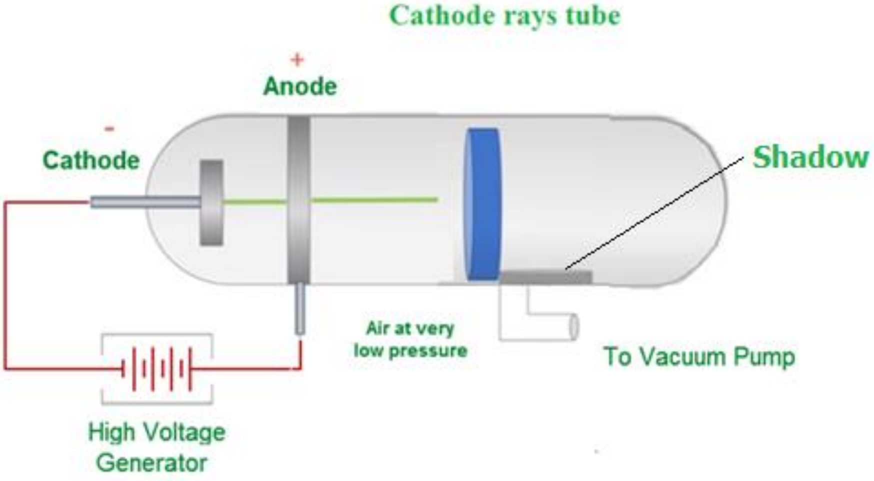
- To his surprise the cathode rays started diverting to the positively charged plate. Then he made a note that the rays are negatively charged because, positive is attracted to the negatively charged bodies vice versa.

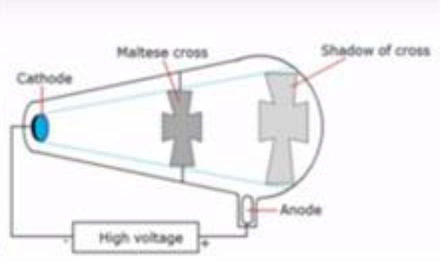
5- J.J. Thomoson thought for a while and was interested whether these rays are made up of particles. So, he kept a pinwheel in that tube.



- The pinwheel started rotating. So, J.J. J.J. Thomoson made a note that these are negatively charged .He also discovered that these were subatomic particles.
- Cathode rays have momentum and kinetic energy. Therefore, they have mass and are particles in nature.

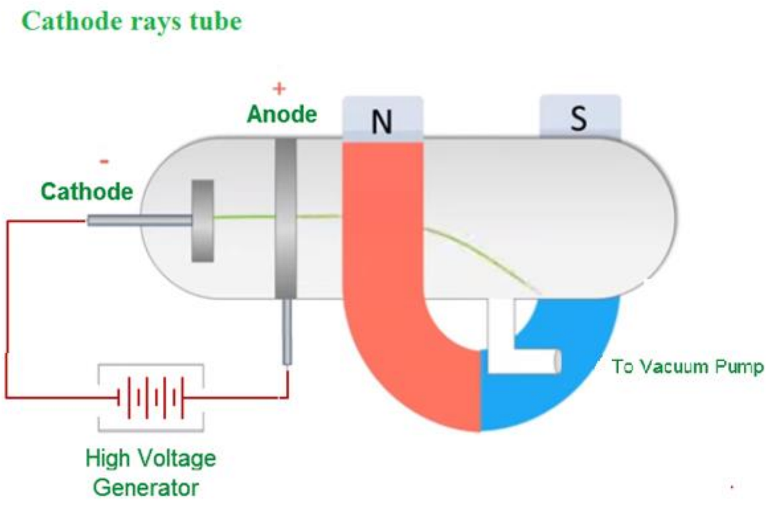
4- A shadow of shape was formed directly behind the anode.





- Cathode rays travel in a straight line and can cast a shadow. Some scientists argued that since waves such as light can produce a similar observation, cathode rays are wave in nature. However, this observation could also be produced by particles.
- Particles are moving from cathode to anode .

3. Experimented with magnets, observing electrons being affected and swerving in response**.**



- Thompson designed an experiment using magnetism to move the electrons in the opposite direction caused by the electric field.
  - This allowed him to calculate a charge to mass ratio for the electrons.

➢ ***With all this evidence, Thompson concluded that cathode rays are made of negatively charged particles called “electrons”.***

Thomson proposed that the shape of an atom resembles that of a sphere having a radius of the order of 10^{-10} m. The positively charged particles are uniformly distributed with electrons arranged in such a manner that the atom is electrostatically stable. Thomson’s atomic model was also called ***the plum pudding model***. The plum pudding atomic theory assumed that the mass of an atom is uniformly distributed all over the atom.

In this model, the atom was believed to consist of a positive material “pudding” with negative “plums” distributed throughout (Figure II-2).

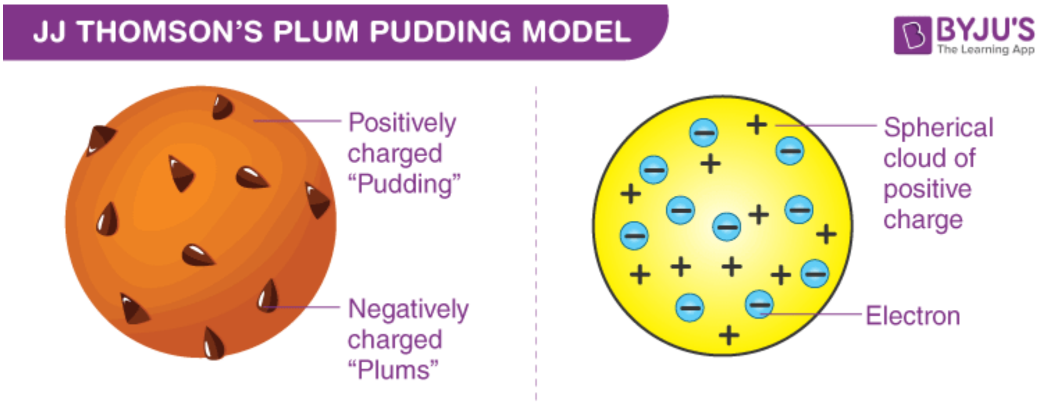
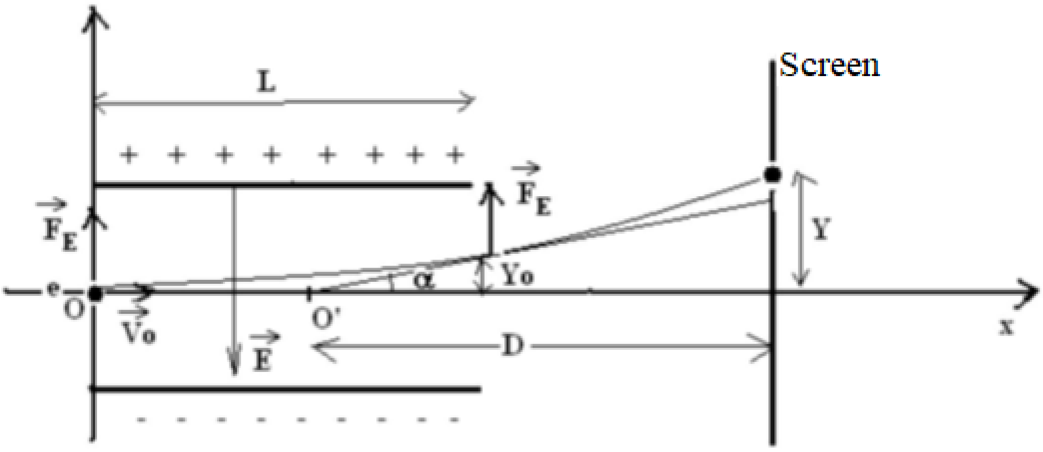


Figure II-2: Thomson's model of an atom

II- Measurement of Charge-to-Mass (e/m) Ratio for the Electron

J.J. Thomson first measured the charge-to-mass ratio of the fundamental particle of charge in a cathode ray tube in 1897. In this experiment you will measure *e*/*m*, the ratio of the charge of an electron to the mass of an electron. The currently accepted value for *e*/*m* is 1.758820 × 10^{11} C/kg.



Preconceptions:

- ✓ *x* : position
- ✓ Velocity is the rate of change of position with respect to time. The velocity function is found by taking the derivative of the position function.

$$v = \frac{dx}{dt}$$

- ✓ Acceleration is the rate of change of velocity with respect to time. The acceleration function is found by taking the derivative of the velocity function.

$$\gamma = \frac{dv}{dt} = \frac{d^2x}{dt^2}$$

- ✓ Newton's law :

$$\sum \vec{F} = m\,\vec{\gamma}$$

Along x-axis :

The force applied to the electron in the direction of the **ox** axis is non-existent:

$$F = m_e\ \gamma_x = m_e\frac{d^2x}{dt^2} = 0 \Rightarrow \frac{dx}{dt} = v_0$$

$$Fx = 0 \Rightarrow \gamma_x = 0 \Rightarrow \frac{d^2x}{dt^2} = 0 \Rightarrow \frac{dx}{dt} = cte$$

$$\frac{dx}{dt} = v \Rightarrow dx = vdt$$

By integration we find :

$$x = vt + c$$

At the moment of time and at the origin $c = 0$

$$\Rightarrow x = v_0 t$$

$\Rightarrow \quad t = \frac{x}{v_0}$ ....(1)

It is the time equation of the movement of the electron towards the **ox** axis and indicates that the movement towards ox is regular.

Along y-axis :

$$F = m_e \ \gamma_y = m_e \frac{d^2y}{dt^2} = e.E \Rightarrow \frac{d^2y}{dt^2} = \frac{e.E}{m_e}$$

$$\int \frac{d^2y}{dt^2} = \int \frac{e.E}{m_e} \Rightarrow \frac{dy}{dt} = \frac{e.E}{m_e}.t = v_y$$

$$dy = \frac{e.E}{m_e}.tdt \Rightarrow \int dy = \frac{e.E}{m_e}\int tdt$$

$$\Rightarrow \quad y = \frac{1}{2}\frac{e.E}{m_e}t^2 \quad \text{.................(2)}$$

From equation (1) $\quad t = \frac{x}{v_0} \quad$, we replace in equation (2)

$$\Rightarrow \quad y = \frac{1}{2}\frac{e.E}{m_e}\left(\frac{x}{v_0}\right)^2$$

X= l

$$\Rightarrow \quad y = \frac{1}{2}\frac{e.E}{m_e}\frac{l^2}{{v_0}^2}$$

This is the deviation that the electron undergoes upon exiting the electric field.

When applying the magnetic field, a force arises that deflects the electron downward, known as the Laplace force.

$$F' = q\, v_0\, B$$

In order for the magnetic force F' to be equal to the electric force F it must be :

$$F' = -F = -(-eE)$$

$$q\, v_0\, B = +e\, E \;\rightarrow\; v_0 = \frac{E}{B}$$

Compensation in the previous relationship

$$\Rightarrow \qquad y = \frac{1}{2}\,\frac{e.E}{m_e}\,\frac{l^2}{v_0{}^2} = \frac{1}{2}\,\frac{e.E}{m_e}\,\frac{l^2}{(\frac{E}{B})^2}$$

$$\Rightarrow \qquad y = \frac{1}{2}\,\frac{e}{m_e}\,\frac{l^2.B^2}{E}$$

$$\Rightarrow \frac{e}{m_e} = \frac{2\,y}{l^2}\,\frac{E}{B^2}$$

Thomson was able to determine the value of *e/m*e as:

$$\frac{e}{m_e} = 1.758820 \times 10^{11}\ \mathrm{C\,kg}{-}1$$

III. Determining the Charge of the Electron: Millikan Experiment:

R.A. Millikan (1868-1953) devised a method known as oil drop experiment , to determine the charge on the electrons. He found that the charge on the electron to be – 1.6 × 10^{-19} C. The present accepted value of electrical charge is – 1.6022 × 10^{-19} C. The mass of the electron (*m*e) was determined by combining these results with Thomson's value of *e*/*m*e ratio

$$m_e = \frac{e}{e/m_e} = \frac{1.6022\times 10^{-19}\,\mathrm{C}}{1.758820\times 10^{11}\,\mathrm{C\,kg^{-1}}}$$
$$= 9.1094\times 10^{-31}\,\mathrm{kg}$$

The apparatus used is illustrated in the following diagram and consists of the following components:

1. A chamber containing air confined between charged plates of a capacitor.
2. An X-ray source (RX) on one side of the chamber.
3. One of the capacitor plates (the upper one) is perforated to allow oil droplets to pass into the ionization chamber.
4. A spray for emitting small oil droplets.
5. A microscope to observe the falling droplets and monitor their speed.

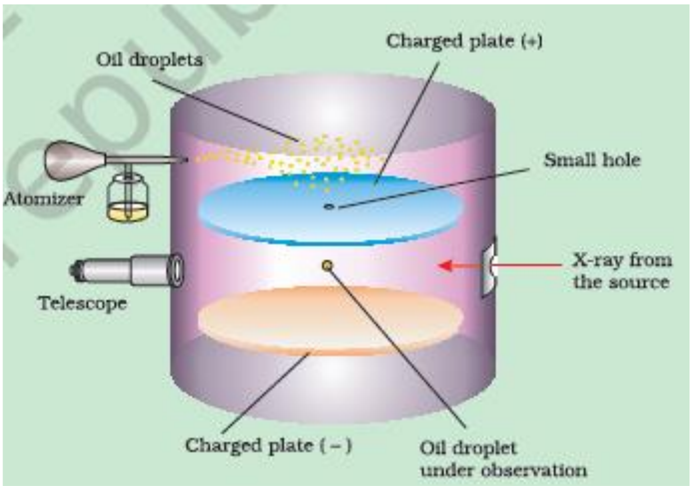


Figure II-3 : The Millikan oil drop apparatus for measuring charge ‘e’.

1. In the absence of an electric field E:

The droplet is subjected to three forces:

1- **The gravitational force (F_G)** (قوة الثقل)

2- **The air resistance force (stokes force) (F_S)** (قوة مقاومة الهواء)

3- **The Archimedean buoyant force (F_A)** (دافعة ارخميدس)

a. The gravitational force, **F_G=mg**,

$$F_G = m.g \Rightarrow \quad F_G = \rho_o.\ V.\ g$$

where:

m: is the mass of the droplet.

g: is the gravitational acceleration.

ρ: is the volume mass density of the droplet (oil).

V: is the volume of the droplet.

And since the droplet has a spherical shape, its shape does not change during movement.

$$F_G = \rho_o.\frac{4}{3}.\pi.r^3.g$$

Where r: is the radius of the sphere.

b. The air resistance force (stokes force) (FS)

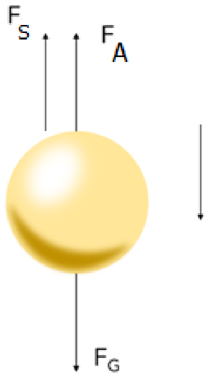
$$F_s = 6\pi.\eta.r.v_0$$

Where: V_0 : Droplet velocity n : Air viscosity coefficient

c. The Archimedean buoyant force:

$$F_A = v.\ \rho_a.g = \frac{4}{3}.\pi.r^3.\rho_a.g$$

Where: ρa : Volume mass density of air



The sum of these forces = 0

$$\vec{F}_G + \vec{F}_S + \vec{F}_A = 0$$

$$F_G - F_S - F_A = 0$$

$$\rho_o.\frac{4}{3}\pi.r^3 g - 6\pi.\eta.r.v_0 - \frac{4}{3}\pi.r^3.\rho_a.g = 0$$

$$\Rightarrow \quad \frac{4}{3}\pi.r^3 g.\left(\rho_o - \rho_a\right) = 6\pi.\eta.r.v_0$$

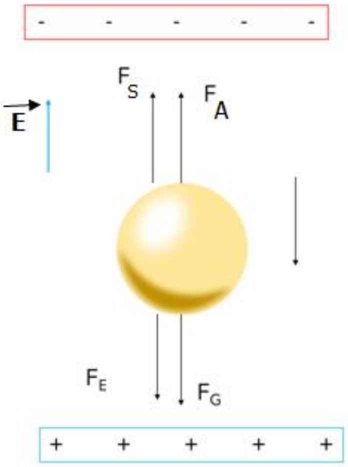
$$\Rightarrow \quad \frac{4}{3}r^2.g.\left(\rho_o - \rho_a\right) = 6.\eta.v_0$$

$$\Rightarrow v_0 = \frac{\frac{4}{3}r^2 g.(\rho_o - \rho_a)}{6\eta} \qquad \textit{equation (1)}$$

2- In the presence of the electric field E:

In this case, we observe the upward and downward movements of the droplets because they become charged due to collisions with air molecules ionized by X-rays. Consequently, each droplet is exposed, in addition to the previous three forces, to the electric Coulomb force.

$$F_e = q.E$$



The previous equation can be written as follows:

$$F_G + F_E - F_S - F_A = 0$$

$$\rho_o . \frac{4}{3}\pi . r^3 g + q.E - 6\pi\eta r v_1 - \frac{4}{3}\pi . r^3 . \rho_a . g = 0$$

$$\Rightarrow \quad \frac{4}{3}\pi . r^3 g . (\rho_o - \rho_a) + q.E = 6\pi\eta r v_1$$

$$\Rightarrow \quad v_1 = \frac{\frac{4}{3}\pi . r^3 g . (\rho_o - \rho_a)}{6\pi\eta r} + \frac{q.E}{6\pi\eta r}$$

$$\Rightarrow \quad v_1 = \frac{\frac{4}{3} r^2 g . (\rho_o - \rho_a)}{6\eta} + \frac{q.E}{6\pi\eta r}$$

In absence of the electric field E, we have : $v_0 = \frac{\frac{4}{3} r^2 g . (\rho_o - \rho_a)}{6\eta}$

$$\Rightarrow \quad v_1 = v_0 + \frac{q.E}{6\pi\eta r}$$

$$\Rightarrow \quad (v_1 - v_0) = \frac{q.E}{6\pi\eta r}$$

$$\Rightarrow \quad q = (v_1 - v_0) . \frac{6\pi . \eta . r}{E} \qquad \textit{equation (2)}$$

The result is that the calculated value q for each droplet does not change except in multiples of the elementary charge, which is the charge of an electron (e). The charge of an electron is the fundamental unit of electric charge.

$$q = e = -1.6 \, x \, 10^{-19} \, coulomb$$

$$m_e = 9.109 \, x \, 10^{-31} \, Kg$$

IV. *Nucleus: Rutherford's experiment, Constitution of the atomic nucleus(Discovery of the Nucleus:1911)*

Before Rutherford’s experiment, the best model of the atom that was known to us was ***<u>the Thomson or “plum pudding” model.</u>***

Later, Rutherford’s alpha-particle scattering experiment changed our perception of the atomic structure. Rutherford directed beams of alpha particles (positive charge) at thin gold foil to test this model and noted how the alpha particles scattered from the foil.

The following figure illustrates the experimental setup:

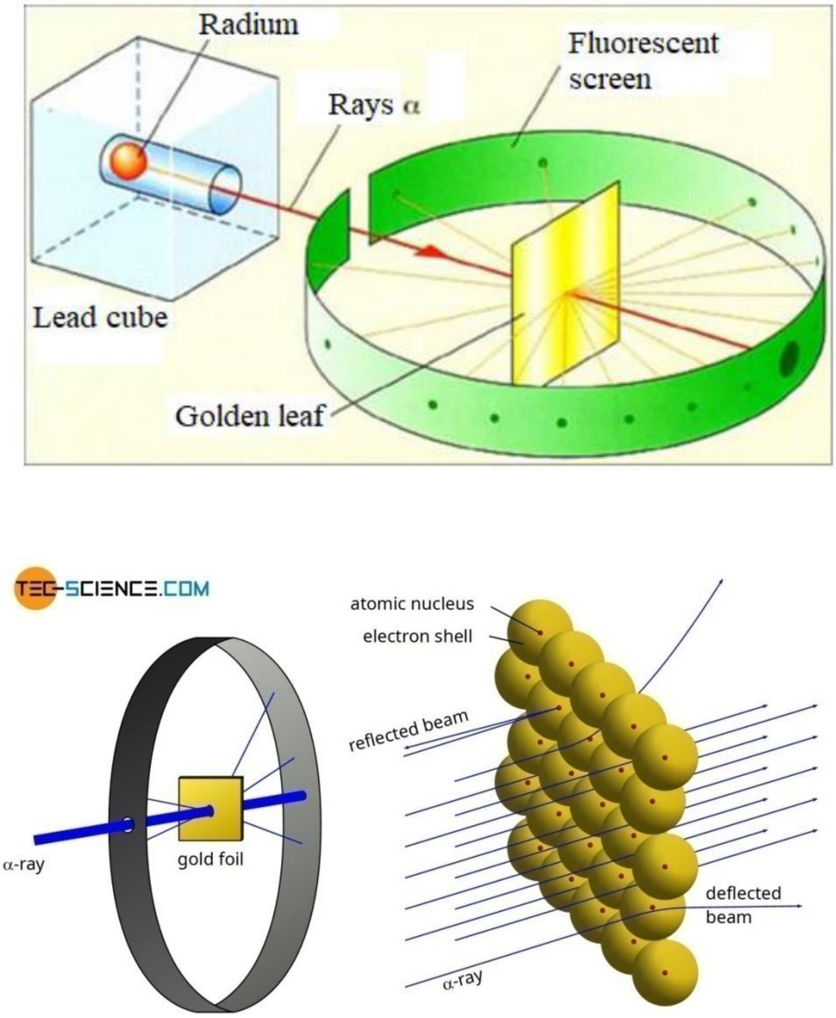


Figure II-4 : Rutherford’s gold foil experiment

Through this experiment, Rutherford made 3 observations as follows:

1) - Highly charged alpha particles (positive charge) went straight through the foil undeflected.

2) - Some alpha particles were deflected back through large angles.
3) - A very small number of alpha particles were deflected backwards

Rutherford's conclusions:

Observation	*Conclusion*
1- Some α particles (positive charge) were deflected by large angles	- There is a concentration of positive charge "the nucleus"
2- Most particles were undeflected	- Most of the atom is empty space
3- Very rarely was the α particle deflected back .	- The nucleus must be very small .

Rutherford atomic model envisioned the atom as ***a miniature solar system***, with electrons orbiting around a massive nucleus, and as mostly empty space, with the nucleus occupying only a very small part of the atom. The neutron had not yet been discovered when Rutherford proposed his model, which had a nucleus consisting only of protons.

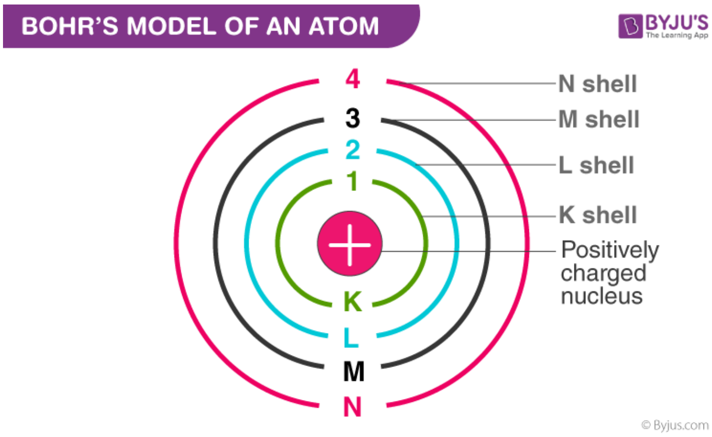
Based on the above observations and conclusions, Rutherford proposed his own atomic structure, which is as follows.

- The nucleus is at the centre of an atom, where most of the charge and mass is concentrated.
- The atomic structure is spherical.
- Electrons revolve around the nucleus in a circular orbit, similar to the way ***planets orbit the sun.***

Bohr's Atomic Theory :

The Bohr model of the atom was proposed by Neil Bohr in 1915. It came into existence with the modification of Rutherford's model of an atom.

Bohr theory modified the atomic structure model by explaining that electrons move in fixed orbitals (shells) and not anywhere in between and he also explained that each orbit (shell) has a fixed energy. Rutherford explained the nucleus of an atom and Bohr modified that model into electrons and their energy levels.



Bohr's model consists of a small nucleus (positively charged) surrounded by negative electrons moving around the nucleus in orbits. Bohr found that an electron located away from the nucleus has more energy, and the electron which is closer to nucleus has less energy.

Postulates of Bohr's Model of an Atom

- In an atom, electrons (negatively charged) revolve around the positively charged nucleus in a definite circular path called orbits or shells.
- Each orbit or shell has a fixed energy and these circular orbits are known as orbital shells.
- The energy levels are represented by an integer (n=1, 2, 3…) known as the quantum number. This range of quantum number starts from nucleus side with n=1 having the lowest energy level. The orbits n=1, 2, 3, 4… are assigned as K, L, M, N…. shells and when an electron attains the lowest energy level, it is said to be in the ground state.
- The electrons in an atom move from a lower energy level to a higher energy level by gaining the required energy and an electron moves from a higher energy level to lower energy level by losing energy.

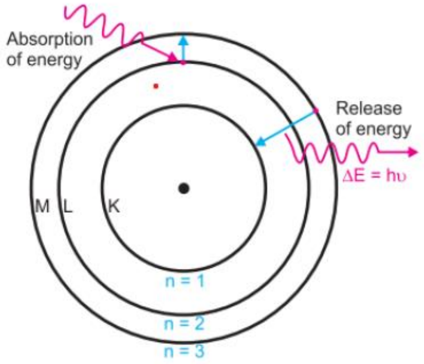
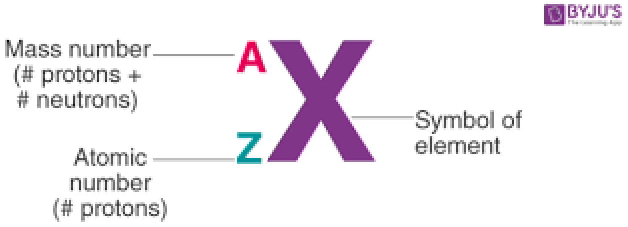


Figure II-5 : The Bohr's atomic model

V. Identification of elements: Representation, Atomic mass, Relative atomic mass

1- Symbolic writing of atom :

The chemical symbol, represented by the letter X in atomic notation, is a one or two-letter abbreviation of an element's name derived from its English or Latin name. For instance, hydrogen is denoted by H, and gold by Au (from 'Aurum')



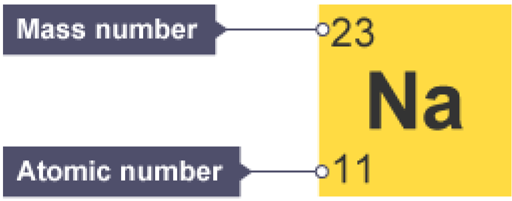
1. **The symbol X:** the usual element symbol
2. **The atomic number Z:** equal to the number of protons (placed as a left subscript)
3. **The mass number A:** equal to the number of protons and neutrons in the isotope (placed as a left superscript)
4. **Neutron Number N** : The number of neutrons is therefore the difference between the mass number and the atomic number: A – Z = number of neutrons (N).

The number of protons in the nucleus of an atom is its **atomic number (Z)**. This is the defining trait of an element: Its value determines the identity of the atom. A neutral atom must contain the same number of positive and negative charges, so the number of protons equals the number of electrons. Therefore, the atomic number also indicates the number of electrons in an atom. The total number of protons and neutrons in an atom is called its **mass number (A)**.

Table 1. Properties of Subatomic Particles

Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)
electron	outside nucleus	-1.602×10^{-19}	1−	0.00055	0.00091×10^{-24}
proton	nucleus	1.602×10^{-19}	1+	1.00727	1.67262×10^{-24}
neutron	nucleus	0	0	1.00866	1.67493×10^{-24}

Exemple : Calculating the number of protons, neutrons and electrons in an atom



This is the chemical symbol for sodium

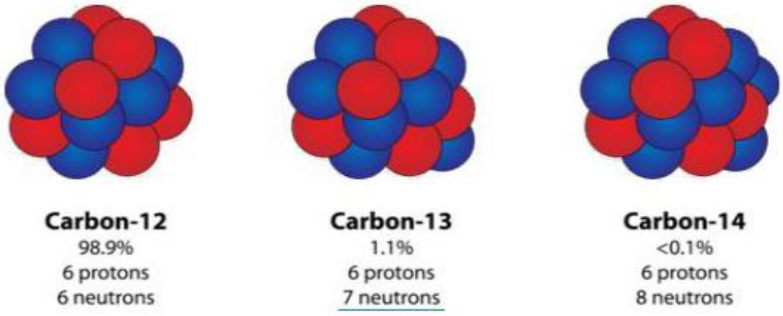
- Mass number = 23

- Atomic number = 11
  - Number of protons in a sodium nucleus = 11
  - Number of electrons orbiting a sodium nucleus = 11.
  - Number of neutrons in a sodium nucleus = 23 – 11 = 12.

2- Isotopes:

Isotopes are atoms with the same atomic number but different neutron numbers and thus different mass numbers. It can be defined as the variants of chemical elements that possess the same number of protons and electrons, but a different number of neutrons. In other words, isotopes are variants of elements that differ in their nucleon (The total number of protons and neutrons) numbers due to a difference in the total number of neutrons in their respective nuclei.

Exemple : isotopes of carbone



Isotope Facts

- All elements have isotopes.
- Isotopes have the same chemical properties, but different physical properties.
- Isotones are atoms that have the same neutron number but different proton number. For example, $^{36}_{16}S$, $^{37}_{17}Cl$, $^{38}_{18}Ar$, $^{39}_{19}K$, and $^{40}_{20}Ca$ are all isotones of 20 since they all contain 20 neutrons.
- Isobars are elements that have the same number of nucleons (sum of protons and neutrons). The series of elements with 40 Mass numbers serve as a good example; $^{40}_{16}S$, $^{40}_{17}Cl$, $^{40}_{18}Ar$, $^{40}_{19}K$, and $^{40}_{20}Ca$. The nucleus of all the above-mentioned elements contain the same number of particles in the nucleus but contain varying numbers of protons and neutrons.There are two main types of isotopes: stable and unstable (radioactive).
- There are 254 known stable isotopes.
- All artificial (lab-made) isotopes are unstable and therefore radioactive; scientists call them radioisotopes.
- Some elements can only exist in an unstable form (for example, uranium).
- Hydrogen is the only element whose isotopes have unique names: deuterium for hydrogen with one neutron and tritium for hydrogen with two neutrons.
- **Average mass of isotopes :**

The mass of an element shown in a periodic table or listed in a table of atomic masses is a weighted, average mass of all the isotopes present in a naturally occurring sample of that element. This is equal to the sum of each individual isotope’s mass multiplied by its fractional abundance divided by 100.

$$\text{Average mass} = \sum i(\text{fractional abundance x isotopic mass})/100$$

Example:

Determine the average atomic mass of carbon (the atomic mass of natural carbon) from its three natural isotopes:

isotope	Relative abundance (%)	Atomic mass (amu)
^{12}C	98.892	12
^{12}C	1.108	13.00335
^{12}C	2.10^{-10}	14.00317

From the above data, the average atomic mass of carbon will out to be :

$$\mathrm{M} = \frac{(98.892\mathrm{x}12 + 1.108\mathrm{x}13.00335 + 2.10^{-10}\mathrm{x}14.00317)}{100} = 12.011\ amu$$

Exercises of chapter 2

Exercise 1:

In J.J. Thomson's experiment, an electron beam is deflected using an electric field E and the deflection Y is measured on the screen. The deflection of the electron beam is cancelled by the opposing action of a magnetic field B.

1- Establish the expression for the mass charge e/m of the electron as a function of the quantities involved in the experiment.

2- Determine the speed of the electrons.

3- Determine the accelerating potential V that must be applied between the cathode and the anode to impart this kinetic energy to the electrons.

4- Calculate the deflection Y_0 undergone by the beam at the capacitor output, knowing that the length of the capacitor is L= 10 cm.

We give: E= 3.6 10^4 V/m; B= 9.10^{-4} Tesla; e=$1.6.10^{-19}$C; me= $9.31.10^{-31}$Kg

Exercise 2 :

In a MILIKAN-type device, the behavior of a single charged droplet is being studied.

a) In the absence of an electric field, the droplet descends 2,61 mm in 12 seconds. Determine the radius and mass of the droplet.

b) An electric field is applied to the droplet, causing it to come to rest. The capacitor plates are 2 cm apart, and the potential difference is 4320 volts. What is the charge of the droplet? Compare this charge to the elementary charge of the electron and draw a conclusion.

c) After remaining in equilibrium for some time, the droplet begins to rise with a speed of V_2 = 2,174 x 10^{-2} cm/s. Calculate the new charge, q'. Given: ρoil = 900 kg/m^3; η = 18 x 10^{-6} MKSA; g = 9,81 m/s^2.

Note : we neglect the Archimedean force (F_A)

Exercise 3:

1- Using Representation of atom , complete the following table :

$$\boxed{{}^{A}_{Z}X}$$

Position in symbol	**Term to describe this information**	**Particle used to determine this :**
X		-
A		
Z		

2- Give the number of protons, electrons, and neutrons in neutral atoms of each of the following atoms:

$${}^{7}_{3}Li\ ,\quad {}^{125}_{52}Te,\quad {}^{31}_{15}P\ ,\quad {}^{77}_{34}Se\ ,\quad {}^{24}_{12}\,Mg^{2+}\quad ,\ {}^{79}_{34}\,Se^{2-}\quad ,\ {}^{9}_{3}\,Li^{+}$$

3- Among the following groups, indicate the isotope groups:

$${}^{34}_{16}X\ ,\ {}^{54}_{26}X\ ,\ {}^{37}_{17}X\ ,\ {}^{58}_{26}X\ ,\ {}^{32}_{16}X\ ,\ {}^{56}_{26}X\ ,\ {}^{35}_{17}X\ ,\ {}^{33}_{16}X\ .$$

Exercise 4:

Iron ($_{26}$Fe) has four naturally occurring isotopes, and their masses and natural abundances are shown in the table below.

Isotope	Isotopic mass	Natural Abundance (%)
^{54}Fe	53.9396	5.845
^{56}Fe	55.9349	91.75
^{57}Fe	56.9354	2.119
^{58}Fe	57.9333	0.282

1. Give the constitution of each of these isotopes.
2. Calculate the relative atomic mass of iron
3. Find the Natural mass of iron (${}^{56}_{26}Fe$)
4. Calculate the mass defect of the nucleus ${}^{56}_{26}Fe$.

Given : The masses of the proton and the neutron value: mn = 1.0086 uma; mp = 1.0073 uma

Exercise 5:

Natural oxygen is composed of 3 isotopes ^{16}O, ^{17}O and ^{18}O with atomic masses of 15.9949 a.m.u, 16.9991 a.m.u and 17.9992 a.m.u respectively. Knowing that the atomic mass of natural oxygen is l5.9994 a.m.u and that the relative abundance of the isotope ^{17}O is 0.037%.

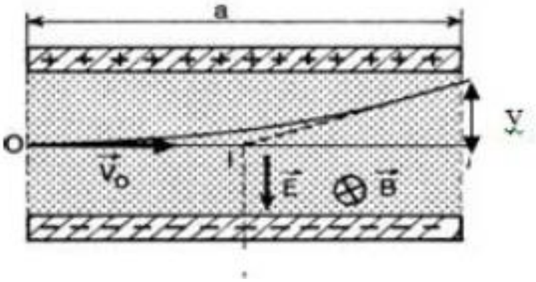
1. What are the relative abundances of the other two isotopes?

2. Which is the most stable isotope?

Given: mp= 1.00727 amu, mn= 1.0086 amu; C=3.10^{8} m/s.

Corrected exercises of chapter 2

Exercise 1: E= 3.6 10^4 V/m; B= 9.10^{-4} Tesla; e=$1.6.10^{-19}$C; me= $9.31.10^{-31}$Kg



1- Establish the expression for the mass charge e/m of the electron as a function of the quantities involved in the experiment.***(These formulas have been demonstrated in the course)***

$$y = \frac{1}{2}\frac{e.E}{m_e}\frac{l^2}{v^{\,2}} \longrightarrow \qquad \frac{e}{m_e} = \frac{2\,y}{l^2}\frac{v^{\,2}}{E}$$

Or in function of B (Magnetic field): $\frac{e}{m_e} = \frac{2\,y}{l^2}\frac{E}{B^2}$

2- Determine the speed of the electrons: E= 3.6 10^4 V/m; B= 9.10^{-4} Tesla

Electric force : F_e = - e.E and magnetic force : F_B = e. v. B

The force due to the electric field is equal to the force due to the magnetic field when the beam is undeflected:

$$FB = -Fe = -(-eE)$$

$$qv\,B = +e\,E \;\rightarrow v = \frac{E}{B}$$

So , $$v = \frac{E}{B} = \frac{3.6\;10^4}{9.10^{-4}} = \;4.10^7\;\frac{m}{s}$$

3- Determine the accelerating potential V that must be applied between the cathode and the anode to impart this kinetic energy to the electrons (me= $9.31.10^{-31}$Kg).

a) - Calculate the kinetic energy (Ec) of the electrons using the formula:

- Ec = ½ m v^2
- Ec = ½ . $9.31.10^{-31}$. $(4.10^7)^2$
- Ec = 7.44 . 10^{-16} J
-

b)- Determine the accelerating potential V

The kinetic energy (Ec) of the electrons is also given by the relationship:

Ec = e. V where V is accelerating potential

$$V = \frac{Ec}{e} = \frac{7.44\ 10^{-16}}{1.6\ 10^{-19}} = 4650\ V$$

4- Calculate the deflection Y_0 undergone by the beam at the capacitor output, knowing that the length of the capacitor is L= 10 cm.

The deflection Y can be calculated using the formula:

$$y = \frac{1}{2}\frac{e.E}{m_e}\frac{l^2}{v^{\ 2}}$$

Substitute the known values:

E =3.6 10^4 V/m; L= 10 cm = 0.1 m ; me = $9.31.10^{-31}$Kg; v= 4 .10^7 m/s ; e= $1.6.10^{-19}$C

$$y = \frac{1}{2}\frac{e.E}{m_e}\frac{l^2}{v^{\ 2}} = \frac{1}{2}\frac{(1.6.10^{-19})(3.6\ 10^4)}{(9.31.10^{-31})}\frac{(0.1)^2}{(4.10^7)^{\ 2}} = 2.03\ 10^{-6}m$$

$$y = 2.03\ \mu m$$

Exercise 2 :
The droplet is subjected to three forces:

4- **The gravitational force (F_G)**

5- **The air resistance force (stokes force) (**FS)

d. The gravitational force, **F_G = mg**,

$$F_G = m.g \Rightarrow \quad F_G = \rho_o.\ V.\ g$$

where: (**m**: is the mass of the droplet., **g**: is the gravitational acceleration, **ρ**: is the volume mass density of the droplet (oil), **V**: is the volume of the droplet).

And since the droplet has a spherical shape, its shape does not change during movement

$$v = \frac{4}{3}.\pi.r^3 \quad : \qquad F_G = \rho_o.\frac{4}{3}.\pi.r^3.g$$

Where r: is the radius of the sphere.

e. The air resistance force (stokes force) (FS)

$$F_s = 6\pi.\eta.r.v_0$$

Where: V_0 : Droplet velocity n : Air viscosity coefficient

The sum of these forces = 0

$$\vec{F}_G + \vec{F}_S = 0 \ , \qquad F_G - F_S = 0$$

$$\rho_o.\frac{4}{3}\pi.r^3g - 6\pi.\eta.r.v_0 = 0, \qquad \rho_o.\frac{4}{3}\pi.r^2g = 6\pi.\eta.v_0$$

$$r^2 = \frac{18\,.\eta.v_0}{4\,g.\rho_o}$$

ρo = 900 kg/m^3; η = 18 x 10^{-6} MKSA; g = 9,81 m/s2. 35316

v= d/t : d= 2,61 mm and t= 12 seconds
v= 2.175 . 10^{-4} m/s

$$r^2 = \frac{18\,.\eta.v_0}{4\,g.\rho_o}$$

$$r^2 = \frac{18\,.(18 \text{ x } 10 - 6).(2.175\,.10 - 4\,)}{4.\,9{,}81.\,900}$$

$r = 1.4\ 10^{-6}$ m

mass of the droplet : $m = \rho_o.\frac{4}{3}\pi.r^3$

$$m = 18\,.10^{-6}.\frac{4}{3}\pi.(1.4\ 10^{-6})^3$$

$$m = 1.03\ \ 10^{-14}\ \text{Kg}$$

b) An electric field is applied to the droplet, causing it to come to rest. The capacitor plates are 2 cm apart, and the potential difference is 4320 volts. What is the charge of the droplet? Compare this charge to the elementary charge of the electron and draw a conclusion.

$$F_G = F_E$$

$F_E = q.E\ ,\qquad F_G = m.g$

$$F_G = F_E\ ,\qquad q.E = m.g$$

$$q = m.\frac{g}{E}$$

The electric field between the plates is given by $E = \frac{V}{d}$

$$q = m.\frac{g}{\frac{V}{d}} = \frac{m.g.d}{V}$$

- q= 4.68.10^{-19} = **3** . 1.6. 10^{-19} c
- (4.68.10^{-19} / 1.6. 10^{-19}) ≅ 3

We conclude that this charge is 3 times the elementary charge e = 1.6. 10^{-19} C.

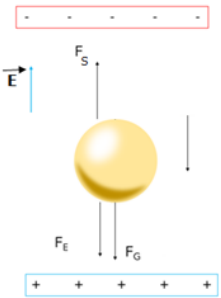
c) After remaining in equilibrium for some time, the droplet begins ***to rise*** with a speed of $V_2 = 2{,}174 \times 10^{-2}$ cm/s.

The previous equation can be written as follows:

$$F_G + F_E - F_S = 0 \ , \ F_G - F_S = -F_E \ , \qquad \mathrm{m.g} - 6\pi\,\eta\,\mathrm{r}\,v \ = -\mathrm{q'.E}$$

$$\mathrm{q'.} = -\,(\mathrm{m.g} - 6\pi\,\eta\,\mathrm{r}\,v \ )/\mathrm{E}$$

$$\mathrm{q'} = -\,9.115\ 10^{-19}\ \mathrm{C}$$



Exercise 3:

4- Using Representation of atom , complete the following table :

$$\boxed{^{A}_{Z}\mathrm{X}}$$

Position in symbol	Term to describe this information	Particle used to determine this :
X	Element symbol	-
A	Mass number	Proton and neutron
Z	Atomoc number	Proton

5- Give the number of protons, electrons, and neutrons in neutral atoms of each of the following atoms:

$^{7}_{3}Li$, $^{125}_{52}Te$, $^{31}_{15}P$, $^{77}_{34}Se$, $^{24}_{12}Mg^{2+}$, $^{79}_{34}Se^{2-}$, $^{9}_{3}Li^{+}$

	Mass number	Number of protons	Number of electrons	Number of neutrons
$^{7}_{3}Li$	7	3	3	4
$^{125}_{52}Te$	12	52	52	73
$^{24}_{12}Mg^{2+}$	24	12	10	12
$^{31}_{15}P$	31	15	15	16
$^{77}_{34}Se$	77	34	34	43
$^{79}_{34}Se^{2-}$	79	34	36	45

$^{9}_{3}\boldsymbol{Li^{+}}$	9	3	2	6

6- Among the following groups, indicate the isotope groups:

$^{34}_{16}\boldsymbol{X}$, $^{54}_{26}\boldsymbol{X}$, $^{37}_{17}\boldsymbol{X}$, $^{58}_{26}\boldsymbol{X}$, $^{32}_{16}\boldsymbol{X}$, $^{56}_{26}\boldsymbol{X}$, $^{35}_{17}\boldsymbol{X}$, $^{33}_{16}\boldsymbol{X}$.

Isotope group 1	**Isotope group 2**	**Isotope group 3**
$^{34}_{16}\boldsymbol{X}$	$^{35}_{17}\boldsymbol{X}$	$^{54}_{26}\boldsymbol{X}$
$^{33}_{16}\boldsymbol{X}$	$^{37}_{17}\boldsymbol{X}$	$^{56}_{26}\boldsymbol{X}$
$^{32}_{16}\boldsymbol{X}$	-	$^{58}_{26}\boldsymbol{X}$

Exercise 4 :

Iron ($_{26}$Fe) has four naturally occurring isotopes, and their masses and natural abundances are shown in the table below.

Isotope	**Isotopic mass**	**Natural Abundance (%)**
^{54}Fe	53.9396	5.845
^{56}Fe	55.9349	91.75
^{57}Fe	56.9354	2.119
^{58}Fe	57.9333	0.282

5. Give the constitution of each of these isotopes.

	Mass number	**Number of protons**	**Number of electrons**	**Number of neutrons**
^{54}Fe	54	26	26	28
^{56}Fe	56	26	26	30
^{57}Fe	57	26	26	31
^{58}Fe	58	26	26	32

6. Calculate the relative atomic mass of iron

$$M = \frac{M1\ \ X\% + M2\ .Y\% + M3\ .Z\% + M4\ .W\%}{100}$$

$$M = \frac{53.9396\ x\ 5.845 + 55.9349\ x\ 91.75 + 56.9354\ x 2.119\ \ + 57.9333\ x\ 0.282}{100}$$

$$M = \ \ 55.8428 \quad uam$$

7. Find the Natural mass of iron ($^{56}_{26}\boldsymbol{Fe}$)

mn = 1.0086 uma; mp = 1.0073 uma , $^{56}_{26}Fe$ **contains : 26 protons and 30 neutrons**

Natural mass of iron ($^{56}_{26}Fe$ **)** : M = (n_p x 1.0073) + (n_n x 1.0086) = 56.4478 uma

M = (26 x 1.0073) + (30 x 1.0086) = 56.4478 uma

8. Calculate the mass defect of the nucleus 2656 $^{56}_{26}Fe$ in uma , in J and in MeV.

Δm = 56.4478 - 55.9349 = 0.5129 uma

Exercise 5:

1. What are the relative abundances of the other two isotopes?

- ^{16}O with an atomic mass of 15.9949 a.m.u
- ^{17}O with an atomic mass of 16.9991 a.m.u (0.037%)
- ^{18}O with an atomic mass of 17.9992 a.m.u

1. Calculate the relative atomic mass of iron

$$M = \frac{M1\ \ X\% + M2\,.Y\% + M3\,.Z\%}{100}$$

The sum of the relative abunda nces must equal 100:

X + 0. 037 + Z = 100

X + 0.037+Z =100 , Thus, X + Z = 99.963 $\longrightarrow$ Z = 99.963- X

$$15.9994 = \frac{15.9949\ \ X\% + 16.9991\,.0.037 + 17.9992\,.\mathrm{Z}\%}{100}$$

$$15.9994 = \frac{15.9949\ \ X + 0{,}6289667 + 17.9992\,.(99.963 - \mathrm{X})}{100}$$

$$15.9994 = \frac{15.9949\ \ X + 0{,}6289667 + 1799{,}25402 - 17.9992\ \mathrm{X})}{100}$$

X = 99,75 % (^{16}O) and Z = 0,205% (^{18}O)

2. Which is the most stable isotope?

ΔE = Δm C^2 , mp= 1.00727 amu, mn= 1.0086 amu; C=3.10^8 m/s ,

1 J= **1 kg.m^2.s^{-2}**

1 MeV =1.6023×10^{-13} J

$1\text{u} = 1.6605 \times 10^{-27}\, Kg$

1) $\boldsymbol{{}^{16}_{8}O \longrightarrow}$ **8 p + 8 n**

$\Delta E = \Delta m\, C^2 = [(8\, m_p + 8\, m_n) - m(^{16}O)]\, .\, C^2$

$= [(8 \times 1.00727 + 8 \times 1.0086) - 15.9949] \times 1.66\, .10^{-27}\, .\, (3.10^8)^2$

= $1.98\, .\, 10^{-11}$ J= 124 Mev , ΔE/ n =124/16=7.75 Mev/ nucleus

2) $\boldsymbol{{}^{17}_{8}O \longrightarrow}$ **8 p + 9 n**

$\Delta E = \Delta m\, C^2 = [(8\, m_p + 9\, m_n) - m(^{17}O)]\, .\, C^2$

$= [(8 \times 1.00727 + 9 \times 1.0086) - 16.9991] \times 1.66\, .10^{-27}\, .\, (3.10^8)^2$

= $2.047\, .\, 10^{-11}$ J= 128 Mev , ΔE/ nucleus =128/17=7.53 Mev/ nucleus

3) $\boldsymbol{{}^{18}_{8}O \longrightarrow}$ **8 p + 10 n**

$\Delta E = \Delta m\, C^2 = [(8\, m_p + 10\, m_n) - m(^{18}O)]\, .\, C^2$

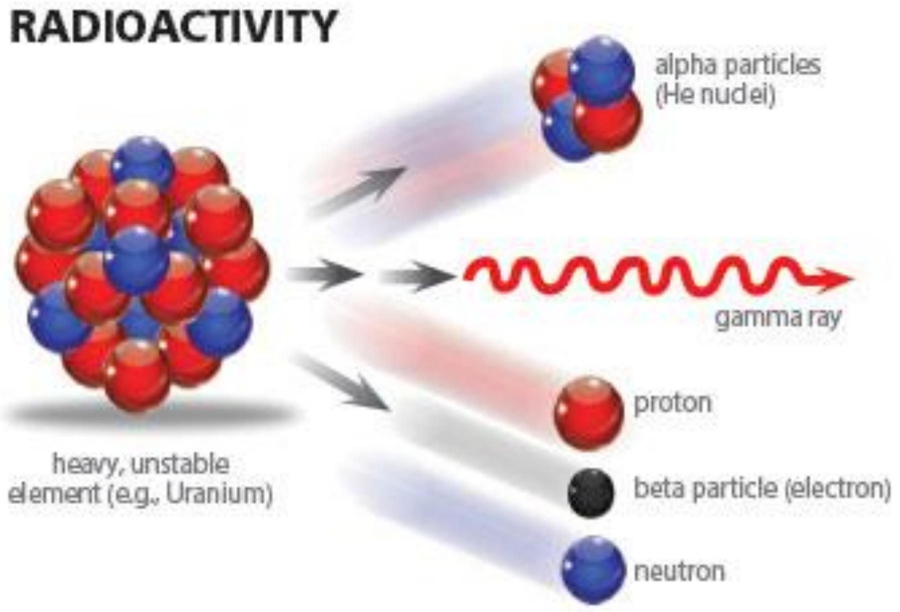
$= [(8 \times 1.00727 + 10 \times 1.0086) - 17.9992] \times 1.66\, .10^{-27}\, .\, (3.10^8)^2$

= $2.175\, .\, 10^{-11}$ J= 136 Mev , ΔE/ nucleus =136/A=7.56 Mev/ nucleus

The most stable isotope is typically the one with the highest binding energy per nucleon. The most stable isotope is $\boldsymbol{{}^{16}_{8}O}$ **, it has** the highest binding energy per nucleon **: ΔE = 7.75 Mev/ nucleus**

Chapter 3

Radioactivity



I. Introduction:

Radioactivity is the process by which unstable atomic nuclei spontaneously decay, emitting energy in the form of particles or electromagnetic radiation. This phenomenon occurs because the nucleus of the atom is unstable due to an imbalance in the number of protons and neutrons. The emission of radiation allows the nucleus to transform into a more stable form, often resulting in the atom becoming a different element.

The spontaneous change of an unstable nuclide into another is **radioactive decay**. The unstable nuclide is called the ***parent nuclide***; the nuclide that results from the decay is known as the ***daughter nuclide***. The daughter nuclide may be stable, or it may decay itself. The radiation produced during radioactive decay is such that the daughter nuclide lies closer to the band of stability than the parent nuclide, so the location of a nuclide relative to the band of stability can serve as a guide to the kind of decay it will undergo.

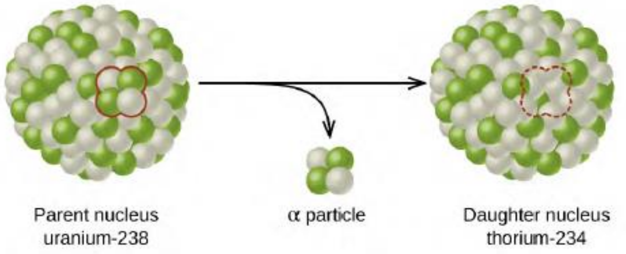


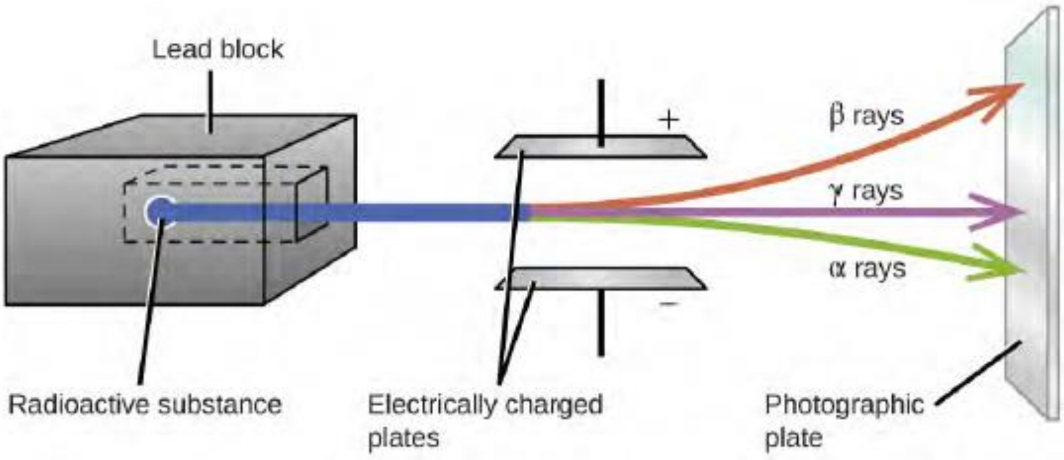
Figure III-1 : A nucleus of uranium-238 (*The parent nuclide*) undergoes α decay to form thorium-234 (*The daughter nuclide*).

II. Natural radioactivity

Natural radioactivity was first discovered by *Henri Becquerel* in 1896 while studying uranium salts. Natural radioactivity refers to the spontaneous decay of unstable atomic nuclei that occur naturally in the environment. This process involves the emission of energy in the form of radiation, such as alpha, beta, or gamma rays, as the unstable nuclei attempt to reach a more stable configuration.

- Types of Natural Radioactive Decay:

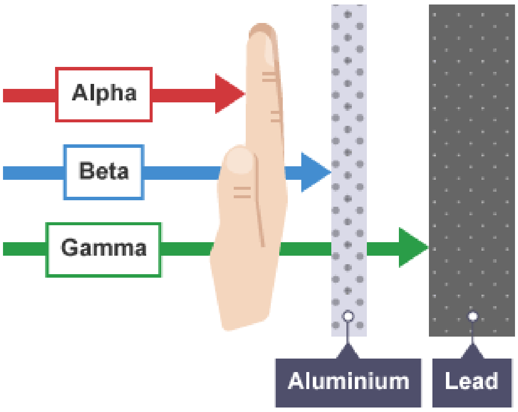
Ernest Rutherford’s experiments involving the interaction of radiation with a magnetic or electric field helped him determine that one type of radiation consisted of ***positively charged*** and relatively massive ***α particles***; a second type was made up of ***negatively charged*** and much less massive ***β particles***; and a third was ***uncharged*** electromagnetic waves, ***γ rays***. We now know that ***α particles*** are ***high-energy*** helium nuclei, ***β particles*** are ***high energy*** electrons, and ***γ radiation*** compose ***high-energy electromagnetic radiation***..



- **Absorption of radiation**

The energy of the three radiations is absorbed by the material through which the radiation passes. The amount of energy which is absorbed depends on the type of radiation and the type of the absorbing material.

- The range of ***the alpha radiation*** in an absorbing material is less than that of beta or gamma. The alpha radiation transfers more energy to an absorber than beta or gamma radiation. Alpha radiation is absorbed by the thickness of the skin or by a few centimeters of air.
- ***Beta radiation*** is more penetrating than alpha radiation. It can pass through the skin, but it is absorbed by a few centimeters of body tissue or a few millimeters of aluminum.
- ***Gamma radiation*** is the most penetrating of the three radiations. It can easily penetrate body tissue. It requires a few centimeters of lead or about 1 meter of concrete to absorb it.



- **Classification of radioactive decay :**

We classify different types of radioactive decay by the radiation produced:

1. **Alpha Decay**: Large, heavy atoms like uranium and radium undergo alpha decay, emitting alpha particles (2 protons and 2 neutrons) and transforming into different elements. ${}^{4}_{2}H_{e} = ({}^{4}_{2}\alpha)$

$${}^{235}_{92}U \rightarrow {}^{4}_{2}He + {}^{231}_{90}Th$$

2. **Beta Decay**: In beta decay, a neutron in the nucleus is converted into a proton or vice versa, accompanied by the emission of a beta particle .

   - ***Transformation of nitron to proton : (β^{-} , ${}^{0}_{-1}e^{-}$)***

$${}^{1}_{0}n \rightarrow {}^{1}_{1}p + {}_{-1}^{0}e^{-}$$

$${}^{A}_{Z}X \rightarrow {}_{Z+1}^{A}Y + {}_{-1}^{0}e^{-}$$

Example : $${}^{14}_{6}C \rightarrow {}^{14}_{7}N + {}_{-1}^{0}e^{-}$$

- ***Transformation of proton to nitron (positrons) : (β^{+} , ${}^{0}_{+1}e^{+}$)***

$$ {}_{1}^{1}p \rightarrow {}_{0}^{1}n + {}_{+1}^{0}e^{+} $$

$$ {}_{Z}^{A}X \rightarrow {}_{Z-1}^{A}Y + {}_{+1}^{0}e^{+} $$

3. **Gamma Decay**: type of radioactivity in which some unstable atomic nuclei dissipate excess energy by a spontaneous electromagnetic process. Gamma rays are emitted as excess energy after alpha or beta decay.

$$ {}_{Z}^{A}X \rightarrow {}_{Z+1}^{A}Y^{*} + {}_{-1}^{0}e^{-} $$

$$ {}_{Z+1}^{A}Y^{*} \rightarrow {}_{Z+1}^{A}Y + {}_{0}^{0}\gamma $$

$$ {}_{Z}^{A}X \rightarrow {}_{Z+1}^{A}Y + {}_{-1}^{0}e^{-} + {}_{0}^{0}\gamma $$

Example: $ {}_{43}^{99m}Tc \rightarrow {}_{43}^{99}Tc + {}_{0}^{0}\gamma $

- **Diagram of stability (NZ GRAPH):**

There are three types of decay, alpha decay, beta decay minus, or beta decay plus.
We can determine whether a radioisotope is likely to decay and what type of decay is likely to occur by looking at its position on a proton-neutron **NZ graph**. This graph shows ***the number of protons*** on **the x-axis** and **the number of neutrons** on **the y-axis**. This is the stability line. If a radioisotope lies left or right of this line it is unstable and likely to decay to become stable.

- A radioisotope that lies to the left of the stability line has too many neutrons and is likely to undergo ***beta minus*** decay, so give off electrons.
- A radioisotope that lies to the right of the stability line has too many protons and is likely to undergo ***beta plus*** decay, so give off positrons.
- If particles have a high number of protons, usually more than **82**, they are likely to undergo ***alpha decay***.

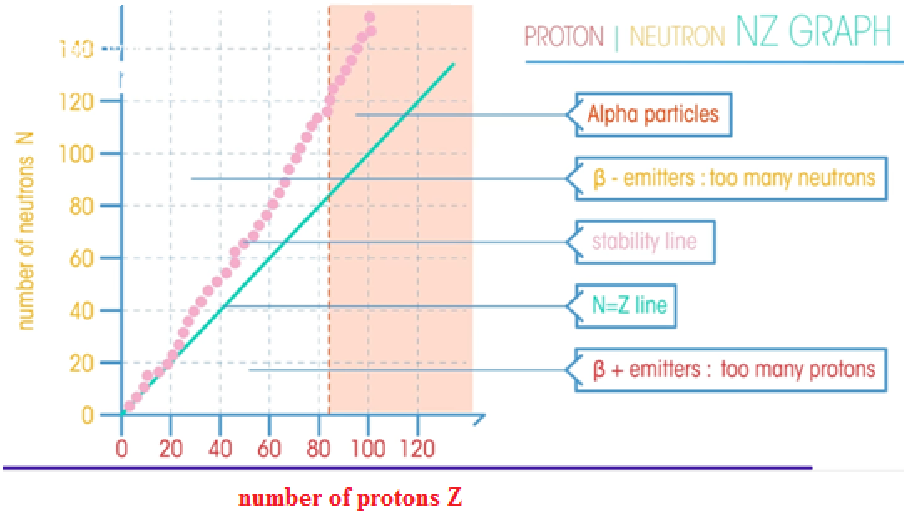


Figure III-1: Diagram of stability (NZ GRAPH)

- **Comparison chart**

Many entities can be involved in nuclear reactions. The most common are protons, neutrons, alpha particles, beta particles, positrons, and gamma rays, as shown in **Table 1**

Name	Symbol(S)	Presentation	Description
Alpha particle	${}^{4}_{2}He$ or ${}^{0}_{1}\alpha$		(High-energy) helium nuclei consisting of two protons and two neutrons
Beta particle	${}^{0}_{1}e$ or ${}^{0}_{-1}\beta$		(High-energy) electrons
Positron	${}^{0}_{+1}e$ or ${}^{0}_{\mp1}\beta$		Particles with the same mass as an electron but with 1 unit of positive charge.
Proton	${}^{1}_{1}H$ or ${}^{1}_{1}P$		Nuclei of hydrogen atoms
Neutron	${}^{0}_{1}n$		Particles with mass approximately equal to that of a proton but with no charge
Gamma ray	γ		Very high-energy electromagnetic radiation.

III. Artificial radioactivity and nuclear reactions: Nuclear fission, Nuclear fusion, Transmutation

Artificial radioactivity refers to the process where normally stable atoms become radioactive through human intervention, typically by bombarding them with subatomic particles such as neutrons, protons, or alpha particles.

It was first discovered by Irene Joliot-Curie and Frederic Joliot in 1934 when they bombarded stable isotopes with alpha particles, creating radioactive isotopes.

Applications:

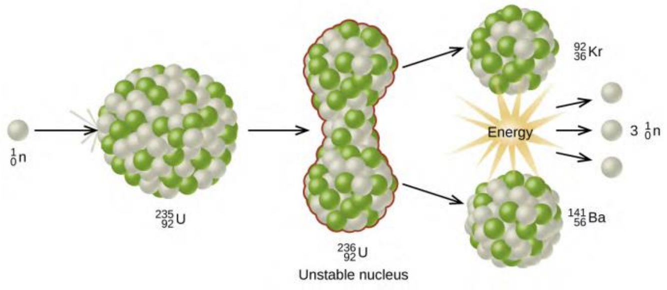
- **Medical Uses**: Radioactive isotopes produced artificially are used in medicine for diagnostic imaging (e.g., PET scans) and cancer treatments (radiotherapy).
- **Industrial Applications**: They are used in industrial gauges, radiography, and sterilization processes.
- **Research**: Artificially produced isotopes are crucial in scientific research, particularly in tracing chemical and biological processes

Nuclear Reactions:

Nuclear reactions involve changes in the nuclei of atoms, resulting in the transformation of one or more atomic nuclei or particles. They are a significant source of energy in nuclear power plants, where controlled fission reactions generate heat used to produce electricity.

Types of Nuclear Reactions:

- **Fission**: Large nuclei split into smaller nuclei, releasing energy (e.g., nuclear power plants)



$$^{235}_{92}U + ^{0}_{1}n \rightarrow \ ^{236}_{92}U \rightarrow \ ^{141}_{56}Ba + \ ^{92}_{36}Kr + 3\ ^{1}_{0}n$$

Uncontrolled nuclear fission ⟶ Atomic bomb

Controlled nuclear fission ⟶ Nuclear power plant

The energy released by this type of reaction is around 200 Mev/atom.

- **Fusion**: Small nuclei combine to form larger nuclei, releasing large amounts of energy (e.g., the sun).

$$ {}_{1}^{2}H + {}_{1}^{3}H \rightarrow {}_{2}^{4}H + {}_{0}^{1}n + energy $$

$$ {}_{1}^{2}H + {}_{1}^{2}H \rightarrow {}_{1}^{1}H_e + {}_{1}^{3}H + energy $$

$$ {}_{1}^{2}H + {}_{3}^{6}Li \rightarrow 2\,{}_{2}^{4}H_e + energy $$

- **Transmutation**: These are reactions that produce nuclei with a mass number approximately equal to or very close to the mass number of the nucleus used as a target. The resulting nuclei can be either stable or radioactive, and the first reaction of this type was observed by Rutherford in 1919.

$$ {}_{2}^{4}H_e + {}_{7}^{14}N \rightarrow {}_{8}^{17}O + {}_{1}^{1}H({}_{1}^{1}p) $$

This is a (α, p) reaction and can be written symbolically as: ${}^{14}{}_{7}N\ (\alpha, P)\ {}^{17}{}_{8}O$

IV. Radioactive decay kinetics: Radioactive decay law: Activity of a radioactive nucleus, Radioactive period or half-life time

- The kinetics of the radioactive decay is measured by the rate at which unstable nuclei disintegrate or decay into more stable nuclei.
- Kinetics gives information about how many nuclei are present at any given instant of time.
- It provides information about how long the unstable nuclei will continue to decay.

1- Law of Radioactive Decay :

Let's assume that X is a radioactive element "parent" that undergoes decay to produce element Y "daughter," as shown in the equation:

	X	$\rightarrow$	Y
$t=0$	N_0		0
$t=t$	N		N_0-N

- 'N_0' is the initial number of nuclei of a radioactive sample present at time $t=0$
- 'N' is the number of nuclei at any time 't'.
- N_0 - N is the number of stable Y atoms formed.

Radioactive decay follows first order kinetics ,

$$-\frac{dN}{dt} = \lambda N \qquad \textit{Equation (1)}$$

Where λ : is the radioactivity constant and represents the probability of collapse per second.
The sign (-) indicates a decrease in the substance.
Rearranging the terms gives :

$$\frac{dN}{N} = -\lambda dt \qquad \textit{Equation (2)}$$

Integrating both sides of the equation gives,

$$\frac{dN}{N} = -\lambda\, dt \quad \Rightarrow \int \frac{dN}{N} = -\lambda \int dt$$

Applying the limits gives :

$$\ln N = -\lambda\, t + \ln N_0 \quad \Rightarrow \quad \ln\frac{N}{N} = -\lambda\, t \qquad \textit{Equation (3)}$$

Raising both sides of the equation to the power 'e' and rearranging the terms we get :

$$N = N_0\, e^{-\lambda\, t} \qquad \textit{Equation (4)}$$

Equation (4) is termed as the law of radioactive decay or the radioactive decay formula. Radioactivity follows the exponential decay law.

This law can be written in another form depending on the mass of the radioactive element studied since:

In 1 mole of the element we have N_A the Avogadro number

$$1 \text{ mol} \longrightarrow N_A$$

$$n \text{ mol } (m/M) \longrightarrow N$$

$$N = n.N_A = \frac{m}{M} .N_A \quad \text{and} \quad N_0 = n_0.N_A = \frac{m_0}{M} .N_A$$

So $N = N_0 \, e^{-\lambda t}$ Can be written: $m = m_0 \, e^{-\lambda t}$

n_0: number of moles of radioactive element at t =0

n: number of moles of radioactive element at t

m: mass of radioactive element at t

m_0: mass of radioactive element at t =0

2- Half-life

- It is the time taken for the number of nuclei (N) of a radioactive material to reduce to half of its initial value. It is denoted by $T_{1/2}$.
- Half-life is the time taken for N to decrease from its initial value N_0 and reach $\frac{N_0}{2}$.

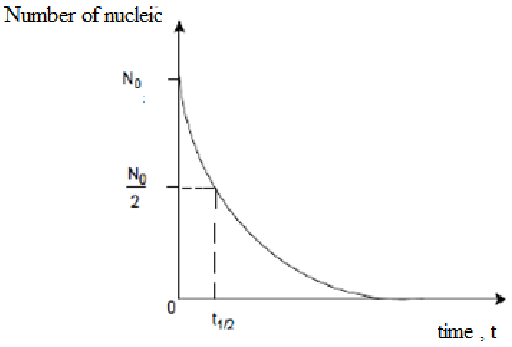
So, when t= $T_{1/2}$, $N = \frac{N}{2}$, Substituting this in equation (4) we get :

$$\frac{N_0}{2} = N_0 \, e^{-\lambda t_{1/2}} \quad \Rightarrow \quad ln\frac{1}{2} = \ln(e^{-\lambda t_{1/2}})$$

$$\Rightarrow \quad ln\frac{1}{2} = -\lambda \, t_{1/2}$$

$$\Rightarrow t_{1/2} = ln\frac{2}{\lambda} \quad , \quad t_{1/2} = T = ln\frac{0.69}{\lambda}$$

The total number of nuclei drops very rapidly at first, and then more slowly (Figure 2) . Graphical representation of the relationship: $N = N_0 \, e^{-\lambda t}$



3- Activity:

Activity is the number of decay per unit time given by definition by the following relationship:

$$A = -\frac{dN}{dT} = +\lambda\, t$$

As $N = N_0\, e^{-\lambda\, t}$ $\lambda\, N = \lambda\, N_0\, e^{-\lambda\, t}$

If we put : $A = \lambda\, N_0 \Rightarrow A = A_0\, e^{-\lambda\, t}$

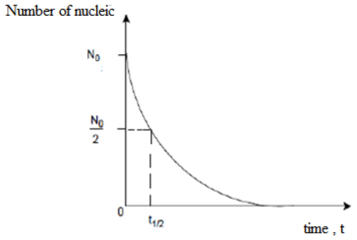
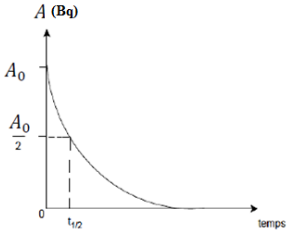
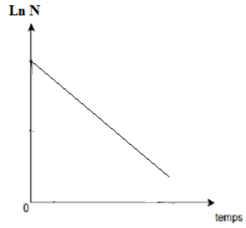
The unit of activity in the international system is the becquerel:

$$Becquerel = Bq = \frac{decqy}{Sec}$$

This unit has replaced another unit, Curie, with its symbol (Ci), where :

$$1\ \mathrm{Ci} = 3{,}7\,.\,10^{10}\ \mathrm{Bq}$$

4- Comparison graphs :

$N = N_0\, e^{-\lambda\, t}$	$A = A_0\, e^{-\lambda\, t}$	$LnN = LnN_0 - \lambda\, t$
		

Exercises of chapter 3

Exercise 1:

1- Complete the abbreviated notations and write the nuclear reactions:

$^{14}_{7}N(\alpha, P)^{17}_{8}O,\qquad ^{12}_{6}C(\text{d}, \text{n})^{13}_{7}N \quad , \quad ^{63}_{29}Cu(\text{p}, \text{n})^{63}_{30}N$

2- Complete the following nuclear reactions. For each equation, indicate the type of reaction involved:

a)- $^{2}_{1}H \quad + \ ^{3}_{1}H \ \rightarrow \ ^{4}_{2}He \ + \cdots$

b)- $^{235}_{92}U \quad + \ \ldots \ \rightarrow \ ^{144}_{54}Xe \ + \ ^{90}_{38}Sr \ + \ 2\,^{1}_{0}n$

c)- $^{2}_{1}H \quad + \ ^{3}_{2}He \ \rightarrow \ ^{4}_{2}He \ + \cdots$

f)- $^{235}_{92}U \ + \ ^{1}_{0}n \ \rightarrow \ ^{92}_{36}Kr \ + \ ^{141}_{56}Ba \ + \ \ldots \ ^{1}_{0}n$

Exercise 2 :

Calculate the energy released in the nuclear reaction given below: $^{10}_{5}B \ + \ _{1}^{2}D \ \rightarrow \ 3\,^{4}_{2}He + Q$

Given that m($^{10}_{5}B$) = 10.01294 u , m($^{2}_{1}D$)= 2.014103 u, and m($^{4}_{2}He$) = 4.002604 u.

Exercise 3 :

Americium-241 is an artificially produced radioactive element that emits α-particles. A sample of americium-241 of mass 5.1 μg is found to have an activity of 5.9×10^{5} Bq.

a)- Determine the number of nuclei in the sample of americium-241.

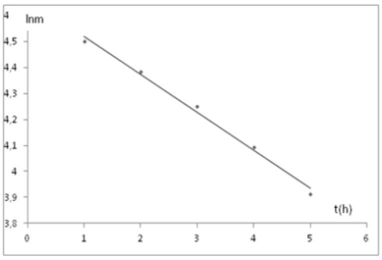
b)- Determine the decay constant of americium-241.

Exercise 4 :

1- From the law of radioactive decay; show the following relations:

$$A = A_0 e^{-\lambda t} \ , \ m = m_0 e^{-\lambda t}$$

2- We give the graph ln m= f(t):



- Express ln m as a function of m_0 and λ.

- Graphically determine the radioactive constant λ in hour^{-1} and the period $T(t_{1/2})$.

- Consider a sample of mass m_0=1g. Calculate the mass m disintegrated after 3 hours.

Corrected exercises of chapter 3

Exercise 1:

a- Complete the abbreviated notations and write the nuclear reactions:

$^{14}{}_{7}$N (α,P) $^{17}{}_{8}$O $^{14}{}_{7}$N + $_{2}{}^{4}$ He $\longrightarrow$ $^{17}{}_{8}$O + $^{1}{}_{1}$H

^{12}C (d,n)^{13}N : $^{12}{}_{6}$C + $_{1}{}^{2}$ H $\longrightarrow$ $^{13}{}_{7}$N + $_{0}{}^{1}$n

^{63}Cu (P, n)^{63}Zn : $^{63}{}_{29}$Cu + $^{1}{}_{1}$H $\longrightarrow$ $^{63}{}_{30}$Zn + $_{0}{}^{1}$n

b- Complete the following nuclear reactions. For each equation, indicate the type of reaction involved:

a)- ${}^{2}_{1}H + {}^{3}_{1}H \rightarrow {}^{4}_{2}He + \cdots$

${}^{2}_{1}H + {}^{3}_{1}H \rightarrow {}^{4}_{2}He + {}^{1}_{0}n$ .... **Fusion**

b)- ${}^{235}_{92}U + \cdots \rightarrow {}^{144}_{54}Xe + {}^{90}_{38}Sr + 2{}^{1}_{0}n$

${}^{235}_{92}U + {}^{1}_{0}n \rightarrow {}^{144}_{54}Xe + {}^{90}_{38}Sr + 2{}^{1}_{0}n$ .... **Fission**

c)- ${}^{2}_{1}H + {}^{3}_{2}He \rightarrow {}^{4}_{2}He + \cdots$

${}^{2}_{1}H + {}^{3}_{2}He \rightarrow {}^{4}_{2}He + {}^{1}_{1}H$.................. **Fusion**

f)- ${}^{235}_{92}U + {}^{1}_{0}n \rightarrow {}^{92}_{36}Kr + {}^{141}_{56}Ba + \ldots {}^{1}_{0}n$

${}^{235}_{92}U + {}^{1}_{0}n \rightarrow {}^{92}_{36}Kr + {}^{141}_{56}Ba + 3{}^{1}_{0}n$ .... **Fission**

<u>Exercise 2 :</u>

Calculate the energy released in the nuclear reaction given below :

$${}^{10}_{5}B + {}_{1}{}^{2}D \rightarrow 3\,{}^{4}_{2}He + Q$$

Given that m(${}^{10}_{5}B$) = 10.01294 u , m(${}^{2}_{1}D$)= 2.014103 u, and m(${}^{4}_{2}He$) = 4.002604 u.

1 u = 1.66054 . 10^{-27} Kg.

1 eV = 1.60217 .10^{-19} J

Δm = (10.01294 u+ 2.014103 u)- (3 x 4.002604 u) = 0.018785 u = 3.1193 .10^{-29} Kg.

E= Δm c² = 3.1193 .10^{-29} x (3.00 x 10^{8})^{2} ,

E= 2.8073 .10^{-12} J = 17.9 10^{6}= 17.9 MeV

<u>Exercise 3 :</u>

- Mass = 5.1 µg = 5.1 × 10^{-6} g
- Molecular mass of americium = 241g/mol
- N_A = **Avogadro constant**

a) Number of nuclei in the sample of americium-241.

$n=m/M= 5.1 \times 10^{-6}/ 241 = 2.11 .10^{-8}$ mol

Number of nucleis = $2.11 .10^{-8} \times 6.02 .10^{23} = 1.27 .10^{16}$

b) The decay constant of americium-241 **: Activity, A = λN**

A : Activity (decays/s)

λ: Decay constant s^{-1}

N : Atom number

$$\lambda = \frac{A}{N} = \frac{5.9 \times 10^5}{1.27 \, . \, 10^{16}} = 4.65 \, . \, 10^{-11} \, S^{-1}$$

1 Becquerel (Bq) = 1 decay/s

1Curie (Ci) = $3.7 \ 10^{10}$ decay/s

Exercise 4 :

1)-

$A = \lambda N, \qquad \Rightarrow \ N = \frac{A}{\lambda} \qquad , \qquad A_0 = \lambda N_0 \Rightarrow \qquad N_0 = \frac{A_0}{\lambda}$

$N = N_0 e^{-\lambda t}$, $\Rightarrow \qquad \frac{A}{\lambda} = \frac{A_0}{\lambda} \, . \, e^{-\lambda t}$

$\Rightarrow \ A = A_0 e^{-\lambda t}$

1 mol $\xrightarrow{\qquad\qquad}$ N_A

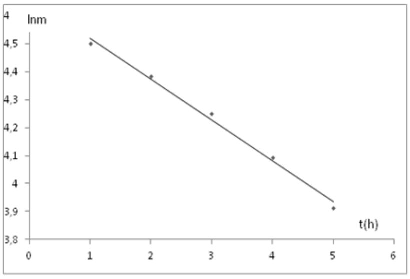
n mol (m/M) $\xrightarrow{\qquad\qquad}$ N

$N= n. N_A = m / M .N_A$ and $N_0= n_0. N_A = m_0 / M .N_A$

n_0: number of moles of radioactive element at t =0

So $N = N_0 e^{-\lambda t}$, Can be written : $m = m_0 e^{-\lambda t}$

2)- Graphically determine the radioactive constant λ in $hour^{-1}$ and the period $T(t_{1/2})$



$\ln m = \ln m_0 - \lambda t$,

from th graphe : y= b – ax

b= $\ln m_0$ and a= λ

$a = \frac{(4.4-4.1)}{(2-4)} = -0.15$ h^{-1}

$\lambda = 0.15$ h^{-1}

Express ln m as a function of m_0 and λ.

b)- $T(t_{1/2}) = \ln 2/\lambda = \ln 2/0.15 = 4.62$ h

- Consider a sample of mass m_0=1g. Calculate the mass m disintegrated after 3 hours.

The mass disintegrated *md* is given by: $m_d = m_0 - m_{(3h)}$

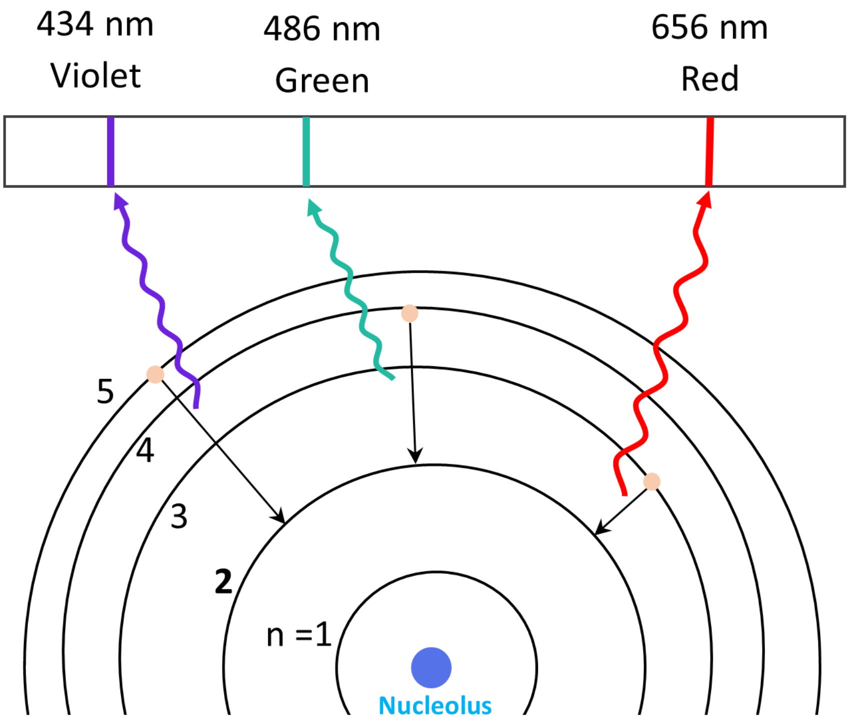
To find md, we need the value of λ from the graph.

$m(3h) = m_0 . e^{-\lambda.3} = 1 . e^{-\lambda.3h} = e^{-0.15 .3} = e^{-0.45} = 0.63$ g

$m_d = m_0 - m_{(3h)} = 1- 0.63 = 0.37$ g

Chapter 4

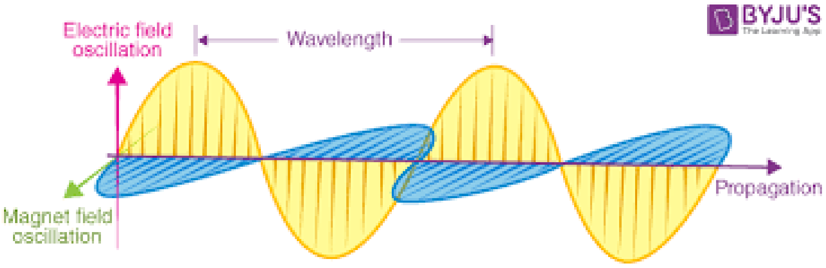
Electronic structure of the atom



I. Electromagnetic radiation

A **wave** is an oscillation or ***periodic movement*** that can transport energy from one point in space to another. Common examples of waves are all around us: Dropping a pebble into a pond causes waves to ripple outward along the water's surface. In this case, kinetic energy is transferred through matter (water) while the matter remains essentially in place.

Electromagnetic waves consist of an electric field oscillating in step with a perpendicular magnetic field, both of which are perpendicular to the direction of travel. These waves can travel through a vacuum at a constant speed of 2.998 × 10^8 m/s, the speed of light (denoted by c).



All waves, including forms of electromagnetic radiation, are characterized by, a wavelength (denoted by λ, the lowercase Greek letter lambda), a frequency (denoted by ν, the lowercase Greek letter nu), and an amplitude. As can be seen in Figure.., the wavelength is the distance between two consecutive peaks or troughs in a wave (measured in meters in the SI system).

The frequency is the number of wave cycles that pass a specified point in space in a specified amount of time (in the SI system, this is measured in seconds). A cycle corresponds to one complete wavelength. The unit for frequency, expressed as cycles per second [s−1], is the **hertz (Hz)**. Common multiples of this unit are megahertz, (1 MHz = 1 × 106 Hz) and gigahertz (1 GHz = 1 × 109 Hz). The amplitude corresponds to the magnitude of the wave's displacement and so, in **Figure 1**, this corresponds to one-half the height between the peaks and troughs. The amplitude is related to the intensity of the wave, which for light is the brightness, and for sound is the loudness

II- Properties of Electromagnetic Radiation:

- **Wavelength (λ)**: The distance between successive crests of the wave, measured in meters.
- **Frequency (υ)**: The number of wave cycles that pass a point in one second, measured in Hertz (Hz).
- **Energy (E)**: Proportional to the frequency and inversely proportional to wavelength; calculated as $E=h\times\upsilon$, where h is Planck's constant.
- **Speed (c)**: All EM waves travel at the speed of light in a vacuum, where $c=\lambda\times\upsilon$.

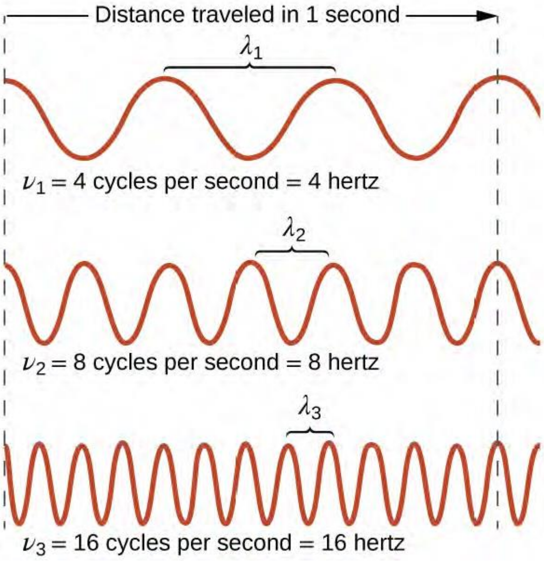


Figure IV- 1: One-dimensional sinusoidal waves show the relationship among wavelength, frequency, and speed..

The wave with the shortest wavelength has the highest frequency

The product of a wave's wavelength (λ) and its frequency (ν), $\lambda\nu$, is the speed of the wave (c) . Thus, for electromagnetic radiation in a vacuum:

$$c = \lambda.\upsilon = 2.998 \times 108 \text{ ms} - 1 = \text{const}$$

$$\lambda = \frac{c}{\upsilon}$$

Wavelength and frequency are inversely proportional: As the wavelength increases, the frequency decreases. The inverse proportionality is illustrated in **Figure 3**.

This figure also shows the electromagnetic spectrum, the range of all types of electromagnetic radiation.

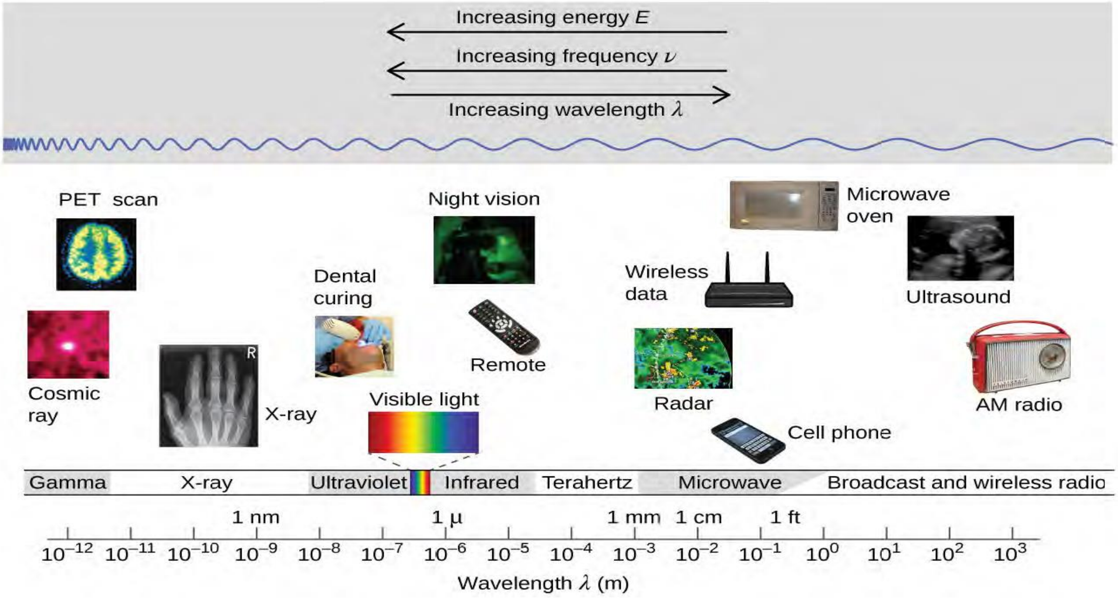


Figure IV- 2: The electromagnetic spectrum , (Wavelength and frequency)

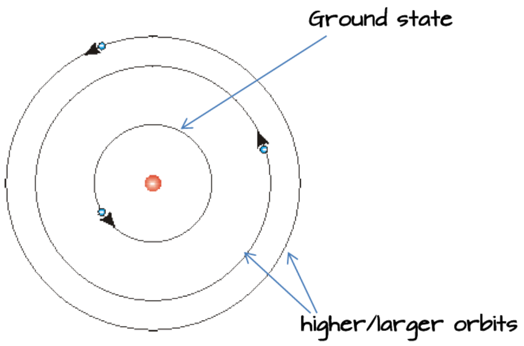
Electromagnetic radiation is classified into a spectrum based on wavelength and frequency. The main types include:

- **Radio Waves**: Longest wavelengths, used in communication.
- **Microwaves**: Used in radar, cooking, and some communication systems.
- **Infrared (IR)**: Experienced as heat; used in thermal imaging and remote controls.
- **Visible Light**: The narrow range of EMR detectable by the human eye.
- **Ultraviolet (UV)**: Beyond visible light, can cause skin tanning or burning.
- **X-rays**: High-energy radiation used in medical imaging.
- **Gamma Rays**: The shortest wavelengths and highest energies; produced by radioactive decay and used in cancer treatment.

III. Production of atomic emission spectra:

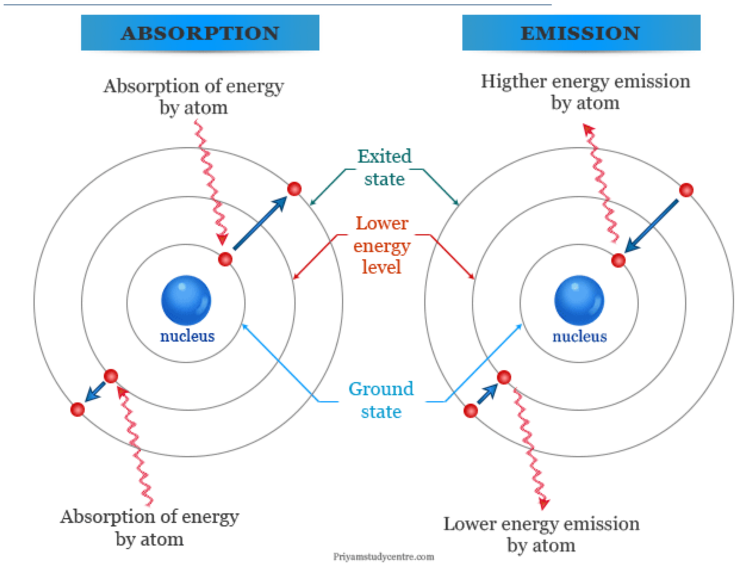
The electrons in an atom tend to be arranged in such a way that the energy of the atom is as low as possible. The **ground state** (The orbital closest to the nucleus) of an atom is the lowest energy state of the atom. When those atoms are given energy, the electrons absorb the energy and move to a higher energy level. These energy levels of the electrons in atoms are quantized, meaning again that the electron must move from one energy level to another in discrete steps rather than continuously. An **excited state** of an atom is a state where its potential energy is higher than the **ground state**. An atom in the excited state is not stable.

When it returns back to the ground state, it releases the energy that it had previously gained in the form of electromagnetic radiation.



- There are several orbitals surrounding the neucleus of an atoms, electrons occupy this orbitals n_1, n_2, n_3, n_4…….. (K,L, M,N ….).
- The energy levels are numbered E_1, E_2, E_3, E_4,…. $E_n = \frac{-13.6 .z}{n^2}$ (eV)

If an electron absorbs a photon, the electron will have the energy to move from a lower to a higher orbital.



$\Delta E = \Delta E_f - \Delta E_i > 0$ $\Delta E = \Delta E_f - \Delta E_i < 0$

Figure IV- 3 : Atomic emission and absorption

- The diagram above shows an electron in energy level 1 absorbing a photon, then moving to energy level. 2 (***Absorption, gain energy***).
- It then shows the electron dropping back down from level 2 to level 1 and emitting a photon (Emission***, lose energy***).

Thus, atoms can absorb or emit photons.

Photons are a form of electromagnetic energy that has a frequency and wavelength. Photons are made by the movement of electrons.

IV - Production of Hydrogen emission spectrum

Atomic emission spectra are produced when electrons in an atom absorb energy, move to higher energy levels, and then release that energy as they return to their ground states. Here’s how it works in detail:

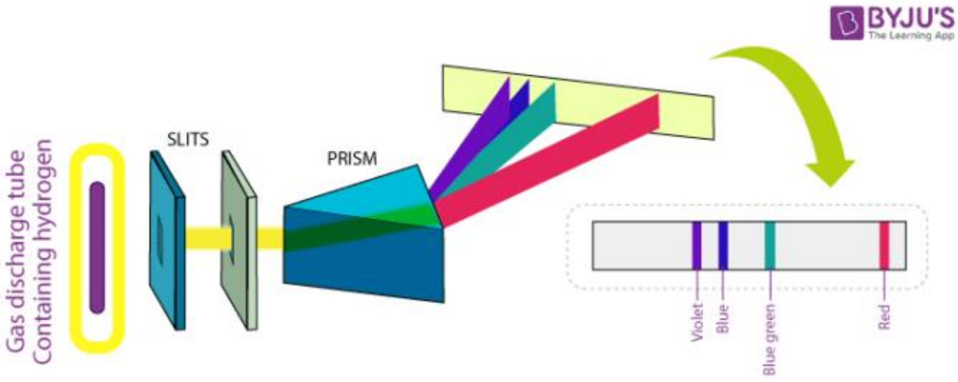


Figure IV- 4 : Hydrogen emission spectrum

1. **Energy Absorption**: When an atom gains energy from an external source (such as heat, electricity, or light), its electrons absorb this energy and jump to higher energy levels, or “excited states.”
2. **Emission of Light**: Excited states are unstable, so electrons quickly return to their lower energy levels. As they drop back down, they release energy in the form of photons (light particles). The energy (and hence, color) of the emitted photon depends on the difference between the energy levels.
3. **Formation of the Emission Spectrum**: When the emitted light passes through a spectroscope, it creates an emission spectrum. This spectrum appears as distinct lines of color, each corresponding to a specific wavelength of light emitted by the atom. The specific colors emitted by an element is called the emission spectrum.

V- The theory of photons: Emission spectrum of the hydrogen atom, Empirical Balmer-Rydberg relation

1- Rydberg relation :

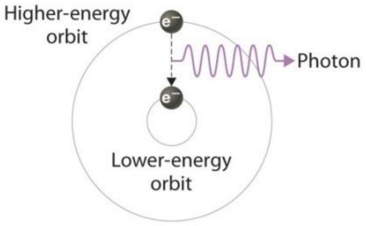
In the year 1885, on the basis of experimental observations, Balmer proposed the formula for correlating the wave number of the spectral lines emitted and the energy shells involved. This formula is given as:

$\frac{1}{\lambda} = \text{RH}\ (\frac{1}{n_1^2} - \frac{1}{n_2^2})$ with $n_2 = 2$, $\frac{1}{\lambda} = \text{RH}\ (\frac{1}{2^2} - \frac{1}{n_2^2})$

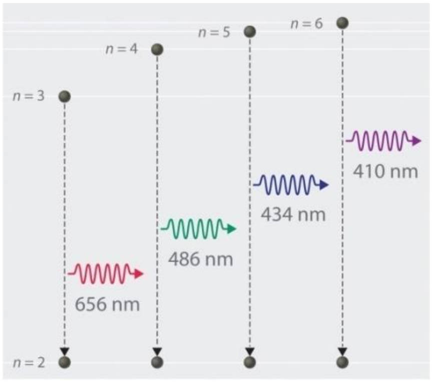
This series of the hydrogen emission spectrum is known as the Balmer series. This is the only series of lines in the electromagnetic spectrum that lies in *the visible region*.

The value, $1.097\ 10^7\ \text{m}^{-1}$, is called the Rydberg constant for hydrogen.

λ : The wavelength of light



(a) Electronic emission transition



(b) Balmer series transitions

$$\frac{1}{\lambda} = \text{RH}\ (\frac{1}{2^2} - \frac{1}{n_2^2})$$

electronic transition Balmer series	Wavelength (nm)	Color
n=3 to n= 2	656	red
n=4 to n =2	486	green
n=5 to n= 2	434	Indigo
n=6 to n =2	410	violet

Similarly, other transitions also have their own series names. Some of them are listed below:

- The transition from the first shell to any other shell – ***Lyman series***
- The transition from the second shell to any other shell – ***Balmer series***
- The Transition from the third shell to any other shell – ***Paschen series***
- The transition from the fourth shell to any other shell – ***Brackett series***
- The transition from the fifth shell to any other shell – ***Pfund series***

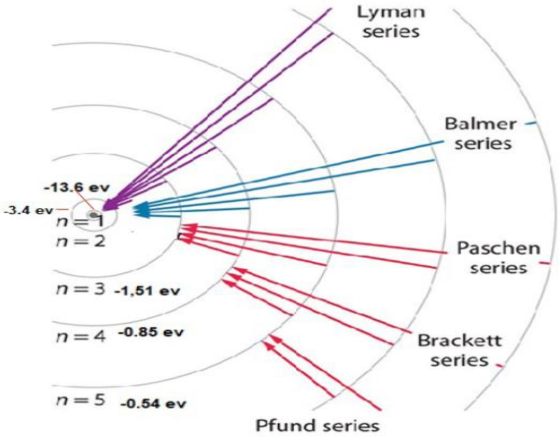


Table IV.1: Series of the emission spectrum of the hydrogen atom

Series	n_1	n_2	Spectral Domain
Lyman (1916)	1	≥ 2 (2,3,4,5….∝)	UV
Balmer (1885)	2	≥ 3 (3,4,5,6….∝)	Visible
Paschen (1908)	3	≥ 4 (4,5, 6, ….∝)	IR Near
Brackett (1922)	4	≥ 5 (5, 6,7,8….∝)	Near IR
Pfund (1924)	5	≥ 6 (6,7,8….∝)	Far IR

2- Bohr's model

The Danish physicist Niels Bohr, who first presented this model of the atom, based it on 3 fundamental postulates.

1- Electrons move around the nucleus in circular non-radiating orbits - called "stationary states".

2- An atom only emits or absorbs electromagnetic radiation when an electron makes a transition from one state to another.

3- In the atom the nucleus is immobile while the electron of mass m moves around the nucleus in a circular orbit of radius (r)

4- the angular momentum of the electron can only take integer values: $m_e\,.V.r = nh/2\pi$ with:

V: speed of the moving electron

h: Planck's constant

r: radius of the atom (electron-nucleus distance)

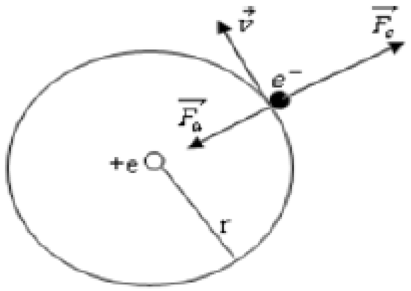
me: mass of the electron

5- During a transition between two stationary states of respective energies E_{n1} and E_{n2} there is emission or absorption of a quantity of energy equal to:

$|E_{n2} - E_{n1}| = \Delta E = h\,\upsilon = \frac{hC}{\lambda}$

With: h: 6.623.10-23 j.s (Planck constant) and υ: frequency of radiation involved.

3- Radius of stationary orbitals



According to Bohr's first postulate, the system is in equilibrium:

$$\overrightarrow{\mathrm{F_a}} + \overrightarrow{\mathrm{F_c}} = 0 \Rightarrow ||\overrightarrow{\mathrm{F_a}}|| = ||\overrightarrow{\mathrm{F_c}}||$$

with

$||\overrightarrow{\mathrm{F_c}}|| = \frac{m_e V^2}{r}$ and $\left|\left|\overrightarrow{\mathrm{F_c}}\right|\right| = K.\frac{|q||q'|}{r^2} = K.\frac{e^2}{r^2}$ with $K = \frac{1}{4\pi\varepsilon_0}$

We have :

$||\overrightarrow{\mathrm{F_a}}|| = ||\overrightarrow{\mathrm{F_c}}|| \Rightarrow \frac{m_e V^2}{r} = K.\frac{e^2}{r^2} \Rightarrow m_e V^2 = K.\frac{e^2}{r^2}$ …………………..(1)

According to Bohr's second postulate describing the quantization of orbital angular momentum, we have:

$m_e V r = \frac{n\,h}{2\pi} \Rightarrow (m_e V r)^2 = (\frac{n\,h}{2\pi})^2 \Rightarrow \; m_e\, V^2 = \frac{n^2\,h^2}{4\pi^2 \quad r^2\; m_e}$ …………………..(2)

Combining equation (1) with (2) leads to the expression of the orbital radius:

$$r = \frac{h^2}{4K\,\pi^2 \quad e^2\; m_e} n^2$$

If π, m_e, h, K and e are constants then r only depends on the value of the positive number n called the principal quantum number, i.e.:

$$r_n = \frac{h^2}{4K\,\pi^2 \quad e^2\; m_e} n^2$$

For n=1, first Bohr radius of the hydrogen atom which we note:

$$a_0 = \frac{h^2}{4K\,\pi^2 \quad e^2\; m_e} .1 = 0.529$$

With :

$K = \frac{1}{4\pi\varepsilon_0} = 9.10^9\,(\frac{N.m^2}{C^2}),\, m_e \; = 9.110\;10^{-31}\,Kg$ and $e = 1.602\;10^{-19}C$

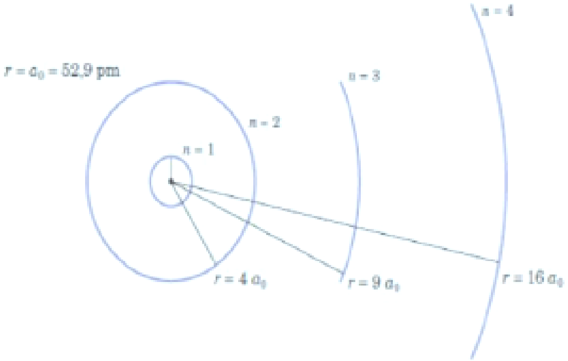
So: $\mathrm{r_n} = \mathrm{r_1} \; = \; 0.5290\; A^0$

$$\mathrm{r_n} = \mathrm{r_1}\,.n^2 = \; a_0.n^2 = 0.5290.n^2\;(A^0\,)$$

For n= 2, 2nd Bohr orbit $r_2 = 4.\ a_0$

For n= 3, 3rd Bohr orbit $r_3 = 9.\ a_0$

For n= 4, 4th Bohr orbit $r_3 = 16.\ a_{0;}$ …..



4- Energy of the electron on a stationary orbital

The total energy of the system considered is the sum of the potential energy EP and the kinetic energy EC:

$$E_T = E_P + E_C = -\frac{1}{2}\frac{1}{4\pi\varepsilon_0}.\frac{e^2}{r} = -\frac{1}{2}.\frac{Ke^2}{r}$$

replacing the expression for r in equation (53), the total energy of the system as a function of n will therefore be:

$$E_T = -\frac{1}{2}.\frac{Ke^2 4K\pi^2 m_e e^2}{h^2}.\frac{1}{n^2} = -\frac{2K^2 e^4 \pi^2 m_e}{h^2}.\frac{1}{n^2}$$

The energy $\mathbf{E_T}$ of the electron in the orbit depends only on n. It is therefore quantized and can only take a few particular values according to the expression:

$$E_n = -\frac{2K^2 e^4 \pi^2 m_e}{h^2}.\frac{1}{n^2}$$

For n= 1,

$$E_n = E_1 = -\frac{2K^2 e^4 \pi^2 m_e}{h^2}.\frac{1}{1^2} = -21{,}76\,.\,10^{-19}\,j = -13{,}6\text{ eV}$$

This value represents the energy of the ground state of the hydrogen atom (-13.6 eV) .

The ground state is the electronic state of minimum energy corresponding to n=1 and the states corresponding to a higher n are called the excitation state of an electron, having received an excess of energy.

We have:

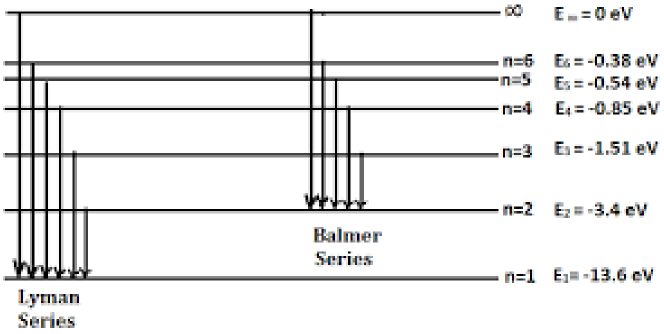
$$En = \frac{E_1}{n^2} = -\frac{13.6}{n^2}(eV)$$

The Bohr model justifies well that the energy of the electron is quantized, it only depends on quantum number n (0,1,2,3,…..,)

The different quantified states of energy are:

$$E1 = \frac{E_1}{1^2}, E2 = \frac{E_2}{2^2}, E3 = \frac{E_3}{3^2}, \ldots.\text{etc}$$

It is possible to calculate the energy of the various states in which the hydrogen atom can be found, these values are represented in the energy diagram below with the different series of the emission spectrum of the hydrogen atom. If we plot the different energies on a diagram, we obtain the energy diagram of the hydrogen atom.



VI- Electron Configurations

1- Shells (energy levels) : is the space around the nucleus in wich an electron can be found . The maximum number of electrons that can be accommodated in a shell is based on the principal quantum number (n). It is represented by the formula 2n2, where ‘n’ is the shell number. The shells, values of n, and the total number of electrons that can be accommodated are tabulated below.

Shell (‘n’ value)	**Maximum electrons present in the shell**
n=1 (K- shell)	2 x 1^2 = 2
n=2 (L- shell)	2 x 2^2 = 8
n=3 (M –shell)	2 x 3^2 = 18
n=4 (N shell)	2 x 4^2 = 32

2- Subshells

Within shells (energy levels) , we have subshells

- When n has a value of 1, one subshell is possible (S).
- When n has a value of 2, two subshells are possible (S, P).
- When n has a value of 3, tree subshells is possible (S, P and d).
- When n has a value of 4, four different subshells are possible (s, p, d, and f).
- Therefore, the s, p, d, and f subshells can accommodate a maximum of 2, 6, 10, and 14 electrons, respectively.

All the possible subshells for values of n up to 4 are tabulated below.

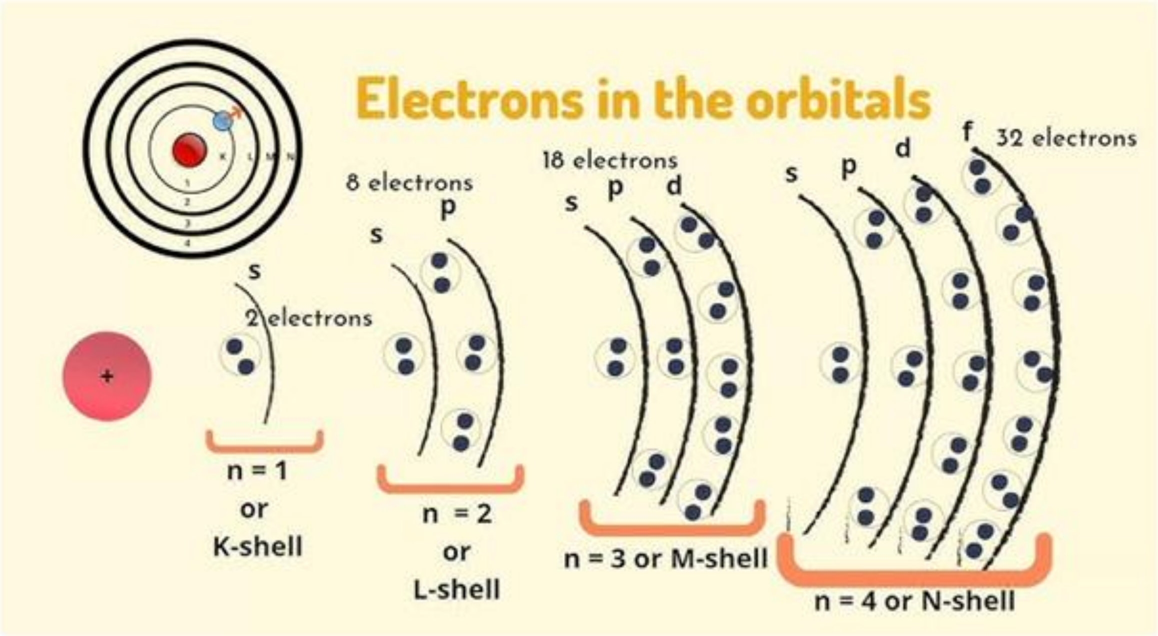
Principle Quantum Number Value	**Resulting Subshell in the Electron Configuration**
n=1	1s
n=2	2s
	2p
n=3	3s
	3p
	3d
n=4	4s
	4p
	4d
	4f

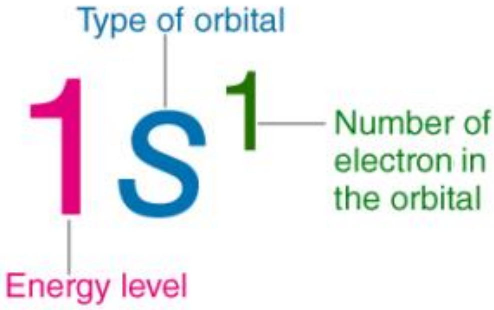
3- Orbitals:

As an atom gains electrons, they fill different orbitals sets according to a specific order. Each orbital can accommodate 2 electrons. The orbital sets are:

- ***The s orbital** set* contains a single orbital, and can hold a maximum of 2 electrons, so each s orbital set can hold 2 electrons.
- ***The p orbital** set* contains 3 orbitals, and thus can hold a total of 6 electrons.
- ***The d orbital** set* contains 5 orbitals, so it can hold 10 electrons.
- ***The f orbital** set* contains 7 orbitals, so it can hold 14 electrons.

subshells	**Number of orbitals**	**Number of electrons**
S	□	**1 x2 = 2 e**
P	□□□	**3 x2 = 6 e**
d	□□□□□	**5 x2 = 10 e**
f	□□□□□□□	**7 x2 = 14 e**





4- Quantum number:

the set of numbers used to describe the position and energy of the electron in an atom are called quantum numbers.

A quantum number is any of four numbers used to classify electrons and their atomic states. In particular, these numbers characterize the electron configuration of an atom. The four quantum numbers are: the principal quantum number, the azimuthal quantum number, the magnetic quantum number, and the spin quantum number.

a- Principal quantum number

The principal quantum number, n (= 1, 2, 3, ...) is used to label electron shells, i.e., to denote the overall size and energy of an electron shell. n may take on integer values from 1 to infinity.

b- Angular (Azimuthal) quantum number

The azimuthal quantum number, l (= 0, 1, 2, ..., n – 1), denotes the angular momentum and shape of an orbital (i.e., the most probable electron distribution). Azimuthal quantum numbers 0, 1, 2, and 3 correspond with the s, p, d, and f orbitals, respectively.

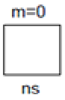
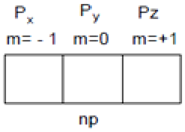
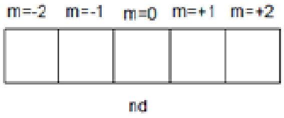
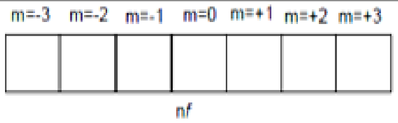
S	l=0
P	l=1
d	l=2
f	l=3

c- Magnetic quantum number

The magnetic quantum number, ml (= –l, –l + 1, ... 0, ... l – 1, l), denotes the orientation of an orbital and labels different orbitals within a subshell.

Its value is a function of the value of the secondary quantum number (l). $-l \leq ml \leq +l$

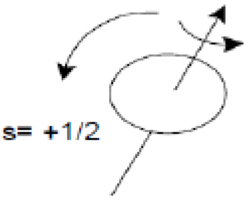
Graphically, this number is represented by a rectangle □ we represent as many rectangles as there are possible values of (m).

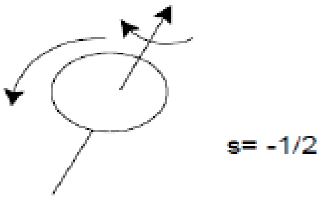
l=0	S	m=0	
l=1	P	m=+1, 0, -1	
l=2	d	m=+2, +1, 0, -1, -2	
l=3	f	m=+3, +2, +1, 0, -1, -2, -3	

d- Spin quantum number

It characterizes the movement of the electron on itself and can take only two different values ms (± ½)

$\boxed{\uparrow\downarrow}$ $S = +\frac{1}{2}: \uparrow$; $S = -\frac{1}{2}: \downarrow$





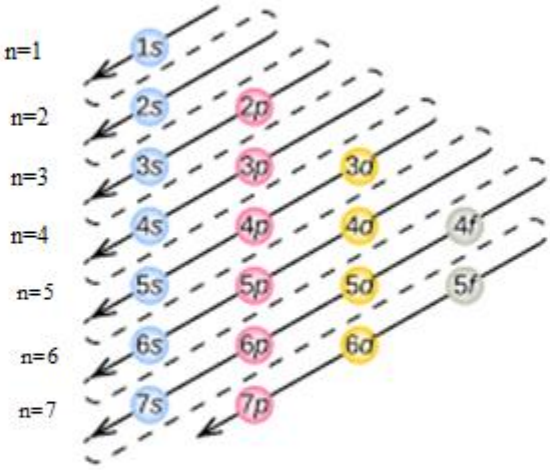
5- Electronic configuration of the elements

Electronic configuration is the arrangement of electrons in an atom's orbitals. It indicates how electrons are distributed among different energy levels and sublevels around the nucleus of an atom. Understanding electronic configuration helps predict an element's chemical properties, reactivity, and the type of bonds it can form. The configuration is constructed by following the following rules of thumb:

a- **Construction (or filling) rule :** Orbitals are filled starting with the lowest energy orbital first, progressing to higher energy orbitals.

b- **Klechkowski's rule :** is a guideline used to determine the order in which atomic orbitals are filled with electrons. It is based on the quantum numbers nnn (principal quantum number) and ℓ (azimuthal quantum number) and is commonly used alongside the Aufbau principle.

A simple way to visualize Klechkowski's rule is by using a **diagonal filling diagram**:



X: 1s 2s 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p……………

Example :

$_{6}C$: $1s^2 2s^2 2p^2$

$_{28}Ni$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$

$_{18}Ar$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

c- Aufbau Principle (Building-Up Principle)

Electrons fill the atomic orbitals in order of increasing energy levels. Lower-energy orbitals are filled first, and then higher-energy orbitals are filled.

The energy of an orbital is determined by the sum $n+\ell$:

- n: Principal quantum number (energy level).
- ℓ: Azimuthal quantum number (sublevel type: s=0, p=1, d=2 , f=3f).

If two orbitals have the same n+ℓ, the orbital with the lower nnn is filled first.

d- Order of orbital filling:

1s < 2s< 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p…....

e- Pauli Exclusion Principle

An atomic orbital can hold a maximum of **two electrons**, and they must have opposite spins. This means that in any given orbital, one electron will have spin +1/2 (up) and the other −1/2 (down).

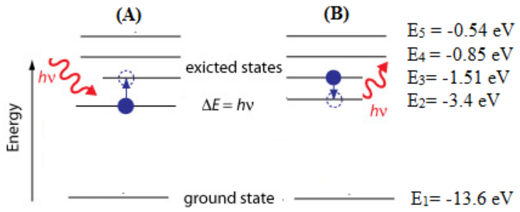
f- Hund’s Rule

For orbitals of the same energy (degenerate orbitals, such as the three ppp-orbitals or five ddd-orbitals), electrons are distributed so that the number of unpaired electrons is maximized. Each orbital is first singly occupied by electrons with parallel spins before any pairing occurs.

Exercises of chapter 4

Exercise 1:

Consider the following energy level diagram



1 - Specify whether the transitions (A) and (B) correspond to emitted or absorbed radiation and explain.

2 - Calculate the wavelength corresponding to of these radiations.

Exercise 2:

A hydrogen atom initially in the ground level absorbs a photon, Which excites it to n=4 level.

1- What is the corresponding change in energy in Joule?

2- Determine the wavelength and frequency of photon.

3- If the wavelength of the first line of the Balmer series of the hydrogen optical spectrum is 6562.8 A^0 , calculate the value of the Rydberg constant (R_H) in m^{-1} .

Data: $C = 3x10^8$ m/s, $h = 6.62x10^{-34}$ J.s

Exercise 3 :

Based on Bohr's atomic model of the circular orbits described by the electron around the nucleus of the hydrogen atom, determine the expression for the radius of the atom, as well as the speed and energy of the electron in its n orbit.

1. Calculate the values of E, v, and r of the ground state and write the expressions of En, v_n , r_n as a function of E_1 , v_1 , r_1 , respectively.

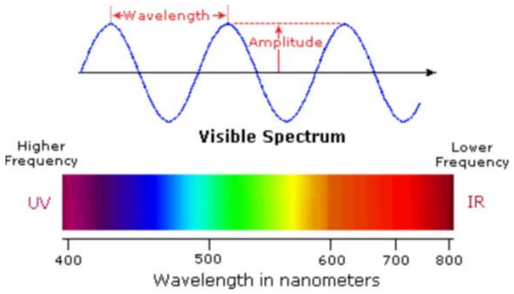
2. Calculate E_n , v_n , r_n for the energy levels: n = 2,3 , 4 (radius in A^0 ", energy in eV).

3. Plot on an energy diagram all the results obtained in the previous questions.

4. Plot on the same diagram the electronic transitions for the Lyman and Balmer spectral series.

Exercise 4 :

In the hydrogen emission spectrum, the four visible lines corresponds to the n =3, 4, 5 and 6 to n=2 electronic transitions. Using Absorption Spectrum diagram presented below, what colors light are these transitions? And deduce the energy of each of these transitions. Data: C = 3x10^8 m/s, h = 6.62x10^{-34} J.s

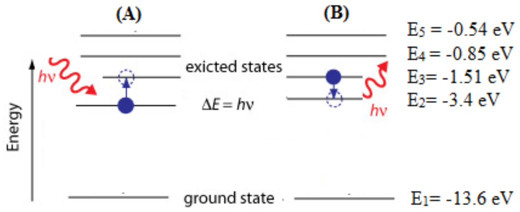


- **Violet:** 400 - 420 nm
- **Indigo:** 420 - 440 nm
- **Blue:** 440 - 490 nm
- **Green:** 490 - 570 nm
- **Yellow:** 570 - 585 nm
- **Orange:** 585 - 620 nm
- **Red:** 620 - 780 nm

Corrected exercises of chapter 4

Exercise 1:

Consider the following energy level diagram



1- (A) corresponds to absorbed radiation.

(B) corresponds to emitted radiation.

For (A) : $\Delta E = E_2 - E_3 = -1.51 - (-3.4) = +\ 1.89$ eV=3.02 10^{-19} J > 0 energy gain

For (B) : $\Delta E = E_3 - E_2 = -3.4 - (-1.51) = -\ 1.89$ eV < 0 energy loss

2-

1 ev = 1.602 10^{-19} J

$$\Delta E = h\,.\frac{c}{\lambda} \quad \Rightarrow \lambda = h\,.\frac{c}{\Delta E} = \frac{6.62.\ 10^{-34}\ 3.10^{8}}{3.02\ .10^{-19}} = 6.57\ 10^{-7}\ m = 657\ nm$$

Exercise 2:

1- $n_1 = 1$ (ground level), $n_2 = 4$

$$\Delta E = E_4 - E_1 = -\frac{13.6}{n_2^2} - \left(-\frac{13.6}{n_2^2}\right) = -\frac{13.6}{4^2} - \left(-\frac{13.6}{1^2}\right)$$

$$\Delta E = 13.6\left(\frac{1}{1^2} - \frac{1}{4^2}\right) = 13.6\ x\ \frac{15}{16} = 12.75\ ev = 12.75\ x\ 1.6\ 10^{-19}$$

$$\Delta \mathbf{E} = \mathbf{20.4\,.\,10^{-19}\ J}$$

2- $\Delta \mathbf{E} = \mathbf{h\,.\,\upsilon = h\,.\frac{c}{\lambda} \Rightarrow \lambda = \frac{h\,.c}{E} = \frac{6.62.\ 10^{-34}\ 3.10^{8}}{20.4\ .10^{-19}} = 9.77\ 10^{-8}\ m = 97.7 nm}$

$$\upsilon = \frac{\mathbf{c}}{\lambda} \Rightarrow \upsilon = \frac{\mathbf{3.10^{8}}}{\mathbf{9.7\ 10^{-8}}} = 3.09\mathrm{x}\ \mathbf{10^{15}}\mathrm{Hz}$$

3- The Balmer series corresponds to electronic transitions in the hydrogen atom where the final energy level is n=2. The first line of the Balmer series corresponds to a transition from n=3 to n=2

We are given λ = 6562.8 Å = 6562.8 x 10^{-10} m. We can plug in the values into the Rydberg formula:

$\frac{1}{6562.8\ \mathrm{x}\ 10^{-10}\ \mathrm{m}} = \mathrm{RH\ x}\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$, RH ≈ 1.097 x 10^7 m^{-1}

Exercise 3:

$$\mathrm{E_n} = \frac{-13.6\,.\,\mathrm{z}}{\mathrm{n}^2}, \mathrm{r_n} = \frac{\mathrm{a_0\ n^2}}{\mathrm{z}}, \mathrm{V_n} = \frac{2{,}18\,.\,10^6\,.\,\mathrm{z}}{\mathrm{n}} \text{ , for hydrogen atom z = 1 ,}$$

$\mathbf{E_n} = \frac{\mathbf{-13.6}}{\mathbf{n^2}}$, $\mathbf{r_n = a_0\ n^2}$ **, with** $\mathbf{a_0 = 0.5290}$ **, so** $\mathbf{r_n = 0.5290\ n^2}$ **,** $\mathbf{V_n} = \frac{\mathbf{2{,}18\ .10^6}}{\mathbf{n}}$

Ground state n=1 : $\mathbf{E_1 = -13.6\ eV}$, $\mathbf{r_1 = 0.5290}$ Å , $\mathbf{V_1 = 2{,}18\,.\,10^6\ \ m/s}$

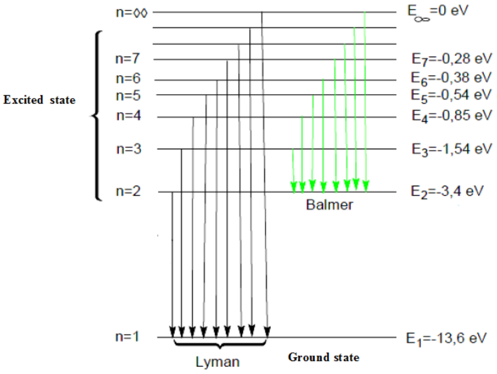
2)

n=2 $\mathbf{E_2} = -3.4$ eV, $\mathbf{r_2} = 0.5290\ 2^2 = 2.12$ Å , $\mathbf{V_2} = 1{,}09\,.\,10^6$ m/s

n=3 $\mathbf{E_3} = -1.51$ eV, $\mathbf{r_3} = 0.5290\ 3^2 = 4.76$ Å , $\mathbf{V_3} = 7{,}3\,.\,10^5$ m/s

n=4 $\mathbf{E_4} = -0.85$ eV, $\mathbf{r_4} = 0.5290\ 4^2 = 8.47$ Å , $\mathbf{V_4} = 8.47\,.\,10^6$ m/s

3)



Exercise 4 :

This result is obtained by calculating the wavelength of the light emitted during the transition using the Rydberg formula.

$$\frac{1}{\lambda} = \mathrm{RH}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) = \frac{1}{\lambda} = 1.09710^7 \ (\frac{1}{3^2} - \frac{1}{4^2})$$

n=3 to n= 2 : $\frac{1}{\lambda} = \mathrm{RH}\left(\frac{1}{2^2} - \frac{1}{3^2}\right) \rightarrow \frac{1}{\lambda} = 1.09710^7 \ \left(\frac{1}{4} - \frac{1}{9}\right), \lambda = 6.56 \,.10^{-7}$ m

n=4 to n= 2 : $\frac{1}{\lambda} = \mathrm{RH}\left(\frac{1}{2^2} - \frac{1}{4^2}\right) \rightarrow \frac{1}{\lambda} = 1.09710^7 \ \left(\frac{1}{4} - \frac{1}{16}\right), \lambda = 4.86 \,.10^{-7}$ m

n=5 to n =2 : $\frac{1}{\lambda} = \mathrm{RH}\left(\frac{1}{2^2} - \frac{1}{5^2}\right) \rightarrow \frac{1}{\lambda} = 1.09710^7 \ \left(\frac{1}{4} - \frac{1}{25}\right), \lambda = 4.34 \,.10^{-7}$ m

n=6 to n= 2 : $\frac{1}{\lambda} = \mathrm{RH}\left(\frac{1}{2^2} - \frac{1}{6^2}\right) \rightarrow \frac{1}{\lambda} = 1.09710^7 \ \left(\frac{1}{4} - \frac{1}{36}\right), \lambda = 4.10 \,.10^{-7}$ m

In the table below we summarize the wavelength, colors and energy results for these lines

electronic transition	**Wavelength**		**Color**	**Energy (J)**
	(m)	**(nm)**		
n=3 to n= 2	6.56 x 10^{-7}	656	*red*	3.03 x 10^{-19}
n=4 to n =2	4.86 x 10^{-7}	486	*blue*	4.09 x 10^{-19}
n=5 to n= 2	4.34 x 10^{-7}	434	*Indigo*	4.57 x 10^{-19}
n=6 to n =2	4.10 x 10^{-7}	410	*violet*	4.84 x 10^{-19}

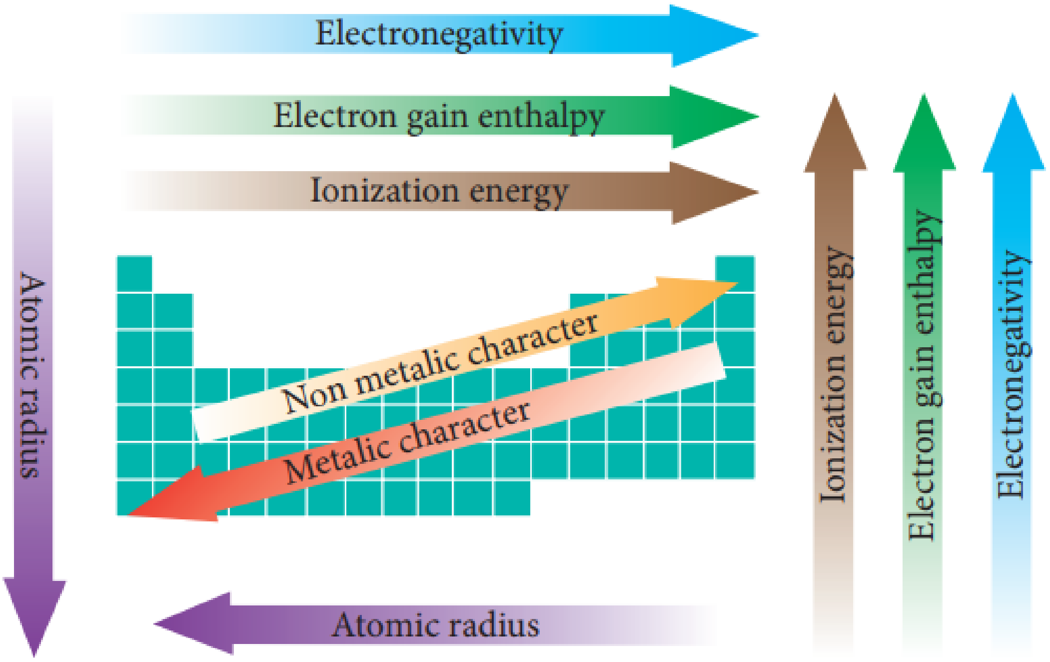
2) Example of Energy calculation:

where h = 6.626 x 10^{-34}J.s, and c = 3 x 10^8 m/s

$$\mathbf{E = h\,\upsilon = h.\frac{c}{\lambda}\ ,\ E = \frac{h\,.c}{\lambda} = \frac{6.62.\ 10^{-34}\ 3.10^{8}}{6.56\ \ 10^{-7}} = 3.027\ 10^{-19}\ J}$$

Chapter 5:

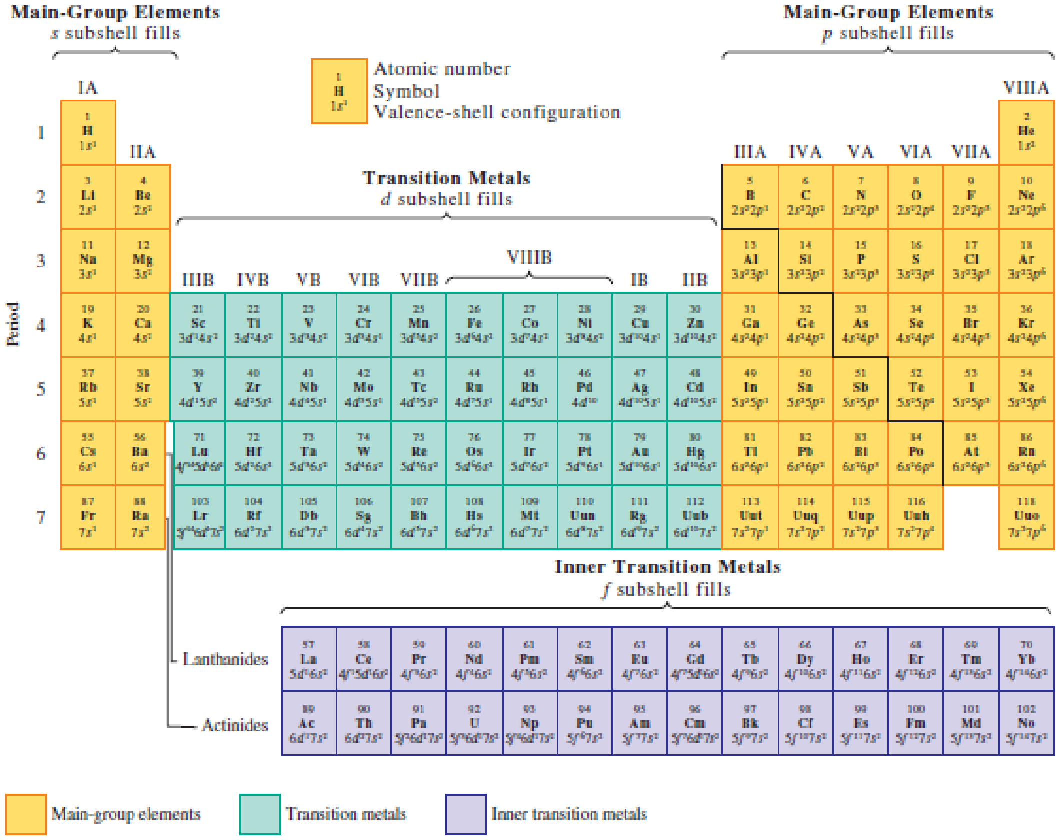
Periodic classification of elements



I. Description of Mendeleev's Periodic Table

As early chemists worked to purify ores and discovered more elements, they realized that various elements could be grouped together by their similar chemical behaviors.

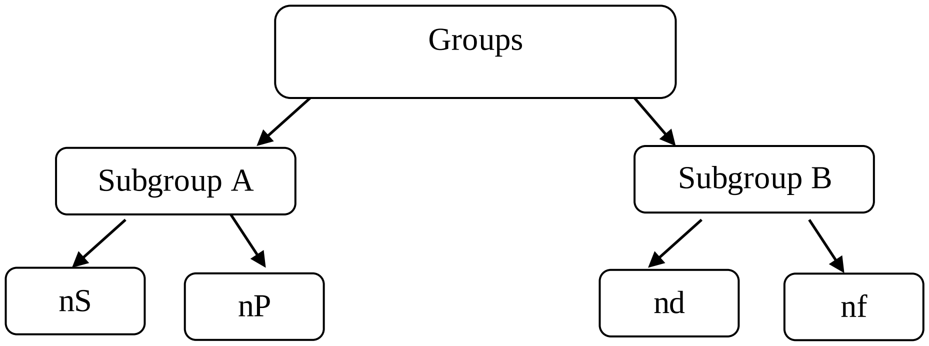
Dimitri Mendeleev in Russia (1869) and Lothar Meyer in Germany (1870) independently recognized that there was a periodic relationship among the properties of the elements known at that time.



1- **Period:** The period (horizontal row) corresponds to an electronic layer identified by its principal quantum number noted (n). There are seven known electronic layers in the ground state, therefore seven periods.

2- **Group**: There are eighteen groups including eight in subgroup A and eight in subgroup B. The number of valence electrons (electrons in the outer shell) corresponds to the group number.

Groups are further divided into two subgroups A and B:



3- **Blocks:**

- ***Block (s):***

- **Elements:** Groups 1 and 2, along with hydrogen (H) and helium (He).
- **Electron Configuration:** Outer electrons in the **s orbital** (ns^1 or ns^2).

- ***Block (p):***

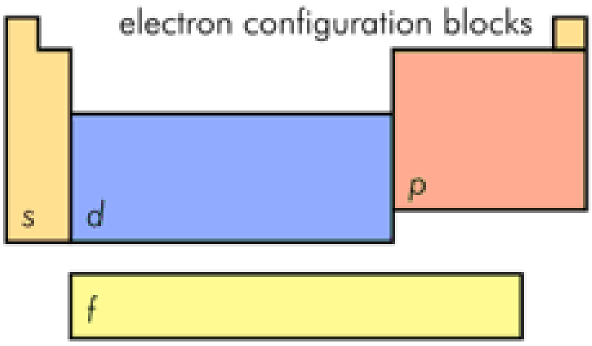
- **Elements:** Groups 13 to 18 (right side of the periodic table).
- **Electron Configuration:** Outer electrons in the **p orbital** (np).

- ***Block (d):***

- **Elements:** Transition metals (Groups 3 to 12).
- **Electron Configuration:** Outer electrons in the **d orbital** ((n-1)d).

- ***Block (f):***

- **Elements:** Lanthanides (58–71) and Actinides (90–103).
- **Electron Configuration:** Outer electrons in the **f orbital** ((n-2)f).



4- Families :

- **Group I$_A$ (Colum 1) :** *(ns^1) (Alkali Metals):* Lithium (Li), Sodium (Na), Potassium (K), etc.
- **Group II$_A$ (Colum 2) :** (ns^2) (*Alkaline Earth Metals*): Magnesium (Mg), Calcium (Ca), Barium (Ba), etc.
- **Group III$_A$ (Colum 13) :** (ns^2 np^1) (*Boron Family*): Boron (B), Aluminum (Al), etc.
- **Group IV$_A$ (Colum 14) :** (ns^2 np^2) (*Carbon Family*): Carbon (C), Silicon (Si), etc.
- **Group V$_A$ (Colum 15) :** (ns^2 np^3) (*Nitrogen Family*): Nitrogen (N), Phosphorus (P), etc.
- **Group VI$_A$ (Colum 16) :** (ns^2 np^4) (*Oxygen Family*): Oxygen (O), Sulfur (S), etc.
- **Group VII$_A$ (Colum 17) :** (ns^2 np^5) (*Halogens Family*): Fluorine (F), Chlorine (Cl), etc.
- **Group VIII$_A$ (Colum 18) :** (ns^2 np^6) (*Noble Gases Family*): Helium (He), Neon (Ne), etc.
- **Group I$_B$ to Group VIII$_B$ (Colum 3 to 12) :** *Transition metals family* : Copper (Cu) , iron (Fe) , etc.
- ***Example***:
- ${}_{26}Fe$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ $\mathbf{4s^23d^6}$
- Principal quantum number n= 4 → ***Period***= 4
- Number of valence electrons = 8 → ***Group***= 8 ($4s^23\underline{\mathbf{d^6}}$ **,** subgroup = B)
- ***Block*** ($4s^23\underline{\mathbf{d^6}}$ **,** block d)
- ***Family*** : transition metals

II- Shorthand electronic configurations:

Shorthand electron configuration of an element starts with the symbol of the noble gas in the previous period, followed by the rest of the electron configuration for the element in question For example, Chlorine has the atomic number 17, the electron configuration of aluminum can be written two ways:

${}_{17}Cl$: $1s^22s^22p^63s^23p^5$. The part of the electron configuration that is blue is the electron configuration of neon, the noble gas in the previous period.

Chlorine is in period 3, and the noble gas at the end of the previous period (period 2) is neon.

Its shorthand electron configuration is [Ne] $3s^23p^5$

Element	Structure	period	Group/subgroup	Bloc	family
$_{10}$Ne	$1s^22s^22p^6$	2	VIII$_A$	P	Noble Gases
$_{17}$Cl	$[Ne]3s^23p^5$	3	VII$_A$	P	Halogens
$_{24}$Cr	$[Ar]\ 4s^13d^5$	4	VI$_A$	d	transition metals

III-Exceptions to Electron Configurations:

Exceptions in electron configurations occur due to the stability achieved by half-filled or fully-filled sublevels.

The electron configurations of most transition metals are straightforward, but chromium and copper present intriguing deviations.

- ***Chromium ($_{24}$Cr)***

Expected Configuration: $[Ar]_{18}\ 4s^2\ 3d^4$

Actual Configuration: $[Ar]_{18}\ 4s^1\ 3d^5$

By moving one electron from the 4s orbital to the 3d orbital, chromium achieves a half-filled 3d sublevel, which provides extra stability.

- ***Copper ($_{29}$Cu)***

Expected Configuration: $[Ar]_{18}\ 4s^2\ 3d^9$

Actual Configuration: $[Ar]_{18}\ 4s^1\ 3d^{10}$

By moving one electron from the 4s orbital to the 3d orbital, copper achieves a fully-filled 3d sublevel, which is more stable.

IV- Electron configurations of cations

– For s- and p-elements, electrons are lost first from the np subshell followed by the ns subshell

$_{38}Sr \rightarrow [Kr]_{36}\ 5s^2$, $Sr^{2+} \rightarrow [Kr]$

$_{31}Ga \rightarrow [Ar]_{18}\ 4s^23d^{10}4p^1$, $Ga^{3+} \rightarrow [Ar]3d^{10}$

$_{26}Fe \rightarrow [Ar]_{18}\ 4s^23d^6$, $_{26}Fe^{2+} \rightarrow [Ar]_{18}\ 4s^03d^6$: $[Ar]_{18}\ 3d^6$

V- Electron configurations of anions:

– Electrons are added until a noble-gas configuration is reached

$_7N \rightarrow [He]_2\ 2s^22p^3$; $_7N^{3-} \rightarrow [He]_2\ 2s^22p^6$

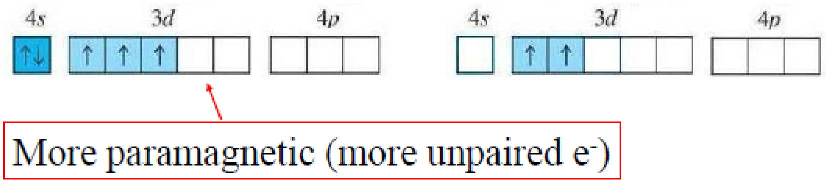
$_{15}P \rightarrow [Ne]_{10}\ 3s^2\ 3p^3$; $_{15}P^{-3}\ [Ne]_{10}\ 3s^2\ 3p^6$

VI- Magnetic properties of atoms and ions:

– Species with unpaired electrons are paramagnetic (attracted by magnetic fields) – Species having all electrons paired are diamagnetic (not attracted by magnetic fields)

Example: Write the electron configurations of V and V^{3+} and determine which species is more paramagnetic.

V → [Ar]$4s^2 3d^3$ **V^{3+} → [Ar]$3d^2$**



More paramagnetic (more unpaired e$^-$)

VII- Periodicity of Properties

The periodicity in Mendeleev's table arises from the recurrence of similar chemical properties at regular intervals when elements are arranged by increasing atomic mass.

a- Atomic Radius: Decreases across a period and increases down a group.

The atomic radius decreases across a period due to increasing nuclear charge pulling electrons closer, and it increases down a group because additional electron shells place outer electrons farther from the nucleus.

Let us take as an example the elements of the same period the third period n=3:

Element	Na	Mg	Al	Si	P	S
Z	11	12	13	14	15	16
Atomic radius (A°)	1.90	1.60	1.46	1.32	1.28	1.27

Let's take as an example the elements of the same group IV$_B$:

Element	Z	Atomic radius (A°)
C	6	0.914
Si	14	1.32
Ge	32	1.37
Sn	50	1.62
Pb	82	1.75

b- **Ionization Energy:** Increases across a period and decreases down a group. Ionization energy increases across a period as the effective nuclear charge strengthens, making it harder to remove an electron. It decreases down a group because outer electrons are farther from the nucleus and experience less attraction, requiring less energy to remove.

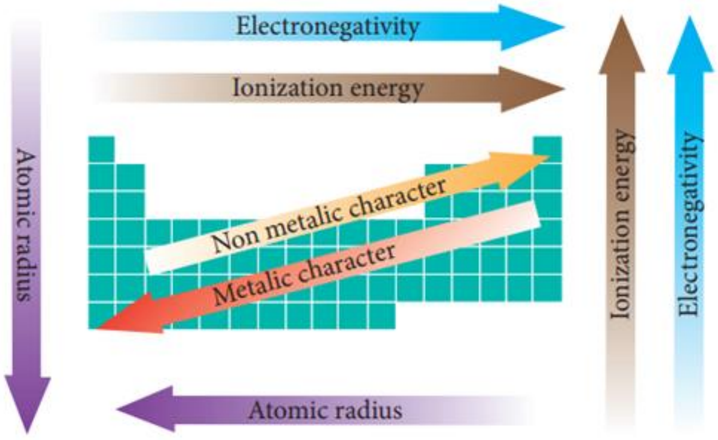
c- **Electronegativity:** Tends to increase across a period and decrease down a group. Electronegativity increases across a period due to the stronger effective nuclear charge, which enhances an atom's ability to attract electrons. It decreases down a group as the increased distance between the nucleus and bonding electrons reduces the attraction.

For the same period:

Element	Na	Mg	Al	Si	P	S
Z	11	12	13	14	15	16
Electronegativity (χ)	0.93	1.31	1.61	1.90	2.19	2.58
Ionization Energy	495.8	737.7	577.5	786.5	1011.8	999.6

For the some group:

Element	Z	Electronegativity (χ)	Ionization Energy
O	8	3.44	1313.9
S	16	2.58	999.6
Se	34	2.55	941.0



Exercises of chapter 5

Exercise 1:

An element X with atomic number Z = 16.

1. Draw up the electronic structure of the corresponding atom in its ground state.
2. Write the shorthand electron configuration
3. Deduce the period and column of the classification to which X belongs.
4. Find the name and symbol of this element.

Exercise 2:

Identify the atoms from the electron configurations given:

(1) $[Ar]4s^2 3d^5$

(2) $[Kr]5s^2 4d^{10} 5p^6$

Exercise 3:

1. Let be the following atoms:

$_{26}Fe$, $_{26}Fe^{2+}$, $_{26}Fe^{3+}$, $_{17}Cl$, $_{17}Cl^-$, $_{29}Cu$, $_{29}Cu^{2+}$, p, p^{-3}

Give the electronic configurations of the atoms and write their shorthand electron configuration.

Exercise 4:

What is the electron configuration and orbital diagram for an aluminum atom? What are the four quantum numbers for the last electron added?

Exercise 5:

Give electronic configuration and deduce the atomic number of the following elements :

- Oxygen: 2nd period, group VI_A (16th column), block p.

- Chlorine: 3rd period, group VII_A (17th column) , block p.

- Manganese: 4th period, group VII_B (7th column), block d.

Exercise 6:

1- Classify the following elements: $_{55}Cs$, $_{9}F$, $_{19}K$, $_{3}Li$, $_{7}N$ by decreasing atomic radius, justifying your answer.

2- Assign to each element among the following $_{37}Rb$; $_{11}Na$; $_{9}F$; $_{8}O$; $_{7}N$; $_{3}Li$; $_{2}He$; $_{1}H$ the electronegativity that corresponds to it: χ=0.9; 4.0; 2.1; 0; 0.8; 3.0; 1.0; 3.5

Corrected exercises of chapter 5

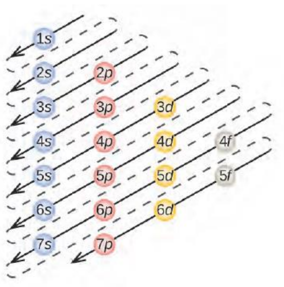
Exercise 1:

An element X with atomic number Z = 16.

5. Draw up the electronic structure of the corresponding atom in its ground state.

- $1s^{2}\ 2s^{2}\ 2p^{6}\ 3s^{2}\ 3p^{4}$

6. Write the shorthand electron configuration : $[\,Ne]_{10}\,3s^{2}\ 3p^{4}$



7. Deduce the period and column of the classification to which X belongs.

- Period :3 , VII_{A}, block p, family : oxygen

8. Find the name and symbol of this element: X is Sulfur “S”

Exercise 2:

Identify the atoms from the electron configurations given :

(1) $[Ar]_{18}\,4s^{2}3d^{5}$: Period :4 , VII_{B}, block d , family : transition metals :

Using table periodic we can Identify the atom , so, it is **:** **<u>Mn</u>**

(2) $[Kr]_{36}5s^{2}4d^{10}5p^{6}$: Period :5 , $VIII_{A}$, block p, family : rare gases : **<u>(b) Xe</u>**

Exercise 3:

1. Let be the following atoms:

$_{26}Fe$, $_{26}Fe^{2+}$, $_{26}Fe^{3+}$, $_{17}Cl$, $_{17}Cl^{-}$, $_{29}Cu$, $_{29}Cu^{2+}$, p, p^{-3}

element	electronic configurations	Shorthand electron configuration.	Classification
$_{26}Fe$	$1s^{2}\ 2s^{2}\ 2p^{6}\ 3s^{2}\ 3p^{6}\ 4s^{2}3d^{6}$	$[Ar]_{18}\,4s^{2}3d^{6}$	Period :4 , $VIII_{B}$, block d, family : transition metals
$_{26}Fe^{2+}$	$[Ar]_{18}\,4s^{0}3d^{6}$: $[Ar]_{18}\,3d^{6}$		
$_{26}Fe^{3+}$	$[Ar]_{18}\,4s^{0}3d^{5}$: $[Ar]_{18}\,3d^{5}$		
$_{17}Cl$	$1s^{2}\ 2s^{2}\ 2p^{6}\ 3s^{2}\ 3p^{5}$	$[Ne]_{10}3\,s^{2}\ 3p^{5}$	Period :3 , VII_{A}, block p, family : halogens
$_{17}Cl^{-}$	$[Ne]_{10}3\,s^{2}\ 3p^{6}$		
$_{29}Cu$	$1s^{2}\ 2s^{2}\ 2p^{6}\ 3s^{2}\ 3p^{6}\ 4s^{1}\ 3d^{10}$	$[Ar]_{18}\,4s^{1}\ 3d^{10}$	Period :4 , I_{B}, block d, family : transition metals
$_{29}Cu^{2+}$	$[Ar]_{18}\,4s^{0}\ 3d^{9}$		
$_{15}p$	$1s^{2}\ 2s^{2}\ 2p^{6}\ 3s^{2}\ 3p^{3}$	$[Ne]_{10}3\,s^{2}\ 3p^{3}$	Period :3 , V_{A}, block P, family : Nitrogen
$_{15}p^{-3}$	$[Ne]_{10}3\,s^{2}\ 3p^{6}$		

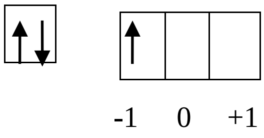
Exercise 3 :

What is the electron configuration and orbital diagram for an aluminum atom ($_{13}$Al)? What are the four quantum numbers for the last electron added?

Electron configuration ($_{13}$Al): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$, $[Ne]_{10}$ **3** s ^{2} 3p^1; Period :**3** , **III**$_A$, block **P**

Quantum numbers for last electron: n=3, l=1, ml=-1, ms=+1/2

Orbital diagram:



Exercise 4 :

Give electronic configuration and deduce the atomic number of the following elements :

- Oxygen: **2nd** period, group **VI**$_A$ (16th column), block **p** . $[He]_2$ $2s^2\ 2p^4$ **: Z= 8 (O)**
- Chlorine: **3rd** period, group **VII**$_A$ (17th column) , block **p**. $[Ne]_{10}$ **3** s ^{2} 3p ^{5} **: Z= 17 (Cl)**
- Manganese: **4th** period, group **VII**$_B$ (7th column), , block **d**. $[Ar]_{18}$ $4s^2\ 3d^5$ **: Z= 25 (Mn)**

Exercise 5 :

1- Classify the following elements: $_{55}$Cs, $_9$F, $_{19}$K, $_3$Li, $_7$N by decreasing atomic radius, justifying your answer.

$_{55}$Cs	$[Xe]_{54}$ **6s^1 , period : 6, I$_A$**
$_9$F	$[He]_2$ **2s^2** 2p^5 **, period : 2, VII$_A$**
$_{19}$K	$[Ar]_{18}$ **4s**^{1}, **period : 4, I$_A$**
$_3$Li	$[He]_2$ **2s^1 , , period : 2, I$_A$**
$_7$N	$[He]_2$ **2s^2** 2p^3 **, period : 2, V$_A$**

Ra ←

	I$_A$	**V$_A$**	**VII$_A$**
period : 1			
period : 2	$_3$Li	$_7$N	$_9$F
period : 4	$_{19}$K		
period : 6	$_{55}$Cs		

$_{55}$Cs > $_{19}$K > $_3$Li > $_7$N > $_9$F

2- Assign to each element among the following $_{37}$Rb; $_{11}$Na; $_9$F; $_8$O; $_7$N; $_3$Li; $_2$He the electronegativity that corresponds to it: χ=0.9; 4.0; 0; 0.8; 3.0; 1.0; 3.5

$_8O$	$[He]_2$ **$2s^2$** $2p^4$ **, period : 2, VI_A**
$_2He$	**$1s^2$, period : 1, $VIII_A$**
$_3Li$	$[He]_2$ **$2s^1$, , period : 2, I_A**
$_7N$	$[He]_2$ **$2s^2$** $2p^3$ **, period : 2, V_A**
$_{37}Rb$	$[Kr]_{36}$ **$5s^1$, period : 5, I_A**
$_{11}Na$	$[Ne]_{10}$ **$3s^1$, , period : 3, I_A**
$_9F$	$[He]_2$ **$2s^2$** $2p^5$ **, period : 2, VII_A**

	I_A	**V_A**	**VI_A**	**VII_A**	**$VIII_A$**
period : 1					$_2He$
period : 2	$_3Li$	$_7N$	$_8O$	$_9F$	
period : 3	$_{11}Na$				
period : 5	$_{37}Rb$				

$_{37}Rb < {}_{11}Na < {}_3Li < {}_7N < {}_8O < {}_9F$

Helium (He) is a noble gas and has no electronegativity: $\chi=0$

$_{37}Rb$	$_{11}Na$	$_3Li$	$_7N$	$_8O$	$_9F$	$_2He$
0.8	0.9	1.0	3	3.5	4	0 (noble gas)