



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

Mohamed Khider University- Biskra

Faculty of sciences and technology

1st year engineer

Elements of chemistry (Chemistry-1)

Course and Exercises

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2023-2024

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Chapter One

Fundamental concepts

Chapter One: fundamental concepts

Definition of matter : Matter makes up all substances that have a mass and occupy space.

States of matter: A substance can exist in different physical forms, depending on its properties, temperature and pressure. These forms are called states of matter. The three best known are solids, liquids and gases. However, there are others. We can think of plasma which is almost a gas, with the difference that it contains free electrons and ions (atoms having lost electrons).

A- Solid state: It is a substance whose shape and size are fixed, meaning that the molecules that make up this substance are semi-fixed, close together, and arranged according to a specific system (for example, table salt).



A solid has a definite shape and a definite volume

B- Liquid state: A liquid is a substance that takes the shape of the container in which it is placed while the volume remains constant. The molecules in a liquid body can move at a limited rate (for example, water).



A liquid has an indefinite shape, it takes the shape of its container and a definite volume

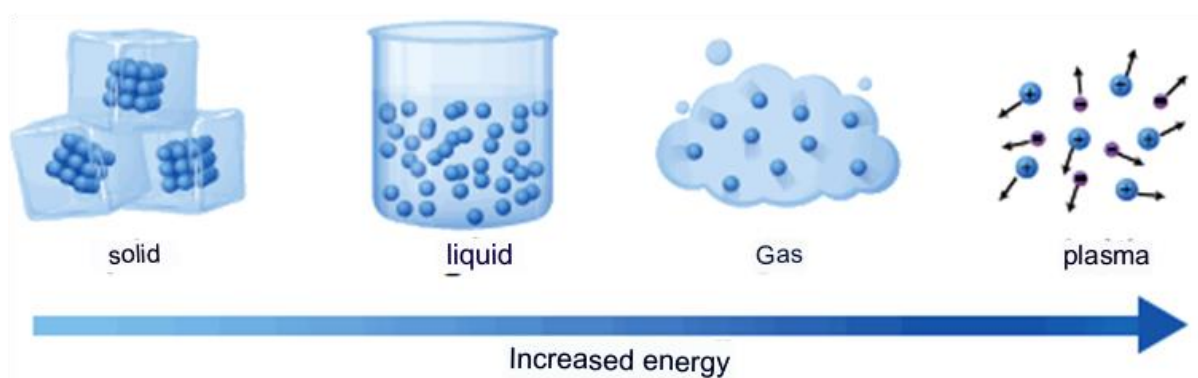
C- The gaseous state: The substance in this state is characterized by the property of expansion, as the gas takes the shape and size of the container in which it is placed, meaning that the molecules that make up this substance are completely free among themselves and are in chaotic movement (for example, oxygen gas



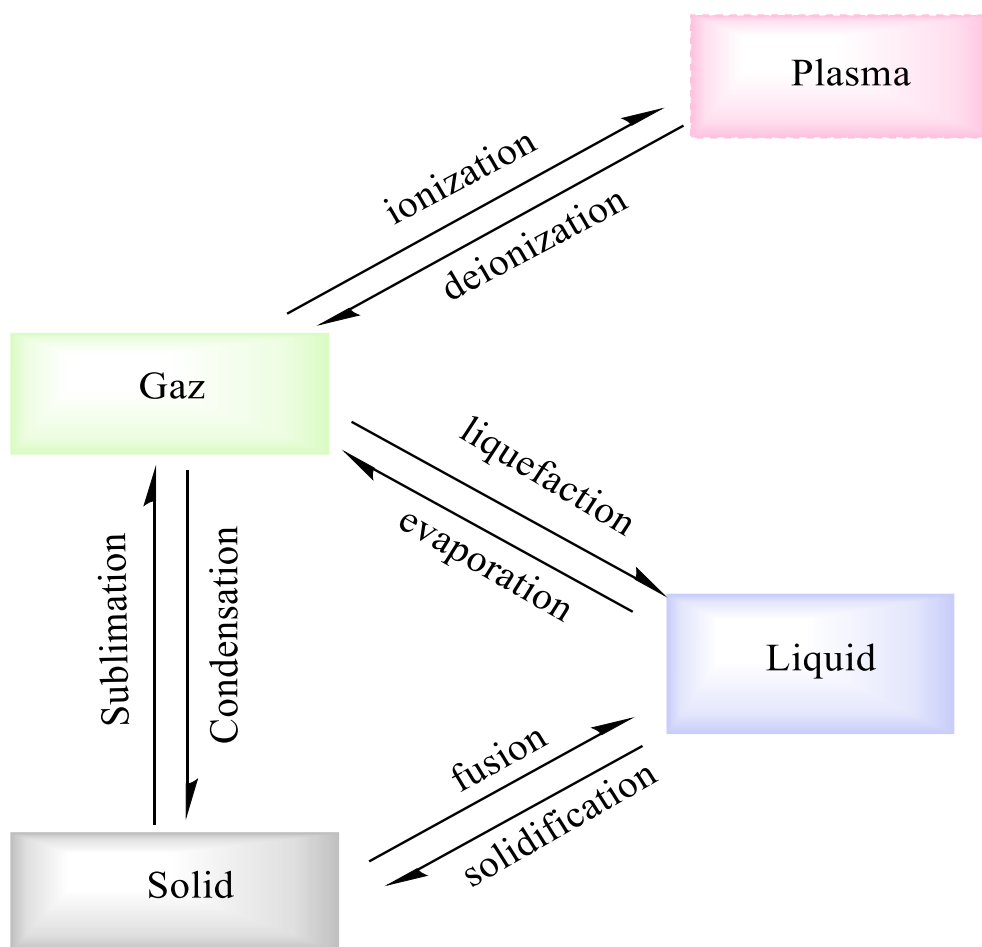
A gas has an indefinite shape and an indefinite volume, it assumes the shape and volume of its container

Changes in states of matter:

Depending on the external conditions (temperature and pressure), the same substance can be either a solid, a liquid or a gas. In fact, the passage of matter from the **solid** state to the **liquid** state occurs by **melting**, from the **liquid** to the **gaseous** state by **vaporization** and from the **solid** to the **gaseous** state by **sublimation**. When particles of matter gain energy, their vibrational movement increases, and the bonding forces between their particles weaken, so matter moves from one state to another.



- A diagram showing the changes between states of matter



III- Classification of matter

Matter exists in the form of mixtures (homogeneous or heterogeneous) and pure substances (simple or compound).

1- Pure substances: They are substances composed of identical molecules. Pure substances are distinguished by their physical properties (melting point, boiling point, density, refractive index, etc.) or chemical properties.

They are divided into two types:

- **Simple pure substances:** They are substances that consist of one type of atom (H_2 , O_2 , Fe, etc.).
- **Pure compound substances:** consisting of molecules with different atoms (H_2O , sodium chloride NaCl, carbon dioxide CO_2 , etc...)

2- Mixtures: They are compounds whose molecules are different, meaning they consist of two or more pure substances. They are divided into two categories:

- **Heterogeneous mixture:** through which it is possible to distinguish, by eye or with the help of a microscope, parts with different properties, meaning that it consists of two or more phases.

Example: A mixture of water and oil.

- **Homogeneous mixture:** It is a mixture of two substances in one phase such that its components cannot be distinguished.

Example: Salt water (a solution of water and salt) is a homogeneous mixture (the salt dissolves in water, so we obtain one phase). , air, steel (alloy), etc

A summary of how to distinguish the different main classifications of matter is presented in the following figure.

Matter

Does it have constant properties and compositions?

No

yes

Mixture

- Physical combination of two or more substances
- Composition can vary
- Properties can vary with composition

Pure substances

- Only one substance present
- Definite and constant composition
- Properties always the same under the same conditions

yes

Is it uniform throughout?

No

Homogeneous

- One visible phase
- Same properties throughout

Heterogeneous

- Two or more visible phases
- Different properties in different phases

Elements

- Cannot be broken down into simpler substances by chemical or physical means
- Building blocks for all other types of matter
- 117 elements known

Compound

- Can be broken down into constituent elements by chemical, but not physical, means
- Chemical combination of two or more elements
- Have definite, constant, elemental



Soft drink



Oil

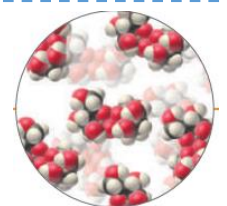
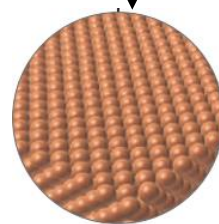
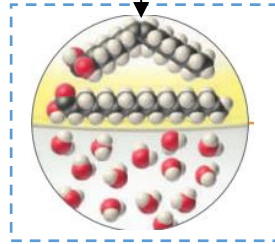
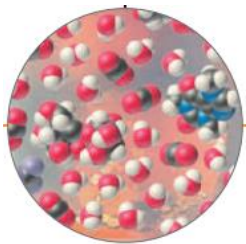
Water



Copper



Sugar



Molecular representation

IV. Concept of atom, molecules, mole and Avogadro's number

IV. 1. The atom: A simple body consists of identical atoms. We can consider that an atom is the smallest particle in a simple body that we can imagine without destroying this simple body. There are more than a hundred atoms. These are based on the same model and differ in their physical and chemical properties. Each element or atom is represented by a symbol. The symbol consists of an uppercase letter or two letters, the first uppercase followed by a lowercase letter. This formalism must be strictly adhered to otherwise the elements are mixed together (e.g. Co = cobalt, while CO = carbon + oxygen).

Example: carbon (C), oxygen (O), copper (Cu), aluminum (Al). The size of an atom is very small in angstroms ($1\text{A}^\circ = 10^{-10}\text{m}$) and its mass is about 10^{-26} kg

IV. 2. The molecule: Molecules are groups or clusters of atoms held together by means of chemical bonding. There are two types of molecule; molecules of an element and molecules of a compound

A molecule is a set of linked atoms. The molecule constitutes the smallest part of a pure body, Such as NaCl, O₃, H₂SO₄, etc..

IV. 3. The mole and Avogadro's number

3.1. Definition and unit: The mole designates the quantity of matter of a substance contained in a system which has as many elementary entities as there are atoms in 12g of carbon 12(¹²C).

The mole is a unit of account used by chemists to express quantities of matter. Its unit is the mole (mol).

3. 2. Avogadro's number: It corresponds to the number of particles found in exactly 12 g of carbon 12. Also, the mole corresponds to the quantity of matter found in 12 g of carbon 12. Avogadro's number, symbolized N_A, corresponds to the number of particles found in a mole, i.e. 6.022×10^{23} particles.

To determine the number of moles from Avogadro's number and the number of particles, we can use the following formula:

$$n = \frac{N}{N_A} \quad \text{ou} \quad N = n \times N_A$$

Where

n : represent the number of moles (mol)

N : represent the number of particles (atoms, molecules, ions, etc.)

N_A : represent Avogadro's number

3. 3. Molar mass: The atomic molar mass of a chemical element is the mass of one mole of atom of this element taken in its natural state. It is expressed in g/mol or g.mol⁻¹ and is denoted by the letter M . The different values of M are read in the periodic table. Example: $M_H = 1\text{g/mol}$, $M_O = 16\text{g/mol}$, $M_{Ag} = 108\text{g/mol}$.

3. 4. The molecular molar mass: or mass of a chemical substance is the mass of one mole of a molecule of that substance. It is determined by taking the sum of all the atomic molar masses of all the chemical elements present in the molecule. **Example:** Determine the molecular molar masses of: H_2O , $\text{C}_{22}\text{H}_{22}\text{O}_{11}$, $\text{C}_6\text{H}_5\text{NO}_2$, $\text{CO}(\text{NH}_2)_2$.

We give in g/mol: C = 12, H = 1, O = 16, N = 14.

3. 5. Atomic mass units (amu): An atomic mass unit is defined as a mass equal to one twelfth the mass of an atom of carbon-12.

$$1\text{uma} = \frac{1}{12} m_{C^{12}}$$

Where $m_{C^{12}}$: represent the mass of one atom of carbon

$$12\text{g} \longrightarrow N_A \text{ atom of carbon}$$

$$m_{C^{12}} \text{ (mass of one atom of carbon)} \longrightarrow 1\text{atom of carbon}$$

$$m_{C^{12}} = \frac{M_{C^{12}}}{N_A}$$

So
$$1\text{mau} = \frac{1}{N_A} = \frac{1}{6,023 \cdot 10^{23}} = 1,666 \cdot 10^{-24}\text{g} = 1,666 \cdot 10^{-27}\text{kg}$$

Note: The key difference between atomic mass unit and atomic mass is that the atomic mass unit is the unit that we use to measure the mass of an atom whereas the atomic mass is the mass of a particular single atom.

3. 6. Molar volume (V_m):

It is the volume occupied by one mole of a substance (chemical element or chemical compound) in the gaseous state and under a certain temperature and pressure under regular conditions ($P = 1\text{atm}$). It is found that the volume equals 22.4 litres.

3. 7. Concept of quantity of matter n : The quantity of material n of a sample is obtained by dividing the mass m of this sample by the molar mass M of the chemical species considered.

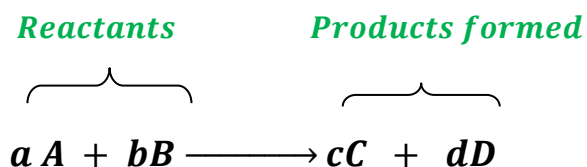
$$n = \frac{m}{M}$$

n (mol), m (g), M (g/mol).

Chemical reaction concept:

Definition: A chemical reaction is a chemical transformation during which pure substances disappear and simultaneously new pure substances are formed. The pure substances which disappear during a chemical reaction are called the reactants and the pure substances which are formed are called the products.

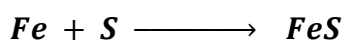
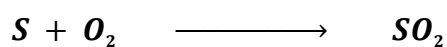
We can write an equation that shows the balance of a chemical reaction:



This balanced chemical equation obeys two laws:

- **Conservation of elements:** During a chemical reaction, there is conservation of atoms (the elements are conserved)

Example



- **Mass conservation:** In a chemical reaction, the mass of the reactants disappeared is equal to the mass of the products trained (Lavoisier's law)

Example

2 H₂ react with 32.00 g of O₂ to give 36.04 g of H₂O



+



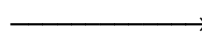
Two hydrogen molecules

One oxygen molecule

two water molecules

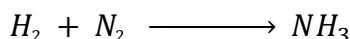
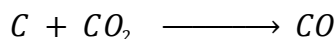
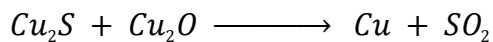
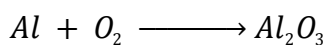


+



Exercise :

Balance the following reactions:

**VI. Qualitative and quantitative aspect of the material**

The main difference between qualitative and quantitative analysis chemistry is that qualitative analysis determines whether or not different chemical components are present in a sample, whereas quantitative analysis determines the amount of different chemical components present in a sample.

VI.1 Solutions

A solution is a homogeneous mixture of two or more components, where the all the components appear as a single phase.

Solvent:

The component that dissolves the other component is called the solvent.

Solute:

The component(s) that is/are dissolved in the solvent is/are called solute(s).

Generally solvent is present in major proportion compared to the solute. The amount of solute is lesser than the solvent.

Example of solution

- Sugar syrup is a solution where sugar is dissolved in water using heat. Here, water is the solvent and sugar is the solute.

VI. 2. Solution concentration units

Because solutions are mixtures, they have a variable composition. Specifying what the composition of a solution is involves specifying solute concentrations.

A concentration is the amount of solute present in a specified amount of solution. Many methods of expressing concentration exist, and certain methods are better suited for some purposes than others.

Molarity: The molarity of a solution is the number of moles of the solute in 1 litre of the solution.

$$\text{molarity } (C) = \frac{\text{moles of solute}}{\text{litres of solution}} = \frac{n \text{ (mol)}}{V(L)}$$

$$n = \frac{\text{mass } (g)}{M}$$

Where: **MM** = molar mass of solute

The equation can also be used to find n (the number of moles of solute) if M (the molarity) and V (the volume in litres) are known, and V if n and M are known

$$n = MV \quad V = \frac{n}{M}$$

Example 1:

A sample of NaNO_3 weighing 0.38 g is placed in a 50.0 mL volumetric flask. The flask is then filled with water to the mark on the neck, dissolving the solid. What is the molarity of the resulting solution?

Solution: To calculate the molarity, you need the moles of solute. Therefore, you first convert grams NaNO_3 to moles.

we have : $n = \frac{\text{mass } (g)}{MM}$

Where: $M = M_{Na} + M_N + 3 M_O = 23 + 14 + 3 \times 16 = 85 \text{ g/mol}$

$$n = \frac{0.38}{85} = 4.47 \times 10^{-3} \text{ mol}$$

So the molarity is: $C = \frac{n \text{ (mol)}}{V(L)} = \frac{4.47 \times 10^{-3}}{50 \times 10^{-3}} = 0.089 \text{ mol/l}$

Example 2:

An experiment calls for the addition to a reaction vessel of 0.184 g of sodium hydroxide, NaOH, in aqueous solution. How many milliliters of 0.150 M NaOH should be added?

Solution:

Because molarity relates moles of solute to volume of solution, you first need to calculate the number of moles of NaOH. Then, you can calculate the volume in liter of solution, using the molarity equation.

$$n = \frac{\text{mass (g)}}{MM}$$

Where: $M_{NaOH} = M_{Na} + M_H + M_O = 23 + 1 + 16 = 40 \text{ g/mol}$

$$n = \frac{0.184}{40} = 4.6 \times 10^{-3} \text{ mol}$$

So the volume is: $V = \frac{n}{M} \implies V = \frac{4.6 \times 10^{-3}}{0.150} = 3.07 \times 10^{-2} \text{ L} = 30.7 \text{ ml}$

exercice:

Exercise:

- Calculate the molarity of a solution of 0.25 mole of NaOH in 5.0 L of solution.
- Calculate the number of moles of citric acid in 250 mL of a 0.400 M solution of citric acid.
- Calculate the volume (in mL) of a 0.355 M NaOH solution which would contain 0.200 mole of NaOH. Calculate the mass of Na₂CO₃ that must be used to make 700 mL of a 0.136 M Na₂CO₃ solution.
- How many moles of sodium chloride should be put in a 50.0-mL volumetric flask to give a 0.15 M NaCl solution when the flask is filled to the mark with water? How many grams of NaCl is this?

Mass Percent Concentration: Percent by mass is the mass of solute in a solution divided by the total mass of solution, multiplied by 100.

$$m \% = \frac{m_{\text{solute}}}{m_{\text{solvent}}} \times 100$$

Example 1: A saline solution with a mass of 355 g has 36.5 g of NaCl dissolved in it. What is the mass/mass percent concentration of the solution?

Solution

We can substitute the quantities given in the equation for mass/mass percent:

$$m\% = \frac{m_{\text{NaCl}}(g)}{m_{\text{solution}}(g)} = \frac{36.5}{355} \times 100\% = 10.3\%$$

Example 2: A dextrose (also called D-glucose, C₆H₁₂O₆) solution with a mass of 2.00 × 10² g has 15.8 g of dextrose dissolved in it. What is the mass/mass percent concentration of the solution?

Normality (N): Represents the number of equivalents contained in one liter solution

$$\text{Normality}(N) = \frac{\text{Number of equivalents of solute}}{\text{volume of solution(liter)}}$$

$$N = \frac{\text{number of equivalent}}{\text{volume}} = \frac{n \text{ eq g}}{V}$$

$$\text{Number of gram equivalent} = \frac{\text{Mass}}{\text{Equivalent weight}}$$

$$\text{Eq. mass} = \frac{\text{Molar. mass}}{\text{Valency}}$$

$$\text{Equivalent weight} = \frac{M}{Z}$$

Z: is the valence factor, where valence factor for acids and bases is the number of H⁺ and OH⁻ ions they release in the solution, respectively

Gram equivalent (Eq.mass or eq. weight): It is the mass of matter corresponding to a mole of particles, whether they are H^+ protons or e^- electrons. There are three families of gram equivalents classified according to fixation or loss of particles.

$$Eq. mass = \frac{Molar. mass}{Valency}$$

$$Equivalent weight = \frac{M}{X}$$

A) - The gram equivalent of the acid:

$$g. eq_{acide} = \frac{M_{acide}}{Z_{H^+}}$$

Example: Gram equivalent of hydrochloric acid HCl,

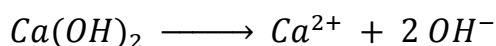


$$Eq. g HCl = \frac{M_{HCl}}{n_{H^+}} = \frac{36,5}{1} = 36,5 g$$

B) Gram equivalent of base

$$g. E_{base} = \frac{M_{base}}{Z_{OH^-}}$$

Example: Gram equivalent of calcium hydroxide $Ca(OH)_2$

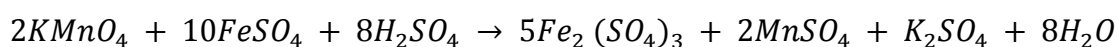


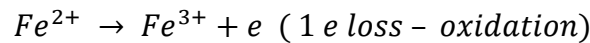
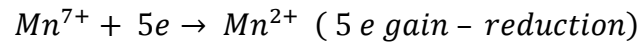
$$Eq. g Ca(OH)_2 = \frac{M_{Ca(OH)_2}}{n_{OH^-}} = \frac{74,1}{2} = 37,05 g$$

c- Gram Equivalent in (oxidation – reduction) reaction (Redox)

$$Eq. g_{ox/red} = \frac{M_{ox/red}}{Z_{e^-}}$$

Example:





$$\text{Eq. g} (\text{KMnO}_4) = \frac{MM}{5} = \frac{157.9}{5} = 31.6$$

The relationship between the various concentration units.

The relationship between molar concentration and mass concentration:

$$C_M(\text{ molar concentration}) = \frac{n}{V} \quad \text{and} \quad C_m(\text{mass concentration}) = \frac{m(\text{ solute})}{V(\text{solvent})}$$

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

$$\Rightarrow C_M = \frac{1}{M} \frac{m}{V} \Rightarrow$$

$$C_M = \frac{C_m}{M}$$

The relationship between molar concentration and regularity:

$$N = \frac{\text{number of g. eq.}}{V}$$

$$\text{number of g. eq.} = \frac{m(\text{mass of solute})}{\text{g. eq}}$$

$$\Rightarrow N = \frac{m}{\text{g. eq.} \cdot V}$$

$$N = \frac{m}{\text{g. eq.} \cdot V} \quad \text{et} \quad \text{g. eq} = \frac{M(\text{molar mass})}{z} \quad \Rightarrow$$

$$N = \frac{m \cdot z}{M \cdot V}$$

$$N = \frac{m \cdot z}{M \cdot V} = n \cdot \frac{z}{V} = C_M \cdot z \quad \Rightarrow \quad N = C_M \cdot z$$

Exercise: We dissolve a mass $m = 20$ g of phosphoric acid H_3PO_4 in a volume of 250 cm^3 , and solution S_1 is formed.

1. Calculate the number of moles of acid dissolved
2. Calculate the molar concentration of solution S_1
3. Calculate the mass concentration of solution S_1

4. Calculate the gram equivalent of phosphoric acid and then deduce the regularity of S1
5. What is the volume of distilled water that must be added to solution S1 to obtain a normality of $N_2 = 0.5$

Molality : is calculated by the following mathematical relationship:

$$\text{Molality (m)} = \frac{\text{number of moles of solute (mol)}}{\text{mass of solvent (Kg)}}$$

Series No. 1**First exercise:**

Among the following compounds, identify which ones are: a homogeneous mixture, a heterogeneous mixture, a compound pure substance, and a simple pure substance.

Methane gas, air saturated with water vapor (smog-filled air), NaCl, distilled water, ozone O₃, sea water, iron Fe, air,

Second exercise:

1) How many moles are there in:

(a) 0.5g of copper (b) 10g of carbon C (c) 2mg of iron. (Fe)

2) What is the number of moles and the number of molecules present in:

(a) 20g of H₂O (c) 15g of CH₃COOH

(b) 45g of NaOH (d) 15g of CO(NH₂)₂.

3) Give the number of moles and the mass in g for the following:

(a) 3,620 10²⁴ atoms of Zn (b) 12,04 10²³ molecules of N₂ (c) 36,12 10²¹ molecules of NaOH

4) What is the mass of one atom of Si in g and in **a m u**.

5) Considering the atom of silver (Ag) is spherical in shape and has a radius $r = 1.44 \text{ \AA}$.

What is the length obtained when we add the atoms in 10 grams of silver side by side in a linear manner?

Third exercise : In 0.6mol of carbon dioxide CO₂, How much do we find:

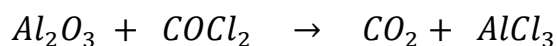
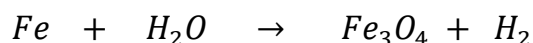
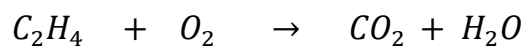
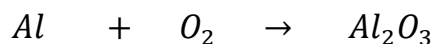
- Gram of carbon dioxide.
- An atom of carbon C and oxygen O.
- A molecule of carbon dioxideCO₂

Fourth exercise.

1) Find the molecular formula of a compound containing 68.4% Cr and 31.6% Oxygen (weight percentages).

2) What is the molecular formula of a compound that contains 67.5% carbon, 12.8% hydrogen and 19.7% nitrogen, noting that its molar mass is 142 g/mol.

Fifth exercise : Balance the following chemical equations



Sixth exercise: $KClO_3$ decomposes to give both KCl and O_2 .

(a) Write the reaction equation

(b) What is the number of moles of $KClO_3$ that produce 1.5 liters of O_2 (under regular conditions of pressure and temperature)

(c) Find the volume of O_2 released as a result of the dissociation of 12.25 g of $KClO_3$ under the same conditions

Seventh exercise: We treat 10g of iron (Fe) with HCl.

(a) Write the reaction equation.

(b) Calculate the mass of HCl reacted.

(c) What is the volume and number of hydrogen (H_2) molecules released (under regular conditions of pressure and temperature).

Eighth exercise

- What is the molarity of the following solutions:
- 3mol of KOH in 4 liters of distilled water.
- 0.5 mol of NaCl in 250ml.
- 70 g of HCl in 3 liters of distilled water.

Ninth exercise:

- What is the mass of sulfuric acid present in 1 liter of a 0.2N solution of this acid?

- Find the molality of a solution of sulfuric acid containing 27% (by mass percentage) of the acid, knowing that its density is 1.198.

H₂SO₄: 98 g/mol

Correction

Ex 01 :

| mixture | | pure substance | |
|----------------------------|--------------------------------|--------------------------------|----------------------------|
| homogeneous mixture | heterogeneous mixture | compound pure substance | simple pure substance |
| air | air saturated with water vapor | NaCl | ozone O₃ |
| sea water | | Methane gas | iron Fe |
| | | distilled water | |

Exo2

MM_{Cu} = 63.5 g/mol , MM_{Fe}=56 , MM_C=12 ; M_{Na}=23 , M_N=14

① The number of moles of :

(a) 0.5 g of copper (Cu) :

$$n = \frac{\text{mass (m)}}{\text{molar mass (MM)}} = \frac{0.5}{63.5} = 7.8789 \times 10^{-3} \text{ mole}$$

(b) 10g of carbone

$$n = \frac{m}{MM} = \frac{10}{12} = 8,333 \times 10^{-1} \text{ mole}$$

(c) 2mg of iron Fe

$$n = \frac{10^{-3} \cdot 2}{56} = 3,5714 \times 10^{-5} \text{ mole}$$

② the number of moles and the number of molecules present in:

(a) 20g of water H_2O

The mole number: $n = \frac{m}{MM} = \frac{20}{18} = 1.111 \text{ mole}$

The number of molecules

1mole \longrightarrow Avogadro's number of molecules ($N_A = 6.023 \times 10^{23}$)

1.111 mole \longrightarrow N molecules

$$N = n \times N_A = 1,111 \times 6,023 \times 10^{23} = 6,69 \times 10^{23} \text{ molecules of water .}$$

b) 45g of sodium hydroxide (NaOH):

The number of moles: $n = \frac{m}{MM} = \frac{45}{(23+16+1)} = \frac{45}{(40)} = 1.125 \text{ mole}$

The number of molecules

$$N = n \times N_A = 1,125 \times 6,023 \times 10^{23} = 6,776 \times 10^{23} \text{ molecules of NaOH}$$

c) 15g of acetic acid (CH_3COOH)

$$n = \frac{m_{CH_3COOH}}{MM_{CH_3COOH}} = \frac{15}{(2 \times 12 + 4 + 16 \times 2)} = \frac{15}{(60)} = 0.25 \text{ mole}$$

$$N = 1,5057 \times 10^{23} \text{ molecules of } CH_3COOH$$

d) 15g of urea $CO(NH_2)_2$

$$n = \frac{15}{2 \times 14 + 16 + 4 + 12} = \frac{15}{60} = 0.25 \text{ mole}$$

$$N = 1,5057 \times 10^{23} \text{ molecules of } CO(NH_2)_2$$

3) the number of moles and the mass in g for the following:

a) 3.62×10^{24} atoms of Zn

1 mole of zn \longrightarrow Avogadro's number of Zn atoms ($N_A = 6.023 \times 10^{23}$)

n mole \longrightarrow $N = 3.62 \times 10^{24}$ atoms of Zn)

$$n = \frac{N}{N_A} = \frac{3.62 \times 10^{24}}{6.023 \times 10^{23}} = 6,0102 \text{ mole}$$

- Mass of Zn

$$n = \frac{m}{MM} \Rightarrow m_{Zn} = MM_{Zn} \times n_{Zn} = 65,37 \times 6,0102 = 392,886g$$

b) 12.04×10^{23} molecules of N_2

$$n_{N_2} = \frac{N_{N_2}}{N_A} = \frac{12.04 \times 10^{23}}{6,023 \times 10^{23}} = 1.999 \text{ mole} \approx 2 \text{ mole}$$

$$m_{N_2} = MM_{N_2} \times n_{N_2} = 28 \times 2 = 56g \text{ of } N_2$$

c) $36,12 \cdot 10^{21}$ Molecules of NaOH

$$n_{NaOH} = \frac{36,12 \cdot 10^{21}}{6,023 \cdot 10^{23}} = 5,997 \times 10^{-2} \text{ mol}$$

$$m_{NaOH} = 40 \times 5,997 = 2,3988g$$

4) Mass of one atom of Si in gram

$$M_{Si} = 28g$$

1 mol \longrightarrow N_A atom \longrightarrow 28g

1 atom \longrightarrow m_{Si}

$$m_{si} = \frac{28}{6,023 \cdot 10^{23}} = 4,65 \cdot 10^{-23}$$

Mass of one atom of Si in amu

$$1 \text{ a m u} = 1,67 \cdot 10^{-27} \text{ Kg}$$

$$1 \text{ a m u} \longrightarrow 1,67 \cdot 10^{-24} \text{ g}$$

$$X \text{ amu} \longrightarrow 4,65 \cdot 10^{-23} \text{ g}$$



$$\Rightarrow X = 27.84 \text{ m a u}$$

5) Calculating the amount of length : $M_{Ag} = 107,9\text{g}$

The number of atom

$$N = n_{Ag} \times \mathcal{N}$$

$$= \frac{m_{Ag}}{M_{Ag}} \cdot \mathcal{N} = \frac{10}{107,9} \times 6,023 \cdot 10^{23}$$

$$N = 5,58 \times 10^{22} \Rightarrow \text{atomes}$$



1 atom corresponds to **2r**

2 atoms corresponds to **4r**

3 atoms corresponds to **6r**

4 atoms corresponds to **8r**

·
·
·
·
·

N atoms corresponds to **2N r**

and therefore

$$\mathbf{L = N \times 2r}$$

$$= 5,58 \cdot 10^{22} \cdot 2 \cdot 1,44$$

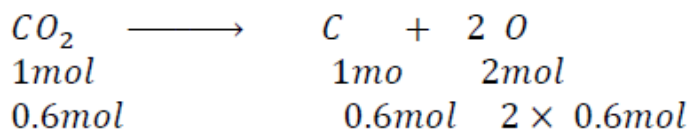
Exercice 3 :

Mass of 0.6 moles of carbone dioxide CO_2

$$M_{\text{CO}_2} = 44 \text{ g/mol}$$

$$m = n \times M = 0.6 \times 44 = 26.4 \text{ g}$$

Number of C and O atoms



Number of carbon atoms:

$$\begin{aligned}
 N_{\text{C}} &= n_{\text{C}} \times N_{\text{A}} = 0.6 \times 6.023 \times 10^{23} \\
 &= 3.6138 \times 10^{23} \text{ atomes}
 \end{aligned}$$

Number of oxygen atoms:

$$\begin{aligned}
 N_{\text{O}} &= n_{\text{O}} \times N_{\text{A}} = 2 \times 0.6 \times 6.023 \times 10^{23} \\
 &= 7.2276 \times 10^{23} \text{ atomes}
 \end{aligned}$$

• Number of CO_2 molecules

$$\begin{aligned}
 N_{\text{CO}_2} &= n_{\text{CO}_2} \times N_{\text{A}} = 0.6 \times 6.023 \times 10^{23} \\
 &= 3.6138 \times 10^{23} \text{ molecules}
 \end{aligned}$$

Ex04 :

1 We put the general formula: $\text{Cr}_x \text{O}_y$

$$\frac{52x}{\text{Cr}\%} = \frac{16y}{\text{O}\%}$$

$$ax = b \cdot y$$

$$\Rightarrow \frac{y}{x} = \frac{a}{b}$$

then

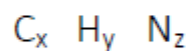
$$\frac{y}{x} = \frac{\frac{52}{\%Cr}}{\frac{16}{O\%}} = 1,5 = \frac{3}{2}$$

If we take $x = 2n$

$$y = 3n$$

The formula $(Cr_2 O_3)_n$

2Let's assume the formula of the compound is :



$$\frac{M}{100\%} = \frac{12x}{C\%} = \frac{Y}{H\%} = \frac{14Z}{Z\%}$$

$$\frac{142}{100\%} = \frac{12X}{67.5} = \frac{Y}{12.8} = \frac{14Z}{19.7}$$

$$X \approx 8, Y = 18, Z \approx 2$$

The formula of the compound is $C_8 H_{18} N_2$

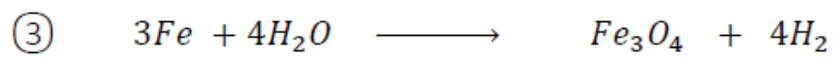
Exo5

$$\text{Al:} \quad x = 2Z$$

$$2Y = 3Z$$

$$2Y = \frac{3}{2}x$$

$$x = 2 \quad \Rightarrow Y = \frac{3}{2}, \quad Z = 1$$



Chapter Two

Components of the Atom

Chapter Two: Components of the Atom

VII. Introduction: Dalton's theory, which was published in 1808, focused mainly on the indivisibility of the atom, but at the end of the 19th century and the beginning of the 20th century, and after these results, research was revealed to reach an explanation of the structure of the atom.

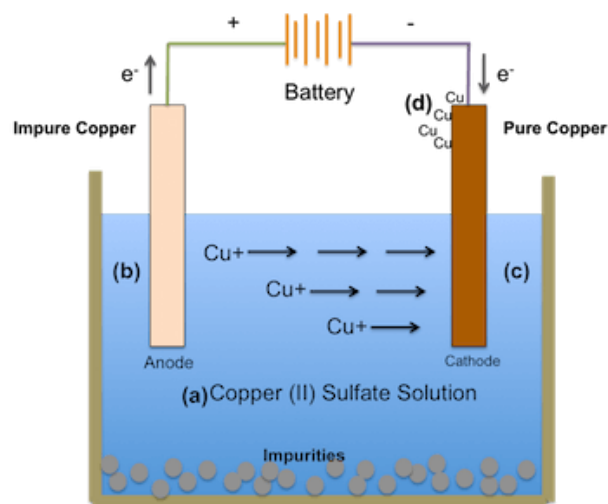
There are different models proposed to explain the structure of the atom, and we will mention the most important of them.

VIII. Discover the components of the atom

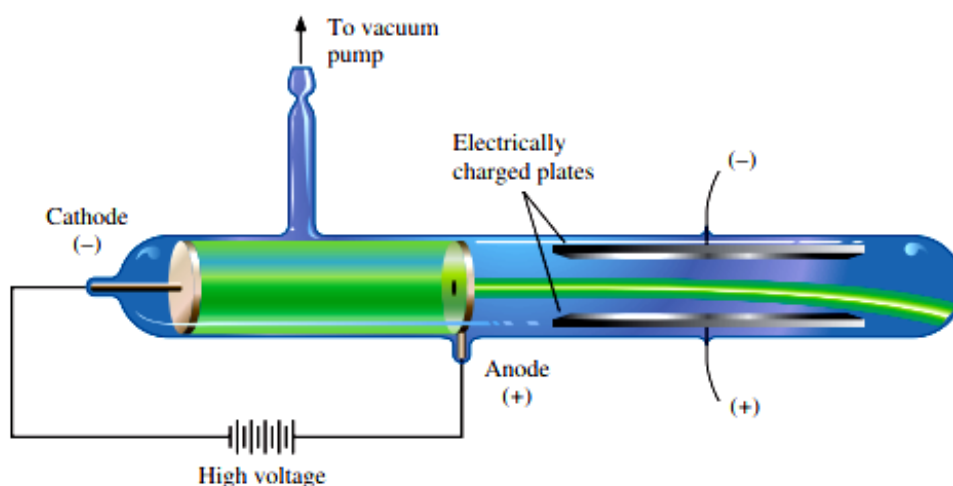
2. Electron discovery

Faraday's experiment: The scientist Faraday carried out the electrolysis of copper sulphate (Cu_2SO_4). He placed this solution in a glass container and then dipped two electrodes in it - an anode (+) and a cathode (-). He observed the following:

- Deposition of copper atoms at the cathode (-)
- Sulphate deposition at the anode (+) Faraday concluded that the atom contains negatively charged particles that were later known as electrons



Crooks experiment: Crooks passed an electric current through rare gases in vacuum tubes, which are closed glass tubes at both ends of which two metal plates (electodes) contain a low-pressure gas. The electrodes are connected to an electric source with a high potential difference of about 15,000 volts.



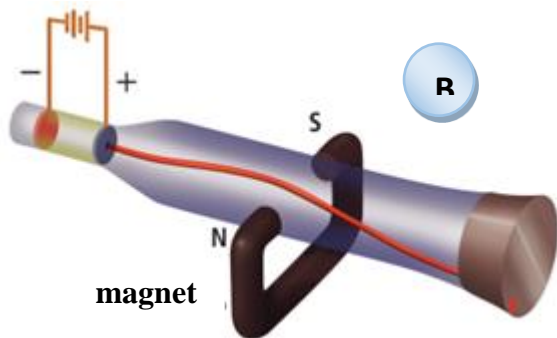
Notes

- When a normal potential difference and normal pressure inside the tube, nothing happened.
- At a large potential difference (V 15000) and emptying the tube of gas to a pressure of 10 mmHg, a glow was observed inside the tube around the cathode.
- When continuing to empty the tube, we notice that the glow has moved away from the circumference of the cathode and spread throughout the entire tube until it reaches the walls of the tube at a pressure of 0.1 mmHg.

Conclusions:

- ❖ At a low potential difference, the gas does not conduct electricity.
- ❖ When there is a high potential difference and a low pressure inside the tube, one of the electrodes sends out small particles, the collision of which with gas molecules causes the rays to shine. Its spread throughout the rest of the tube, increases when the number of gas particles decreases until these particles begin to collide with the walls of the tube.
- ❖ These rays travel in straight lines and form a shadow for objects in their path, which indicates that they have the properties of rays (Figure A)
- ❖ These rays are deflected when exposed to a magnetic field, which confirms that they carry a charge (Figure B)
- ❖ They are deflected if they pass through an electric field, and the deflection toward the positive pole indicates that they are composed of negatively charged particles (Figure C).

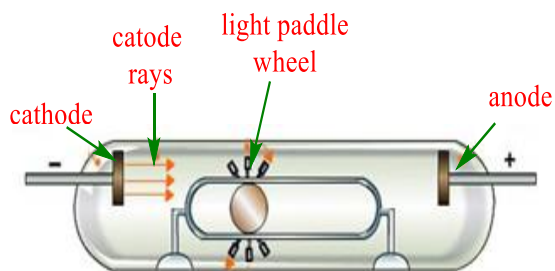
- ❖ Since changing the metal composing the electrodes or changing the gas in the tube does not affect the resulting cathode rays, scientists have concluded that the negatively charged particles of cathode rays exist in all forms of matter, and are called electrons.



When the cathode rays are exposed to a magnetic field, they deviate, which indicate that these particles have a charge.

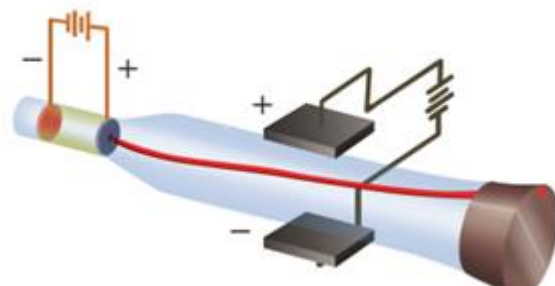


Shadow of the object is formed due to cathode rays travelling in straight lines



cathode rays can rotate a high paddle wheel placed in their path

The light paddle wheel rotates due to the collision of cathode rays on them



In the presence of an electric field, the cathode rays are deflected towards the positive core, which indicates that these particles have a negative charge.

Characteristics of these particles:

- 1) When placing a barrier with a specific geometric shape between the cathode and the anode, it was observed that the shadow of the barrier fell from the side of the anode, so it was concluded that it is particles emanating from the cathode and traveling in straight lines called the cathode rays
- 2) When a fan is placed between the two poles in the path of cathodic radiation, the latter begins to rotate, so it is concluded that this cathodic radiation consists of particles with kinetic energy, and therefore it has a mass
- 3) It was noticed that these cathode rays are attracted to the positive plate inside a condenser, so it turns out that these particles carry negative charges. These negative mass minutes were called electrons by the scientist SToney.

$$F_e = q E$$

where:

F_e represents electrical force,

q: electric charge

E: electric field

4. **Action of the magnetic field:** The electron beam is deflected downward due to the magnetic force $F_m = q v B$

Where:

F_m: is the magnetic force,

q: the electric charge,

V : represent the beam speed,

and **B** is the magnetic field.

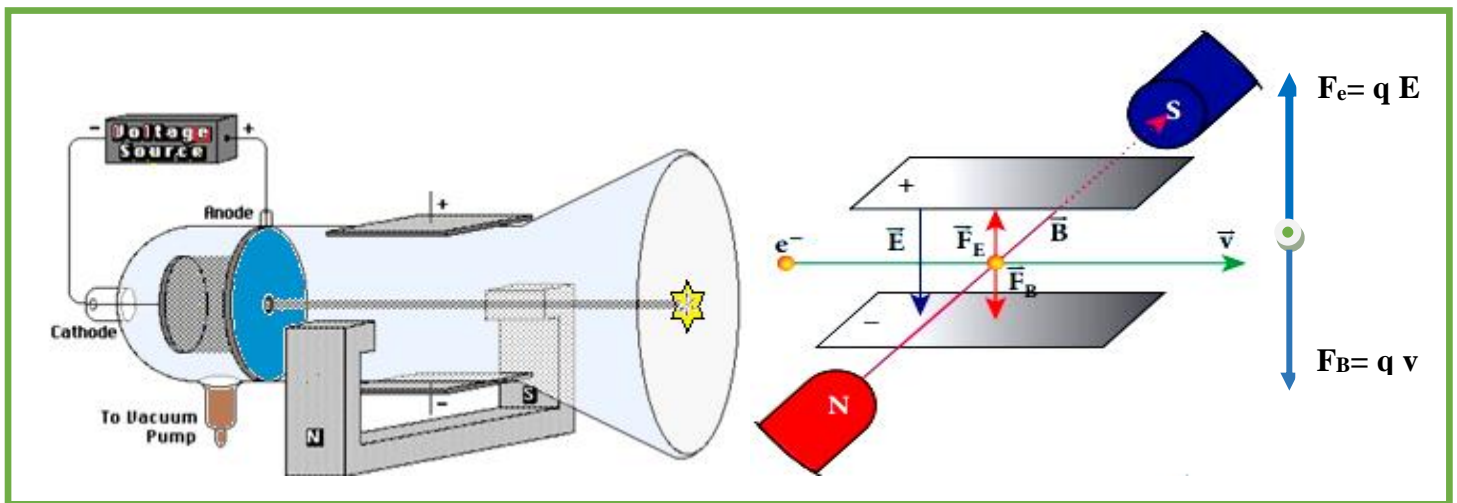
5. Balanced fields

6. Determination of velocity of cathode rays

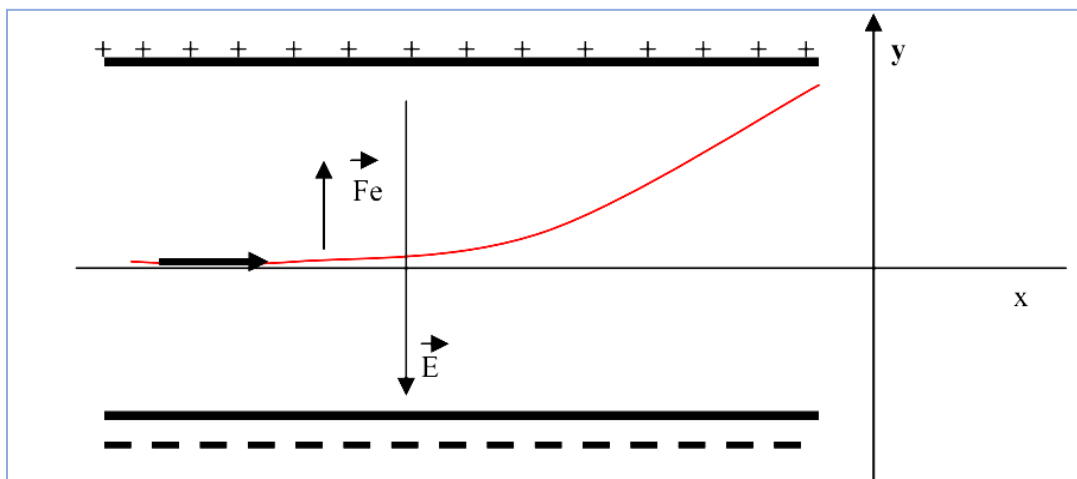
The beam passes through the fields without being deflected when the electric force F_e is equal and opposite to the magnetic force F_m on each electron (as shown in Figure 2). The two fields may be said to be ‘balanced’ because they produce zero resultant force on each electron. In this situation, the speed of each electron is given by:

$$\|\vec{F}_e\| = \|\vec{F}_m\|$$

$$qE = qvB \implies v = \frac{E}{B}$$



When the magnetic field is removed, the beam is deflected by the electric field E towards the positive plate. When it leaves the electric field, it collides with the screen at point M



The total forces acting on the beam are given as follows:

$$\sum F_{ext} = m_e \gamma \implies F_e = qE = m_e \gamma \implies \gamma = \frac{qE}{m_e}$$

Where m_e is the mass of the electron and γ is the acceleration of its motion

If we divide the motion at the entrance to the condenser, we get uniformly varying motion along the (oy) axis and uniform along the (ox) axis

$$\gamma = \gamma_y + \gamma_x$$

$$\gamma_x = 0 \text{ (uniform)} ; \gamma_y = qE/m \text{ (uniformly accelerated)}$$

So the motion along the x axis is a uniform straight line:

$$X = v t$$

So the motion along the x axis is a uniform straight line:

$$t = \frac{X}{v}$$

Motion along the y axis is uniformly accelerated motion:

$$Y = \frac{1}{2} \gamma t^2 = \frac{1}{2} (qE / m) t^2$$

We substitute the time expression into y

$$Y = \frac{1}{2} \frac{qE}{m_e} \left(\frac{X}{v} \right)^2$$

Since the charge q is the charge of the electron, we replace it in the equation with e , and replace x with d (the length of the capacitor) and y with OM , so the expression for y becomes as follows:

$$OM = Y = \frac{1}{2} \frac{eE}{m_e} \frac{d^2}{v^2}$$

Thus, the ratio of the electron's charge to its mass can be deduced

$$\frac{e}{m_e} = 2 \cdot (OM) \cdot \frac{v^2}{Ed^2}$$

By substituting the experimental values d , B , E , y , the ratio between the charge of the electron to its mass e/m can be calculated, which represents a constant quantity, where:

$$\frac{e}{m_e} = 1.6589 \times 10^{11} \text{ coulomb/Kg}$$

Millikan's experiment (1911):

Millikan's experiment aims To determine the elementary electron charge by observing the motion of charged oil drops in an electric field.

The apparatus he used is shown schematically in the following Figure.

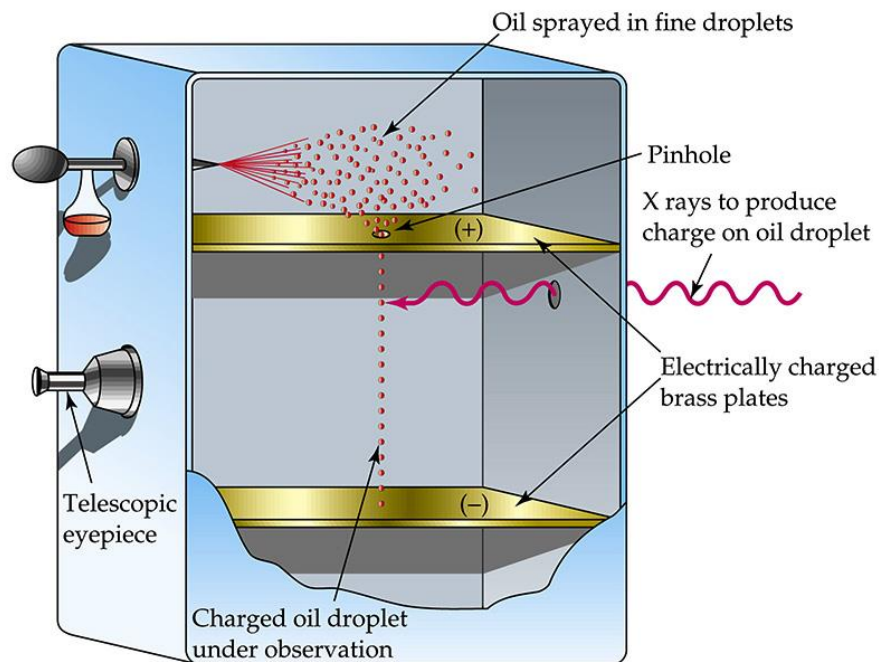


Figure 4: Illustration of Millikan's oil drop experiment

The Millikan device consists of a constant-temperature chamber, in which the air pressure can be changed using a pump connected to the chamber. The spray device emits oil droplets (small in size). At the bottom of the chamber there are two condenser plates, and the upper plate is perforated to allow the oil droplets to fall into the condenser. The air inside the device is ionized by rays. X. After some air molecules are ionized into positive and negative ions, these ions stick to the droplets and become positively or negatively charged. Kinetic study of droplets in an electric field allows the charge carried by each droplet to be determined. We perform the experiment in the following two cases:

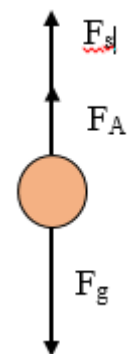
1- In the absence of an electric field (Determination of radius of the droplet) :

First, with zero applied electric field, the droplet falls, with a limit speed v_0 , and is under the influence of the following forces:

- **The force of gravity F_g is directed downwards where:**

$$F_g = m.g$$

m: is the mass of the droplet



g: the gravity of the earth

$$\rho_{\text{oil}} = \frac{m}{V} \Rightarrow m = \rho_{\text{oil}} V$$

ρ_{oil} : represent the volumetric mass of the oil

V: represents the volume of the droplet

Considering that the droplet is spherical in shape

$$V = \frac{4}{3}\pi r^3$$

Therefore, the expression for gravity force becomes as follows

$$F_g = \frac{4}{3}\pi r^3 \cdot \rho_{\text{oil}} g$$

7. The buoyancy force (Archimedes' principle **F_A**)

$$F_A = m' g$$

$$\rho_{\text{air}} = \frac{m'}{V} \Rightarrow m' = \rho_{\text{air}} \cdot V$$

$$F_b = \frac{4}{3}\pi r^3 \cdot \rho_{\text{air}} g$$

where ρ_{air} is the volumetric mass of the air.

8. **The Stoke's force F_s (drag force)**: It is the force of air resistance to the motion. this force is related only to the velocity of the droplet, using Stoke's law is:

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

Here: **r** is the radius of the drop and **η** , the viscosity constant of air.

$$\sum \mathbf{F}_{\text{ext}} = m\boldsymbol{\gamma}$$

When the droplet reaches terminal velocity

$$\vec{F}_g + \vec{F}_A + \vec{F}_S = m\vec{\gamma} = 0$$

$$P - F_A - F_S = 0$$

$$P - F_A = F_S$$

$$\frac{4}{3}\pi r^3 (\rho_{\text{oil}} - \rho_{\text{air}})g = 6\pi\eta r v_0$$

$$r^2 = \frac{9\eta v_0}{2(\rho_{\text{oil}} - \rho_{\text{air}})g}$$

$$r = \sqrt{\frac{9\eta v_0}{2(\rho_{\text{oil}} - \rho_{\text{air}})g}}$$

$$r = 3\sqrt{\frac{\eta v_0}{2(\rho_{\text{oil}} - \rho_{\text{air}})g}}$$

9. **archimedes force in front of gravity is usually neglected** And that's because :

$$\rho_{\text{air}} \ll \rho_{\text{oil}}$$

hence the expression for the radius becomes

$$r = 3\sqrt{\frac{\eta v_0}{2\rho_{\text{oil}}g}}$$

2- In the presence of an electric field:

- The droplet is rising

When an electric field is applied, the droplet rises with a speed v_1

$$\vec{F}_e + \vec{p} + \vec{F}_A + \vec{F}_S = \vec{0}$$

$$\vec{F}_e + \vec{F}_A - \vec{p} - \vec{F}_S = \vec{0}$$

$$qE + \frac{4}{3}\pi r^3 \rho_{oi} - mg - 6\pi\eta r v_1 = 0$$

$$qE = mg + 6\pi\eta r v_1 - \frac{4}{3}\pi r^3 \rho_{oil}$$

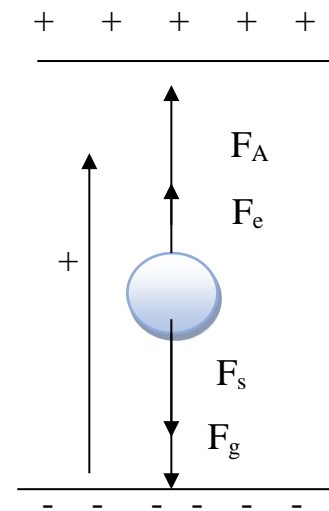
we have

$$mg = 6\pi\eta r v_0 + \frac{4}{3}\pi r^3 \rho_{oil}$$

$$\Rightarrow qE = 6\pi\eta r v_0 + \frac{4}{3}\pi r^3 \rho_{oil} + 6\pi\eta r v_1 - \frac{4}{3}\pi r^3 \rho_{oil}$$

$$qE = 6\pi\eta r (v_0 + v_1)$$

$$q = \frac{6\pi\eta r (v_0 + v_1)}{E}$$



Using this technique Millikan calculated the charge of an electron equal to:

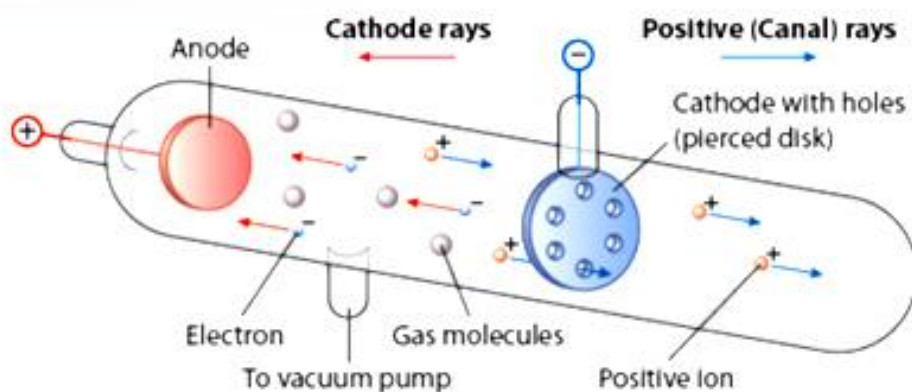
$$e = 1.602 \times 10^{-19} \text{ Coulombs}$$

From these data he calculated the mass of an electron:

$$m = 9.109 \times 10^{-28} \text{ g}$$

The Nucleus discovery:

Goldstein (1886) repeated the experiment with a discharge tube filled with a perforated cathode and found that new type of rays came out through the hole in the cathode.



When this experiment is conducted, a faint red glow is observed on the wall behind the cathode. Since these rays originate from the anode, they are called anode rays (canal rays)

Properties of Anode Rays:

10. Anode rays travel in straight lines in a direction opposite to the cathode rays.
11. Anode rays are material particles.
12. Anode rays are positively charged.
13. Anode rays may get swung by external magnetic field.
14. Its mass varies by changing the type of gas in the discharge tube.
15. Anode rays also affect the photographic plate. These anode rays produce flashes of light on ZnS screen

Nucleus components:

16. Proton (rutherford 1919):

Ernest Rutherford bombarded nitrogen atoms with helium nuclei (alpha particles), and hydrogen atoms (protons) were formed as a result. From this, he concluded that nitrogen nuclei contained protons

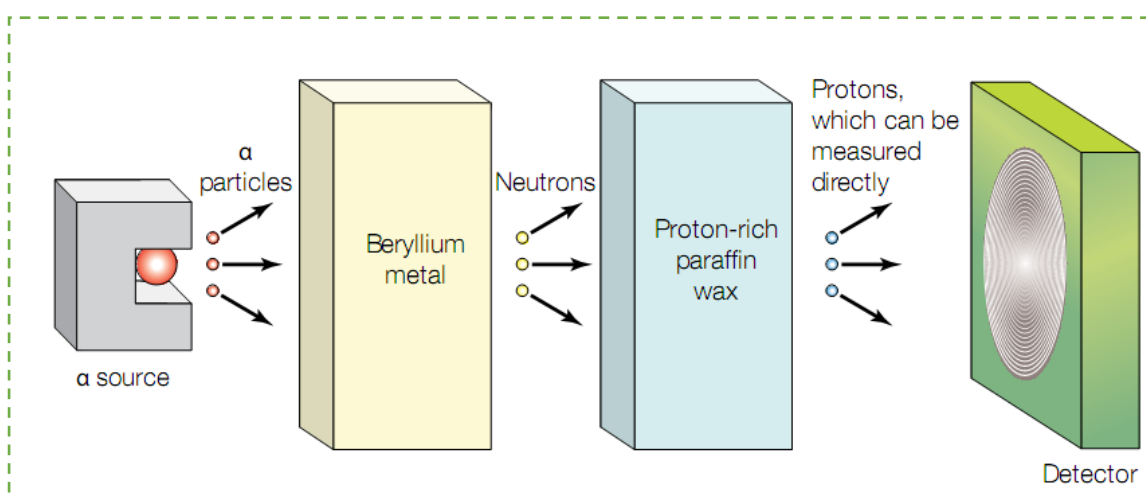


Protons Characteristics

- The charge on a proton is positive.
- The nucleus contains protons.
- The rest mass of a proton is $m_p = 1.6726210 \cdot 10^{-27} \text{ kg}$, i.e., 1836 times that of an electron.
- One proton has a charge of $q_p = 1.602 \cdot 10^{-19} \text{ C}$.

17. Neutron:

In 1920, Rutherford predicted the existence of the neutron because he realized that the nucleus would be unstable if it were only made up of protons. Chadwick confirmed this prediction in 1932 by bombarding beryllium (Be) with alpha particles (He) and observing the production of neutral radiation which not affected by magnetic field and electric field. These radiations consisted of neutral particles, which were called neutrons



Experimentation of neutron discovery

Neutrons Characteristics

- Neutrons do not carry any charge
- The mass of a neutron is approximately the same as that of a proton. A neutron has an absolute mass of $1.674927 \times 10^{-27}\text{kg}$.

Note: Protons and neutrons are composed of quarks.

There are six different types of quarks: **Up, down, top (truth), bottom (beauty), charm, and strange.**

Quarks have an electrical charge, for example, up quarks have a charge of $+\frac{2}{3}$ while

down quarks have a charge of $-\frac{1}{3}$

The Nuclear Model of the Atom

(C) Thomson's Model of the atom :

An atom is electrically neutral. It contains positive charges (due to the presence of protons) as well as negative charges (due to the presence of electrons). It assumes that mass is equally distributed in the atom. Hence, J.J. Thomson assumed that an atom is a uniform sphere of positive charges with electrons embedded in it.

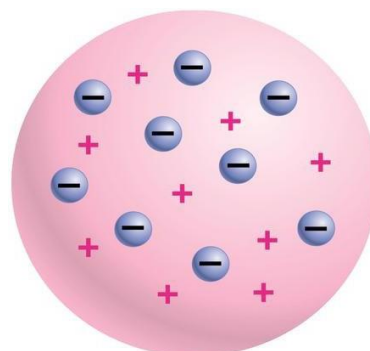
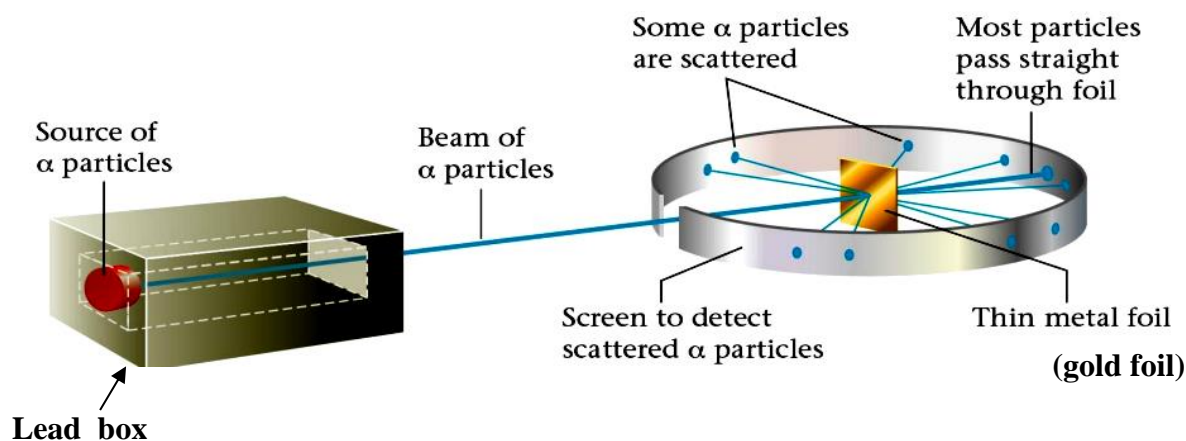


Figure-4 Thomson's atomic model

(D) Rutherford's Model of the atom

Rutherford's Experiment: Rutherford conducted a series of experiments using α -particles. A beam of α -particles was directed against a thin foil of gold, platinum, silver, or copper. The foil was surrounded by a circular fluorescent zinc screen. Whenever an α -particle struck the screen, it produced a flash of light.



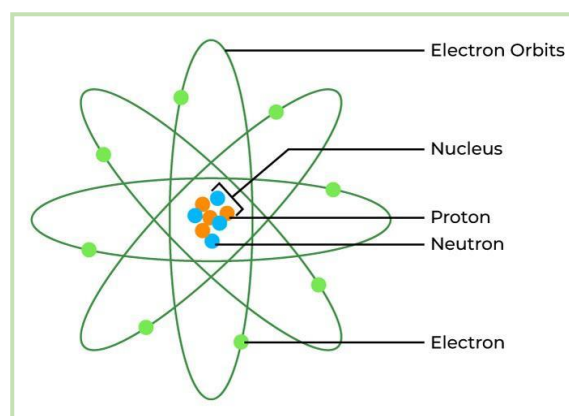
Rutherford's observations and conclusions:

1. Most of the α -particles passed straight through the gold foil without being deflected, indicating that there must be very large empty space within the atom.
2. A few of them were deflected through small angles, while a very few were deflected to a large extent, indicating that whole of the positive charge is concentrated in a space called nucleus. It is proposed to be present at the center of the atom.

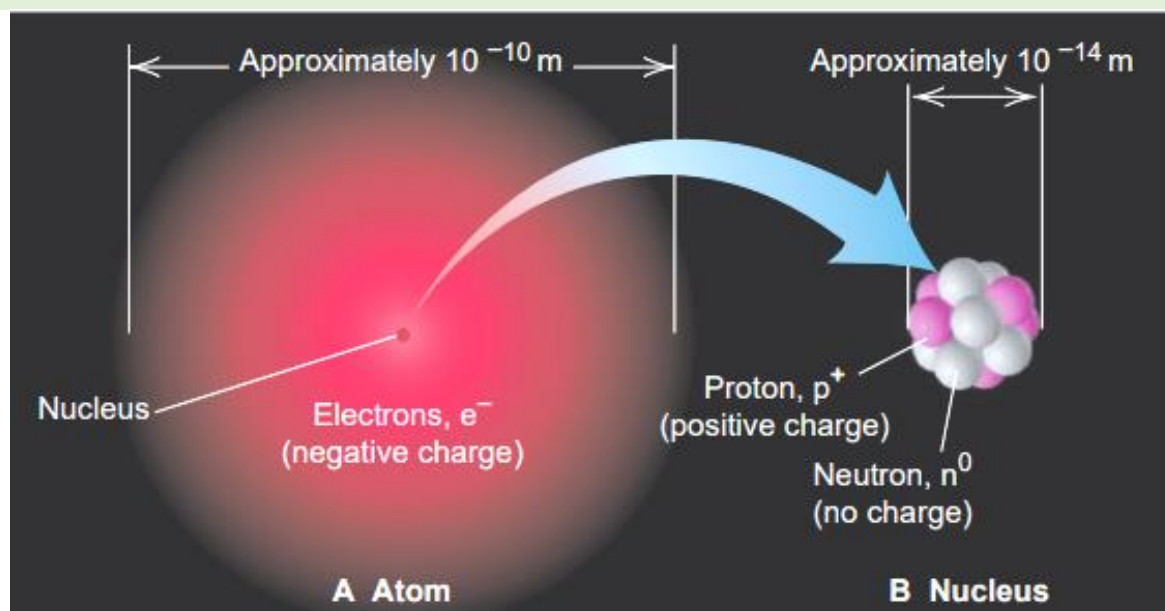
3. A very small percentage (1 in 10000) was deflected through angles ranging from nearly 180° , indicating that the nucleus is rigid and α -particles recoil due to direct collision with the positively charged heavy mass.

Rutherford's nuclear concept of the atom

- 1) The atom of an element consists of a small positively charged 'nucleus' which is situated at the center of the atom and which carries almost the entire mass of the atom.
- 2) The electrons are distributed in the empty space of the atom around the nucleus in different concentric circular paths, called orbits.
- 3) The number of electrons in orbits is equal to the number of positive charges (protons) in the nucleus. Hence, the atom is electrically neutral.
- 4) The volume of the nucleus is negligibly small as compared to the volume of the atom.



$$\frac{\text{Diameter of atom}}{\text{Diameter of the nucleus}} = \frac{r_a}{r_n} = \frac{10^{-10}}{10^{-14}} = 10^4$$

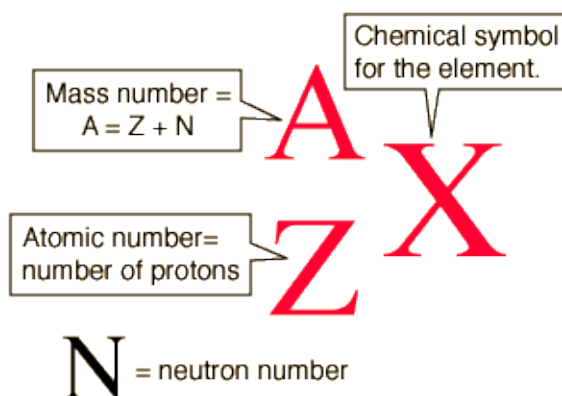


Atomic Number, Mass Number, and Atomic Symbol

The **atomic number** (Z) of an element equals the number of protons in the nucleus of each of its atoms.

The total number of protons and neutrons in the nucleus of an atom is its **mass number** (A).

The nuclear mass number and charge are often written with the **atomic symbol** (or *element symbol*). The atomic number (Z) is written as a left **subscript** and the mass number (A) as a left **superscript** to the symbol:



The number of neutrons equals the mass number minus the atomic number:

$$\text{number of neutrons } (N) = A - Z$$

Isotopes, isobars, isotones and isodiaphers:

Isotopes

Isotopes are atoms of the **same element** that have **different numbers of neutrons** (they also have **different mass numbers**). They have the same number of protons, which determines the element's identity, but a different number of neutrons, which determines the isotope's mass number.

Average Atomic Mass

The average atomic mass of an element is the sum of the masses of its isotopes, each multiplied by its natural abundance (the decimal associated with percent of atoms of that element that are of a given isotope).

$$\begin{aligned} \text{Average atomic mass} &= \frac{\sum \text{abundance} \times \text{atomic mass of isotope}}{100} \\ &= \frac{\sum M_i A_i}{100} \end{aligned}$$

Where: **M** represent atomic mass

and **A** represent Relative Abundance

Isobars are atoms of different chemical elements (different atomic numbers) with equal atomic mass values. ${}^{40}_{18}\text{Ar}$, ${}^{40}_{19}\text{K}$, ${}^{40}_{20}\text{Ca}$

Isotones are atoms of different chemical elements with an equal number of neutrons in the atomic nucleus.

Isodiaphers are the atoms of different element which have the same difference of the number of Neutrons and protons (the same n-p)

Example: Naturally occurring chlorine consists of ${}^{35}\text{Cl}$ (mass 34.96885 amu) and ${}^{37}\text{Cl}$ (mass 36.96590 amu), with an average mass of 35.453 amu. What is the percent composition of Cl in terms of these two isotopes?

Solution: The average mass of chlorine is the fraction that is ${}^{35}\text{Cl}$ times the mass of ${}^{35}\text{Cl}$ plus the fraction that is ${}^{37}\text{Cl}$ times the mass of ${}^{37}\text{Cl}$.

$$\text{Average atomic mass} = \frac{\sum \text{abundance} \times \text{atomic mass of isotope}}{100} = \frac{\sum M_i A_i}{100}$$

average mass

$$= \frac{(\text{abundance of } {}^{35}\text{Cl} \times \text{mass of } {}^{35}\text{Cl}) + (\text{abundance of } {}^{37}\text{Cl} \times \text{mass of } {}^{37}\text{Cl})}{100}$$

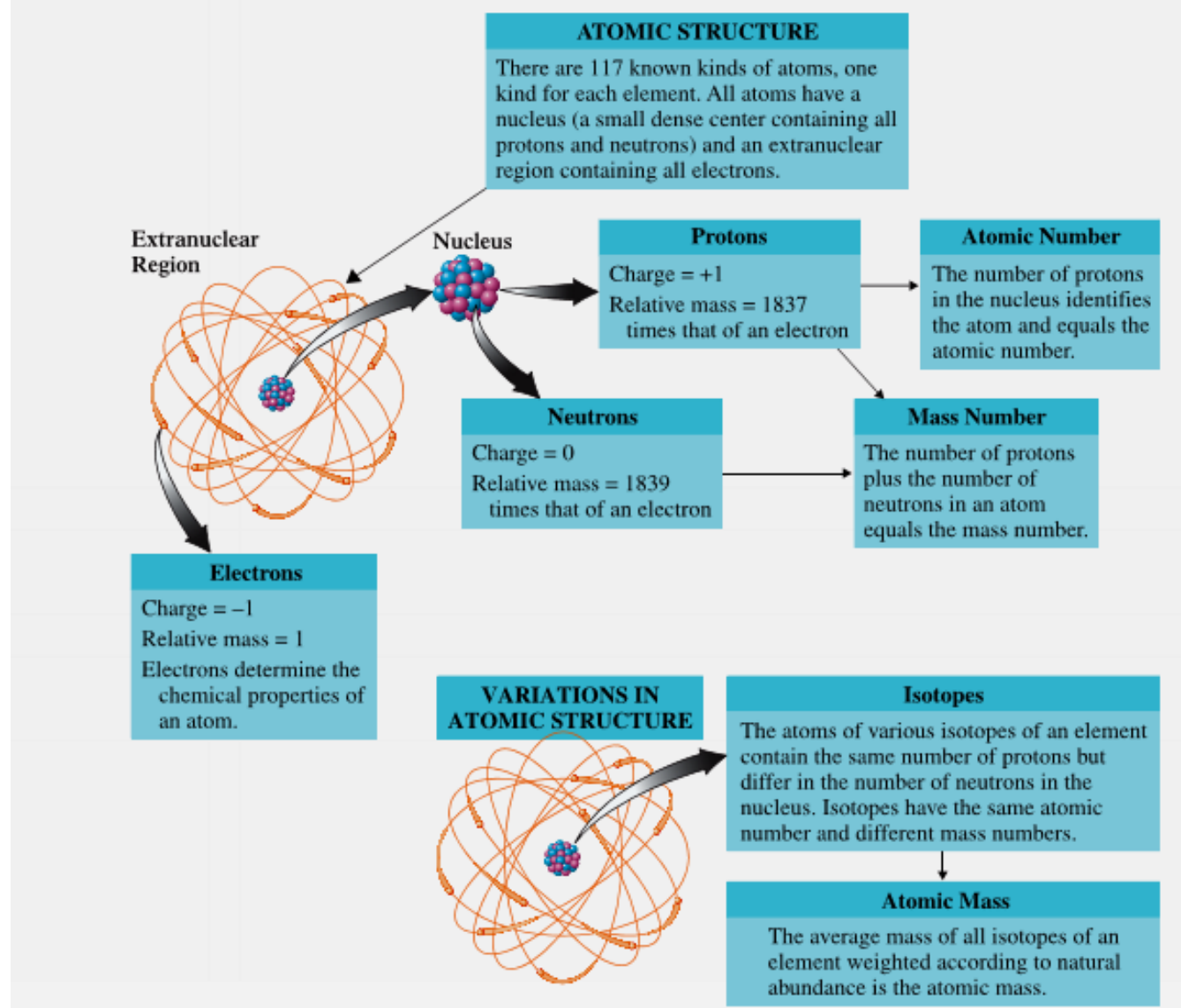
If we let x represent the fraction that is ${}^{35}\text{Cl}$, then the fraction that is ${}^{37}\text{Cl}$ is represented by $1-x$. (The fraction that is ${}^{35}\text{Cl}$ + the fraction that is ${}^{37}\text{Cl}$ must add up to 1, so the fraction of ${}^{37}\text{Cl}$ must equal $1.00 -$ the fraction of ${}^{35}\text{Cl}$.) Substituting this into the average mass equation, we have:

$$\begin{aligned}
 35.453 \text{ amu} &= (x \times 34.96885 \text{ amu}) + [(1 - x) \times 36.96590 \text{ amu}] \\
 &= 34.96885x + 36.96590 - 36.96590x \\
 &= 1.99705x + 36.96590 - 36.96590x \\
 &= 1.99705x - 36.96590x + 36.96590 \\
 &= -34.96885x + 36.96590 \\
 &= 1.99705x - 34.96885x + 36.96590 \\
 &= -32.9718x + 36.96590 \\
 &= 1.99705x = 1.99705x - 32.9718x + 36.96590 \\
 &= -30.97475x + 36.96590 \\
 &= 1.99705x = -30.97475x + 36.96590 \\
 &= 32.9718x = 36.96590 \\
 &= x = \frac{36.96590}{32.9718} = 1.12418 \\
 &= x = 0.7576
 \end{aligned}$$

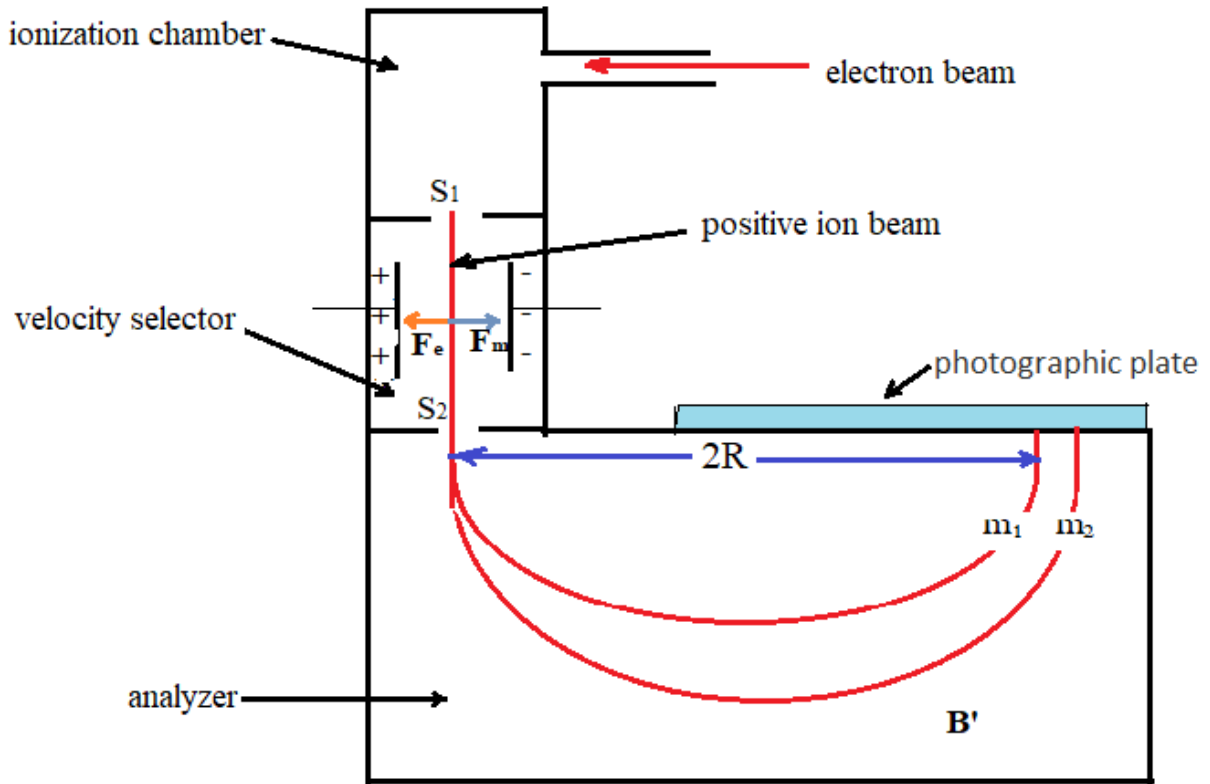
So solving yields: $x = 0.7576$, which means that $1 - 0.7576 = 0.2424$.

Therefore, chlorine consists of 75.76% ^{35}Cl and 24.24% ^{37}Cl

Atomic Structure



Bainbridge mass spectrometer - Determination of isotopic masses of nuclei



Ions are formed at ionization chamber and pass through a slit S_1 . They then travel between two plates, which a potential (V) is applied. A magnetic field (strength B) is applied at right angles to the electrostatic field and so the electrostatic and electromagnetic forces act in opposite directions to each other.

A charged particle with velocity v will only pass through the next slit if the total force on it is zero, meaning that it is traveling in a straight line.

$$\begin{aligned} \vec{F}_e &= \vec{F}_m \\ qE &= Bqv \\ v &= \frac{E}{B} \end{aligned}$$

Only ions with velocity v will pass through the velocity selector and slit S3 to the analyzer, where they encounter another uniform magnetic field B . Here, they are deflected along a circular path of radius R to strike the photographic plate. The centripetal force is provided by this magnetic field. Therefore,

$$F_m = \frac{mv^2}{R}$$

$$B'qv = \frac{mv^2}{R}$$

$$R = \frac{mv}{B'q}$$

Substituting $v = \frac{E}{B}$

$$R = \frac{mE}{qBB'}$$

$$R_1 = \frac{m_1E}{qBB'}$$

$$R_2 = \frac{m_2E}{qBB'}$$

The distance (d) between the two points of impact of the two ions on the photographic plate

$$d = D_2 - D_1 = 2R_2 - 2R_1 = 2 \left(\frac{m_2E}{qBB'} - \frac{m_1E}{qBB'} \right) = 2 \frac{E}{qBB'} (m_2 - m_1)$$

Since q , B , B' , E and R are known, the mass of the positive ions and hence isotopic masses can be calculated.

Mass defect, nuclear binding energy

Mass defect: According to nuclear particle experiments, the total mass of a nucleus (m_{nuc}) is **less** than the sum of the masses of its constituent nucleons (protons and neutrons). The mass difference, or **mass defect**, is defined as the difference between the predicted mass of an atom's nucleus (the sum of the masses of its protons and neutrons), and its actual mass.

The mass defect can be calculated using equation

$$\Delta m = [Z(m_p + m_e) + (A - Z)m_n] - m_{atom}$$

where: Δm : mass defect [atomic mass unit (amu)];

m_p : mass of a proton (1.007277 amu);

m_n : mass of a neutron (1.008665 amu); m_e = mass of an electron (0.000548597 amu);

m_{atom} : mass of nuclide X Z A (amu); Z = atomic number (number of protons); and A = mass number (number of nucleons).

This difference is linked to the existence of an energy called **cohesion energy** or **binding energy of the nucleus** which is linked to the mass defect by the mass-energy equivalence based on Einstein's relationship:

$$\Delta E = \Delta m \cdot c^2$$

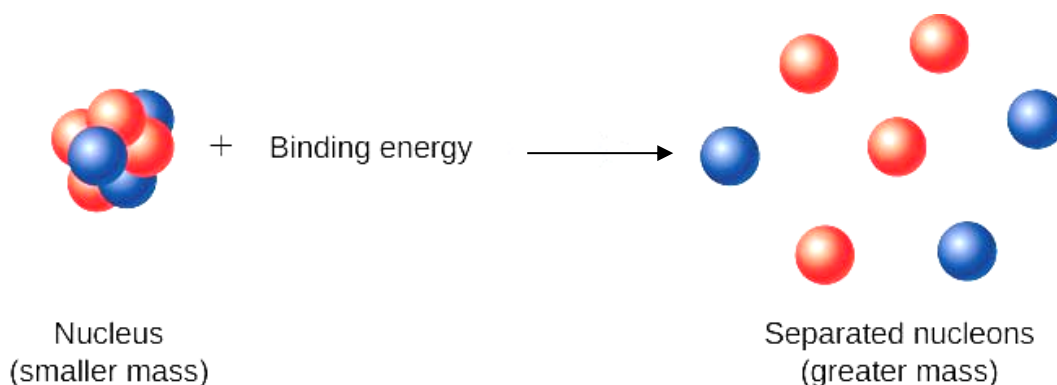
where ΔE is energy released,

Δm is mass defect,

and c is the speed of light.

Nuclear Binding energy (E):

Nuclear binding energy is the energy required to separate an atomic nucleus completely into its constituent protons and neutrons, or, equivalently, the energy that would be liberated by combining individual protons and neutrons into a single nucleus.



Binding Energy per Nucleon

In nuclear physics, one of the most important experimental quantities is the **binding energy per nucleon (BEN) or average binding energy**, which is defined as the total binding energy divided by the number of nucleons (protons and neutrons), .

$$BEN = \frac{\Delta E}{A}$$

Binding energy per nucleon is a good indication of nuclear stability

Energy equivalent of one atomic mass unit:

Example: Find the energy equivalent of one atomic mass unit, first in Joules and then in MeV.

Using this, express the mass defect of ${}^{16}_8\text{O}$ in MeV.

Solution : As we have, $1\text{amu} = 1.66 \times 10^{-27} \text{ kg}$

To convert it into energy units, we use the Einstein's mass energy equivalence given by

$$\Delta E = \Delta m \times c^2$$

Where: c = velocity of light, $c = 2.9979 \times 10^8 \text{ m/s}$

We have $m = 1\text{amu} = 1.6605 \times 10^{-27} \text{ kg}$

Hence we can write

$$\Delta E = 1.6605 \times 10^{-27} \times (2.9979 \times 10^8)^2 \text{ kg} \frac{\text{m}^2}{\text{s}^2} = 1.4923 \times 10^{-10} \text{ J}$$

We have $1\text{eV} = 1.6 \cdot 10^{-19} \text{ J}$

$$\Delta E = \frac{1.4923 \times 10^{-10}}{1.602 \times 10^{-19}} \text{ eV} = 0.9315 \times 10^9 \text{ eV}$$

$$1\text{MeV} = 10^6 \text{ eV}$$

$$\Delta E = 931.5 \text{ MeV}$$

Hence we have proved that energy equivalent to $1\text{amu} = 931.5 \text{ MeV}$

Example: Calculate the energy released in MeV needed to separate $^{16}_8\text{O}$, into its constituents

$m_n = 1.00867 \text{ amu}$, $m_p = 1.00728 \text{ amu}$, $m_e = 0.000549 \text{ amu}$, $m_o = 15.999 \text{ amu}$

$$\Delta m = [Z(m_p + m_e) + (A - Z)m_n] - m_{atom}$$

$$\Delta m = [8(1.00728 + 0.000549) + (16 - 8) \times 1.00867] - 15.999$$

$$\Delta m = 0,132992 \text{ amu}$$

We have $1 \text{ amu} = 931.5 \text{ MeV}$

$$\Delta E = 0.132992 \times 931.5 = 123,882 \text{ MeV}$$

The energy needed to separate $^{16}_8\text{O}$ into its constituents is thus $123,882 \text{ MeV}$.

Series No. 2

Exercise 1: Calculate the speed of an electron after it has been accelerated from rest through a potential difference of 5000 V.

The mass of an electron = 9.11×10^{-31} kg. $e = 1.60 \times 10^{-19}$ C.

Exercise2: A beam of electrons is directed horizontally into a uniform electric field which acts vertically downwards. The electric field is due to a potential difference of 4500 V applied between two parallel deflecting plates 60 mm apart. With the electric field on, a uniform magnetic field is applied to the beam perpendicular to the electric field and to the initial direction of the beam. The magnetic flux density is adjusted to a value of 2.4×10^{-3} T so the beam is undeflected.

1. State the direction of:
 - a. the electric force
 - b. the magnetic force.
2. Calculate the speed of the electrons as they pass through the electric and magnetic fields.

$$e = 1.60 \times 10^{-19} \text{ C}$$

Exercise 3: Electrons moving at a constant speed are directed horizontally into a uniform electric field due to two parallel plates of length 60.0 mm. The plates are spaced 50.0 mm apart and have a potential difference of 3000 V between them. The electrons are deflected by 22.0 mm as a result. When a uniform magnetic field of flux density 2.05×10^{-3} T, is applied at right angles to the beam and the electric field, the beam is undeflected.

1. Calculate the speed of the beam
2. Calculate the specific charge e/m of the electron.

Exercise 4: A charged oil droplet was held stationary between oppositely charged metal plates 5.0 mm apart when the potential difference between the plates was 490 V. When the potential difference was switched off, the droplet fell at a terminal speed of 1.3×10^{-4} m/s . Calculate:

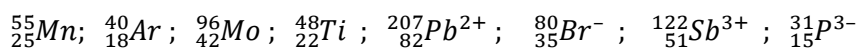
1. the mass of the droplet
2. the charge of the droplet

$$\text{Data: } \rho_{\text{oil}} = 960 \text{ Kg/m}^3 ; g = 9.81 \text{ m/s} ; \eta = 18 \times 10^{-6} \text{ (MKSA)}$$

Exercise 5: Millikan's apparatus is used to study the movement of a charged droplet of oil in the air.

1. In the absence of an electric field, the droplet falls a distance of 2.66 mm in 12 S.
- Calculate the radius of the droplet and deduce its mass.
2. The droplet remains stable when we subject it to an electric field. The distance between the capacitor leads is 2cm and the difference in potential between them is equal to 432 volts.
- What is the charge of the droplet acquired? Compare this charge with the elementary charge.
What do you conclude? Data: $\rho_{oil} = 900 \text{ Kg/m}^3$; $g = 9.81 \text{ m/s}^2$; $\eta = 18 \times 10^{-6} \text{ (MKSA)}$

Exercise 6: How many neutrons, protons, and electrons are present in each of the following atoms or ions?



Exercise 7: Naturally occurring iron (${}_{26}\text{Fe}$) consists of four stable isotopes (listed with their atomic mass and abundance, respectively): ${}^{54}\text{Fe}$ (atomic mass= 53,953, abundance 6,04 %), ${}^{57}\text{Fe}$ (atomic mass= 56,960, abundance 2,11 %), ${}^{56}\text{Fe}$ (atomic mass= 55,94891, abundance 57 %), ${}^{58}\text{Fe}$ (atomic mass= 57,9590, abundance 28 %)

Calculate the average atomic mass of natural iron

Exercise 8: 1). The atomic mass of lithium is 6.943 amu and it has two isotopes with atomic masses of 6.017 amu and 7.018 amu.

Give the formula of the two isotopes in the form ${}^A_Z\text{X}$ and determine the natural abundance of the two isotopes.

2) Classify the following nuclei according to the following division: -1 Isotopes -2 Isobars -3 Isotones: ${}^{15}_8\text{O}$, ${}^{14}_7\text{N}$, ${}^{13}_6\text{C}$, ${}^{12}_5\text{B}$, ${}^{17}_8\text{O}$, ${}^{15}_7\text{N}$, ${}^{16}_8\text{O}$, ${}^{18}_9\text{F}$, ${}^{17}_9\text{F}$

Exercise 9: Using the Bainbridge mass spectrograph, we separate two isotopes, carrying a positive elementary charge (+e): ${}^{20}\text{Ne}^+$ and ${}^{21}\text{Ne}^+$. Their speed at the entrance to the analyzer $V = 10^4 \text{ m/s}$.

In this spectrograph part and under the effect of the magnetic induction field $B = 0.1 \text{ Tesla}$, the trajectories of these ${}^{20}\text{Ne}^+$ and ${}^{21}\text{Ne}^+$ isotopes are respectively semi-circles of radii R_1 and R_2 . Calculate R_1 and R_2 and deduce the distance (d) between the two points of impact of the two ions on the photographic plate.

Data: $1 \text{ uma} = 1,67 \cdot 10^{-27} \text{ Kg}$

Correction

Exercise 1 :

Here we need to find the velocity of travelling electrons using the given stopping potential.

We know that

$$eU = \frac{1}{2}mv^2$$

$$v^2 = \frac{2eU}{m} \Rightarrow v = \sqrt{\frac{2eU}{m}}$$

The charge(e) and mass(m) of electron are also given as, $e=1.602 \times 10^{-19}$ and $m=9.11 \times 10^{-31}$ Kg

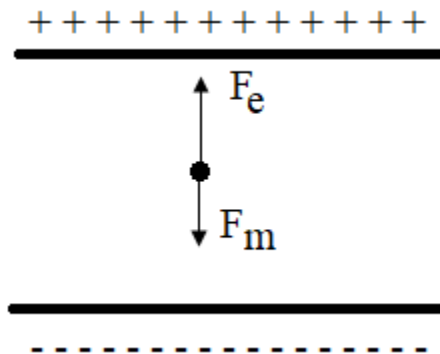
Now substituting the values of e, m, U.

$$v = \sqrt{\frac{2 \times (1.602 \times 10^{-19}) \times 5000}{9.11 \times 10^{-31}}}$$

$$v = 1.33 \times 10^6 \text{ m/s}$$

Exercise 2: 1.

- a. The electric force is vertically upwards (in the opposite direction to the electric field lines because the electron has a negative charge).



- b. The magnetic force is vertically downwards (in the opposite direction to the electric force).
2. Using $F_e = F_m \implies qE = qvB$
 We have $E = \frac{U}{d}$

so $q \frac{v}{d} = qvB \implies$ the speed $v = \frac{v}{dB} = \frac{4500}{60 \times 10^{-3} \times 2.4 \times 10^{-3}} = 3.2 \times 10^7 \text{ m/s}$

Exo3:

1. The speed of the beam

$$\|\vec{F}_e\| = \|\vec{F}_m\|$$

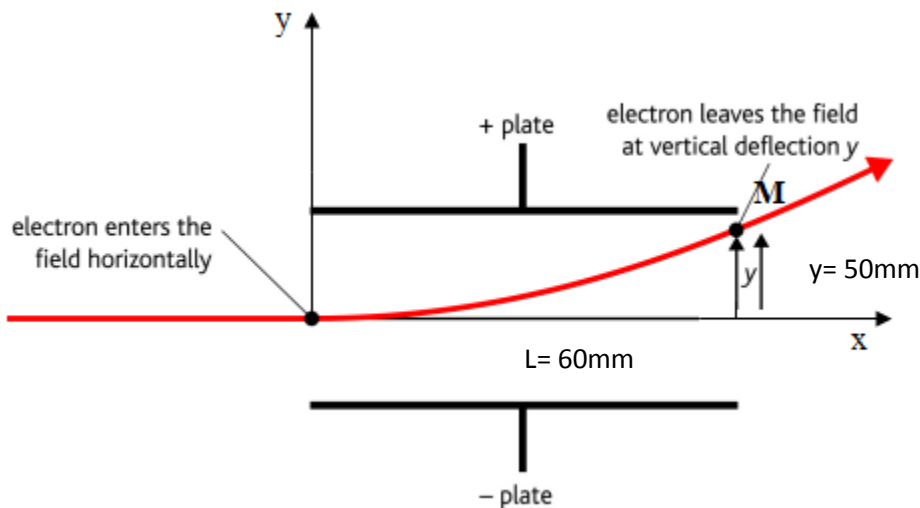
$$qE = qvB \implies v = \frac{E}{B}$$

we know that $E = \frac{U}{d}$

$$\implies v = \frac{U}{Bd} = \frac{3000}{2.05 \times 10^{-3} \times 50 \times 10^{-3}} = 2.93 \times 10^7 \text{ m/s}$$

2. Calculate the specific charge e/m of the electron

When the magnetic field is deleted, the beam is deflected by the electric field E towards the positive plate. When it leaves the electric field, it collides with the screen at point M



The total forces acting on the beam are given as follows:

$$\sum F_{ext} = m_e \gamma$$

$$F_e = qE = m_e \gamma$$

By projecting the force F_e onto the x axis

$$\begin{cases} F_{e_x} = 0 \\ F_e = m \gamma_x \end{cases} \implies m \gamma_x = 0 \implies \gamma_x = 0$$

$$\Rightarrow v = Cst$$

So the movement along the ox axis is a uniform straight line:

$$\begin{aligned} \mathbf{x} &= \mathbf{v} \mathbf{t} \\ \mathbf{t} &= \frac{\mathbf{x}}{\mathbf{v}} \end{aligned}$$

By projecting the force F_e onto the (oy) axis we find:

$$\begin{cases} F_{e_y} = qE \\ F_e = m\gamma_y \end{cases} \Rightarrow m\gamma_y = qE \Rightarrow \gamma_y = \frac{qE}{m}$$

Therefore the movement along the (oy) axis is uniformly accelerated motion:

$$\mathbf{Y} = \frac{1}{2} \boldsymbol{\gamma} t^2 = \frac{1}{2} (qE / m) t^2$$

We substitute the time expression into y

$$y = \frac{1}{2} \frac{qE}{m_e} \left(\frac{x}{v} \right)^2 \Rightarrow \frac{q}{m_e} = 2 \cdot (y) \cdot \frac{dv^2}{UL^2}$$

$$\frac{q}{m_e} = 2 \cdot (22 \times 10^{-3}) \cdot \frac{50 \times 10^{-3} (2.93 \times 10^7)^2}{3000 \times (60 \times 10^{-3})^2} = 1.748 \times 10^{11} \text{ Coulomb/kg} - 1$$

Exo 4: density of the oil = 960 kg m^{-3} , viscosity of air $\eta = 1.8 \times 10^{-5} \text{ N s / m}^2$

First, with zero applied electric field, the droplet falls, with a limit speed v_0

$$\sum \mathbf{F}_{\text{ext}} = m\boldsymbol{\gamma}$$

When the droplet reaches terminal velocity ($v = \text{constant}$)

$$\rightarrow \mathbf{F}_g + \mathbf{F}_s = m\boldsymbol{\gamma} = 0$$

$$F_g - F_S = 0$$

$$P = F_S$$

$$\frac{4}{3} \pi r^3 \rho_{\text{oil}} g = 6 \pi \eta r v_0$$

$$r^2 = \frac{9 \eta v_0}{2 \rho_{\text{oil}} g}$$

$$r^2 = \frac{9 \eta v}{2 \rho g} = \frac{9 \times 1.8 \times 10^{-5} \times 1.3 \times 10^{-4}}{2 \times 960 \times 9.81} = 1.12 \times 10^{-12} \text{ m}^2$$

$$\text{hence } r = 1.06 \times 10^{-6} \text{ m}$$

$$m = \frac{4}{3} \pi r^3 \rho = \frac{4}{3} \pi (1.06 \times 10^{-6})^3 \times 960 = 4.75 \times 10^{-15} \text{ kg}$$

When an electric field is applied, the droplet was held stationary ($v=0$)

$$\rightarrow \rightarrow \rightarrow \rightarrow$$

$$\mathbf{F}_e + \mathbf{p} = \mathbf{0}$$

$$\mathbf{F}_e - \mathbf{p} = \mathbf{0}$$

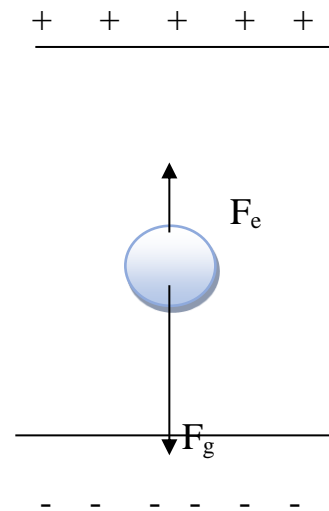
$$qE = mg$$

$$q = \frac{mg}{E}$$

we have $E = \frac{U}{d}$

$$q = \frac{mgd}{U}$$

$$= \frac{4.75 \times 10^{-15} \times 9.81 \times 5.0 \times 10^{-3}}{490} = 4.8 \times 10^{-19} \text{ C}$$



Exo5:

$$d = v_0 t \Rightarrow v_0 = \frac{d}{t} = \frac{2.66 \times 10^{-3}}{12} = 2.26 \times 10^{-4} \text{ m/s}$$

First, with zero applied electric field, the droplet falls,

$$\sum \mathbf{F}_{\text{ext}} = m\mathbf{y}$$

When the droplet reaches terminal velocity ($v = \text{constant}$)

$$\begin{aligned} \rightarrow F_g + F_S &= m\mathbf{y} = 0 \\ F_g - F_S &= 0 \\ P &= F_S \\ \frac{4}{3} \pi r^3 \rho_{\text{oil}} g &= 6 \pi \eta r v_0 \\ r^2 &= \frac{9 \eta v_0}{2 \rho_{\text{oil}} g} \end{aligned}$$

$$r = 3 \sqrt{\frac{\eta v_0}{2 \rho_{\text{oil}} g}} = 3 \sqrt{\frac{18 \times 10^{-6} \times 2.26 \times 10^{-4}}{2 \times 900 \times 9.81}} = 1.43 \times 10^{-6} \text{ m}$$

$$m = \frac{4}{3} \pi r^3 \cdot \rho_{\text{oil}} = \frac{4}{3} \pi (1.43 \times 10^{-6})^3 = 12,248 \times 10^{-18} \text{ kg}$$

2- In the presence of an electric field:

- The droplet is rising

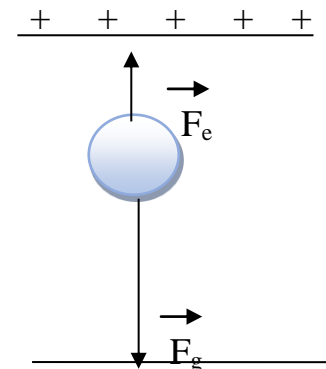
When an electric field is applied, the droplet rises with a speed v_1

droplet was held stationary ($v=0$)

$$\begin{aligned} \rightarrow \rightarrow \\ \mathbf{F}_e + \mathbf{F}_g &= \mathbf{0} \\ \mathbf{F}_e - \mathbf{p} &= \mathbf{0} \\ qE &= mg \\ q &= \frac{mg}{E} \end{aligned}$$

we have $E = \frac{U}{d}$

$$q = \frac{mgd}{U} = \frac{12.248 \times 10^{-18} \times 9.81 \times 2 \times 10^{-2}}{4320}$$



itre
A-

| | | | | | |
|------------------|----|-----|----|-----|----|
| Ar | 18 | 40 | 18 | 22 | 18 |
| Mo | 42 | 96 | 42 | 54 | 42 |
| Ti | 22 | 48 | 22 | 26 | 22 |
| Pb ²⁺ | 82 | 207 | 82 | 125 | 80 |
| Br ⁻ | 35 | 80 | 35 | 45 | 36 |
| Sb ³⁺ | 51 | 122 | 51 | 71 | 48 |
| P ³⁻ | 15 | 31 | 15 | 16 | 18 |

Exo7: It simply involves calculating the weighted average of the four atomic masses:

$$\text{Average atomic mass of Fe} = \frac{\sum \text{abundance} \times \text{atomic mass of isotope}}{100} = \frac{\sum M_i A_i}{100}$$

$$A_{\text{moy}} = \frac{53,953 (6,04) + 55,948 (91,57) + 56,960 (2,11) + 57,959 (0,28)}{100} = 55,854 \text{ Cette valeur est}$$

conforme à la masse atomique expérimentale : 55,85 (cf. Annexe E)

Exo8 : I. If x and y are the isotopic percentages of ${}^6\text{Li}$ and ${}^7\text{Li}$, such that

$$x + y = 1(100\%)$$

and

$$6.017 x + 7.018 y = 6.943,$$

we find: ${}^6\text{Li}$: 7.4%; ${}^7\text{Li}$: 92.6%.

II.

| Isotopes (have same atomic number Z) | Isobars (have same weight number A) | Isotones (have same number of neutron N) |
|--|---|---|
| ${}^{15}_8\text{O}, {}^{16}_8\text{O}, {}^{17}_8\text{O}$ ${}^{18}_9\text{F}, {}^{17}_9\text{F}$ ${}^{15}_7\text{N}, {}^{14}_7\text{N},$ | ${}^{15}_7\text{N}, {}^{15}_8\text{O},$ ${}^{17}_9\text{F}, {}^{17}_8\text{O}$ | ${}^{15}_7\text{N}, {}^{17}_9\text{F}, {}^{16}_8\text{O}$ ${}^{18}_9\text{F}, {}^{17}_8\text{O}$ ${}^{12}_5\text{B}, {}^{13}_6\text{C}, {}^{14}_7\text{N}, {}^{15}_8\text{O}$ |

Calculation of radii R_1 and R_2

$$F_e = F_m$$

$$qE = Bqv$$

$$v = \frac{E}{B}$$

$$B'qv = \frac{mv^2}{R}$$

$$R = \frac{mv}{B'q} = \frac{mE}{qBB'} = \frac{19.9924 \times 1.66 \times 10^{-27} \times 10^4}{1.6 \times 10^{-19} \times 0.1} = 0.02074m = 2.074m$$

$$R = \frac{mE}{qBB'} = \frac{20.9939 \times 1.66 \times 10^{-27} \times 10^4}{1.6 \times 10^{-19} \times 0.1} = 0.02178m = 2.178m$$

The distance (d) between the two points of impact of the two ions on the photographic plate

$$d = D_2 - D_1 = 2(R_2 - R_1) = 2.08 \times 10^{-3}m$$

Chapter three

Radioactivity

Chapter three: Radioactivity

VIII. Introduction

The phenomenon of radioactivity was discovered by Antoine Henri Becquerel in 1896. He discovered that photographic plates develop bright spots when exposed to uranium minerals, and he concluded that the minerals give off some sort of radiation.

IX. Stable and unstable nuclides

A nuclear reaction is a reaction in which changes occur in the nucleus of an atom.

Nuclear reactions are not considered to be ordinary chemical reactions. The governing principles for ordinary chemical reactions deal with the rearrangement of electrons; this rearrangement occurs as the result of electron transfer or electron sharing.

No simple rule exists for predicting whether a particular nuclide is radioactive.

However, considering some observations about those nuclides that are stable is helpful in understanding why some nuclides are stable and others are not.

The stability of nuclides, or atomic nuclei, is governed by following main rules:

1. **Average binding energy per nucleon:** Nuclides with higher average binding energies are generally more stable than those with lower average binding energies.
2. **The total number of nucleons:** When the atomic number (Z) is greater than 83, the electrostatic repulsion between protons becomes so immense that almost all isotopes are unstable and hence can undergo radioactive decay..
3. **The N/P ratio:** The ratio of neutrons (N) to protons (P) in a nucleus plays a crucial role in its stability. Nuclei with a N/P ratio close to 1 are generally more stable, while those with significantly more or fewer neutrons tend to be radioactive.
4. **Magic Numbers:** Certain numbers of neutrons and protons are particularly stable. These “magic numbers” are the same for neutrons and protons. All the magic numbers are even. They are: 2, 8,

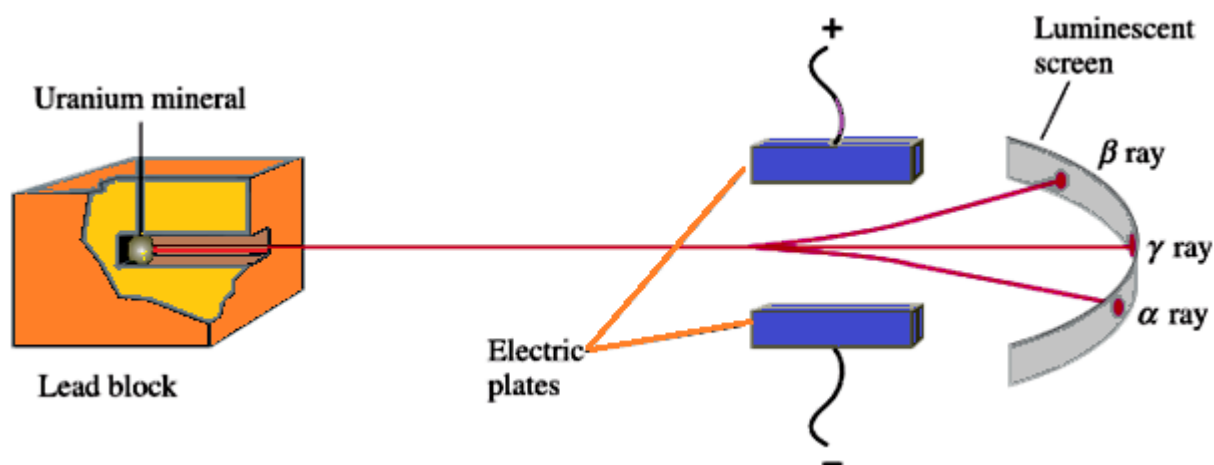
20, 28, 50, 82, 126. On the chart of the nuclides, the magic numbers are bounded with extra heavy lines.

X. Types of Radioactivity:

Some nuclei are unstable; hence they emit sub-atomic particles or electromagnetic radiation in a phenomenon known as radioactivity. Radioactivity can be of two types namely, natural and artificial radioactivity

XI. 1. Natural radioactivity: This is a type of radioactivity that occurs spontaneously, emitting electromagnetic radiation and particles which include three main types α , β , γ , and in a few cases, spontaneous fission.

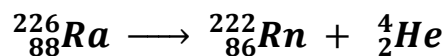
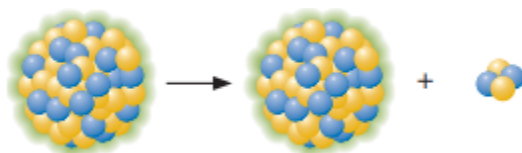
The ability to separate these different types of radiation using electric and magnetic fields was a significant discovery in understanding the nature of radioactivity.



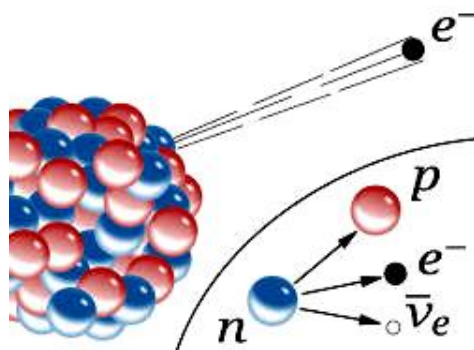
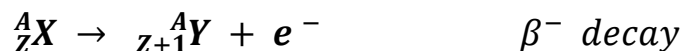
- **Alpha emission (abbreviated α):** emission of a ${}^4_2\text{He}$ nucleus, or alpha particle, from an unstable nucleus.



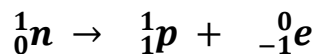
An example is the radioactive decay of radium-226



- **Beta decay:** Beta decay is a common mode of radioactive decay, which includes three types: emission of electron e^- (β^-) or positron e^+ (β^+), and electron capture.

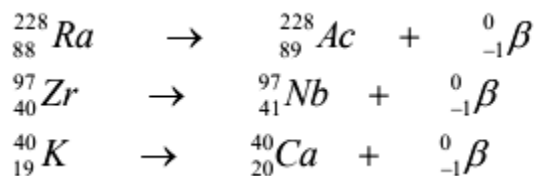


A **beta minus (β^-)** also called negatron is an electron ejected from the nucleus when a neutron is converted into a proton:

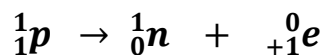


Beta particle β^- emission leads to an increase in the number of proton in the nucleus and a simultaneous decrease in the number of neutrons.

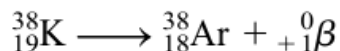
Examples:



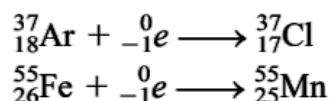
The **positron** (β^+) comes from the transformation of a proton into a neutron following the equation



An example of positron emission is



Electron capture is the capture of an electron—usually a 1s electron—by the nucleus. The captured electron combines with a proton to form a neutron so that the atomic number decreases by one while the mass number remains the same. This process has the same net effect as positron emission:



- **γ decay** : Gamma rays are high energy radiation, emitted when an unstable nucleus undergoes a rearrangement of its constituent particle to give more stable, lower energy nucleus. Often follows α or β decay.

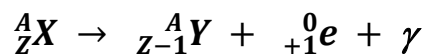
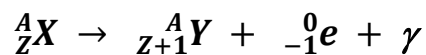
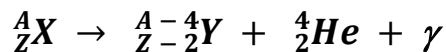
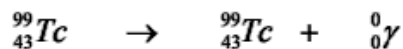
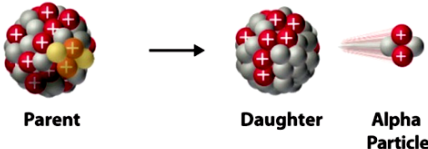
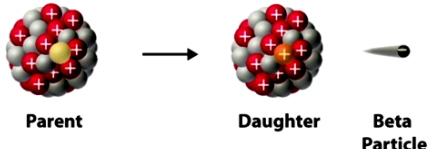
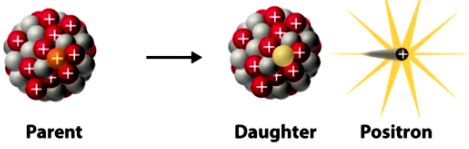
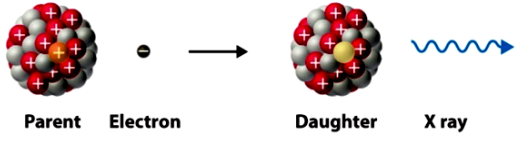
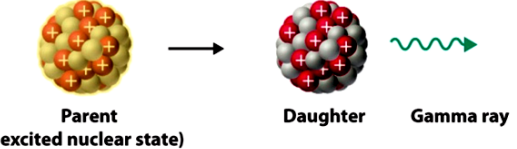
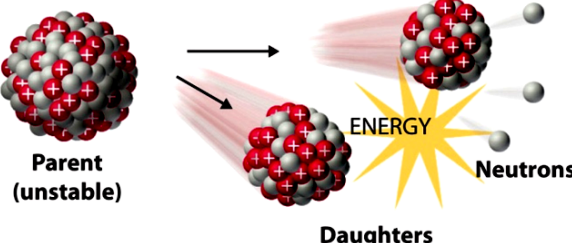
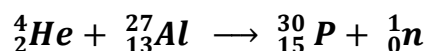


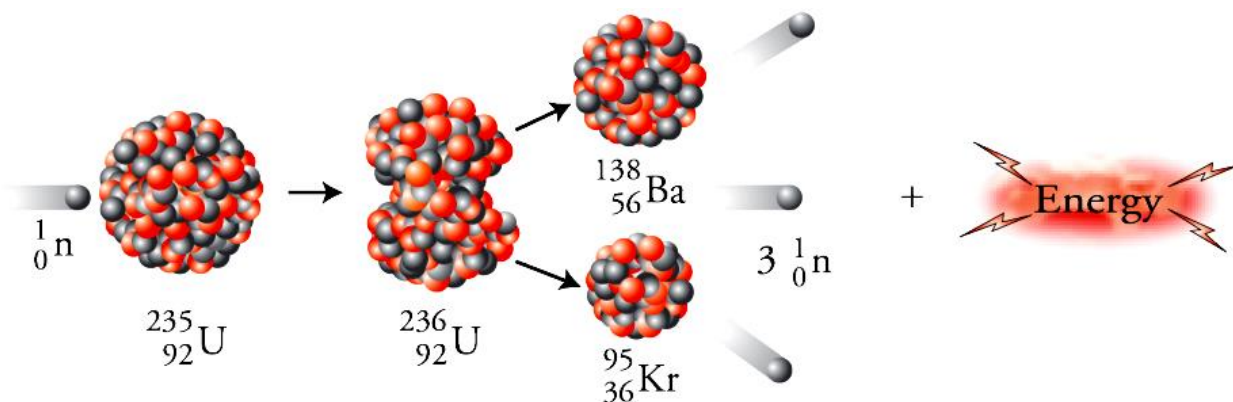
Table 1: Common Modes of Nuclear Decay: The different types of decay are alpha, beta, positron emission, electron capture, gamma emission, and spontaneous fission.

| Decay Type | Radiation Emitted | Generic Equation | Model |
|---------------------|-------------------|---|--|
| Alpha decay | ${}^4_2\alpha$ | ${}^A_ZX \longrightarrow {}^{A-4}_{Z-2}X' + {}^4_2\alpha$ |  <p>Parent → Daughter + Alpha Particle</p> |
| Beta decay | ${}^0_{-1}\beta$ | ${}^A_ZX \longrightarrow {}^A_{Z+1}X' + {}^0_{-1}\beta$ |  <p>Parent → Daughter + Beta Particle</p> |
| Positron emission | ${}^0_{+1}\beta$ | ${}^A_ZX \longrightarrow {}^A_{Z-1}X' + {}^0_{+1}\beta$ |  <p>Parent → Daughter + Positron</p> |
| Electron capture | X rays | ${}^A_ZX + {}^0_{-1}e \longrightarrow {}^A_{Z-1}X' + \text{X ray}$ |  <p>Parent + Electron → Daughter + X ray</p> |
| Gamma emission | ${}^0_0\gamma$ | ${}^A_ZX^* \xrightarrow{\text{Relaxation}} {}^A_ZX' + {}^0_0\gamma$ |  <p>Parent (excited nuclear state) → Daughter + Gamma ray</p> |
| Spontaneous fission | Neutrons | ${}^{A+B+C}_Z X \longrightarrow {}^A_Z X' + {}^B_Y X' + C {}^1_0 n$ |  <p>Parent (unstable) → Daughters + ENERGY + Neutrons</p> |

III. 2. Artificial radioactivity: It is a non-spontaneous form of radioactivity which requires effect of bombardment of the nuclei with sub-atomic particles (proton, neutron or electron). It is otherwise known as anthropogenic or induced radioactivity. Juliot and his wife, Irene Curie discovered artificial radioactivity in 1934 in which aluminum nuclei is bombarded with He nuclei to form new element with emission of electromagnetic ray or particles:

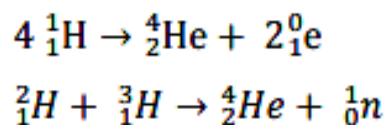


Nuclear Fission: Isotopes of unstable nuclei with atomic number greater than 80 are capable of undergoing a nuclear reaction called nuclear fission, in which they split into nuclei of intermediate masses and emit one or more neutrons. The energy generated is called atomic energy. Some fission reactions are spontaneous while some are not spontaneous; hence, the non-spontaneous require activation energy from bombardment. A given nucleus is split in many ways liberating enormous energy a typical example is shown below



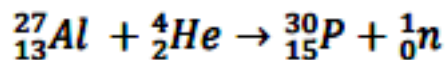
Nuclear fusion

During this type of reaction, two light nuclei will fuse to give a heavier atom and various



Transmutation

These reactions produce nuclides with a mass number equal to or very close to that of the nuclide which served as the target. The nuclides formed are stable or radioactive



XII. Kinetics of radioactive decay

The decomposition of a radioactive element can be very rapid, a few milliseconds, or extremely long, several or thousands of years. If we call the number of radioactive nuclei N , the number of decaying atoms per unit time is dN/dt . It is found that this rate is constant in time and it is proportional to the number of nuclei themselves:

$$\frac{dN}{dt} \propto N$$

$$\frac{dN}{dt} = -\lambda N \implies \frac{dN}{N} = -\lambda dt$$

Therefore: $\ln \frac{N}{N_0} = -\lambda t$

Where N is the number of nuclei

λ : radioactive constant of the element considered. These units: s^{-1} , min^{-1} , h^{-1} , day^{-1}

XIII. Activity of a radioactive nucleus:

Activity is represented by the number of decays that occur per second.

$$A = \lambda N_t \implies A = \lambda N_0 e^{-\lambda t} \implies A = A_0 e^{-\lambda t}$$

$$A_0 = \lambda N_0$$

The units of radioactivity are: – disintegration per second (d.p.s)

Curie (Ci) ; $1 \text{ Ci} = 3,7 \times 10^{10} \text{ dps}$

Radioactive half-life or half-life time: The period or half-life time is the time after which half of the initial nuclei have undergone decay. It is obtained by replacing

$$t = T \text{ and } N_t = N_0 / 2$$

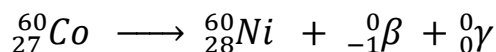
$$\int_{N_0}^N \frac{dN}{N} = -\lambda \int_0^T dt \implies [\ln N]_{N_0}^{\frac{N_0}{2}} = -\lambda [t]_0^T$$

$$\ln \frac{N_0}{2} - \ln N_0 = -\lambda T \implies \ln 2 = \lambda T$$

$$t_{1/2} = T = \frac{\ln 2}{\lambda}$$

Example

A cobalt-60 (^{60}Co) nucleus decays with the emission of beta particle-gamma rays with half life of 5.27 years:



How much of a 3.42 mg sample remain 30.0 years

Solution

$$t_{1/2} = \frac{\ln 2}{\lambda} \implies \lambda = \frac{\ln 2}{t_{1/2}} = \frac{\ln 2}{5.27} = 0.131 \text{ years}^{-1}$$

We have: $A = \lambda N_t$

$$A_0 = \lambda N_0$$

$$\implies A = \lambda N_0 e^{-\lambda t} \implies A = A_0 e^{-\lambda t}$$

$$\ln \frac{A_0}{A_t} = \lambda t = 0.131 (30.0) = 3.93$$

Consider the inverse of two sides

$$\frac{A_0}{A_t} = e^{3.93}$$

$$A_t = \frac{A_0}{e^{3.93}} = \frac{3.41}{51} = 0.067 \text{ mg of } {}^{60}\text{Co}$$

Applications of Radioactivity

Radioactivity has a wide range of applications in science, medicine, and industry. Some examples include:

- **Medicine:** In the realm of medicine, radioactivity empowers a range of diagnostic tools, including X-rays, CT scans, and PET scans, providing physicians with invaluable insights into the human body's inner workings. Beyond diagnosis, radioactivity serves as a potent weapon against cancer, selectively targeting and eliminating malignant cells.
- **Industry:** Industry, too, harnesses the power of radioactivity to enhance various processes. Sterilization techniques employing radioactivity ensure the safety of medical equipment and food products. Nondestructive testing methods utilizing radioactivity allow for the inspection of materials without compromising their integrity. Food irradiation, powered by radioactivity, extends the shelf life of perishable goods, reducing waste and ensuring a stable food supply.
- **Science:** In the pursuit of knowledge, science embraces radioactivity as an indispensable tool. Unraveling the structure and properties of matter at the atomic level is made possible by radioactive isotopes. Moreover, radioactivity provides a means to date archaeological artifacts and geological formations, offering a glimpse into the depths of time.

Safety Considerations

Radioactivity poses risks if mishandled. Employing safety measures, when dealing with radioactive substances is crucial. These measures encompass:

- **Shielding:** Employing shielding aids in safeguarding individuals from radiation exposure.

Various types of shielding are employed to block distinct forms of radiation.

- **Distance:** Maintaining distance from a radioactive source can mitigate radiation exposure.

Greater distance from the source diminishes the level of radiation exposure.

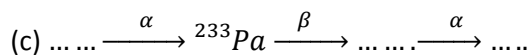
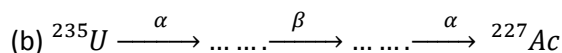
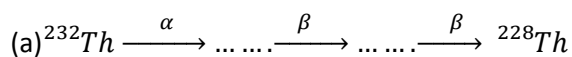
- **Time:** Minimizing the duration of proximity to a radioactive source can curtail radiation exposure.

Series No. 3Exercise 1:

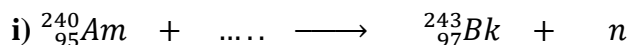
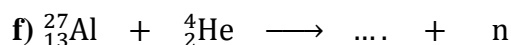
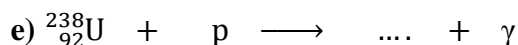
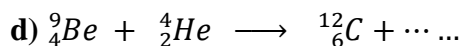
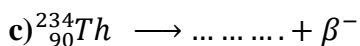
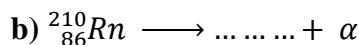
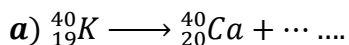
- 1) Calculate the mass defect during the formation of the nuclei of both plutonium ${}_{94}^{239}\text{Pu}$ and barium ${}_{56}^{137}\text{Ba}$.
- 2) Calculate the binding energy and binding energy per nucleon for each.
Data : Ba : 136.8748 ; Pu : 239.0005 ; m_p : 1.00727 ; m_n : 1.00866 (uma),
 m_e : 0.000548 amu

Exercise 2:

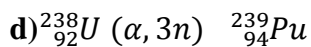
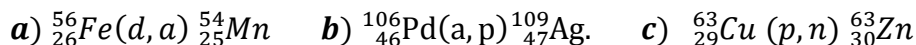
Fill in the blanks in these radioactive decay series:



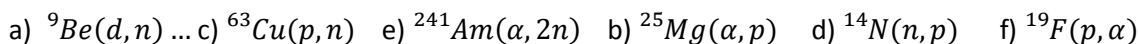
Exercise 3: Complete and Write the abbreviated notation for the following nuclear reaction,



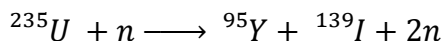
Exercise 4: 1) Write the balanced equation for the nuclear reaction



2) What nuclide is formed in the following reactions?



Exercise 5 : One of the reactions reflecting the fission of uranium-235 is:



What is the energy released (in joule) by fission according to this stoichiometric equation
 1 g of uranium-235?

The masses to be taken into account are as follows:

$$\begin{array}{ll}
 {}^{235}\text{U}: 3.9017 \cdot 10^{-25} \text{ kg} & {}^{95}\text{Y}: 1.5754 \cdot 10^{-25} \text{ kg} \\
 \text{n}: 1.6747 \cdot 10^{-27} \text{ kg} & {}^{139}\text{I}: 2.3058 \cdot 10^{-25} \text{ kg}
 \end{array}$$

Exercise 6: Radiocarbon Dating A piece of wood taken from a cave dwelling in New Mexico is found to have a carbon-14 activity (per gram of carbon) only 0.636 times that of wood cut today. Estimate the age of the wood. The half-life of carbon-14 is 5730 years.

Exercise 7: A sample of uranium ore is found to contain 4.64 mg of ${}^{238}\text{U}$ and 1.22 mg of ${}^{206}\text{Pb}$. Estimate the age of the ore. The half-life of ${}^{238}\text{U}$ is 4.51×10^9 years.

Exercise 8: A sample of the isotope ${}^{131}\text{I}$, which has a half-life of 8.04 days, has an activity of 5.0 mCi at the time of shipment. Upon receipt of the sample at a medical laboratory, the activity is 2.1 mCi. How much time has elapsed between the two measurements?

Exercise 9:

The thyroid gland produces hormones essential to various functions in the body, from dietary iodine.

To check the shape or functioning of this gland, a thyroid scintigraphy is carried out using the isotopes ${}^{131}_{53}\text{I}$ or ${}^{123}_{53}\text{I}$ of iodine. For this scintigraphy, a patient ingests a mass $m = 1.00 \mu\text{g}$ of the isotope ${}^{131}_{53}\text{I}$, or 4.60×10^{15} atoms. The half-life of the isotope ${}^{131}_{53}\text{I}$ is 8.0 days.

- 1- Recall the law of radioactive decay by involving N_0 and the radioactive constant λ .
- 2- Define the half-life time $t_{1/2}$ of a radioactive sample. Deduce the relation $\lambda \times t_{1/2} = \ln 2$.
- 3- Trace, the shape of the curve corresponding to the evolution over time of the number of radioactive nuclei in the sample, justifying the reasoning used. We will correctly place the points corresponding to times $t_{1/2}$, $2t_{1/2}$ and $3t_{1/2}$.

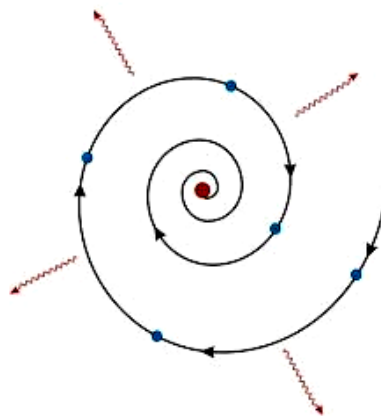
Chapter four
Electronic structure of the
atom

Chapter four: Electronic structure of the atom

Introduction:

Rutherford was unable to form a convincing idea about the location of electrons within the atom; he assumed that electrons orbit around the nucleus at a sufficient speed to generate a central force that balances the nucleus's attractive force. Consequently, the electrons rotate at a fixed distance from the nucleus.

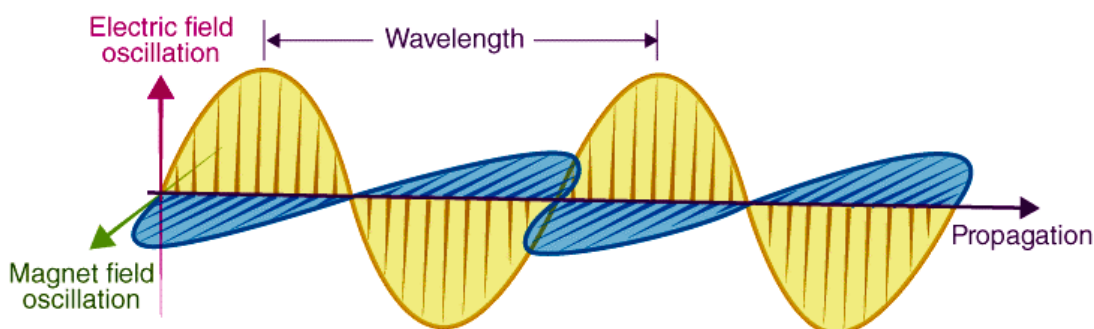
However, according to wave theory, when a charged particle moves under the influence of an attractive force, it loses energy continuously in the form of electromagnetic radiations. Thus, the electron which moves in an attractive field (created by protons present in the nucleus) will emit radiations. As result the electron will lose energy at every turn and move closer and closer to the nucleus following a spiral path and finally fall into the nucleus, thereby making the atom unstable. But the atom is quite stable meaning the electrons do not fall into nucleus, thus this model does not explain the stability of the atom. . Hence, according to Rutherford's conception, the atomic structure collapses.



Electromagnetic Radiations

The nature of light

Visible light is one type of electromagnetic radiation (also called electromagnetic energy or radiant energy). Other familiar types include x-rays, microwaves, and radio waves. All electromagnetic radiation consists of energy propagated by means of electric and magnetic fields that alternately increase and decrease in intensity as they move through space.



Each electromagnetic radiation can be characterized by its **wavelength** and/or **frequency**.

The wavelength (λ) is the distance from one crest to the next in the wave. It is measured in units of distance (**cm, nm, Å**).

The frequency (ν) is the number of complete cycles per second, expressed in **s⁻¹ or Hz**. The frequency ν of a wave is inversely proportional to its wavelength, λ :

$$\nu = \frac{c \text{ (m/s)}}{\lambda \text{ (m)}} \text{ (s}^{-1} \text{ or Hz)}$$

All electromagnetic waves travel with the same velocity, that is, **C = 3 × 10⁸ m/s**.

Wave number: This is the reciprocal of wavelength, i.e. the number of wavelength per centimeter. It is denoted by the symbol $\bar{\nu}$. $\bar{\nu} = \frac{1}{\lambda} \text{ (cm}^{-1} \text{ or m}^{-1} \text{)}$.

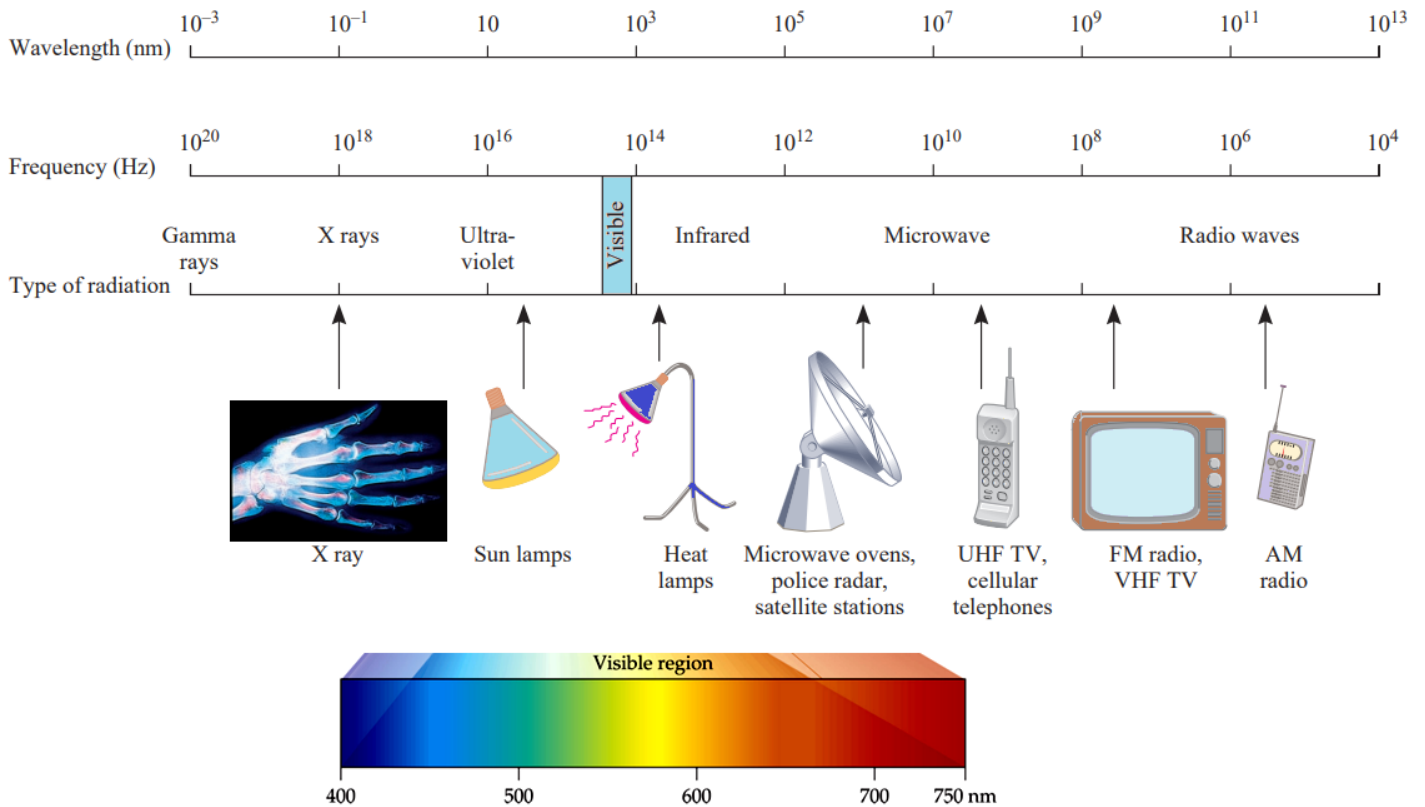
The period (T): In physics, the period of a wave is the amount of time it takes for a wave to complete one wave cycle or wavelength, which is the distance from peak to peak or trough to trough.

Spectrum of electromagnetic radiation:

The electromagnetic spectrum is the distribution of electromagnetic radiation according to energy or, equivalently, according to the wavelength or frequency.

There are various types of electromagnetic waves (radiation) which differs from one another in wavelengths.

Ex: Cosmic Rays, γ -rays, X-rays, U.V, visible, I.R, Micro, Radio.



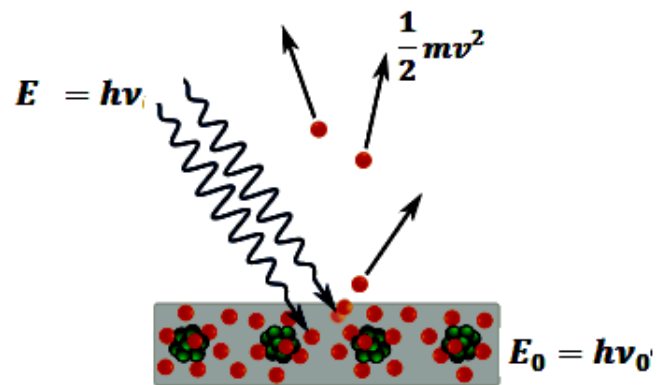
Planck's quantum theory:

Light was made up of tiny packets of energy called photons which carries an amount of energy: $E = h\nu$

Where h: planck's constant $h = 6,62 \times 10^{-34} \text{ j.s}$

Photoelectric effect :

The photoelectric effect is the ejection of electrons from the surface of a metal or from another material when light shines on it (as shown in the Figure 1). Electrons are ejected, however, only when the frequency of light exceeds a certain threshold value characteristic of the particular metal called **frequency threshold or photoelectric threshold**.



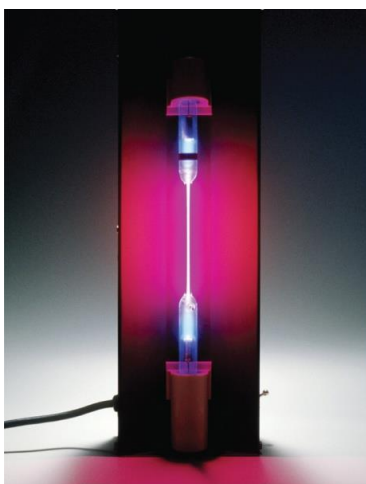
- Only light of frequency $\nu > \nu_0$ cause electron emission.
- If a photon of energy $E = h \cdot \nu \geq E_0 = h \cdot \nu_0$ is absorbed, the emitted electron will reach maximum kinetic energy :

$$E - E_0 = E_c = \frac{1}{2} \cdot m \cdot V^2 = h(\nu - \nu_0)$$

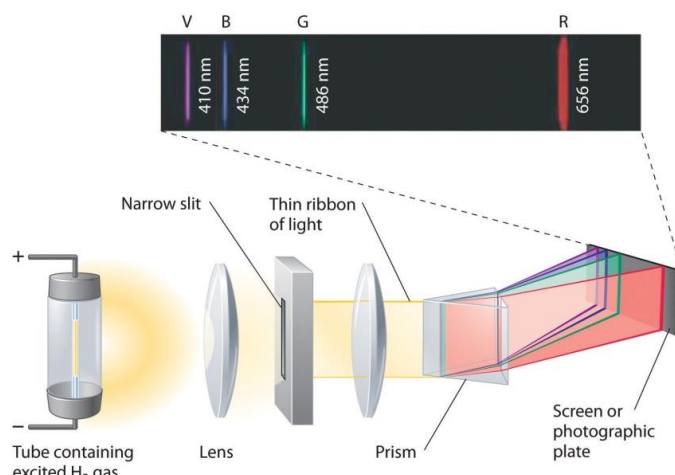
Atomic Spectrum of Hydrogen

Line spectra of the hydrogen atom

By using an electrical discharge inside a tube containing hydrogen gas under low pressure, and when applying a high voltage difference, the hydrogen atoms are excited (activated), causing them to glow to produce light.



(a)



(b)

If this light is analyzed through a prism into its primary compounds, intermittent lines appear on the photographic plate in the regions of the Following spectrum. : **Ultraviolet UV, Visible VIS, Infrared IR.**

When gaseous hydrogen in a glass tube is excited by a 5000-volt electrical discharge, four lines are observed in the visible part of the emission spectrum: **red** at 656.3 nm, **blue-green** at 486.1 nm, **blue violet** at 434.1 nm and **violet** at 410.2 nm:

The emission spectrum of the hydrogen atom and the laws of line distribution:

Groups of lines representing the emission spectrum of a hydrogen atom are called **series**. The **wave number** of each line can be calculated:

$$\bar{\nu} = \frac{1}{\lambda} = R_H \left[\frac{1}{n_1^2} - \frac{1}{n_2^2} \right] \quad n_2 > n_1$$

Where $\bar{\nu}$ wave number.

R_H is the Rydberg constant and its experimental value $R_H = 1.09677 \times 10^5 \text{ cm}^{-1}$

n_1, n_2 : two natural numbers different from zero, n_1 denotes the serie, while n_2 denotes a line within the series.

When a photon is emitted, $n_1 > n_2$. Consequently the term in parentheses is negative and ΔE is negative (energy is lost to the surroundings). When energy is absorbed, $n_1 < n_2$ and the term in parentheses is positive, so ΔE is positive. Each spectral line

Table 7.1 The Various Series in Atomic Hydrogen absorption Spectrum

| Series | n_1 | n_2 | Spectrum Region |
|-----------------|-------|--------------|-------------------------|
| Lyman | 1 | 2, 3, 4, ... | Ultraviolet |
| Balmer | 2 | 3, 4, 5, ... | Visible and ultraviolet |
| Paschen | 3 | 4, 5, 6, ... | Infrared |
| Brackett | 4 | 5, 6, 7, ... | near Infrared |
| Pfund | 5 | 6, 7, | Far Infrared |

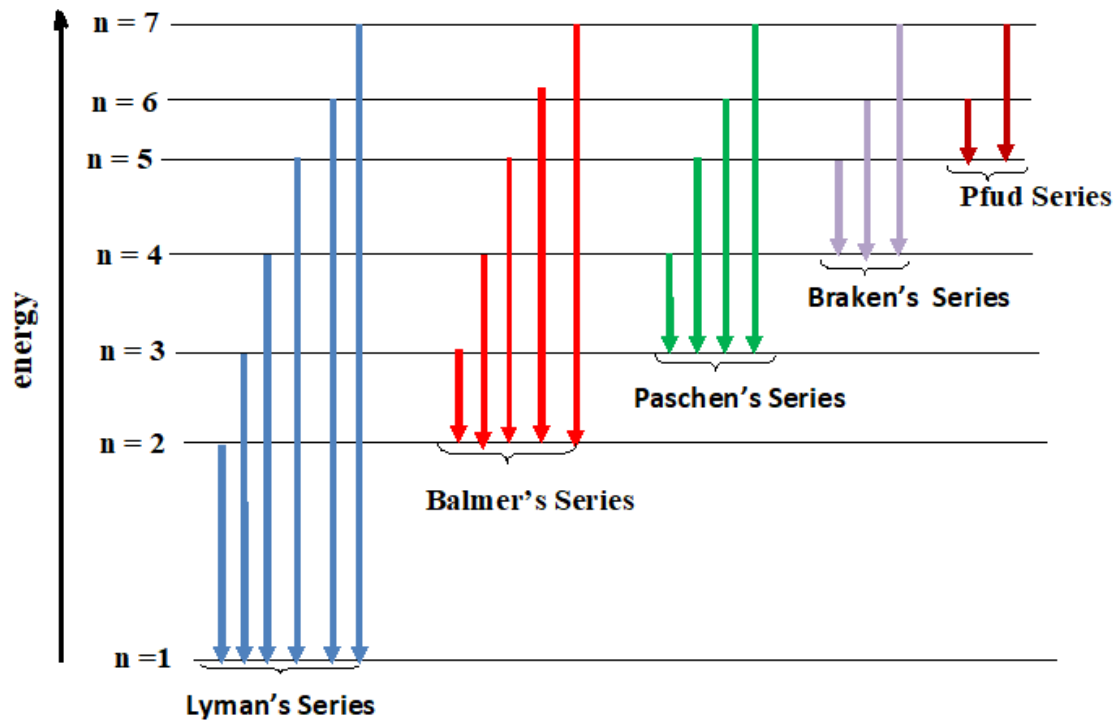


Fig.1 :Energy Level Diagram for the Six Lowest Energy Levels of Hydrogen.

2 Emission spectrum of hydrogen-like atoms ions:

We call hydrogen-like atoms, all the atoms that carry a Z proton in their nucleus and one electron orbiting around them, or they are all the atoms that It loses its electrons and retains only one.

Example: He^+ , ${}_{4}\text{Be}^{3+}$ ${}_{3}\text{Li}^{2+}$

The Ritz-Palmer relationship for hydrogen-like atoms is given as follows

$$\bar{\nu} = \frac{1}{\lambda} = R_H Z^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2} \right] \quad n_2 > n_1$$

Bohr's model (For Hydrogen and H-Like Atoms).

Neils Bohr (1913) was the first to explain quantitatively the general features of Hydrogen atom structure and its spectrum. Bohr's model for Hydrogen atom is based on the following postulates:

1. Electrons in hydrogen atoms can move in circular paths of specific radii and energies around the nucleus. These paths, also known as **orbits, shells, energy levels, or energy states**, are concentrically arranged around the nucleus.
2. The electron in an atom revolves around the nucleus in certain discrete orbits (permissible orbits). Such orbits are known as **stable orbits** or **non – radiating** or **stationary orbits**.
3. Angular momentum of the electron in these stationary orbit is always an integral multiple of $\frac{h}{2\pi}$. i.e : $mvr = n \frac{h}{2\pi}$

Where, m = mass of electron; v= velocity of the electron; r= radius of the orbit; n is an integer ($n = 1, 2, 3$ etc. for the first, second and third orbits respectively). and h= plank's constant

4. Electrons circulate within these orbits under the influence of the **Coulomb force of attraction**, which draws them towards the nucleus, and the **centrifugal force**, which arises from their circular motion. These opposing forces maintain a balance, allowing the electrons to maintain their orbits.
5. The emission or absorption of radiation by the atom takes place when an electron jumps from one stationary orbit to another.

The radiation is emitted or absorbed as a single quantum (photon) whose energy $h\nu$ is equal to the difference in energy ΔE of the electron in the two orbits involved. Thus, $E = h\nu$; Where 'h' =Planck's constant, ν frequency of the radiant energy. Hence the spectrum of the atom will have certain fixed frequency.



Radius of Orbits

Now, the stability of the atom requires that the centrifugal force acting on the revolving electron be balanced by the electrostatic pull exerted by the positively charged nucleus on the electron.

Centrifugal force = pull force (Coulomb's force)

$$\frac{mv^2}{r} = \frac{Ze^2}{4\pi\epsilon_0 r^2} \longrightarrow \textcircled{1}$$

Where: ϵ_0 : permittivity of free space and e : electron charge

Also, according to Bohr's postulates, $mv = n \frac{h}{2\pi} \longrightarrow \textcircled{2}$.

From equation (1) and (2) radius of nth orbit :

$$r_n = \left(\frac{h^2 \epsilon_0}{\pi m e^2} \right) \times \frac{n^2}{Z}$$

From this equation we can see that value of possible radii is discrete then the above equation can be rewrite:

$$r_n = r_1 \times \frac{n^2}{Z}$$

Where r_n is the radius of n^{th} orbit and r_1 is the radius of first orbital

$$r_1 = 0.52 \text{ \AA} = 0.52 \times 10^{-10} \text{ m}$$

1-6-2 Velocity of Electron in the Orbits

In the same way by substituting equation of radius in the Bohr's equation we will find:

$$v_n = \frac{e^2}{2h\epsilon_0} \times \frac{Z}{n} = v_1 \times \left(\frac{Z}{n} \right)$$

Where v_n is the velocity of revolving electron in the n^{th} orbit and v_1 is the velocity of electron in the first orbital

$$v_1 = 2.1818 \times 10^6 \text{ m/s}$$

1-6-3 Electron Energy for Each Orbit

There two type of energy for electron in Hydrogen atom:

1- **Kinetic Energy** ($K.E = \frac{1}{2}mv^2$) since electron is a moving mass.

By substituting v_n in the kinetic energy (K.E) equation we will have

$$K.E = \frac{m e^4}{8h^2 \epsilon_0^2} \times \frac{Z^2}{n^2} \text{ Joules}$$

2- **Potential Energy** (P.E) since electron is a negative charge move in electrical field of positive nucleus.

$P.E = \frac{-Ze^2}{4\pi\epsilon r}$ by substituting $r_n \left(\left(\frac{h^2 \epsilon_0}{\pi m e^2} \right) \times \frac{n^2}{Z} \right)$ in this equation we will obtain

$$P.E = -\frac{m e^4}{4h^2 \epsilon_0^2} \times \frac{Z^2}{n^2} \text{ Joules}$$

Therefore the total energy E will be:

$$E_n = P.E + K.E = \frac{-me^4}{8h^2 \epsilon_0^2} \times \frac{Z^2}{n^2}$$

$$E_n = \frac{-me^4}{8h^2 \epsilon_n^2} \times \frac{Z^2}{n^2} = E_1 \times \frac{Z^2}{n^2}$$

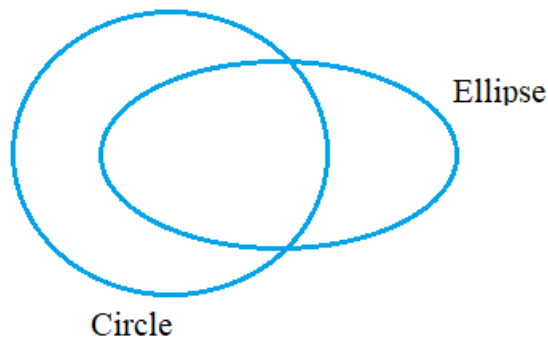
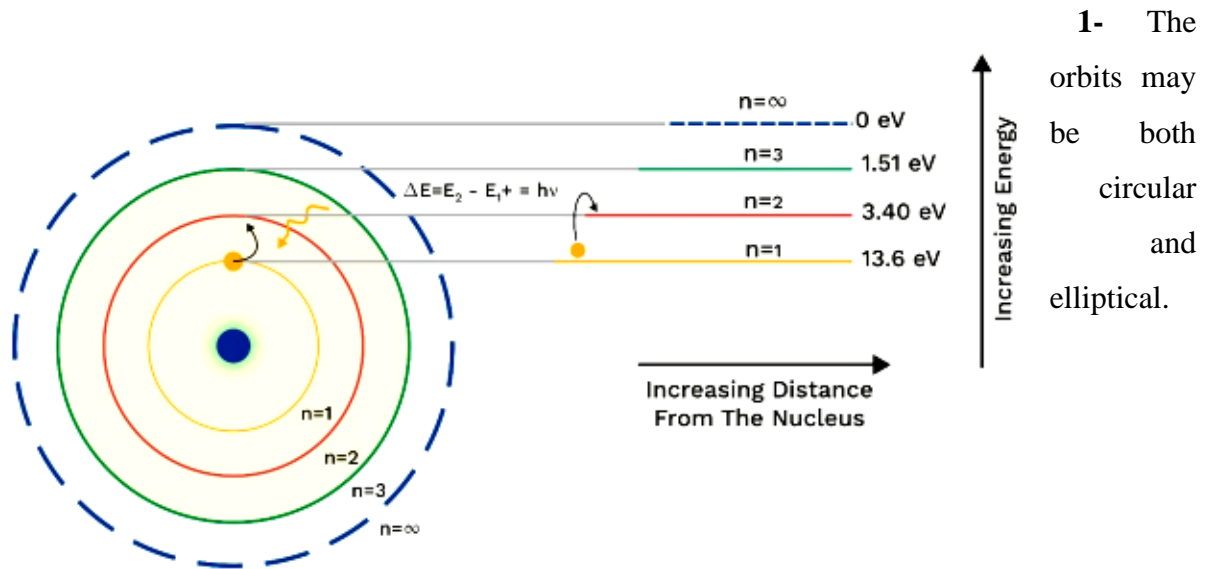
$$E_n = -21.7 \times 10^{-19} \times \frac{Z^2}{n^2} \text{ joule}$$

Where E_n is the total energy of electron in the n_{th} orbital or energy of the n_{th} level Energy can found in term of e.V:

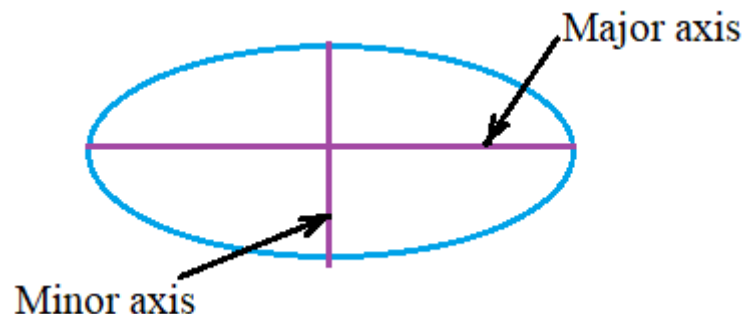
$$E_n = -13.6 \frac{Z^2}{n^2} \text{ eV}$$

Now we can found energy for each orbital

Sommerfeld atomic model This model explains the fine spectrum of Hydrogen atom. The important postulates of Sommerfeld atomic model are:



2- When path is elliptical, then there are two axis – major axis & minor axis. When length of major & minor axis becomes equal then orbit is circular.

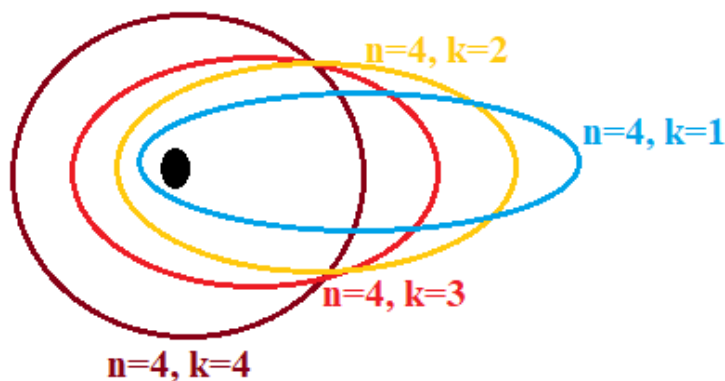


3- The angular momentum of electron moving in an elliptical orbit is $(k \frac{h}{2\pi})$.

Where k is an integer except zero. Value of $k = 1, 2, 3, 4, \dots$

$$(n/k) = \frac{\text{length of major axis}}{\text{length of minor axis}} .$$

With increase in value of k, ellipticity of the orbit decreases. When $n=k$, then orbit is circular.



- 4- Sommerfeld suggested that orbits are made up of sub energy levels. These are s,p,d,f. These sub shells possess slightly different energies.
- Bohr gave a quantum number 'n', which determines the energy of electron.
 - Sommerfeld introduced a new quantum number called Orbital or Azimuthal Quantum number (ℓ) which determines the orbital angular momentum of electron.

Values of $\ell = 0$ to $(n-1)$ ($0 < \ell < n-1$)

For, $n=1$; $\ell=0$; 1s sub shell

$n=2$; $\ell=0,1$; 2s , 2p sub shell

$n=3$; $\ell=0,1,2$; 3s , 3p , 3d sub shell

$n=4$; $\ell=0,1,2,3$; 4s , 4p , 4d , 4f sub shell

THE FOUR QUANTUM NUMBERS The quantum numbers are parameters that describe the distribution of electrons in the atom, and therefore its fundamental nature. They are:

5. **PRINCIPAL QUANTUM NUMBER (n)** - Represents the main energy level, or shell, occupied by an electron. It is always a positive integer, that is $n = 1, 2, 3 \dots$

$n = 1$, is known as **K-shell**.

$n = 2 \Rightarrow$ L –shell

$n = 3 \Rightarrow$ M- shell

$n = 4 \Rightarrow$ N –shell (highest energy)

6. SECONDARY QUANTUM NUMBER (ℓ) -

Represents the energy sublevel, or type of orbital, occupied by the electron. The value of ℓ depends on the value of n such that $0 \leq \ell \leq n - 1$

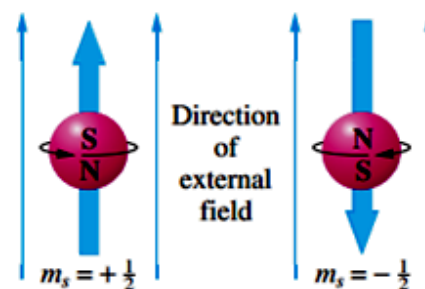
This quantum number characterizes the subshell occupied by the electron

$\ell = 0 \rightarrow$ subshell s

$\ell = 1 \rightarrow$ subshell p

$\ell = 2 \rightarrow$ subshell d

$\ell = 3 \rightarrow$ subshell f



One electron has $m_s = +\frac{1}{2}$;
the other has $m_s = -\frac{1}{2}$.

7. MAGNETIC QUANTUM NUMBER (m) -

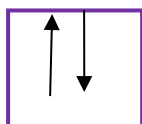
Represents the number of possible orientations in space for each type of orbital. Since the type of orbital is determined by ℓ , the value of m ranges between -1 and $+1$ ($-\ell < m < +\ell$) such that $m = -\ell, \dots, 0, \dots, +\ell$.

This number divides the subshell into individual orbitals which hold the electrons; there are $2\ell+1$ orbitals in each subshell. Thus the s subshell has only one orbital, the p subshell has three orbitals, and so on.

8. SPIN QUANTUM NUMBER (m_s or S)

Represents the two possible orientations that an electron can have in the presence of a magnetic field, or in relation to another electron occupying the same orbital. Only two electrons can occupy the same orbital, and they must have opposite spins.

When this happens, the electrons are said to be paired. The allowed values for the spin quantum number m_s are $+1/2$ and $-1/2$.



Representation of atomic orbitals

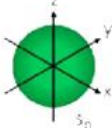
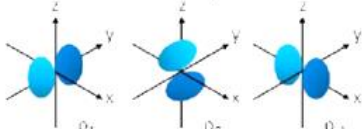
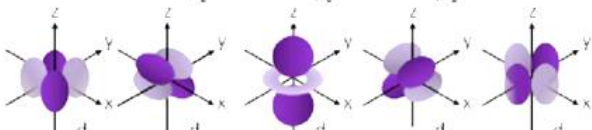
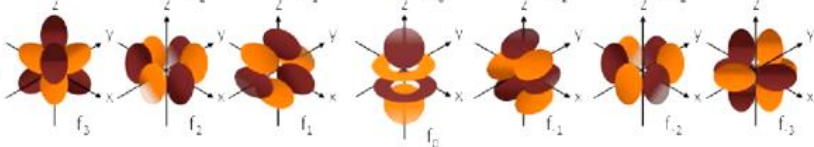
Orbital is a Region of space around nucleus where there is a high probability of finding electron (it represents a discrete energy level)

The number of orbitals and electrons in n principal quantum number is given as

Number of orbitals = n^2

Number of electrons = $2n^2$

Different shapes of orbitals are drawn depending on the possibility of finding electrons around the nucleus. These are s, p, d and f.

| n | l (0 to $n-1$) | | |
|-----|-------------------|-------|--|
| 1 | 0 | $l=0$ | 's' subshell |
| | | |  |
| 2 | 0, 1 | $l=1$ | 'p' subshell |
| | | |  |
| 3 | 0, 1, 2 | $l=2$ | 'd' subshell |
| | | |  |
| 4 | 0, 1, 2, 3 | $l=3$ | 'f' subshell |
| | | |  |

The electronic structure of the atom: It is the distribution that the e^- takes within the atomic subshell. This distribution depends on principles and rules.

4- The principle of stability: In the basic state, the e^- occupies the subshell with the lowest energy level, which gives the atom the lowest energy and maximum stability.

Example: The hydrogen atom electron occupies the 1S station before the 2S station.

5- Aufbau principle

Electrons fill orbitals starting at the lowest available energy state before filling higher states.

The order in which these orbitals are filled is given by the **$n+l$** rule (also known as the **Klechkowski** rule), where orbitals with a **lower $n+l$** value are filled before those with **higher $n+l$** values. The rule is based on the total number of nodes in the atomic orbital, **$n+l$** , which is related to the energy. In the case of **equal $n+l$** values, the orbital with a **lower n value is filled first**.

Example : comparison between E_{4s} and E_{3d}

$$E_{4s}: n + l = 4 + 0 = 4$$

$$E_{3d}: n + l = 3 + 2 = 5$$

$$E_{3d} > E_{4s}$$

comparison between 3p, 3d, 4s

$$n + l = 4 \quad 3p \begin{cases} n = 3 \\ l = 1 \end{cases} \Rightarrow n + l = 4 \quad 4s \begin{cases} n = 4 \\ l = 0 \end{cases}$$

$$3d \begin{cases} n = 3 \\ l = 2 \end{cases} \quad n + l = 5$$

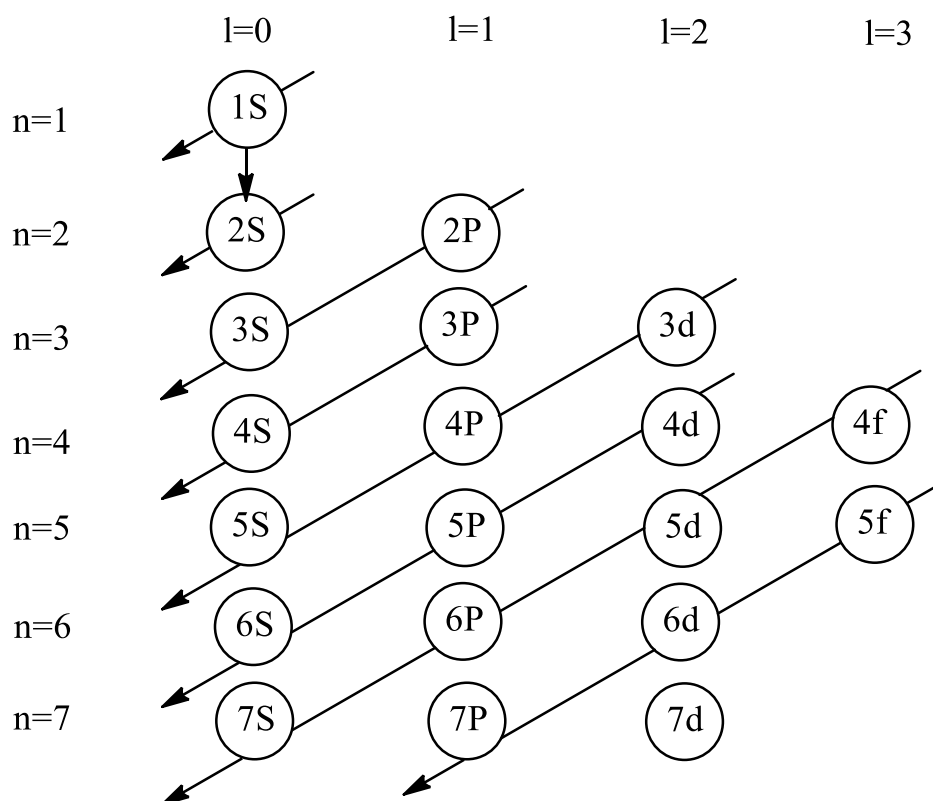
$$E_{3p} < E_{4s} \leq E_{3d}$$

Comparison between 4f, 5d

$$n + l = 7 \quad 3d \begin{cases} n = 5 \\ l = 4 \end{cases}$$

$$\Rightarrow 4f < 5d \quad 3d \begin{cases} n = 4 \\ l = 3 \end{cases} \quad n + l = 7$$

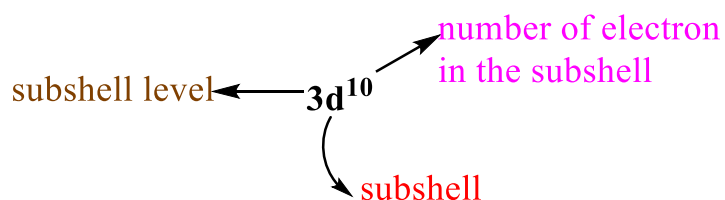
The following figure shows a simplified picture of how to fill the energy sublevels, according to the direction of the arrow



The reading is as follows:

1S 2S 2p 3S 3p 4S 3d 4p 5S 4d 5p 6S 4f 5d 6f 7s 5f 6d
7p

- We write the number of electrons that occupy the orbital as a superscript after the orbital type.



Example: $3d^{10} \rightarrow$ orbital number

6- Pauli's exclusion principle: No two electrons in an atom can have exactly the same set of all the four quantum numbers (n, l, m_l, m_s).

By giving the values of n , ℓ , and m_ℓ , we also specify a particular orbital (e.g., $1s$ with $n = 1$, $\ell = 0$, $m_\ell = 0$). Because m_s has only two possible values ($+\frac{1}{2}$ or $-\frac{1}{2}$),

Example: For an atom ${}_2\text{He}$, the four quantum numbers for the two electrons are as follows: ${}_2\text{He}: 1S^2$

The first electron: $n=1, \ell=0, m=0, s=+\frac{1}{2}$

The second electron: $n=1, \ell=0, m=0, s=-\frac{1}{2}$

If an orbital is occupied by only one e^- , we say that it is a single electron, and it is represented by an arrow inside a box quantum. However, when it is occupied by $2e^-$ in the same quantum box or the same orbital, we say that they form an electronic pair (electron pair).

7- Hund's rule: Hund's rule states that the orbitals of a given subshell are first filled singly, and then the pairing of electrons in each orbital begins. This means that the subshell will have the maximum number of unpaired electrons, resulting in a lower energy state.

Example:

$2p^3$

| | | |
|---|---|---|
| ↑ | ↑ | ↑ |
|---|---|---|

 CORRECT

$2p^3$

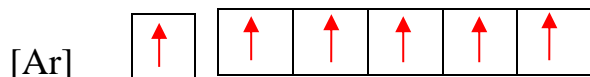
| | | |
|----|---|--|
| ↑↓ | ↑ | |
|----|---|--|

 INCORRECT

Some Anomalous Electron Configurations These rules apply almost all of the time, but you should be aware of some exceptions.

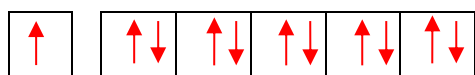
Example: Cr should be $[\text{Ar}] 4s^2 3d^4$ but actually is $[\text{Ar}] 4s^1 3d^5$

Cr : $[\text{Ar}] 4s^1 3d^5$



Also, Cu should be $[\text{Ar}] 4s^2 3d^9$ but is actually $[\text{Ar}] 4s^1 3d^{10}$

$[\text{Ar}] 4s^1 3d^{10}$



The difference arises since there is an enhanced stability associated with **half** or **fully filled** subshells.

| Element | Predicted Electron Configuration | Actual Electron Configuration |
|----------------------|----------------------------------|------------------------------------|
| copper, Cu | $[\text{Ar}] 3d^9 4s^2$ | $[\text{Ar}] 3d^{10} 4s^1$ |
| silver, Ag | $[\text{Kr}] 4d^9 5s^2$ | $[\text{Kr}] 4d^{10} 5s^1$ |
| gold, Au | $[\text{Xe}] 4f^{14} 5d^9 6s^2$ | $[\text{Xe}] 4f^{14} 5d^{10} 6s^1$ |
| palladium, Pd | $[\text{Kr}] 4d^8 5s^2$ | $[\text{Kr}] 4d^{10}$ |

| | | |
|----------------|--------------------------------------|--------------------------------------|
| chromium, Cr | [Ar] 3d ⁴ 4s ² | [Ar] 3d ⁵ 4s ¹ |
| molybdenum, Mo | [Kr] 4d ⁴ 5s ² | [Kr] 4d ⁵ 5s ¹ |

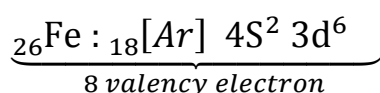
Abbreviated electron configuration

We can abbreviate the electron configuration of any element by putting brackets around the symbol of the noble gas that occurs just before it on the Periodic Table and continuing with the valence shell from there. Some simple examples:

Li as 1s² 2s¹ becomes [He] 2s¹

Mg as 1s² 2s² 2p⁶ 3s² becomes [Ne] 3s²

3- **Valence electrons:** They are the electrons that occupy the subshell that come after the inert gas in the electronic configuration.



Series No. 4

Exercise 1: A monochromatic light beam has a frequency of 5.1×10^{14} Hz.

- Calculate: the wavelength (λ), the wave number ($\bar{\nu}$), the period (T) and the photon energy carried by this optical radiation.
- This radiation hits a zinc metal plate (Zn) where: $E_0(\text{Zn}) = 3.4$ eV
Does photoelectric effect occur?
- This radiation hits a cesium metal plate (Cs) where: $E_0(\text{Cs}) = 1.9$ eV

Does photoelectric effect occur?

If so, calculate the kinetic energy of the released electron, and then calculate the speed of the electrons released from the cesium metal.

Given: $h = 6.62 \times 10^{-34}$ J.s , $C = 3 \times 10^8$ m/s , $m_e = 9.1 \times 10^{-31}$ Kg

Exercise 2: Given for a spectral series for an ion similar to a hydrogen atom:

- The wavelength corresponding to the first line is 3.3×10^7 m⁻¹
- The wavelength corresponding to the limit line is 4.4×10^7 m⁻¹

1. Identify this series.
2. Calculate the atomic number Z for this ion.

Exercise 3:

1. Calculate the wavelength corresponding to the first line and the limit line of the Lyman, Balmer, Paschen, and Brackett series.

2. Calculate the ionization energy of a hydrogen atom in its first excited state.

3. Calculate the radius (r_n), energy (E_n), and speed (V_n) for $n = 1, 2, 3, 4$ according to the Bohr model.

4. Sketch an energy level diagram showing the transitions corresponding to the first line of the Lyman series and the second line of the Paschen series.

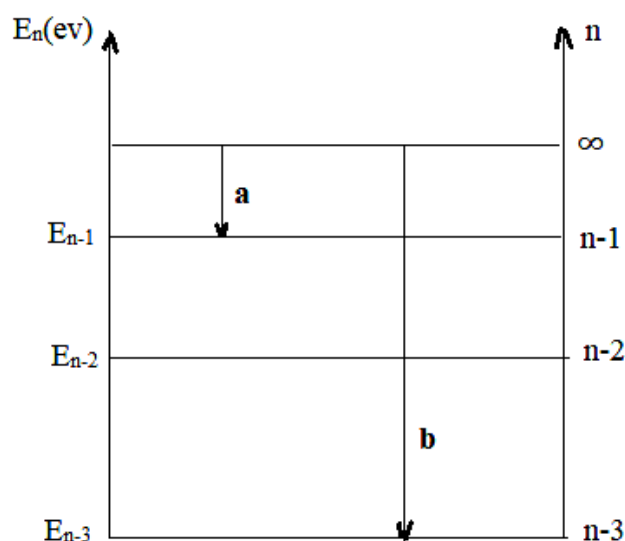
5. Let the transitions b and a shown on the diagram below: If $\frac{\Delta E(b)}{\Delta E(a)} = \frac{25}{9}$, where ΔE is the difference in energy.

- Identify the number n shown on the energy diagram

- Name each spectral series corresponding to each transition

Exercise 4: The energy diagram for different energy levels of an atom is as follows:

Where: a, b, c, d, and e represent the different transitions of the electron of this atom.

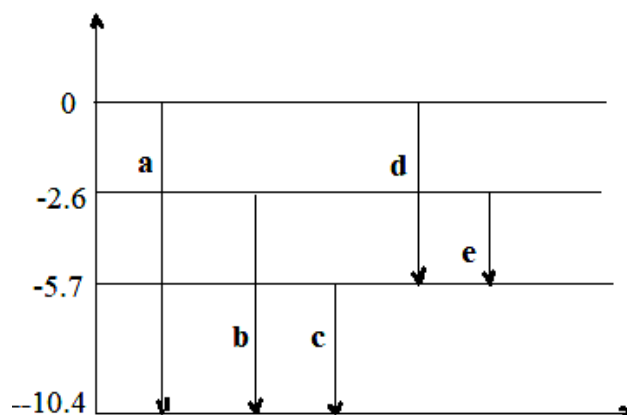


1. Calculate the frequencies ν_c , ν_d and ν_e , corresponding to the transitions c, d, and e.

Then, deduce the frequencies of the transitions ν_a , ν_b .

2. An electron collides with this atom in its first excited state.

- Calculate the kinetic energy that causes ionization of the atom.



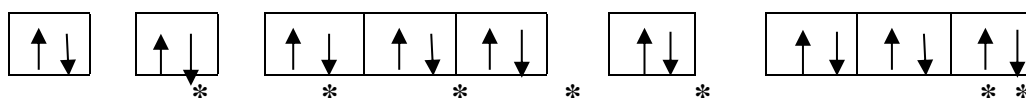
Exercise 5: 1) Compare, for a multi-electron atom, the energy levels of the following orbitals:

A- 4s, 3d B- 4s, 4p C- 6s, 5p,

4d

2) What is the maximum number of electrons that can be occupied by the following orbitals: 3p, 4s, 3d, 2p, 4f?

3) The electronic configuration of an atom can be represented as:



What are the four quantum numbers (n , l , m , s) for the indicated electrons?

Exercise 6: Give the electronic distribution and configuration of the element rubidium ${}^{85}_{37}\text{Rb}$

What are the four quantum numbers for the outer electron of this atom?

Deduce the electronic structure of the ${}^{85}_{37}\text{Rb}^+$ ion

Exercise 7: Let the following elements be ${}_{11}\text{Na}$; ${}_{49}\text{In}$; ${}_{52}\text{Te}$:

Give the electron representation for each element

Locate each element on the periodic table

Arrange these elements in order of increasing ionization energy

Arrange these elements in order of increasing electronegativity.

Correction

Exercise 1:

1. Wavelength

$$\lambda = \frac{c}{\nu} = (3 \times 10^8) / (5.1 \times 10^{14}) = 5.9 \times 10^{-7} \text{ m} = 590 \text{ nm}$$

Wave number

$$\bar{\nu} = \frac{1}{\lambda} = \left(\frac{1}{5.9 \times 10^{-7}} \right) = 1.69 \times 10^7 \text{ m}^{-1}$$

Period

$$T = \frac{1}{\nu} = \frac{1}{5.1 \times 10^{14}} = 1.96 \times 10^{-15} \text{ s}$$

Photon energy

$$E = h\nu = (6.62 \times 10^{-34}) \times (5.1 \times 10^{14}) = 3.38 \times 10^{-19} \text{ J} = 2.17 \text{ eV}$$

2. Photoelectric effect

In the zinc metal, the **threshold energy** (work function) is $E_0(\text{Zn}) = 3.4 \text{ eV}$.

If the energy of the photon is greater than or equal to the work function, then the photoelectric effect will occur.

In this case, the energy of the photon is $E_{ph} = 2.17 \text{ eV}$, which is smaller than the threshold energy ($E_0(\text{Zn})$). Therefore, the photoelectric effect does not occur.

3. Photoelectric effect

In the cesium metal, the work function is $E_0(\text{Cs}) = 1.9 \text{ eV}$.

If the energy of the photon is greater than or equal to the work function, then the photoelectric effect will occur.

In this case, the energy of the photon is $E_{ph} = 2.17 \text{ eV}$, which is greater than the work function $E_0(\text{Cs})$.

Therefore, the photoelectric effect will occur.

The kinetic energy of the released electron is given by the following equation:

$$KE = E_C = E - E_0$$

where

- $KE = E_C$ is the kinetic energy of the released electron
- E is the energy of the photon
- E_0 is the work function

In this case,

$$KE = 2.17 \text{ eV} - 1.9 \text{ eV} = 0.27 \text{ eV}$$

Speed of the released electrons

The speed of the released electrons is given by the following equation:

$$\text{We have } KE = E_C = \frac{1}{2} m_e v^2 \Rightarrow v = \sqrt{\frac{2 \times E_C}{m_e}}$$

where

- v is the speed of the released electrons
- m_e is the mass of the electron

In this case,

$$v = \left(\frac{2 \times 0.27 \times 1.6 \times 10^{-19} \text{ Joule}}{9.1 \times 10^{-31} \text{ kg}} \right)^{1/2} = 3.08 \times 10^5 \text{ m/s}$$

Solution exercise 2:

The first line of a spectral series for a hydrogen-like ion is given by the following equation:

$$\bar{\nu} = \frac{1}{\lambda} = R_H \times \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

where:

- $\bar{\nu}$ wave number
- λ is the wavelength of the line
- R_H is the Rydberg constant
- n_1 is the principal quantum number of the initial state
- n_2 is the principal quantum number of the final state

The wave number corresponding to the first line is $3.3 \times 10^7 \text{ m}^{-1}$

$$\bar{\nu}_1 = R_H Z^2 \left(\frac{1}{n_1^2} - \frac{1}{(n_1+1)^2} \right) = 3.3 \cdot 10^7 \quad \textcircled{1}$$

The wave number corresponding to the boundary line $4.4 \times 10^7 \text{ m}^{-1}$ is

$$\nu_L = R_H Z^2 \left(\frac{1}{n_1^2} - \frac{1}{\infty} \right) = R_H Z^2 \frac{1}{n_1^2} = 4.4 \cdot 10^7 \quad \textcircled{2}$$

by dividing 2 by 1 and then simplifying, we find

$$3n_1^2 - 2n_1 - 1 = 0$$

It is a quadratic equation

The equation has two solutions

$$\begin{cases} n_1 = 1 \\ n_1 = -\frac{1}{3} \text{ rejected solution} \end{cases}$$

Hence, it belongs to the **Lyman series**.

- Calculating the atomic number Z for this ion

$$- Z = \sqrt{\frac{\nu_L n^2}{R_H}} = \sqrt{\frac{4.4 \cdot 10^7}{1.1 \cdot 10^7}} = 2$$

Exercise 5:

The arrangement of orbitals on the basis of energy is based upon their **(n+l)** value. **Lower** the value of **(n+l)**, the **lower** is the **energy**. For orbitals having the **same** values of **(n+l)**, the orbital with a **lower** value of **n** will have **lower energy**:

A. 4s vs. 3d:

$$4s \rightarrow \begin{cases} n = 4 \\ l = 0 \end{cases} \Rightarrow n + l = 4$$

$$3d \rightarrow \begin{cases} n = 3 \\ l = 2 \end{cases} \Rightarrow n + l = 5$$

Thus: $4s < 3d$

B. 4s vs. 4p:

$$4s \rightarrow \begin{cases} n = 4 \\ l = 0 \end{cases} \Rightarrow n + l = 4$$

$$4p \begin{cases} n = 4 \\ l = 1 \end{cases} \Rightarrow n + l = 5$$

$4s < 4p$

C. 6s, 5p, and 4d:

$$6s \rightarrow \begin{cases} n = 6 \\ l = 0 \end{cases} \Rightarrow n + l = 6$$

$$5p \begin{cases} n = 5 \\ l = 1 \end{cases} \Rightarrow n + l = 6$$
$$4d \rightarrow \begin{cases} n = 4 \\ l = 2 \end{cases} \Rightarrow n + l = 6$$

Each orbital has a different principal quantum number:

$n = 6$ for 6s

$n = 5$ for 5p

$n = 4$ for 4d

The energy of an orbital decrease with increasing n value.

Therefore, the energy levels are ordered as follows: $6s < 5p < 4d$.

Chapter five

Periodic Classification Of Elements

Introduction

In 1804, physicist John Dalton proposed the atomic theory of matter, aiding scientists in determining the masses of known elements. Concurrently, chemists Sir Humphry Davy and Michael Faraday developed electrochemistry, facilitating the discovery of new elements. By 1829, chemist Johann Wolfgang Doberiner observed that certain elements with similar properties occurred in groups of three, such as chlorine, bromine, and iodine; calcium, strontium, and barium; sulfur, selenium, and tellurium; iron, cobalt, and manganese. However, due to the limited number of known elements at the time and confusion between molecular and atomic weights, chemists did not fully grasp the significance of Doberiner's triads.

In 1859, physicists Robert Wilhelm Bunsen and Gustav Robert Kirchoff discovered spectroscopy, enabling the identification of numerous new elements. This breakthrough provided scientists with the means to elucidate relationships between elements. Consequently, in 1864, chemist John A. R. Newlands arranged elements in increasing order of atomic weights, noting that a given set of properties recurred every eighth place. He termed this pattern the "law of octaves."

Mendeleev's Periodic Table

Mendeleev's periodic table is a systematic arrangement of elements based on their properties, grouping together those with similar characteristics. Notably, Mendeleev introduced the concept of leaving gaps or blank spaces for yet-to-be-discovered elements, anticipating their existence based on the periodicity of known elements. These placeholders, corresponding to specific atomic masses (e.g., 44 for scandium, 68 for gallium, 72 for germanium, and 100 for technetium), allowed for the logical grouping of elements with similar properties.

Table 6.3: Mendeleev's updated periodic table

| Groups | I | | II | | III | | IV | | V | | VI | | VII | | VIII | | | | | |
|--------------------|--------------|---|-----------------------|---|--|---|------------------------------------|---|--|---|------------------------------------|---|-------------------------------------|---|-------------------|--|--|--------------|--|--------------|
| Oxides Hydrides | RO RH | | RO RH ₂ | | R ₂ O ₃ RH ₃ | | RO ₂ RH ₄ | | R ₂ O ₅ RH ₃ | | RO ₃ RH ₂ | | R ₂ O ₇ RH | | RO ₄ | | | | | |
| Periods ↓ | A | B | A | B | A | B | A | B | A | B | A | B | A | B | Transition series | | | | | |
| 1 | H 1.008 | | | | | | | | | | | | | | | | | | | |
| 2 | Li 6.939 | | Be 9.012 | | B 10.81 | | C 12.011 | | N 14.007 | | O 15.999 | | F 18.998 | | | | | | | |
| 3 | Na 22.99 | | Mg 24.31 | | Al 29.98 | | Si 28.09 | | P 30.974 | | S 32.06 | | Cl 35.453 | | | | | | | |
| 4 | K 39.102 | | Ca 40.08 | | Sc 44.96 | | Ti 47.90 | | V 50.94 | | Cr 50.20 | | Mn 54.94 | | Fe 55.85 | | | Co 58.93 | | Ni 58.71 |
| | Cu 63.54 | | Zn 65.37 | | Ga 69.72 | | Ge 72.59 | | As 74.92 | | Se 78.96 | | Br 79.909 | | | | | | | |
| 5 | Rb 85.47 | | Sr 87.62 | | Y 88.91 | | Zr 91.22 | | Nb 92.91 | | Mo 95.94 | | Tc 99 | | Ru 101.07 | | | Rh 102.91 | | Pd 106.4 |
| | Ag 107.87 | | Cd 112.40 | | In 114.82 | | Sn 118.69 | | Sb 121.75 | | Te 127.60 | | I 126.90 | | | | | | | |
| 6 | Cs 132.90 | | Ba 137.34 | | La 138.91 | | Hf 178.49 | | Ta 180.95 | | W 183.85 | | | | Os 190.2 | | | Ir 192.2 | | Pt 195.09 |
| | Au 196.97 | | Hg 200.59 | | Tl 204.37 | | Pb 207.19 | | Bi 208.98 | | | | | | | | | | | |

MODERN PERIODIC TABLE

The Modern Periodic Table, based on the modern periodic law, is an extended version of Mendeleev's table, known as the Long Form of the Periodic Table. It distinguishes subgroups A and B, which were combined in Mendeleev's original table.

The table comprises 18 vertical columns, each referred to as a group, numbered from 1 to 18 in Arabic numerals. Elements within the same group share similar electronic configurations and possess the same number of valence electrons.

There are seven horizontal rows, or periods, in the periodic table. Each period consists of elements with consecutive atomic numbers and is numbered from 1 to 7 in Arabic numerals.

The first period is the shortest, containing only two elements: hydrogen (H) and helium (He).

The second and third periods are considered short periods, each containing 8 elements.

The fourth and fifth periods are classified as long periods, each containing 18 elements.

The sixth and seventh periods are categorized as very long periods, each containing 32 elements.

Description of the Mendeleev Periodic Table

The periodic table is a consequence of electronic configurations. The periodic classification is based on the formation of groups made up of elements (with atomic number Z) with analogous properties.

The periodic table is made up of 4 blocks: s, p, d, and f.

The elements in the same horizontal row of the periodic table constitute a period. There are 7 of them.

The elements in the same column with the same outer shell electronic configuration constitute a family or group. The periodic table is made up of 18 columns divided into 8 groups.

- Subgroup A: contains elements whose outer shell is $ns\ np$.
- Subgroup B: contains atoms that have a d state.

The indices I, II, III, ... indicate the number of electrons in the outer shell, called valence electrons.

The main families of the periodic table are:

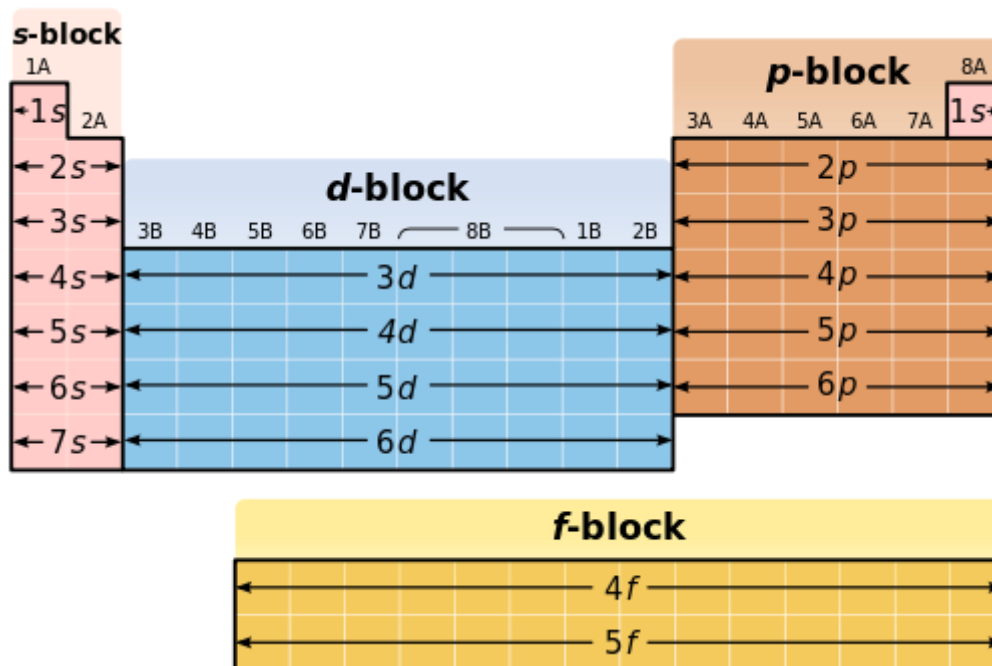
- **The s-block** of the periodic table consists of the combined columns of Group IA and Group IIA.

Group I_A, also known as alkali metals, consists of elements with an electron configuration of ns^1 . These alkali metals readily undergo ionization to form M^+ . Notably, hydrogen, despite its unique properties, does not belong to the alkali metal group or the general metals category.

Group II_A, referred to as alkaline earth metals, comprises elements with an electron configuration of ns^2 . These alkaline earth metals easily undergo ionization to form M^{2+} .

- **The p block** is made up of columns III_A, IV_A, V_A, VI_A, VII_A, and VIII_A, characterized by the filling of the following respective outer subshells: $ns^2 / np^1\ np^2\ np^3\ np^4\ np^5\ np^6$ and whose names are:

- trivalent metals for III_A
 - metalloids from IV_A to VI_A
 - halogens that easily ionize to M⁻: VII_A
 - noble gases, chemically inert VIII_A.
- **The d block** is made up of 10 columns characterized by the filling of the d subshell. Called transition elements, they are all metals and are good conductors of electricity. They are characterized mainly by the fact that they generally have several ionization degrees and incomplete d orbitals. These are the families I_B → VIII_B.
- **The f block** contains the lanthanides and actinides. It is the filling of the subshells $ns^2 (n-1)d^1 (n-2)f^1 \rightarrow 14$ Note: For the f block, we will accept the following filling rule: $ns^2 (n-1)d^1 (n-2)f^1 \rightarrow 14$



Classification of the Elements

The elements can be classified in several ways. The two most common classification systems are

1. A system based on selected physical properties of the elements, in which they are described as metals or nonmetals.
2. A system based on the electron configurations of the elements, in which elements are described as *noble-gas*, *representative*, *transition*, or *inner transition elements*.

The classification scheme based on electron configurations of the elements is depicted in the Figure below.

A **noble-gas element** is an element located in the far right column of the periodic table. These elements are all gases at room temperature, and they have little tendency to form chemical compounds. With one exception, the distinguishing electron for a noble gas completes the p subshell; therefore, noble gases have electron configurations ending in p^6 . The exception is helium, in which the distinguishing electron completes the first shell—a shell that has only two electrons. Helium's electron configuration is $1s^2$.

A **representative element** is an element located in the s area or the first five columns of the p area of the periodic table. The distinguishing electron in these elements partially or completely fills an s subshell or partially fills a p subshell. The representative elements include most of the more common elements.

A **transition element** is an element located in the d area of the periodic table. Each has its distinguishing electron in a d subshell.

An **inner transition element** is an element located in the f area of the periodic table.

Each has its distinguishing electron in an f subshell. There is very little variance in the properties of either the $4f$ or the $5f$ series of inner transition elements.

| Representative elements | | | | | | | | | | Noble-gas elements | | | | | | | |
|---------------------------|----------|---------------------|-----------|-----------|-----------|-----------|-----------|-----------|-----------|--------------------|----------|-----------|-----------|-----------|-----------|----------|----------|
| 1 H | | | | | | | | | 2 He | | | | | | | | |
| 3 Li | 4 Be | | | | | | | | | 5 B | 6 C | 7 N | 8 O | 9 F | 10 Ne | | |
| 11 Na | 12 Mg | Transition elements | | | | | | | | 13 Al | 14 Si | 15 P | 16 S | 17 Cl | 18 Ar | | |
| 19 K | 20 Ca | 21 Sc | 22 Ti | 23 V | 24 Cr | 25 Mn | 26 Fe | 27 Co | 28 Ni | 29 Cu | 30 Zn | 31 Ga | 32 Ge | 33 As | 34 Se | 35 Br | 36 Kr |
| 37 Rb | 38 Sr | 39 Y | 40 Zr | 41 Nb | 42 Mo | 43 Tc | 44 Ru | 45 Rh | 46 Pd | 47 Ag | 48 Cd | 49 In | 50 Sn | 51 Sb | 52 Te | 53 I | 54 Xe |
| 55 Cs | 56 Ba | 57 La | 72 Hf | 73 Ta | 74 W | 75 Re | 76 Os | 77 Ir | 78 Pt | 79 Au | 80 Hg | 81 Tl | 82 Pb | 83 Bi | 84 Po | 85 At | 86 Rn |
| 87 Fr | 88 Ra | 89 Ac | 104 Rf | 105 Db | 106 Sg | 107 Bh | 108 Hs | 109 Mt | 110 Ds | 111 Rg | 112 — | 113 — | 114 — | 115 — | | | |
| Inner transition elements | | | | | | | | | | | | | | | | | |
| | | 58 Ce | 59 Pr | 60 Nd | 61 Pm | 62 Sm | 63 Eu | 64 Gd | 65 Tb | 66 Dy | 67 Ho | 68 Er | 69 Tm | 70 Yb | 71 Lu | | |
| | | 90 Th | 91 Pa | 92 U | 93 Np | 94 Pu | 95 Am | 96 Cm | 97 Bk | 98 Cf | 99 Es | 100 Fm | 101 Md | 102 No | 103 Lr | | |

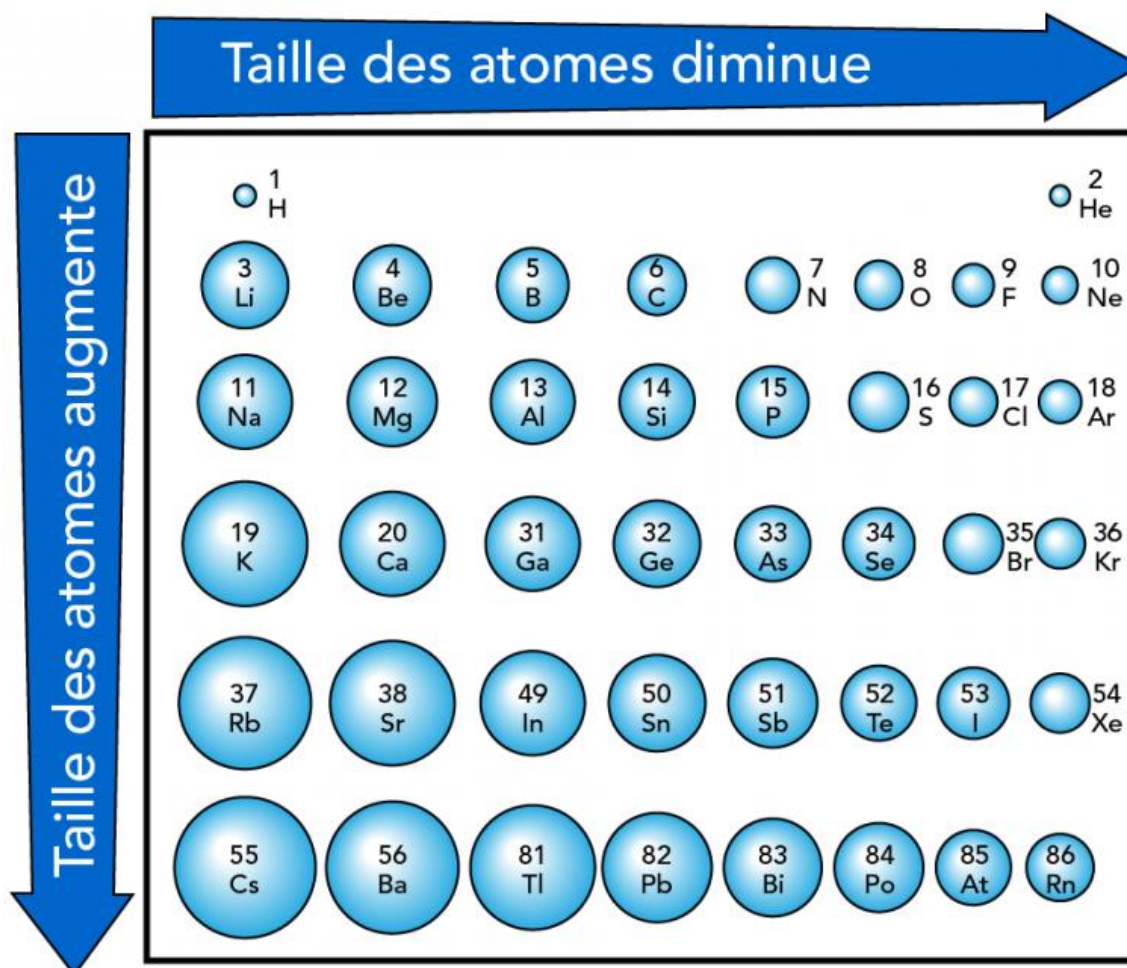
Periodicity of properties

B. Atomic Size

Atomic size, also known as atomic radius, refers to the distance between the nucleus's center and the outermost shell of an isolated atom. This measurement is crucial in understanding various properties of atoms and is typically expressed in picometers (pm), where 1 pm equals 10^{-12} meters. The atomic size plays a significant role in determining numerous other atomic properties.

Variation of atomic size in periodic table.

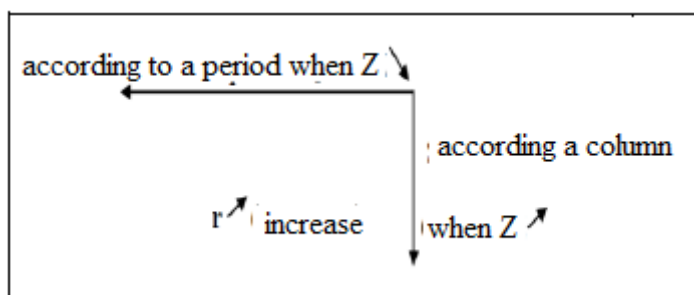
The size of atoms decreases from left to right **in a period** but increases from top to bottom **in a group**.



In a period the atomic number and therefore the positive charge on the nucleus increases gradually. As a result, the electrons are attracted more strongly and they

come closer to the nucleus. This decreases the atomic size in a period from left to right.

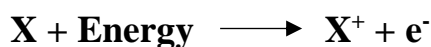
In a **group** as one goes down, a new shell is added to the atom which is farther away from the nucleus. Hence electrons move away from the nucleus. This increases the atomic size in a group from top to bottom.



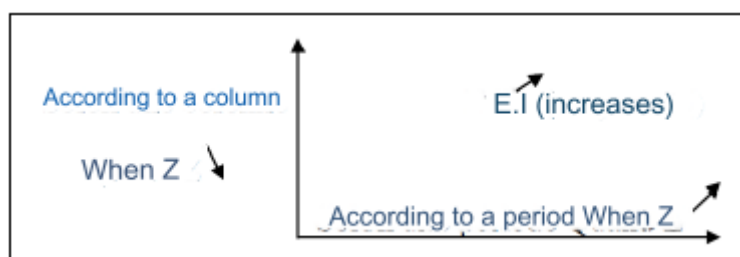
B. Ionization energy (IE)

Ionization energy, denoted (IE), is the energy required to remove an electron from an atom.

The equation for ionization energy is:



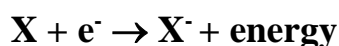
In the periodic table the ionization energy decreases from top to bottom and increases from left to right.



C. Electron affinity (EA)

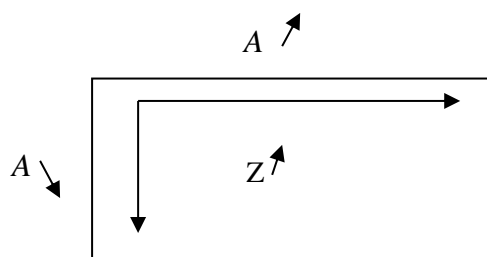
Electron affinity, denoted EA, is the energy released when an atom captures an electron.

The equation for electron affinity is:



The electron affinity of the elements generally increases as you move from left to right across the periodic table. This is because the number of protons in the nucleus increases as you move from left to right.

The electron affinity of the elements also generally decreases as you move down a group in the periodic table. This is because the electrons are further from the nucleus as you move down a group.

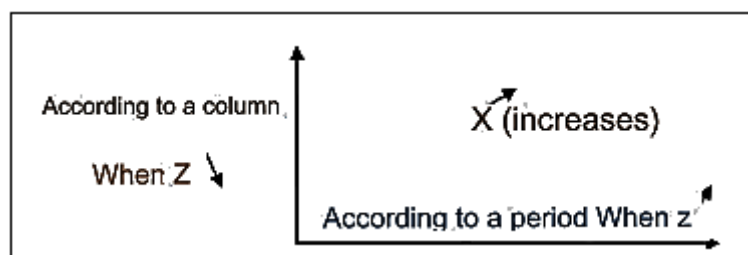


D. Electronegativity

Electronegativity is the ability or tendency of an atom to attract electrons when forming a chemical bond with another element.

Electronegativity generally increases as you move from left to right across the periodic table. This is because the number of protons in the nucleus increases as you move from left to right.

Electronegativity also generally decreases as you move down a group in the periodic table. This is because the electrons are further from the nucleus as you move down a group.



The electronegativity EN of an element X can be defined according to several scales:

a- Mulliken scale:

The electronegativity of an element on the Mulliken scale is equal to the arithmetic mean of the first ionization energy, E_{I1} , and the electron affinity, AE.

$$EN = (EI_1 + AE) / 2$$

b- Pauling scale:

This scale is based on the bond dissociation energies of simple diatomic molecules:

E_{A-A} , E_{B-B} , and E_{A-B} : energies in $\text{kJ}\cdot\text{mol}^{-1}$ to break the bonds A-A, B-B, and A-B.

The difference in electronegativity between elements B and A is given by:

$$|EN_{(B)} - EN_{(A)}| = \Delta^{1/2} \quad \Delta = \text{resonance energy, } \Delta \text{ is expressed in eV}$$

$$\Delta = Ed (\text{real bond}) - Ed (100\% \text{ covalent bond})$$

$$Ed (100\% \text{ covalent bond}) = \sqrt{E_{A-A} \cdot E_{B-B}}$$

As:

1eV is equivalent to $96.48 \text{ kJ}\cdot\text{mol}^{-1}$; $(96.48)^{-1/2} = 0.102$

$$|EN_{(B)} - EN_{(A)}| = 0.102 \cdot \sqrt{E_{A-B} - \sqrt{E_{A-A} \cdot E_{B-B}}} \text{ (in } eV^{1/2}\text{)}$$

The Pauling scale is a relative scale: "EN" of F = 4.0 (in $eV^{1/2}$)

c. Mulliken scale:

The Mulliken scale is the oldest and simplest electronegativity scale. It was proposed by Robert Mulliken in 1934. The Mulliken scale is based on the idea that the electronegativity of an atom is related to the energy required to remove an electron from that atom.

- Pauling scale:

The Pauling scale was proposed by Linus Pauling in 1932. The Pauling scale is based on the idea that the electronegativity of an atom is related to the bond dissociation energy of a diatomic molecule that contains that atom.

The Pauling scale is more widely used than the Mulliken scale. This is because the Pauling scale is more intuitive and easier to use.

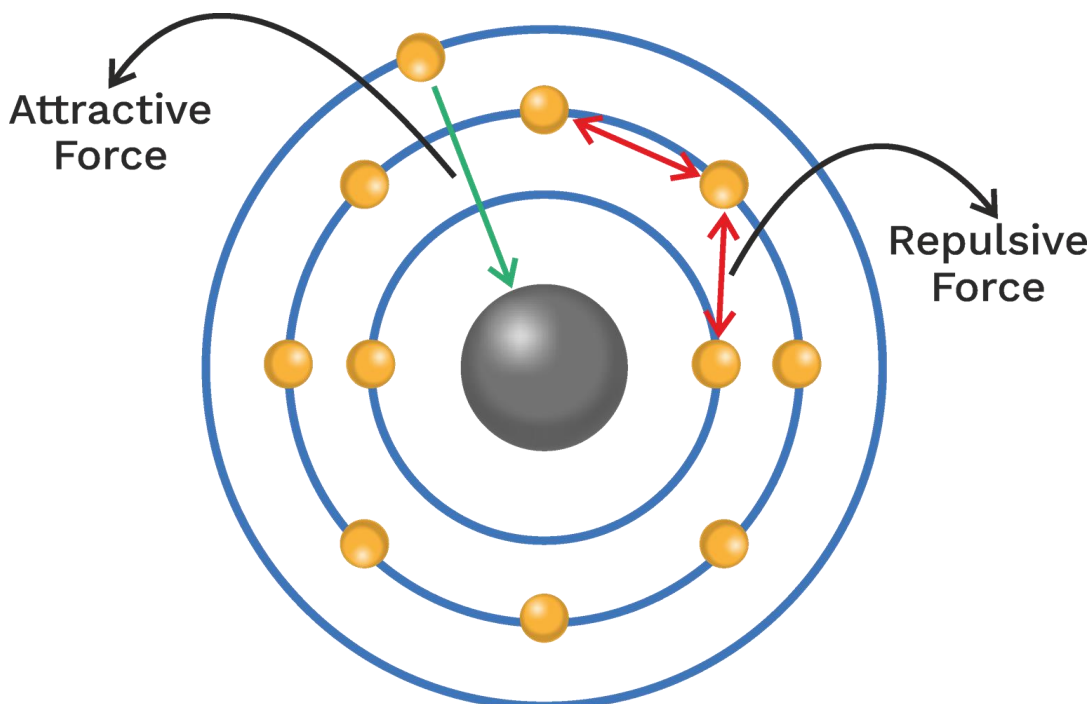
D. Z_{eff} (Effective nuclear charge)

$$Z_{\text{eff}} = Z - \sigma$$

Where,

Z is number of protons

σ is shielding/screening constant



Shielding effect:

(1) It is the phenomenon in which nuclear charge present on outermost electron is decreased by presence of inner or outer electron.

(2) The shielding of electron present in s-orbital is most effective, p-orbital is effective, d-orbital is poor and f-orbital is negligible.

$s > p > d > f$ (shielding effect).

Slater's Rule :

(Calculation of σ and z_{eff})

(1) For ns/np electron.

$$n^{\text{th}} \Rightarrow \sigma = 0.35$$

$$(n-1) \Rightarrow \sigma = 0.85$$

$(n-2)$ or lower $\Rightarrow \sigma = 1.0$

(2) For nd/nf electron.

nd/nf $\Rightarrow \sigma = 0.35$

(3) For 2 electron species (Ex : He, Li⁺, Be⁺⁺ etc.)

$\sigma = 0.3$

Examples :

(1) Calculate the σ and z_{eff} for Na.

Solution: Na = $1s^2 2s^2 2p^6 3s^1$

$\sigma = 8 \times 0.85 + 2 \times 1.0 = 8.85$

$Z_{\text{eff}} = Z - \sigma = 11 - 8.85 = 2.15$

(2) Calculate the σ and z_{eff} for 3s electron of Mg.

Solution: Mg = $1s^2 2s^2 2p^6 3s^2$

$s = 1 \times 0.35 + 8 \times 0.85 + 2 \times 1.0 = 9.15$

$Z_{\text{eff}} = Z - \sigma = 12 - 9.15 = 2.85$

- From top to bottom Z_{eff} value is

Li = 1.30

Na = 2.20

K = 2.20

Rb = 2.20

Metallic and Non-metallic Character

The tendency of an element to lose electrons to form cations is called electropositive or metallic character of an element. Alkali metals are most electropositive. The tendency of an element to accept electrons to form anions is called electronegative or non-metallic character of an element.

Variation of Metallic Character in a Group

Metallic character increases from top to bottom in a group as tendency to lose electrons increases. This increases the electropositive character and metallic nature.

The variation can best be seen in group 14 as shown below.

Example: Metallic character of groups 14 elements

| Element | Nature |
|---------|-----------|
| C | Non-metal |

| | |
|----|-----------|
| Si | Metalloid |
| Ge | Metalloid |
| Sn | Metal |
| Pb | Metal |

Along a period:

- The electropositive metallic character decreases when Z increases.

Along a column:

- The electropositive metallic character increases when Z increases.

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