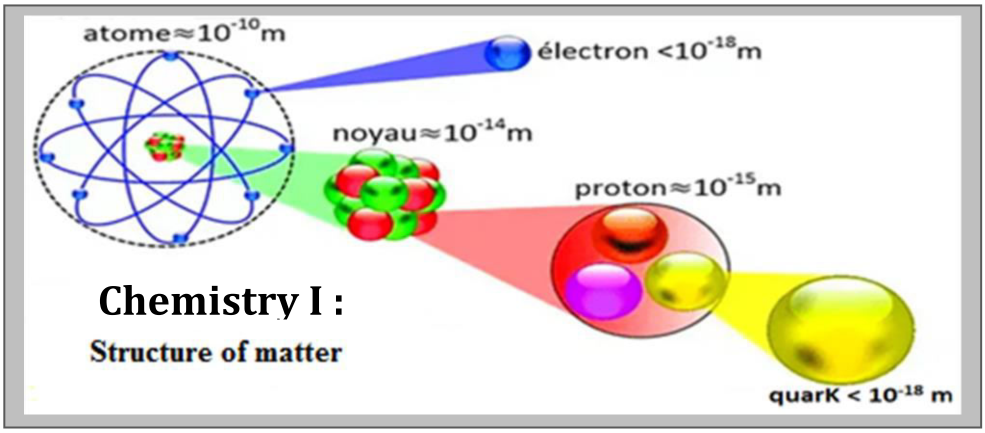
الجمهورية الجزائرية الديمقراطية الشعبية
People's Democratic republic of Algeria
وزارة التعليم العالي والبحث العلمي
Ministry of Higher Education and Scientific Research
جامعة عبد الحميد بن باديس مستغانم
AbdelhamidIbnBadis University – Mostaganem
كلية العلوم والتكنولوجيا
Faculty of Sciences and Technology
قسم العلوم والتقنيات
Department of Science and Techniques

Course Handout

Chemistry I

Course Material and Exercises with Answer Keys



Prepared by :

Dr. Mohammed Cherif TERKHI ***Senior Lecturer « A »***

Reviewed by : Pr GHEZZAR M.R. and Pr AZIZ Z.

Academic Year : 2024-2025

FOREWORD

This handout, written for the first time in English at the University of Mostganem, is intended for first-year science and technology students of the L.M.D. (Bachelor-Master-Doctorate) program. It includes a reminder of the Chemistry Module I (Structure of Matter), illustrated by real photos and simplified diagrams. Students will follow the new undergraduate program easily by acquiring the fundamental concepts of matter structure. The course is made up of six chapters:

The first chapter is devoted to reminders of fundamental concepts. The second chapter deals with the different components of matter, namely the proton, the neutron and the electron and the nuclear symbol. The third chapter introduces radioactivity. The fourth chapter describes two atomic models: the classical model and the quantum (or wave) model. The goal of the fifth chapter is to explain the principles of the periodic classification of chemical elements and their physicochemical properties.

The last chapter is devoted to chemical bonding and the different types of bonds. This handout also includes sample exercises with answer keys to help the student better assimilate the module.

Summary

Chaptre I « General definition »

Chaptre II « Atomic model evolution »

Chaptre III « Radioactivity »

Chapter I

General definition

I- Some definitions :

1-Chemistry : Is the science that studies the structure of matter and its transformations.

2-Matter : Is anything that occupies a volume in space and has a mass.

3-Mass : Is the amount of particles in a substance measured in kg.

4-Volume : Is the space occupied by a substance measured in cm^3 or ml.

II- Properties of matter :

The three most widely recognized states of matter are shown in Fig. I.1 :

Figure I.1 : The state of matter

II-1-Solids :

-Particles are very close to each other.

-Existence of a strong attractive force between particles.

-Has a defined mass (doesn't change).

-Has a defined volume.

-Has a defined shape (does not change)

II-2-Liquids :

-The particles are close to each other, but there's space between them.

-The force of attraction between particles is weaker than in solids. Particles move easily.

-Has a defined mass..

- Has a defined volume,

- Particles taking the shape of their container. Their shape is therefore indefinite.

- Their surface are horizontal. Particles move easily

II-3-Gas :

- The particles are far apart. The force of attraction is very weak (small in value).

- Has an undefined volume.

- Has indefinite mass (particles are always in motion).

- Has an undefined form because it takes the container form.

III- Changes in the state of matter (water) :

Un changement d'état est une opération de transition de phase en passant d'un état de matière à un autre.

III-1- Physical change :

The physical state changes, but the species remains the same (Fig. I.2).

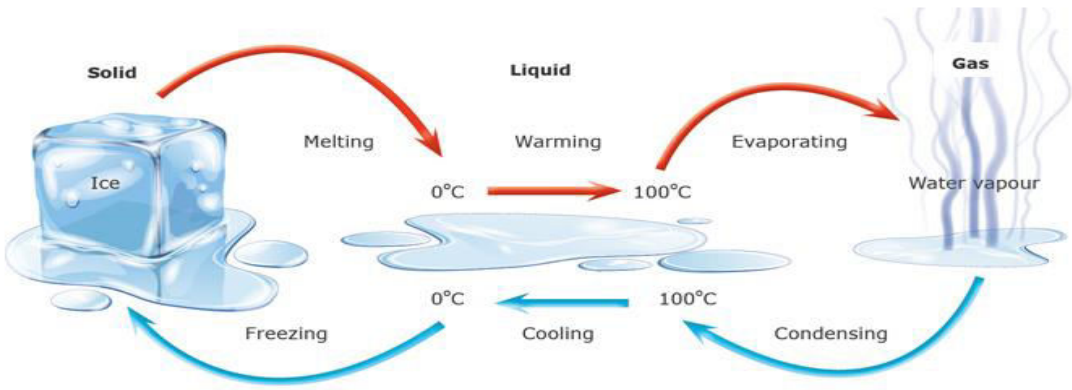


Figure I.2 : change of physical state of water

III-2- Chemical change :

Is a transformation that changes the nature of a substance by means of a chemical reaction (Fig. I. 3).

Figure I.3 : The chemical reaction that causes rust

IV- Matter organization :

IV-1-Pure substance :

IV-1-1- Element : A pure substance composed of only one kind of atom. E.g.: H_2 (See figure I.4).

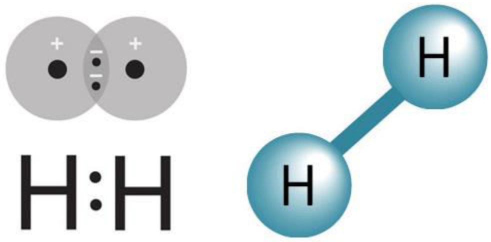


Figure I.4 : Chemical structure of (H_2)

IV-1-2- Compound : A pure substance composed of several types of atoms that we can separate by chemical techniques. E.g.: H_2O (See Fig. I.5).

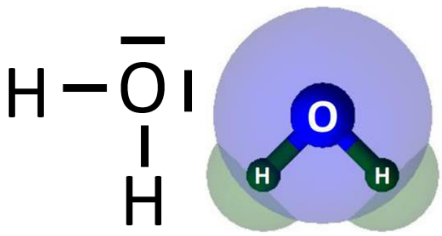


Figure I.5 : Chemical structure of (H_2O)

IV-2- Mixtures :

Several substances, at least two types of particles. The mixtures are divided into two groups :

IV-2-1- Heterogeneous :

Mixture in which at least two substances can be distinguished with the naked eye.

E.g.: water + oil

IV-2-2- Homogenous :

Mixture in which no more than one substance can be distinguished with the naked eye: E.g.: water + sugar Homogeneous mixtures are divided into two categories: Colloids and Solutions

A/-Colloïds :

Homogeneous mixture in which at least two substances can be distinguished under the microscope. e.g.: Whey

B/-Solutions :

Homogeneous mixture in which substances cannot be distinguished even with the aid of a microscope.

Solutions are divided into two components:

a) Solute : Substance dissolved in solvent. There may be more than one solute in a solvent. If the solute is solid, it's called dissolution; if it's liquid, it's called dilution.
b) Solvent : Substance capable of dissolving the solute.

C/- Preparation of solutions:

Use the volumetric flask to prepare a solution from a liquid or solid, or to make a dilution. The gauge line is located on the neck, which is which is longer and narrower than the base to allow precise volume adjustment (Fig. I.6).

Figure I.6 : Volumetric flask (5, 10, 25, 50 and 100 ml)

The organization of matter can be summarized as follows (Fig. I.7) :

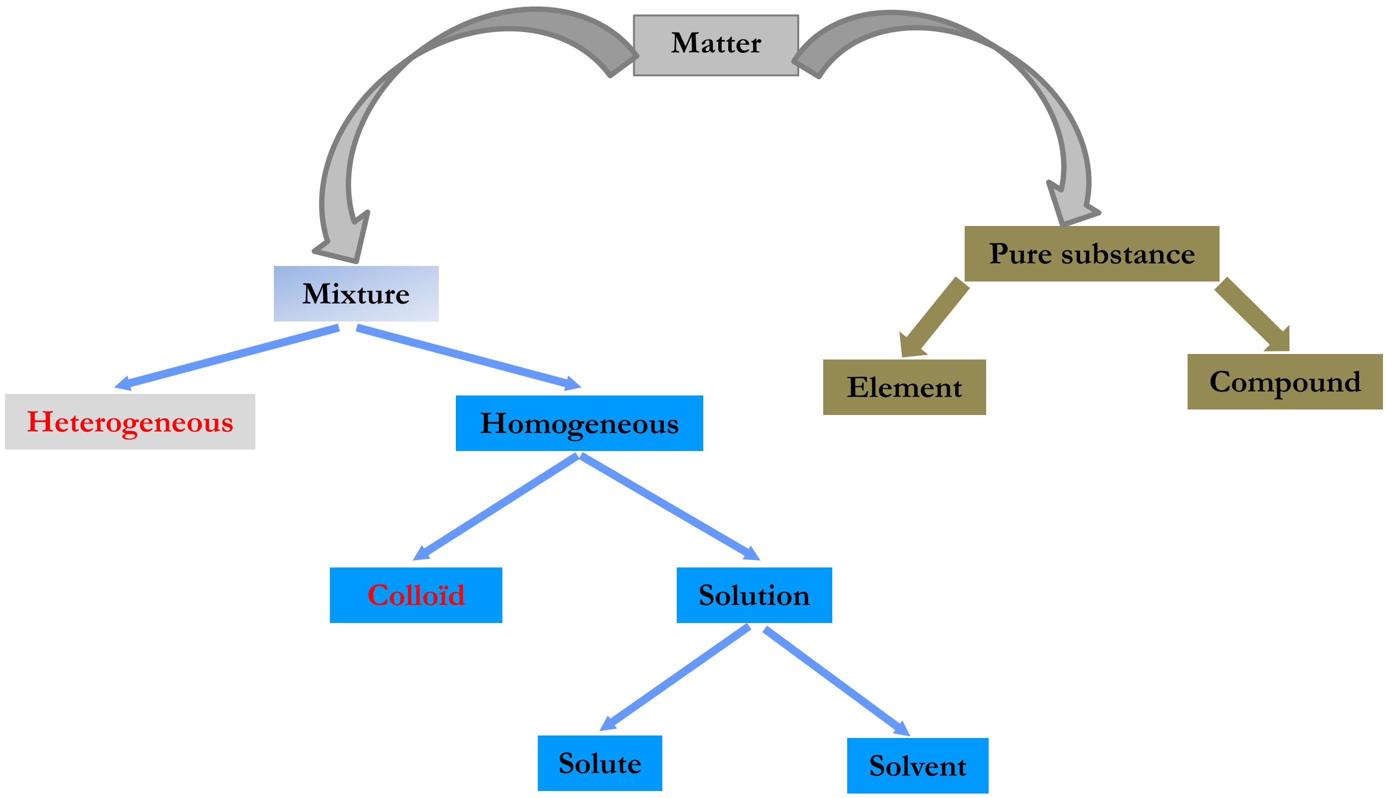


Figure I.7 : Matter organization

V- Concentration expressions :

V-1- Definition of concentration :

The concentration expresses the quantity of substance per unit volume. We distinguish three concentrations:

-*Mass concentration (g/l)* : Concentration in grams of solute per liter of solution.

$$C = \frac{m}{v}$$

-*Molar Concentration (mol/l)* : Concentration in moles of solute per liter of solution or Molarity.

$$C' = \frac{n}{v}$$

-*Molality mol/kg* : expresses the quantity of solute contained in 1000g of solvent..

-*Normality N (eqg/l)* : expresses the number of gram equivalents of solute per liter of solution.

.

Eqg= mass of chemical substance/number of equivalents

-*Percentage % of a solution*: indicates the mass of substance per 100g of solution. This is a weight-for-weight comparison

-*The mole fraction x* : indicates the ratio between the number of moles and the total number of moles in the solution. (dimensionless) $X_i = \frac{n_i}{n_T}$; $\sum X_i = 1$

VI-Mole :

1 mole is the quantity of a substance that contains the d'Avogadro number of particuls.

1mol(of atoms, ions, molecules) = 6.023 x10^{23}**(of atoms, ions, molculs)**

VII-The molar mass :

The mass of a mole of particles in g/mol.

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

VIII- Atoms Mass Unit (amu) :

The unit of mass of atoms (amu) is a measurement unit, used to measure the mass of atoms and molecules. It is defined as 1/12 of the mass of an atom of the ^{12}C.

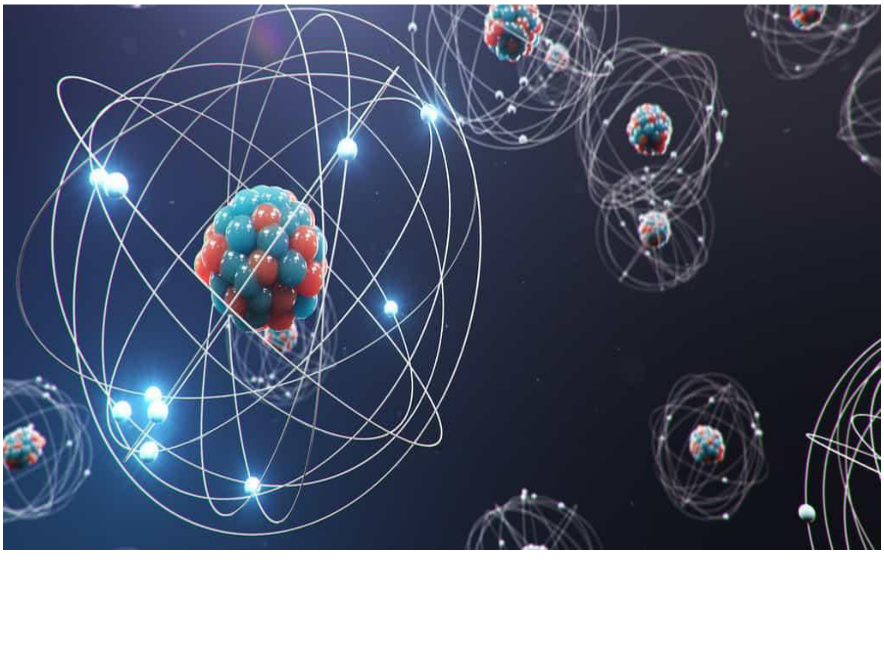
We have : 1 amu= (1/12) . atomic mass C (g)

With : atomic mass C=12/N_A (g)

Therefore : 1amu=(1/12). 12/N_A (g) = 1/ N_A (g) =1.66.10^{-24} (g)

Chapter II

Atomic model evolution



I-Introduction :

The evolution of the atomic model has gone through several models over several eras (Fig. II.1).

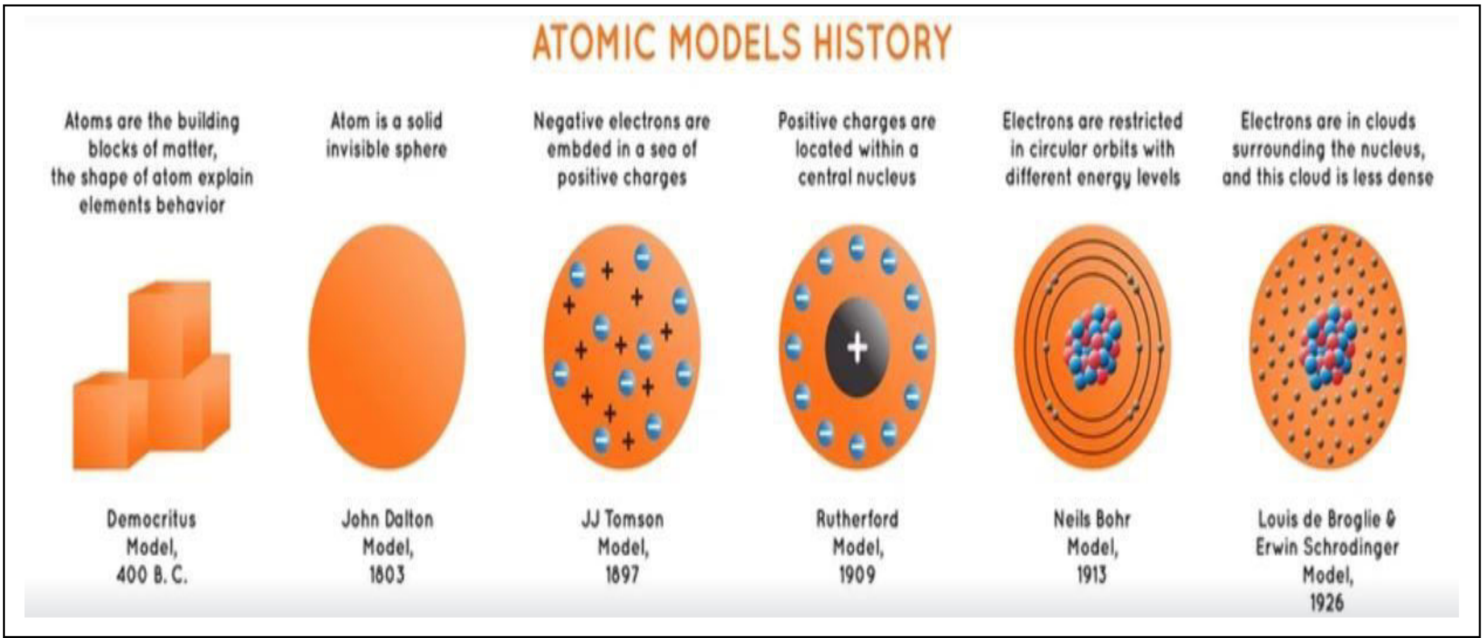


Figure II.1 : The evolution of the atomic model

I-1-Aristote 384-322 B.C. :

According to the Greeks : Matter is continuous, no vacuum, no atoms.

Matter consists of 4 elements : 1- Air ; 2- Earth ; 3- Fire ; 4-Water.

I-2-Démocrite 460-370 B.C. :

The Democrites say: " Matter is made up of small indivisible particles called atoms."

I-3-Dalton 1766-1844 :

Dalton proposed the following theory : “matter is made up of indivisible, indestructible atoms, and an atom of a given element always has the same mass.”

II- Electron discovery :

II-1- Crookes Experiments :

Crookes tube (Fig. II.2) was invinted by the British physician William Crookes between 1869 and 1875 during which cathode rays (electrons) were discovered.

Cathode rays are only produced when the cathode and anode of the cathode-ray tube are connected to across a high-voltage source.

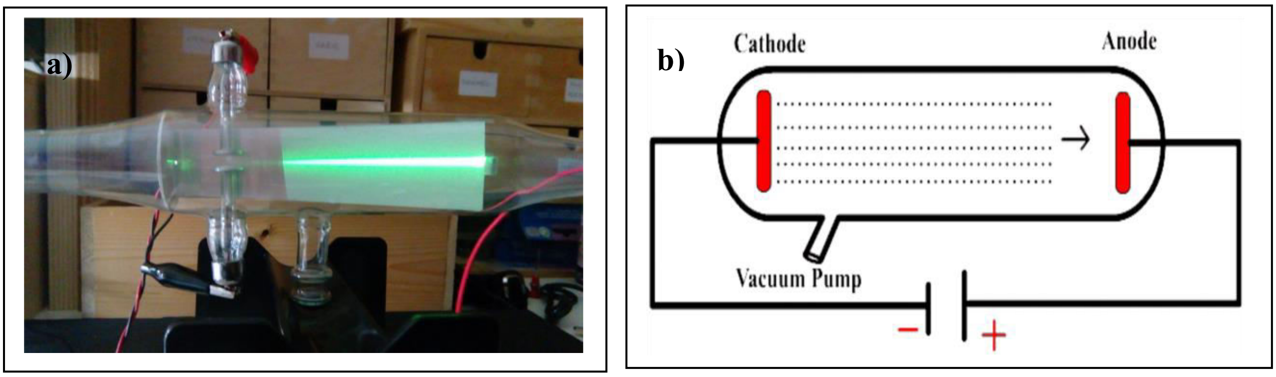


Figure II.2 : Tube de Crookes : **a)** Real photo and **b)** Simplified drawing

- Cathodic rays behavior:

The three main results obtained after several experiments are as follows:

1- When blades are placed in the center of the tube, we can see that they rotate.

This tells us that cathode rays are made up of particles with mass and kinetic energy..

2- The beam is deflected by a magnet.

This tells us that cathode rays are made up of charged particles.

3- The beam is deflected by an electric field.

This tells us that cathode rays are made up of negatively charged particles.

II-2-J.J. Thomson experiments :

The experiment (1895) is shown in Fig. II.3. The aim of this experiment is to determine the ratio between the electron's elementary charge and its mass (e/m).

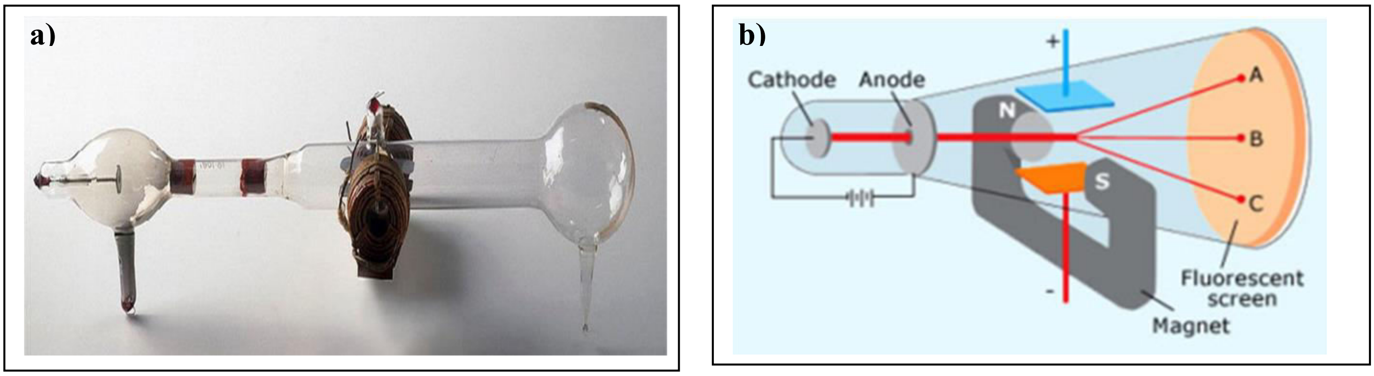


Figure II.3 : J.J. Thomson experiments : **a)** Real photo and **b)** Simplified drawing

-J.J. Thomson's experiment is composed of three steps :

A/ Simultaneous action of both fields E et B (Fig. II.4) :

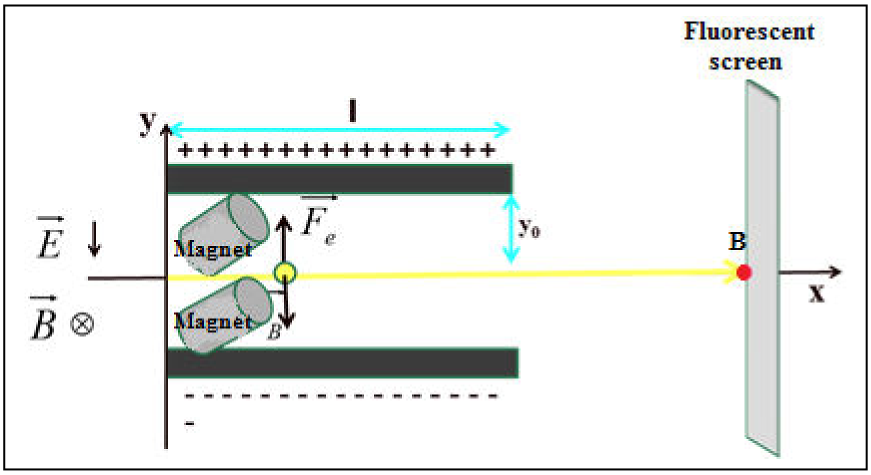


Figure II.4 : Simultaneous action of both fields E and B

Under the action of the two fields, the electron beam (yellow) does not deflect. We have rectilinear uniform motion up to the point **B** $\Rightarrow \vec{\gamma} = 0$

$$\sum \vec{F} = m.\vec{\gamma} = 0 \Rightarrow F_e - F_B = 0 \Rightarrow F_e = F_B \Rightarrow q.E = q.V.B \Rightarrow V = \frac{E}{B} \quad (1)$$

B/ Action of E field alone (Fig. II.5).

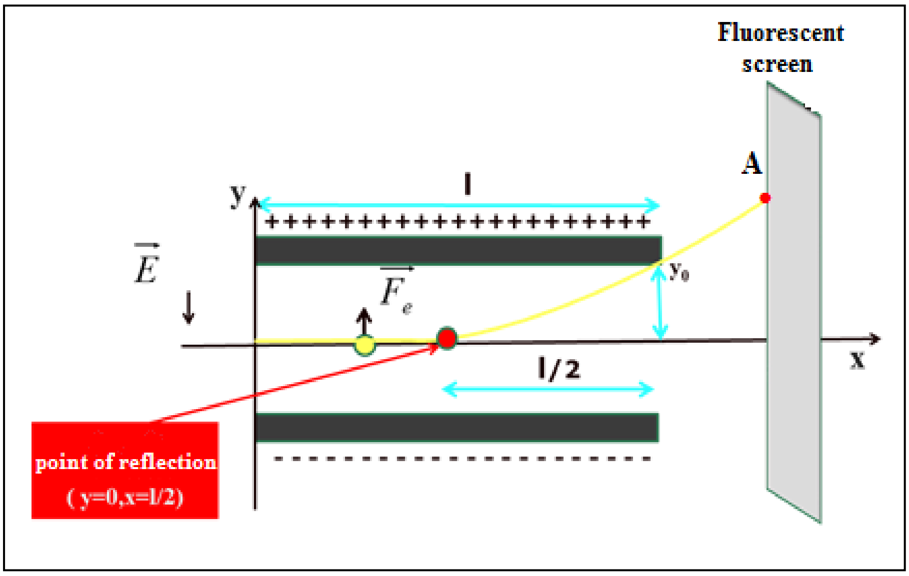


Figure II.5 : Action of E field alone

Under the action of the E field alone, the electron beam (yellow) deflects towards the plate +.

Along OX axis: no force : $on\ a \sum \vec{F} = m.\vec{\gamma}$

$$\Rightarrow \sum \vec{F} = m.\vec{\gamma} = 0 \Rightarrow \gamma_x = 0 \Rightarrow x = v.t \quad (2)$$

So the electron beam has a uniform rectilinear motion.

Along OY axis:

$$on\,a \sum \vec{F} = m\vec{\gamma} \Rightarrow F_e = m\gamma_y = q.E \Rightarrow \gamma_y = \frac{q.E}{m} = cst \quad (3)$$

So the electron beam has a uniformly accelerated motion:

$$\Rightarrow y = \frac{1}{2}\gamma_y t^2 \quad (4)$$

(2) and (3) in (4) :

$$\Rightarrow y = \frac{1}{2}\gamma_y \left(\frac{x}{V}\right)^2 \Rightarrow y = \frac{1}{2}\left(\frac{q.E}{m}\right)\left(\frac{1}{V}\right)^2 .x^2 \Rightarrow y = f\left(x^2\right)$$

⇒ The electron trajectory is parabolic.

Remark: at the condenser exit: ($x=l$, $y=y_0$) et $V=E/B$:

$$\Rightarrow y_0 = \frac{1}{2}\left(\frac{q}{m}\right)\left(\frac{B^2}{E}\right).l^2$$

This is the deflection equation at the output of the condenser, and we calculate $\frac{q}{m}$.

B/ Action of the only field B (Fig. II.6) :

Under the action of the only field B, the electron beam (yellow) describes a circular arc of radius R..

$$on\ a \sum \vec{F} = m.\vec{\gamma}$$

Following γ_T : no force :

$$\Rightarrow \sum \vec{F} = m.\vec{\gamma} = 0$$

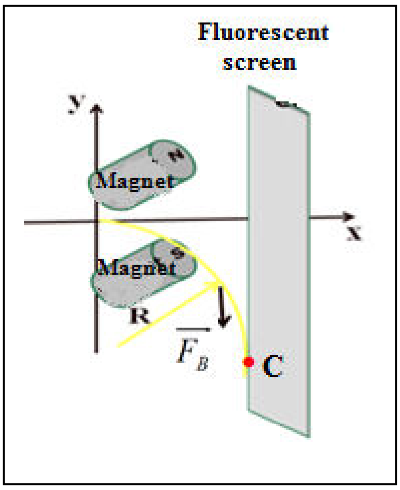


Figure II.6 : Action of the only field B

Following γ_N :

$$\Rightarrow \sum \overrightarrow{F} = F_B = m.\gamma_N \Rightarrow q.V.B = m.\frac{V^2}{R} \Rightarrow q.B = m.\frac{V}{R}$$

$$\Rightarrow R = \frac{m}{q}.\frac{V}{B} \Rightarrow R = \frac{1}{\frac{q}{m}}.\frac{E}{B^2}$$

So, the Atomic Model according to J.J. Thomson:

-Electricity does exist

-He discovered the electron

-The atom is therefore divisible

- Raisin bread model (Fig. II.7)

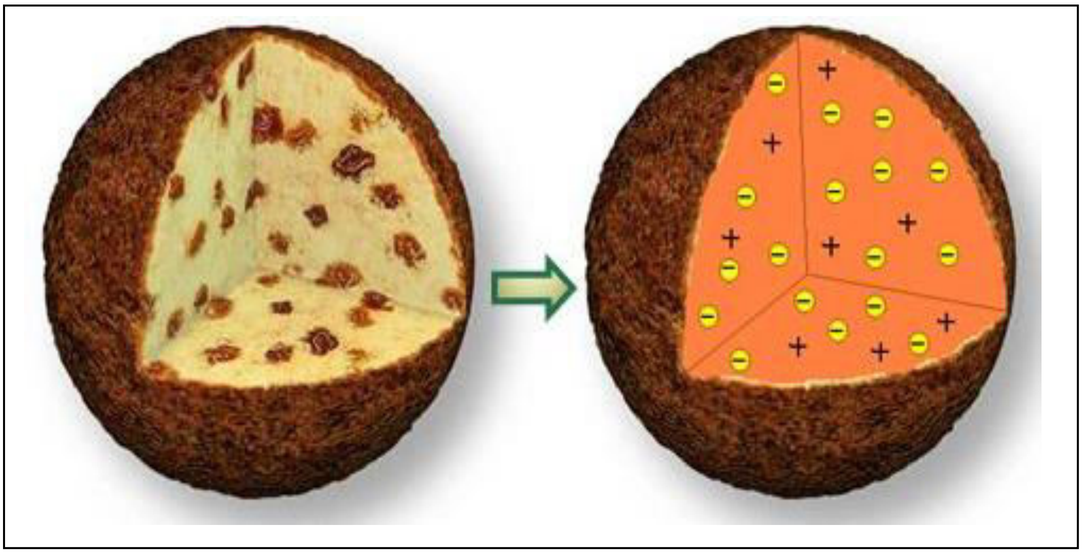


Figure II.7 : .J. Thomson's raisin bread model.

<u>II-3- Millikan's experiments :</u>

The aim of this experiment is to measure the electron's elementary charge.

- is the study of the movement of ionised oil droplets between the plates of a horizontal condenser (Fig. II.8).

- with each ionisation, the oil droplet acquires a new charge.

- The speed at which the droplet rises and falls is determined.

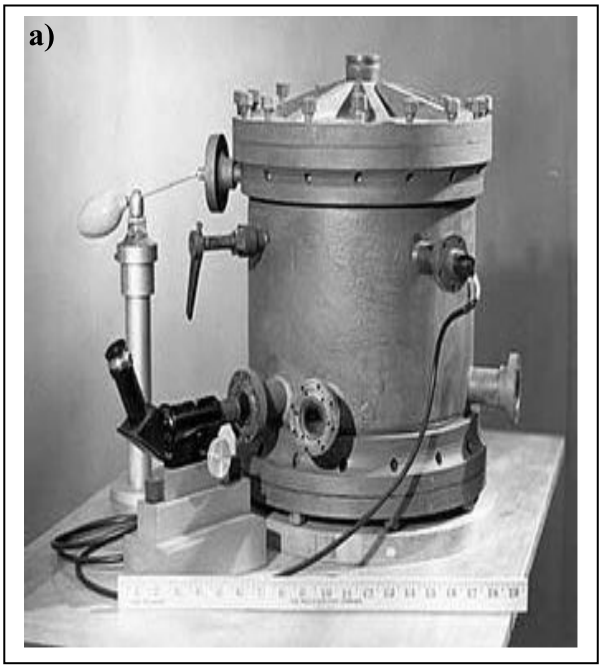
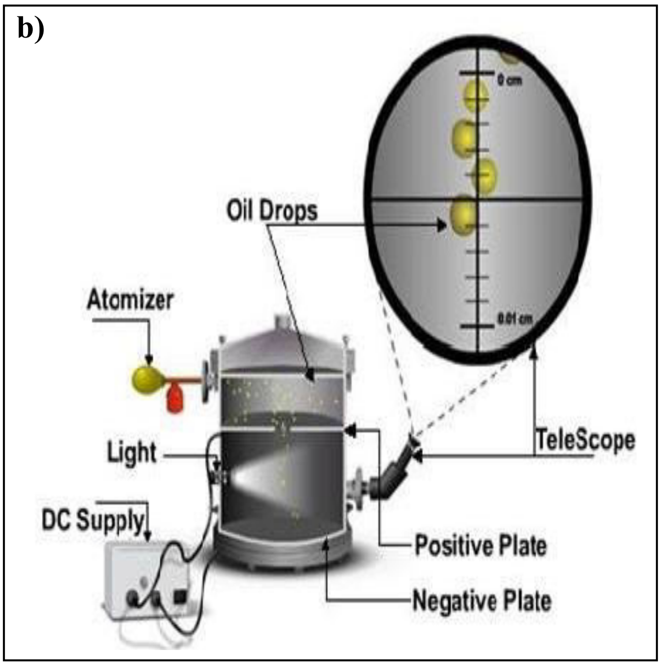



Figure II.8 : Millikan's experiments: **a)** Real photo and **b)** Simplified drawing

Millikan's experiment is composed of three steps :

A/ In the absence of the electric field E:

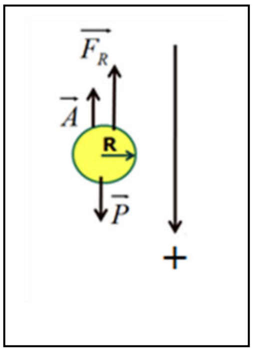


Figure II.9 : Free-fall of the droplet

The forces acting on the droplet are (Fig. II.9) :

- The Weight:

$$P = m.g = (\rho.V).g = \left(\rho.\frac{4}{3}\pi.R^3\right).g$$

- Friction force (or Stokes force) :

$$F_R = 6.\pi.\eta.R.v_0$$

- Archimedean thrust:

$$A = \left(\rho_0.\frac{4}{3}\pi.R^3\right).g$$

With:

ρ : Oil density.

ρ_0 : air density.

R : radius of the oil droplet (assimilated to a sphere of radius R.

V_0 : speed of oil droplet.

η : air viscosity coefficient.

Calculation of the droplet radius (R):

Because the oil droplet reaches very quickly its speed limit

$$\Rightarrow v = cst \Rightarrow \gamma = 0 \Rightarrow \sum \vec{F} = m.\vec{\gamma} = 0.$$

$$So : P - (A + F_R) = 0 \Rightarrow P - A = F_R \quad (1)$$

$$\Rightarrow \left(\rho.\frac{4}{3}\pi.R^3\right).g - \left(\rho_0.\frac{4}{3}\pi.R^3\right).g = 6.\pi.\eta.R.v_0$$

$$\Rightarrow (\rho - \rho_0).\left(\frac{4}{3}\pi.R^3\right).g = 6.\pi.\eta.R.v_0$$

$$\Rightarrow R = 3.\sqrt{\frac{\eta.v_0}{2(\rho-\rho_0).g}}$$

B/ In the presence of the electric field E1 (Fig. II.10) :

Calculation of droplet charge (q):

$$We\ have \sum \vec{F} = m.\vec{\gamma} = 0$$

Therefore :

$$P + F_{R_1} - (A + F_e) = 0 \Rightarrow P - A = F_e - F_{R_1} \quad (2)$$

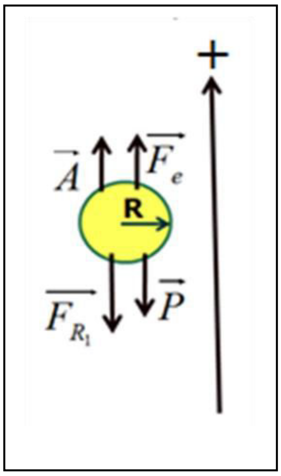


Figure II.10 : upwelling of the droplet

(1) dans (2) $\Rightarrow F_R = F_e - F_{R_1} \Rightarrow F_R + F_{R_1} = F_e$

$$\Rightarrow 6.\pi.\eta.R.v_0 + 6.\pi.\eta.R.v_1 = q.E$$

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

Therefore :

In the presence of the electric field E1 $\Rightarrow q_1 = \dfrac{6.\pi.\eta.R}{E_1}.(v_0 + v_1)$

In the presence of the electric field E2 $\Rightarrow q_2 = \dfrac{6.\pi.\eta.R}{E_2}.(v_0 + v_2)$

In the presence of the electric field Ei: $\Rightarrow q_i = \dfrac{6.\pi.\eta.R}{E_i}.(v_0 + v_i)$

Numerical application :

E1 →	$q = 4.8 \times 10^{-19}$ C	$q = 3 . (1.602 . 10^{-19})$
E2 →	$q = 3.2 \times 10^{-19}$ C	$q = 2 . (1.602 . 10^{-19})$
E3 →	$q = 6.4 \times 10^{-19}$ C	$q = 4 . (1.602 . 10^{-19})$
E4 →	$q = 1.6 \times 10^{-19}$ C	$q = 1 . (1.602 . 10^{-19})$

Millikan has shown that all values of q were a multiple of 1.602 . 10-19 C. This value therefore represents the smallest electrical charge that an oil droplet can carry: this is the elementary charge(e).

So: q= n. e

III- Rutherford's experiments :

In 1911,Rutherford carried out experiments which consisted of bombarding a gold foil with alpha particles beam. (Fig. II.11).

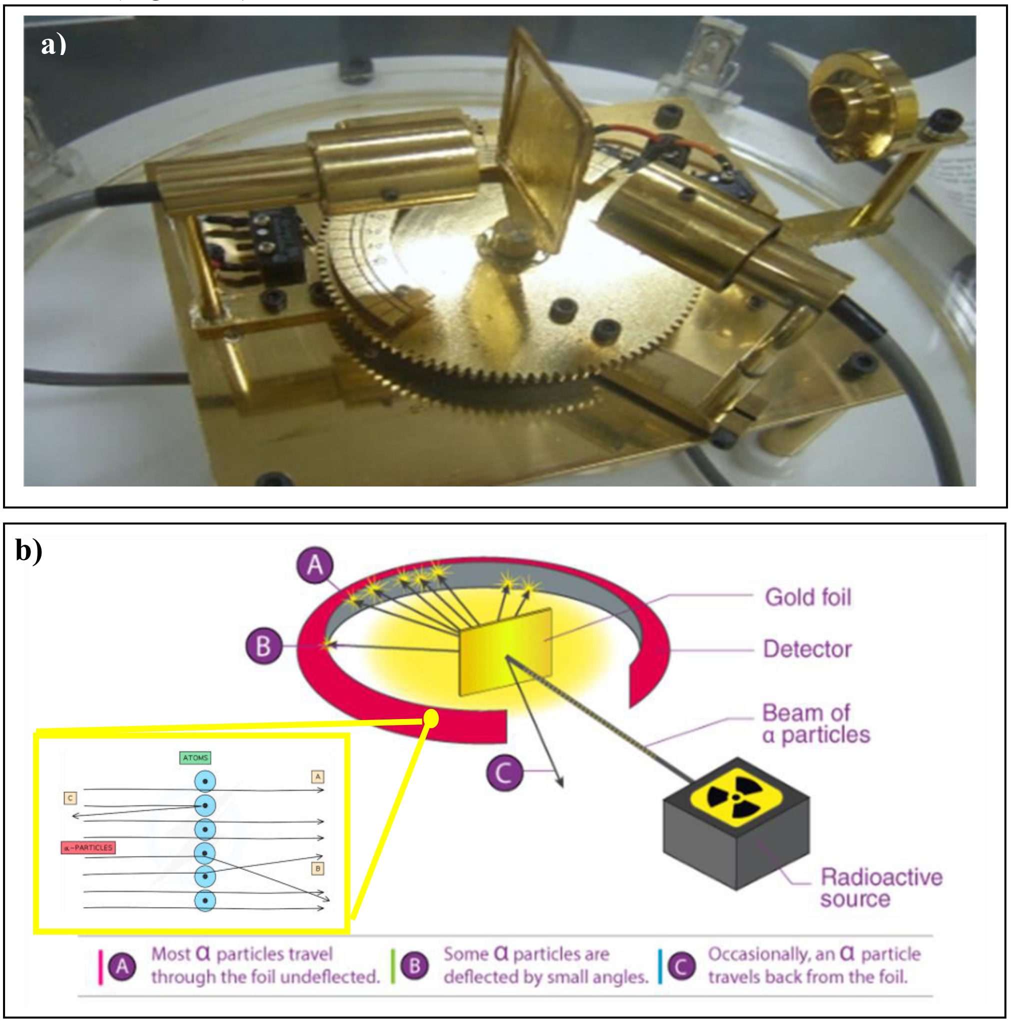


Figure II.11 : Rutherford's experiments : **a)** Real photo and **b)** Simplified drawing

Observation of Rutherfordde expériment :

-99 percent of alpha particles passed through atoms as if they found almost nothing in their path, so Rutherford concluded that "solid matter is transparent"

- Rutherford concluded that the atom contained a massive core with a positive electric charge, capable of repelling alpha.

-The experiment that proved the existence of a nucleus in the atom.

IV- Structure of an Atom :

- Each atom possesses : - A nucleus of overall charge(+) made up of (Fig. II.12) :

- Protons: electrical charge (p+)

- Neutrons: neutral electrical charge(n^0)

Nucleus

- Electrons (e-) orbiting around the nucleus.

- 99.9% of atomic mass is represented by the nucleus = (protons + neutrons)

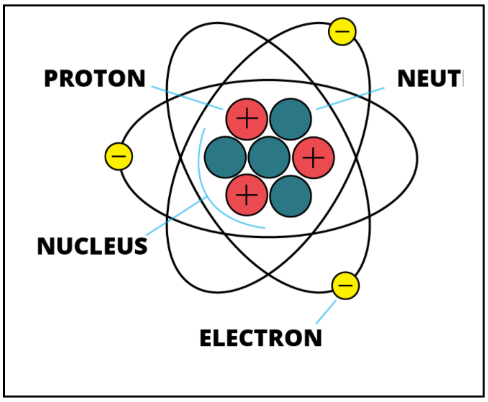


Figure II.12 : Schematic representation of the structure of an atom

IV-1- The protons :

In 1911 : bombardment of nitrogen by α-particles (^4_2He) was highlighted by Rutherford.

$$He_2^4 + N_7^{14} \rightarrow O_8^{17} + H_1^1 \quad ; \quad H_1^1 \equiv \text{proton} \equiv p$$

IV-2-The neutrons :

In 1932 : bombardment of an element Be by α particles (^4_2He) was highlighted by par James Chadwick.

$$He_2^4 + Be_4^9 \rightarrow C_6^{12} + n_0^1 \quad ; \quad n_0^1 \equiv \text{neutron} \equiv n$$

Proton	Neutron
$e = +1{,}60\ 10^{-19}$ Coulomb	0
$m_p = 1{,}7\ 10^{-27}$ kg	$m_n = 1{,}7\ 10^{-27}$ kg

IV-3- The electrons :

- Each electron carries a single negative charge.
- They are all identical.
- They form an electronic cloud.
- They have a very low (negligible) mass : $m(e^-) = 9{,}1.10^{-31}$ kg.

IV-4- The nuclides :

X : atom, element or nuclide. $\left({}^{A}_{Z}X\right)^{\text{charge}}$

-The number of protons in the nucleus is called the atomic number Z or charge number.

-The number of nucleons (protons + neutrons) is called A. It is called the mass number..

-The number of neutrons is denoted N.

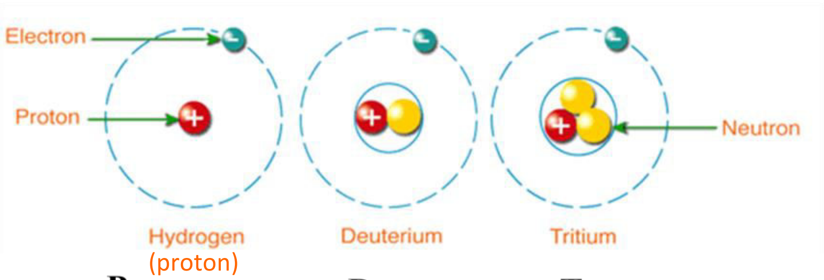
We have : N=A-Z

-The electrons : é= (Z ± the charge).

V-Isotopes :

V-1-Définition : Atoms with the same atomic number Z but a different mass number A are called isotope atoms.

i.e 1 : → ${}^{A1}_{Z}X$ ${}^{A2}_{Z}X$ ${}^{A3}_{Z}X$



P $H^1{}_1$ **D** $H^2{}_1$ **T** $H^3{}_1$

i.e 2 :

The image shows two samples, each with a mass of 100g. The sample on the left contains water (H_2O) and the one on the right contains heavy water (D_2O). Because of the difference in density, the volume occupied by heavy water is 11% less than that of normal water (Fig. II.13).

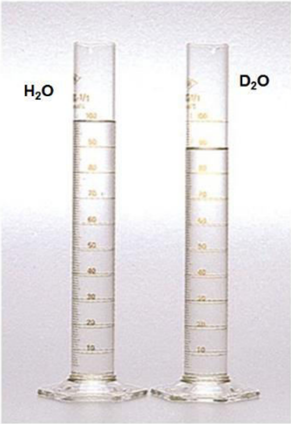
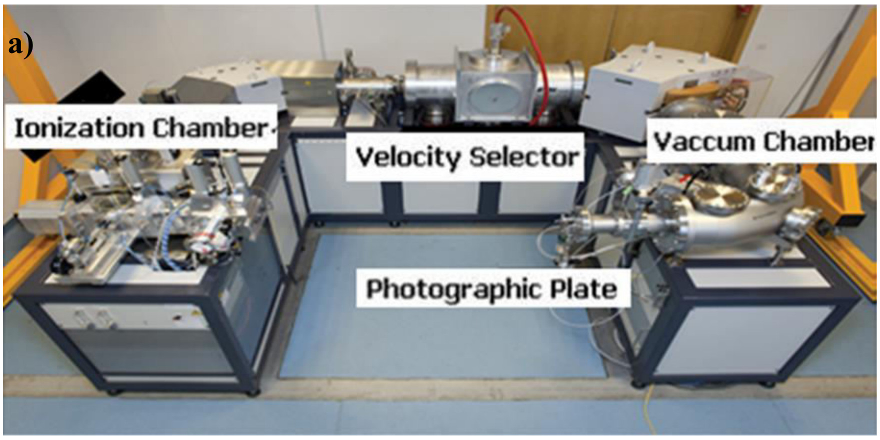


Figure II.13 : Experiment showing the difference in volumes occupied by the same compound as a function of density

V-2- Bainbridge experiment :

Aim: The mass spectrometer developed by Bainbridge in 1933 allows the mass of isotopes to be measured (Fig. II.14).



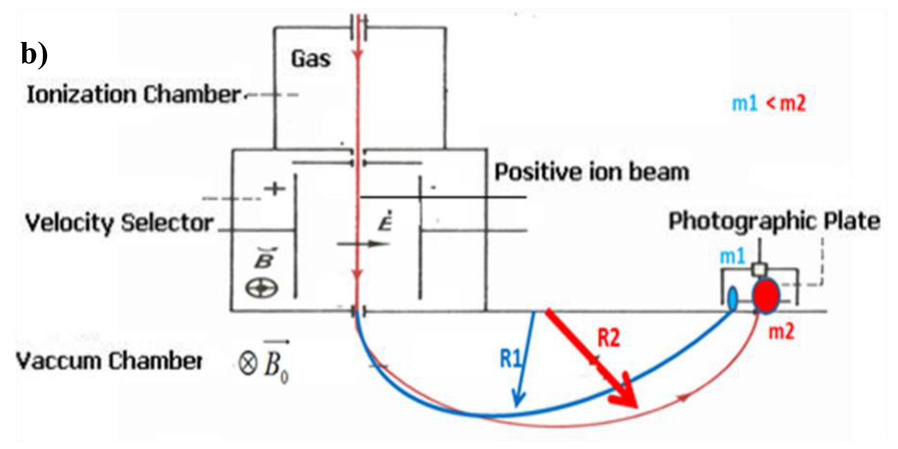


Figure II.14 : Bainbridge experiment : **a)** Real photo and **b)** Simplified drawing

A mass spectrograph is used, with 4 parts :

1. Ionizer : Obtaining positive ions by gas bombardment.

2. Speed filter: Ions accelerated by a capacitor (condenser) and introduction of a speed filter so that the ions enter the analyser at the same speed.

3. Analyser : Ions deflected by a constant magnetic field B_0.

4. Detector : electronic counting of impacts.

<u>In the speed filter :</u>

Simultaneous action of the two fields E and B on the ions : no deflection.

$$\sum \vec{F} = m\vec{\gamma} = 0 \Rightarrow F_e - F_B = 0 \Rightarrow F_e = F_B \Rightarrow q.E = q.V.B \Rightarrow V = \frac{E}{B}$$

<u>In the analyser :</u>

Action of the magnetic field Bo alone: the ion is deflected along a circle of radius (R) such that: :

$$We\ have\ \sum \vec{F} = m.\vec{\gamma}$$

Following γ_T: no force : $\Rightarrow \sum \vec{F} = m.\vec{\gamma} = 0$

Following γ_N:

$$\Rightarrow \sum \vec{F} = F_{B_0} = m.\gamma_N \Rightarrow q.V.B_0 = m.\frac{V^2}{R} \Rightarrow q.B_0 = m.\frac{V}{R}$$

$$\Rightarrow R = \frac{m}{q}.\frac{V}{B_0} \Rightarrow \frac{m}{R} = \frac{q.B_0.B}{E} = cst$$

$$We\ have: \frac{m}{R} = \frac{q.B_0.B}{E} = cst \Rightarrow \frac{m_1}{R_1} = \frac{m_2}{R_2} = \frac{m_3}{R_3} = ... = \frac{m_i}{R_i} = \frac{q.B_0.B}{E} = cst$$

<u>V-3- Mean atomic mass :</u>

The average atomic mass is the average of the mass of an atom.

$M_{moy} = \Sigma\ (\ . mi)$	$M_{moy} = \Sigma\ (x_i.\ mi)/100$
With : $0 \leq x_i \leq 1$	With : $0 \leq x_i \leq 100\%$
and $\Sigma\ x_i = 1$	and $\Sigma\ x_i = 100\ \%$

We have : - x_i : *abondance isopique i* (percentage of existence).

- *mi : isotope mass i*

i.e : *Uranium* **U_{92}**, this is a mixture of three isotopes:

^{234}U (x_1=99,275%)

^{235}U (x_2=0,720%)

^{238}U (x_3=0,005%)

<u>VI- Equivalency : mass - nucleus binding energy (Einstein-1905) :</u>

The mass of a real nucleus is always less than the sum of the masses of its constituents.

M (real nucleus) <M (noyau)

With M (nucleus)=Z.m(p)+N.m(n).

<u>VI-1- Mass defect :</u>

The mass defect (Δm) is shown in Figure II.15.

$$\Delta m = M_2 - M_1.$$

M_1 : The actual nucleus mass ;
M_2 : Theoretical nucleus mass = $Z.m_p+N.m_n$
m_p : proton mass
m_n : neutron mass

With : $M_1<M_2$

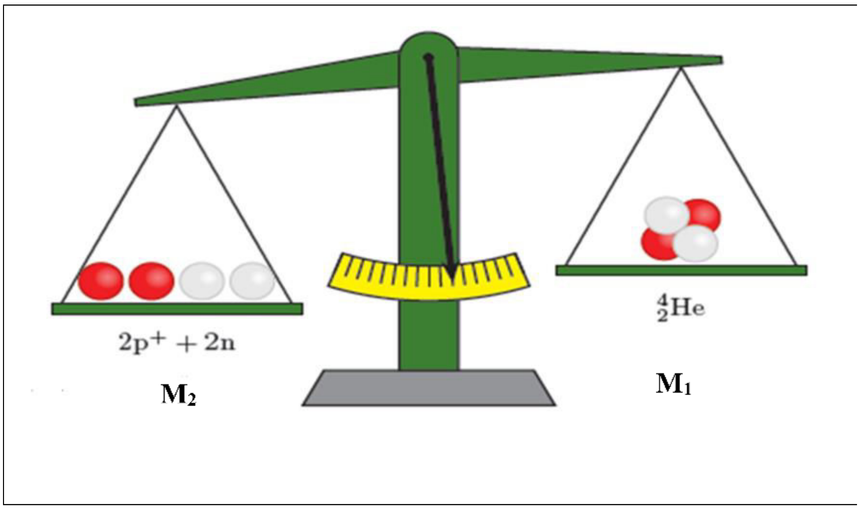


Figure II.15 : Explanatory diagram of the mass defect (example of $_2He$)

What happened to this mass (Δm) ?

-The mass defect is transformed into binding energy, which can be calculated using Einstein's relation :

$$\Delta E = \Delta m\, c^2.$$

VI-2-Binding energy per nucleon :

This is the energy required to tear a nucleon from the nucleus, increase from ^{1}H to ^{56}Fe and decreases above ^{56}Fe (Fig. II.16).

Remark: The greater the ΔE_A, the more stable the nucleus.

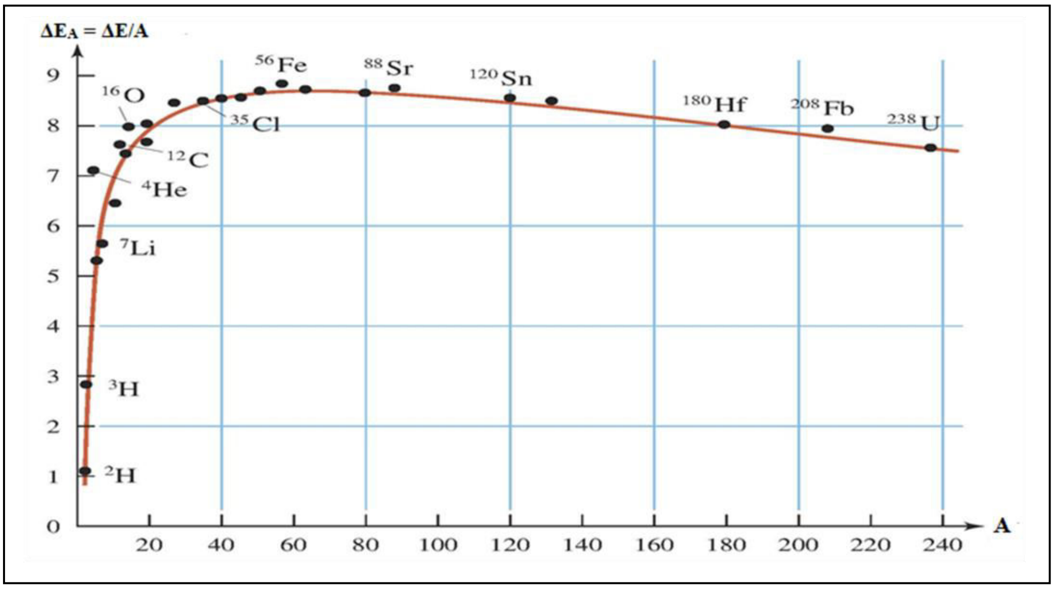


Figure II.16 : Variation of the binding energy per nucleon as a function of mass number (A).

VII- Mass and Energy units :

- In the S.I: the mass in kg and the energy in Joules (J) ;
- In nuclear physics : the mass in atomic mass uni (amu) is the twelfth of the mass of a carbon 12 atom :

$$1amu= (1/12)\,.\,(12.10^{-3}/N_A) = (1/12).(12.10^{-3}/\ 6{,}023.\ 10^{23}) = 1{,}66.10^{-27}\ Kg$$

Mass of a proton = 1,00727 uma and Mass of a neutron =1,00867 uma

-Energy in joule (J) and Electron-volt (eV) :

$\Delta E\ (J) = \Delta m\ (Kg)\ .\ c^2\ (m^2/s^2)$

and $\Delta E\ (Mev) = \Delta m\ (uma)\ .\ 931{,}5$

Demonstration of the energy relationship

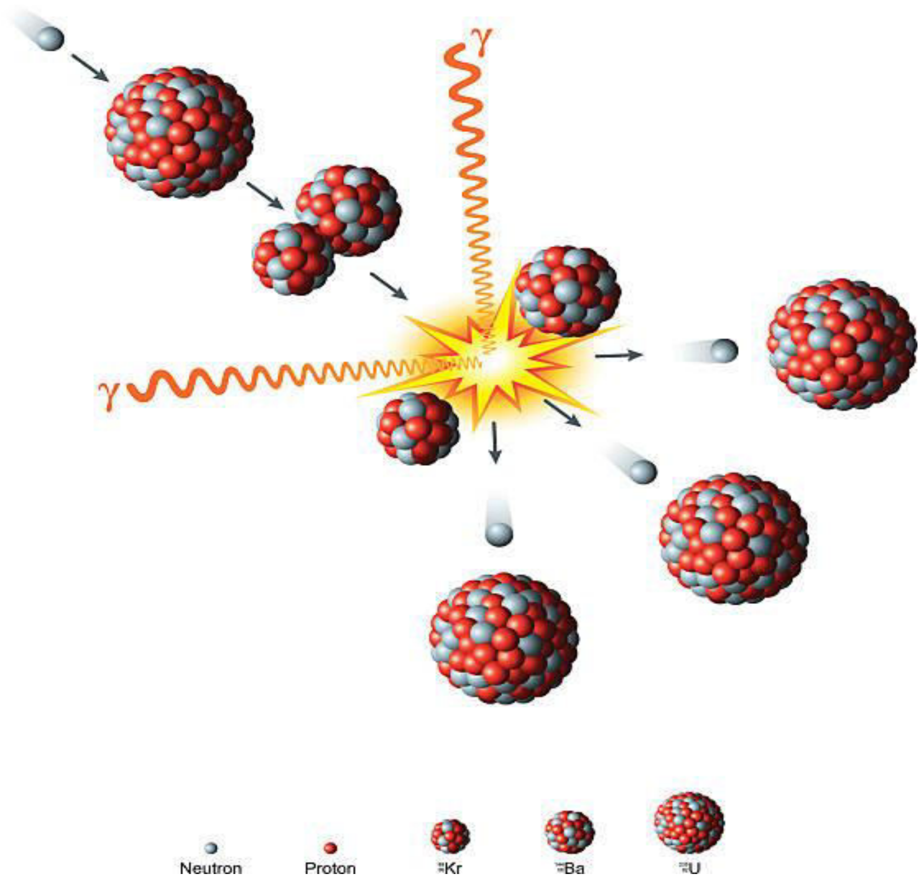
ΔE (J)= Δm(Kg). c^2 (m^2/s^2) = $\Delta m. (3.10^8)^2 / 1,60. 10^{-19}$ (ev) = $\Delta m. (3.10^8)^2 / 1,60.10^{-13}$ (Mev)

= $\Delta m. (1.66. 10^{-27})$ (uma). $(3.10^8)^2 / 1,60.10^{-13}$ (Mev) $\Rightarrow \Delta E$ (Mev) = Δm (uma) . 931,5

We have : 1 eV = $1,60.10^{-19}$ J ; 1 MeV = $1,60.10^{-13}$ J ; 1 MeV = 10^6 eV.

Chapter III

Radioactivity



I-Historique :

- ✓ Marie Curie and her husband Pierre used chemical methods to isolate two new radioactive element :

  - Polonium (July 1898)

  - Radium (december 1898)

- ✓ It was found that temperature, pressure or chemical state had no effect on radioactivity

- ✓ It became clear that radioactivity was due to an unknown process occurring inside atoms.

In 1896, Becquerel discovered radioactivity by accident, while researching the fluorescence of uranium salts.

II- Why are some nuclei radioactive :

The nuclei richest in protons and neutrons are the most radioactive. Possessing high energy, the nucleus is thus unstable, it tends to decompose (radioactive decay) (Fig. III.1).

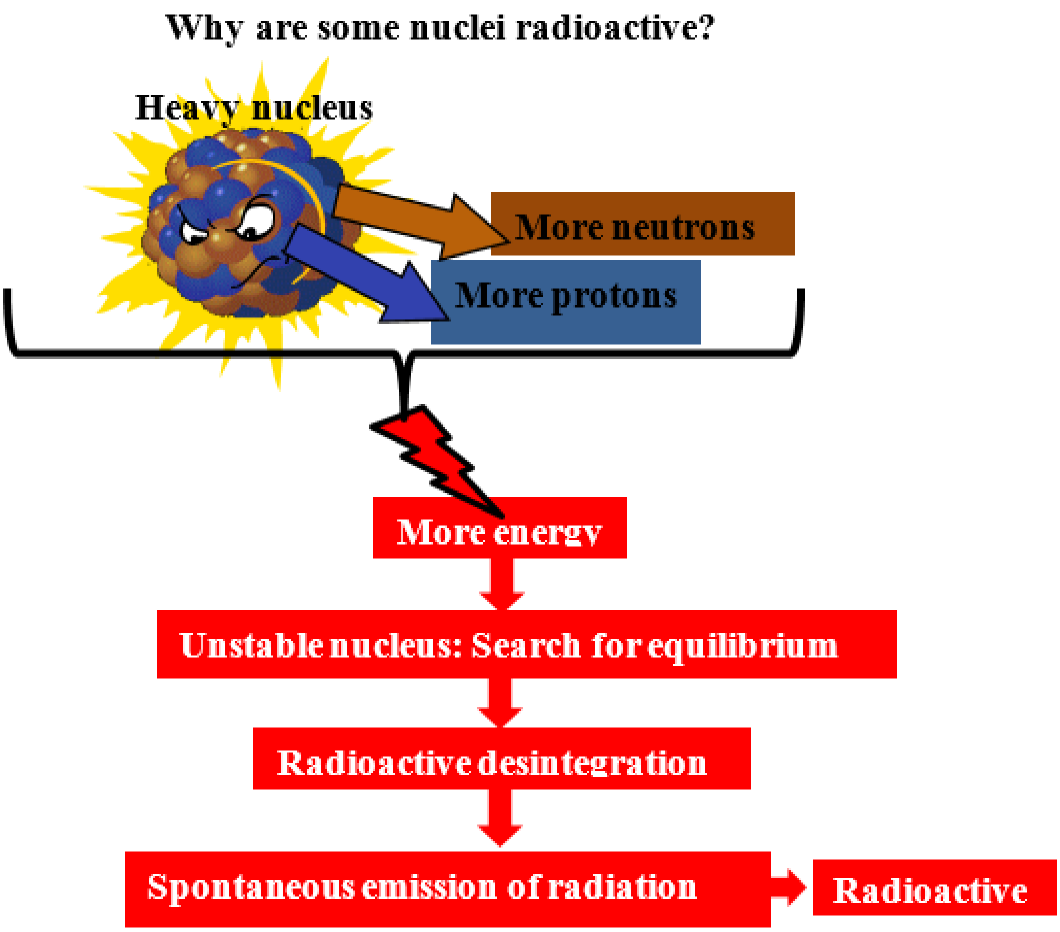


Figure III.1 : Explanatory diagram of radioactive decay.

III- Natural radioactivity :

-Due to the presence of naturally radioactive elements in our environment
-Natural radioactive elements are elements whose ratio : $\frac{A-Z}{Z} \geq 1{,}5$

For example : - Cosmic rays ;

-Soil radioactivity ;

-Radioactivity in water ;

-Radioactivity in the human body ;

-Radioactivity in the air.

IV- Emitted radiation :

During radioactivity, several types of radiation appear (Fig. III.2)

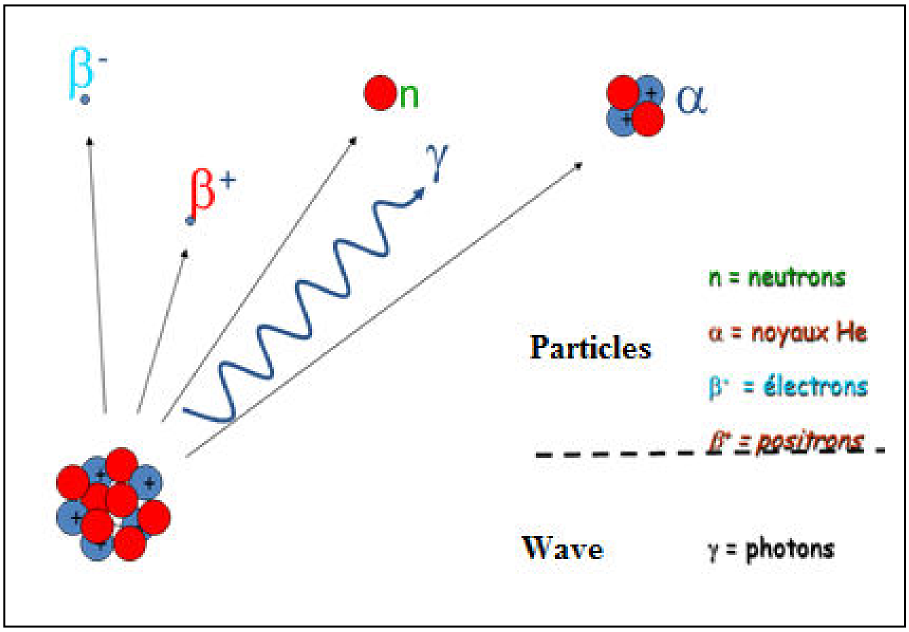
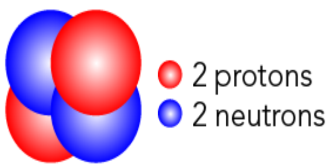


Figure III.2 : Emitted radiation

IV-1- Alpha Particuls :

- ✓ Rutherford discovered that alpha particles are in fact helium atoms He_2^4;
- ✓ With the symbole (α),
- ✓ **Réaction α**



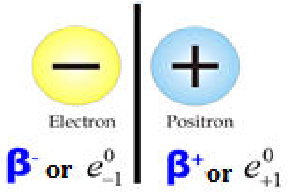
$$^{A}_{Z}X \rightarrow {}^{A-4}_{Z-2}Y + {}^{4}_{2}He$$

- ✓ Example : Equation for nuclear reaction α

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

Particule alpha

IV-2- Bêta particules:



- ✓ With the symbole (β) ;
- ✓ Positive or negative charge.

- ✓ **Réaction β**

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

- ✓ Example : Equation of the nuclear reaction β.

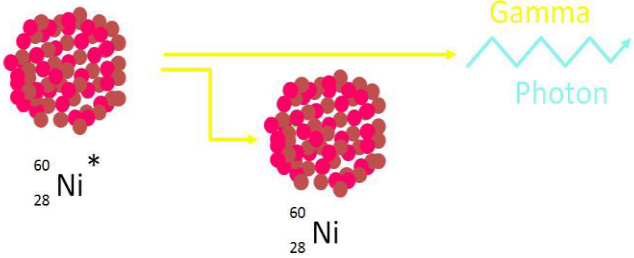
$$ {}^{14}_{6}C \rightarrow {}^{14}_{7}N + \beta^{-} $$
$$ {}^{12}_{7}N \rightarrow {}^{12}_{6}C + \beta^{+} $$

IV-3-Gamma radiation :

- ✓ In most cases, both alpha and beta radiations are accompanied by the emission of gamma radiation (γ).
- ✓ It's very penetrating.
- ✓ **Réaction γ :**

$$ {}^{A}_{Z}X^{*} \rightarrow {}^{A}_{Z}X + \gamma $$

- ✓ Example : Equation of the nuclear reaction gamma (γ).



IV-4- Neutron radioactivity :

- ✓ In addition to natural alpha, beta and gamma radioactivity, there is neutron radioactivity and electron capture.
- ✓ this type of radioactivity is very common in excited nuclei.
- ✓ **Réaction (n):**

$$ {}^{A}_{Z}X^* \longrightarrow {}^{A-1}_{Z}X + {}^{1}_{0}n $$

- ✓ Example : Equation of the nuclear reaction **n.**

$$ {}^{17}_{7}N \longrightarrow {}^{17}_{8}O^* + {}^{0}_{-1}e \qquad {}^{17}_{8}O^* \longrightarrow {}^{16}_{8}O + {}^{1}_{0}n $$

V- Radiation pathways in air:

Radiation interactions with the human body

Ionizing radiation, such as gamma rays, X-rays and some part of ultraviolet light (shortwave UVC), can ionize atoms due to its high energy content.
Damage can then occur to the DNA of our body as well as cellular changes (Fig. III.3).

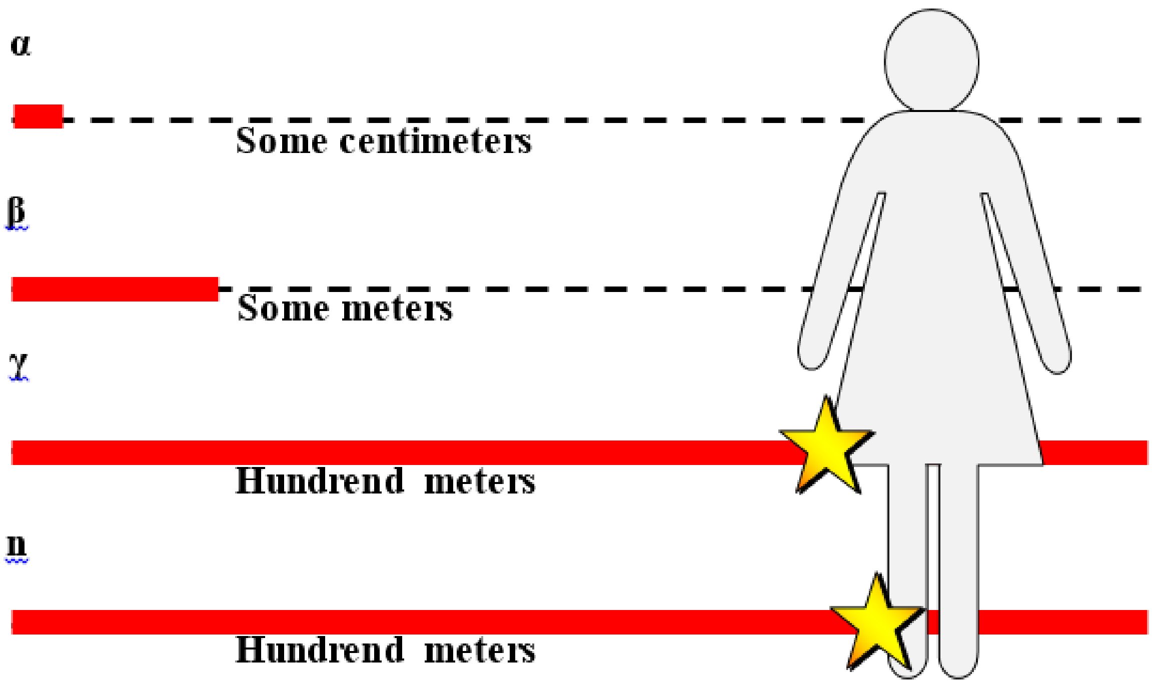


Figure III.3 : Radiation interactions with the human body

VI- The law of radioactivity :

The number of disintegrations per unit of time is proportional to the number of radioactive atoms at time t.

With : $\frac{dN}{dt} = -\lambda.N$

$\frac{dN}{dt}$: is the rate (speed) of disintegration ;

λ : is the disintegration constant ; The constant of desintégration (decay) is independent of pressure and temperature conditions

N : is the number of radioactive atoms remaining at time t.

$$We\ have: \frac{dN}{dt} = -\lambda.N \Rightarrow \frac{dN}{N} = -\lambda.dt \Rightarrow \int_{N0}^{N} \frac{dN}{N} = \int_{0}^{t} -\lambda.dt \Rightarrow [\ln]_{N0}^{N} = -\lambda.[t]_{0}^{t} \Rightarrow \ln\frac{N}{N_0} = -\lambda.t \Rightarrow N = N_0.e^{-\lambda.t}$$

Law of radioactive decay: $N = N_0 e^{-\lambda t}$ (1)

Where N_0 is the number of radioactive atoms at $t_0 = 0$.

- For each molecular weight M (g/mol) : 1mol ⟶ M ⟶ N_A

n(mol) ⟶ m(gram) ⟶ N?

We find: $N = \frac{m}{M}.N_A$ et $N_0 = \frac{m_0}{M}.N_A$ dans (1) on obtient :

Law of radioactive decay: $m = m_0 e^{-\lambda t}$ (2)

- *Half-life or time period (T ou $t_{1/2}$) :*

Time required for half of a given stock of nuclei to disintegrate (SeeTable III.1).

If : t = T and N = $N_0/2$ in (1) :

$$\frac{N_0}{2} = N_0 e^{-\lambda.T} \Rightarrow \frac{1}{2} = e^{-\lambda.T} \Rightarrow -\ln 2 = -\lambda.T$$

$$\Rightarrow T = \frac{\ln 2}{\lambda} = \frac{0.693}{\lambda} = t_{1/2} = t_{0.5} \quad (3)$$

- The table below shows examples of radioactive periods of some isotopes.

Table III.1 : radioactive periods of some isotopes

Parent	Decays into:	Half life (years)
Carbon-14	Nitrogen-14	5,730
Aluminum-26	Magnesium-26	740,000
Iodine-129	Xenon-129	17 million
Uranium-235	Lead-207	704 million
Potassium-40	Argon-40	1.3 billion
Rubidium-87	Strontium-87	49 billion

VII- The activity (A) :

Activity (A) is defined by the rate of decay, i.e. by the number of disintegrations per second :

$$A(t) = \frac{-dN}{dt} = \lambda.N = \lambda.N_0 e^{-\lambda t} = A_0.e^{-\lambda t} \Rightarrow A(t) = A_0.e^{-\lambda t} \quad (4)$$

Unity : becquerel (Bq) with : 1 Bq = **d**esintégration **p**er **s**econd (**dps**)

1 kBq = 1 000 Bq
1 MBq = 1000 000 Bq
1 GBq = 1 000 000 000 Bq

Previous unit: curie (Ci) with : 1 Ci = 3,7 . 10^{10} Bq = 3,7 . 10^{10} **dps**.

VIII- Artificial radioactivity :

-Linked to human activities such as civil or military nuclear energy.

- Artificial radioelements are obtained by bombarding stable elements (aluminum, beryllium, iodine, etc.) with neutrons, protons and helium. Historically : In 1934 Frédéric Joliot-Curie and Irène Joliot-Curie discovered this phenomenon by producing Phosphorus 30 by bombarding Aluminium 27 with an α particle usually from a Radium source. ${}^{27}_{13}Al + {}^{4}_{2}He \longrightarrow {}^{30}_{15}P + {}^{1}_{0}n$

Irène Joliot-Curie is the daughter of Pierre and Marie Curie. She and her wife Frédéric Joliot-Curie won the Nobel Prize in Chemistry in 1935 for the discovery of artificial radioactivity. Irène Joliot-Curie died on 17 March 1956 of acute leukemia related to her exposure to polonium and X-rays, the same disease that had taken her mother Marie Curie, in July 1934.

VIII-1- Mass variation and energy balance of a nuclear reaction :

Consider the following nuclear reaction : ${}_{Z_1}^{A_1}A + {}_{Z_2}^{A_2}B \longrightarrow {}_{Z_3}^{A_3}C + {}_{Z_4}^{A_4}D$

- As with any other nuclear transformation, the mass number A and the charge number Z are conserved, mass number A and charge number Z.

So : $A_1 + A_2 = A_3 + A_4$ and $Z_1 + Z_2 = Z_3 + Z_4$

- mass variation of a nuclear reaction : Δm (reaction)=∑m (Products) −∑m (Reagents)

- The energy of a nuclear reaction : $\Delta E = \Delta m.\ C^2$

VIII-2- Simplified representation of the nuclear reaction:

recall that:

$$P \equiv {}_1^1H \qquad \alpha \equiv {}_2^4He \qquad n \equiv {}_0^1n$$

$$d \equiv {}_1^2H \qquad B^- \equiv {}_{1-}^{0}e \qquad \gamma \equiv {}_0^0e$$

$$T \equiv {}_1^3H \qquad B^+ \equiv {}_{1+}^{0}e$$

Example : ${}_{13}^{27}Al + {}_2^4He \longrightarrow {}_{15}^{30}P + {}_0^1n$

We can write:

$$\underbrace{{}_{13}^{27}Al(\propto}_{\text{Reagents}},\underbrace{n){}_{15}^{30}P}_{\text{Products}}$$

Reagents Products

VIII-3- Artificial nuclear Reactions :

VIII-3-1- Nuclear transmission :

This reaction produces nuclei with an atomic mass equal to or close to that of the target nucleus.

The nucleus formed may be stable or radioactive..

Example :

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

$$^{6}_{3}Li + ^{2}_{1}H \longrightarrow ^{7}_{3}Li + ^{1}_{1}H \Leftrightarrow Li^{6}_{3}(d, P)Li^{7}_{3}$$

VIII-3-2-Fission :

- Production of energy, 2 to 3 free neutrons and 2 new nuclei (fission products).

- The fission products are BETA and GAMMA emitters.

- The neutrons released can create other fissions (chain reaction).

Example :

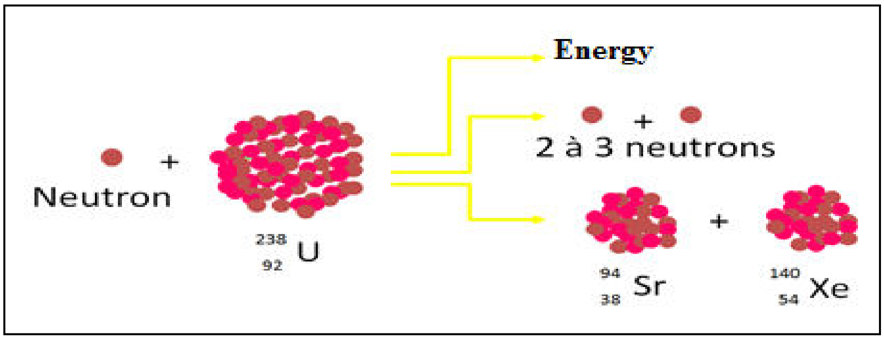


Figure III.4 : Reaction of Uranium fission

Note : one gram of uranium releases the same energy as the combustion of 1.8 tons of oil (See Fig. III.5).

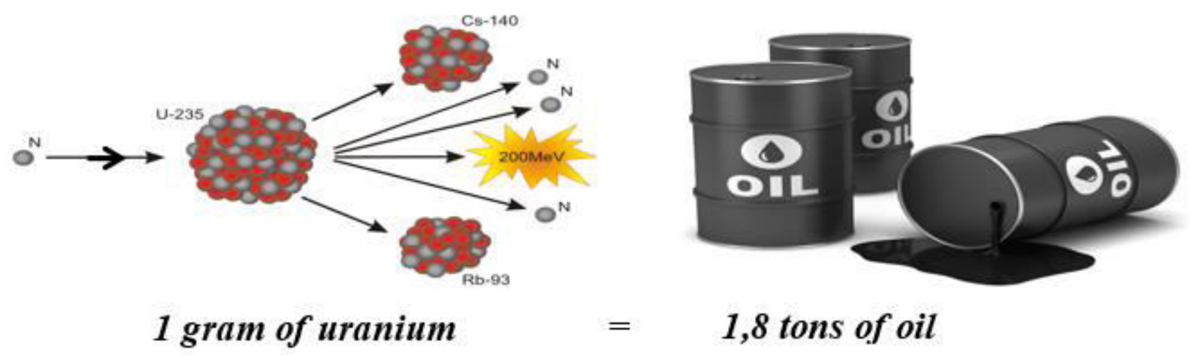


Figure III.5 : The energy released by one gram of Uranium

VIII-3-3-Fusion :

- In a high-temperature environment, 2 light bodies collide. They fuse, releasing energy.

- The energy released is greater than that released by a fission reaction.

Example :

- ✓ $D + D \rightarrow {}^{3}He + n + 3{,}3\ MeV$
- ✓ $D + D \rightarrow T + p + 4{,}0\ MeV$
- ✓ $T + T \rightarrow {}^{4}He + 2n$
- ✓ ${}^{3}He + D \rightarrow {}^{4}He + p$
- ✓ ${}^{6}Li + n \rightarrow T + {}^{4}He$
- ✓ ${}^{7}Li + n \rightarrow T + {}^{4}He + n$

-Other interesting fusion reactions include :

a-*The hydrogen bomb (See Fig. III.6)* :

$D + T \rightarrow {}^{4}He + n + 17{,}6\ MeV.$

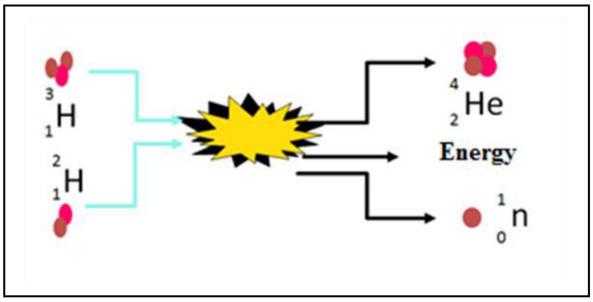


Figure III.6 : Reaction of the hydrogen fusion

- H bombs ≡ 14 000 tons of TNT

- H-bombs are 1,000 times more powerful than the Hiroshima bomb.

- The explosion would sweep away most of the buildings over an area with a radius of around 32.6 km.

- The thermal radiation would extend over an area of 73.7 km radius (3rd degree burns).

- The human toll of such an attack is estimated at 6.87 million dead and 3.94 million injured.

b-Note :

-The fusion of one gram of tritium releases the same energy as the combustion of 13.5 tons of oil.

-The tritium tube photographed here (See Fig. III.7) had been glowing for a year and a half (half-life of tritium = 12.5 years.

Figure III.7 : The tritium tube

- Example of radiation interactions with the human body (Fig. III.8) :

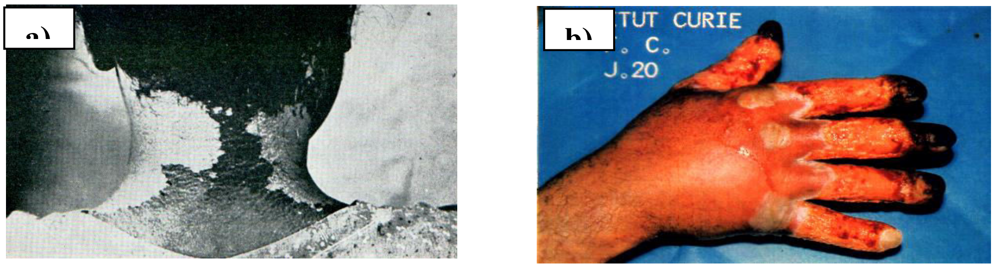


Figure III.8 : **a)**Beta burn on the neck, one month after exposure ; **b)** Skin contact with cobalt 60

<u>IX- Application of radioactivity : C14 dating :</u>

<u>Principle of carbon-14 dating :</u>

- 14C its abundance in the atmosphere is constant : $\frac{N_{^{14}C}}{N_{^{12}C}} = 1.3\times10^{-12}$

- When a living organism dies, its 14C abundance begins to fall as it no longer absorbs CO2 from the atmosphere.

We can calculate the dating of a substance (bone, wood) by the following equation :

$$We\ have: A(t) = A_0.e^{-\lambda t} \Rightarrow t = \frac{\ln(\frac{A_0}{A_t})}{\lambda} = \frac{\ln(\frac{A_0}{A_t})}{0.000124487\ ans^{-1}} = \ln(\frac{A_0}{A_t}).8033\ ans = age$$

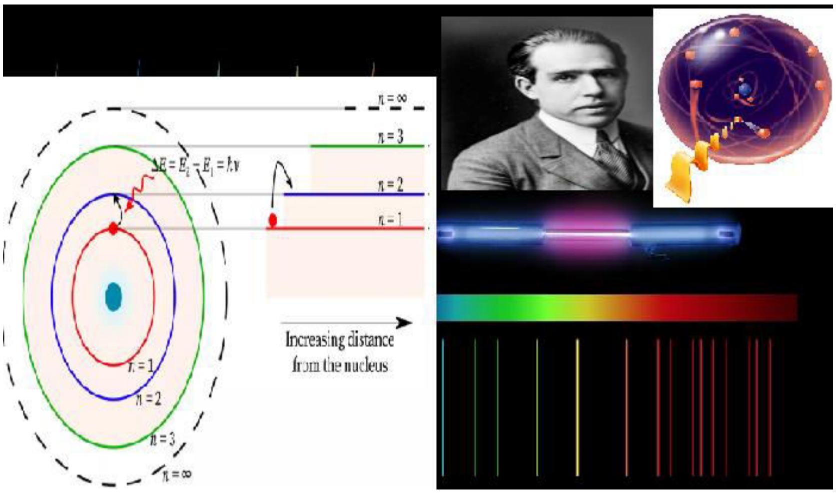
T : period of C14=5568 years $\Rightarrow \lambda = \frac{\ln 2}{T} = 0.000124487 \quad ans^{-1}$

A_t : carbon activity of the sample (archaeological sample) ;

A_0 **:** modern carbon activity (standard reference sample).

Chaptre IV

Structure of the atom



I-Introduction :

I-1- Classical and quantum mechanics :

- Classical (Newtonian) mechanics describes the movement of macroscopic objects when their speed is low compared with that of light.

- Quantum mechanics studies physical systems at the atomic and subatomic scales (<atom).

- It was developed at the beginning of the 20th century by a dozen American and European physicists in order to solve various problems that classical physics had failed to explain, such as the existence of spectral lines, the photo-electric effect, or blackbody radiation.

I-2- Blackbody radiation :

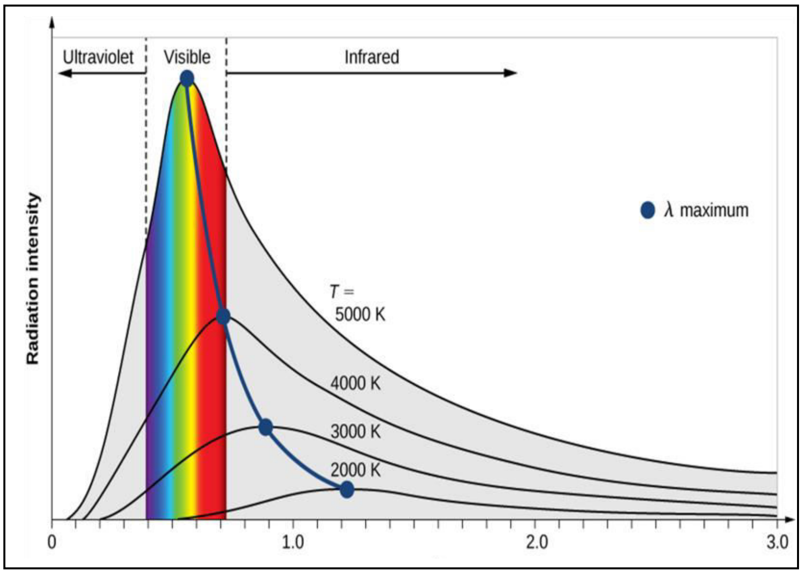
- Everyone has had the experience of heating a piece of metal.

- A black body radiates its energy in "bursts" rather than continuously (Fig. VI.1) :

If heated to : -500 C^0 its color turns red ;

-1000 C^0 its color turns yellow ;

-1200 C^0 its color turns white.



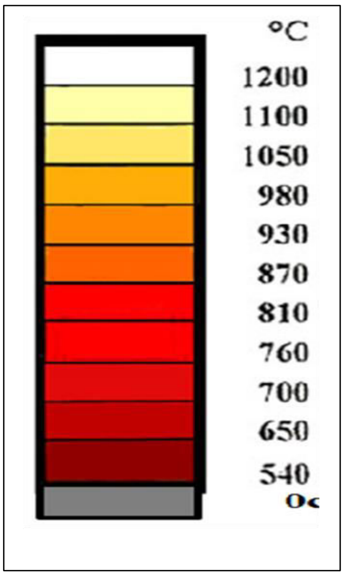


Figure VI.1 : Radiation from a black body (a piece of iron)

-Planck pense que ces sauts sont une propriété «interne» des atomes.

II- Quantum energy and photon :

- Max Planck (1900) concluded that : a light source does not emit electromagnetic waves continuously, but exchanges of light exchange of light energy can only take place in discontinuous in discontinuous "packets" or quanta (Quantum) (Fig. VI.2).

- In 1905 Einstein hypothesized that these "packets" of electromagnetic energy electromagnetic energy are carried by photons (particles of particles of zero mass that travel at the speed of light ($c = 3.00 \times 10^8$ m / s in a vacuum).

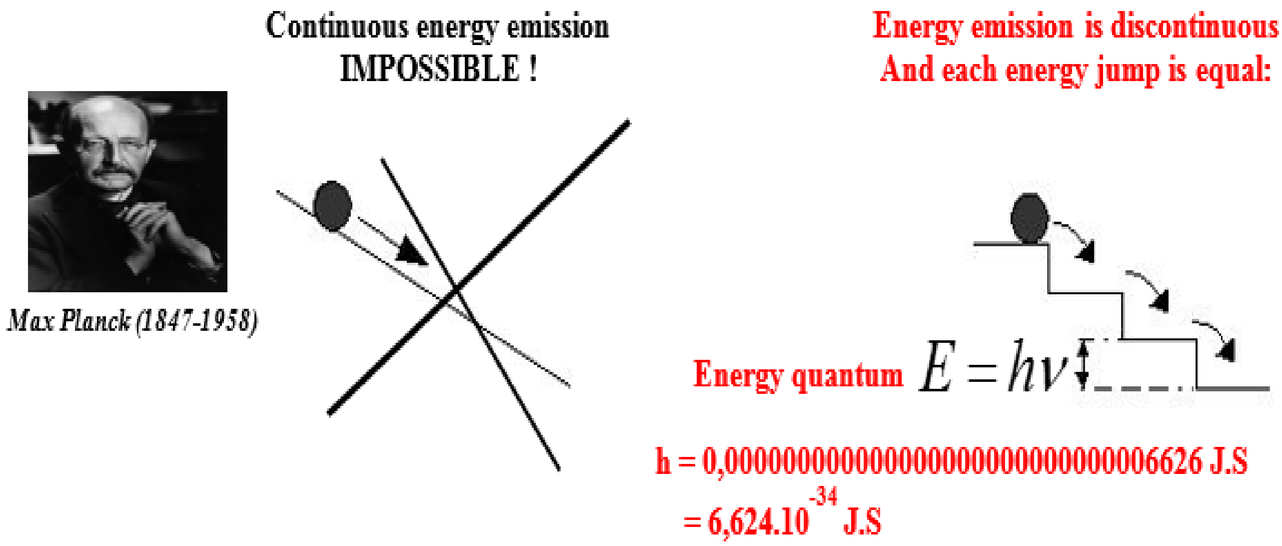


Figure VI.2 : Quantum energy according to Planck

- The energy of a photon associated with an electromagnetic wave of frequency **ν** and wavelength **λ** in vacuum is :

QUANTUM=Energy of one PHOTONS=Packet of energy= $E = h\nu = \frac{hc}{\lambda}$

h : Planck constant $=6,624.10^{-34}$ J.s ; *c* : speed of light $=3.10^8$ m/s.

III- Wave-corpuscle duality :

III-1- The light :

Light is both a undulatory phenomenon, hence the concept of wavelength λ, and a corpuscular phenomenon, as evidenced by photons (mass and speed).

$$E = h\nu = mc^2 \qquad \Rightarrow h.\frac{c}{\lambda} = mc^2 \quad \Rightarrow \lambda = \frac{h}{mc}$$

- Amplitude (A) : is the maximum height reached by the wave from the resting position.

- Frequency (ν) : is the number of cycles per unit of time.

- It is measured in Hertz ($Hz=s^{-1}$). $\nu = c/\lambda$

- Wavelength (λ) : distance between two successive maxima (Fig. IV.3) :

- Wave number : $\bar{\nu} = \frac{1}{\lambda}$

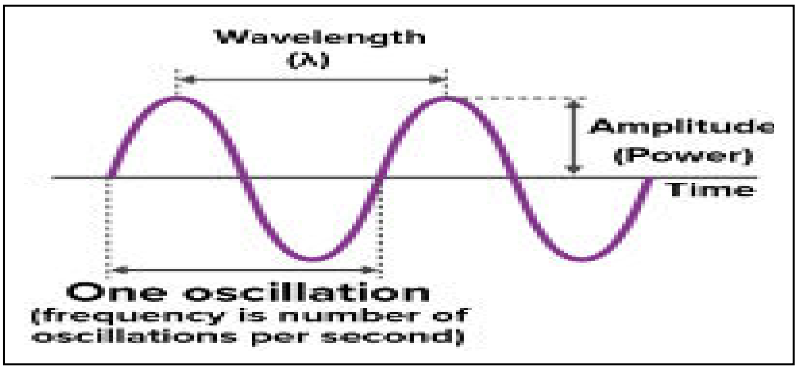


Figure IV.3 : Wave-corpuscle duality

III-2- Matter (Louis de Broglie relation) :

In 1924, de Broglie proposed that if light had a wave/particle duality, why not matter too?

In 1925, de Broglie generalized the previous relationship $\lambda = \frac{h}{m.c}$ any particle of mass "m", moving

at speed "V" is associated with a wave, of wavelength "λ", defined by the relation : $\lambda = \frac{h}{m.V}$

λ= Wavelength =h/p with : p= motion quantity = m.v

IV- The photoelectric effect (Einstein 1905) :

- When the plate is illuminated with an (intense) red light, nothing happens ; If you shine a blue light, a current flows : (Fig. IV.4).

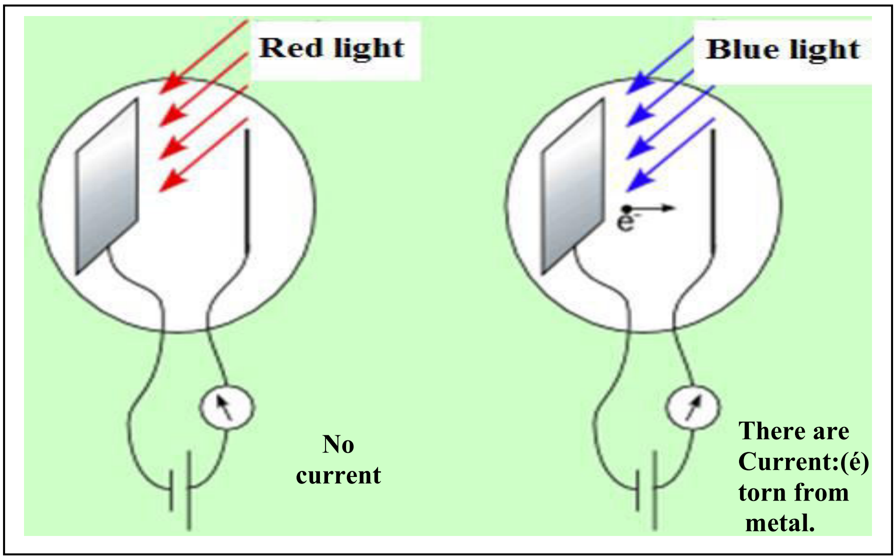


Figure IV.4 : The photoelectric effect (Einstein 1905)

Explication :

- We can't explain this effect with a wave : the energy of a wave is proportional to its intensity!
- On the other hand, if we consider that light is made up of small particles (Fig. IV.5) :

- ✓ At low frequency (red), the energy of each particle is too small to pull an electron from the plate ;

- ✓ At a higher frequency (blue), the energy exceeds the minimum necessary to pull out an electron, resulting in an electric current.

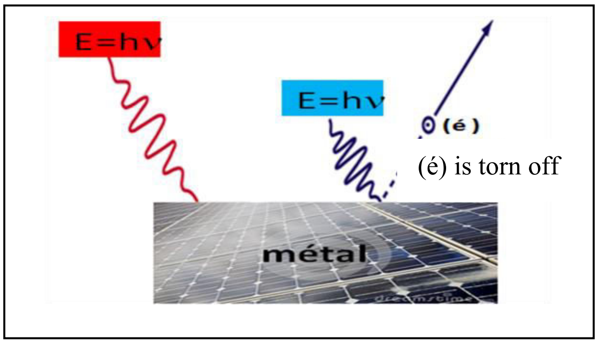


Figure IV.5 : Explanatory diagram of electron extraction

- Albert Einstein explained the photoelectric effect in this way, and proposed the name photon for this tiny speck of light. The result is a simple equation :

Photon energy=Extration energy (threshold)+kinetic energy of the electron

$$\Rightarrow E = E_0 + E_c \Rightarrow h\nu = h\nu_0 + \frac{1}{2} m_e V^2$$

- The photoelectric effect requires : $E > E_0 \Rightarrow h\nu > h\nu_0 \Rightarrow \nu > \nu_0 \Rightarrow \lambda_0 > \lambda$

<u>V- Electromagnetic spectra :</u>

<u>V-1-The visible spectrum :</u>

- The human eye is only sensitive to a very small part of this spectrum : the visible spectrum (Fig. IV.6).

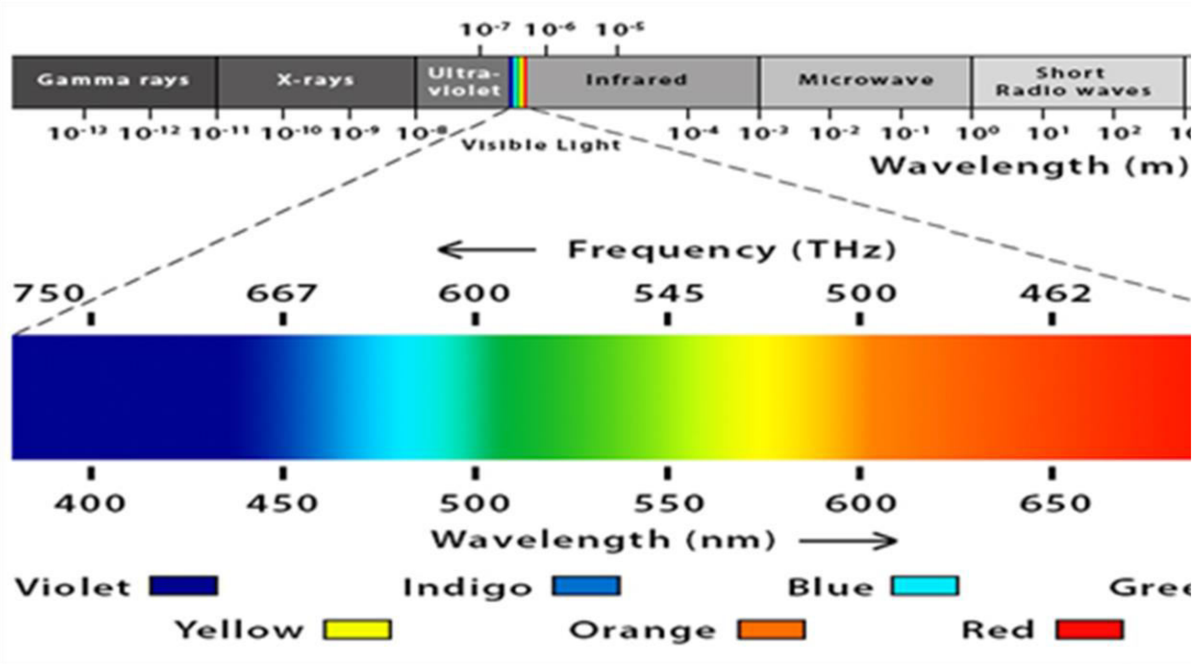
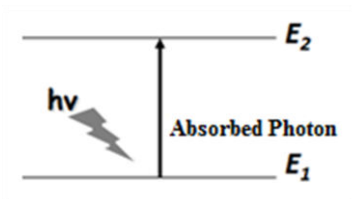


Figure IV.6 : Logarithmic scale of electromagnetic waves.

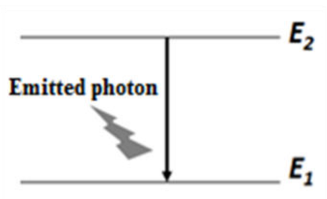
V-2- Absorption and emission spectra :

- The spectrum is the set of photons absorbed or emitted by a substance.

- A substance absorbs or emits a photon (Fig. IV.7) when one of its particles passes from one energy level to another (jumps).



Absorption



Emission

Figure IV.7 : Diagram of an electronic transition

-To study these jumps, simply measure the spectra of light emitted by atoms that absorb or emit energy.

V-2-1-Absorption spectra :

By placing hydrogen in front of a white light source, we can obtain an absorption spectrum of the element with black lines (discontinuous spectrum) (Fig. IV.8).

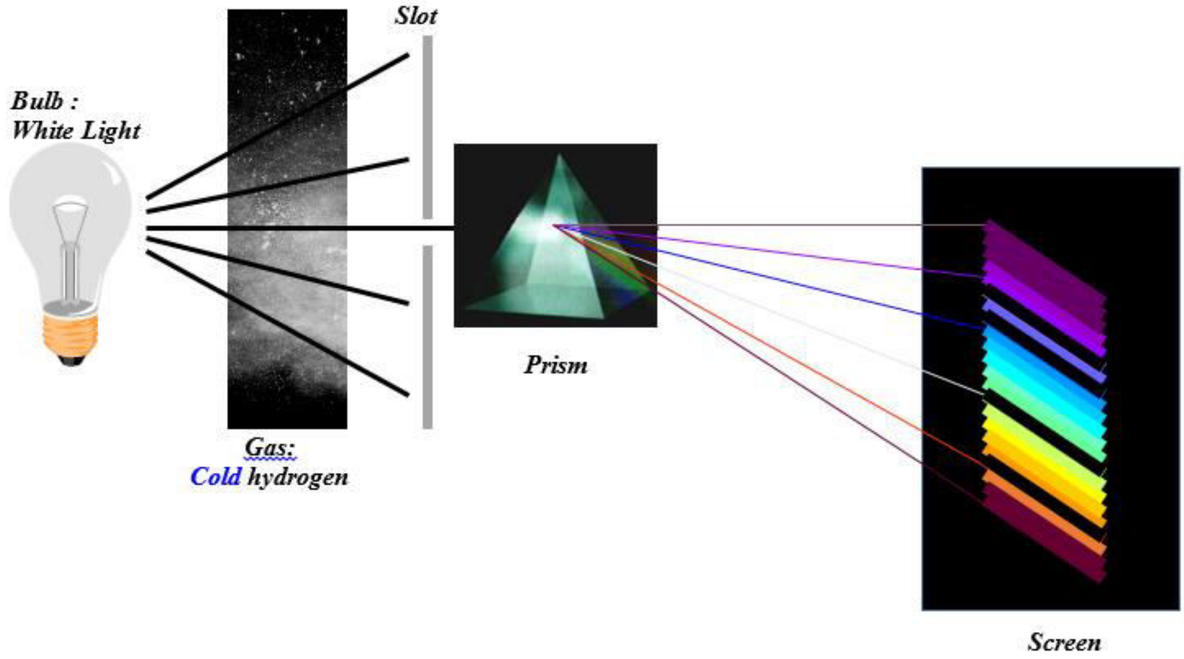


Figure IV.8 : Absorption spectra (discontinuous spectrum)

V-2-2- Emission spectra :

The H emission spectrum is made up of 4 fine lines with specific wavelengths λ : this is a (non-continuous) line spectrum. (Fig. IV.9).

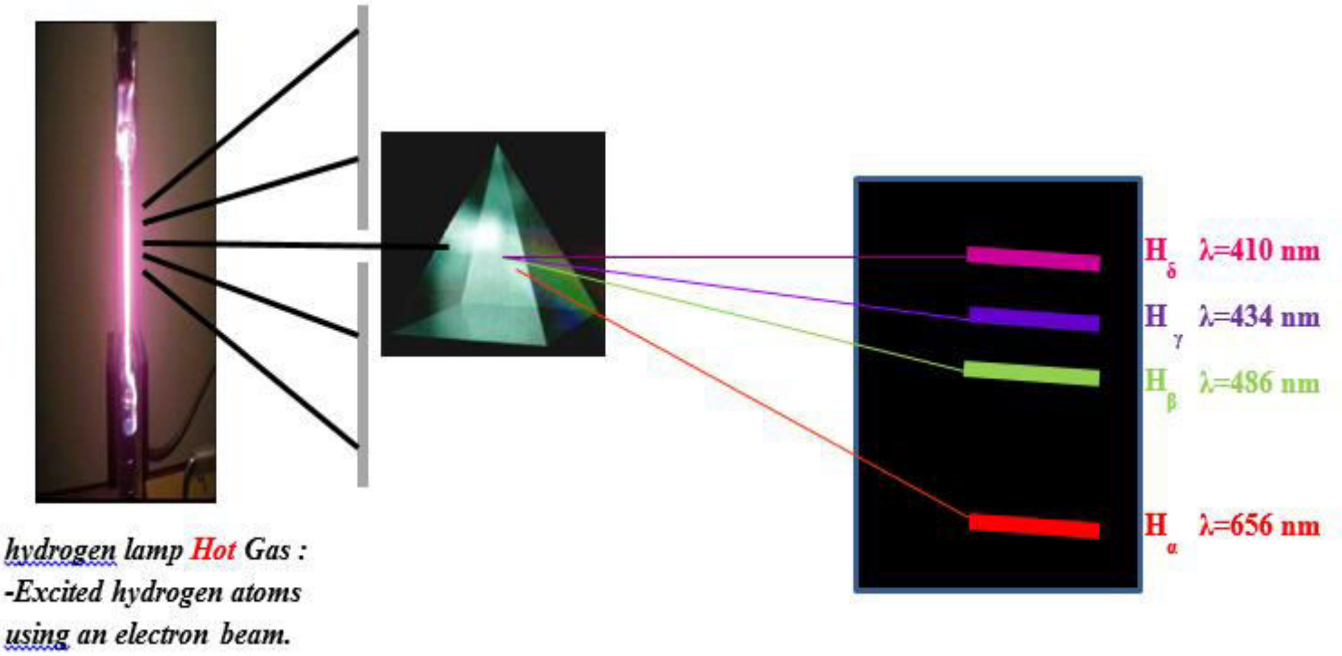


Figure IV.9 : Emission spectra (line spectrum).

- Remark : Each atom has its own electronic signature, which is characteristic of the jumps its electrons can make (Fig. IV.10).

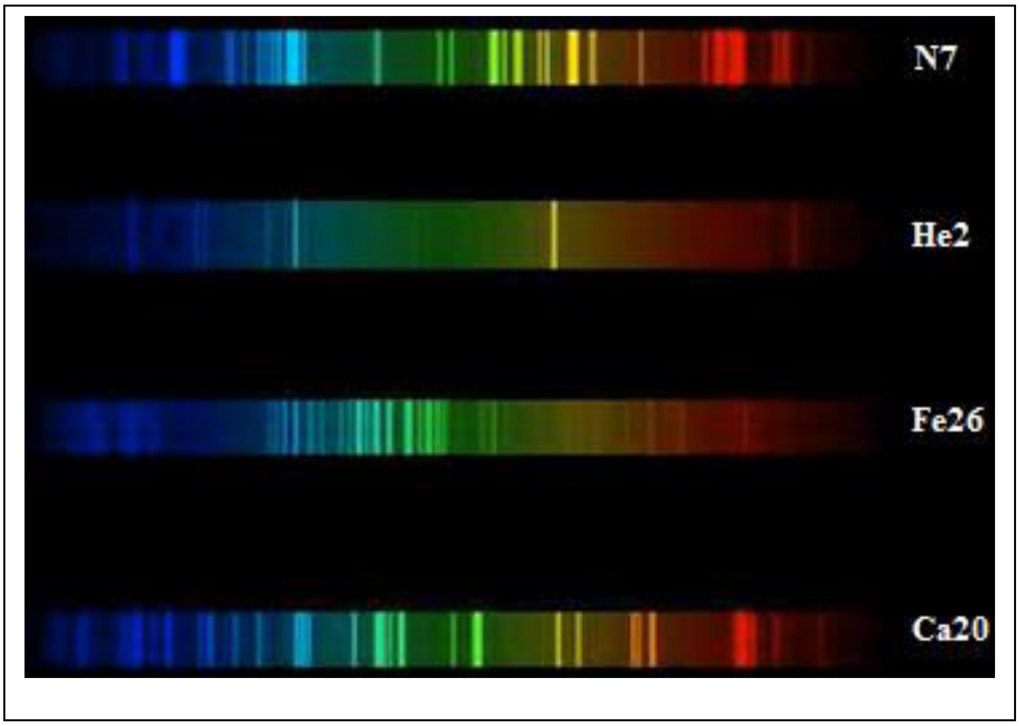


Figure IV.10 : emission line spectra of nitrogen, helium, iron and calcium

V-3- Balmer Relation :

Balmer's formula combines the wavelengths of the spectral lines of the hydrogen atom in the visible. They correspond to transitions between excitation levels n_f greater than 2 (Fig. IV.11).

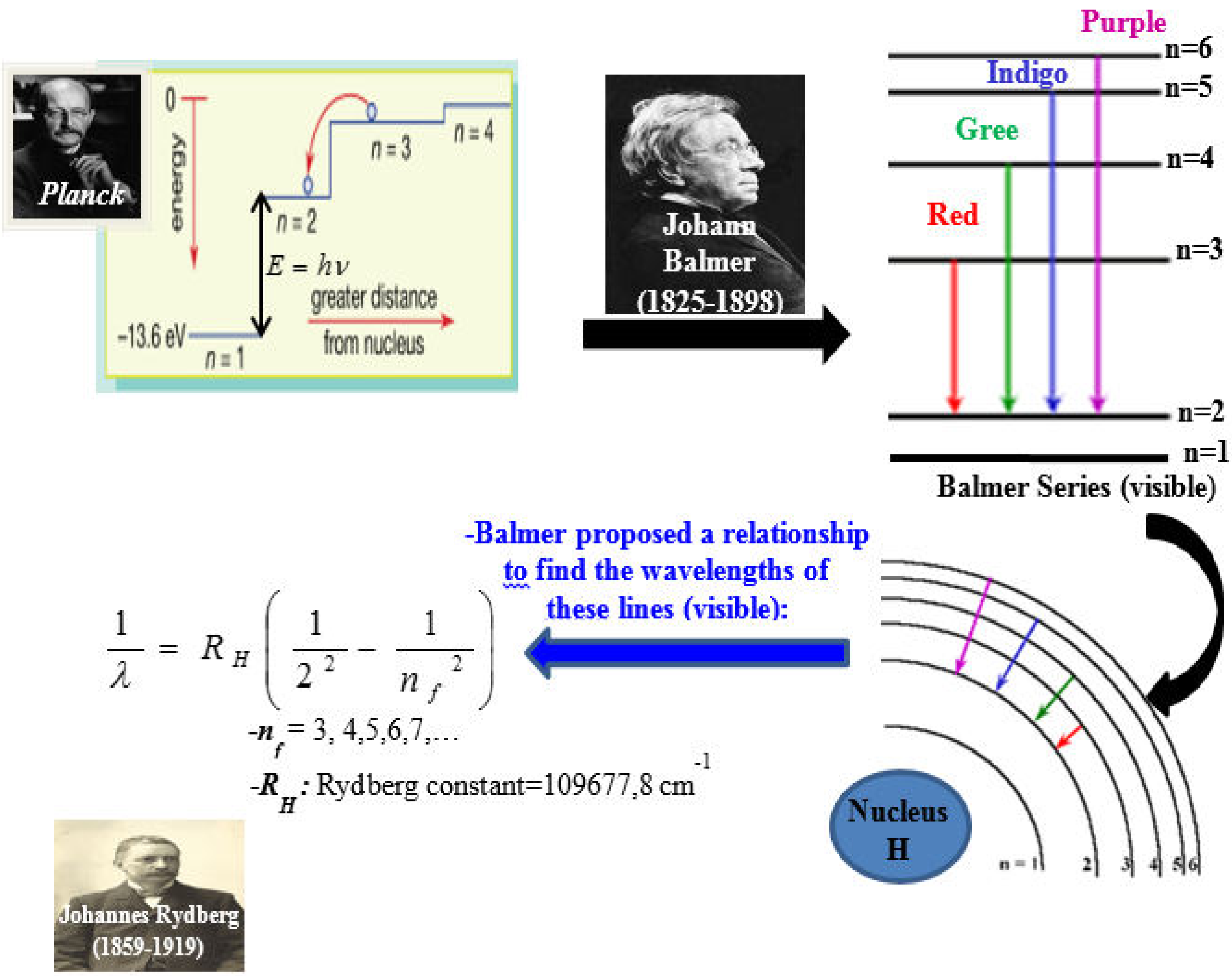


Figure IV.11 : Explanatory diagram of the Balmer relationship

VI- Other spectrum series :

VI-1- Ritz Formula :

Ritz generalized Balmer's formula to all lines in the atomic emission spectrum of hydrogen :

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

With : n_i and n_f are integers with $n_i < n_f$.

Table IV.1 shows the different series and their corresponding spectral domains.

Tableau IV.1 : Series of hydrogen transition spectra and corresponding region.

Série	n_i (letter)	n_i (number)	n_f	Region
Lyman	K	1	2, 3, 4, 5…	UV
Balmer	L	2	3, 4, 5, 6…	Visible
Pashen	M	3	4, 5, 6, 7…	Near IR
Brackett	N	4	5, 6, 7, 8…	far IR
Pfund	O	5	6, 7, 8, 9…	far IR

VI-2- Spectral ray :

We have : $\frac{1}{\lambda} = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$

1st ray (α) : $n_i \rightarrow n_f = (n_i+1)$

2nd ray (β) : $n_i \rightarrow n_f = (n_i+2)$

3rd ray (γ) : $n_i \rightarrow n_f = (n_i+3)$

Limited ray: $n_i \rightarrow n_f = (n_i+\infty)$

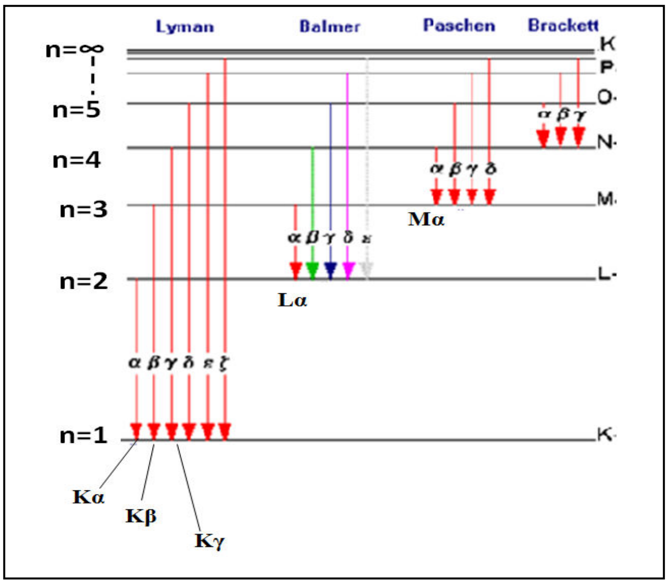


Figure IV.12 : The spectral ray of the hydrogen atom

Exemple :

-1st Lyman ray (K_α) $\Rightarrow n_i = 1 \rightarrow n_f = 2$

-2nd Lyman ray (K_β) $\Rightarrow n_i = 1 \rightarrow n_f = 3$

-3rd Lyman ray ($K\,\gamma$) $\Rightarrow n_i = 1 \rightarrow n_f = 4$

VI-3- Calculation of energy ΔE:

ΔE can be calculated by the following equation:

$$\Delta E = h\nu = h\frac{c}{\lambda} = h.c.R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

1^{st} Excited state : $n_i \rightarrow n_f = n_i+1$

2^{nd} Excited state : $n_i \rightarrow n_f = n_i+2$

3^{rd} Excited state : $n_i \rightarrow n_f = n_i+3$

Ionisation state : $n_i \rightarrow n_f = n_i+\infty$

-We know that the shorter the wavelength, the greater the energy:

$$\Delta E \rangle\rangle\rangle 0 \Rightarrow \lambda \langle\langle\langle 0 \Rightarrow transiton\ (n_i = 1 \rightarrow n_f = \infty)$$

- We know that the longer the wavelength, the lower the energy:

$$\Delta E \langle\langle\langle 0 \Rightarrow \lambda \rangle\rangle\rangle 0 \Rightarrow transiton\ (n_i = 1 \rightarrow n_f = 2)$$

VI-4- Hydrogen energy diagram :

The energy level diagram (Fig. IV.13) shows the possible energy states for the different atoms.

Example : H

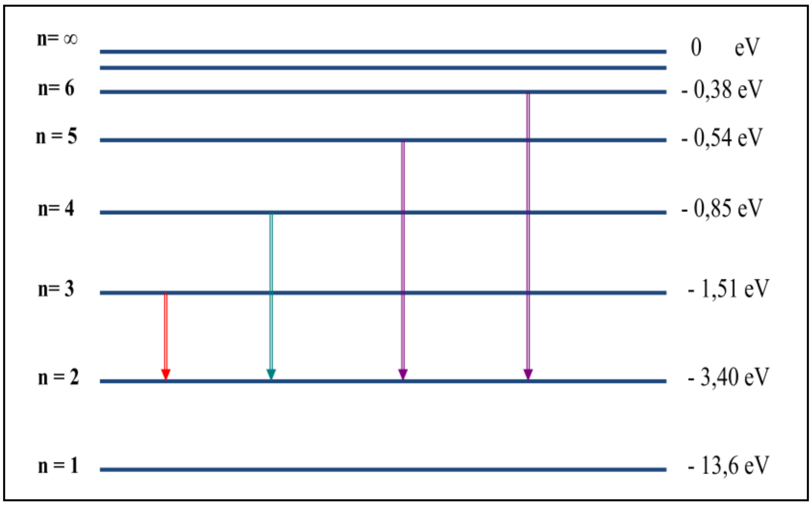


Figure IV.13 : Energy diagram of hydrogen

VI-5- Ionization energy (ΔEi) :

ΔEi : is the energy an atom needs to extract an electron.

$$H \rightarrow H^{+} + 1\acute{e}$$

1^{st} calculation method ΔEi :

Ionisation state ($n_i \rightarrow n_f = n_i+\infty$) $\Rightarrow \Delta E_i = h\nu = h\frac{c}{\lambda} = h.c.R_H\left(\frac{1}{1^2} - \frac{1}{\infty^2}\right) = 13{,}6\ eV$

2^{nd} calculation method ΔEi :

$$\Delta E_i = \Delta E_\infty - \Delta E_1 = \left(-\frac{1}{\infty^2}.13{,}6\right) - \left(-\frac{1}{1^2}.13{,}6\right) = 13{,}6\ eV$$

VII-The Bohr atom :

Bohr's model is based on 3 postulates :

-The first mechanical postulate : The electron moves only in certain circular orbits of radius r called "stationary states".

-The second optical postulate : Radiation is emitted only if the electron moves from a higher orbit to a lower one. $\Delta E = h\nu = h\frac{c}{\lambda}$

and ΔE : final energy(f) - initial energy(i)

-The third postulate of angular momentum : The electron's angular momentum can only take on integer values that are multip $\hbar$ of : .

$$mvr = n\hbar = \frac{nh}{2\pi}$$

$\hbar = \frac{h}{2\pi}$: Planck's reduced constant ;

From these 3 postulates, Bhor was able to calculate the radii of circular orbits and the possible energies of é.

VII-1-Orbit radius calculation :

The forces applied to the nucleus in the hydrogen atom are (figure IV.14) :

Attraction Force (Coulomb Law) : $F_e = -k\frac{e^2}{r^2}$

Centrifugal forces : $F_C = \frac{mV^2}{r}$

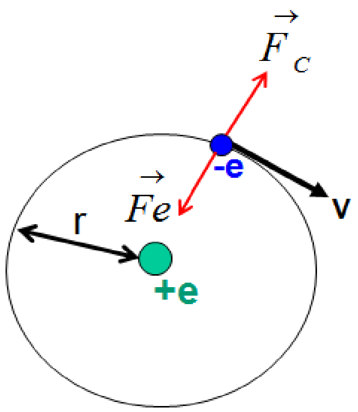


Figure IV.14 : The forces applied in the hydrogen atom.

To prevent the é from falling onto the nucleus, it is necessary to :

$$|F_C| = |F_e| \Rightarrow \frac{mV^2}{r} = k\frac{e^2}{r^2} \Rightarrow \quad mV^2 = k\frac{e^2}{r} \quad (1)$$

We have Bohr hypothesis : $mVr = n\frac{h}{2\pi} \Rightarrow V = n\frac{h}{2\pi mr}$ (2)

(2) in (1) :

$$r_n = \frac{h^2 n^2}{4\pi^2 ke^2 m} = 5{,}29\ .10^{-11} .n^2 (m) = 0{,}529\ .n^2 (A^0) \quad (3)$$

Given : K : Coulomb constant = 9. 10^9 MKSA

m: Mass of the é = 9,1 . 10^{-31} Kg

h: Planck's constant = 6,62 . 10^{-34} j.s

e: charge of é= 1,6 . 10^{-19} c

VII-2- Calculating the electron's total energy :

Energie conservation : E = E kinetic + E potential

$$E = \frac{mV^2}{2} + k\frac{-e^2}{r} \quad (4)$$

(1) in (4) $\rightarrow E = \frac{k\frac{e^2}{r}}{2} + k\frac{-e^2}{r} = -\frac{1}{2}k\frac{e^2}{r} \quad (5)$

We have :

$$r_n = \frac{h^2 n^2}{4\pi^2 ke^2 m}$$

(3) in (5):

$$E_n = \frac{-2\pi^2 k^2 e^4 m}{h^2}\frac{1}{n^2}$$

$$\Rightarrow E_n = -2{,}17.10^{-19}\frac{1}{n^2}(j) = -13.6\frac{1}{n^2}(eV) \quad (6)$$

So the electron's total energy is quantized : n is the first quantum number, called the principal quantum number.

VII-3- Calculating the electron speed :

$$mVr = n\frac{h}{2\pi} \Rightarrow V = n\frac{h}{2\pi mr} \quad (3)$$

$$r_n = \frac{h^2 n^2}{4\pi^2 ke^2 m} \quad (2)$$

$$(3) \text{ in } (2) \Rightarrow V_n = n\frac{h}{2\pi mr} = \frac{2\pi ke^2}{h}\frac{1}{n} = 2{,}18.10^6.\frac{1}{n}\left(m.s^{-1}\right) \quad (7)$$

Exercise : calculate the radius and energy of orbits 1, 2, 3 ……….

We find:

n	=	1	2	3	4	5	6	7
n^2	=	1	4	9	16	25	36	49
r_n	$= n^2 (0{,}528)$ Å	0,528	2,11	4,75	8,40	13,28	19,0	25,87
E_x	$= - 13{,}6 / n^2$ eV	- 13,6	- 3,4	- 1,5	- 0,85	- 0,54	- 0,37	- 0,27

VII-4- Expression of Rydberg constant (R_H):

-A/ When é absorbs radiation (energy), it passes from level n_1 to another level, n_2.

We have :

$$\Delta E = E_2 - E_1 = h\nu = h\frac{c}{\lambda} = hc\frac{1}{\lambda}$$

$$et : \Delta E = \frac{-2\pi^2 k^2 e^4 m}{h^2}\frac{1}{n_2^2} - \left(\frac{-2\pi^2 k^2 e^4 m}{h^2}\frac{1}{n_1^2}\right)$$

$$\Rightarrow \Delta E = \frac{2\pi^2 k^2 e^4 m}{h^2}\left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right) = h\nu = h\frac{c}{\lambda} = hc\frac{1}{\lambda}$$

$$\Rightarrow \frac{1}{\lambda} = \frac{2\pi^2 k^2 e^4 m}{h^3 c}\left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right)$$

$$et\ on\ a : \frac{1}{\lambda} = R_H\left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right)$$

$$donc\ R_H = \frac{2\pi^2 k^2 e^4 m}{h^3 c} = 109500 cm^{-1}$$

-B/ When an é emits radiation (energy), it transits from level n_2 to level n_1.

$$On\ a : \Delta E = E_1 - E_2$$

$$et : \Delta E_{émission} = -\frac{2\pi^2 k^2 e^4 m}{h^2}\left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right) = -\Delta E_{absorption}$$

<u>VIII- Application of Bhor theory on hydrogenoids :</u>

An hydrogenoid is an atom having one electron. It then a structure similar to hydrogene atom. It is therfore a cation from which all but one electron has been stripped.

Exemple : ${}^{9}_{4}Be^{3+}$ ${}^{4}_{2}He^{+}$

<u>VIII-1-Calculating the orbit radius</u> :

The forces applied to the nucleus in the hydrogen atom are (figure IV.15) :

Attraction force : $F_e = -k\frac{Ze^2}{r^2}$

Centrifugal forces : $F_C = \frac{mV^2}{r}$

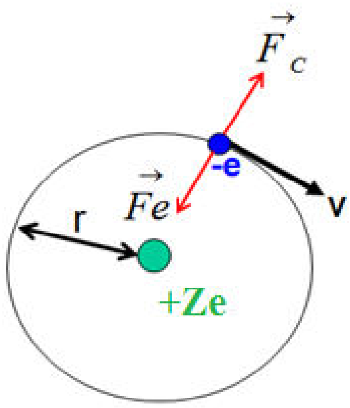


Figure IV.15 : Diagram of the hydrogenoids atom

To prevent the e from falling onto the core, it is necessary for :

$$|F_C| = |F_e|$$

*Same calculation as part (**VI-1**) : we replace Z by Ze :*

$$\Rightarrow r_n = \frac{h^2 n^2}{4\pi^2 kZe^2 m} = 0.529.\frac{n^2}{Z}\left(A^0\right)$$

VIII-2- Electron energy calculation :

Energy Conservation : E = E kinetic+E potential

$$\Rightarrow E = \frac{mV^2}{2} + k\frac{-Ze^2}{r}$$

*Same calculation as part (VI-2) : replace Z by Ze **:***

$$\Rightarrow E_n = \frac{-2\pi^2 k^2 Z^2 e^4 m}{h^2}\frac{1}{n^2} = -13{,}6.\frac{Z^2}{n^2}(eV)$$

We have :

$$\Delta E = E_2 - E_1 = \frac{2\pi^2 k^2 Z^2 e^4 m}{h^2}\left(\frac{1}{n_1{}^2} - \frac{1}{n_2{}^2}\right)$$

$$et \Delta E = h\nu = h\frac{c}{\lambda} = hc\frac{1}{\lambda}$$

So $\frac{1}{\lambda} = R_H Z^2\left(\frac{1}{n_1{}^2} - \frac{1}{n_2{}^2}\right)$

IX- Application of Bhor theory on polyelectronic atoms :

IX-1-Generalized Balmer Formula :

Same calculation, but we replace Z by Zeffectif

We found that :

1) $r_n = 0,529 . \frac{n^2}{Z_{eff}} (A^0)$

2) $E_n = -13,6 . \frac{Z^2{}_{eff}}{n^2} (eV)$

3) $\frac{1}{\lambda} = R_H Z^2{}_{eff} \left(\frac{1}{n_1{}^2} - \frac{1}{n_2{}^2}\right)$

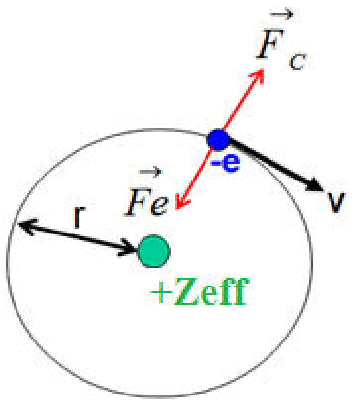


Figure IV.16. Schematic diagram of a

IX-2- Calculation of Z effectif (Zeff) according to Slater :

In a polyelectronic atom, the **other electrons** form a screen between the **nucleus** and the **electron under study** (Figure IV.17).

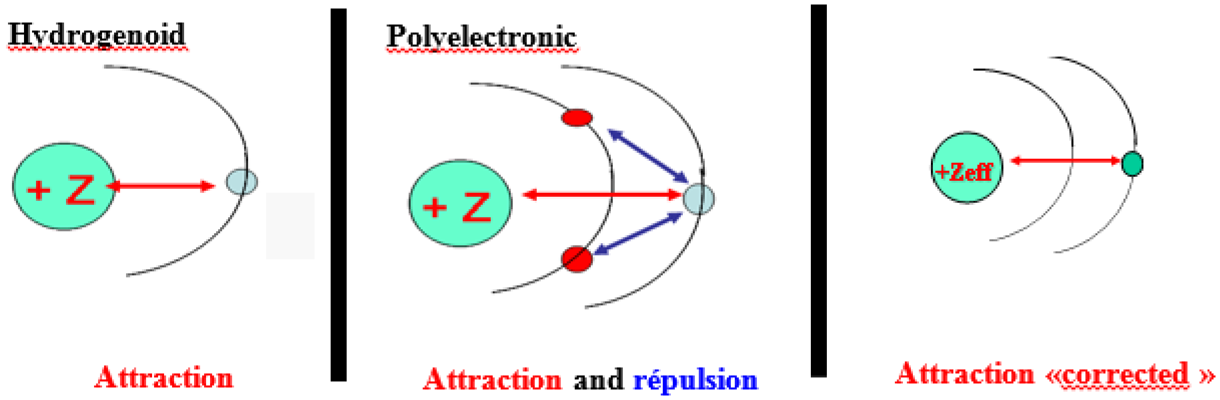


Figure IV.17 : Diagram explaining the screen effect

Modèle of Slater ≡ Attraction «corrected » So the real charge Z is replaced by the charge Zeff.

We have : Zeff = Z-δ

δ: The screen constant depende on the number of electrons located between the considered electron and the nucleus. It is calculated from Slater rule.

IX-3- Comparison with Moseley law :

- The analysis of the emission rays of the X-ray spectrum of a certain number of elements allowed Moseley to establish an empirical law linking the frequency of the emitted radiation v and the atomic number Z

$$On\ a\ la\ loi\ de\ Moseley : \sqrt{\nu} = a.Z_{eff} = a.(Z-\delta)$$

$$\Rightarrow \nu = a^2.Z_{eff}^2 \Rightarrow \frac{c}{\lambda} = a^2.Z_{eff}^2 \Rightarrow \frac{1}{\lambda} = \frac{a^2}{c}.Z_{eff}^2 \quad (1)$$

$$On\ a : \frac{1}{\lambda} = R_H Z_{eff}^2\left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right) \quad (2)$$

Let's compare (1) and (2) $\Rightarrow a = \sqrt{c.R_H\left(\frac{1}{{n_1}^2} - \frac{1}{{n_2}^2}\right)}$

X-Heisenberg uncertainty principle :

If a particle like an electron has an important undulatory behavior, how can we describe its movement ?

Heisenberg has proposed the uncertainty principle: "It is impossible to know simultaneously and with certainty the speed and exact and exact position of a particle.

$$\Delta X.\Delta P \geq \frac{h}{2\pi} \quad (1)$$

$$On\ a : P = m.V \Rightarrow \Delta P = \Delta(m.V) = m.\Delta V$$

in (1) $\Rightarrow \Delta X.\Delta V \geq \dfrac{h}{2.\pi.m}$

The uncertainty relationship also applies to the time-energy couple :

$$\Delta E.\Delta t \geq \frac{h}{2\pi}$$

XI- Failure of the Bohr model :

The disadvantages of this model include :

- Explaining the emission spectra of atoms having several electrons.

- Using a single quantum number (n).

- Using the classical mechanics and the mechanics simultaneously,

- Simple model not explaining certain observations : for example Zeeman effect.

XI-1- Zeeman effet :

The subdivision of the energy levels of atoms immersed in a magnetic fiels.

Sommerfeld interpreted this new phenomenon (Zeeman effet). Sommerfeld modifies the Bohr of the atom by adding two degrees of freedom on the electron :

1-The possibility of elliptical orbits (l = orbital quantum number) ;

2-The possibility for these orbits to change trajectoire in the presence of a magnetic field (m = magnetic quantum number).

XII- Quantum numbers:

Explains the behavior of the electron in a atom, that is to say:

Its energy, its movement around the necleus and the shape of the orbit are described by specifying its 4 quantum numbers.

The four quantum numbers are: n, l m et s.

XII-1- The principal quantum number (the layer) (n):

n value determines the orbital energy; determines the average distance between an electron in a given orbit and the nucleus.

$$We\ have : n \geq 1$$

if n = 1→ The layer K
if n = 2→ The layer L
if n = 3→ The layer M
…….

XII-2- The secondary quabtum number (the sublayers) (l):

indicates the shape (the geometry) of the orbita.

$$We\ have: 0 \leq l \leq n-1$$

if n = 1, l=0 → la sous-couche S
if n = 2, l=0, 1→ la sous-couche S, P
if n = 3, l=0, 1, 2 → la sous-couche S, P, d
if n = 4, l= 0, 1, 2, 3 → la sous-couche S, P, d, f
…….

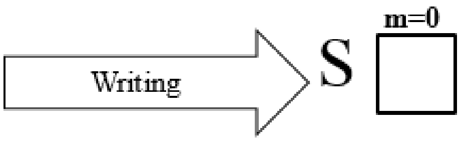
XII-3- The magnetic quantum number (m) :

m determines the orbital orientation.

$$We\ have : -l \leq m \leq +l$$

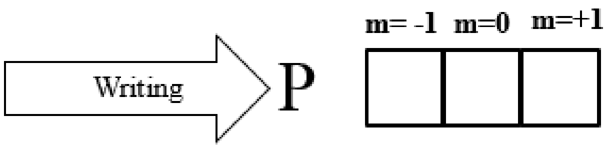
If : l = 0 (S) ⇒ **m = 0**→1 orientation only →

→1 quantum box → (contains maximum 2é)



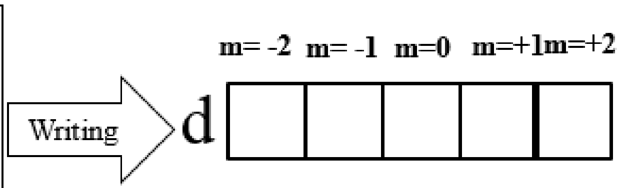
If : l = 1(P) ⇒ **m = -1, 0, +1**→ 3 orientations→

→3 quantic boxes→ (contains maximum 6é)



If : l = 2 (d) ⇒**m = -2, -1, 0, +1, +2** → 5 orientations

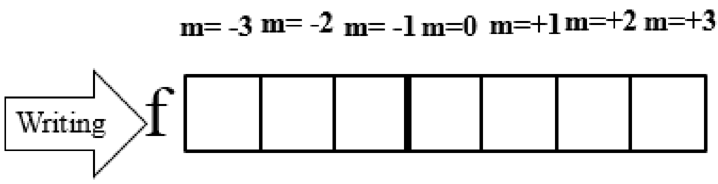
→5 quantic boxes→ (contains maximum 10é)



If: l = 3⇒ **m = 3-, -2,-1, 0, +1, +2, +3**

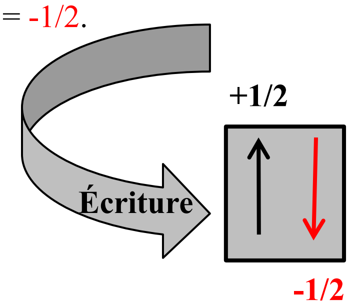
→ 7 orientation

→7 quantic boxes→ (contains maximum 14 é)



XII-4- The spin magnetic number (s) :

La valeur (s) détermine l'orientation de l'électron sur lui –même. Ces valeurs sont : s = +1/2 et s = -1/2.



XIII- Electronic capacity (nombre Max number of é in a layer) :

The Max number of (e^-) in a layer is based on the distribution shown in Table IV.2 :

Tableau IV.2 : Max number of (e^-) in a layer.

Shell number (*n*)	Number of subshells	Subshell symbol	Maximum number of electrons in subshell	Maximum number of electrons in shell = $2(n)^2$
1	1	*s*	2	2
2	2	*s*	2	8
		p	6	
3	3	*s*	2	18
		p	6	
		d	10	
4	4	*s*	2	32
		p	6	
		d	10	
		f	14	

Remarque : So, we can clearly see that it is no longer a question of describing electrons as "planets" or according to an organization in successive layers as was still the case in Bohr's atom. But it is about orbital, zach orbital represents only a volume of probability of presence for an electron. (figure IV.18) :

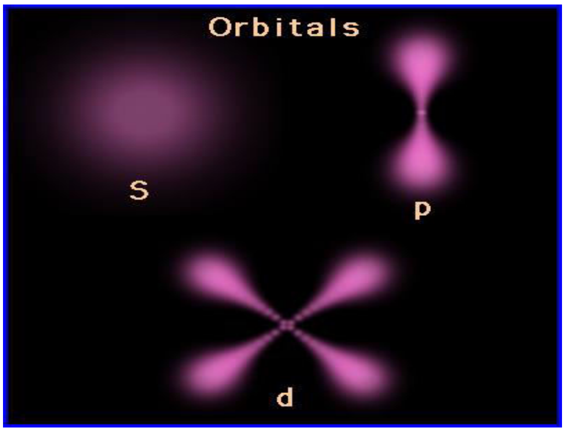


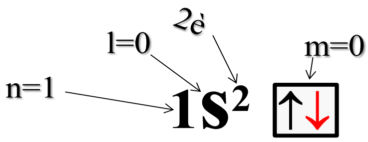
Figure IV.18 : Probability of presence for an electron (s, p, d)

XIV-The ectronic configuration :

Recall that each atomic orbital has three quantum numbers : n, l, m and that each electron has four quantum numbers with the notation : (n, l, m, s).

The electronic configuration of an atom indicates how electrons are distributed in the various atomic orbitals. The fundamental level is the electronic configuration that leads to the lowest possible energy (stable ground state)

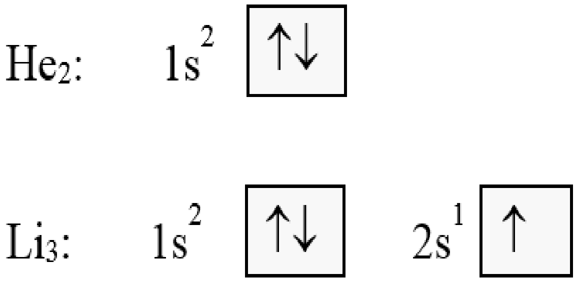
- The electronic configuration on an element is obtained by writing it as follow:



XIV-1- Pauli exclusion principle :

- An atomic orbital can only contain two electrons : One electron necessarily has spin +1/2 and the other necessarily has spin spin -1/2.

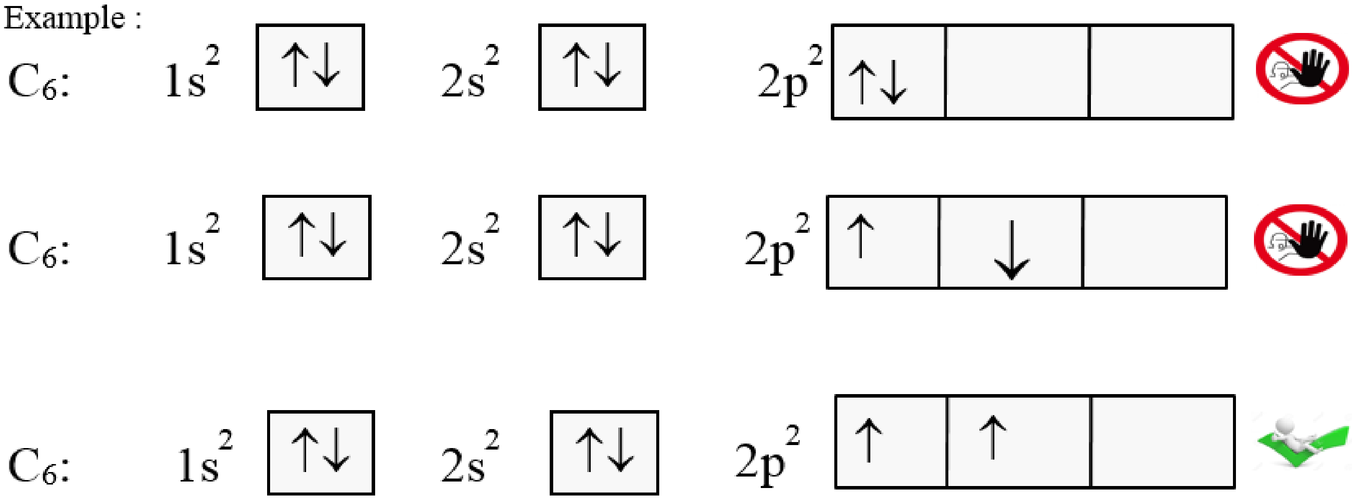
Example :



XIV-2- Hund Principle :

- The HUND rule specifies that : When an underlay is not complete, the (é) occupy the maximum of orbital with spin in the same direction.

Example :



XIV-3- Klechowski principle :

Following the red arrow (figure IV.19), here is the filling order of the sublayers one after the other following the Klechkowski rule (orbitals of increasing energy).

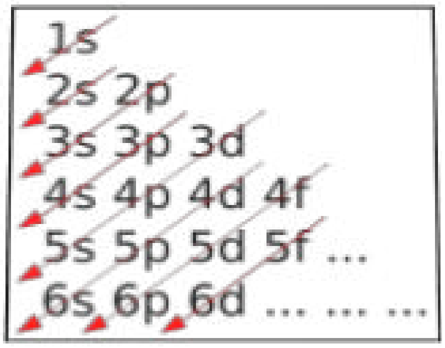


Figure IV.19 : Probability of presence for an electron (s, p, d)

So following the red arrow, we then obtain the electronic configuration for any element X :

X: 1S 2S 2P 3S 3P 4S 3d 4P 5S 4d 5P 6S 4f 5d 6P 7S……………

XIV-4- Aufbau principle :

Aufbau (German word) :" stacking construction". To make the electronic configuration more compact, we use the inert gas carcass, by placing the rare gas between brackets which precedes the element in question.

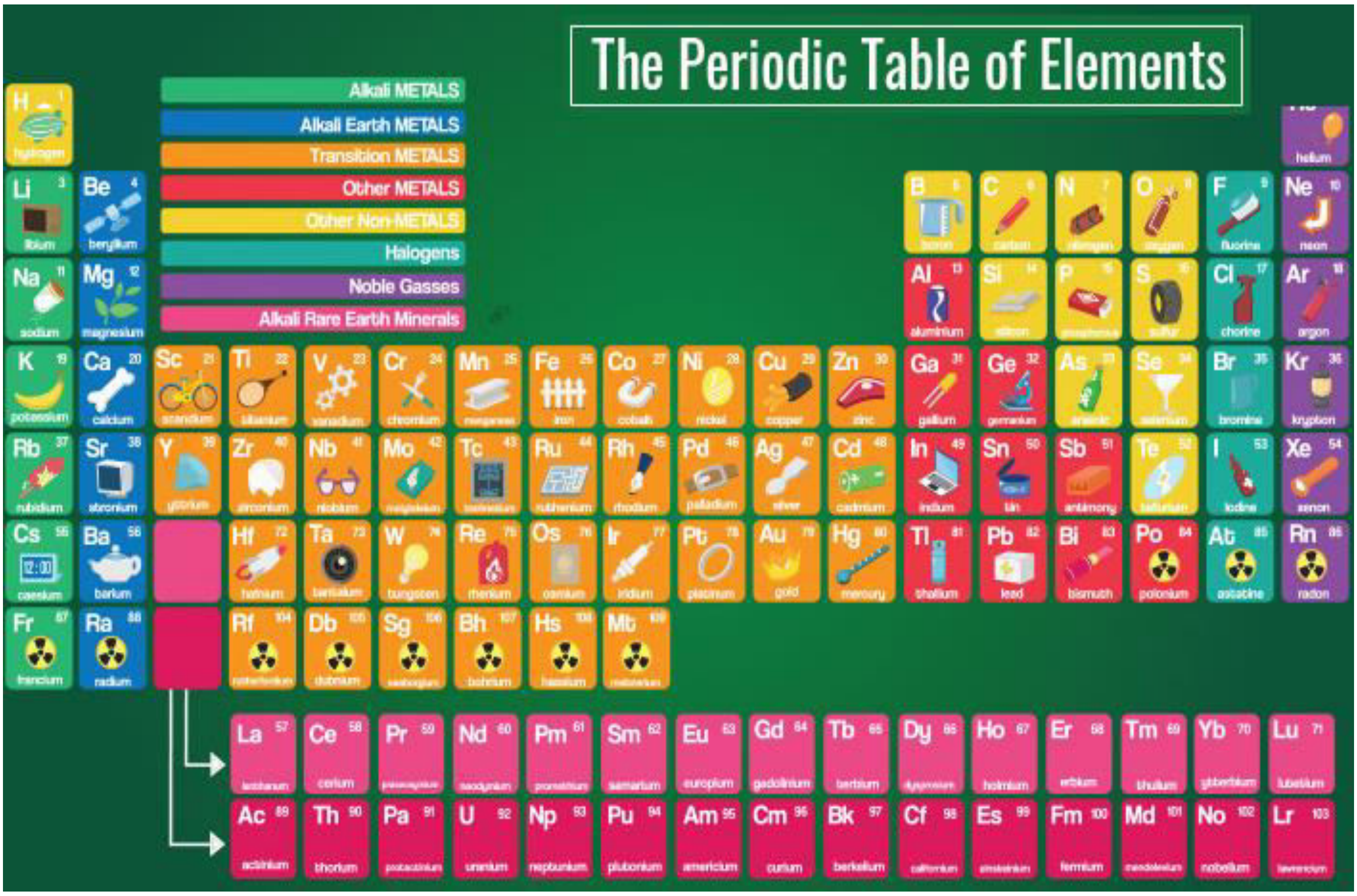
Example:

Al_{13} : $1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^1$

Al_{13} : [Ne] $3S^2 3P^1$

Chapter V

Periodic Classification of elements.



I- Historical :

Mendeleïev is Russian chemist, was the first to propose in 1869 the most complete and elaborate classification of the 66 chemical elements known at the time. Mendeleïev left some boxes empty, which were subsequently filled in as the corresponding elements were discovered. Four "superheavy" chemical elements have been added to the periodic table. They were discovered by Japanese, Russian and American teams and published on Tuesday, January 5, 2016.

Example :

- The ununtrium (Uut^{113}) is one of the elements predicted by Mendeleïev (éka-thallium).

Mendeleïev classified the elements in ascending order of atomic mass, going to the next line, so that atoms with similar properties are found one below the other, forming a family.

II- The periodical table :

The periodic table includes :

- 7 lignes called the 7 PERIODS.

- 18 colomns called FAMILIES and designated from left to right by a number from 1 à 18 or by Roman numerals followed by the symbol A or B.

-103 elements, 90 of them naturally occurring and 13 artificially created, defined primarily by the element's symbol, atomic number (Z) and atomic mass.

- Elements of the same family have the same number of valence electrons.

- Valence electrons are the electrons of the outermost layer.

- The family number corresponds to the number of valence electrons.

Example :

IA family elements have 1 valence electron ;
IIA family elements have 2 valence electrons ;
IIIA family elements have 3 valence electrons ;
IVA family elements have 4 valence electrons ;
VA family elements have 5 valence electrons ;
VIA family elements have 6 valence electrons ;

VIIA family elements have 7 valence electrons ;

VIIIA family elements have 8 valence electrons.

II-1- Blocks of the periodic table :

The periodic table is composed of four blocks (S, B, P and f) (figure V.1).

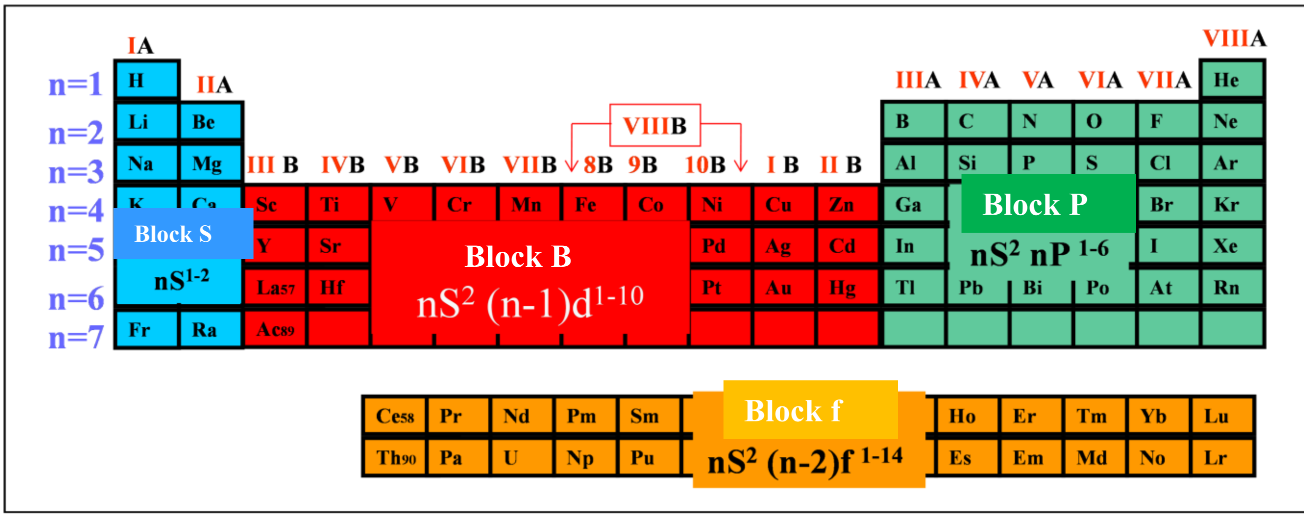


Figure V.1 : The blocks of the periodic table

We have : **- n=Ligne = period=layer=1 , 2 , 3 , 4 , 5 , 6 et 7.**

- Colonne = famille= groupe= (chiffre + lettre A ou B)

=IA→VIIIA

=IB→VIIIB

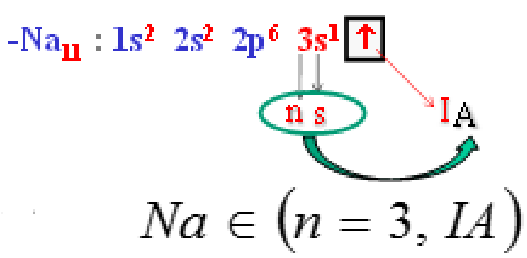
II-2- Location of elements in the periodic table :

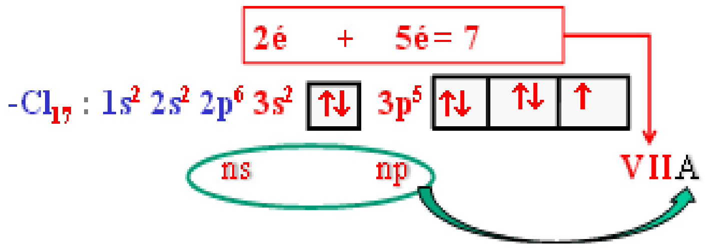
To locate an element in a periodic table, you need to know its group and period :

1-Make the eletronic structure of the elements.

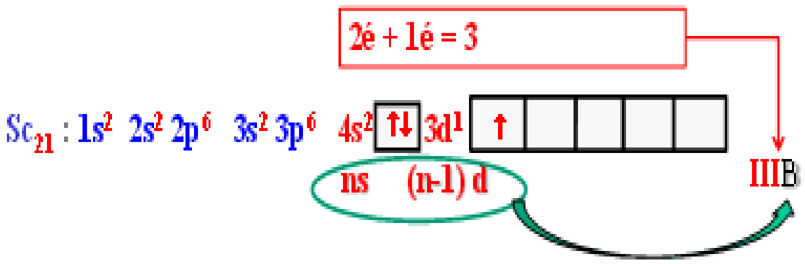
2-Search for the layer and the valence electron (outer layer).

Example : Na_{11}, Cl_{17} and Sc_{21}.





$$Cl \in (n = 3, VIIA)$$



$$Sc \in (n = 4, IIIB)$$

So The location of the elements Na_{11}, Cl_{17} and Sc_{21} in the periodic table is therefore as follows (figure V.2):

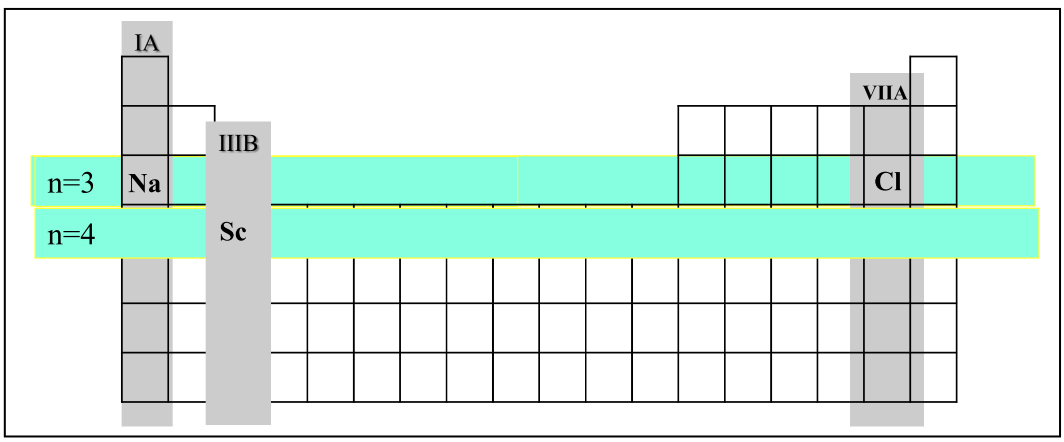


Figure V.2 : Position of the elements Na_{11}, Cl_{17} and Sc_{21} in the periodic table

II-3- Families of elements :

The 103 elements of the periodic table are classified into 11 families as shown in figure V.3 below.

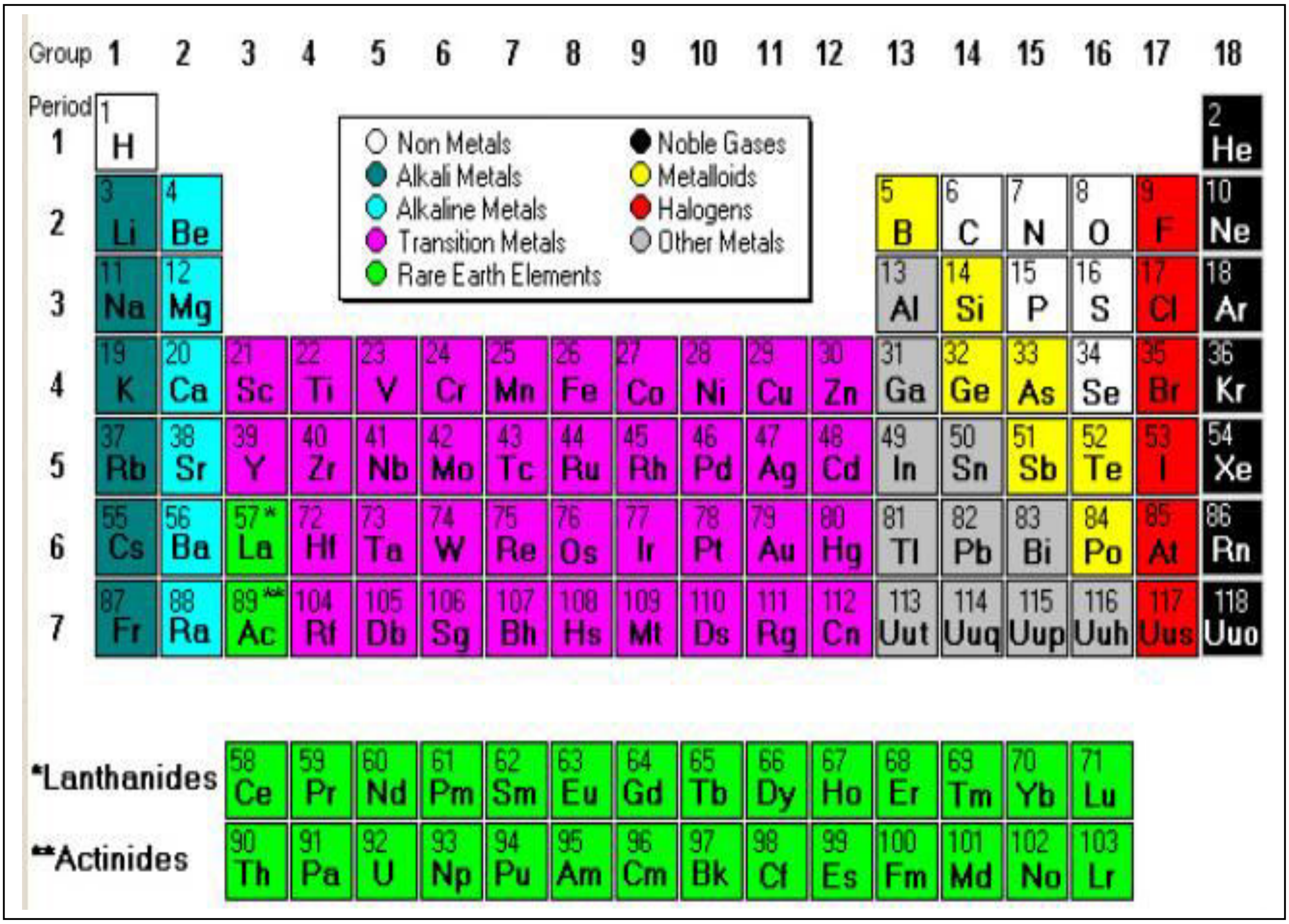


Figure V.3 : Families of elements in the periodic table

II-4-Metals :

The 103 elements of the periodic table are classified into 3 categories according to their properties but most of chemical elements are metals (figure V.4).

- **Metalloïds** are difficult to classify as metal or non-metal, they are at the boundary (step ligne) that separates metals from non-métals.

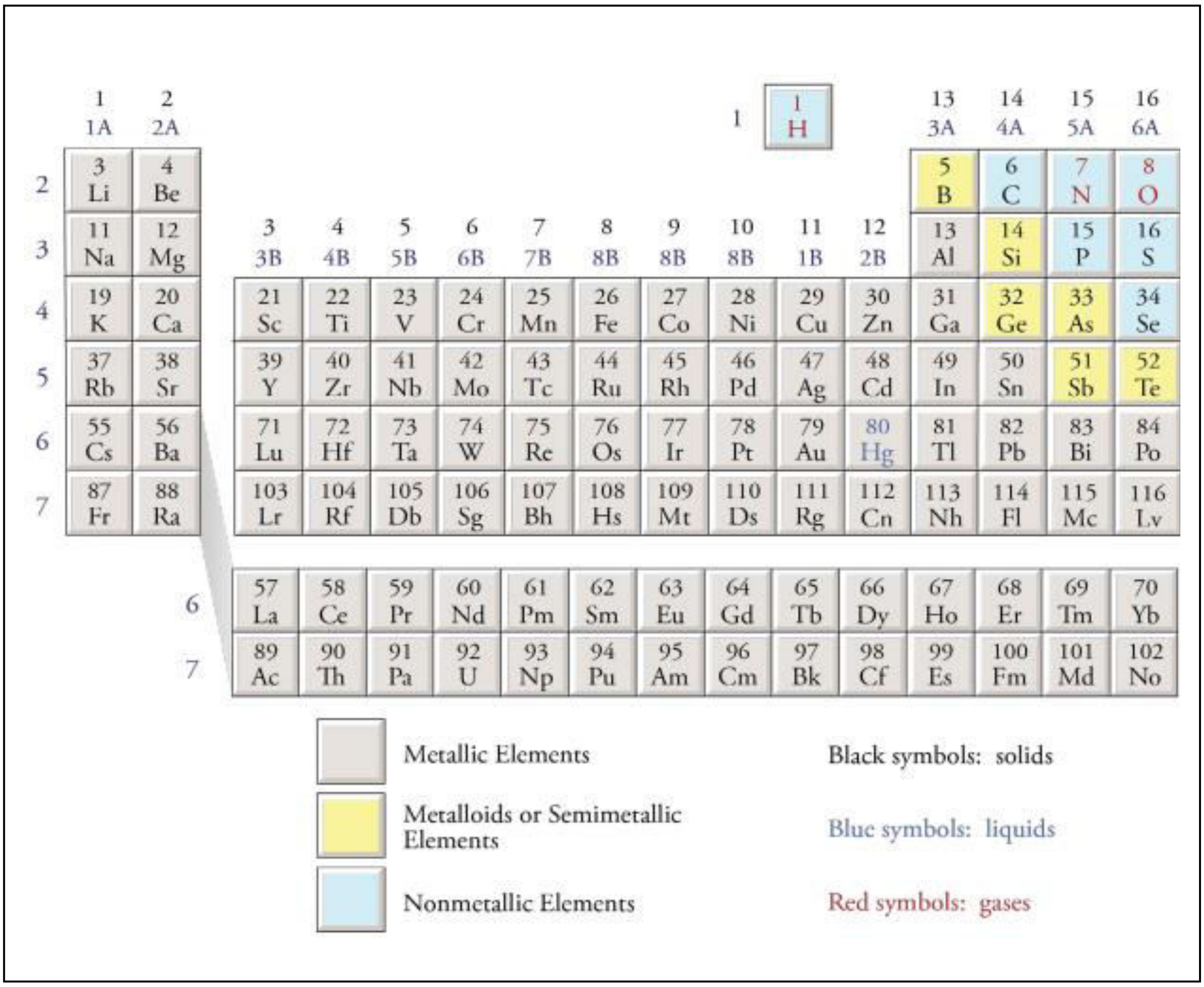


Figure V.4 : Metalloids in yellow separate metals from non-metals in the periodic table.

II-5-Properties of elements :

II-5-1- The atomic radius or the covalent atomic radius (r) :

It is half the distance between the nuclei of the simple body (diatomic) (figure V.5).

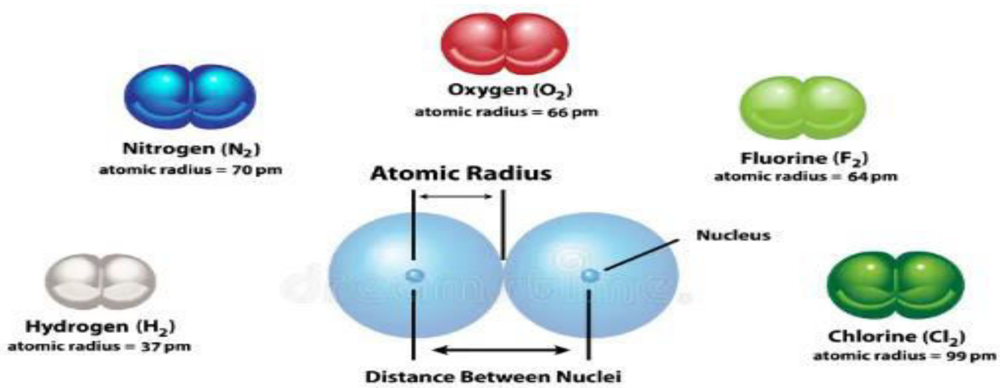


Figure V.4 : Atomic Radius of diatomic Molecules

II-5-2-Ionization energy :

It is the necessary energy to remove one or more electrons from the atom in the gaseous state.

-First ionization energy (**Ei** $_1$) : It is the necessary energy to remove one electron from the atom in the gaseous state.

$$x \rightarrow x^{+} + e^{-} \qquad \mathbf{Ei}_{1} = E(x^{+}) - E(x)$$

-Second ionization energy (**Ei** $_2$) : It is the necessary energy to remove a second electron from the atom in the gaseous state.

$$x^{+} \rightarrow x^{2+} + e^{-} \qquad \mathbf{Ei}_{2} = E(x^{2+}) - E(x^{+})$$

Example : Béryllium Atom Be^{4+}

$Ei_1 = 9{,}28$ eV ; $Ei_2 = 18{,}1$ eV ; $Ei_3 = 155$ eV ; $Ei_4 = 217$ eV

II-5-3-Electronic affinity :

This is the energy that is released when an electron is added to the atom in its gaseous state.

-First electronic Affinity ($\mathbf{A_{E\,1}}$) :

$$x + e^{-} \rightarrow x^{-} \qquad \mathbf{A_{E\,1}} = E(x^{-}) - E(x)$$

-Second electronic Affinity ($\mathbf{A_{E\,2}}$) :

$$x^{-} + e^{-} \rightarrow x^{-2} \qquad \mathbf{A_{E\,1}} = E(x^{-2}) - E(x^{-})$$

II-5-4-The electronegativity :

-Measure the trend of atoms whithin moleculs to attract electrons from another atom to form bonds.

-Let A and B be two chemical elements :

$$A + B \rightarrow A^{+\delta} — B^{-\delta}$$

A easily gives one (é) and **B** easily captures one (é).

<u>With :</u> **+δ :** Positive partial charge ;

–δ : Negative partial charge.

Example HF molecule :

H easily gives one (é) and **F** easily captures one (é) (figure V.5).

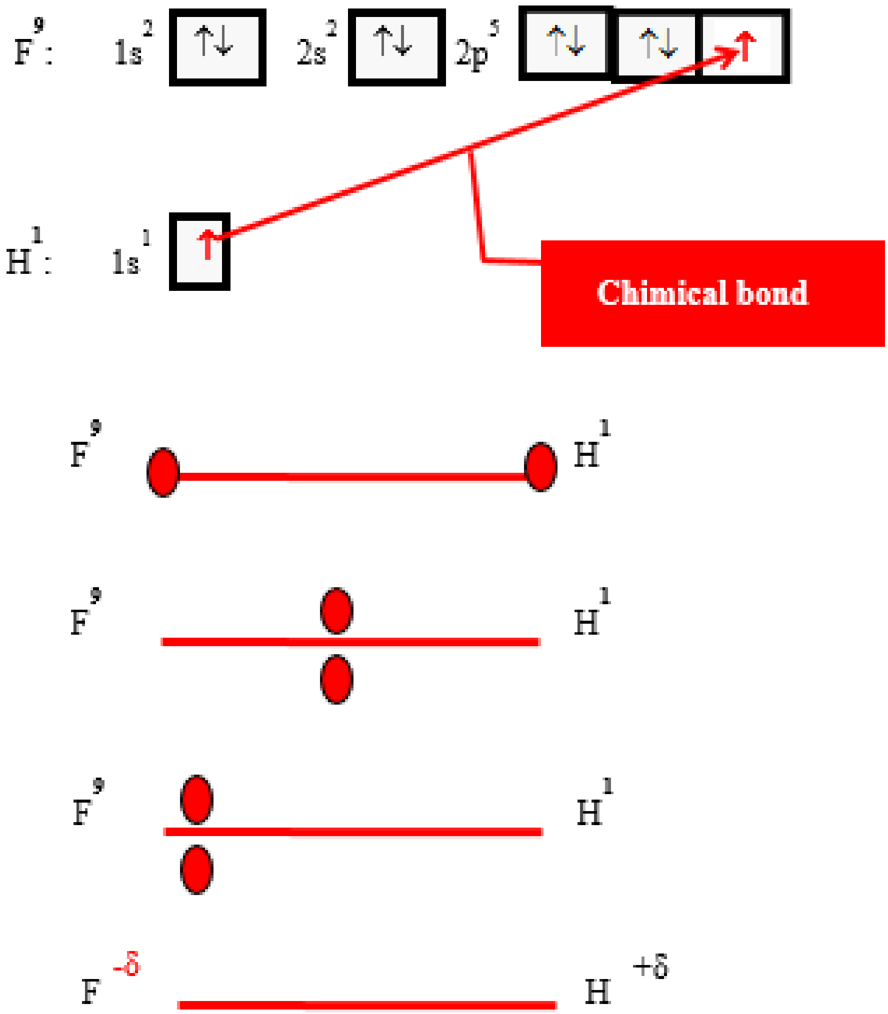


Figure V.5 : Chimical band between H and F

Notice :

$0,7 < \mathbf{En} \leq 1,2$ ……….Strong metals
$1,5 \leq \mathbf{En} \leq 2,0$ ……...Weak metals
$2,1 \leq \mathbf{En} \leq 4,0$………. Non-metals

-The Halogens (VIIA) are the most electronegative elements.

-Fluorine (F) = 4, the most electronégative.

III- Electronegativity calculation :

III-1- Mulliken scale :

According to this scale, electronegativity is calculated on the basis of ionization energy and electron affinity.

$$E_n = \frac{E_i + |A_E|}{2} (eV)$$

III-2- Pauling scale :

According to this scale, electronegativity is calculated on the basis of the bond dissociation energies of diatomic molecules.

If : D_{A-B}, D_{A-A} and D_{B-B} in Kcal/mol, we find :

$$|E_{nA} - E_{nB}| = 0{,}208\sqrt{D_{A-B} - \sqrt{D_{A-A} . D_{B-B}}}\ (eV)$$

If : D_{A-B}, D_{A-A} *et* D_{B-B} in Kj/mol, we find :

$$|E_{nA} - E_{nB}| = 0{,}102\sqrt{D_{A-B} - \sqrt{D_{A-A} . D_{B-B}}}\ (eV)$$

III-3- Allred Rochow scale :

According to this scale, electronegativity is calculated on the basis of covalent radius and effective Z (Zeff).

$$E_n = 0{,}359.\frac{Z_{eff}}{r^2_{\text{cov}\,alent}} + 0{,}744\,(eV)$$

With:
r : Covalent ray in A^0.

Z_{eff} : Effective charge units.

$$Z_{eff} = (Z - \delta)$$

δ: The screen constant which depends on the number of electrons located between the electron considered and the nucleus.It is calculated by Slater's rule.

VI- Calculation of effective charge units Zeff using Slater's rule :

- Slater stated the rules which allow these screening effects δ between electrons to be expressed.

- Write the electronic configuration of the element using the following groups and in the following order :

⇒Slater groups : [1s] ; [2s, 2p] ; [3s, 3p] [3d] ; [4s, 4p] [4d] [4f] ; [5s, 5p] [5d] ; [5f]...

- To calculate the screen effect (δ), use Slater's screen coefficients (Slater's triangle) (figure V.6) :

1s	0,3				
2s2p	0,85	0,35			
3s3p	1	0,85	0,35		
3d	1	1	1	0,35	
4s4p	1	1	0,85	0,85	0,35

Figure V.6 : Slater's triangle

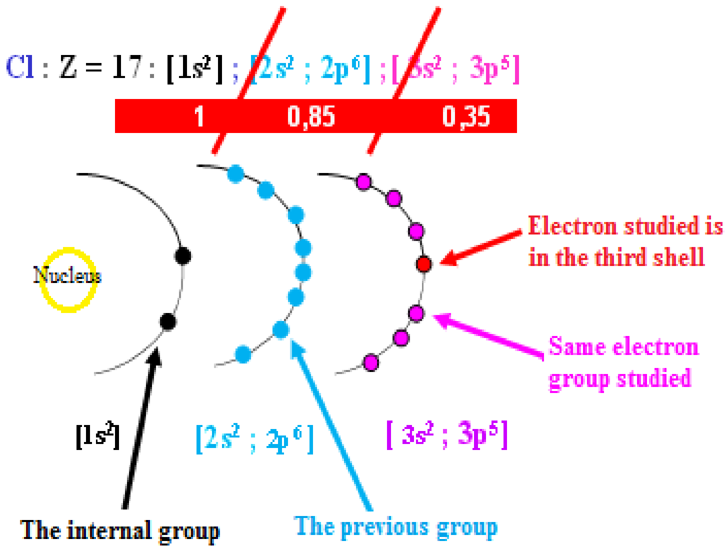
We have :

$$Zeff = Z - \delta \;\; avec\; \delta = \sum \delta_i$$

We have :

$$\delta = \sum \delta_i = [(n\; électron\; de\; la\; couche\; étudié\; - 1électron\; étudié)\, .\, coeff\; slater$$
$$+\; n\; électron\; de\; la\; couche\; précédente\, .\, coeff\; slater$$
$$+\; n\; électron\; de\; la\; couche\; pécédente\, .\, coeff\; slater + \cdots \ldots]$$

-Example of Z_{eff} calculation for chlorine (Cl) :



Zeff = 17 - ((7-1) * 0.35) + (8 * 0,85) + (2 * 1)) = 6,1

-Example of Ei calculation for chlorine (Cl) :

Cl : Z = 17 : [1s²] ; [2s² ; 2p⁶] ;[3s² ; 3p⁵]

Cl⁺ : Z = 17 : [1s²] ; [2s² ; 2p⁶] ;[3s² ; 3p⁴]

$Cl \rightarrow Cl^+ + e^-$

Ei= E$_{Cl^+}$ - E$_{Cl}$

$$\Rightarrow E_i = E_{Cl^+} - E_{Cl}$$

$$\Rightarrow E_i = \left(2\cancel{E_{1s^2}} + 8\cancel{E_{2s^2 2p^6}} + 6E_{3s^2 3p^4}\right) - \left(2\cancel{E_{1s^2}} + 8\cancel{E_{2s^2 2p^6}} + 7E_{3s^2 3p^5}\right)$$

$$\Rightarrow E_i = \left(6E_{3s^2 3p^4}\right) - \left(7E_{3s^2 3p^5}\right)$$

$$\diamond\ E_{3s^2 3p^4} = \frac{-13{,}6.Z^2_{eff\ cl^+}}{n^2} = \frac{-13{,}6}{3^2}.Z^2_{eff\ cl^+}$$

$$\diamond\ E_{3s^2 3p^5} = \frac{-13{,}6.Z^2_{eff\ cl}}{n^2} = \frac{-13{,}6}{3^2}.Z^2_{eff\ cl}$$

Zeff cl = 17 - ((7-1) * 0.35) + (8 * 0,85) + (2 * 1)) = 6,1

Zeff cl⁺= 17 - ((6-1) * 0.35) + (8 * 0,85) + (2 * 1)) =……

-Example of A$_E$ calculation for chlorine (Cl) :

Cl : Z = 17 : [1s²] ; [2s² ; 2p⁶] ;[3s² ; 3p⁵]

Cl⁻ : Z = 17 : [1s²] ; [2s² ; 2p⁶] ;[3s² ; 3p⁶]

$Cl + e^- \rightarrow Cl^-$

A$_E$ = E$_{Cl^-}$ - E$_{Cl}$

$$\Rightarrow A_E = E_{Cl^-} - E_{Cl}$$

$$\Rightarrow A_E = \left(2\cancel{E_{1s^2}} + 8\cancel{E_{2s^2 2p^6}} + 8E_{3s^2 3p^6}\right) - \left(2\cancel{E_{1s^2}} + 8\cancel{E_{2s^2 2p^6}} + 7E_{3s^2 3p^5}\right)$$

$$\Rightarrow A_E = \left(8E_{3s^2 3p^6}\right) - \left(7E_{3s^2 3p^5}\right)$$

$$\diamond\ E_{3s^2 3p^6} = \frac{-13{,}6.Z^2_{eff\ cl^-}}{n^2} = \frac{-13{,}6}{3^2}.Z^2_{eff\ cl^-}$$

$$\diamond\ E_{3s^2 3p^5} = \frac{-13{,}6.Z^2_{eff\ cl}}{n^2} = \frac{-13{,}6}{3^2}.Z^2_{eff\ cl}$$

Zeff cl = 17 - ((7-1) * 0.35) + (8 * 0,85) + (2 * 1)) =……

Zeff cl⁻ = 17 - ((8-1) * 0.35) + (8 * 0,85) + (2 * 1)) =……

Notice :

So, wa can calcule **the electronegativity (Mulliken scale)** For the chlore (Cl):

$$E_n = \frac{E_i + |A_E|}{2} (eV)$$

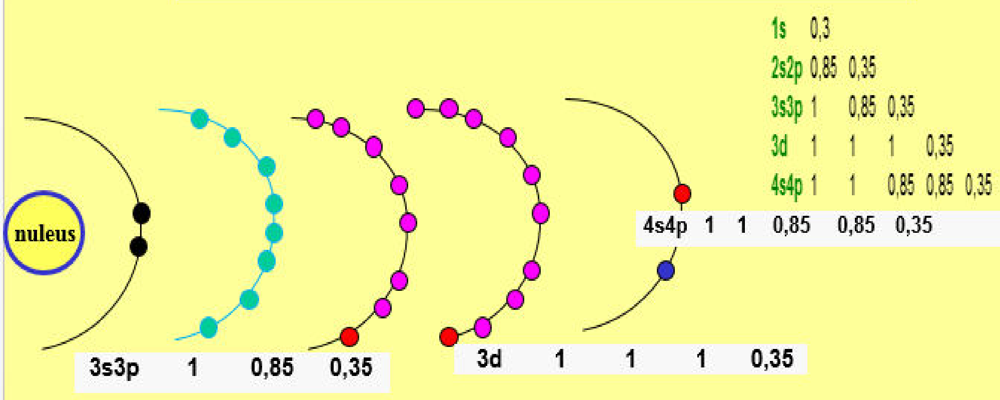
-Example of Zeff calculation for Zinc (Zn) :

Zn : Z = 30 : $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^{10}$

Slater groups:

Zn : Z = 30 : [$1s^2$] ; [$2s^2$; $2p^6$] ;[$3s^2$; $3p^6$] ; [$3d^{10}$] ; [$4s^2$]

1 1 0,85 0,85 0,35



(10 * 0,85 + 8 * 0,85)

$Zeff_{4s}$ = 30 - ((2-1) * 0,35) + (18 * 0,85) + (8 * 1) + (2 * 1)) = 4,35

$Zeff_{3d}$ = 30 - ((10-1) * 0,35) + (8 * 1) + (8 * 1) + (2 * 1)) = 8,85

$Zeff_{3s;3p}$ = 30 - ((8-1) * 0,35) + (8 * 0,85) + (2 * 1))= 18,75

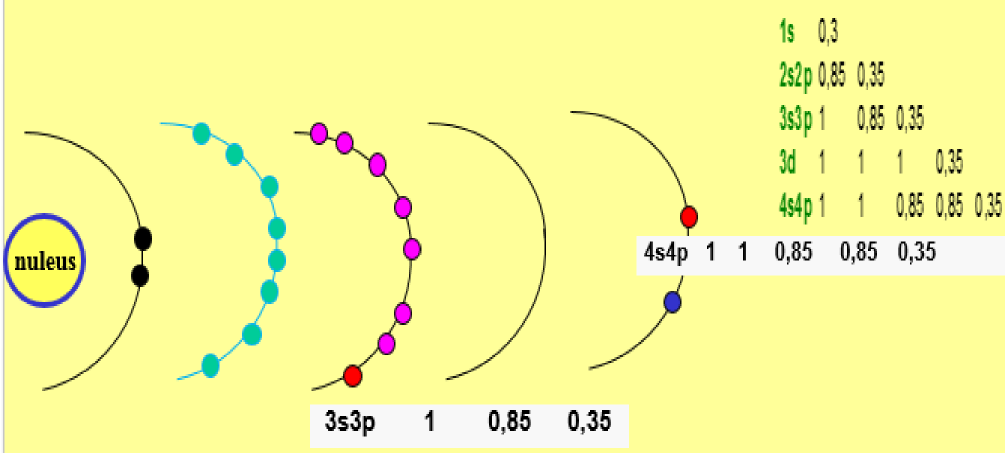
-Example of Zeff calculation for Calcium (Ca) :

Ca : Z = 20 : $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$

Slater groups:

Ca : Z = 20 : [$1s^2$] ; [$2s^2$; $2p^6$] ; [$3s^2$; $3p^6$] ; [$3d^0$] ; [$4s^2$]

1 1 0,85 0,85 0,35



(0 * 0,85 + (6+2) * 0,85)

$Zeff_{4s}$ = 20 - ((2-1) * 0,35) + (8 * 0,85) + ((6+2) * 1) + (2 * 1)) = 2,85

$Zeff_{3s;3p}$ = 20 - ((8-1) * 0,35) + (8 * 0,85) + (2 * 1) = 8,75

V-Variation of En, AE, r and Ei in the périodic table :

Generally speaking, according to the periodic table, the further to the right you go in the same period, the smaller the atomic radius (r) and the smaller the electron affinity (A_E) (figure V.7). The further down the same column you go, the greater the r and A_E. The further to the right over the same period, the greater the ionization energy (Ei) and electronegativity (En). The further you go down the same column, the smaller the Ei and En.

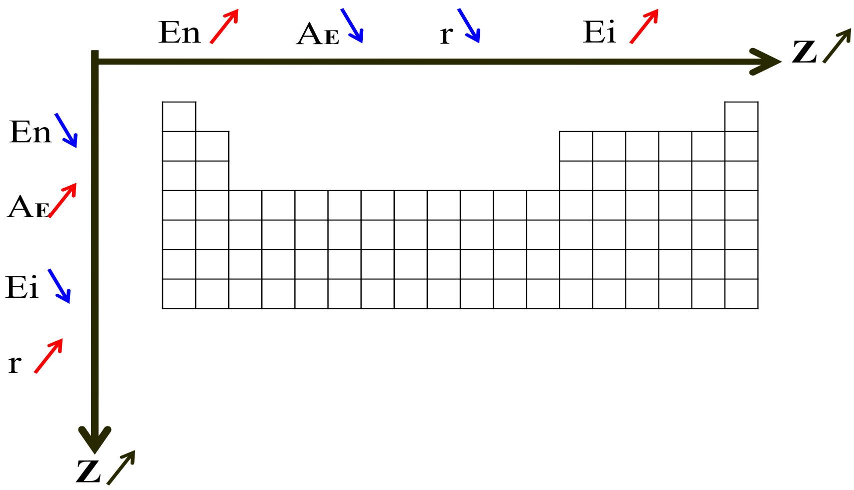


Figure V.7 : Variation of En, AE, r and Ei in the periodic table

Example 1 :

Ei comparison for two elements Li and Rb in the same IA group (figure V.8).

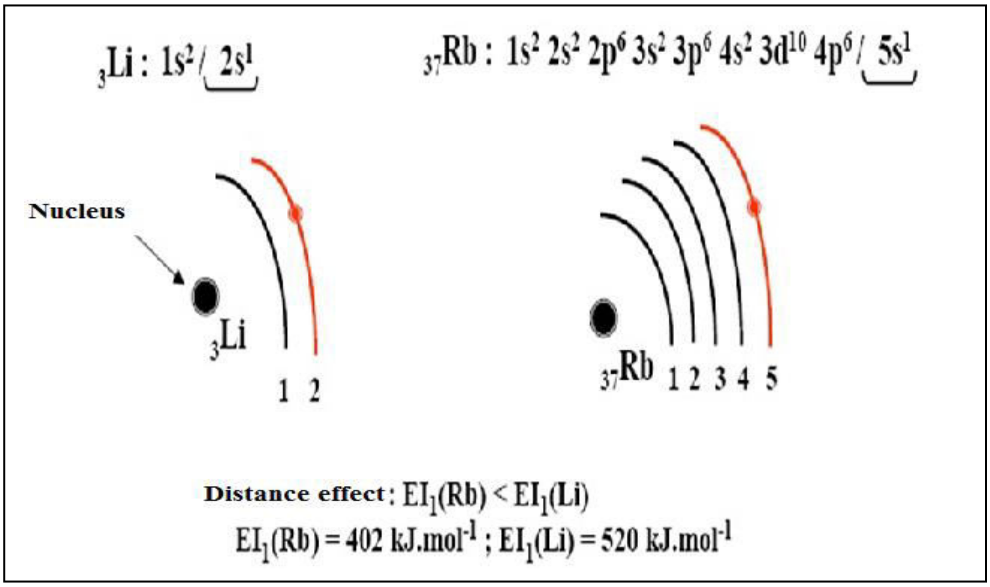


Figure V.8 : Variation of Ei for Li and Rb elements in the same group.

Example 2 :

Ei comparison for two elements Li and F in the same period (n=2) (figure V.9).

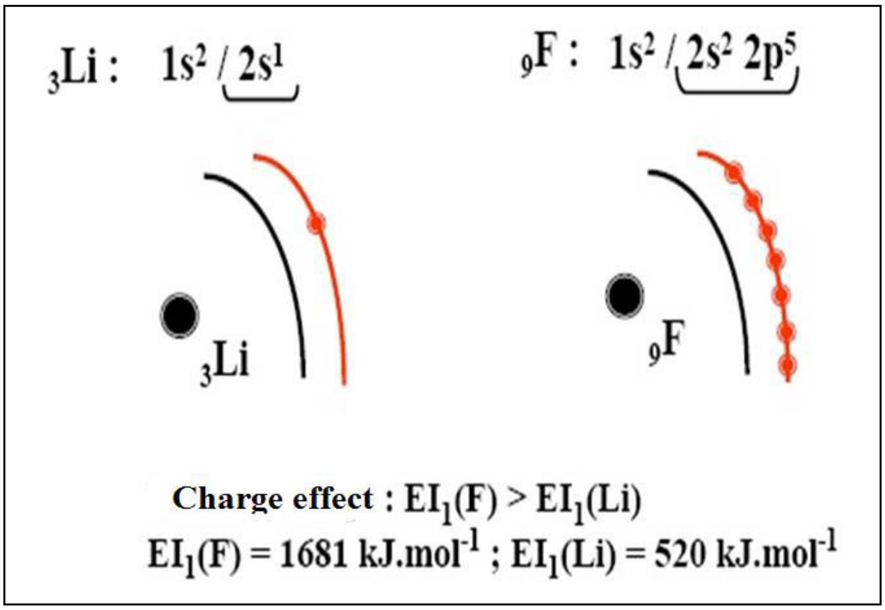
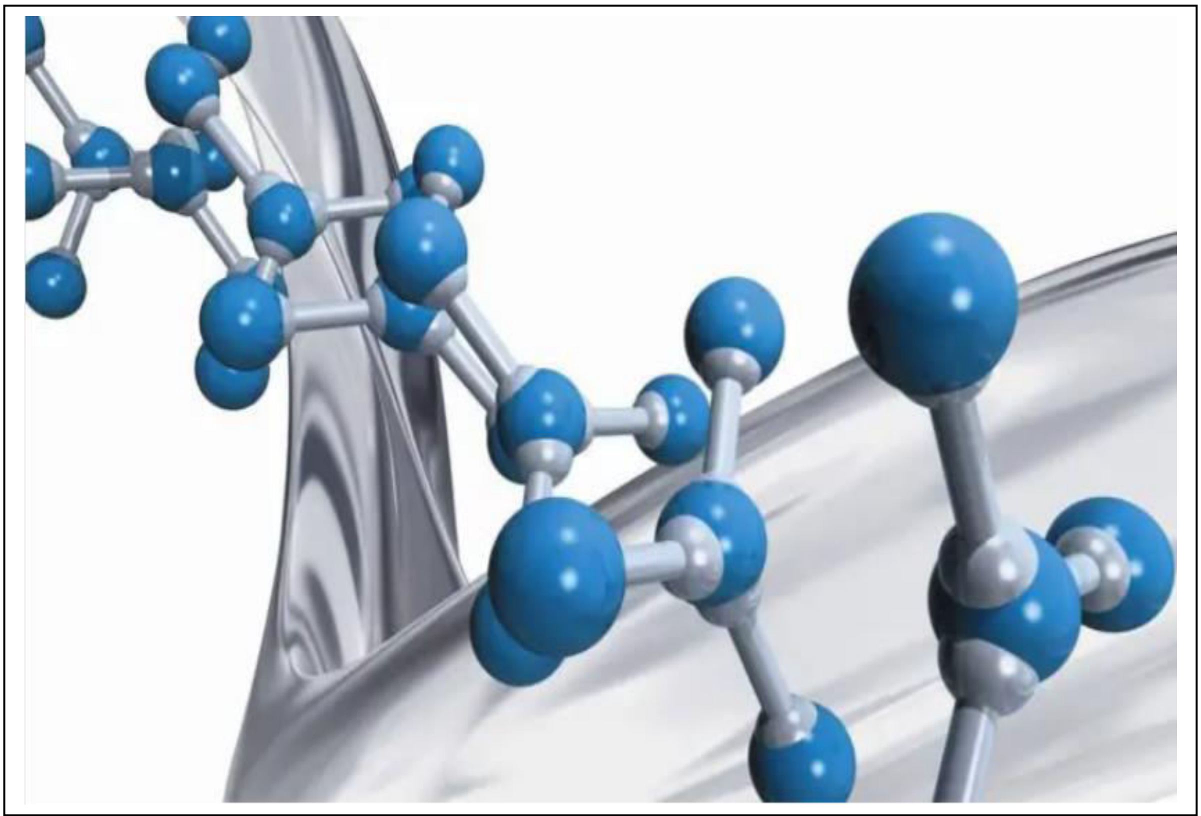


Figure V.9 : Change in Ei for Li and Rb over the same period.

Chapter VI

Chemical bonds



I- Lewis's theory of chemical bonding :

I-1- Electrons :

Valence electrons, in particular, play a crucial role in chemical bonding. In combinations composed solely of non-metallic atoms, bonded atoms share one or more doublets of valence electrons, giving rise to covalent bonds. Exp : H_2O, CH_4, NH_3

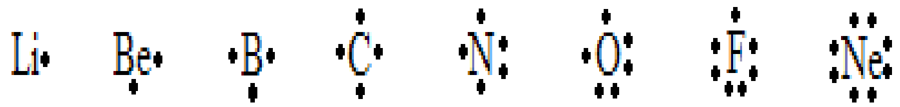
I-2- The octet rule :

Atoms combine to have eight electrons in their valence shell, giving them the same electronic structure as a noble or rare gas (groupVIIIA).

I-3- Lewis notations :

The valence electrons of a chemical element are identified by dots distributed around the symbol.

Example :



II- Nature and different types of bonds :

The bonding doublet is symbolized by a dash between the 2 atoms.

We have two atoms : A•   •B

We write : A—B

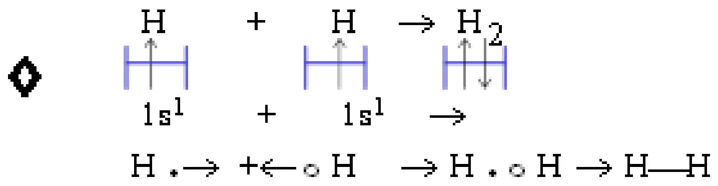
II-1- The covalent bond :

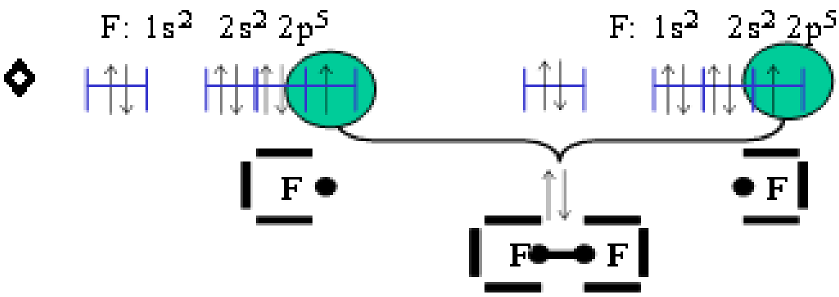
It results from the formation of a pair of opposite-spin electrons belonging in common to both atoms ; this is the theory of covalent bonds :

II-1-1- The simple covalent bond :

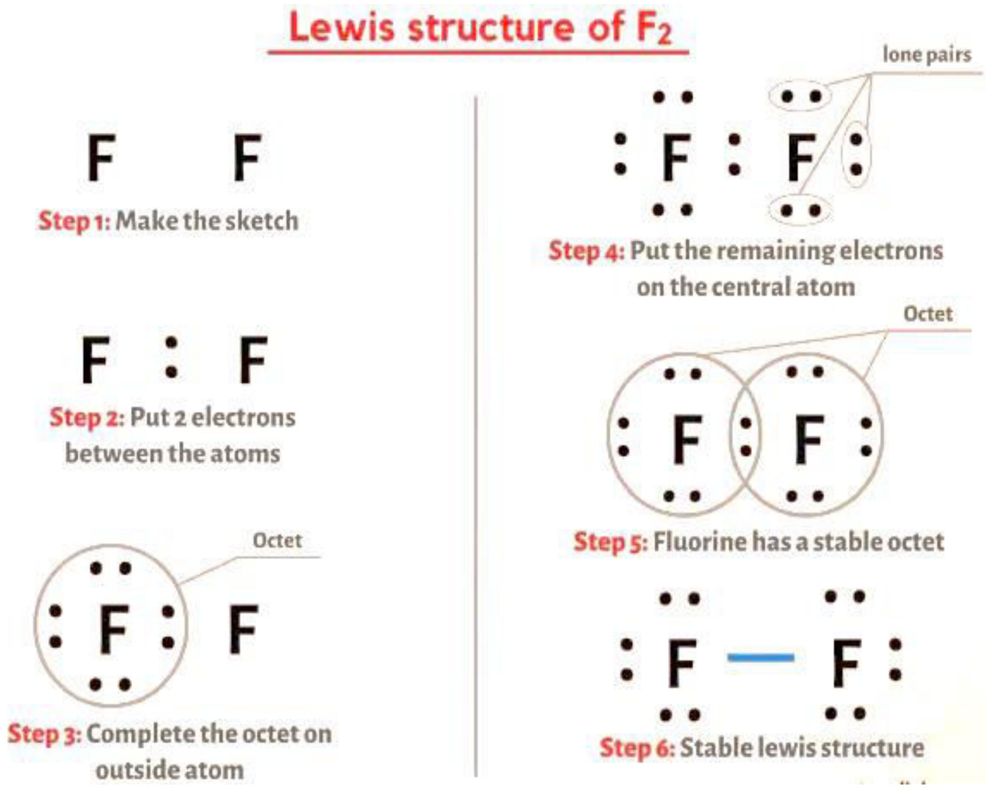
A bond in which two atoms A and B share two electrons from their outer layer (each atom contributes one electron).

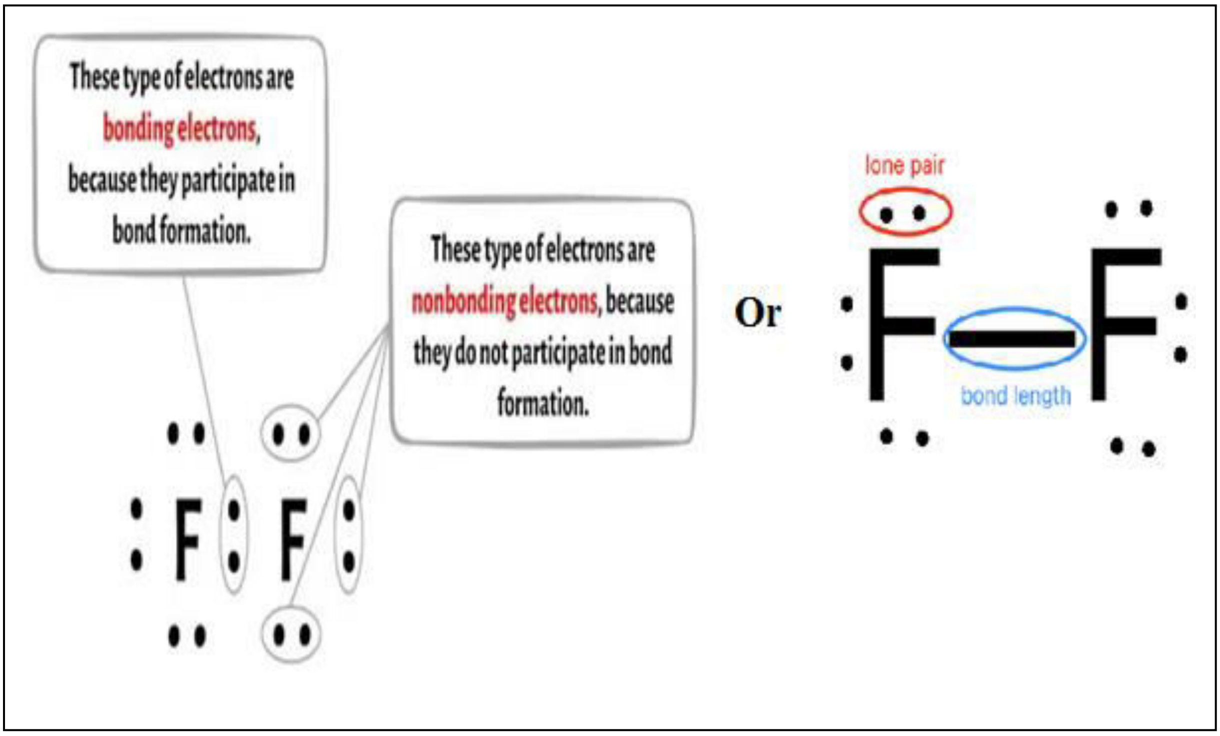
-Example H_2 and F_2 :



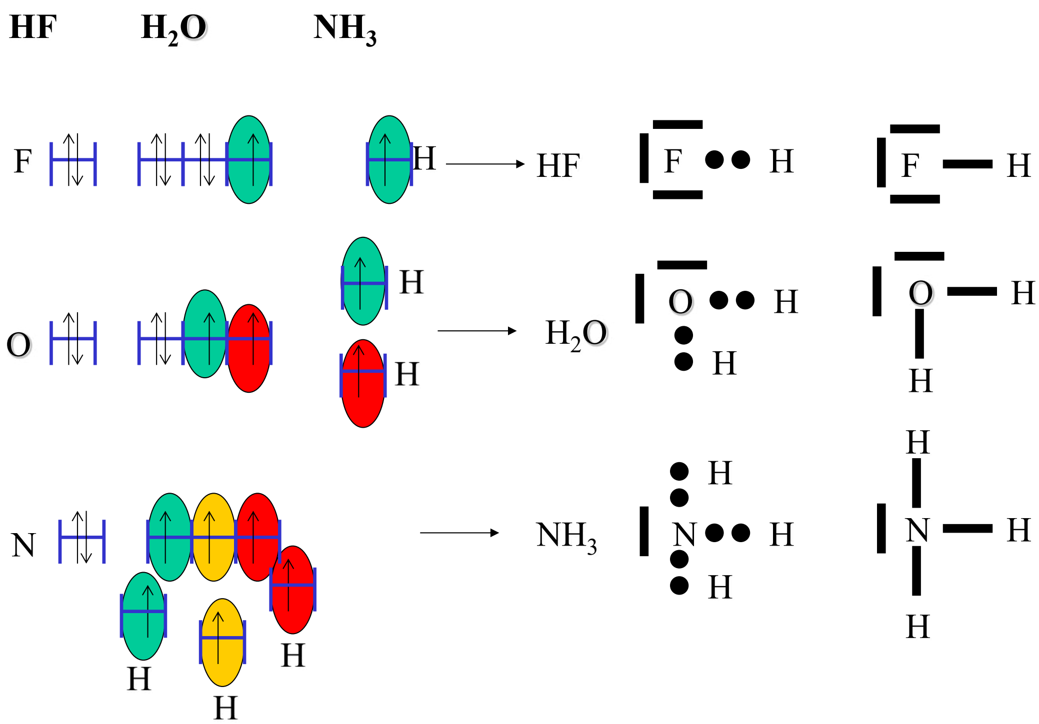


So for molecule F_2 we have :





Other examples : HF, H_2O, NH_3

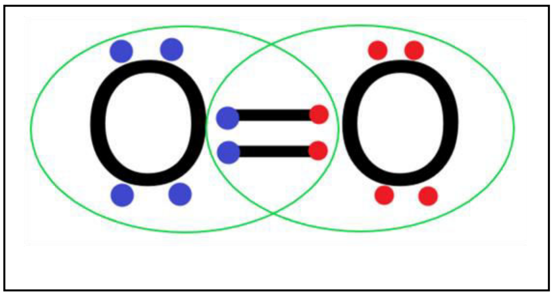


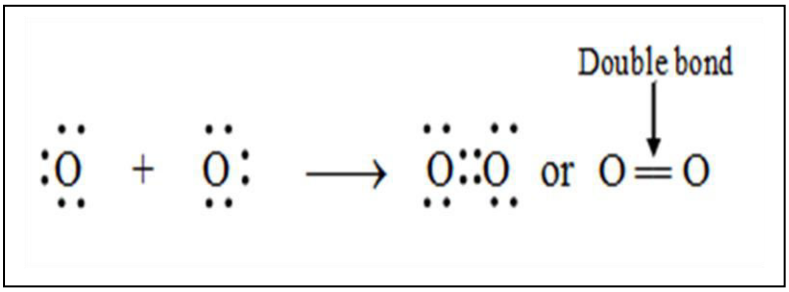
II-1-2- The multiple covalent bond :

Two bonded atoms can share more than one electron doublet ; this is known as a multiple bond. There are double bonds and triple bonds. In chemistry, a triple bond is a chemical bond between two atoms involving six valence electrons instead of two in a single covalent bond.

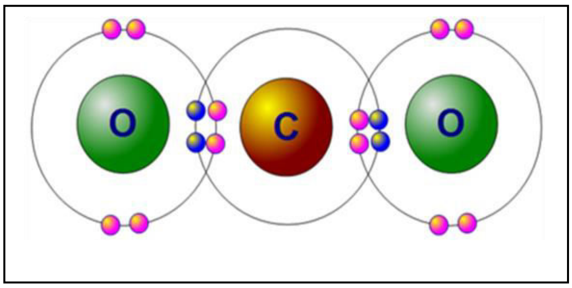
Example of a double bond O_2 et CO_2 :

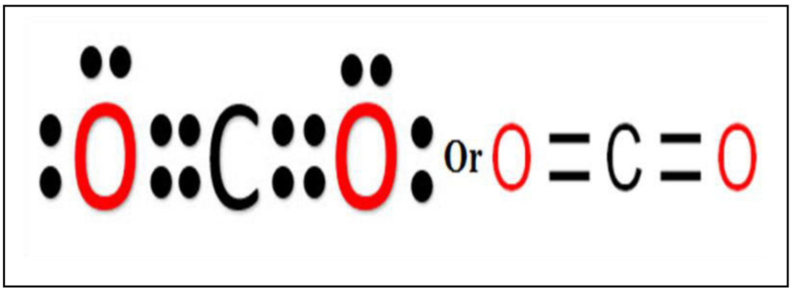
O_2



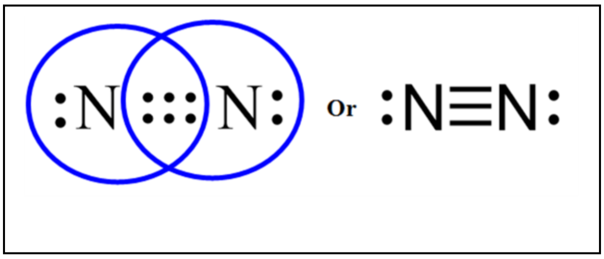


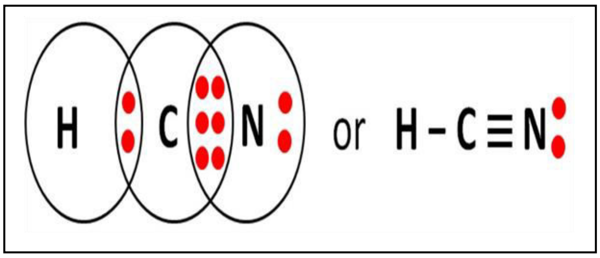
CO_2





Example of a triple bond N_2 et HCN :





III- The ionic bond :

Ionic bonding results from purely electrostatic attractions between ions of opposite signs. This is observed when the two atoms have different electronegativities ; in other words, the more electronegative atom attracts the bonding doublet towards itself, thus polarizing the bond.

Two atoms A and B with B more electronegative give the following ionic bond : $A^{\delta+}$— $B^{\delta-}$

Example : NaCl ($Na^{\delta+}$, Cl^{-})

The reaction between sodium and chlorine to form NaCl is as follows :

$$2\ Na_{(s)} + Cl_{2(g)} \rightarrow 2\ NaCl_{(s)}$$

In this reaction, sodium loses an electron to take on the electronic configuration of neon ; it then becomes the ion Na^{+} (figure VI.1).

Chlorine acquires an electron to take on the electronic configuration of Argon ; it then becomes the ion Cl^{-}. The result is a pair of ions with opposite signs that are strongly attracted to each other.

The electrostatic forces of attraction that hold cations and anions together are called ionic bonds. The solid, highly structured assembly of ions is called an ionic crystal.

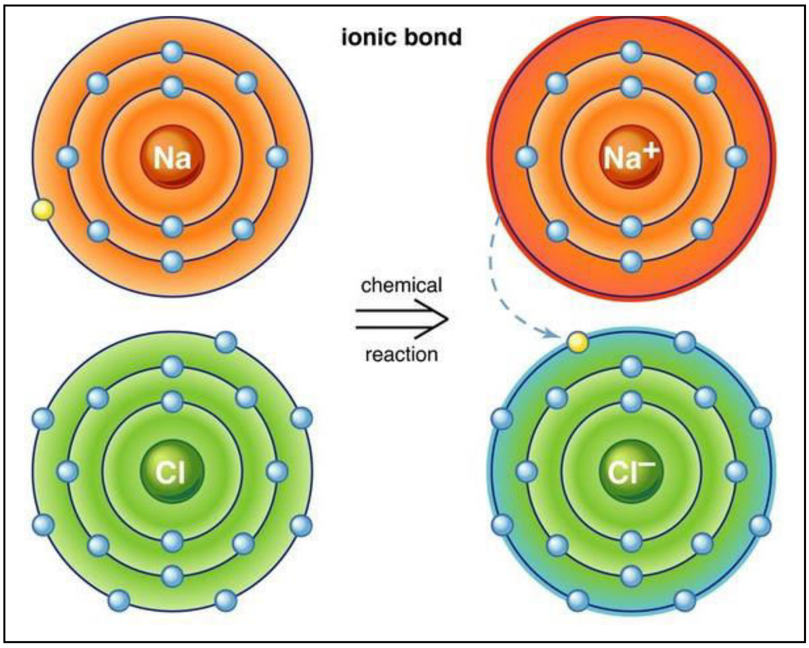
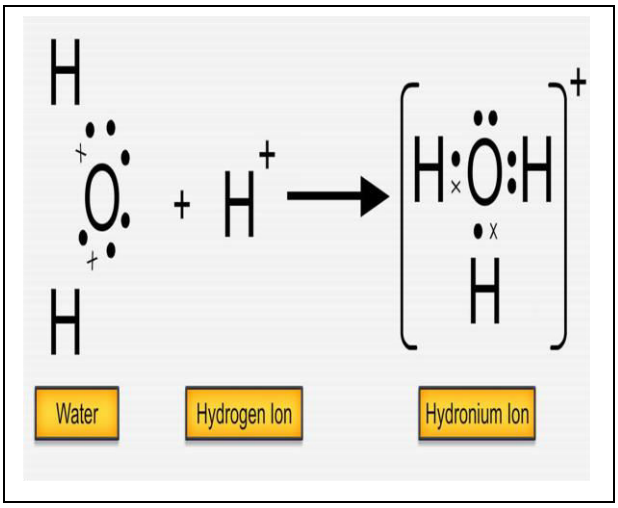


Figure VI.1 : Formation of the ionic bond between the Na and Cl atoms.

IV- The coordinating bond :

When the same atom uses its two free electrons to form a covalent bond.

Example : Formation of the hydronium ion, H_3O^{+}, in acidic aqueous media (figure VI.2).



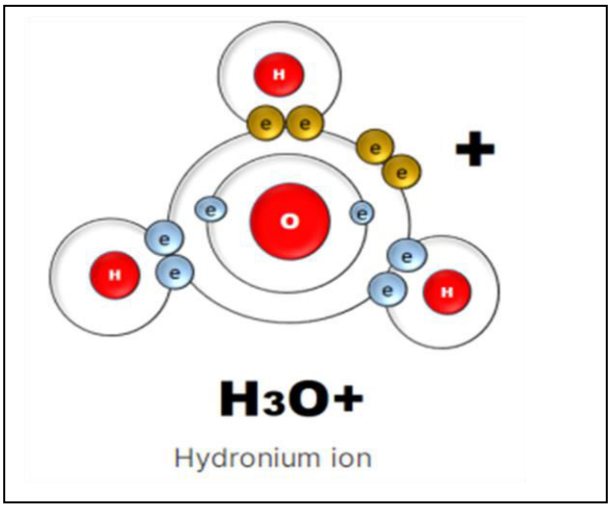
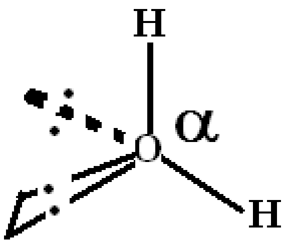


Figure VI.2 : Formation of the coordinative bond for the hydronium ion.

V- Orbital hybridization theory :

V-1-Introduction :

Gillespie's theory or VSEPR: (valence electron doublet repulsion) which expresses that the bonding doublets shared or not (binding and non-binding doublets) around the central atom undergo mutual repulsion to increase the distances between them and thus decrease the interaction energy and give the molecule more stability. That is 4 doublets around O, which will undergo mutual repulsions and give the molecule a spatial configuration of ($\alpha > 90°$) :



The same applies to NH_3, PCl_3, ….

VSEPR is therefore very useful for predicting the geometry of a molecule. To calculate the number of free doublets (m), we can use the following relationship: n+m = 1/2 (number of valence electrons in the central atom + number of single bonds - number of double bonds) + 1/2 (number of negative charges - number of positive charges).

V-2- Different types of hybridization :

V-2-1- SP^3 hybridization :

This is known as sp^3 hybridization of carbon. The 4 sp^3 hybrid orbitals correspond to more stable energy states for the valence electrons. They give 4 identical bonds at 109.5°. Since the S orbitals are spherical and the P orbitals point along the three Cartesian axes, we can't simply explain the tetrahedral molecule.

The atomic orbitals involved in its formation are the 2s and 2p orbitals of the carbon atom, which overlap with the 1s orbitals of hydrogen : $1s + 3p \rightarrow 4\ sp^3$

Sp^3 hybridization helps explain the spatial configuration of molecules such as CH_4 (figure VI.3).

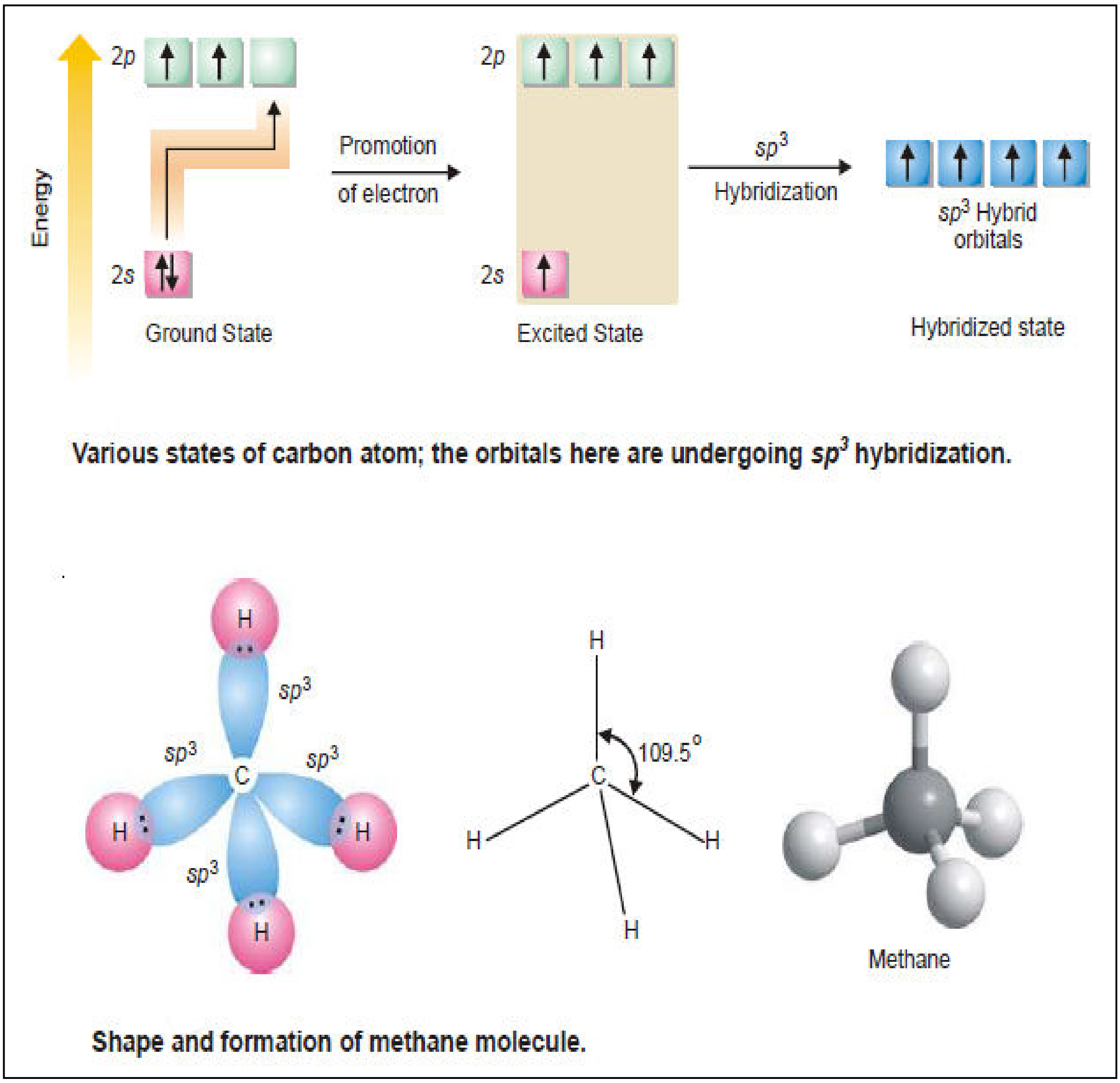
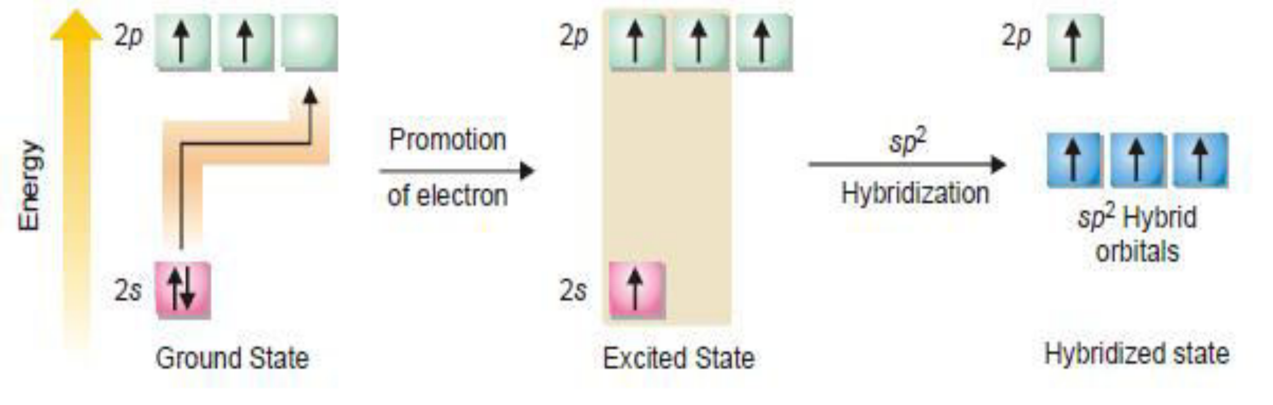


Figure VI.3 : sp^3 hybridization for the CH_4 molecule

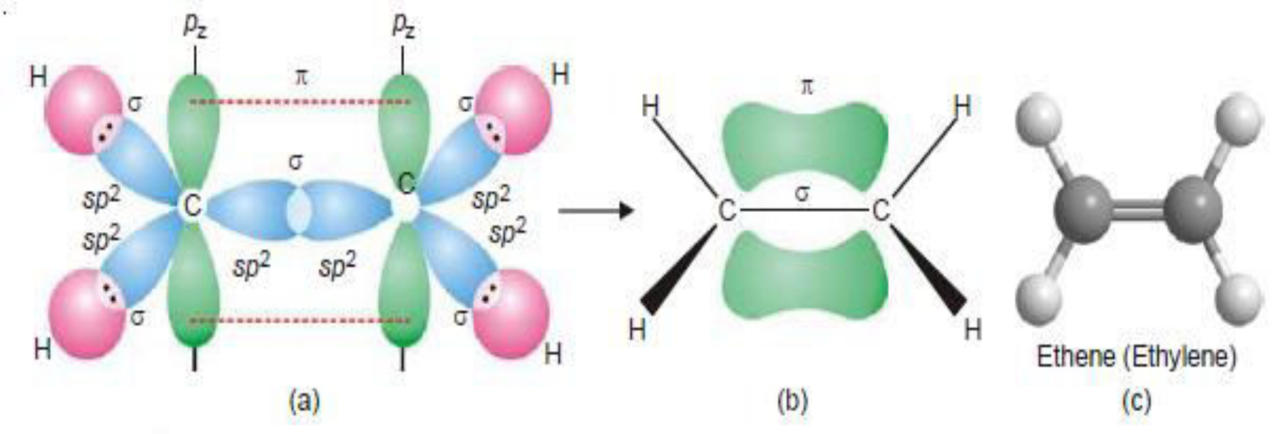
V-2-2- SP² hybridization :

The exponent 2 here indicates that only two p orbitals will be combined with the s orbital to obtain this type of hybrid orbital. This leaves one p-orbital on each carbon atom : $1s + 3p \rightarrow 3\ sp^2 + 1p$.

These sp^2 hybrid orbitals will point in the directions of an equilateral triangle centered on the carbon atom. sp^2 hybridization helps explain the spatial configuration of molecules such as C_2H_4 (figure VI.4).



Excited carbon atom undergoing sp^2 hybridization leaves a pure $2p_z$ orbital.



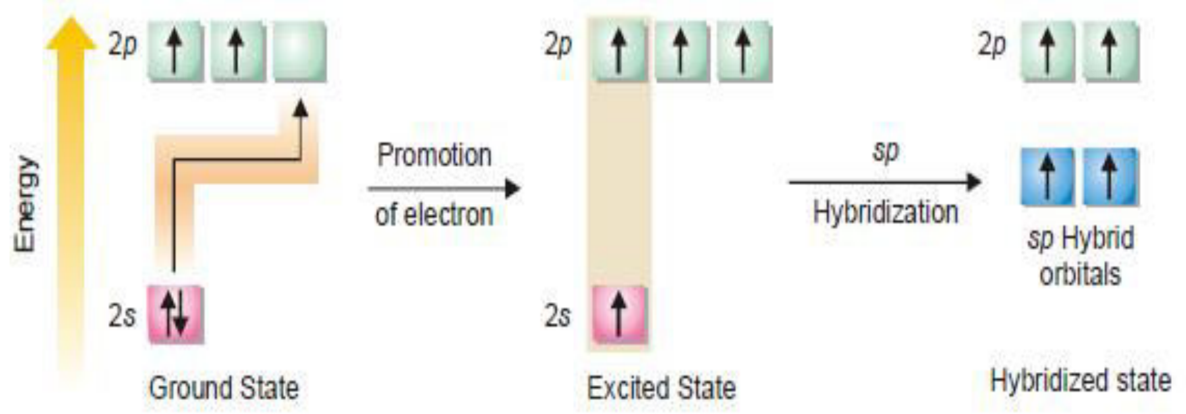
Orbital model of ethene molecule. (*a*) shows scheme of overlaps; (*b*) shows the bonds, sigma bonds indicated by straight lines; and (*c*) shows ball-and-stick model of ethene (ethylene).

Figure VI.4 : sp^2 hybridization for the C_2H_4 molecule

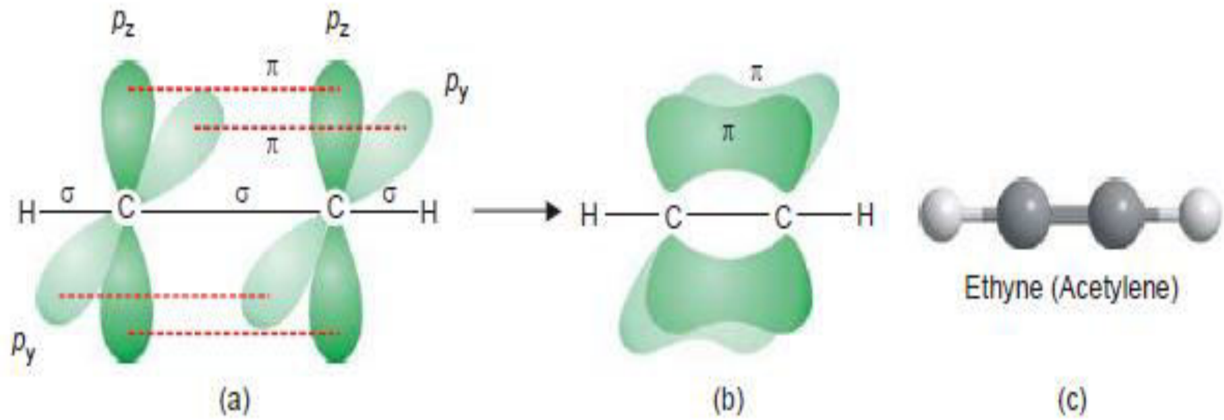
V-2-3- SP hybridization :

To describe the C_2H_2 molecule, we'll use hybrid atomic orbitals obtained by linearly combining the 2s atomic orbital and just one of the 2p atomic orbitals. We thus obtain 2 sp hybrid atomic orbitals and 2 atomic orbitals that remain pure 2p.

These hybrid orbitals are labelled sp and point 180° apart. sp hybridization helps explain the spatial configuration of molecules such as acetylene C_2H_2 (figure VI.5).



sp Hybridization of the orbitals of excited carbon.

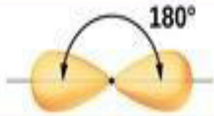
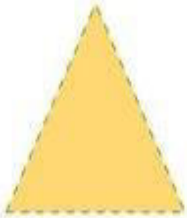
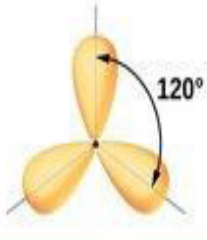
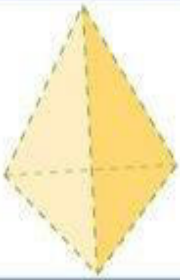
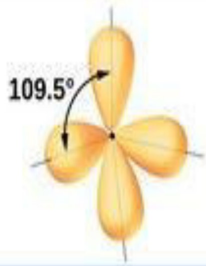
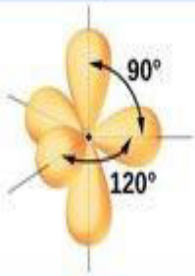
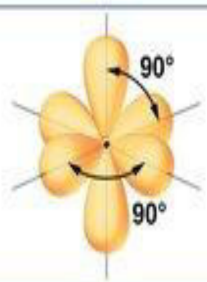


Formation of ethyne molecule. (*a*) shows scheme of overlaps; (*b*) shows the bonds, σ bonds being indicated by straight lines; and (*c*) shows ball-and-stick model of ethyne (acetylene)

Figure VI.5 : sp hybridization for the C_2H_2 molecule

- The most frequent cases of hybridization of atomic orbitals correspond to the geometric shapes already described in the VSEPR method, and are presented in the table VI.1 below :

Tableau VI.1 : The most frequent cases of atomic orbital hybridization.

Regions of Electron Density	Arrangement		Hybridization	
2		linear	*sp*	
3		trigonal planar	sp^2	
4		tetrahedral	sp^3	
5		trigonal bipyramidal	sp^3d	
6		octahedral	sp^3d^2	

Application exercises

Preparation of solutions

Exercise 1:

Give the definition of : pure substance, mixture, solvent, solution, reactant, mole numbers, molar mass, molecular weight, density, specific gravity, mass concentration, molar concentration (molarity), normality, molality, mole fraction.

Exercise 2 :

1. Explain how to prepare 100 mL of 1g/L solution of NaOH (sodium hydroxide)?
2. Explain how to prepare 2 L of a 1mol/L solution of NaOH?
3. Explain how to prepare 500 mL of a 1N solution of NaOH?

   (Molar mass of sodium M_{Na} *= 23 g/mol; Molar mass of oxygen* M_O *= 16 g/mol; Molar mass of hydrogen* M_H *= 1 g/mol)*
4. What is the volume of 96% (w/w) sulfuric acid (H_2SO_4), with a specific gravity of 1.84 required to prepare 250 mL of 1 N sulfuric acid?

   (Molar mass of sulfur M_S *= 32 g/mol)*

Exercise 3:

1. How many copper atoms are in 0.25 mol of copper?
2. What is the number of glucose molecules $C_6H_{12}O_6$ in 0.75 mol of glucose?
3. What is the number of atoms in 0.75 mol of glucose?

   (*Avogadro's number : NA=6.023* x 10^{23} mol^{-1})

Exercise 4:

A mass of 51.3 g of aluminum sulfate $Al_2(SO_4)_3$ is dissolved in 500 mL of water (H_2O).

1. Write the dissolution equation.
2. Determine the molarity, molality, mass concentration and normality of the prepared solution.
3. Determine the molar concentration of each ion species in the solution.
4. What is the volume to be taken from the initial solution to make 100 mL of 0.15 mol/L $Al_2(SO_4)_3$ solution?

   (Molar mass of Aluminum M_{Al} *= 27 g/mol)*

Exercise 1: (J.J. Thomson's Experiment)

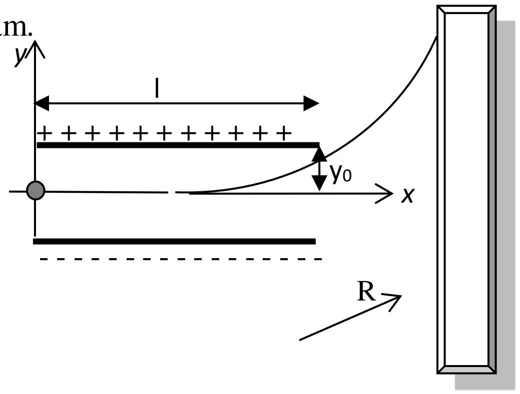
A- A beam of electrons of velocity **v** flows between two oppositely charged plates of a capacitor where a uniform electric field **E** is perpendicular to the direction of the beam and a magnetic field **B** perpendicular to both **E** and **v**.

1- Show the direction of vectors v_0, E, Fe, B and F_B. (define Fe, F_B, v_0)

B- The two fields E and B act simultaneously and their intensity is adjusted so that the beam passes undeflected.

1- Derive an expression for the velocity **v** of the electron beam.

2- Calculate the charge-to-mass ratio **e/m** of the electron.



C- The action of field B causes the electron beam to move in a path with a radius of curvature R.

1-What is the relationship between e/m and R? Calculate R.

Data: $E = 5.10^4$ volts.m^{-1} $B = 10^{-3}$ Tesla

l =10 cm $y_0 = 1.76$ cm

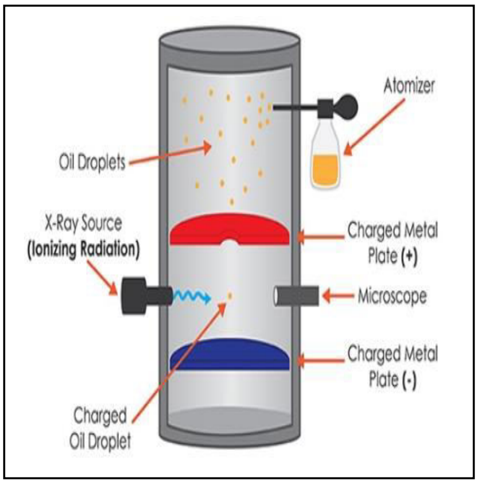
Exercise 2: (Millikan's experiment)

1. Millikan's apparatus is used to track the free fall of an oil droplet in air. The velocity of the droplet is 0.32 mm/s. While neglecting the upward thrust due to the surrounding air, calculate the radius and mass of the droplet.

2. When the droplet is subject to an electric field E equal to 4.5 10^5 volts.m^{-1} across the parallel plates, it rises and its velocity reaches 0.118 mm/s. Calculate the charge **e** carried by the droplet.

3. During the experiment the charge of the droplet varies and can be immobilized between the plates when the electric field E' is equal to $2.47.10^5$ volts/m.

- Calculate the new charge of the droplet.
- From all these measurements, find the value of the elementary charge.



Data:

-Specific mass of oil used $\varphi = 1.26 \times 10^3$ kg/m^3

-Acceleration of gravity $g = 9.81$ m/s^2

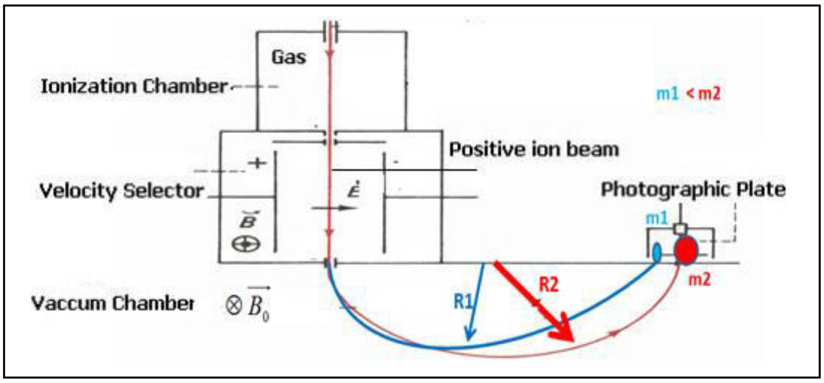
-Coefficient of viscosity of air:

$\eta = 1.82 \times 10^{-5}$ decapoise (SI).

Exercise 3: (Bainbridge spectrograph, isotopes)

The Bainbridge spectrograph is used to separate isotopes of the element **x**. The carbon atom $^{12}_{6}C$ is used as a reference element knowing that it is lighter than the isotopes of **x**. Four points of impact of the ionized element **x** on the photographic plate are observed at distances of 32 cm, 64 cm, 66.64 cm and 69.32 cm, respectively.

1. Find the relationship between the distance between the impact points and the velocity selector outlet and the isotope masses according to the diagram below.
2. Calculate the mass of each isotope.
3. Calculate the average atomic mass of element **x**, knowing that the relative abundance of the isotopes is : 77.60 %, 11.11 % and 11.29 % from the lightest to the heaviest isotope, respectively.



Composition of the atom

Exercise 1:

Give in table form the mass number A, atomic number Z, number of electrons e^-, number of neutrons N of the following chemical species:

$$^{40}_{18}Ar \; ; \; ^{16}_{8}O^{2-} \; ; \; ^{35}_{17}Cl^{-} \; ; \; ^{137}_{56}Ba \; ; \; ^{23}_{11}Na^{+} \; ; \; ^{207}_{82}Pb^{2+}$$

Exercise 2:

Find the average atomic mass of chlorine, knowing that two isotopes of chlorine occur in nature. Their abundances and masses are listed in the following table.

Isotope	Abundance (%)	Mass (amu)
$^{35}_{17}Cl$	75.4	34.96
$^{37}_{17}Cl$	24.6	36.96

Exercise 3 :

1. What is the composition of the nucleus of the uranium isotope 235 (U-235), symbol $^{235}_{92}U$?
2. Calculate the mass defect of this nucleus, in amu and kg.
3. What is the binding energy of this nucleus in joules and MeV?
4. Calculate its energy per nucleon.
5. Compare the stability of the U-235 nucleus with that of radium 226 (Ra-226), whose binding energy per nucleon is 7.66 MeV.

Data: Atomic mass of uranium 235 **:** m ($^{235}_{92}U$) = 234.99332 amu ; mass of neutron: m_n = 1.00866 amu ; mass of proton: m_p = 1.00728 amu ; speed of light: c = 3 x 10^8 m/s.

Radioactivity

Exercise 1:

A sample of radioactive radon $^{222}_{86}Rn$ weighs 1 g. The decay of a nucleus of radon involves the loss of an α particle. The half-life of $^{222}_{86}Rn$ is T= $t_{1/2}$ = 3.8 days.

1. Write balanced nuclear equation for the decay of $^{222}_{86}Rn$.
2. Calculate the radioactive decay constant of $^{222}_{86}Rn$.
3. How many radioactive nuclei are present in the initial sample?
4. What is the activity of this sample? What will it be after 15 days?

Exercise 2:

The reaction of a lithium nucleus with a deuterium nucleus is given by:

$$Li_3^6 + D \rightarrow 2\,\alpha$$

1. Write the balanced equation for this reaction. What type of nuclear reaction is it: Transmutation, Fusion or Fission?
2. Calculate the energy released in joules during this reaction.

Data: ^2H = 2.0135 amu, ^4He = 4.0015 amu, ^6Li = 6.0135 amu, 1amu =1.66.10^{-27} kg, c = 3.108 m s^{-1}.

Exercise 3:

Thorium $^{237}_{90}Th$ is radioactive and undergoes α decay. It has a half-life of 18 days. Initially (at t = 0), we have a sample of thorium with mass m_0 = 1 μg from a radioactive source.

1. Write balanced nuclear equation for the decay of $^{237}_{90}Th$ knowing that it yields radium Ra.
2. Calculate the radioactive constant λ of thorium.
3. Calculate the mass of thorium remaining at t_1 = 36 days and t_2 = 6 months.
4. How long does it take for Thorium to decay to a mass = 0.0156 μg?

The Bohr Model of the Atom

Exercise 1:

When a light beam of wavelength 300 nm strikes a metal surface, electrons are emitted from the metal. The kinetic energy of the electrons emitted is 0.74 eV.

1. What is the maximum speed of the electrons emitted?
2. Calculate the extraction energy in joules and eV.
3. Calculate the threshold frequency ν_0 and threshold wavelength λ_0 .

Given: h = 6.62×10^{-34} J s; m_e= 9.1×10^{-31} kg; c = 3×10^{8} m s^{-1}

Exercise 2:

Bohr's theory is applied to an electron revolving around the nucleus of the hydrogen atom in the third circular orbit (n = 3).

1. Calculate the radius of Bohr's orbit in Å.
2. Calculate the energy level of the atom in eV.
3. Calculate the ionizing wavelength and deduce the corresponding energy in eV and in joules.

Given : Rydberg constant : R_H= 1.1×10^{7} m^{-1}

Exercise 3 :

1. Calculate the radius of the first orbit (n = 1) for the Li^{2+} ion.
2. Calculate the energy of the third ionization of the lithium atom.
3. Calculate the wavelength and the frequency of the corresponding line between energy levels n = 4 and n = 1.
4. What is the minimum energy value that the Li^{2+} ion must absorb to move from the ground state to the excited state?

5. What is the frequency for: a) the limiting line in the Lyman series; b) the second Balmer line?

Periodic table

Exercice 1 :

1. Are the following series of quantum numbers characterizing an electron possible ?

A/ n = 2, ℓ = 0, m = 0 ; **B/** n = 2, ℓ = 1, m = 1 ; **C/** n = 2, ℓ = 2, m = 0 ; **D/** n = 3, ℓ = 2, m = -1 ; **E/** n = 1, ℓ = 0, m = 1.

2. What is the maximum number of electrons described by the following (n ; ℓ ; m ; s):

A/ n = 4. **B/** n = 3 et ℓ = 2. **C/** n = 2 et ℓ = 1. **D/** n = 0, ℓ = 0 et m = 0. **E/** n = 2, ℓ = 1 et m = -1.

Exercice 2 :

Give the position (period and group) of the following elements in the periodic table :

Na_{11}, Al_{13}, Sr_{38}, Sc_{21}, Cu_{29}, Zn_{30}, Ge_{32}, I_{53}.

Exercice 3 :

An element X belongs to the boron column (B_5) and the potassium period (K_{19}).

1. Determine its atomic number ?
2. Give the quantum numbers for the valence electrons of X ?
3. Classify these elements (X, B_5 and K_{19}) in order of increasing radius ?
4. Calculate the radius and energy of Chlorine (Cl_{17}) according to Slater ?

Chemical bonds

Exercise1 :

Represent, according to the Lewis model, the following elements :

H, He, C, N, F, Ne, Li, Be.

Exercise 2 :

The electronegativity of the H, F, Cl and K atoms are successively : (2.2), (4), (3.1), (0.8). Predict the main character (ionic, polar, covalent) of the molecules' bonds :

K-F ; H-F ; K-Cl ; H-Cl et H-H.

Exercise 3 :

The butadiene molecule $CH_2 = CH\text{-}CH = CH_2$ is planar.

1. What are the hybridization states of the carbon atoms in this molecule ?

2. Respecting the values of the angles between the bonds, what are the two possible geometric forms of butadiene ?

Answers to application exercises

Preparation of solutions

Exercise 1:

***Pure substance*:** It is a substance made up of only one kind of particle and has a constant structure.

a) Element: It is a substance that consists of only one kind of atom. Example: hydrogen $\mathbf{H_2}$.

b) Compound: It is a substance made up of different elements in definite proportions that can be separated using chemical techniques. Example: water H_2O**.**

Mixture: **It is a combination of two or more substances in any proportion.** Mixtures can be classified into two main categories: homogeneous and heterogeneous.

a) Homogeneous mixture: It is a mixture of two or more chemical substances (elements or compounds) where the different components cannot be visually distinguished. They can be classified into two categories: colloids and solutions.

  a)-1 Colloid: It is a mixture appearing homogeneous but in which we can distinguish at least two substances under a microscope. Example: whey (لبن).

  a)-2 Solution: It is a homogeneous mixture in which the components cannot be distinguished even under a microscope . Example: water + sugar. See Figure 1.

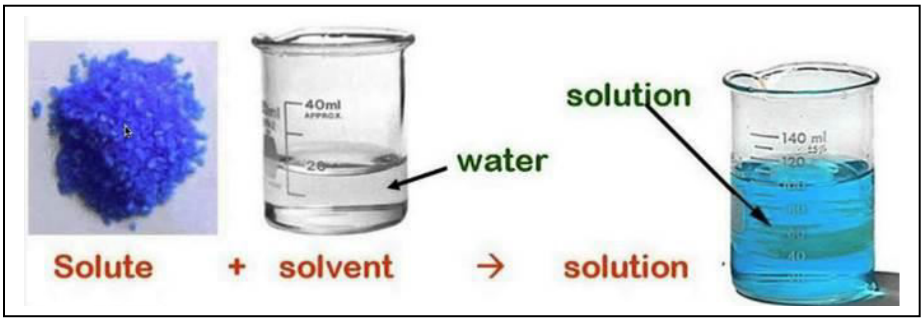


Figure 1. Example of a solution (homogeneous mixture)

Solutions can be separated into two components;

a) Solute (s): A substance dissolved in the solvent (minor component).
b) Solvent: A substance capable of dissolving one or more solutes (major component).

If the solute is solid, the process is called dissolution, if it is liquid, it is called dilution.

A solvent is usually a liquid that is capable of dissolving and diluting other substances without modifying them chemically and without modifying itself. Example: water.

A reactant is a starting material in a chemical reaction that is consumed to make products. Their amount in the system decreases with reaction time. In a chemical reaction, reactants reorganize by breaking one set of chemical bonds and forming a new set (formation of new molecules).

Mole: The mole is defined as the amount of a substance that contains exactly 6.02214179 x 10^{23} elementary entities (atoms, ions, molecules, etc.). The number 6.02214179 x 10^{23} is called Avogadro's number.

Molar mass: It is the mass of one mole of a chemical compound. Unit: g/mol or g mol^{-1}.

The number of moles n (mol) = given mass m (g) / molar mass M (g/mol).

Example: oxygen O^{16} : M = 16 g/mol: i.e., the mass of 1 mole of oxygen atoms is 16 g.

Molecular mass: It is the mass of one molecule. It is equal to the sum of the atomic masses of the atoms in the molecule. It is expressed in unified atomic mass units (1 amu: 1/12 of the mass of one atom of the isotope carbon-12). 1 amu = 166 x 10^{-27} kg.

Density (of a substance): It is the mass per unit volume of that substance. Symbol: $\boldsymbol{\rho}$. $\boldsymbol{\rho = \frac{m}{V}}$.

Example: ρ_{water} = 1 g/mL, i.e., the mass of 1 mL of water (temperature ?) is 1 g.

Specific gravity (dimensionless) : It is the ratio of the density of a substance with that of water (999.975 kg/m^3 almost 1000 kg/m^3 or kg/L) at 4 °C. $SG = \frac{\rho\ substance}{\rho\ water}$

Mass concentration: The mass concentration of a solution is the mass of a solute per unit volume of solution. $\boldsymbol{\rho_i(\frac{g}{l}) = \frac{m_i}{V} = \frac{mass\ of\ solute}{volume\ of\ solution}}$ **;** SI unit : kg m^{-3}, SI equivalent units : g L^{-1}, g dm^{-3}, mg mL^{-1}.

Molar concentration (or Molarity): It is the number of moles of solute present in one liter of solution. Symbol : M . $\boldsymbol{M\ (\frac{mol}{L}) = \frac{n\ (mol)}{V\ (L)}}$

Normality : It is the number of gram-equivalents (equivalent weights) of solute present in one liter of solution. Symbol : N. One gram-equivalent is the amount of a substance that can react with one mole of protons (for a base), or with one mole of hydroxide ions OH^- (for an acid), or with one mole of electrons (for an oxidizing agent or reducing agent).

$N\left(\frac{Eq}{L}\right) = M\ .\ K$ **where K is the number of equivalents per mole and M the molarity.**

$$\text{If } K = 1, \text{then normality} = \text{molarity}.$$

For simple salts, such as KCl, Na_2SO_4, $AlCl_3$ and $Ca_3(PO_4)_2$, the value of K is the total valency of the cation or the anion. For KCl, K = 1; for $AlCl_3$, K = 3; for $Ca_3(PO_4)_2$, K = 6.

Molality: It is the number of moles of a solute dissolved in 1000 g of solvent. Symbol: m

Mole fraction x (dimensionless): It is the number of moles ni of component **i** in a solution divided by the total number of moles n_t of all the components in that solution.

$$x_i = \frac{n_i}{n_t} \quad ;\ note\ that \qquad \sum x_i = 1$$

Exercise 2 :

1. **Preparation of 100 mL of a 1 g/L solution of NaOH (sodium hydroxide)**

Since the mass concentration of the solution is 1 g/L, we first convert the volume from L to mL:

1 L= 1000 mL

To prepare a volume of 1000 mL of the solution, we dissolve 1 g of sodium hydroxide and to prepare 100 mL (1/10 of one liter), we have to place a mass of NaOH m = $\frac{100\ mL\ x\ 1\ g}{1000\ mL}$ = $0.1\ g$ in a 100 mL volumetric flask, dissolve with water and make up to the mark with the same solvent (water).

2. **Preparation of 2 L of a 1 mol/L solution of NaOH**

Converting from mol/L of NaOH to g/L:

$$C_{molar}\left(\frac{mol}{L}\right) = \frac{Mass\ concentration\ \left(\frac{g}{L}\right)}{Molar\ mass\ \left(\frac{g}{mol}\right)} \Rightarrow Mass\ concentration$$

$$= C_{molar}\ x\ Molar\ mass \ = 1\ \frac{mol}{L}\ x\ 40\frac{g}{mol} = 40\ \frac{g}{L}$$

1 L of solution contains 40 g, to prepare 2 L, we weigh a mass of NaOH $m = \frac{2\ L\ x\ 40g}{1\ L} = 80\ g$, dissolve it in water and make up the volume with water.

3. **Preparation of 500 mL of a 1 N solution of NaOH**

Converting from normality to mass concentration knowing that the equivalent weight (gram equivalent) of NaOH is 40 g/eq since it dissociates into Na^+ and OH^- where the number of equivalents of OH^- is K = 1 (a mass of 40 g of NaOH reacts with one mole of protons or acidity of NaOH is 1):

$$N\left(\frac{eq}{L}\right) = Molar\ concentration\ x\ \ K = \frac{Mass\ concentration\left(\frac{g}{L}\right)}{Molar\ mass\ \left(\frac{g}{mol}\right)}\ x\ K \Rightarrow$$

$$Mass\ concentration\ = N\ x\ \frac{Molar\ mass}{K} = 1\ N\ x\ \frac{40\ g/mol}{1} \Rightarrow Mass\ conentration\ = 40\ g/L.$$

To prepare 1 L of a 1 N solution of NaOH, we weigh 40 g of NaOH and to prepare 500 mL of this solution, we need to find the mass of NaOH required:

40 g of NaOH → 1 L =1000 mL of solution

m of NaOH → 500 mL of solution?

$$\Rightarrow m = \frac{500\, mL\, x\;\; 40g\;\; \text{NaOH}}{1000\, mL} = 20\; g\;\; \text{NaOH}$$

Therefore, we weigh a mass of 20 g of sodium hydroxide, dissolve it in water and make up to 500 mL with water.

4. **Volume of 96% (w/w) sulfuric acid (H_2SO_4), with a specific gravity of 1.84 required to prepare 250 mL of 1 N sulfuric acid**

Likewise, converting from normality to mass concentration:

$$N\left(\frac{eq}{L}\right) = Molar\ concentration\ x\ \ K = \frac{Mass\ concentration\left(\frac{g}{L}\right)}{Molar\ mass\left(\frac{g}{mol}\right)}\ x\ K \Rightarrow$$

$Mass\ concentration\ = N\ x\ \frac{Molar\ mass}{K} = 1\ N\ x\ \frac{98\ g/mol}{2} \Rightarrow Mass\ conentration\ = 49\ g/L$, as K = 2 for H_2SO_4

We can first determine the normality of the concentrated solution:

In a 100 g of solution, there are $\frac{96}{49}$ = 1.959 equivalent weights of H_2SO_4, corresponding to a volume of solution of $\frac{100}{1.84} = 54.347\ mL.$

N = $\frac{1.959}{54.347}\ x\ 1000 = 36.04\ eq/L$

To find the volume required, we can use the law of equivalence: $N_1V_1 = N_2V_2$.

Where N_1 = 36.04 N ; V_1 = ? N_2 = 1 N ; V_2 = 250 mL

$V_1 = \frac{N_2\ V_2}{N_1} = \frac{1\ x\ 250}{36.04} = 6.94\ mL\ \approx\ 7\ mL$

So, place about 200 mL of (distilled) water in a volumetric flask, add to this 7 mL of concentrated acid (96%), stir and make up to 250 mL with water.

Exercise 3:

4. **Number of copper atoms in 0.25 mol of copper**

$6.022\ x\ 10^{23}$ atoms of Cu → 1 mol of Cu

N of atoms of Cu ? → 0.25 mol of Cu

N = $\frac{0.25\ x\ 6.022\ x\ 1023}{1}$ = $1.505\ x\ 10^{23}$ atoms.

5. **Number of glucose molecules $C_6H_{12}O_6$ in 0.75 mol of glucose**

1 mol of glucose → $N_A = 6.022\ x\ 10^{23}$ molecules

0.75mol of glucose → N = ? ⇒ N= $4.517\ x\ 10^{23}$ molecules.

6. Number of atoms in 0.75 mol of glucose

First, we sum up the numbers of atoms in one molecule of $C_6H_{12}O_6$: 6 C atoms + 12 H atoms + 6 O atoms = 24 atoms.

Nt = N x 24 = 4.517 x 10^{23} = 1.084 x10^{25} atoms

Exercise 4 :

A mass of 51.3 g of aluminum sulfate $Al_2(SO_4)_3$ is dissolved in 500 mL of water (H_2O).

5. Dissolution equation

$Al_2(SO_4)_3$ (s) → 2 Al^{3+} (aq) + 3 SO_4^{2-} (aq)

6. Molarity, molality, mass concentration and normality of the prepared solution.

a) Molarity (mol/L)

Molar mass of $Al_2(SO4)_3$ = M ($Al_2(SO_4)_3$) = 2 (27) +3 (32 + 4 x 16) = 342 g/mol

Number of moles: n = m/M = 51.3/342= 0.15 mol

Mass of solution: mass of solute + mass of solvent: 51.3 + 500 = 551.3 g (aqueous solution assumed to be same as water) so:

ρ = m/v ⇒ v = m/ρ = 551.3 g/1 g mL^{-1} = 551.3 mL = 0.5513 L

⇒ Molarity : $\frac{n}{V} = \frac{0.15}{0.5513}$ = 0.27 mol L^{-1} = 0.27 M

b) Molality (mol/kg)

Moles of solute in 1000 g of solvent

n = 0.15 mol in 500 mL of water, i.e., 500 g of water

⇒ Molality = 0.15 mol/ 500 g = 0.3 mol kg^{-1} = 0.3 m

c) Mass concentration (g/L)

Mass of solute in 1 L of solution

51.3 g → 0.5513 L

C $= \frac{51.3}{0.5513}$ = 93.05 g/L

d) Normality

N = number of equivalent weights (gram equivalents) of solute present in 1 liter of solution.

Aluminum sulfate dissociates in aqueous solution as follows:

$Al_2(SO_4)_3$ (s) → 2 Al^{3+} (aq) + 3 SO_4^{2-} (aq)

For $Al_2(SO_4)_3$ K, the number of equivalents per mole is 6 since the total valency of the cation (or the anion) is 6.

So N = M x K = 0.27 (mol L^{-1}) x 6 eq/mol= 1.62 N

7. Molar concentration of each ion species in the solution.

$$\mathbf{1}\ Al_2(SO_4)_3\,(s) \rightarrow \mathbf{2}\ Al^{3+}\,(aq) + \mathbf{3}\ SO_4^{2-}\,(aq)$$

1 mol $Al_2(SO_4)_3$ → **2** mol Al^{3+} and **3** mol SO_4^{-2}

So : $[Al^{3+}]$ = 2 M = 2 x 0.27= 0.54 mol/L

$[SO_4^{2-}]$ = 3 M = 3 x 0.27= 0.81 mol/L

8. Volume to be taken from the initial solution to make 100 mL of 0.15 mol/L $Al_2(SO_4)_3$ solution

We can use the dilution equation:

C_1 x V_1 = number of moles = C_2 x V_2 where :

C_1 is the concentration of the original undiluted solution.

V_1 is the volume of the original undiluted solution.

C_2 is the concentrated of the diluted solution after it has been mixed with a solvent (water).

V_2 is the volume of the diluted solution after it has been mixed with a solvent (water).

$$V_1 = \frac{C_2\ x\ V_2}{C_1} = = \frac{0.15\ x\ 0.100}{0.27} = 0.05555 \text{ L} = 55.55 \text{ mL}$$

Components of Matter

Exercise 1 : J.J. Thomson Experiment

A. The direction of vectors v_0, E, Fe, B and F_B is shown in Figure 1.

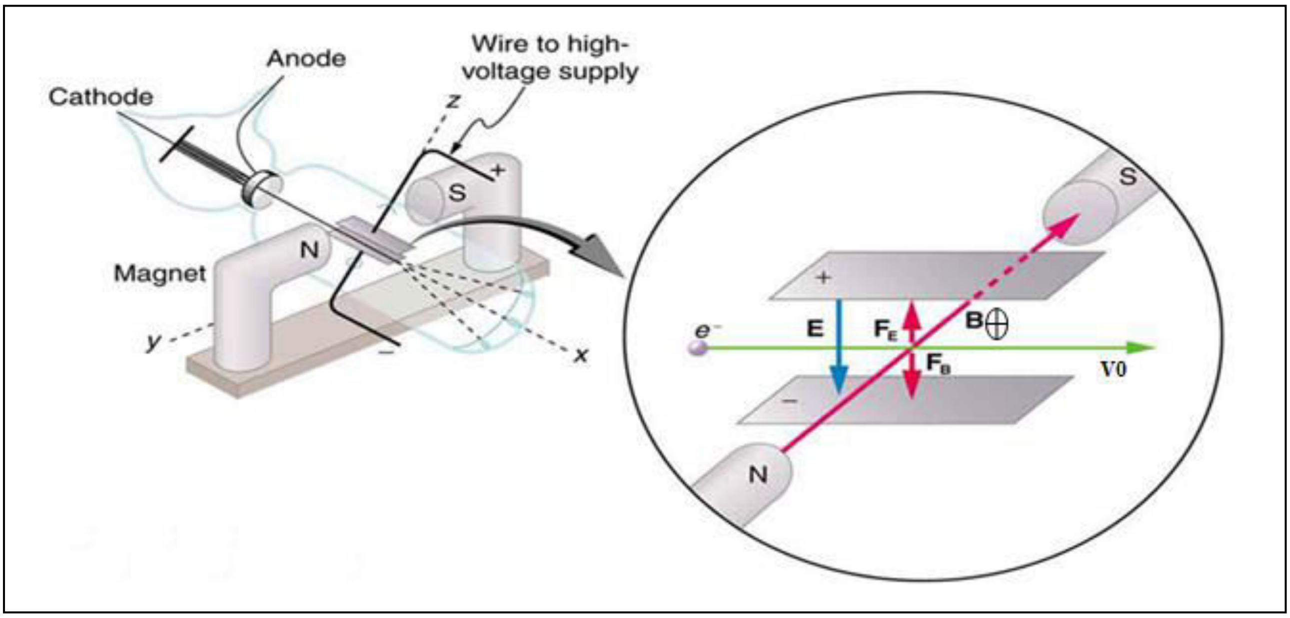


Figure 1. Direction of Vectors v_0, E, Fe, B and F_B

B.

1. Expression of the electron beam velocity v :

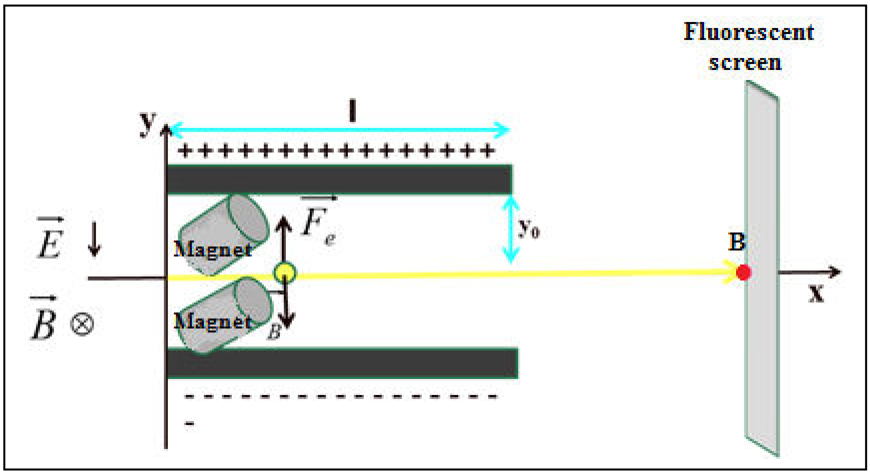


Figure 1 : Simultaneous action of the electric and magnetic fields E and B on the electron beam

The electric and magnetic fields are adjusted so that the electron beam does not deviate (yellow line) → Motion is uniform rectilinear (zero acceleration) up to point **B** $\Rightarrow \vec{a} = 0$. Therefore, the electric and magnetic forces are equal and opposite.

Magnetic force = F_B = qvB ; Electric force F_e= qE

$F_B = F_e$ therefore qvB = qE so v = $\frac{E}{B}$

$$\sum \vec{F} = m.\vec{a} = 0 \Rightarrow F_e - F_B = 0 \Rightarrow F_e = F_B \Rightarrow q.E = q.V.B \Rightarrow v = \frac{E}{B} \quad (1)$$

2- Charge-to-mass ratio e/m of the electron

Electric field E on ; Magnetic field off (Figure 2)

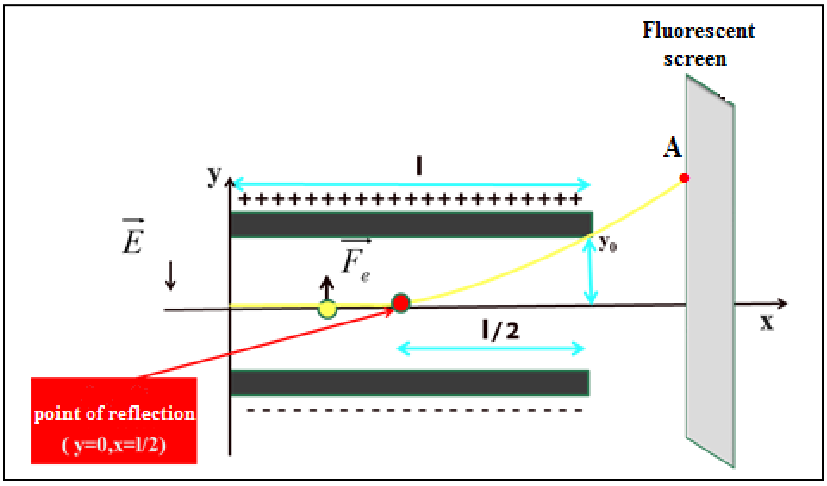


***Figure 2** : Action of electric field only*

The electron beam is deflected towards the positive plate (+) by the electric field E (yellow line).

In the x-direction, no force is exerted. Applying Newton's second law: the net force equals mass times acceleration:

$$\sum \vec{F} = m.\vec{a}$$

$$\sum \vec{F} = m.\vec{a} = 0 \Rightarrow a_x = 0 \Rightarrow x = v.t \qquad (2)$$

Therefore, the motion is uniform rectilinear.

In the y-direction:

$$\sum \vec{F} = m.\vec{a} \Rightarrow F_e = m.a_y = q.E \Rightarrow a_y = \frac{q.E}{m} = CST \qquad (3)$$

Since acceleration is constant, the motion of the electron beam is uniformly accelerated:

$$\Rightarrow y = \frac{1}{2} a_y \ t^2 \qquad (4)$$

Substituting (2) and (3) into (4)

:

$$\Rightarrow y = \frac{1}{2} a_y \left(\frac{x}{v}\right)^2 \Rightarrow y = \frac{1}{2}\left(\frac{q.E}{m}\right)\left(\frac{1}{v}\right)^2 .x^2 \Rightarrow y = f\left(x^2\right)$$

⇒ the beam moves in a parabolic path.

Note that at the exit of the capacitor x = l, y = y_0 and v=E/B.

$$\Rightarrow y_0 = \frac{1}{2}\left(\frac{q}{m}\right)\left(\frac{B^2}{E}\right).l^2$$

This is the deflection equation at the exit of the capacitor, from which the charge to mass ratio $\frac{q}{m}$ can be calculated:

Therefore, $y_0 = \frac{1}{2}\left(\frac{q}{m}\right)\left(\frac{B^2}{E}\right).l^2 \Rightarrow \frac{q}{m} = \frac{2y_0 E}{.l^2 B^2} = \frac{2.1.76.10^{-2}.5.10^4}{.(10^{-1})^2.(10^{-3})^2} = 1.76\,x10^{11}\ C/kg$

C- Electric field E off ; Magnetic field on

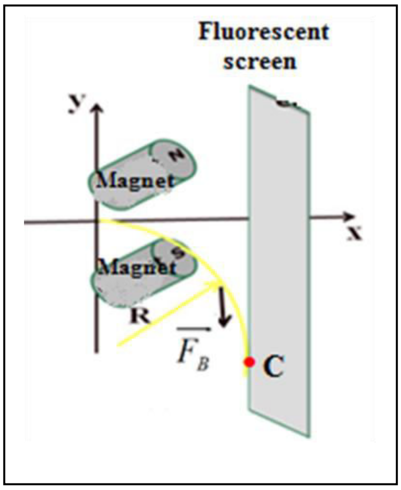
The magnetic force acts as a centripetal force and causes the beam to move in a circular path with a radius of curvature R (Figure 3).

1-Relationship between q/m and R

Action of the B field alone (figure 3) :

- The magnetic field B causes the electron beam to move (yellow line) in a circular path with radius R. Applying Newton's Second law:

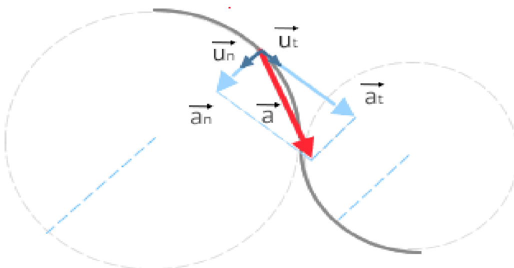
$$\sum \vec{F} = m.\vec{a}$$



a_N is the normal or centripetal acceleration of the beam.

a_T is the tangential acceleration.

***Figure 3**: Action of Magnetic field only*.



For a_T : no force is exerted.

$$\Rightarrow \sum \vec{F} = m.\vec{a} = 0$$

For a_N : We can compute the value of R as follows:

$$\Rightarrow \sum \vec{F} = F_B = m.a_N \Rightarrow q.v.B = m.\frac{v^2}{R} \Rightarrow q.B = m.\frac{v}{R}$$

$$\Rightarrow R = \frac{m}{q}.\frac{v}{B} \Rightarrow R = \frac{1}{\frac{q}{m}}.\frac{E}{B^2} \Rightarrow R = \frac{1}{1.76.10^{11}}.\frac{5.10^4}{(10^{-3})^2} = 0.28\,m$$

Exercise 2 : (Millikan's oil drop experiment)

1. Calculation of radius and mass of droplet

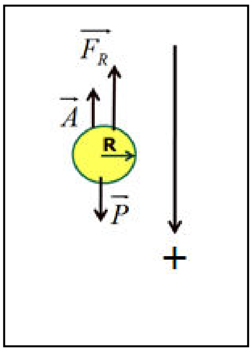


Figure 1 : représentation schématique de la chute libre de la gouttelette

a) Radius of the droplet

As the electric field E is off, the droplet does not experience an electric force and begins to fall.

- **Weight (W or p)**

$$p = m.g = (\rho.V).g = \left(\rho.\frac{4}{3}\pi.R^3\right).g$$

- **Viscous drag force** F_R (calculated using Stokes'law) : $F_R = 6.\pi.\eta.R.v_0$

- **Buoyant force** (Archimedes principle) : $A = \left(\rho_0.\frac{4}{3}\pi.R^3\right).g$

where :

ρ : is the density of the oil,

ρ_0 : the density of the surrounding air,

R : the radius of the oil droplet (assumed to be a sphere of radius R),
v_0 : Terminal velocity of the droplet,

η : the coefficient of viscosity (or viscosity) of the oil.

The droplet reaches rapidly its terminal velocity:

$$\Rightarrow v = CST \Rightarrow a = 0 \Rightarrow \sum \vec{F} = m.\vec{a} = 0.$$

$$So : p - (A + F_R) = 0 \Rightarrow W - A = F_R \qquad (1)$$

$$\Rightarrow \left(\rho.\frac{4}{3}\pi.R^3\right).g - \left(\rho_0.\frac{4}{3}\pi.R^3\right).g = 6.\pi.\eta.R.v_0$$

$$\Rightarrow (\rho - \rho_0).\left(\frac{4}{3}\pi.R^3\right).g = 6.\pi.\eta.R.v_0$$

$$\Rightarrow R = 3.\sqrt{\frac{\eta.v_0}{2(\rho - \rho_0).g}} = 3.\sqrt{\frac{1.82.10^{-5}.\,0.32\,.10^{-3}}{2.\left(1.26.10^3 - 0\right).9.81}} == 1.45.10^{-6}\,m$$

b) Mass of the droplet

$$m = \rho.V = \rho\,\frac{4}{3}.\pi.R^3 = 1.26\ x\,10^3\ x\frac{4}{3}x\ 3.14\ x\left(1.45\ x10^{-6}\right)^3 = 1.61.x\,10^{-14}\,kg$$

2. Calculation of the charge on the droplet

The effect of the electric field E on the rising droplet with velocity v_1

(Figure 2), results in a net force:

$$So : \quad \sum \vec{F} = m.\vec{a}$$

$$\Sigma F = -p - F_{R_1} + A + F_e = -\frac{4}{3}\pi R^3 \rho g - 6\pi\eta\,Rv + \frac{4}{3}\pi R^3 \rho_0 g + qE$$

In equilibrium, the net force is zero, $\sum F = 0$ and the drop has a terminal

velocity v_1:

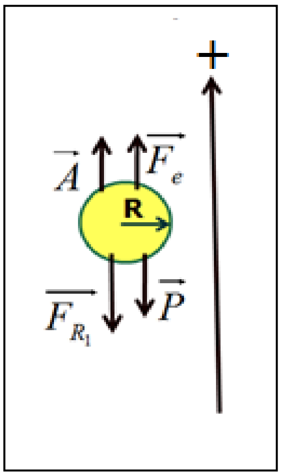


Figure 2 : Schematic representation of the rise of the droplet

$$0 = -\left(\frac{4}{3}\pi R^3 \rho g - \frac{4}{3}\pi R^3 \rho_0 g\right) - 6\,\pi R\,\eta\,v_1 + qE = -6\,\pi\,\eta\,R\,v_0 - 6\,\pi\,\eta\,R\,v_1 + qE$$

$$\Rightarrow 6.\pi.\eta.R.v_0 + 6.\pi.\eta.R.v_1 = q.E$$

$$\Rightarrow q = \frac{6.\pi.\eta.R}{E}.(v_0 + v_1) = \frac{6\ x\ 3.14.x\ 1.82x\ 10^{-5}\ x\ 1.45x\ 10^{-6}}{4.5\ x10^{5}}.(0.32 + 0.118)x10^{-3} = 4.8.x\ 10^{-19}\ C$$

3. Calculation of the new charge q' on the droplet and the elementary charge

- ***New charge on droplet***

The droplet is held stationary (Figure 3) ⇒ No friction (drag force) $F_R = 0$

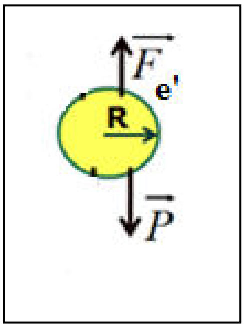


Figure 3 : Schematic representation of a droplet held stationary

Therefore

$$p - F_{e'} = 0 \Rightarrow p = F_{e'} \Rightarrow m.g = q'.E'$$

$$\Rightarrow q' = \frac{m.g}{E'} = \frac{1.61\ x\ 10^{-14}\ x\ 9.81}{2.47.x\ 10^{5}} \approx 6.4\ 10^{-19}\ C$$

- ***Value of the elementary charge*** **:**

(If charge is quantized, the charges on the droplet (or series of droplets) should each be an integer multiple n of the electron charge e (elementary charge).

So the ratio $\frac{q}{e}$ **should be an integer.**

From the first experiment q = 4.8 x 10^{-19} C → $\frac{q}{e} = \frac{4.8\ x\ 10^{-19}}{e}$

From the second experiment q' = 6.4 x 10^{-19} C → $\frac{q'}{e} = \frac{6.4\ x\ 10^{-19}}{e}$

n = $\frac{q}{e}$ **and n' =** $\frac{q'}{e}$

$= \frac{4.8\ x\ 10^{-19}\ n}{6.4\ x\ 10^{-19}\ n'}$

$\frac{n}{n'} = 0.75 = \frac{3}{4}$

In the first experiment, the charge on the droplet is 3 times the elementary charge and in the second experiment, the charge is 4 times the elementary charge.

We calculate the elementary charge to be : 4.8 x 10^{-19}/3 = 1.6 x 10^{-19} C.)

$$\Rightarrow q' - q = 6.4.10^{-19}\ C - 4.8.10^{-19}\ C = 1.6.10^{-19}\ C$$

So :

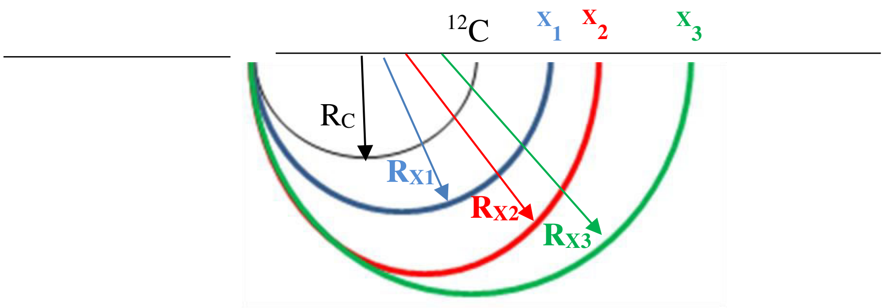
$$\Rightarrow q' = 6.4.10^{-19}\ C = 4\left(1.6.10^{-19}\right)C = 4.e$$

$$and \Rightarrow q = 4.8.10^{-19}\ C = 3.\left(1.6.10^{-19}\right)C = 3.e$$

The droplet carries 3 to 4 elementary charges (e = 1.6. 10^{-19} C).

Exercise 3 : (Bain Bridge spectrograph, isotopes)

1. ***Find the relationship between the distance between the impact points and the velocity selector exit and the isotope masses according to the diagram below.***



In the analyzer :

Only magnetic field B_0 is on : The ion is deflected and moves in a circular path of radius R :

$$we \quad have \quad \sum \vec{F} = m.\vec{a}$$

In tangential direction a_T: No force is exerted: $\Rightarrow \sum \vec{F} = m.\vec{a_T} = 0$

In normal direction a_N $\Rightarrow \sum \vec{F} = F_{B_0} = m.a_N \Rightarrow q.V.B_0 = m.\frac{V^2}{R} \Rightarrow q.B_0 = m.\frac{V}{R}$

$$\Rightarrow R = \frac{m}{q}.\frac{V}{B_0} \Rightarrow \frac{m}{R} = \frac{q.B_0.B}{E} = CST$$

$$So: \frac{m}{R} = \frac{q.B_0.B}{E} = CST \Rightarrow \frac{m_C}{R_C} = \frac{m_{X1}}{R_{X1}} = \frac{m_{X2}}{R_{X2}} = \frac{m_{X3}}{R_{X3}} = CST$$

2. Calculation of the mass of each isotope

$$We\ have \quad \frac{m_C}{R_C} = \frac{12}{32cm} = \frac{m_{X1}}{64cm} = \frac{m_{X2}}{66.64cm} = \frac{m_{X3}}{69.32cm}$$

$$\Rightarrow m_{X1} = 24\ g/mol \quad ; \quad m_{X2} = 25\ g/mol \quad ; \quad m_{X3} = 26\ g/mol$$

3. Calculation of the average atomic mass of element x, knowing that the relative abundances of the isotopes are 77.60 %, 11.11 % and 11.29 % from the lightest to the heaviest, respectively.

$$Maverage = \frac{\sum Mi.Xi}{\sum Xi} = \frac{M_1.X_1 + M_2.X_2 + M_3.X_3}{100} = \frac{24 \text{ x } 77.60 + 25.11 \text{ x } 11 + 26 \text{ x } 11.29}{100} = 24.33\ g/mol.$$

Composition of the atom

Exercise 1:

The symbol for a chemical element **X** is : ${}^{A}_{Z}X$ where A is the mass number or number of nucleons, Z is the atomic number or number of protons. The number of protons (positively charged) Z is equal to the number electrons e^- (negatively charged). The number of neutrons N is A-Z. In an extended symbol, additional details may be added as superscripts or subscripts for an ion, isotope, etc. The charge is represented as a superscript on the right side of the element symbol: ${}^{A}_{Z}X^{charge}$. The number of electrons in the species e^- = Z ± charge on the ion. The following table regroups the detailed information asked for:

	A	Z	N = A-Z	e^- = Z ± charge on the ion
${}^{40}_{18}Ar$	40	18	40-18=22	18±0=18
${}^{16}_{8}O^{2-}$	16	8	16-8=8	8+2=10
${}^{35}_{17}Cl^{-}$	35	17	35-17=18	17+1=18
${}^{137}_{56}Ba$	137	56	137-56=81	56±0=56
${}^{23}_{11}Na^{+}$	23	11	23-11=12	11-1=10
${}^{207}_{82}Pb^{2+}$	207	82	207-82=125	82-2=80

Exercise 2 :

The average atomic mass is calculated as follows:

$$M_{average} = \sum \frac{x_i \, . M_i}{100} = \frac{x_1 \, . \, M_1 + x_2 \, . \, M_2}{100}$$

where x_i is the abundance of isotope i and M_i its atomic mass. Therefore, the average atomic mass of chlorine is:

$$M_{average} = \frac{75.4 \, x \, 34.96 + 24.6 \, x \, 36.96}{100} = 35.45 \, amu$$

Exercise 3 :

1. *Composition of the nucleus of* ${}^{235}_{92}U$:

A = 235; Z = 92; N = A-Z = 235-92 = 143; e^- = Z ± charge on the ion= 92 ± 0 = 92.

2. *Mass defect of* ${}^{235}_{92}U$ *nucleus :*

$$\Delta m = |m_{theroetical} - m_{real}|$$

$$m_{theoretical} = Z.m_{proton} + N.m_{neutron} = 92 \, x \, 1.00728 + 143 \, x \, 1.00866$$

$$= 236.90814 \, amu$$

$$\Delta m = |236.90814 - 234.99332| = 1.91482 \, amu = 1.91482 \, x \, 1.66 \, . \, 10^{-27} \, kg$$

$$= 3.1786 \, x \, 10^{-27} \, kg/nucleus.$$

3. *Binding energy of the nucleus*

1) *In joules*

$$E_b = \Delta m \, . C^2 = 3.1786 \;\; x\; 10^{-27} \; x \; (3 \; x \; 10^8)^2 = 2.86 \; x \; 10^{-10} \; J/nucleus$$

2) *In MeV* :

Converting from joules to MeV:

Knowing that 1 eV = 1.66 x 10^{-19} J,

we then have:

$$E_b = \frac{2.86\,.\,10^{-10}}{1.66\,.\,10^{-19}} = 1.72289 \; x \; 10^9 \; eV = 1722.89 \; MeV$$

4. *Binding energy per nucleon of the nucleus of U-235:*

$$E_{b\;per\;nucleon} = \frac{E_b}{A} = \frac{1722.89 \; MeV}{235} = 7.33 \; MeV$$

5. *Comparison of nucleus stability*

The nucleus of uranium-235 is less stable than that of radium-226 as the binding energy per nucleon of Ra-226 (7.66 MeV) is larger than that of U-235 (7.33 MeV).

Radioactivity

Exercise 1:

1. *Balanced nuclear equation for the decay of* $^{222}_{86}Rn$

Nuclear reactions always obey two fundamental conservation laws:

Electric charge is conserved.

Total number of nucleons is conserved.

$$^{222}_{86}Rn \rightarrow {}^{A}_{B}X + \alpha$$

$$\Rightarrow {}^{222}_{86}Rn \rightarrow {}^{A}_{B}X + {}^{4}_{2}He \text{ ; So } \begin{cases} 222 = A + 4 \Rightarrow A = 218 \\ 86 = B + 2 \Rightarrow B = 84 \end{cases} \Rightarrow {}^{222}_{86}Rn \rightarrow {}^{218}_{84}X + {}^{4}_{2}He$$

Therefore : $^{218}_{84}X \equiv {}^{218}_{84}Po \; (Polonium)$

2. *The radioactive decay constant of* $^{222}_{86}Rn$

$$T = t_{1/2} = \frac{ln2}{\lambda} \Rightarrow \lambda = \frac{ln2}{t_{1/2}} = \frac{ln2}{3.8 \; x \; 24 \; x \; 3600} \approx 2.11 \; x \; 10^{-6} \; s^{-1} \text{ (Bq , dps)}$$

3. *Radioactive nuclei present in the initial sample*

$$N_0 = \frac{m \, . \, N_A}{M} = \frac{1g \; x \; 6.02 \; x \; 10^{23}}{222 \, \frac{g}{mol}} = 2.71 \; x \; 10^{21} \; nuclei$$

4. *Activity of the sample, Activity after 15 days*

Activity of a radioactive sample is the number of disintegrations per second (or the number of unstable atomic nuclei that decay per second in a given sample).

The following radioactive decay law is given by:

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

Where N_0 is the initial number of nuclei and N is the number of radioactive nuclei remaining after time t. λ *is* the radioactive decay constant for the particular nucleus.

Defining the **initial activity A_0** as $\boldsymbol{A_0 = \lambda \,.\, N_0}$, we have

$A = A_0 \,.\, e^{-\lambda.t}$ where A is the activity decreasing exponentially with time.

$\Rightarrow A_0 = \lambda \,.\, N_0 = 2.11 \; x \; 10^{-6}\, S^{-1} \, x \, 2.71 \; x \, 10^{21} \, nuclei = 5.72 \, x \, 10^{15} \, dps$

$= 5.72 \,.\, 10^{15} \, Bq$

-Activity after 15 days

$We \; have : A = A_0 \,.\, e^{-\lambda.t} = 5.72 \, x \, 10^{15} \, Bq \,.\, e^{-(2.11 \,.\, 10^{-6} .24 .3600. \; 15)} = 3.71 \, x \, 10^{14} \, Bq$

Exercise 2 :

1. *Balanced equation for the reaction. Type of nuclear reaction*

$^{6}_{3}Li + D \rightarrow 2 \,\alpha \quad \Rightarrow \quad ^{6}_{3}Li + \, ^{2}_{1}H \; \rightarrow 2 \; ^{4}_{2}He$

Type of nuclear reaction: Fusion

2. *The energy released*

Mass defect Δm is the difference between the actual atomic mass and the predicted mass calculated by adding the mass of protons and neutrons present in the nucleus.

We use the energy release formula converting mass into energy : $\Delta E = |\Delta m| \text{ x } c^2$

$where \;\; \Delta m = \sum m_{products} - \sum m_{reactants} = 2 \text{ x } 4.0015 - (6.0135 + 2.0135)$

$= -0.024 \;\; \text{amu}$

Therefore : $\Delta E = |\Delta m| \text{ x } c^2 = 0.024 \text{ x } 1.66 \text{ x } 10^{-27} \text{ kg x } (3 \text{ x } 10^8)^2 = 3.58 \text{ x } 10^{-12} \text{ J}$

Exercise 3 :

1. *Balanced nuclear equation for the decay of* $^{237}_{90}Th$

$^{237}_{90}\text{Th} \rightarrow \, ^{A}_{B}\text{X} + \alpha \quad \Rightarrow \quad ^{237}_{90}\text{Th} \rightarrow \, ^{A}_{B}\text{X} + \, ^{4}_{2}\text{He} \quad \Rightarrow \, ^{237}_{90}\text{Th} \rightarrow \, ^{233}_{88}\text{X} + \, ^{4}_{2}\text{He}$

$where : \, ^{233}_{88}X \equiv \, ^{233}_{88}Ra$

2. *Radioactive constant λ*

$\lambda = \frac{\ln 2}{T} = \frac{\ln 2}{18} = 0.0385 \, d^{-1}$ (d = day)

3. *Mass of thorium remaining*

$at\ t_1 = 36$ days: $m = m_0 \times e^{-\lambda t} = 1\mu g \times e^{-36\ d \times 0.0385\ d^{-1}} = 0.25\ \mu g$

at $t_2 = 6$ months $= 180$ days : $m = m_0 \times e^{-\lambda t} = 1\mu g \times e^{-180\ d\ \times\ 0.0385\ d^{-1}} =$
$9.8 \times 10^{-4}\ \mu g$

4. *Time t for thorium to decay to a mass = 0.0156 μg*

m = 0.0156 μg ; $m_0 = 1\ \mu g$

$$m = m_0 . e^{-\lambda t} \Rightarrow \frac{m}{m_0} = .e^{-\lambda t} \Rightarrow t = \frac{-\ln\frac{m}{m_0}}{\lambda} = \frac{-\ln\frac{0.0156}{1}}{0.0385} = 108\ d$$

The Bohr Model of the Atom

Exercise 1:

When a light beam of wavelength 300 nm strikes a metal surface, electrons are emitted from the metal. The kinetic energy of the electrons emitted is 0.74 eV.

1. Maximum speed of the electrons emitted

$$We\ have: E_C = \frac{1}{2} m_e V^2 \Rightarrow V = \sqrt{\frac{2 E_C}{m_e}} = \sqrt{\frac{2 \times 0.74 \times 1.6 \times 10^{-19} J}{9.1 \times 10^{-31} kg}} = 5.1 \times 10^5 \frac{m}{s}$$

2. Extraction energy in joules and eV

$We\ have: E = E_0 + E_C$

$where: E\ is\ the\ energy\ of\ the\ incident\ photon,$

$E_0\ is\ the\ extraction\ energy,$

$E_C\ is\ the\ kinetic\ energy.$

$$Therefore: E = E_0 + E_C \Rightarrow E_0 = E - E_C = h.\nu - E_C = h.\frac{c}{\lambda} - E_C$$

$$= 6.62 \times 10^{-34}\ J.s^{-1}.\frac{3 \times 10^8 \frac{m}{s}}{300 \times 10^{-9}\ m} - 0.74 \times 1.6 \times 10^{-19} J$$

$$= 6.62 \times 10^{-19} J - 1.18 \times 10^{-19} J = 5.44 \times 10^{-19}\ J = 3.4\ eV$$

3. Threshold frequency ν₀

$$as: E_0 = h\nu_0 \quad so\ \nu_0 = \frac{E_0}{h} = \frac{5.44 \times 10^{-19} j}{6.62 . \times 10^{-34}\ J.s^{-1}} = 8.22 . 10^{14}\ s^{-1} = 8.22 . 10^{14}\ Hz$$

4. Threshold wavelength λ₀ :

$$We\ have: \nu_0 = \frac{c}{\lambda_0} \Rightarrow \lambda_0 = \frac{c}{\nu_0} = \frac{3 . 10^8 \frac{m}{s}}{8.22 \times 10^{14}\ s^{-1}} = 364 \times 10^{-9}\ m = 364\ nm.$$

Exercise 2:

Bohr's theory is applied to an electron revolving around the nucleus of the hydrogen atom in the third circular orbit (n = 3).

1. Radius of Bohr's orbit in Å

$r_n = 0.53\ x\, n^2\,(\text{Å}) \Rightarrow r_3 = 0.53\,x\ 3^2 = 4.77\ \text{Å}$

2. Energy of the electron in eV

$as: E_n = \dfrac{-13.6}{n^2}\ (eV) \Rightarrow E_3 = \dfrac{-13.6}{3^2} = -1.51\ eV$

3. Ionizing wavelength $((n_f = \infty)$

a) <u>First method :</u>

$We\ have: \dfrac{1}{\lambda} = R_H\,.\left(\dfrac{1}{n_i^2} - \dfrac{1}{n_f^2}\right) = R_H\,.\left(\dfrac{1}{3^2} - \dfrac{1}{n_f^2}\right) \xrightarrow{\overline{IONIZATION}} \dfrac{1}{\lambda} = R_H\,.\left(\dfrac{1}{3^2} - \dfrac{1}{\infty^2}\right)$

$\Rightarrow \dfrac{1}{\lambda} = R_H\,.\left(\dfrac{1}{3^2}\right) \Rightarrow \lambda = \dfrac{9}{R_H} = \dfrac{9}{1.1\text{ x } 10^7\ m^{-1}} = 8.1818\,x\ 10^{-7}m = 818.18\,x\ 10^{-9}\ m$

$\Rightarrow \lambda = 818.18\ nm$

Corresponding energy

$\Delta E = h\,.\nu = h\,.\dfrac{c}{\lambda} = 6.62\text{ x } 10^{-34}\ \text{J}.s^{-1}.\dfrac{3\ x\ 10^8\ \frac{m}{s}}{818.18\,x\ 10^{-9}m} = 2.43\ x\ 10^{-19}J = 1.51\ eV$

b) <u>Second method :</u>

$E_i = E_\infty - E_n = E_\infty - E_3 = \dfrac{-13.6}{\infty^2} - \left(\dfrac{-13.6}{3^2}\right) = -\left(\dfrac{-13.6}{3^2}\right) = 1.51\ eV = 2.43\ x\ 10^{-19}\ J$

$We\ also\ have: \Delta E = E_i = h\,.\nu = h\,.\dfrac{c}{\lambda} \Rightarrow \lambda = \dfrac{h.c}{E_i} = \dfrac{6.62\text{ x } 10^{-34}\ \text{J}.s^{-1}\ x\ 3\ x\ 10^8\ \frac{m}{s}}{2.43\ x\ 10^{-19}\ J}$

$= 8.17\ x\ 10^{-7}\ m = 817\ nm$

Exercise 3 :

Consider the hydrogenoid Li^{2+}, it has the atomic number Z = 3.

1. Radius of the first orbit (n = 1)

For a hydrogenoid, we have:

$r_n = 0.53\,.\dfrac{n^2}{Z}\,(\text{Å}) \Rightarrow r_1 = 0.53\ x\ \dfrac{1^2}{3} = 0.177\ \text{Å}$

2. *Energy of the third ionization of the lithium atom:*

First method :

$$as \quad E_n = \frac{-13.6\ .\ Z^2}{n^2}\ (eV) \Rightarrow E_1 = \frac{-13.6\ .\ 3^2}{1^2} = -122.4\ eV\ , therefore, E_i = 122.4\ eV$$

Second method:

$$E_i = E_\infty - E_n = E_\infty - E_1 = \frac{-13.6\ .\ 3^2}{\infty^2} - \left(\frac{-13.6\ .\ 3^2}{1^2}\right) = -\left(\frac{-13.6\ .\ 3^2}{1^2}\right) = 122.4\ eV$$

3. *Wavelength and frequency of the corresponding line between energy levels n = 4 and n = 1*

$$we\ have: \begin{cases} \frac{1}{\lambda} = R_H\ .\ Z^2.\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \\ where: n_i = 1 \quad and \quad n_f = 4 \end{cases}$$

$$\Rightarrow \frac{1}{\lambda} = 1.1\ .\ 10^7\ m^{-1}\ x\ \ 3^2\ x\ \left(\frac{1}{1^2} - \frac{1}{4^2}\right) \Rightarrow \lambda = 1\ .10^{-8}\ m = 100\ x\ \ 10^{-10}\ m = 100\ \text{Å}$$

$$and: \ \nu = \frac{c}{\lambda} = \frac{3\ x\ \ 10^8\ \frac{m}{s}}{100\ x\ \ 10^{-10}\ m} = 3\ x\ 10^{16}\ S^{-1} = 3\ x\ \ 10^{16} Hz$$

4. Minimum energy value that the Li^{2+} ion must absorb to move from the ground state to the excited state

   That is $n_i = 1\ et\ n_f = 2$ donc :

Fisrt method :

$$using: \begin{cases} \frac{1}{\lambda} = R_H\ .\ Z^2.\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \\ n_i = 1 \quad and \quad n_f = 2 \end{cases}$$

$$\Rightarrow \frac{1}{\lambda} = 1.1\ \text{x}\ \ 10^7\ m^{-1}\ x\ \ 3^2\ x\ \left(\frac{1}{1^2} - \frac{1}{2^2}\right) \Rightarrow \lambda = 13\ x\ 10^{-9}\ m = 13\ nm$$

Corresponding energy :

$$\Delta E = h\,.\,\nu = h\,.\frac{C}{\lambda} = \ 6.62\,.\,10^{-34}\ \mathrm{J}.\,sS^{-1}.\frac{3\ x\ 10^{8}\ \frac{m}{s}}{13\,.\,x\ 10^{-9}m} = 1.53\ x\ 10^{-17}\,J = 92.16\ eV$$

$$\approx 92\ eV$$

<u>Second method :</u>

$$Using\ \Delta E = \ E_2 - E_1 = \frac{-13.6\ x\ \ 3^2}{2^2} - \left(\frac{-13.6\ \ x\ \ 3^2}{1^2}\right) = -30.6 - (-122.4) = 91.8\ eV$$

$$\approx 92\ eV$$

6. ***Frequency for:***

a) The limiting line in the Lyman series :

$$Using:\ \begin{cases} \frac{1}{\lambda} = R_H\ .\ Z^2.\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \\ where:\ \ for\ the\ Lyman\ \text{series}\ n_i = 1\ and\ the\ limiting\ line:\ n_f = \infty \end{cases}$$

$$\Rightarrow \frac{1}{\lambda} = 1.1\ \mathrm{x}\ \ 10^7\ m^{-1}\ x\ \ 3^2 x\ \left(\frac{1}{1^2} - \frac{1}{\infty^2}\right) \Rightarrow \lambda = 1\ x\ \ 10^{-8}\ m = 10\ x\ \ 10^{-9}\ m = 10\ nm$$

$$and:\ \nu = \frac{c}{\lambda} = \frac{3\ x\ \ 10^{8}\ \frac{m}{s}}{10^{-8}\ m} = 3\ x\ \ 10^{16}\ S^{-1} = 3\ x\ \ 10^{16} Hz$$

b) The second Balmer line :

$$Now:\ \begin{cases} \frac{1}{\lambda} = R_H\ .\ Z^2.\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \\ where:\ \ n_i = 2\ in\ the\ case\ of\ the\ Balmer\ series\ and\ the\ second\ line\ corresponds\ to\ \ n_f = 4 \end{cases}$$

$$\Rightarrow \frac{1}{\lambda} = 1.1\ \mathrm{x}\ \ 10^7\ m^{-1}\ \ x\ 3^2\ x\ \left(\frac{1}{2^2} - \frac{1}{4^2}\right) \Rightarrow \lambda = 53\ x\ 10^{-9}\ m = 53\ nm$$

$$et:\ \nu = \frac{c}{\lambda} = \frac{3\ x\ \ 10^{8}\ \frac{m}{s}}{53\ x\ \ 10^{-9}\ m} = 5.57\ x\ \ 10^{15}\ S^{-1} = 5.57\ x\ \ 10^{15} Hz$$

Periodic table

Exercise 1 :

1/ <u>Quantum numbers:</u>

Explains the behavior of the electron in an atom, i.e. its energy, its motions around the nucleus and the shape of the orbital, specifying its 4 quantum numbers (***n, l, m ,s***).

a- <u>The principal quantum number (the layer) (*n*):</u>

The value of n determines the energy of the orbital and the average distance between an electron in a given orbital and the nucleus.

$$\boldsymbol{On\ a : n \geq 1 \Rightarrow n = 1\ ou\ 2\ ou\ 3\ ou\ 4\ ou\ 5\ ou \ldots\ldots \infty}$$

b- <u>Secondary quantum number (sublayers) (*l*):</u>

The value of (***l***) indicates the shape (geometry) of the orbital.

$$\boldsymbol{We\ have : 0 \leq l \leq n - 1}$$

$if: l = 0\ \ c'est\ la\ sous\ couche\ S$

$if: l = 1\ \ c'est\ la\ sous\ couche\ P$

$if: l = 2\ \ c'est\ la\ sous\ couche\ d$

$if: l = 3\ \ c'est\ la\ sous\ couche\ f$

c- <u>The magnetic quantum number (***m***) :</u>

The value of (m) determines the orientation of the orbital.

$$\boldsymbol{We\ have : -l \leq m \leq +l}$$

d- <u>The magnetic spin quantum number (***s***) :</u>

The value (s) determines the orientation of the electron on itself. These values are ***s* = +*1/2*** et ***s* = -*1/2***.

For : **A/** n = **2**, ℓ= **0** et m = **0** :

$$if\ n = \mathbf{2} \Rightarrow l = \begin{cases} \mathbf{0} \Rightarrow m = \mathbf{0} \\ 1 \Rightarrow m = -1; 0; 1 \end{cases}\ it's\ right$$

Pour : **B/** n = **2**, ℓ= **1** et m = **1** :

$$if\ n = \mathbf{2} \Rightarrow l = \begin{cases} 0 \Rightarrow m = 0 \\ \mathbf{1} \Rightarrow m = -1; 0; \mathbf{1} \end{cases}\ it's\ right$$

For : **C/** n = **2**, ℓ= **2** et m = **0** :

$$if\ n = \mathbf{2} \Rightarrow l = \begin{cases} 0 \Rightarrow m = \mathbf{0} \\ 1 \Rightarrow m = -1; \mathbf{0}; 1 \end{cases}\ it's\ wrong\ because\ \ell = 0\ et\ 1; \ell \neq\ \mathbf{2}$$

For : **D/** n = **3**, ℓ= **2** et m = **-1**:

$$if\ n = \mathbf{3} \Rightarrow l = \begin{cases} 0 \Rightarrow m = 0 \\ 1 \Rightarrow m = -1; 0;\ +1 \\ \mathbf{2} \Rightarrow m = -2;\ -\mathbf{1}; 0;\ +1;\ +2 \end{cases} \qquad it's\ right$$

For : **E/** n = **1**, ℓ= **0** et m = **1** :

$$if\ n = \mathbf{1} \Rightarrow l = \mathbf{0} \Rightarrow m = 0 \qquad it's\ wrong\ because\ m = 0; m \neq \mathbf{1}$$

2/ The maximum number of electrons described by the (n ; ***l*** ; m ; s) :
Each box contains a maximum of $2e^-$:

if ***l***=0⇒S

m=0
↑↓

if ***l***=1⇒P

m=-1	m=0	m=1
↑↓	↑↓	↑↓

if ***l***=2⇒d

m=-2	m=-1	m=0	m=1	m=2
↑↓	↑↓	↑↓	↑↓	↑↓

if ***l***=3⇒f

m=-3	m=-2	m=-1	m=0	m=1	m=2	m=3
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

<u>For : **A/** n = **4** :</u>

$$if\ n = \mathbf{4} \Rightarrow l = \begin{cases} 0 \Rightarrow m = 0\ \square \\ 1 \Rightarrow m = -1; 0;\ +1\ \square\square\square \\ 2 \Rightarrow m = -2;\ -1; 0;\ +1;\ +2\ \square\square\square\square\square \\ 3 \Rightarrow m = -3; -2;\ -1; 0;\ +1;\ +2;\ +3\ \square\square\square\square\square\square\square \end{cases}$$

In the case of n=4 we have found 16 boxes, each box contains 2 e- so there is a maximum of 32e-.

<u>For : **B/** n = **3** et ℓ= **2** :</u>

$$if\ n = \mathbf{3} \Rightarrow l = \begin{cases} 0 \Rightarrow m = 0\ \square \\ 1 \Rightarrow m = -1; 0;\ +1\ \square\square\square \\ \mathbf{2} \Rightarrow m = -2;\ -1; 0;\ +1;\ +2\ \square\square\square\square\square \end{cases}$$

In the case of n=3 and l=2 we have found 5 boxes, each box contains 2 e- so there are 10e- at maximum.

<u>For : **C/** n = **2** et ℓ= **1** :</u>

$$if\ n = \mathbf{2} \Rightarrow l = \begin{cases} 0 \Rightarrow m = 0\ \square \\ \mathbf{1} \Rightarrow m = -1; 0; 1\ \square\square\square \end{cases}$$

In the case of n=2 and l=1 we have found 3 boxes, each box contains 2 e- so there are 6e- at maximum.

<u>For : **D/** n = 0, ℓ= 0 et m = 0 :</u>
This is false because n≠0

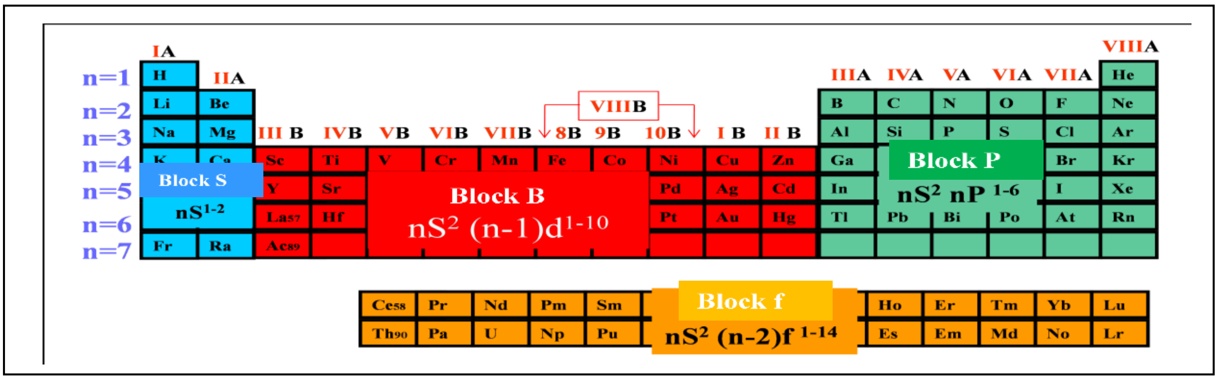
<u>For : **E/** n = **2**, ℓ= **1** et m = **-1** :</u>

$$if\ n = \mathbf{2} \Rightarrow l = \begin{cases} 0 \Rightarrow m = 0\ \square \\ \mathbf{1} \Rightarrow m = -\mathbf{1}; 0; 1\ \square\square\square \end{cases}$$

In the case of n=2, l=1 and m=-1 we have found 1 box that contains 2 e- so there are 2e- in maximum.

Exercise 2 :
<u>The position (period and group) of elements in the periodic table :</u>

The periodic table is composed of four blocks (S, B, P and F) :



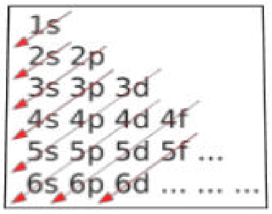
We have : n= Line = period=layer =1, 2, 3, 4, 5, 6 et 7.
Column = family= group= (figure + letter A or B)
= IA→VIIIA
Or = IB→VIIIB

<u>Locating elements on the periodic table :</u>

1- Make the electronic structure of the elements.

2- Find the valence layer and electrons (the outer layer).

To make the electronic structure of the elements, here's the order in which the sublayers are filled, one after the other, according to Klechkowski's rule (orbitals of increasing energy) following the red arrow :



We obtain the electronic configuration for any element X:

X $_{\text{number of e-}}$: **$1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ $4S^2$ $3d^{10}$ $4P^6$ $5S^2$ $4d^{10}$ $5P^6$ $6S^2$ $4f^{14}$ $5d^{10}$ $6P^6$ $7S^2$**…

$1e^-$

Na_{11} : $1S^2$ $2S^2$ $2P^6$ / $3S^1$ → IA (nS) So : Na_{11} ∈ (n=3 ∩ IA)

$2e^- + 1e^- = 3e^-$

Al_{13} : $1S^2$ $2S^2$ $2P^6$ / $3S^2$ $3P^1$ → IIIA (nS nP) So : Al_{13} ∈ (n=3 ∩ IIIA)

$2e^-$

Sr_{38} : $1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ $4S^2$ $3d^{10}$ $4P^6$ / $5S^2$ → IIA (nS) So : Sr_{38} ∈ (n=5 ∩ IIA)

$2e^- + 1e^- = 3e^-$

Sc_{21} : $1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ / $4S^2$ $3d^1$ → IIIB (nS (n-1)d) So : Sc_{21} ∈ (n=4 ∩ IIIB)

$2e^- + 9e^- = 11e^-$

Cu_{29} : $1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ / $4S^2$ $3d^9$ → 11B (nS (n-1)d) So : Cu_{29} ∈ (n=4 ∩ 11B)

Or *:*

$1e^-$

Cu_{29} : $1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ / $4S^1$ $3d^{10}$ → IB So : Cu_{29} ∈ (n=4 ∩ IB), because d^{10} is saturated.

nS (n-1)d

$2e^- + 10e^- = 12e^-$

Zn_{30} : $1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ / $4S^2$ $3d^{10}$ → 12B (nS (n-1)d) So : Zn_{30} ∈ (n=4 ∩ 12B)

Or *:*

$2e^-$

Zn_{30} : $1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ / $4S^2$ $3d^{10}$ → IIB (nS (n-1)d) So : Zn_{30} ∈ (n=4 ∩ IIB), because d^{10} is saturated.

$2e^- + 2e^- = 4e^-$ → IVA

Ge_{32} : $1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^6$ / $4S^2\ 3d^{10}\ 4P^2$ IVA So : $Ge_{32} \in (n=4 \cap IVA)$, because d^{10} is saturated.

$4S^2\ 4P^2$ (nS nP) → IVA

$2e^- + 5e^- = 7e^-$ → VIIA

I_{53} : $1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^6 4S^2\ 3d^{10}\ 4P^6$ / $5S^2\ 4d^{10}\ 5P^5$ VIIA So : $I_{53} \in (n=5 \cap VIIA)$,

$5S^2\ 5P^5$ → VIIA

Exercise 3 :

An element X belongs to the boron column (B_5) and the potassium period (K_{19}).

1/ The atomic number of X :

- The Boron (B_5) column:

$2e^- + 1e^- = 3e^-$ → IIIA

B_5 : $1S^2$ / $2S^2\ 2P^1$ IIIA So : $B_5 \in (n=2 \cap IIIA)$

(nS nP) → IIIA

- The potassium (K19) period:

$1e^-$ → IA

K_{19} : $1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^6$ /$4S^1$ IA So : $K_{19} \in (n=4 \cap IA)$

(nS) → IA

So: X ∈ (période du K_{19} ∩ colonne de B_5)

⇒ X ∈ (n=4 ∩ IIIA)

⇒ the electronic structure of X then ends with: $4S^2\ 4P^1$

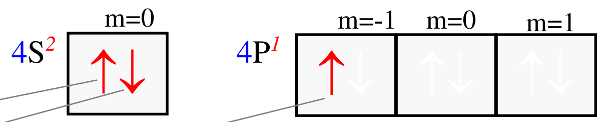
⇒X: ……………………….$4S^2\ 4P^1$ ⇒X: $1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^6$ / $4S^2\ 3d^{10}\ 4P^1$ ⇒ X_{31}

⇒ le numéro atomique de X est 31

2/ The quantum numbers for the valence electrons of X:

X_{31}: $1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^6$ / $4S^2\ 3d^{10}\ 4P^1$

⇒ The outer or valence layer is: $4S^2\ 4P^1$; it contains 3e- of valence:

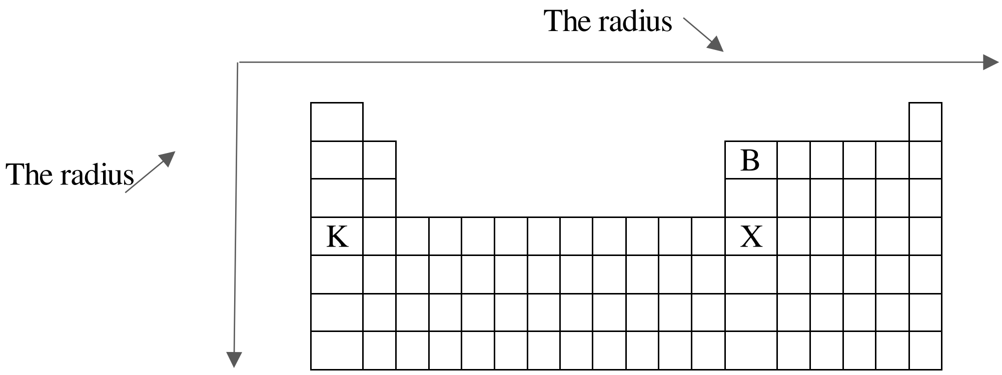


(4, 0, 0, +1/2)

(4, 0, 0, -1/2)

(4, 1, -1, +1/2)

3/ Classification of these elements (X_{31}, B_5 and K_{19}) in order of increasing radius :



According to the table, we have : radius de K_{19} > radius de X_{31}

And : radius de X_{31} > radius de B_5

So : radius de K_{19} > radius de X_{31} > radius de B_5

4/ The radius (r_n) and energy (E_n) of chlorine (Cl_{17}) according to Slater :

<u>SLATER rules (Zeff calculation) :</u>

- Slater has set out the rules for expressing these shielding effects (δ) between electrons.

We have :

$$Zeff = Z - \delta \;\; with\; \delta = \sum \delta_i$$

$$\delta = \sum \delta_i = [(n\; electron\; of\; the\; layer\; studied\; - 1\; electron\; studied)\,.\,coeff\; slater$$
$$+\; n\; electron\; of\; the\; previous\; layer\,.\,coeff\; slater$$
$$+\; n\; electron\; of\; the\; previous\; layer.\,coeff\; slater + \cdots]$$

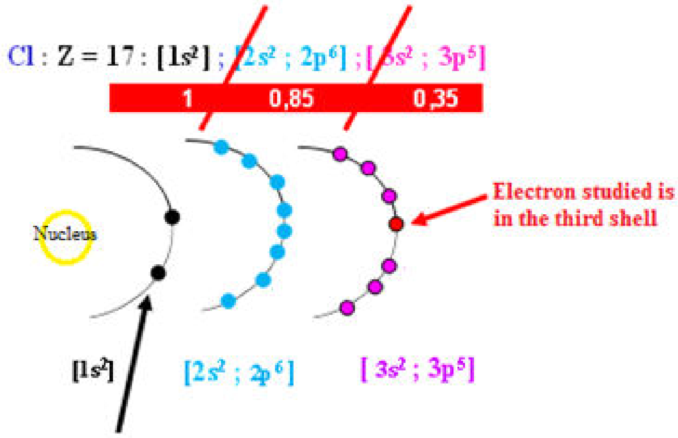
- The electronic configuration of the element must be written using the following groups in the following order :

X : [1s] ; [2s,2p] ; [3s,3p] [3d] ; [4s,4p] [4d] [4f] ;[5s,5p] [5d] ; [5f]...

- Slater screen coefficients (Slater triangle) are as follows :

1s	0,3				
2s2p	0,85	0,35			
3s3p	1	0,85	0,35		
3d	1	1	1	0,35	
4s4p	1	1	0,85	0,85	0,35

- Zeff calculation for chlorine (Cl):



Zeff = 17 - ((7-1) . 0.35) + (8 . 0,85) + (2 . 1)) = 6,1

$$-The\ radius\ (rn):\ r_n = 0.529\,.\frac{n^2}{Z_{eff}}\ (\text{Å}) \Rightarrow r_3 = 0.529\,.\frac{3^2}{6.1} = 0.78\ (\text{Å}).$$

$$-Energy\ (En): E_n = -13.6\,.\frac{Z_{eff}^2}{n^2}\ (eV) \Rightarrow E_3 = -13.6\,.\frac{(6.1)^2}{3^2} \approx\ 56.23\ (eV).$$

Chemical bonds

Exercise 1:

The electronic structure of the elements are:

H_1: $1S^1$ [↑]

He_2: $1S^2$ [↑↓]

Li_3: $1S^2\ 2S^1$ [↑]

Be_4: $1S^2\ 2S^2$ [↑↓]

B_5: $1S^2\ 2S^2$ [↑↓] $2P^1$ [↑][ ][ ]

C_6: $1S^2\ 2S^2$ [↑↓] $2P^2$ [↑][↑][ ]

N_7: $1S^2\ 2S^2$ [↑↓] $2P^3$ [↑][↑][↑]

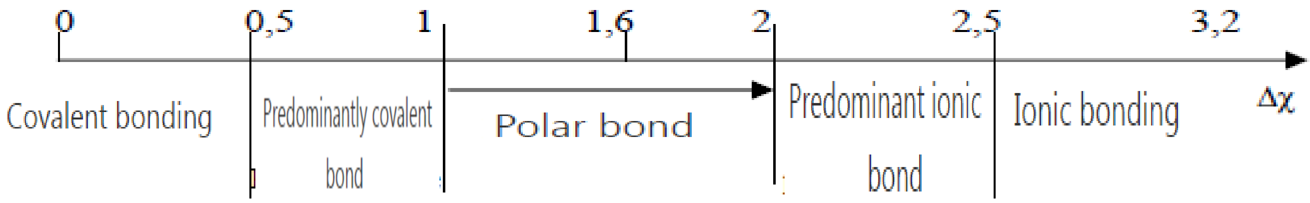
F_9: $1S^2$ $2S^2$ [↑↓] $2P^5$ [↑↓ | ↑↓ | ↑]

Ne_{10}: $1S^2$ $2S^2$ [↑↓] $2P^6$ [↑↓ | ↑↓ | ↑↓]

According to the Lewis model ; we have :

Exercise 2:

The difference in electronegativity $\Delta\chi$ between the most and least electronegative elements is given in the diagram below :



For KF, We have : $\Delta\chi = 3{,}2 \Rightarrow$ The bond is ionic (K^+, F^-).

For KCl, We have : $\Delta\chi = 2{,}3 \Rightarrow$ The bond is predominantly ionic.

For HF, We have : $\Delta\chi = 1{,}8 \Rightarrow$ The bond is polar. Atoms carry partial charges : $H^{\delta+} - F^{\delta-}$.

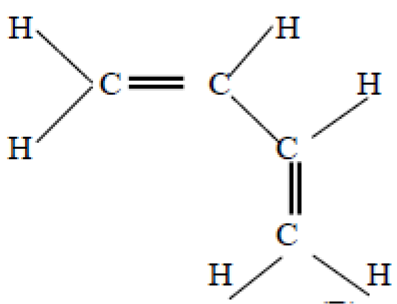
For HCl, We have : $\Delta\chi = 0{,}9 \Rightarrow$ The bond is polar but less marked.

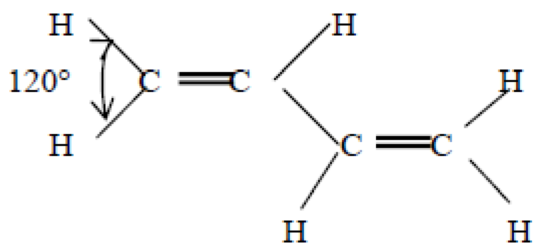
and For HH, We have : $\Delta\chi = 0 \Rightarrow$ The bond is purely covalent.

Exercise 3:

1. All carbons 1, 2, 3 and 4 in molecule : Butadiène $C^{(1)}H_2 = C^{(2)}H\text{-}C^{(3)}H = C^{(4)}H_2$ form three bonds s. They are therefore all hybridized to Sp^2.
2. The two possible geometric forms are:

Form A:





Référence :

1-J. Chem. Educ. 1926, 3, 10, 1110 Publication Date:October 1, 1926 ; https://doi.org/10.1021/ed003p1110

2- Brown, Theodore L.; LeMay, H. Eugene Jr.; Bursten, Bruce E.; Murphey, Catherine J.; Woodward, Patrick M.; Stoltzfus, Matthew W.; Lufaso, Michael W. (2018). "Introduction: Matter, energy, and measurement". Chemistry: The Central Science (14th ed.). New York: Pearson. pp. 46–85. ISBN 978-0134414232.

3-"Alchemy", entry in *The Oxford English Dictionary*, J.A. Simpson and E.S.C. Weiner, vol. 1, 2nd ed., 1989, ISBN 0-19-861213-3.

4- Hill, J.W.; Petrucci, R.H.; McCreary, T.W.; Perry, S.S. (2005). General Chemistry (4th ed.). Upper Saddle River, New Jersey: Pearson Prentice Hall. p. 37.

5- Glaser, Christopher (1663). Traite de la chymie. Paris. as found in: Kim, Mi Gyung (2003). Affinity, That Elusive Dream – A Genealogy of the Chemical Revolution. The MIT Press. ISBN 978-0-262-11273-4.

6- Chang, Raymond (1998). Chemistry, 6th Ed. New York: McGraw Hill. ISBN 978-0-07-115221-1.

7- Winter, Mark. "WebElements: the periodic table on the web". The University of Sheffield. Archived from the original on 4 January 2014. Retrieved 27 January 2014.

8- Ihde, Aaron John (1984). *The Development of Modern Chemistry*. Courier Dover Publications. p. 164. ISBN 978-0-486-64235-2.

9 -S. Meziane, Chimie générale - Structure de la matière, 4ème BERTI édition (2021).

10-R. Ouahès, B. Devallez, Chimie Generale. 4ème édition corrigée, édition L'Esprit frappeur (1997).

11-M. Guymont, Structure de la matière : Atomes, liaisons chimiques et cristallographie éditeur : Belin(2003).

12-G. Germain, R. Mari, D. Burnel, Chimie générale, édition Elsevier / Masson, (2007).