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Faculty of Medecine

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وزارة التعليم العالي و البحث العلمي
جامعة الجيلالي اليابس بسيدي بلعباس
كلية الطب

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GENERAL CHEMISTRY COURSE

First year LMD: ST, SM
Medicine and Pharmacy

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Foreword

This modest work is intended for LMD students of (science of technology, science of materials, public works, electrical engineering and biology) and for first-year Medicine, Pharmacy and Dentistry students.

This course material includes nine chapters primarily aimed at educating students on the most important concepts in chemistry.

Its purpose is to bring together the fundamental notions related to chemistry and enabling a better understanding of general chemistry because it brings together the essential information in structural chemistry, thermodynamics and solution chemistry: acids and bases and chemical equilibria. The reader wishing to delve deeper into the concepts covered can refer to the works cited in the references.

This work is the result of several years of teaching and represents a synthesis of experience gained through my pedagogical task as a General Chemistry instructor at the Faculty of Sciences, the Faculty of Engineering in the Department of Electrical Engineering, and the Faculty of Medicine in the Departments of Pharmacy and Medicine at Djilali Liabès University of Sidi Bel Abbès.

Finally, I would welcome with pleasure, any comments and suggestions from users, colleagues, teachers and students, as they demonstrate interest in this educational aid.

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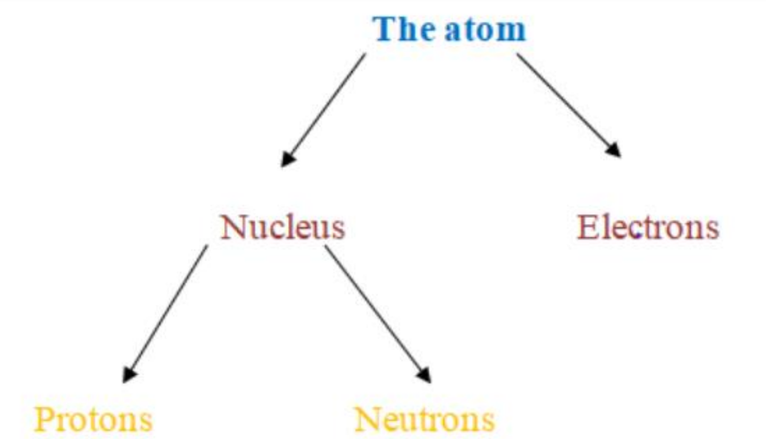
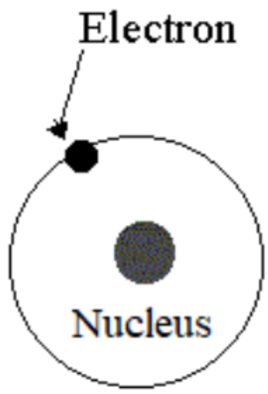
I- Identification of the constituents of the matter:

The matter, whatever its physical state, is made up of fine invisible and indivisible particles called atoms.

The atom, in turn, is made up of a nucleus surrounded by an electronic cloud. The nucleus encompasses a number of protons and neutrons.





The mass: The mass of the atom represents the total mass of these constituents.

- Electron : m_{e^-}= 9,108 x 10^{-31} kg.
- Neutron: m_n= 1, 6749 x 10^{-27}kg.
- Proton : m_p = 1,672 x 10^{-27} kg.

m_n=1839 x m_e m_p= 1836x m_e

The mass of the electrons is negligible compared with that of the protons and neutrons. The mass of the atom is therefore concentrated in the nucleus.

The charge: The charge of the atom is the sum of the charges of these constituents.

- Electron: $|e^-|$ = -1,602x 10^{-19} c.
- Neutron: $|n|$ = 0.
- Proton: $|e^+|$ = 1,602 x 10^{-19} c.

⇨ The atom contains the same number of proton and electrons, so the atom is electrically neutral.

⇨ The distribution of electrons around an atom's nucleus (electronic configuration) determines the chemical properties of the element. On one hand, electrons are the main agents responsible for chemical interactions.

The number of valence electrons and their arrangement determine how an element will bond with other elements. Metals tend to lose electrons and become cations, while non-metals tend to gain electrons and become anions.

Electrons, through their configuration and dynamics, largely determine how elements interact chemically, thus forming the basis of the chemical properties of elements.

The radius:

The radius of the nucleus $r_N = 10^{-14}$ m

The radius of the atom $r_A = 10^{-10}$ m

⇒ $\frac{rN}{rA} = \frac{10^{-14}}{10^{-10}} = \frac{1}{10^4} = 10^{-4}$ $r_A = 10^4 \times r_N$

The atomic radius is a measure of the size of an atom, defined as the average distance between the nucleus of an atom and the outer boundary of its outermost electron shell. The atomic radius is related to atomic stability:

A smaller atomic radius means that the valence electrons (the electrons in the outer shell) are closer to the nucleus. This results in stronger electrostatic forces between the electrons and the nucleus, higher binding energies, which means more energy is required to remove an electron from the atom. This contributes to the overall stability of the atom.

If the atomic radius is larger, the valence electrons are farther from the nucleus and are less strongly attracted to it, they tend to lose their valence electrons more easily, thus increasing the reactivity of the atom. Conversely, atoms with smaller atomic radii have a strong tendency to gain electrons to achieve a stable electron configuration, also increasing their reactivity.

In summary, the atomic radius directly influences the stability and reactivity of an atom. A smaller atomic radius tends to increase stability and decrease reactivity, while a larger atomic radius tends to increase reactivity and decrease stability.

I-1-Characteristics of the atom:

The atom is symbolised by the first or first and second letters of its Latin name $^{A}_{Z}X$.

- A: the mass number, is the atomic mass.
- Z: the charge number or proton number.
- N: the neutron number. N = A – Z.

Example :

$^{1}_{1}H$ A = 1 Z = 1 N=0

$^{27}_{13}Al$ A= 27 Z = 13 N=14

I-2- Isotope concept :

Isotopes are atoms of the same element with the same atomic number Z but a different mass number A.

Example:

$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$	$^{16}_{8}O$	$^{17}_{8}O$	$^{18}_{8}O$
Z = 1	Z = 1	Z = 1	Z = 8	Z = 8	Z = 8
A= 1	A = 2	A = 3	A = 16	A = 17	A = 18
N= 0	N = 1	N = 2	N= 8	N = 9	N = 10

Isotopes are therefore atoms of the same element that differ in the number of neutrons. These isotopes exist in different proportions, known as "isotopic abundance".

The natural average atomic mass is calculated using the following expression:

$$\boldsymbol{Mavg = \mathrm{Mnat} = \frac{\sum(\mathrm{Mi} \times \mathrm{Xi})}{100}}$$

With Xi is the percentage of each constituent and $\sum Xi = 100\%$.

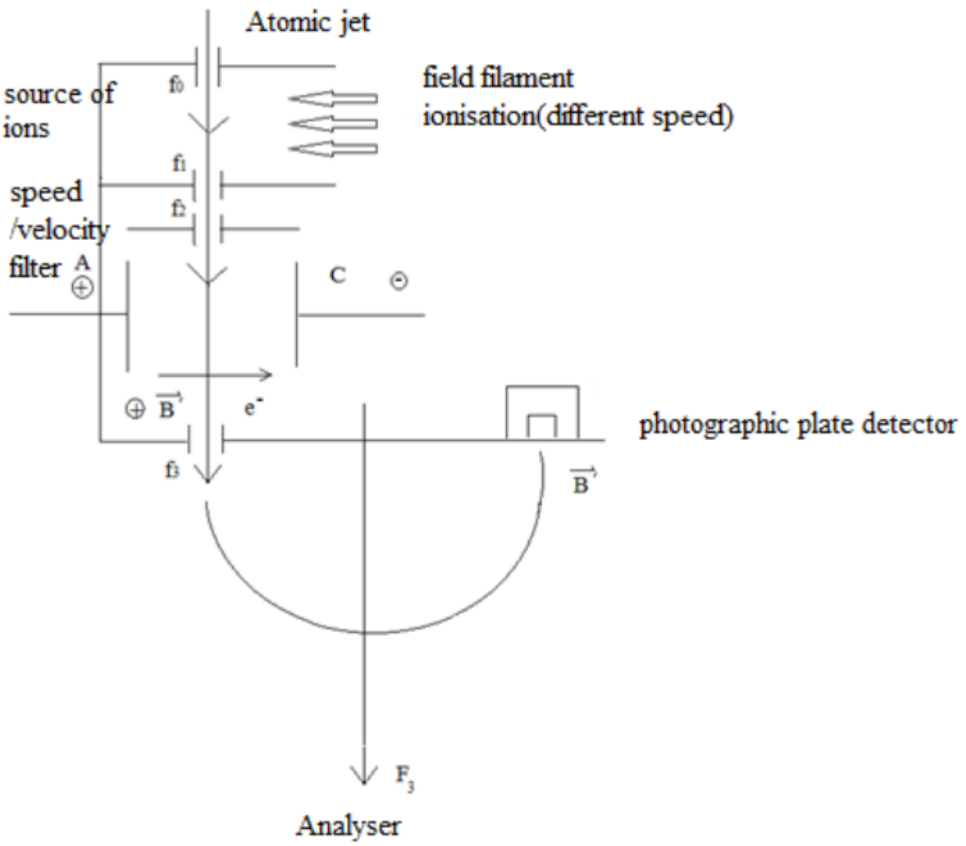
Example :

$^{16}_{8}O$	$^{17}_{8}O$	$^{18}_{8}O$
99,76%	0,037%	0,205%

$$Mnat = \frac{16 \times 99.76 + 17 \times 0.037 + 18 \times 0.205}{100}$$

I-3- Isotope separation:

To measure the mass of an atom in a mixture of isotopes, the most commonly practiced method is to measure the ratio q/m, or mass charge, using a device called a "BAINBRIDGE mass spectrometer", where q is the charge and m is the mass.



The principle consists of ionising a jet of atoms in the ionisation chamber, which are subjected to an electric fiel d$\vec{E}$ where they are accelerated, then deflected along circular trajectories by a magnetic field $\vec{B}$ perpendicular to the electric field.

Electric force

Magnetic field (Force) $\overrightarrow{F_2} = q . \overrightarrow{B} . V$

The passage through the f_3 slit results the equality of the two forces $|\overrightarrow{F_2}| = |\overrightarrow{F_1}|$

$$q . E = q .B.V \quad \text{donc : } E = B . V$$

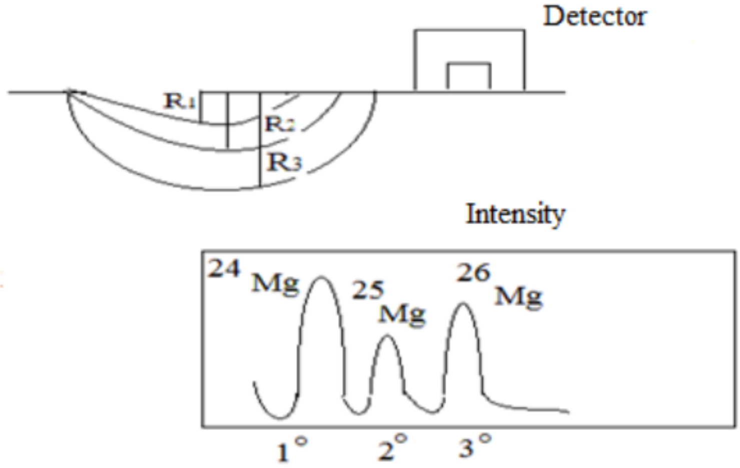
V: speed(speed=velocity) of ionised atoms m/s. $V = \frac{E}{B}$

The ions are deflected in a semicircle form at constant velocity in a uniform circular motion, leading to a centrifugal force F3 which is balanced by the magnetic force.

$$F_3 = m\frac{V^2}{R} = q \cdot V \cdot B_0 \Rightarrow \frac{q}{m} = \frac{V}{R \cdot B_0}$$ with R: radius (m)

Depending on the number of isotopes existing, we have the equality:

$$\frac{q}{m_1} = \frac{q}{m_2} = \frac{q}{m_3} = \frac{q}{mi}$$



Example: Magnesium Mg

^{25}Mg: 10,11% ^{24}Mg: 78,6% ^{26}Mg: 11,29%

I-4- Stability of an atomic nucleus

The stability of an atomic nucleus is governed by two fundamental forces: the strong nuclear force and the weak nuclear force, which play important roles in the behavior of atomic nuclei.

I-4-1- Strong Nuclear Force

This is the nuclear interaction force, the most powerful force in the universe, but it operates over very short distances (approximately 10^{-15} meters, roughly the size of an atomic nucleus). It is the attractive force that acts between nucleons (protons and neutrons) in the nucleus, holding them together despite the repulsive force between positively charged protons. The energy associated with the strong nuclear force is known as nuclear binding energy. It is the energy required to break apart a nucleus into its individual protons and neutrons, and it explains the mass defect observed in nuclei.

I-4-2- Weak Nuclear Force

The weak nuclear force, or weak interaction, is much weaker than the strong nuclear force. The weak force operates over a very short distance (less than 10^{-18} meters). It is responsible for processes such as beta decay, which are crucial for nuclear reactions and stability. The weak force is also responsible for changes that can transform a proton into a neutron or vice versa, which is involved in radioactive decay processes. For example, in beta decay, a neutron decays into a proton, an electron, and an antineutrino through the weak interaction. This process is vital for the synthesis of elements in stars and the regulation of energy production in nuclear reactors.

The balance between the strong nuclear force and the electromagnetic force determines the stability of a nucleus. While the strong force binds protons and neutrons together, the electromagnetic force causes repulsion between protons. If the repulsive force outweighs the attractive strong force, the nucleus becomes unstable and may undergo decay.

To illustrate the concepts of the strong nuclear force and the weak nuclear force, here are some concrete examples:

⇨ **Example of the Strong Nuclear Force: Nuclear Fusion in Stars**

Nuclear fusion is the process by which two light atomic nuclei fuse to form a heavier nucleus. This process releases a huge amount of energy, and it is how stars, including our Sun, produce their energy.

- In the core of the Sun, temperatures and pressures are sufficiently high. Protons (hydrogen nuclei) come close enough for the strong nuclear force to bind them together, forming helium nuclei. The energy released by this process is what powers the Sun and sustains life on Earth.

⇒ Example of the Weak Nuclear Force: Beta Decay

Beta decay is a radioactive decay process in which a neutron in an unstable nucleus can decay into a proton, an electron, and an antineutrino. This process occurs through the weak nuclear force. For example, in the beta decay of carbon-14, a neutron transforms into a proton, and carbon-14 becomes nitrogen-14.

This process helps maintain a balance between the number of protons and neutrons in the nucleus, and it plays an essential role in natural processes such as radioactivity and the synthesis of elements in stars.

⇒ Example of the Balance between the Strong Nuclear Force and the Electromagnetic Force: Uranium-235 and Nuclear Fission

Uranium-235 is a fissile isotope used in nuclear reactors and atomic bombs. Nuclear fission is the process by which a heavy nucleus splits into two lighter nuclei, releasing energy.

The strong and weak nuclear forces play fundamental roles in various atomic and nuclear processes, influencing the stability of nuclei, the production of energy in stars, and the practical applications of nuclear fission.

I-5-Application exercise

Exercise N°01:

1. Indicate the number of protons, neutrons, and electrons that participate in the composition of the following structures and identify those that present isotopes:

$^{18}_{8}O$, $^{18}_{8}O^{-2}$, $^{35}_{17}Cl^{-}$, $^{56}_{26}Fe^{+2}$, $^{40}_{20}Ca$, $^{40}_{20}Ca^{+2}$, $^{32}_{16}S^{-2}$, $^{27}_{13}Al^{+3}$, $^{59}_{28}Ni$, $^{52}_{27}Co$

1) Complette the folowing table :

Elément		H		
Core	$^{1}_{1}H$			
Z				
A				40
Number of neutrons		1	6	21

Exercise N°02:

1. Natural copper is composed of two stable isotopes with respective atomic masses of 62.929 and 64.927. The atomic number of copper is Z = 29. Indicate the composition of the two isotopes:
   - One isotope has 29 neutrons and 34 protons, the other isotope has 29 protons and 34 neutrons.
   - One isotope has 30 protons and 33 neutrons, the other isotope has 29 protons and 36 neutrons.
   - One isotope has 29 protons and 34 neutrons, the other isotope has 29 protons and 36 neutrons.
   - One isotope has 29 protons and 35 neutrons, the other isotope has 29 protons and 37 neutrons.
2. Potassium exists in the form of two stable isotopes, 39K and 41K. The relative abundances of these two isotopes are 93.09% for one and 6.91% for the other. Given that the molar mass of natural potassium is 39.10 g/mol, assign the natural abundance to each isotope.
3. Magnesium (Z = 12) has three stable isotopes: 24Mg, 25Mg, and 26Mg. Their natural abundances are 78.6%, 10.1%, and 11.3% respectively. Calculate the approximate atomic molar mass of magnesium and explain why the obtained result is only approximate.

Solution

Solution N°01:

$^{A}_{Z}X$ $\quad$ $^{A'}_{Z'}X$ ion Charge
ion

Atome

A ou A': mass number = Proton number + neutron number.

Z ou Z': atomic number.

Z: Proton number = electron number

Z': Proton number.

electron number e^- = Z' – ion charge.

1)
a)

	A	Z (protons)	A-Z (Neutrons)	Z- ion charge (Electrons)
$^{18}_{8}O$	18	8	10	8
$^{18}_{8}O^{-2}$	18	8	10	8
$^{35}_{17}Cl^{-}$	35	17	18	18
$^{56}_{26}Fe^{+2}$	56	26	30	24
$^{40}_{20}Ca$	40	20	20	20
$^{40}_{20}Ca^{+2}$	40	20	20	18
$^{32}_{16}S^{-2}$,	32	16	16	18
$^{27}_{13}Al^{+3}$	27	13	14	10
$^{59}_{28}Ni$	59	28	31	28
$^{52}_{27}Co$	52	27	25	27

b) The structures that present isotopes are: $^{18}_{8}O$ et $^{52}_{27}Co$

O: ^{13}O, ^{14}O, ^{15}O, ……..^{18}O Cl: ^{31}Cl, …………….. ^{43}Cl

Fe: ^{49}Fe, …………….^{62}Fe Ca: ^{37}Ca, ……………^{50}Ca

S: ^{29}S, ………………..…^{40}S Al: ^{23}Al, …………... ^{34}Al

Ni: ^{53}Ni, ……………..^{68}Ni Co: ^{53}Co, …………….^{64}Co

2)

Elément	H	H	C	K
Core	$^{1}_{1}H$	$^{2}_{1}H$	$^{12}_{6}C$	$^{40}_{19}K$
Z	1	1	6	19
A	1	2	12	40
NeutronNumber	0	1	6	21

Solution N°02:

1)

$^{63}_{29}Cu$

Z = proton number P =electron nombre e^- = 29

Number N = A-Z = 34

29 e_s^-, 29p, 34n

$^{65}_{29}Cu$

Z = proton number = e^- nombre = 29

Number N = A-Z = 36

29 e_s^-, 29p, 36n

C This is the correct composition.

2)

$_{x\%}{}^{39}K$ $\quad$ $^{41}K_{\ y\%}$

$\underbrace{\qquad\qquad\qquad}$

$^{39,1}K_{\ 100\%}$

93,09%

06,91% $\qquad$ x, y ?

The first Méthod: $M = A = \sum x_i A_i$

1st case : x = 93,09%, y = 06,91%

$M = A = \sum x_i A_i = (0{,}9309 .\ 39) + (0{,}0691 .\ 41) = 39{,}13$ g/mole.

2nd case: x = 06,91%, y = 93,09%

$M = A = \sum x_i A_i = (0{,}0691 .\ 39) + (0{,}9309 .\ 41) = 41$ g/mole.

Thus, hypothesis No. 01 is correct. x = 93,09%, y = 06,91%

The second Méthod:

One could also have solved the system:

$M = A = \sum x_i A_i \Rightarrow 39{,}1 \approx \begin{cases} \alpha 41 + \beta 39 \\ \alpha + \beta = 1 \end{cases}$

either: $\alpha = 0{,}05$, $\beta = 0{,}95$

so 95% of ^{39}K and 5% de ^{41}K.

These approximate results correspond to the experimental values.

3) We can also notice that one isotope is overwhelmingly predominant while the other is practically negligible. The average molar mass of 39.1 g/mol is therefore very close to that of the most abundant isotope. Here, it is the isotope 39K which represents 93%.
4) ^{24}Mg, ^{25}Mg, ^{26}Mg

$M = \sum x_i A_i$

$M_{Mg} = (0{,}101\ 25) + (0{,}113\ 24) + (0{,}786\ 26) \approx 24{,}3$ g/mole

Approximations:

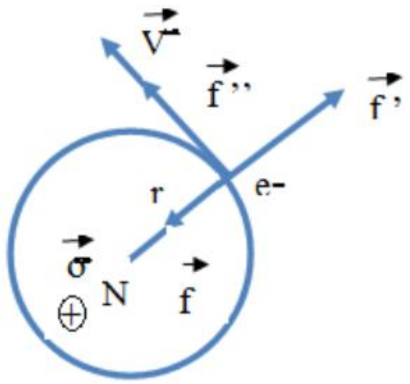
1- mp ≈ mn ≈ 1 u.m.a
2- mass of electrons - negligible.

Mass defect - negligible.

I- BOHR's theory applied to the hydrogen atom:

The Bohr theory, developed by Niels Bohr in 1913, is a crucial step in understanding atomic structure and electron behavior.

BOHR's model assumes that the electron of the hydrogen atom revolves around the nucleus and describes circular orbits around it.



Centrifugal force $\vec{f}'$

Attractive force $\vec{f}$

(Force of attraction)

Tangential force $\vec{f}''$

The electron of mass m_e orbits the nucleus with a constant velocity v

⟹ $\vec{p} = m_e.\vec{v}$ its angular momentum is $\sigma = m_e.\, v.\, r$

p : Momentum (kg.m/s).

v : The speed/velocity of the electron (m/s).

σ: the angular momentum ($kg.m^2/s$).

m_e :The mass of the electron (kg).

The angular momentum σ is perpendicular to the circumference.

I-1- BOHR's postulates :

- **1st postulate** :

  The angular momentum m.v.r can only take discrete values equal to integer multiples of $\frac{h}{2\pi}$.

$\sigma = m_e.v.r = n\frac{h}{2\pi} = n\,\hbar$ n = 1, 2, 3…… natural number (integer)

According to the 1st postulate, only certain orbits are stable, therefore only certain values of radius r are allowed.

We can say that angular momentum is quantized.

- **2nd postulate :**

When the electron moves into one of these permitted orbits, the atom does not radiate; these orbits are called stationary orbits.

- **3rd postulate** :

An electron only emits or absorbs light radiation when it jumps from one stationary orbit to another.

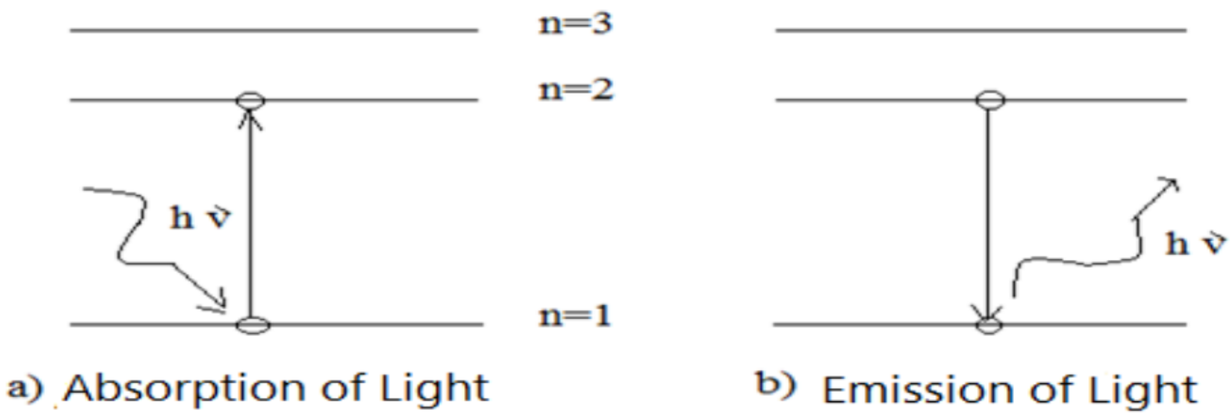


Figure 1 : Absorption ad Emission of light

With: level n=1 the ground state level n=2,3, the excited state

Imagining that a photon of light energy encounters a hydrogen atom in its ground state. The electron will passe from the initial level n=1 to the higher-level n=2, and the atom has absorbed the incident radiation of frequency γ.

The excited state in which the electron finds itself is unstable, the electron will return to the level n = 1, so it emits the frequency γ.

The change in energy during an electronic transition is written as:

$E_2 - E_1 = h\upsilon$ h: PLANCK's constant = $6{,}62\times10^{-34}$j.s

υ: frequency in s^{-1}.

I-2- BOHR's calculation:

I-2-1 Radius calculation:

For the electron to be in the circular orbit, the centrifugal force f' must be equal to the coulombic attraction force f.

f': centrifugal force = $m_e \frac{v^2}{r}$ (N)

f : coulombic attraction force = $\frac{ke^2}{r^2}$ (N)

v: tangential speed /velocity (m/s).

m_e: mass of the electron.

r: proton-electron distance (radius).

K : Coulomb attraction constant (k= 9×10^9SI)

According to Bohr's first postulate :

$$m_{e^-}.v.r = n\,\frac{h}{2\pi} \;......(1)$$

where n is an integer designating the number of the stationary orbit"

According to Bohr's second postulate :

f'= f c.à.d. $m_{e^-}\frac{v^2}{r} = \frac{ke^2}{r^2}$ so $m_{e^-}v^2 = \frac{ke^2}{r}$.......(2)

From (1) we derive the speed and replace it in (2) to obtain:

$m_{e-}\frac{n^2h^2}{4\pi^2 m_{e-} r^2} = \frac{ke^2}{r}$ after simplification, we find :

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

Hence the final expression for the radius ⟹ $r = \frac{n^2h^2}{4\pi^2 m_{e-} ke^2}$

We define $\frac{h^2}{4\pi^2 m_{e-} ke^2}$= A=cste = 0,53°A (angström) 1 A° = 10^{-10}m

It can be written as :

$r = A.n^2 = 0,529.\ n^2 = 0,53.n^2 A° = 0,53.10^{-10} n^2$ m

With m_{e-}: mass of the electron = $9,1x10^{-31}$kg

e : charge of the electron = $1,6x10^{-19}$j

h : Planck's constant = $6,62x10^{-34}$js

π = 3,14

I-2-2 Energy calculation :

The total energy E of the e-in the atom is equal to the sum of the kinetic energy Ek and the potential energy Ep.

$E = E_k+E_p$ Ek : kinetic energy, Ep : potential energy

With $E_k = \frac{1}{2} m_{e-} v^2$ and E_p= -eV

V is the electrostatic potential undergone by the electron towards to the proton.

If f is the coulomb force of attraction.

$dE_p = f\ dr = \frac{ke^2}{r^2}$ dr the potential energy is proportional to the radius if r ↗Ep ↗

We integrate $\int dEp = \int f dr = \int \frac{ke^2}{r^2} dr$

⟹ $E_p = -\frac{ke^2}{r} + cst$ by convention E_p=0 when r→ ∞ so the cst =0.

Then the total energy becomes E= ½ $mv^2 - \frac{ke^2}{r}$

As $m_{e-}\frac{\upsilon^2}{r} = \frac{ke^2}{r^2}$ so $m_{e-}\upsilon^2 = \frac{ke^2}{r}$

$E = \frac{1}{2}\frac{k\,e^2}{r} - \frac{ke^2}{r} = -\frac{ke^2}{2r} = -\frac{me\!-\!v^2}{2}$

According to Bohr's first postulate $m_{e-}vr = n\frac{h}{2\pi}$, the speed/velocity $v = \frac{nh}{2\pi m_{e-} r}$

If we replace the expression for v in the relation : $E = -\frac{ke^2}{2r} = -\frac{m_{e-}v^2}{2}$

We obtain the expression $E = -\frac{m}{2}\frac{n^2h^2}{4\pi^2 m_{e-}{}^2 r^2} = -\frac{ke^2}{2r}$

$E = \frac{1}{2}\frac{n^2h^2}{4\pi^2 m_{e-} r} = \frac{ke^2}{2}$

Replacing the expression of the radius in the relation: $E = -\frac{ke^2}{2r}$

We obtain the final expression of energy:

$$\mathbf{E = \frac{-2\pi^2 m_{e-} k^2 e^4}{n^2h^2}}$$ r and E are quantified.

If we put: $\frac{2\pi^2 m_{e-} k^2 e^4}{h^2} = B$ = cste= 13,6 eV = 21,76x10^{-19}j

The expression becomes: $E = \frac{-13,6}{n^2}$eV or $\mathbf{E = \frac{-21,76x10^{-19}}{n^2}}$j

Generally speaking, we can write: $\mathbf{E = \frac{-B}{n^2}}$

I-3- Case of hydrogénoïdes :

the hydrogenoid atom is any atom from which (Z-1) electrons have been removed. Behaves like a hydrogen atom, it is a monoaelectronic ion possessing only one electron like hydrogen. It therefore has a structure similar to that of the hydrogen atom: $1s^1$

He → He^++1e- Z=2 then : $f = \frac{2ke^2}{r^2}$

Li → Li^{2+}+2e- Z=3 then : $f = \frac{3ke^2}{r^2}$

Be $\rightarrow$ Be^{3+} +3e- Z=4 then : $f = \frac{4ke^2}{r^2}$

Generally speaking, the expression for the force of attraction becomes:

$$\mathbf{f} = \frac{Zke^2}{r^2}$$

Hence the final expression for the radius:

$$r_{Hn+} = \frac{n^2h^2}{4\pi^2 m_{e-} Zke^2}$$

The relationship between the radius of a hydrogenoid (hydrogen-like) and the hydrogen radius:

$$r_{Hn+} = \frac{r_H}{Z}$$

The expression of energy is :

$$E_{Hn+} = \frac{-2\pi^2 m_{e-} Z^2 k^2 e^4}{n^2h^2}$$

Generally speaking $E_{A+} = Z^2 x E_H$

⟹ for He^+ Z=2 we have : $r_{He+} = \frac{r_H}{2}$ and $E_{He+} = 2^2.E_H$

⟹ for Li^{2+} Z=3 we have : $r_{Li2+} = \frac{r_H}{3}$ and $E_{Li2+} = 3^2.E_H$

⟹ for Be^{3+} Z=4 we have : $r_{Be3+} = \frac{r_H}{4}$ and $E_{Be3+} = 4^2.E_H$

I-4- Electronic transition :

The electronic transition represents the passage of an electron from one energy level to another,this passage is accompanied by an absorption or emission of energy.

An electronic transition is characterised by:

A change in energy : **ΔE**.

A wavelength:**λ**

A wave number :**σ**

A frequency:***ν***

A period : **T**

Applying BOHR's 3rd postulate, we obtain:

$\Delta E = E_2 - E_1 = h\upsilon = \frac{2\pi^2 mk^2 e^4}{h^2}(\frac{1}{n_1^2} - \frac{1}{n_2^2})$ with $n_2 > n_1$

Knowing that the wavelength $\lambda = \frac{C}{\upsilon}$

λ: wavelenght (m)

c: speed/velocity of light= 3×10^8 m/s.

ν : frequency (s^{-1})

$$\Delta E = E_2 - E_1 = h\upsilon = B(\frac{1}{n_1^2} - \frac{1}{n_2^2}) \quad \text{with} \quad B = \frac{2\pi^2 mk^2 e^4}{h^2}$$

From the expression $\Delta E = h\frac{c}{\lambda} = h\upsilon$

We can deduce the frequency $\upsilon = \frac{\Delta E}{h}$

$$\upsilon = \frac{B}{h}(\frac{1}{n_1^2} - \frac{1}{n_2^2})$$

$$h\frac{c}{\lambda} = \frac{2\pi^2 mk^2 e^4}{h^2}(\frac{1}{n_1^2} - \frac{1}{n_2^2}) \quad \frac{1}{\lambda} = \frac{2\pi^2 mk^2 e^4}{Ch^3}(\frac{1}{n_1^2} - \frac{1}{n_2^2})$$

Hence $\frac{2\pi^2 mk^2 e^4}{Ch^3}$ = R_H the RYDBERG constant.

And $\frac{1}{\lambda}$= σ wave number (cm^{-1}).

$\frac{1}{\lambda} = \sigma = R_H(\frac{1}{n_1^2} - \frac{1}{n_2^2})$ RITZ’s formula.

$R_H = 109738\ cm^{-1} = 1{,}09 x 10^7 m^{-1}$ calculated from m, e, c and h (theoretical value).

I-5- Validity of BOHR's theory :

This theory has proved incapable for atoms other than the hydrogen atom. We now know that the angular momentum σ of hydrogen in the n=1 (ground) state is zero, whereas the BOHR model predicts a momentum equal to m.υ.r, which shows that even for hydrogen this model is inaccurate.

For this, it was necessary to develop this theory even further to apply it to other atoms, so quantum mechanics had to be used.

II- BOHR's model and quantum mechanics:

To understand the foundations of quantum mechanics, all we need to do is interpret the photoelectric effect.

Photoelectric effect:

When, under certain conditions, light rays strike the surface of a metal, electrons leave the metal. To interpret this phenomenon, Einstein assumed that the energy transported by a light wave does not flow continuously, but moves in packets or quanta of energy, also known as photons. These energy quanta are distinct from one another.

W is the energy of a photon.

υ the frequency of the light radiation.

We have:

$\Delta E = W = h\upsilon$

$\Delta E = W = mc^2$ m: equivalent mass.

c: speed/velocity of light

$mc^2 = h\upsilon \qquad \upsilon = \frac{mC^2}{h} = \frac{C}{\lambda}$

So $\lambda = \frac{h}{mC}$

- Louis DE BROGLIE (1924) associated with any particle of mass m, moving at speed v, a wave of wavelength λ such that :

$$\lambda = \frac{h}{mv}$$

- Order of magnitude of λ of the wave associated with the electron: if an electron is subjected to a potential difference V (volts), it has a kinetic energy
- $\frac{1}{2}mv^2 = eV$ hence $v = \sqrt{\frac{2eV}{m}}$ we replace in the expression of λ

$$\lambda = \frac{h}{mv} = \frac{h}{\sqrt{2meV}}$$

If V = 100 volts, we obtain:

$$\lambda = \frac{6{,}62 \times 10^{-34}}{(2 \times 0{,}91 \times 10^{-30} \times 1{,}6 \times 10^{-19} \times 100)^{\frac{1}{2}}} = 1{,}2 \times 10^{-10}\ m = 1{,}2\ A° \text{ (angstrom).}$$

- Therefore, in terms of electron waves, BOHR's 1st postulate means that the wavelength of a stationary orbit must be equal to an integer number of wavelengths for a stationary wave to be established:

$2\pi r = n\lambda = n\frac{h}{mv}$ where $mvr = n\frac{h}{2\pi}$

We thus find the quantization condition.

- The radius BOHR's orbit is given by:

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

If n = 1 ⟹ $r = \frac{n^2h^2}{4\pi^2 m\, ke^2} = 0{,}529$ A°, this value is called the BOHR radius or the radius of the hydrogen atom.

If n=2 ⟹ r = 4x0,529A° ⟹ $E = \frac{-13.6}{4}$eV

If n=3 ⟹ r = 9x0,529A° ⟹ $E = \frac{-13.6}{9}eV$

If n=4 ⟹ r = 16x0,529A° ⟹ $E = \frac{-13.6}{16}eV$

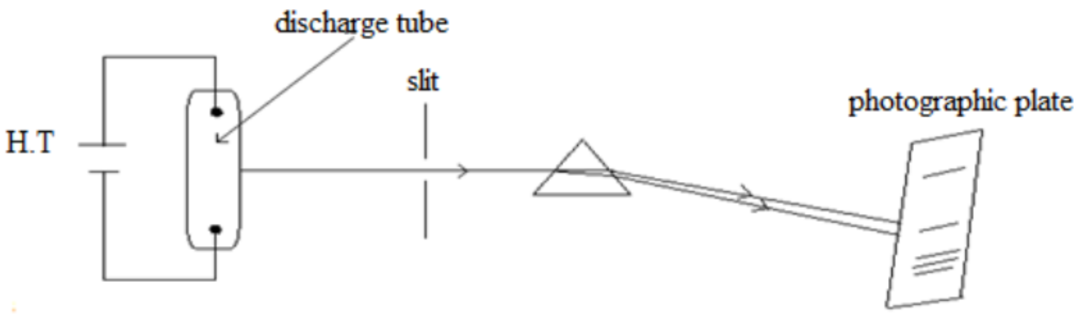
If n=5 ⟹ r = 25x0,529A° ⟹ $E = \frac{-13.6}{25}eV$

If $n \rightarrow \infty$ ⟹ $r \rightarrow \infty$ ⟹ $E \rightarrow 0$

III- Emission spectrum of the hydrogen atom :

To observe this spectrum, an electrical discharge is produced in a tube containing hydrogen under low pressure; the excited atoms emit red light.

A prism or network is used to analyze the light emitted.



The third postulate of Bohr highlights the electronic transition from one energy level to another (absorption or emission of energy).

The photographic plate shows several groups of luminous lines against a dark background. Together, these lines form a hydrogen emission spectrum. Each group of lines is called a series.

- The spectrum includes a number of series, each series contains a number of lines so :

- LYMAN series $n_1 = 1$ ULTRA-VIOLET.

  $n_2 = 2, 3, 4, 5 \ldots\ldots \infty$

- BALMER series $n_1 = 2$ VISIBLE.

  $n_2 = 3, 4, 5, 6 \ldots\ldots\ldots \infty$

- PASCHEN series $n_1 = 3$ Near Infra Red.

  $n_2 = 4, 5, 6 \ldots\ldots\ldots\ldots \infty$

- BRACKETT series $n_1 = 4$ Infra Red.

  $n_2 = 5, 6, 7 \ldots\ldots\ldots\ldots \infty$

- PFUND series $n_1 = 5$ Far Infra Red.

  $n_2 = 6, 7, 8 \ldots\ldots\ldots\ldots \infty$

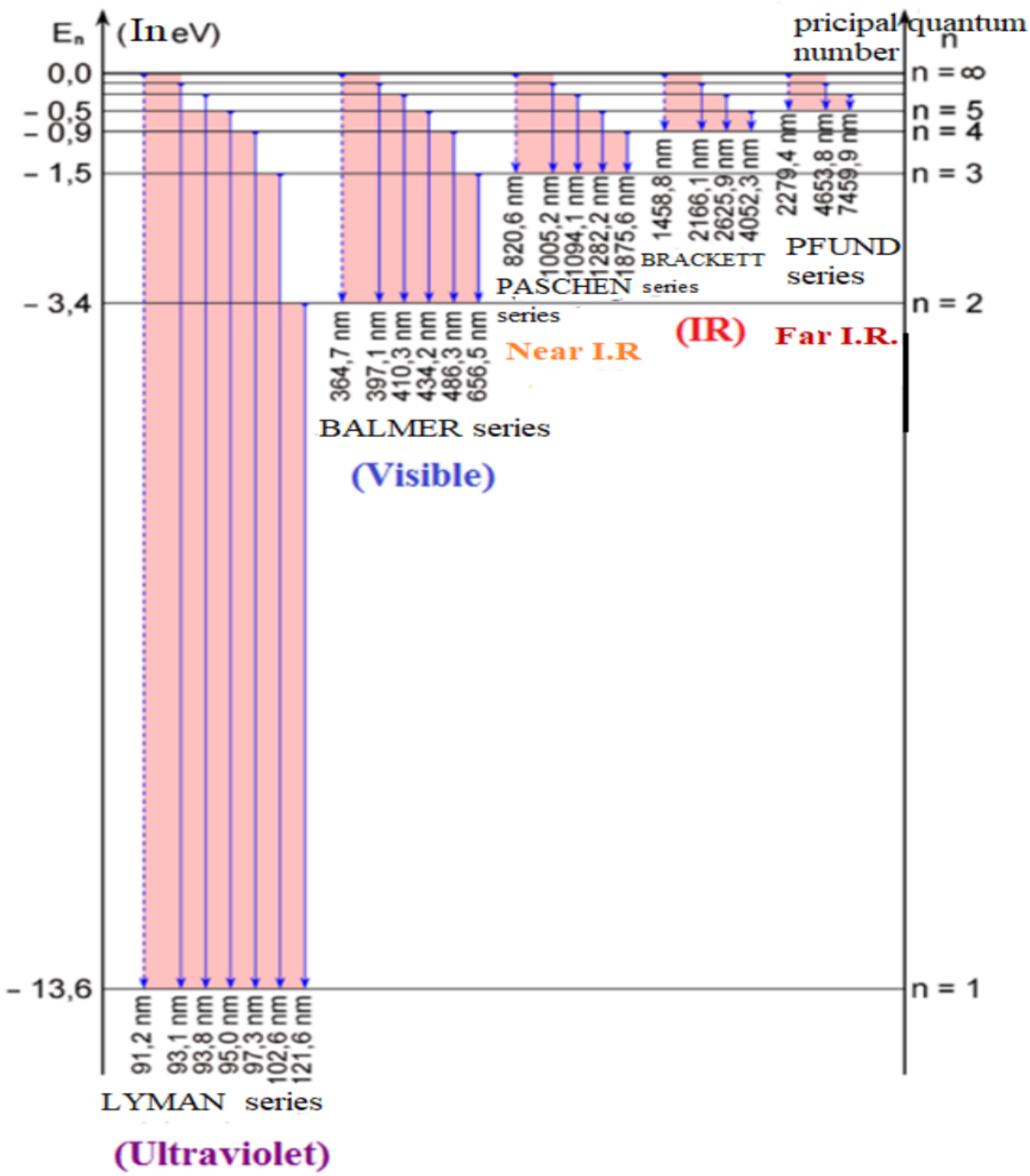


Figure:1 The emission spectrum of the hydrogen atom.

The Bohr theory has varied and fundamental practical applications in many fields of science and technology. Here are some concrete examples demonstrating the practical applications of this theory:

1. **In Atomic Spectroscopy:** The Bohr theory explains the spectral lines of elements. For example, the Bohr model helps to understand the emission lines of hydrogen, visible in the electromagnetic spectrum.
2. **Lasers:** Lasers operate on the principle of stimulated emission, where electrons in atoms or molecules are excited to a higher energy level and, when they return to a lower level, they emit coherent light.
3. **In Plasmas:** In plasmas, atoms are often ionized, and free electrons can move between different energy levels. The Bohr theory helps to explain the interactions between these electrons and nuclei, as well as the energy transitions that occur in these conditions.
4. **In Astrophysics:** The Bohr theory is used in astrophysics to analyze the chemical composition of stars and nebulae by observing their light spectra.
5. **Scanning Tunneling Microscopes (STM):** STMs use the principle of quantum tunneling, which is related to quantum mechanics described by the Bohr theory.
6. **In Electronics and Semiconductors:** Understanding the energy levels of electrons in atoms and materials is essential for developing electronic devices such as transistors and diodes.

IV- The atom – Modern theory:

The electrons are not in a circular orbit, as in the BOHR model, but in an atomic orbital, which is an area of space where the probability of the presence of e- is the highest. Orbitals are defined in wave mechanics by the SCHRÖDINGER equation.

In 1926, SCHRÖDINGER postulated that the wave function ψ (x, y, z, t) is a solution of the following equation :

$$\frac{h^2}{8\pi^2 m}\left[\frac{\partial^2\psi}{\partial x^2}+\frac{\partial^2\psi}{\partial y^2}+\frac{\partial^2\psi}{\partial z^2}\right]+V\Psi = -\frac{h}{2\pi i}\frac{\partial\psi}{\partial t}$$

m: mass of the particle.

V: potential energy = - $\frac{e^2}{r}$

This very general equation admits particular solutions in which the variables of time and space are separated, the solutions are of the form:

Ψ (x, y, z, t) = φ (x, y, z) e^{-iwt}

These functions represent standing waves.

If we substitute a stationary solution in the equation.

$$-\frac{h^2}{8\pi^2 m}\left[\frac{\partial^2\psi}{\partial x^2}+\frac{\partial^2\psi}{\partial y^2}+\frac{\partial^2\psi}{\partial z^2}\right] e^{-iwt} + V\varphi\, e^{-iwt} = \frac{h}{2\pi} w\, \varphi\, e^{-iwt}$$

E = hυ and w = 2πυ

$$\frac{\cancel{h}}{2\cancel{\pi}} 2\cancel{\pi}\frac{E}{\cancel{h}} = E$$

Or: w = 2πυ = 2π $\frac{E}{h}$ because E = hυ

So:

$$-\frac{h^2}{8\pi^2 m}\left[\frac{\partial^2\psi}{\partial x^2}+\frac{\partial^2\psi}{\partial y^2}+\frac{\partial^2\psi}{\partial z^2}\right] + V\psi = E\psi$$

This equation can also be written as:

$$\Delta\psi + \frac{8\pi^2 m}{h^2}(E - V)\,\psi = 0$$

Where Δ is the Laplacian operator.

$$\left[\frac{\partial^2}{\partial x^2}+\frac{\partial^2}{\partial y^2}+\frac{\partial^2}{\partial z^2}\right]$$

Ψ is the part of the wave function that depends only on the space coordinates.

All quantum mechanics consists of solving the SCHRÖDINGER equation, the unknown of which is the wave function ψ(x, y, z).

Solving this equation is a very difficult operation. It therefore requires a computer. The result, from the wave function ψ(x, y, z) therefore defines the space òu in which the electron is located.

IV-1- Quantum Numbers:

The results of solving the SCHRÖDINGER equation are the four quantum numbers, which define the state of an electron:(n, l, m, s s is a number that is defined experimentally).

n : principal quantum number:

Designates the shell (layer) and the energy level n takes the values ***n= 1, 2, 3, 4,……***∞ and defines the energy of the e- placed in a given orbit.

n=1 the 1st shell (K)

n=2 the 2nd shell (L)

n=3 the 3rd shell (M)

n=4 the 4th shell (N)

n=5 the 5th shell (O)

ℓ: secondary quantum number:

Designates the sub-shell (sub-layer),Tapez une équation ici. defines the shape of the orbital on which the e- is located.

For each value of n we have ***$0 \leq l \leq n-1$***

If n = 1 l=0 (s)

If n = 2 $\begin{cases} l = 0(s) \\ l = 1(p) \end{cases}$

If n = 3 $\begin{cases} l = 0(s) \\ l = 1(p) \\ l = 2(d) \end{cases}$

If n = 4 $\begin{cases} l = 0(s) \\ l = 1(p) \\ l = 2(d) \\ l = 3(f) \end{cases}$

m : magnetic quantum number :

Designates the orbital (the number of values of m designates the number of orbitals of the same nature and their orientation in space.

For each value of l, we have ***-l≤m≤+l***

If n = 1 l=0 (s) m=0

If n = 2 l=0 (s) m=0

l= 1(p) m=-1,0,+1

If n = 3 l=0 (s) m=0

l=1(p) m=-1,0,+1

l=2 (d) m=-2,-1,0,+1,+2

If n = 4 l=0 (s)m=0

l=1(p) m=-1,0,+1

l=2 (d) m=-2,-1,0,+1,+2

l=3 (f) m=-3,-2,-1, 0, +1, +2, +3

s : spin quantum number or spin moment

Characterises the rotation of the electron around an axis passing through its centre of gravity (rotation of the electron around itself), and defines the electron's own kinetic energy as it rotates around itself.It takes the values s = ± ½.

IV-2- Orbitals and energy level :

A value n of the principal quantum number corresponds to n^2 wave functions or orbitals and, depending on the value of l, we distinguish between orbitals **s, p, d and .**

The set of $\mathbf{n^2}$ orbitals corresponding to a given value of "n" forms a Shell.

The same value of l corresponds to 2l+1 orbitals, which give a single or multiple sub-shell, depending on whether it contains one or more orbitals.

Each sub-Shell contains 02 more orbitals than the previous one. In a shell n, there are a total of: $1+3+5+7+\ldots+(2n-1)= n^2$ orbitals.

- In the hydrogen atom and hydrogen-like ions ($z = {}_2He^+$; $z = {}_3Li^{++}$; $z = {}_4Be^{+++}$ …), the orbitals in the same shell have the same energy value, since the latter depends on n, but not on l and m. The energy levels are said to be degenerate.
- In a poly-electron atom, on the other hand, the energy corresponding to each orbital depends on both n and l. There is a lifting of the degeneracy of the energy levels.

Tables showing some values for n and l:

Number n	Shell	Number l	Sub-Shell	Number of O's in each Sub-Shell	Number of O's in each Shell
1	**K**	0	S	1	1
2	**L**	0	s	1	
		1	p	3	4
3	**M**	0	s	1	
		1	p	3	
		2	d	5	9
4	**N**	0	s	1	
		1	p	3	
		2	d	5	
		3	f	7	16
5	**O**	0	s	1	
		1	p	3	
		2	d	5	
		3	f	7	25
		4	g	9	

Note :

A quantum orbital can accommodate a maximum of 02 electrons.

Shell K n = 1 2 e^-

L	n = 2	8 e^-
M	n = 3	18 e^-

-In a shell **'n'** it exists **n** sub-shells.

-In a sub-shell **'l '**it exists**'2l+1'** orbital.

-In a shell **'n '** it exists **'n^2 '**orbitals.

-In a shell **'n '** it exists **'$2n^2$ '**electrons.

V- Electronic structure of atoms :

V-1- Pauli's exclusion principle :

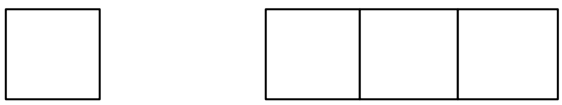
When an atom contains several electrons, their behavior obeys the PAULI exclusion principle.

Two e-'s of an atom cannot have two identical wave functions.

Example:

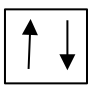
If two electrons have the same values for n, l, and m or the same spatial wave function, they must differ in their spin quantum number, represented by the symbol "s.".

- An orbital is defined by the three quantum numbers n, l, and m. In energy diagrams, it is preferable to represent an orbital using a quantum box:

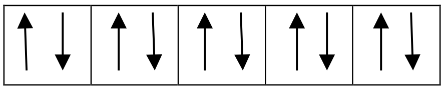


1s 3p

A quantum box can only contain a maximum of 2 e- of opposite spins.

(s=+ ½, s=-½). 

A sub-shell, characterized by the quantum number << l >>, which contains (2l+1) quantum boxes, i.e. a maximum of 2(2l+1) e-.



Sub-shell ‘d’ saturated.

A shell of principal quantum number ‘n’ contains ‘n’ sub-shell i.e :

$1+3+5+......+(2n-1) = n^2$ quantum boxes and a maximum of $2n^2$ electrons.

V-2- Hund's rule:

The rules pertain to the arrangement of electrons in the ground state of atoms.

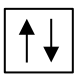
- When electrons occupy a multiple subshell, they fill the maximum number of orbitals.
- In the same subshell, unpaired electrons have parallel spins. In conclusion, we can say that the filling of quantum boxes obeys the principles of PAULI and HUND.

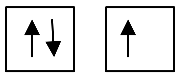
Two electrons in the same orbital can never have identical four quantum numbers; they differ at least by the spin quantum number. In each 'p' or 'd' orbital, the electrons with the same spin are filled first, and then electrons with opposite spins are filled in parallel.

V-3-The principle of stability:

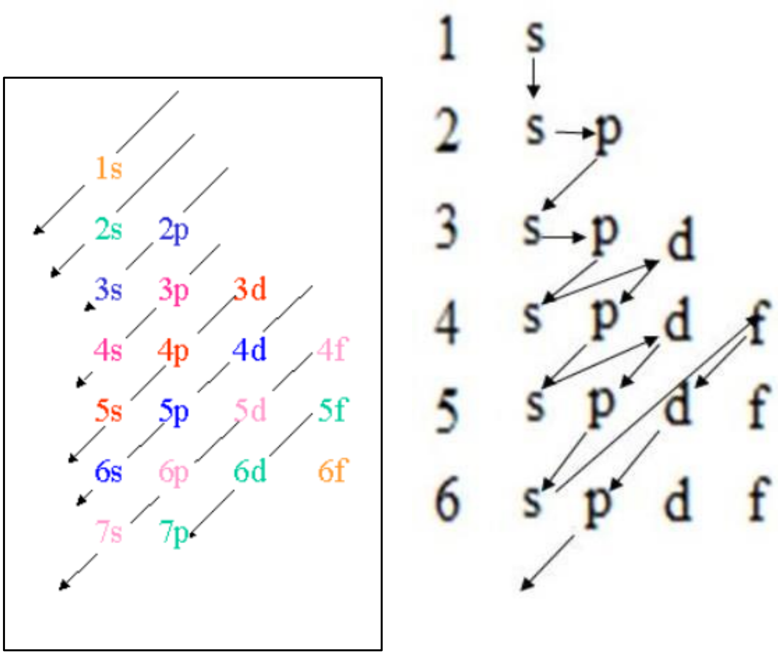
In this case, we consider the ground state of atoms, which is the most stable state. Electrons occupy the lowest energy levels available, up to the limit of available spaces.

Example:

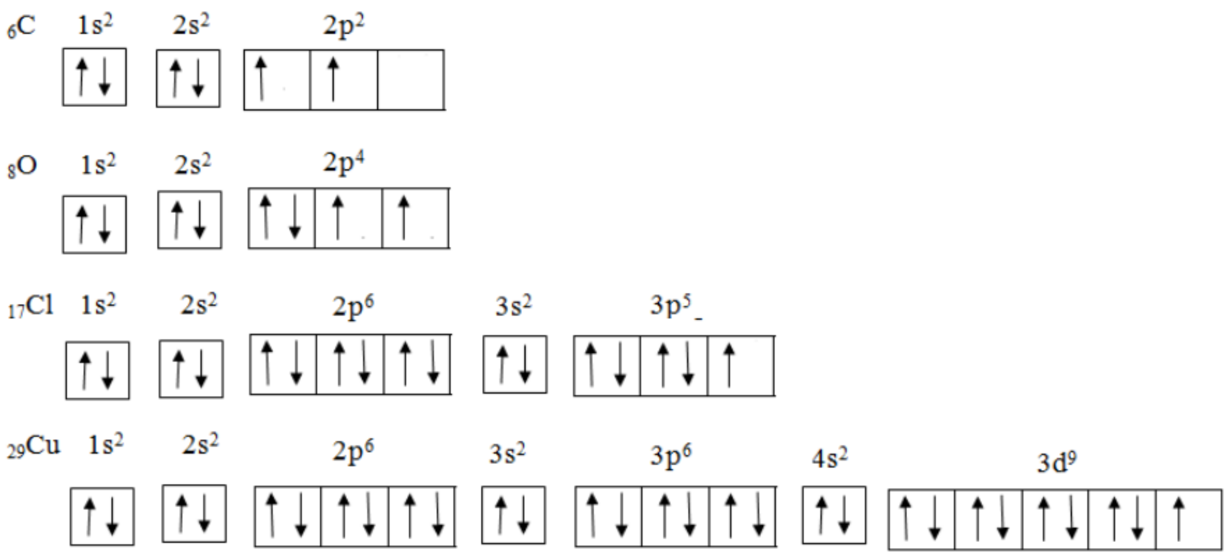
$_2He$ $z = 2\ e^-$ $1s^2$ 

$_3Li$ $z = 3\ e^-$ $1s^2$ $2s^1$ 

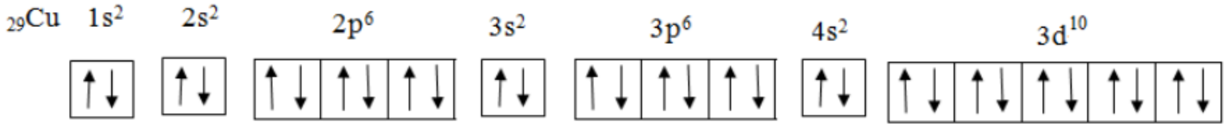
The order of filling sub-shells: KLECHKOVSKY’s rule n l=0, l=1, l=2, l=3 so-called rule of $\sum(n+l)$ minimum



Example:



We can write:



Exceptions to klechkowsky rules

The sub-shells d are stable when half or fully filled.

$(n+1)\ s^2\ \ nd^4 \longrightarrow (n+1)\ S^1\ nd^5$

$(n+1)\ s^2\ \ nd^9 \longrightarrow (n+1)\ S^1\ nd^{10}$

Example: $_{29}Cu$: $1s^2\ \ 2s^2 2p^6\ 3s^2\ 3p^6 4s^2\ 3d^9$ (instable)

$1s^2\ \ 2s^2 2p^6\ 3s^2\ 3p^6 4s^1\ 3d^{10}$ (stable)

$_{24}Cr$: $1s^2\ 2s^2 2p^6\ 3s^2\ 3p^6 4s^2\ 3d^4$ (instable)

$1s^2\ \ 2s^2 2p^6\ 3s^2\ 3p^6 4s^1\ 3d^5$ (stable)

V-4- SLATER's rule:

Slater's method is used to transform a polyelectronic atom into a hydrogen-like atom (hydrogenoid). To simulate a one-electron an electronic system , we calculate i the effective nuclear charge perceived by each electron:

$$Z^* = Z - \sigma$$

Z : The effective nuclear charge.

σ : represents the shielding/screening effect produced by electrons near or close to the nucleus.

The shielding/screening effect σj on electron j is the sum of the shielding effects σj→I exerted on electron j by any electron i.

The following procedure can be followed:

a) The electron configuration, and ordering it according to ;

(1s) (2s,2p)(3s,3p) (3d)(4s,4p) (4d) (4f) (5s,5p)

b) Choosing the electron for which we are looking for the effective charge

♦ Each other electron in the same group as electron j exerts a shielding/ screening effect equivalent to 0.35, with the exception of group 1s for which the shielding of one electron on the other is 0.30.

♦ An electron j in a group ns np undergoes a screening effect of 0.85 by each electron of principal quantum number (n-1) , as well as a screen of 1 by each electron of lower main quantum number.

♦ An electron j of group nd or nf undergoes a screening effect of 1 by each electron of a lower group, either with a lower value, or with the same value of n as the electron j and a lower value of l..

The values of screening constants:

	ej screening electron/shielding electron								
ei studied electron	σij	**1s**	**2s,2p**	**3s,3p**	**3d**	**4s,4p**	**4d**	**4f**	**5s,5p**
	1s	0,30							
	2s,2p	0,85	0,35						
	3s,3p	1	0,85	0,35					
	3d	1	1	1	0.35				
	2s,2p	1	1	0,85	0,85	0,35			
	4d	1	1	1	1	1	0,35		
	4f	1	1	1	1	1	1	0,35	
	2s,2p	1	1	1	1	0,85	0,85	0,85	0,35

Example: Calculation of the effective charge for the last electron

► **$_{17}$Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$** **$(1s)^2$ $(2s,2p)^8$ $(3s,3p)^7$**

$$\sigma = 6x\sigma_{3s,3p} + 8x\sigma_{2s,2p} + 2x\sigma_{1s} = 6x0,35 + 8x0,85 + 2x1$$

This gives an effective charge of: $Z^* = 17 - 10,9 = 6.1$

► **$_{30}$Zn: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$** **$(1s)^2$ $(2s,2p)^8$ $(3s,3p)^8$ $(3d)^{10}$ $(4s)^2$**

$$\sigma = 22\sigma_{4s} + 10x\sigma_{3d} + 8x\sigma_{3s,3p} + 8x\sigma_{2s,2p} + 2x\sigma_{1s}$$

= (2x0,35) + (18x0,85) + (10x1) = 25,65

This gives an effective charge of: $Z^* = 30 - 25.65 = 4.35$

VI- Application exercise

ExerciseN°1

The yellow line in the spectrum of a sodium vapor lamp has a frequency of 5.08×10^{14} sec^{-1}. Calculate:

1. The wavelength of the line.
2. The associated wavenumber.
3. The energy of the emitted photons.

ExerciseN°2

In the Balmer series of the hydrogen spectrum, a line has λ=4340.5 A°. Which transition does it correspond to? Calculate the frequency and energy of the electron in Joules during this transition. Given:

- RH=1.0967×107 m^{-1} and C=3×108 m/ s

ExerciseN°3

In the hydrogen atom, the energy of the electron in its ground state is equal to −13.6 eV. a) What is, in eV, the smallest amount of energy it must absorb to:

- transition to the first excited state?
- transition from the first excited state to the ionized state? b) What are the wavelengths of the emission spectrum lines corresponding to the return:
- from the ionized state to the first excited state?
- from the first excited state to the ground state?

I- Constitution of the matter:

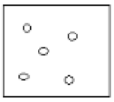
I-1- Different phases of matter:

At the macroscopic scale, a phase is a homogeneous quantity of matter, and we distinguish between the gaseous, liquid, and solid phases.

I-1-1- Gaseous phase (gas phase):

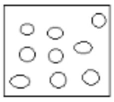
Molecules are practically independent; they move due to thermal agitation in a random motion. Gases form only homogeneous mixtures between each other, for example: air, O_2, H_2, N_2.

The interactions are Weak.



I-1-2- Liquid phase:

The molecules can move relative to each other; the interactions are stronger than in gases.



A liquid mixture can result in a single phase, two miscible phases, or an emulsion.

⇨ **Miscible Phases:** Miscible phases are liquids that can mix together in any proportion without separating into distinct layers.

Examples include :

- **Water and Ethanol:** These two liquids mix in any ratio to form a homogeneous solution.
- **Acetone and Water:** Acetone dissolves completely in water, creating a single phase.

⇨ **Emulsion:** An emulsion is a mixture of two immiscible liquids where one liquid is dispersed in the form of small droplets within the other. Emulsions require an emulsifier to stabilize them and prevent the droplets from coalescing.

Examples include:

- **Oil in Water (O/W) Emulsion:** Mayonnaise and milk are common examples where oil droplets are dispersed in water.
- **Water in Oil (W/O) Emulsion:** Butter and some lotions are examples where water droplets are dispersed in oil.

The stability and properties of emulsions depend on factors like the type of emulsifier used, the ratio of the components, and the processing methods.

I-1-3- Solid phase:

The molecules are arranged in specific, almost immobile patterns.

Mixture of solids ⟶ as many phases as solids.

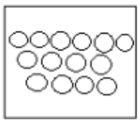
Mixtures of solids can vary widely in their composition and properties.

Here are a few examples:

1. **Sand and Salt:** When mixed, sand and salt form a heterogeneous mixture where the particles of each component are distinct and separable.
2. **Granite:** A natural example of a solid mixture, granite is composed of minerals like quartz, feldspar, and mica, each of which retains its own properties within the rock.
3. **Trail Mix:** A combination of nuts, dried fruit, and chocolate or candy pieces. This mixture is often enjoyed as a snack and can be separated into its individual components.
4. **Powdered Spices:** Mixtures such as curry powder or chili powder, where various spices are combined into a single blend.

5. **Concrete:** A mixture of cement, sand, gravel, and water. The solid components are mixed together to create a durable building material.

In these mixtures, the individual solid components retain their physical properties and can often be separated by physical methods like sieving or filtration.



Example1 :

$H_{2(g)} + \frac{1}{2} O_{2(g)} \longrightarrow H_2O_{(l)}$ number of phases $\varphi = 2$.

$SO_{3(g)} \longrightarrow SO_{2(g)} + \frac{1}{2} O_{2(g)}$ number of phases $\varphi = 1$.

$H_2O_{(l)} \longrightarrow H_2O_{(g)}$ number of phases $\varphi = 2$.

$2Fe_{(s)} + H_2O_{(g)} \longrightarrow Fe_2O_{(s)} + H_{2(g)}$ number of phases $\varphi = 3$.

This chapter discusses the gaseous (gas) and liquid states

II- The gaseous state:

II-1- Characteristics of the gaseous (gas) state:

A gas is composed of molecules that are quite far apart from each other and exhibit random movements. The gas is characterized by its state variables: volume, pressure, and temperature.

a) The Volume: gases occupy volumes that depend on pressure and temperature.

The volume can be expressed in mL or cm^3, L or m^3.

$1L = 10^3 mL = 10^3 cm^3$.

$1m^3 = 10^3 L$.

b) The Pressure: It is a force exerted on a unit area. $\vec{F} \perp \vec{A}$. It is expressed in Pascal (Pa), atm, mmHg, cmHg. Pressure is inversely proportional to volume.

$P = \frac{F}{A}$ (F⊥A)

- F : Newton force (N).
- A : area (m^2).
- P : pressure (Pa) Pascal

1 atm = 760 mmHg = 76 cmHg

1atm = 101325 Pa = $1{,}012 \times 10^5$ Pa.

1 atm = 1,013 10^6 dynes/ cm^2

1 cmHg = 1,33 10^4 dynes/ cm^2

c) The temperature: expressed either in degrees Celsius or in kelvin

T (k)= T°c + 273,15

d) Standard conditions of temperature and pressure (STP):

These standard conditions correspond to the absolute temperature T = 0°C and atmospheric pressure P = 1 atm.

The gas law is the law that relates the volume of a gas with the pressure and temperature.

II-2-The ideal gases (perfect gases):

In the case of perfect gases, the interactions between the gas molecules are neglected. Gases that follow these laws behave like a perfect gas or an ideal gas. Note: these laws only apply to gases that do not undergo chemical transformations when the pressure and/or temperature changes.

P.V = n.R.T

P: Pressure in atm, Pa , mmHg. n: Number of moles in mol.

V: Volume in L, m^3.

T: Temperature in kelvin.

R: Ideal gas constant.

The value of the ideal gas constant:

P	V	$R = \frac{PV}{nT}$
Atm	L	$\frac{1 \, x \, 22{,}4}{1 \, x \, 273}$ = **0,082 l.atm.mol^{-1}.K^{-1}**
mmHg	L	$\frac{760 \, x \, 22{,}4}{1 \, x \, 273}$ = **62,36 l.mmHg.mol^{-1}.K^{-1}**
Pa (1N/m^2)	m^3	$\frac{1{,}013 \, x \, 10^5 \, x \, 22{,}4 \, x 10^{-3}}{273}$ = **8,31j.K^{-1}.mol^{-1}**

II-3- The laws of gases:

II-3-1- BOYLE's law (Mariotte's Law):

When the temperature is kept constant, the volume of a given mass of an ideal gas varies inversely with the pressure to which the gas is subjected.

$$(V \text{ x } P)_T = C^{ste}$$

Consider a gas undergoing a transformation from an initial state to a final state

$P_1V_1 = P_2V_2 = cst$

We have: $\mathbf{PxV = C^{ste}}$

II-3-2- CHARLES's law:

At constant pressure, the volume of a given mass of gas varies proportionally with the temperature.

V = kT

$$\left(\frac{V}{T}\right)_{initial} = \left(\frac{V}{T}\right)_{final}$$

$$\left(\frac{V}{T}\right)_P = C^{ste} \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

II-3-3- GAY-LUSSAC's law:

At constant volume, the pressure of a gas varies proportionally with temperature.

$$\left(\frac{P}{T}\right)_V = C^{ste} \; \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

II-3-4- DALTON 's law of partial pressures:

Consider a mixture of gases under total pressure, each component is characterized by its own pressure, known as the partial pressure. The total pressure of the mixture is equal to the sum of the partial pressures of all the gaseous constituents.

The partial pressure (P_i) of a component in a gas mixture is the pressure that this gas would exert if it alone occupied the entire volume.

The expression for the partial pressure P_i is given by the following relationship:

$$P_i = P_T . x_i$$

With: P_T is the total pressure is : $P_T = \sum_{i=1} Pi \;\; = P_1 + P_2 + P_3 + ... + P_i$

$x_i = \frac{n_i}{\sum n_i}$ Molar fraction

Total number of moles: $n_T = n_1 + n_2 + n_3 + ... + n_i = \sum_{i=1} ni$

$\sum_{i=1} xi = 1$

Example :

At atmospheric pressure, we mix 0,5g d'O_2, 0,8g d'H_2 and 3,6g d'N_2 calculate the partial pressure of each constituent in the mixture.

We give : H=1g/mol, O=16g/mol, N=14g/mol

1) Calculation of the number of moles of each gas :