MOHAMED KHIDER UNIVERSITY OF BISKRA.

FACULTY OF EXACT SCIENCES AND NATURAL AND LIFE SCIENCES

DEPARTMENT OF BIOLOGY

COURSE TITLE: General Chemistry I

CHAPTER I

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1. Generality

1.1. Atom, nucleus, isotopy,…

1.1. Stability and cohesion of the nucleus, binding energy per nucleon,…

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1. Generality

Matter is made up of atoms, themselves made up of a nucleus surrounded by an electronic cloud (Fig. 1 and Table. 1). The central spherical nucleus is composed of A nucleons distributed into:

- Z protons (charge q = + e; mass mp).
- N neutrons (charge q = 0; mass mn).

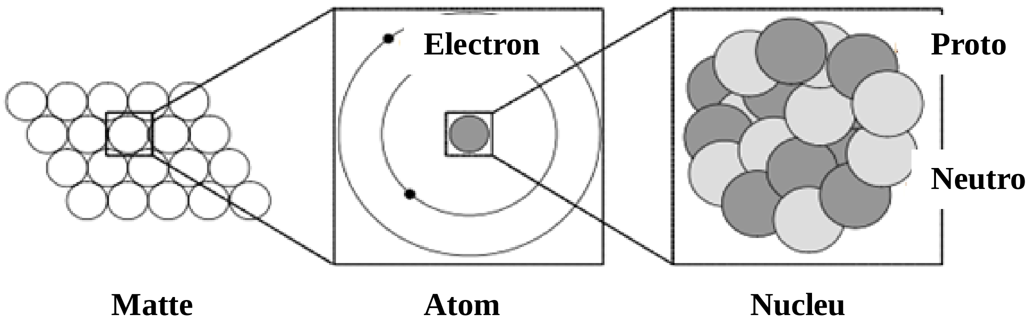

Figure: Schematic view of the structure of matter

1.1. Atom, nucleus, isotopy,

1.1.1. <u>Atoms and subatomic particles:</u>

- Atoms are mostly empty space.
- Protons and neutrons occupy small, dense nucleus.
- Electrons move around outside of the nucleus.

The proton and neutron are each more than 1800 times heavier than the electron.

Charge of proton = −(charge of electron) = e

Table : Main characteristics of the electron, proton and neutron

Particules and symboles		Autors	Charge*	Masses (m_e, m_p, m_n)
Electron	**e –**	J. J. Thomson (1897) R. A. Millikan (1911)	$-1,6.10^{-19}$ C	$9,11.10^{-31}$ kg)
Proton	**p**	E. Rutherford (1910)	$+1,6.10^{-19}$ C	$1,67.10^{27}$ kg)
Neutron	**n**	J. Chadwick (1932)	0	$1,67.10^{-27}$ kg)

1.1.2. Nucleus characteristic $^{A}_{Z}X^{q}$

The nucleus always has a positive electric charge "q" which is an integer multiple of the elementary charge e: q= Z.e (Z∈N, N integer).

The nucleus is characterized by two numbers Z and A:

Z: is called atomic number: characteristic of the atom, represents the number of electrons.

Z: designates the number of charge: characteristic of the nucleus represents the number of protons.

A: designates the number of a mass, it also represents the number of protons and neutrons constituting the nucleus (the number of nucleons); it is the whole number closest to the atomic mass evaluated in atomic mass unit (u) such that: $A = (m_p+m_n) - m_{nucleus}$.

Definition: The atomic mass unit u is equal to 1/12 of the mass of the isotope 12 of carbon 12C (including the mass of the gram atom "atg" of 12C = 12.00000g). So :

$$1u = \frac{1}{12} * \frac{12*10^{-3}}{\mathcal{N}} = \frac{10^{-3}}{6.0221367*10^{23}} = 1.6605402*10^{-27}Kg$$

Or: $\mathcal{N}$: Avogadro number=$6.022167*10^{23}$ mol^{-1}.

Avogadro's constant represents the number of particles per mole.

1.1.3. Concept of isotopy

Isotopes of the same element are species which have the same number of protons (same Z) and a different number of neutrons (therefore different A).

Example :	*Carbon*	*Hydrogen*
	$^{12}_{6}C$	$^{1}_{1}H$
	$^{13}_{6}C$	$^{2}_{1}H$
	$^{14}_{6}C$	$^{3}_{1}H$

a. Relative abundance of different isotopes

By natural abundance we designate the percentage in number of atoms of each of the isotopes present in the natural mixture.

The natural abundance of each isotope is always the same regardless of the origin of the sample studied.

Example: Carbon has two natural stable isotopes: commonly called Carbon 12 and Carbon 13. Their natural abundances are as follows:

Number of Mass: ^{12}C ^{13}C
Abundance: 98.9 % 1.1%

b. Molar Mass of the element (Average or isotopic)

An element is made up of a mixture of various isotopes and the proportions of these various isotopes are constant. We will be able to define for each element an average molar mass which will take into account its composition.

$$M_A = \frac{\sum_{i=0}^{n} X_i * M_i}{\sum_{i=0}^{n} X_i}$$

Xi: designating the natural abundance of isotope i with molar mass Mi.

Mi: Molar mass of isotope i.

Example: Calculate the average isotopic mass of carbon:

$M_C = 98.89 * M(^{12}C) + 1.1 * M(^{13}C)/100$

If we do not need extreme precision we can assimilate the molar masses of each of the isotopes to their mass number.

$M_C = 98.89 * 12 + 1.1 * 13/100 = 12.02\ g.mol^{-1}$

Note: For certain elements, there are also unstable natural or artificial isotopes called radioactive. Due to their instability, their abundance varies over time and is therefore never specified.

1.1.4. Stability and cohesion of the nucleus, binding energy per nucleon

a. Mass and energy: Einstein's relationship

During a nuclear transformation (natural or induced), the mass of the products is always a little lower than the mass of the reactants. The mass loss is denoted Δm. Associated with this loss of mass, there is a release of energy whose value ΔE is given by the Einstein relation:

$$\Delta E = (\Delta m)\, C_0^2$$

C_0: is the speed of light in a vacuum. The convention adopted here is to represent Δm and ΔE by positive quantities.

b. Cohesive energy of a nucleus

The mass of an atomic nucleus is always a little less than the sum of the masses of its nucleons. The difference Δm between the mass of the nucleons and that of the nucleus is called mass defect. The corresponding quantity of energy ΔE given by the relation $\Delta E = (\Delta m).C_0^2$ is the energy which would be released if the nucleus was formed from nucleons.

A nucleus is all the more stable as this quantity ΔE is large. This is why ΔE is called cohesion energy (or binding energy) of the nucleus. To compare the stabilities of the nuclei of various elements, we compare the cohesion energies referred to 1 nucleon, and expressed in MeV/nucleon* (*: 1 eV = $1.602.10^{-19}$ J; 1 MeV= $1.602.10^{-19} \times 106 = 1.602.10^{-13}$ J).