Chapter I

**General**

In antiquity, the atom was considered the elementary particle of matter (Democritus). This theory was only due to intuitions and not to theories or experimental observations. It was not until the end of the 19th century that this theory be confirmed by experiments which made it possible to demonstrate the existence of constituent particles of the atom: electrons

**1.1. Constitution of the atom**

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

- Z Protons) charge q = + e ; mass mp).

- N neutrons (charge q = 0; masse mn).



**Fig.1: Schematic view of the structure of matter**

 Nuclear cohesion is ensured by very short-range nuclear forces, known as strong interactions.

The electron cloud of a neutral atom is composed of Z electrons (charge q = -e; me <<mp and mn). The elementary charge, in coulomb, is e = 1.602×10-19 C.

 The masses of the proton and the neutron are close to each other, and equal to 1836 times the mass of the electron (Table. 1). The mass of an atom is therefore concentrated in its nucleus. The size of the atoms is of the order of 10–10 m; the size of the nuclei, on the order of 10–15 to 10–14m.

* + 1. **The electron**

 The electron is a particle which intervenes in many physical phenomena and, especially in all electrical phenomena (example: passage of electric current in a conductor), the electron is designated by the symbol e. It is a very low mass, negatively charged particle.

**1.1.2. Core structure**

 Protons and neutrons (called nucleons) are located in the nucleus of the atom, while the electrons form the procession or electronic cloud around the nucleus. The mass being concentrated in the nucleus. Since atoms are electrically neutral, there is as many electrons as protons.

 **Neutrons** $n \_{0}^{1}$

 These are neutral particles (from the electrical point of view zero charge) whose mass is close to that of the proton.

**Protons (p)**

 These are positively charged particles whose charge is equal to that of electrons but with the opposite sign.

**Kernel characteristic**

 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 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 = ( mp + mn ) - mnucleus.

**Definition**

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

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Or: Avogadro number=6.022167\*1023 mol-1. Avogadro's constant represents the number of particles per mole.

**Concept of isotopy**

**Definition:**

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

**Example**:

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**Example** :

|  |  |
| --- | --- |
| **Carbone** | ***Hydrogène*** |
| *12**6C* | *1**1H* |
| *13**6C* | *2**1H* |
| *14**6C* | *3**1H* |

**Relative abundance of different isotopes**

We denote by natural abundance the percentage in number of atoms of each of the isotopes present in the natural mixture. This abundance is equivalent to the fraction molar of each stable isotope. This natural abundance could be measured and can be found in tables. 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:

**Mass Number :** 12 13

**Abundance** 98,9 %1,1%

**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 a molar mass for each element average which will take into account its composition .

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

Mi : Molar mass of isotope i

**Example:**

Calculate the average carbon isotope mass

MC = 98,89 \* M(12C) + 1,1\* (M13C)/100

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

MC = 98,89 \* 12 + 1,1\* 13/100 = 12,02 g.mol-1

**Noticed:**

For some elements, there are also natural or artificial isotopes unstable called radioactive. Due to their instability, their abundance varies over the course of the time and is therefore never specified. Stability and cohesion of the nucleus, binding energy per nucleon

**Mass and energy: Einstein's relationship**

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

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

**Cohesion 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 ground fault. The corresponding quantity of energy ΔE given by the relation ΔE = (Δm) C0 2 is the energy that 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 them the stabilities of the nuclei of various elements, we compare the cohesion energies related to 1 nucleon, and expressed in MeV/nucleon 1 eV = 1,602.10–19 J; 1 MeV= 1,602.10–19 × 106 = 1,602.10–13 J).

***Orbitale et nombre quantique magnétique (m)***

**Electronic configuration of atoms**

**Introduction to quantum numbers**

 The state of an electron in an atom that is to say: its energy, its movements around of the nucleus, the shape of the orbital, is defined by 4 parameters called quantum numbers.

**Electronic layer and principal quantum number (n)**

In 1913, BOHR developed a theory based on the hypothesis that "the energy of a electron does not vary continuously, but in leaps. When the electron is at its level of lowest energy, it is said to be at its base or ***fundamental level***.

 These energy levels of electrons or electronic layers are numbered 1, 2, 3...7 from the basic level

and the order number of the different levels is called ***number main quantum***.

The number ***n, principal quantum number*** defined by:

***n > 0*** *⇒* n = 1, 2, ….7

An energy level signified a layer or period or row in the table periodic. The energy levels corresponding

 to n = 1, n = 2... n = 7 are called K layers, L, M, N, O, P, Q.

 Quantifies and defines an energy level of the electron



 Calculate the number of OAs: n2

 The number of electrons in the different levels is limited to 2n2.

**Example**: there will be no more than 2 electrons and 1 atomic orbital at the lowest level because

(n = 1 ⇒ 2n2 e- = 2\*12 = 2 e- et n2 = 12 = 1 OAs).

Electronic subshell and secondary quantum number (l) it tells you the irregularities in the refill of the couches or the energy supplies. Our electronic devices are available when they exist, on a couch, in our homes, different energy levels. Each layer must be considered as a set of sub-layers, (l) whose energies differ. The number of sublevels is given by the value of the number principal quantum n.

The number l, secondary quantum number defined by:

0 ≤ ***l*** ≤ n–1

A sublevel of energy means an underlayer

The corresponding energy sublevels has = 0, n = 1, n = 4 are called sublayers s, p, d, f.

 characterizes the “shape” of the orbital; it defines an electronic sublayer, or a sub-level of energy.

**Orbital and magnetic quantum number (m)**

 Furthermore, for each sublevel, the magnetic quantum number (m) characterizes the various electronic orbitals, each of which has the same electronic energy but a different spatial orientation.

There are, for each subshell, [2 k-1] orbitals of forms close but of different orientation.

The number **m, magnetic quantum number** defined by:

-*l* ≤ ***m*** ≤ +*l*

Defines the orientation of the orbital

**Example:**



*l = 0* ⇒ m = 0 ⇒ 1 only orientation ⇒ 1 orbital s ⇒ 1 quantum box

*l =1* ⇒ m= -1; 0 ; 1 ⇒ 3 orientations ⇒ 3 p orbitals of the same energy ⇒ 3 boxes

quantum

**Number of electrons per orbital and spin (s)**

 In addition to the three quantum numbers necessary to describe the position in space of the electron in relation to the nucleus, there exists a fourth quantum number, internal to the electron itself, which describes the rotation (we often retain the English word "spin") of the electron on itself. Given the fact that this rotation can only be carried out in the positive direction or negative direction, this number can only take two values.

**↑:** electron of S= +1/2, **↓:** electron of

**Example**

🖎 Determine the quantum numbers associated with the principal quantum number n=3?

🖎 How many atomic orbitals are associated with n=3? How many electrons?

Answer: We have: **n** > 0 et 0 ≤ ***l*** ≤ n–1, -*l* ≤ **m** ≤ +*l,* ***S****=±1/2*

***n*** *> 0 : n*=3 ⇒ Layer M (Line 3 or period 3)

0 ≤ ***l*** ≤ n–1 ⇒ 0 ≤ ***l*** ≤ 2

***l=0*** ⇒ Undercoat s ⇒ 1 OA or 1 quantum box ⇒