Chapter II
The components of matter

I-introduction

In the atomic theory it was considered that the atom was an indestructible entity, at the beginning of this century a large number of experiments follow one another which obliged to revise this theory.

In this chapter we will examine all the experiments that have served to affirm the theory of atomic structure

These experiments were carried out in three stages:

1- Discovery of the electrical nature of matter
2- Discovery that the atom consists of a nucleus and particles
3- Discovery of mechanical laws that show the behaviour of elements in the atom

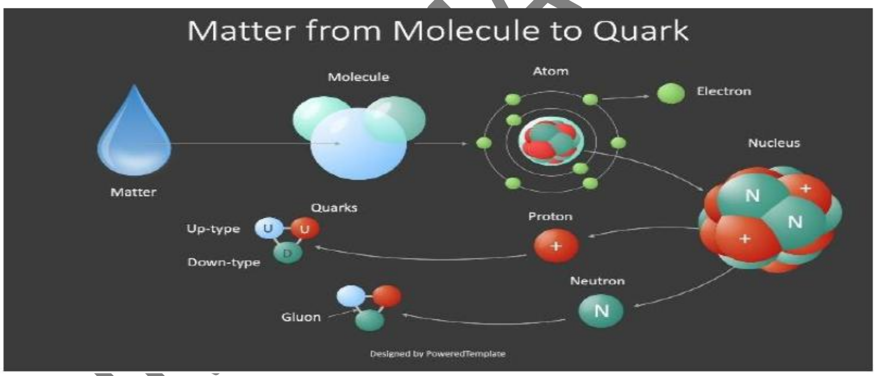


Figure I: The Various Constituents of Matter

II- Discovery of the electric nature of matter: Faraday's experience, The relationship between matter and electricity

The purpose of the FARADAY experiment is the conclusion of the existence of a relationship between matter and electricity.

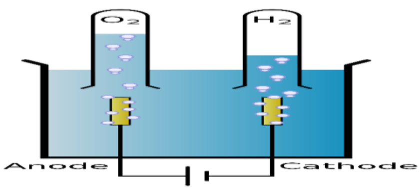


Figure II: Electrolysis of H_2O

II- Matter constituents

In 1807, John Dalton proposed a theory in which he described atoms as indivisible and stable particles. Towards the end of the 19th century, however, new experiments suggested that an atom has an internal structure: it consists of subatomic particles: the electron, the proton and the neutron.

II-1- Electron

II-1-1-Experience of J.J Thomson (1897): The discovery of electron

- In a first part, a beam of cathodic ray is subjected to the action of an electric field (2 plates of a capacitor); it is noted that there is appearance of a luminous point C on the screen.

This explains why these rays contain negatively charged particles (attraction of charges (–) towards the plate (+) of the capacitor).

- In a second part of the experiment, we use a magnetic field (magnet) we notice the appearance of a luminous point A on the screen. This further proves the existence of charges (-) in the cathodic ray (deviation of radiation towards the pole N which is the pole (+) of the magnet.
- In a third experiment the two fields are submitted at the same time, it is noted that there is no deviation of cathode radiation (appearance of a luminous point at point B)

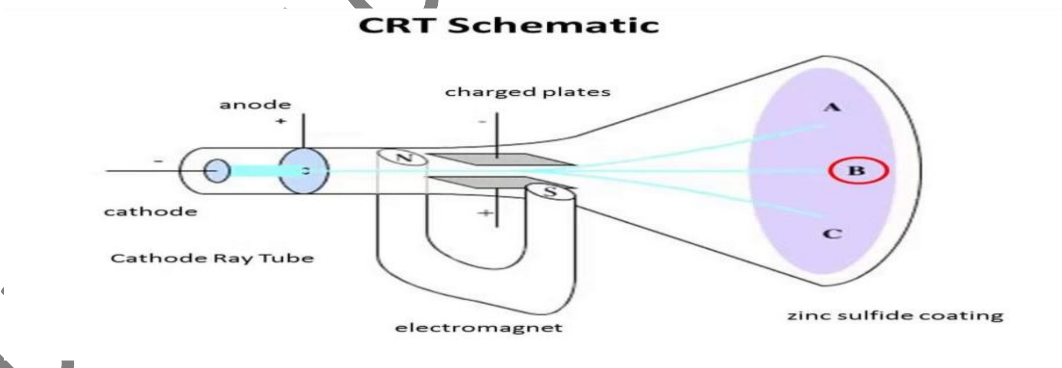


Figure II-1-1: Schematic of the cathode ray tube by J.J. Thomson

Conclusion

Electrons are part of the constitution of all atoms.

By combining the intensity of the electric and magnetic fields, the cathode ray beam may not deviate from its original path.

Thomson thus determined the charge-to-mass ratio of the particles:

$$\mathbf{e/m_e = 1.76 \times 10^{11} \quad C/Kg}$$

II-1-2-Millikan's experience (1909): measurement of the electron charge

The purpose of this experiment is to determine the charge of the electron

- Sprayer: injects droplets of non-volatile liquid
- Microscope: used to measure droplet speed between capacitor plates

By means of X-rays these droplets are charged, when the electric field is applied the droplet rises and when the field is cut it goes down.

There are several forces that are present in each type of movement

1- In a climb E is different from 0
2- In a descent E is equal to 0

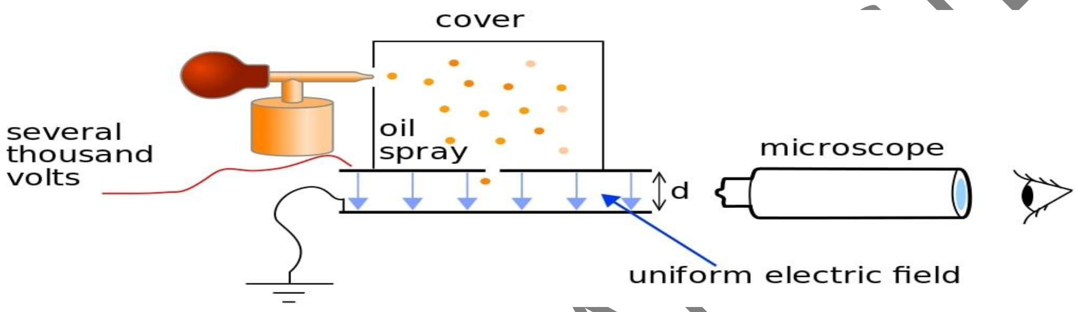


Figure II-1-2: Experimental scheme of the Millikan experiment

From the force of the electric field necessary to cancel the force of gravity (weight) on the droplets, Millikan was able to determine the values of the particle charges.

As each oil droplet contains several additional electrons, it assumed the smallest charge difference between two droplets

Modern value is **-e** with **$e = 1.602 \times 10^{-19}$ C**

C: coulomb, the SI unit of electric charge

It is considered that **–e** is a "unit" of negative charge, and that **e**, called fundamental charge, is a "unit" of positive charge.

The electron mass was calculated by combining this value with the ratio $e/m_e = 1{,}76 \times 10^{11}$ C/Kg, calculated by Thomson we find: **$m_e = 9.109 \times 10^{-31}$ Kg**

III- Planetary Atomic Model of Rutherford (1909)

III-1- Discovery of the nucleus

III-1-1- Gold foil experiment

A thin gold leaf is bombarded by α (charged +) particles emitted by a radium source. The mass of each atom of the gold leaf must be evenly distributed over the entire atom. Therefore, when particles α hit the leaf, they are expected to slow down and change direction only by small angles as they pass through the leaf.

However, in doing so, Rutherford and his assistants make these observations:

- Most particles pass through the gold leaf without deviation as if they had never met the gold atoms.
- Several particles α are slightly deflected during the gold leaf traverse.
- Some particles α (1 in 20.000 to 30.000) appear to be back.

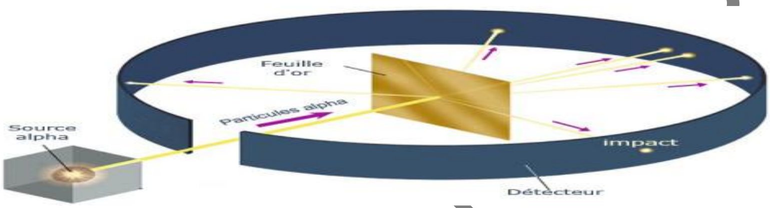


Figure II-1-1: Rutherford experience

Conclusion

-The particles have passed through the leaf without being deflected, most of the space in the atom is empty.

-The deviation of some positively charged α particles must be due to the enormous repulsion force. This suggests that the positive charge is not uniformly distributed throughout the atom as proposed by Thomson. The positive charge must be concentrated in a very small volume to deflect the positively charged α particles.

-Rutherford calculations show that the volume of the nucleus is very small in relation to the total volume of the atom and that the radius of an atom is about 10^{-10} m, while that of the nucleus is 10^{-15} m.

III-1-2-Composition of the Atomic Nucleus

The nucleus is composed of stable elementary particles called nucleons, which exist in two forms when free: the neutron and the proton.

Protons have a positive charge: $+e = 1.602\times10^{-19}$ C

Proton mass: $m_p = 1.673\times10^{-27}$ kg

Neutrons have zero charge, and their mass is: $m_n = 1.675\times10^{-27}$ kg.

All the mass of the atom is concentrated in the nucleus.

IV- Presentation and Characterization of The Atom

Atoms are the particles that make up matter. In the center of the atom, there is a nucleus composed of protons and neutrons. Around this nucleus are very fast moving particles, electrons. In an atom there are as many positively charged protons as negatively charged electrons, an atom is electrically neutral, characterized by a number of electrons that revolve around the nucleus and a number of nucleons that make up its nucleus

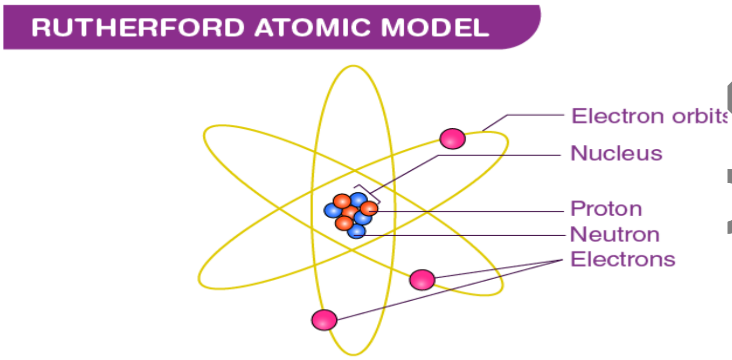
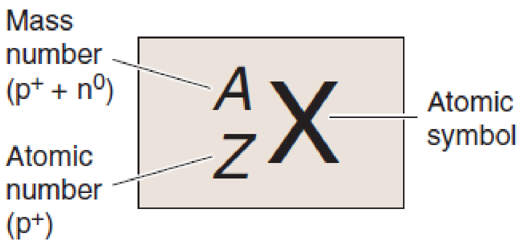


Figure IV-1: Planetary Model (Rutherford) of the Atom

Matter is composed of elementary units called atoms, and there are 112 atoms or elements that have been discovered, an atom or chemical element may be presented by the following symbol:



The atom X is characterized by two numbers:

Z: The charge number or atomic number, represents the number of protons or electrons. For any element, the charge of the nucleus is +Ze, and the charge of the electrons is -Ze.

A: The mass number, represents the number of nucleons (the sum of protons Z and neutrons N) with **A = Z + N.**

A determines the number of **neutrons** N, that an atom has.

If the element is an **anion** (negative charge X^{-Ze}), we must add the charge number to the number of protons. **number of electrons = Z+Ze**

If the element is a cation (positive charge X^{+Ze}), we must subtract the charge number from the number of protons. **number of electrons = Z- Ze**

V- Isotopes and Relative Abundance of Different Isotopes

There are two types of atoms: non-isotopic atoms and isotopes.

V-1- Non-Isotopic Atom

The atomic mass is equal to the sum of the masses of the atom's constituents:

$$\mathbf{M_{atom} = z \times m_{e-} + m_{Nucleus}}$$

Z: Atomic number or charge number, representing the number of protons.

me-: Mass of the electron

mNucleus: Mass of the nucleus

$$\mathbf{m_{Nucleus} = Z \times mp + n \times mn}$$

mp: Mass of the proton

n: Number of neutrons

mn: Mass of the neutron

V-2- Isotopes

When two elements of the same symbol have the same atomic number or number of protons (Z) but the mass number (A) is different they are called isotopes.

Example:

Natural hydrogen consists of three isotopes:

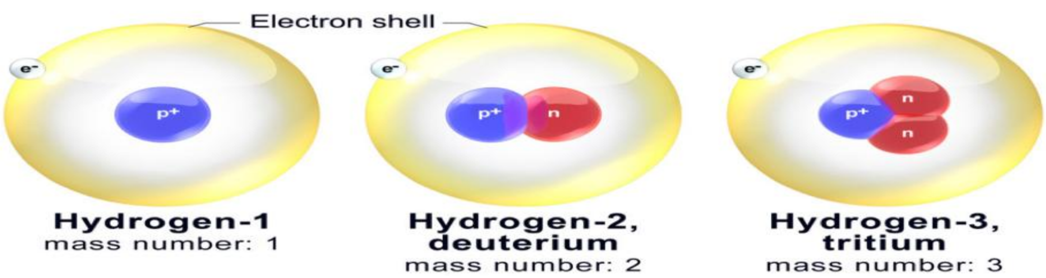


Figure V-2 : Isotopes of Hydrogen

V-3-Relative Abundance of Different Isotopes and atomic mass

In the general case, an element has one or more isotopes, so the atomic mass will be the sum of the relative proportions of each isotope.

The natural abundance (x) or natural isotope content represents the percentage in number of atoms of each of the isotopes present in the natural mixture.

$\mathbf{m = \Sigma\, x_i m_i}$ **(amu)**

Where: M_i : atomic mass of isotope i

X_i : natural abundance of isotope i

V-4- Mass Spectrometer: Bainbridge Spectrograph

to measure the mass of an atom within a mixture of isotopes, the most practical technique is to determine the q/m ratio of an ionized atom, where q denotes the ion's charge and m its mass. This is done using devices called mass spectrometers, such as the Bainbridge spectrograph. These spectrometers are able to distinguish between an element's isotopes and assess their abundances.

This Bainbridge spectrograph consists of four parts:

- A source of ions,
- The speed filter,
- The analyzer
- The ion detector (photographic plate).

An electron beam emitted by a heated filament positively ionizes the atoms of a gas. It receives ions moving at different velocities.

These ions are subjected to the simultaneous action of an orthogonal electric field and a magnetic field. Therefore, the ionic beam exiting the velocity filter is monocinetic.

In this part of the apparatus, an ion with mass m is subjected to a constant magnetic field with induction B directed perpendicular to its trajectory. The ion is then deflected in a circular path. The deflected ion impacts a photographic plate. The detector collects the ions and amplifies the signal, after which a computer system transforms the received information into a mass spectrum.

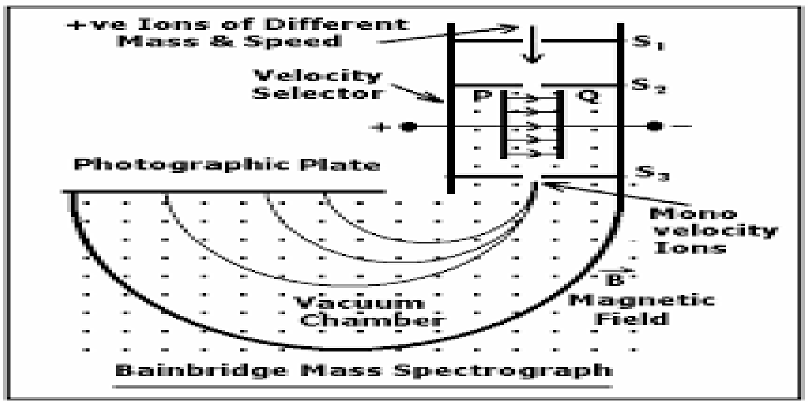


Figure V-4: Principle of the Bainbridge Spectrometer

VI- Nuclear Binding and Cohesion Energy

VI-1- Mass Defect

The formation of a nucleus from its separate nucleons is accompanied by a loss of mass Δm, also known as the mass defect. The mass defect Δm is always positive. Its expression is:

$$\Delta m = (Z \cdot m_p + (A-Z) \cdot m_n) - m_{nucleus}$$

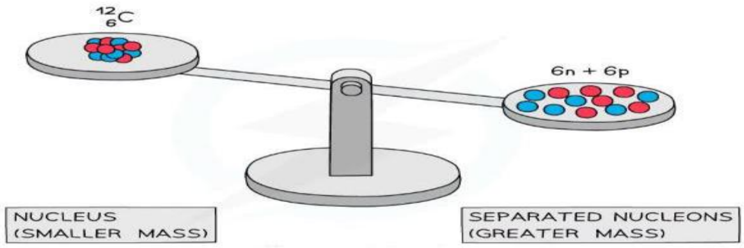


Figure VI-1: Representation of a Mass Defect

VI-2-Binding Energy

According to Einstein, the mass loss is transformed into energy that allows the nucleons to remain bound within the nucleus. $\mathbf{\Delta E = \Delta m\, C^2}$

Binding energy is the energy required to dissociate a nucleus into its nucleon particles.

$\mathbf{E_b = \Delta m\, C^2}$

$\mathbf{E_b}$: binding energy of the nucleus (in J, eV, or MeV)

$\mathbf{\Delta m}$: mass defect of the nucleus (in kg)

C: speed of light in a vacuum (m/s)

This energy is positive because it is received by the considered system (nucleus).

VI-3- Cohesion Energy

Cohesion energy is the amount of energy required to destroy a nucleus into its individual neutrons and protons. This energy is typically negative, indicating that work must be done to break the nucleus apart.

It is related to the total nuclear energy (E) by the equation: $\mathbf{E_b = -E}$

Where:

$\mathbf{E_b}$: represents the binding energy (the energy required to break the nucleus into free nucleons).

E : represents the total nuclear energy, which is negative due to the attractive forces that bind the nucleons together.

VII- Nuclear Stability

VII-1- Determination of Cohesive Energy (Aston's Curve)

Aston Curve: The Aston curve illustrates how the binding energy per nucleon varies with the mass number of different nuclei. It is a crucial tool in nuclear physics for comparing nuclear stability. The binding energy per nucleon is defined as the total binding energy Eb divided by the number of nucleons (A). It is expressed in MeV/nucleon.

The curve shows that:

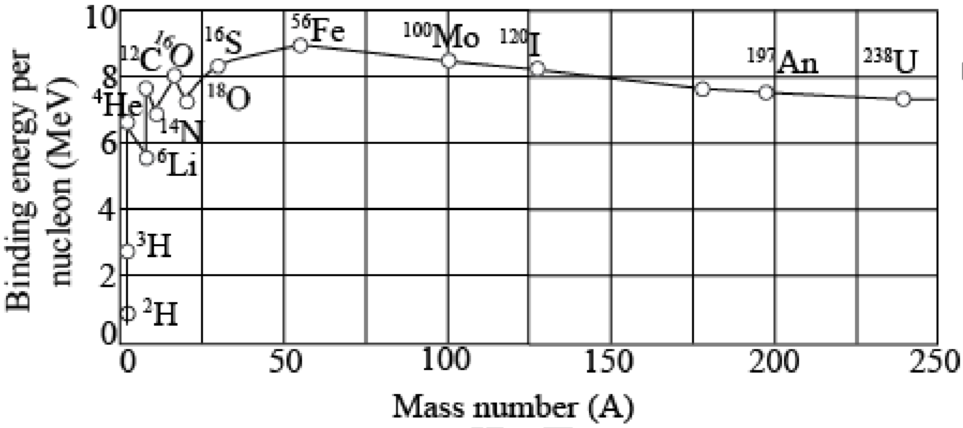


Figure VII-1: Aston Curve

VII-2- Stability And Number of Nucleons

According to Rutherford, the atom is mainly made up of vacuum and all its mass is focused in the nucleus which constitutes protons and neutrons so the stability of a nuclide will be influenced by the number of protons and neutrons. This effect is represented by a stability diagram (see figure) which shows the variation of the neutron number as a function of the atomic number Z in the nucleus.

The shape of this curve shows that stability is:

- important for all elements with Z=N (shown diagonally);
- less important (not stable) between $20 \leq Z \leq 84$, N>Z ;
- very low, the nucleus are very unstable, for $Z \geq 84$,

they are radioactive (α emission).

For the first elements with Z < 30, we observe that stable isotopes have a number of neutrons roughly equal to that of protons, $Z \approx N$.

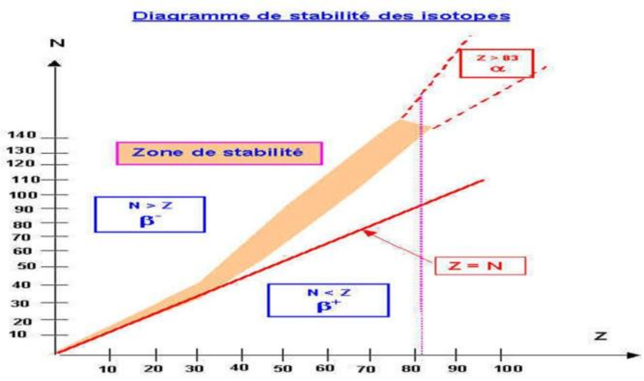


Diagramme de stabilité des isotopes
N
Z > 83
α
Zone de stabilité
N > Z
β⁻
Z = N
N < Z
β⁺
Z
140
130
120
110
100
90
80
70
60
50
40
30
20
10
10 20 30 40 50 60 70 80 90 100