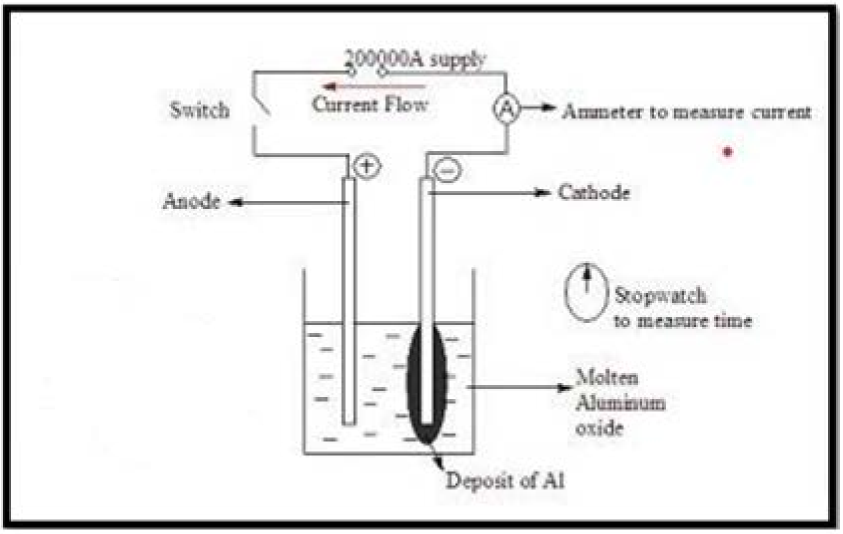
## 1. Introduction

In 1800, the english chemist **John Dalton** suggested that all matter is made up of atoms, which were **indivisible**. Later, scientists discovered particles inside the atom that proved the atoms are divisible. These particles are called subatomic particles.

## 2. Discovery of the electron

### 2.1. Electrical nature of the matter (Experiment of Faraday (1833))

One of the most significant early indications in discovering the electrical nature of matter and the relationship between matter and electricity emerged from the experimental research conducted by scientist Faraday in 1833 in the field of electrochemical analysis (electrolysis).




**At anode (+) (Oxidation) :**

$O^{2-}_{(l)} \longrightarrow ½\ O_{2(g)} + 2e^-$

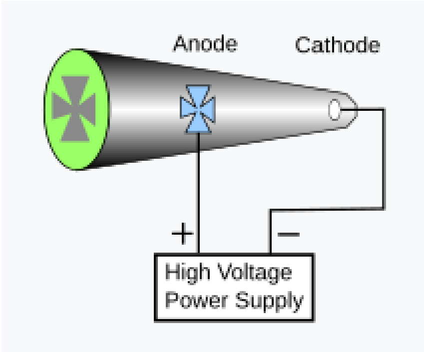
**At cathode (-) (Reduction) :**

$Al^{3+}_{(l)} + 3e^- \longrightarrow Al_{(s)}$

From the experiments conducted by Faraday in the field of electrolysis, it has been found that the weight of the substance appeared at one of the electrodes is proportional to the amount of electricity passing through the solution. This result led Faraday **to assume that electricity is composed of elementary charges, and that atoms contain such charges**.

### 2.2. Cathode Ray Experiment (Experiment of Crooks (1879))

The first ideas about electrons came from experiments with cathode-ray tubes. A cathode-ray tube is a partially evacuated glass tube with electrodes placed in each end. lf a high electrical voltage is applied to the electrodes, an electrical discharge can be created between them. This discharge appears to be a stream of particles emanating from the cathode (there are rays emerging from the cathode towards the anode). These rays were called “Cathode Rays”.



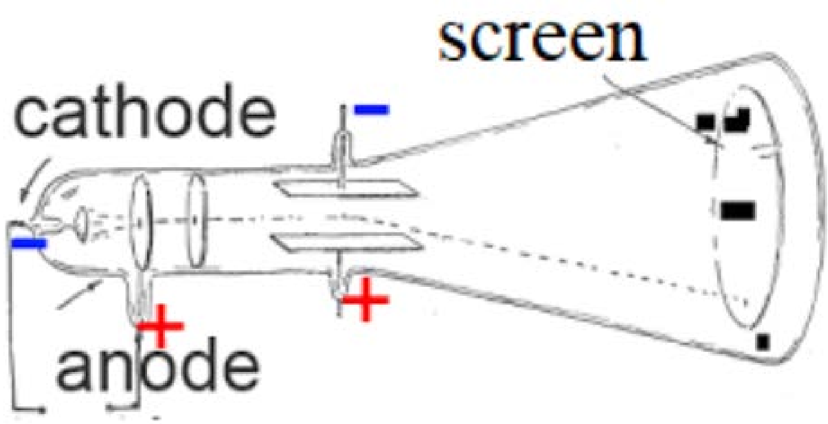

**2.3. Determination of the (|*e*|/m) ratio (Experiment of J.J. Thomson (1897))**

Thomson used cathode ray tube to demonstrate that the cathode ray responds to both magnetic and electric fields. Thomson determined that the cathode ray must be composed of negatively charged particles called "**electrons**" since the ray was deflected toward the positive plate of the capacitor.

The determination of the **(|*e*|/m)** ratio is based on the measurement of the deflection of the cathode ray under an electric or magnetic fields, where:

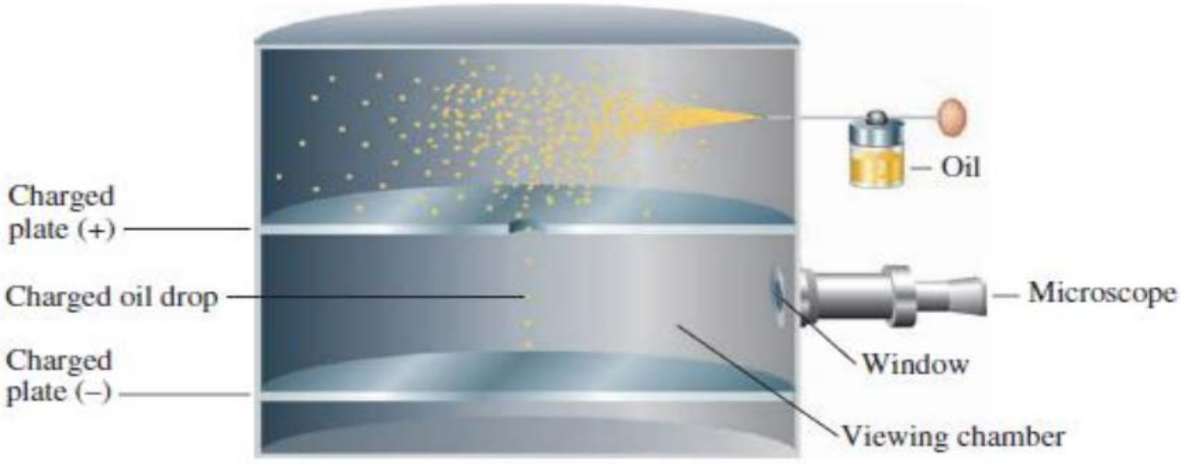
**$|e|/m_e = 1{,}759.10^{11}$ Coulombs/Kg**

m : mass of electron (kg).
e : charge of electron (coulombs).




**2.4. Determination of the charge |e| of the electron and deduction of its mass (Experiment of Millikan (1909))**

Robert Millikan performed a series of experiments in which he obtained the charge of the electron by observing how a charged drop of oil falls in the presence and in the absence of an electric field.




The charge of an electron is the fundamental unit of electric charge.

$q = e = -1{,}60218.\ 10^{-19}\ Coulomb$

With a good estimate of the charge of an electron and the ratio ($|e|/m$):

$|e|/m = 1{,}759.\ 10^{11}\ coulomb/Kg$

We can calculate the mass of the electron as follows:

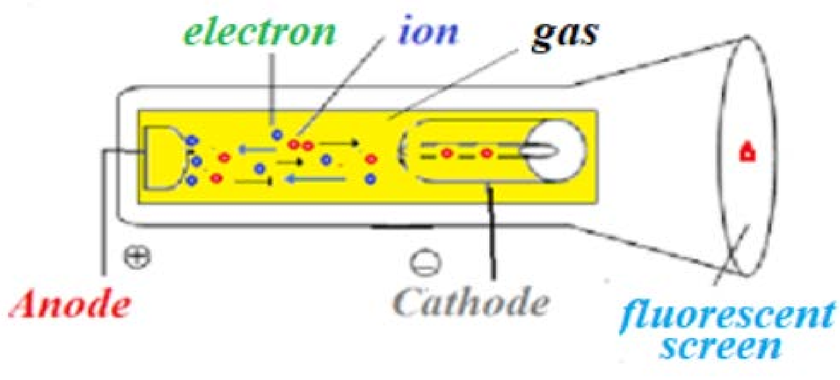
m = $|e|/(|e|/m) = 1{,}602.10^{-19} / 1{,}759.10^{11} = \mathbf{9{,}108.10^{-31}}$ **Kg**

$\mathbf{m_e = 9{,}10939.10^{-31}}$ **Kg**

## 3. Discovery of the proton

### 3.1. Experiment of Goldstein (1886)

In the experiment, Goldstein applied high voltage across a cathode ray tube which had a perforated cathode. A faint luminous rays were seen extending from the holes in the back of the cathode. These rays were called “Anode Rays”. They are composed of positively charged particles.




### 3.2. Experiment of Rutherford (1918)

Rutherford was shooting **alpha** particles ($^{4}_{2}He$) through air, which is mostly nitrogen, with some oxygen and traces of other gasses. He saw twinkles that occurred when the nuclei of hydrogen hit the detector.

He repeated the experiment with vials of pure nitrogen gas and pure oxygen gas. ***He found out that the nitrogen was the source of the hydrogen nuclei signal.*** He didn't see it at all with oxygen.

$^{14}_{7}N + ^{4}_{2}He \longrightarrow ^{17}_{8}O + ^{1}_{1}H$ **(proton : $^{1}_{1}H$)**

<u>*Charge of the proton</u>

**$q_p$= +e = +1,60218. $10^{-19}$ Coulomb**

<u>*Mass of the proton</u>

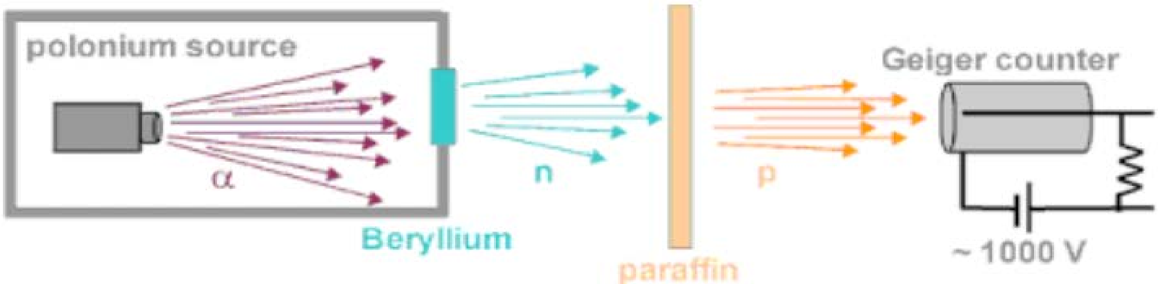
**$m_p$ = 1836 $m_e$ = 1,67262. $10^{-27}$ Kg**

## 4. Discovery of the neutron

### 4.1. Experiment of Chadwick (1932)

Chadwick bombarded beryllium metal (Be) with alpha particles ($^{4}_{2}He$). A strongly penetrating radiation was obtained from the metal which emit protons when it collides with paraffin rich in hydrogen. He showed that this penetrating radiation consists of neutral particles called « **neutrons** ».

$^{9}_{4}Be + ^{4}_{2}He \longrightarrow ^{12}_{6}C + ^{1}_{0}n$ **(neutron : $^{1}_{0}n$)**




<u>*Charge of the neutron</u>

**$q_n$= 0**

<u>*Mass of the neutron</u>

**$m_n$ = 1838 $m_e$ = 1,67493.$10^{-27}$ Kg**

The properties of the subatomic particles are summarized in the following table:

| Particle | Mass (kg) | Charge (C) | Mass (amu)* | Charge (e) |
|---|---|---|---|---|
| Electron | $9.10939 \times 10^{-31}$ | $-1.60218 \times 10^{-19}$ | 0.00055 | −1 |
| Proton | $1.67262 \times 10^{-27}$ | $+1.60218 \times 10^{-19}$ | 1.00728 | +1 |
| Neutron | $1.67493 \times 10^{-27}$ | 0 | 1.00866 | 0 |

**<u>Note</u>**

**<u>Atomic mass unit (amu)</u>**

The atomic mass unit (a.m.u) is a mass unit equal to exactly one-twelfth (1/12) the mass of a carbon-12 atom. It is also called the dalton (Da).

1 a.m.u = (1/12)*(the mass of a ${}_{6}^{12}C$ atom)

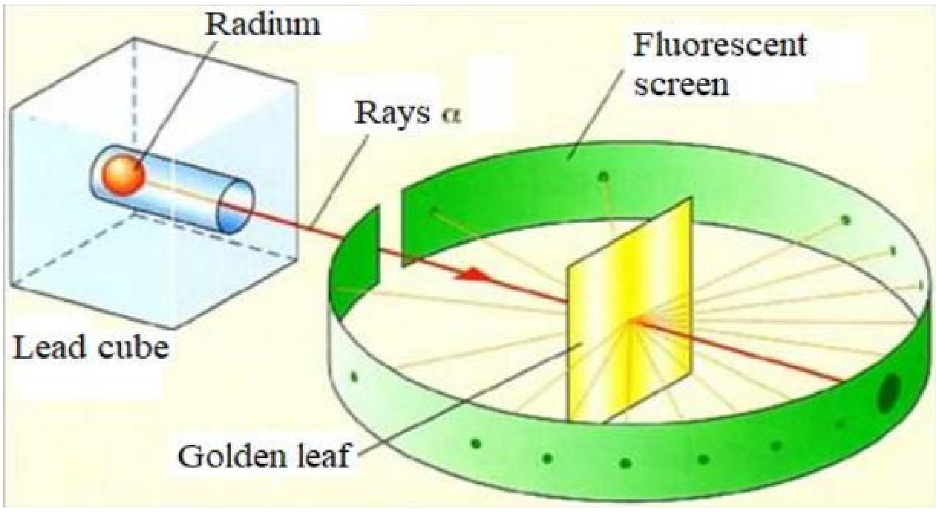
$$\left.\begin{array}{l} \text{1 mole of } {}_{6}^{12}C \text{ atoms} \longrightarrow \mathbf{N_A} \text{ of } {}_{6}^{12}C \text{ atoms} \longrightarrow 12g \\ \text{1 atom of } {}_{6}^{12}C \longrightarrow m_{atom} \end{array}\right\} \Longrightarrow m_{atom} = 12/N_A$$

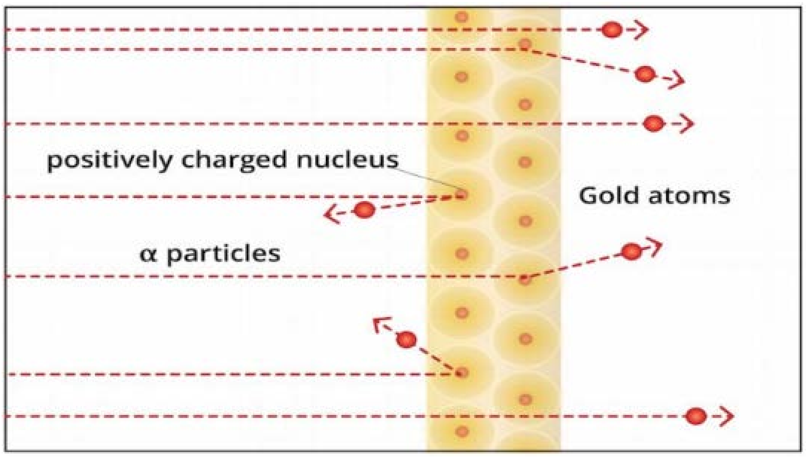
$1\ a.m.u = (1/12)*(12/N_A) = 1/N_A = 1/\ 6.023.\ 10^{23}\ (g)$

$1\ a.m.u = 1.66030.10^{-24}\ g = 1.66030.10^{-27}\ Kg$

## <u>5. Rutherford's model :(Rutherford's experiment (1911))</u>

Rutherford took a thin gold foil and made alpha particles ($He^{2+}$) fall on it. Behind the foil sat a fluorescent screen for which he could observe the alpha particles impact.




**<u>Observations:</u>**

*Most of the alpha particles passed through the gold foil without getting reflected.

*A few alpha particles (1 in 100) got reflected after hitting the gold foil.

*1 in 100000 alpha particles completely rebound (reflected by an angle of 180°) after hitting the gold foil.

**Conclusions**

• Since most rays passed through, Rutherford concluded that most of the space inside the atom is empty.

• A few rays got reflected because of the repulsion of its positive charge with some other positive charges inside the atom.

• 1/100000th of the rays got strongly reflected because of a very strong positive charge in the center of the atom. He called this strong positive charge "nucleus".

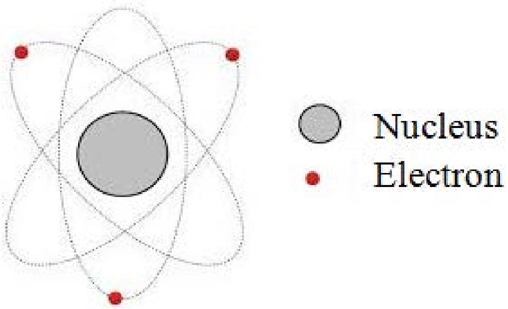
**Rutherford's structure of Atom**

Based on the above observations and conclusions, Rutherford proposed his own atomic structure, which is as follows.

• The majority of the atom's mass is concentrated in small positively charged center called **nucleus**.

• Electrons revolve around the nucleus in a circular orbit, similar to the way planets orbit the sun.

• The atomic structure is spherical.




## 5. Symbolic writing of atom

The atom can be represented as follows: $^{A}_{Z}X$

* **Z** : **Atomic Number** = Number of protons.

* **A** : **Mass Number** = Number of nucleons = Number of protons + Number of neutrons **(Atomic Molar Mass (g/mol))**.

**Example :** $^{16}_{8}O$

Z = 8 , A = 16

$A = Z + N \longrightarrow N = A - Z \longrightarrow N=16-8 = 8$

This atom contains: Z=8 electrons, Z=8 protons, N=8 neutrons

**Note**

An atom is normally **electrically neutral**, so it has as many **electrons** as the nucleus has **protons**.

***Ion :** Ion is formed from an atom by the loss or gain of one or more electrons.

**Anion:** is an atom that has gained one or more electrons, and it carries a negative charge.

**Cation:** is an atom that has lost one or more electrons, and it carries a positive charge.

**Examples :**

| Ion | protons | neutrons | electrons |
|---|---|---|---|
| $^{63}_{29}Cu^{2+}$ | 29 | 34 | 27 |
| $^{35}_{17}Cl^{-}$ | 17 | 18 | 18 |

## 6. Isotopes

### 6.1. Definition

Isotopes of the same element are atoms that have the same atomic number (Z) but different mass numbers (A). They differ only in the number of neutrons.

Examples :

Isotopes of $_1H$ : $^1_1H$, $^2_1H$, $^3_1H$ ; Isotopes of $_{17}Cl$ : $^{35}_{17}Cl$, $^{37}_{17}Cl$

### 6.2. Isotopic abundance (relative abundance) :

Is the mass percentage of isotope « i » in the natural element, defined as follows: $\mathbf{\Sigma x_i = 100\%}$.

Example : $^{35}_{17}Cl$ (75.77%) and $^{37}_{17}Cl$ (24.23%)

### 6.3. Average atomic mass of isotopes

Since an element is made up of a mixture of various isotopes with constant proportions (abundances), we can define an average atomic mass for each element.

$$\mathbf{M_{average} = \frac{\sum m_i x_i}{100}}$$

$\mathbf{m_i}$ : mass of the isotope « i », ($\mathbf{m_i \approx A_i}$ $\mathbf{A_i}$: Mass number of the isotope « i »).

$\mathbf{x_i}$ : Isotopic abundance of the isotope « i ».

Example : Calculate the average atomic mass of Chlorine.

| isotope | $^{35}Cl$ | $^{37}Cl$ |
|---|---|---|
| $x_i$ (%) | 75.77 | 24.23 |
| $m_i$ (amu) | 34.969 | 36.966 |

$\mathbf{M_{average} = ((34.969 \times 75.77) + (36.966 \times 24.23))/ 100 = 35.452\ amu}$

## 7. Nuclear binding energy

### A/ Definition

The binding energy is the energy needed to separate the nucleus of an atom into its nucleons (protons and neutrons).

### B/ Mass defect (mass loss)

The mass defect (Δm) is the difference between the total mass of the individual protons and neutrons before they combine to form a nucleus and the mass of the nucleus itself.

$\mathbf{\Delta m = [Z.m_p + (A\text{-}Z).m_n] - m_{nucleus}}$

### C/ Principle of mass-energy equivalence (Einstein's relationship)

The mass defect is related to the binding energy according to Einstein equation :

$\mathbf{\Delta E = \Delta m.\ C^2}$

C : Speed of light in the vaccum (C = 3. $10^8$ m/s)

**D/ Units of binding energy**

***Joule (J)**

***Electron-Volt (eV) : 1 eV = 1,6. $10^{-19}$ J**

***Mega életron-Volt (MeV) : 1 MeV = $10^{6}$ eV ; 1 MeV = 1,6. $10^{-13}$ J**

**Example :** Calculate the mass defect and the binding energy of helium (${}_2^4He$), knowing that the mass of helium is 4.0026 amu.

***Mass defect**

$\Delta m = [Z.m_p + (A-Z).m_n] - m_{nucleus} = (2*1,0074 + 2*1,0086) - 4,0026$

$\Delta m$ = 0,0294 amu

$\Delta m = 0,0294*1,6605.10^{-27} = 4,98.10^{-29}$ Kg

***Binding energy**

$\Delta E = \Delta m.C^2 = 4,98.10^{-29}.(3.10^8)^2 = 4,39.10^{-12}$ joule

$\Delta E = 4,39.10^{-12}$ joule $= 27,45.10^6$ eV = 27,45 MeV

**E/ Energy equivalent of atomic mass unit**

$\Delta m = 1$ amu $= 1,66.10^{-27}$ Kg

$\Delta E = \Delta m.C^2 = 1,66.10^{-27}.(3.10^8)^2 = 1,49.10^{-10}$ J = **931.5 MeV**

The following formula can also be used :

**$\Delta E = \Delta m. 931,5$** $\Delta$**E** in **Mev**, and, $\Delta$**m** in **a.m.u**

## 8. Nuclear stability

### 8.1. Stability and binding energy per nucleon

The binding energy per nucleon is the ratio of the nucleus's binding energy to its mass number.

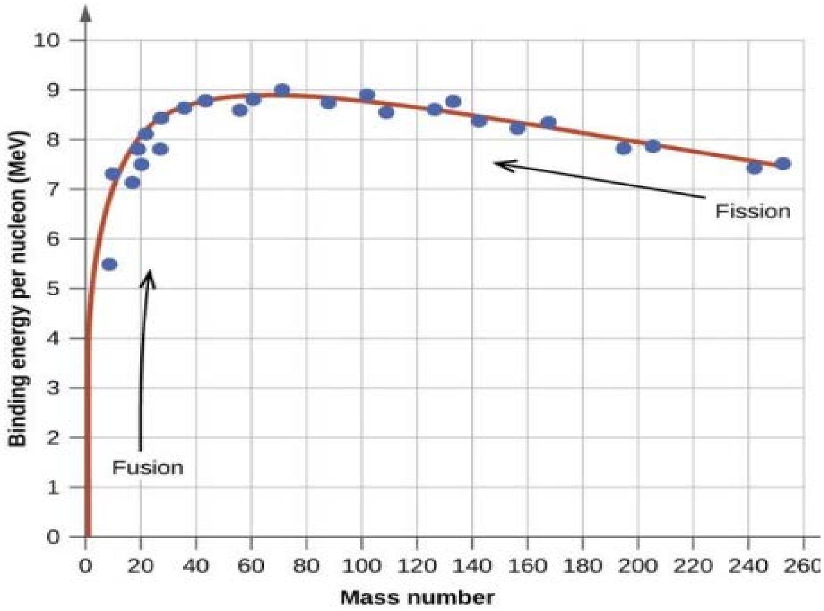
$$\frac{\Delta E}{A} = \frac{\Delta m.931,5}{A} \left(\frac{MeV}{nucleon}\right)$$

The binding energy per nucleon is a measure of the stability of an atomic nucleus. Nuclei with higher binding energies by nucleon are generally more stable, while those with lower binding energies per nucleon are less stable.

**Example :** ${}^{56}Fe$ : 8.79 MeV/nucleon ; ${}^{238}U$ : 7.57 MeV/nucleon.

Iron (${}^{56}Fe$) is more stable than uranium (${}^{238}U$).

**Aston Curve**

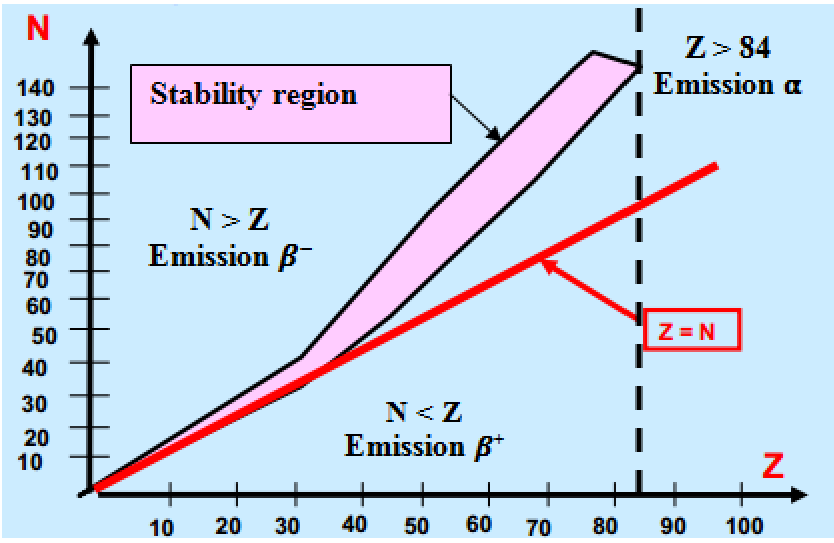



•For 50<A<80: the curve represents a maximum, which corresponds to the most stable nuclei. The ends of the curve correspond to unstable nuclei whose binding energy per nucleon is low ($\Delta E$ /A < 8 MeV/nucleon). These nuclei tend to stabilize through two different processes:

• **Fission**: heavy nuclei are split into lighter nuclei.

• **Fusion:** two light nuclei combine to form a more stable nucleus.

**8.2. Stability and Number of Nucleons**




* For **Z < 30**, **stable nuclei** with N=Z.

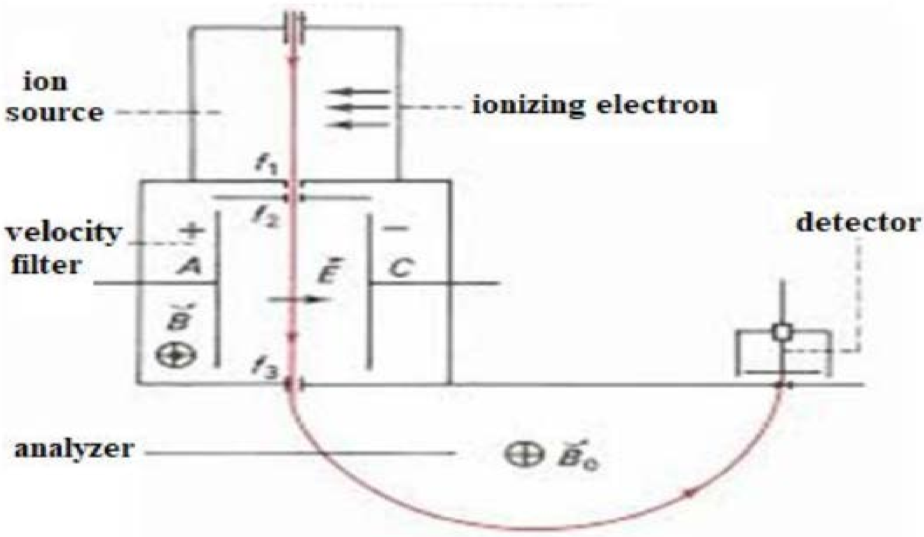
* For **Z > 30**, **stable nuclei** with N>Z are found in the stability region.

Therefore, unstable nuclei exist in three regions for **Z < 84** :

**<u>*Unstable nuclei with an excess of neutrons (N > Z) :</u>** a neutron transforms into a proton and a negaton ($^{0}_{-1}e$), which is called $\beta^-$ rays. ($^{1}_{0}n \rightarrow {}^{1}_{1}H + {}^{0}_{-1}e$)
**<u>* Unstable nuclei with an excess of protons (N < Z) :</u>** a proton transforms into a neutron and a positon ($^{0}_{+1}e$), which is called $\beta^+$ rays. ($^{1}_{1}H \rightarrow {}^{1}_{0}n + {}^{0}_{+1}e$)
**<u>* Heavy nuclei (Z > 84)</u>**
The nucleus has excess of neutrons and protons, so emission of an α particle (rays $_2^4He$) occurs.

## 9. Isotopes separation (Bainbridge Spectrometer)

Mass spectrometry allows the identification of different isotopes of an element by determining their masses and isotopic abundances. There are several types of spectrometers, among them: the Bainbridge spectrometer.




### A. Ionization Chamber

The gas atoms are ionized by electrons; the formed ions move at different speeds as they enter the velocity filter.

### B. Velocity Filter

Ions are subjected to the simultaneous action of electric field **E** and magnetic field **B**; the forces applied to an ion with charge (q) and velocity (v) are:
*The electric force (Fe): Fe = q.E
*The magnetic force (Fm): Fm = q.v.B
B and E are orthogonal; Fe and Fm have parallel directions and opposite senses.
**Fe = Fm → q.E = q.v.B → v = E/B**

### C. Analyzer

The ion of mass **m** is subjected to a constant magnetic field $\mathbf{B_0}$, and it deviates in a circular path of radius: **r**.
**Fm = Fc →** $\mathbf{q.v.B_0 = m.v^2/r \rightarrow m/r = q.B_0/v}$
**v = E/B ;** <u>So:</u> $\mathbf{m/r = q.B.B_0/E}$ **= constante**

**D. Detector**

The ions are collected separately according to their masses and their abundances.

**Example :** The separation of oxygen isotopes from the air yields the following results:

| Isotope | Mass (amu) | Abundance (%) |
|---|---|---|
| $^{16}O$ | 15.995 | 99.76 |
| $^{17}O$ | 16.999 | 00.04 |
| $^{18}O$ | 17.999 | 00.20 |