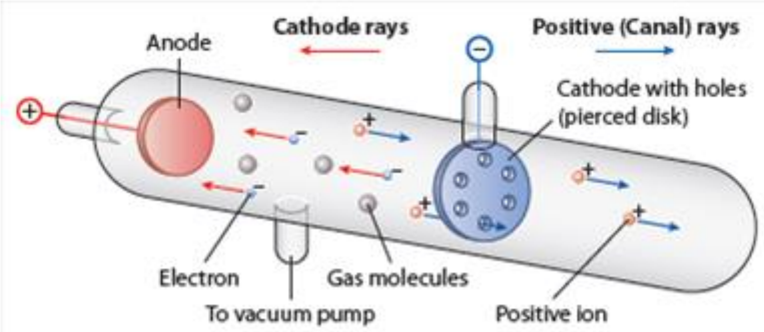
2- Goldstein's experiment 1886

The experiments conducted by the scientist Goldstein in 1886 led to the proposal that positively charged particles also exist in electrical discharge tubes.



"The same Crookes tube apparatus was used, but the orientation of the cathode and anode was reversed, with the addition of a small amount of hydrogen gas in the tube. When an electric current was passed through, it was observed that there was radiation on the screen, which can be explained as follows:

The emitted electrons from the cathode (-) are attracted towards the anode (+). Since there is hydrogen gas in the tube, the electrons collide with the neutral hydrogen atoms.

If the electrons have enough energy, they can dislodge other electrons away from the neutral atoms. As a result, when the negative electrons are repelled, positive particles remain in their place. Most of these positive particles are captured by electrons and become neutral, while a small portion of these positive particles passes through the holes into the region behind the cathode (hence they are called canal rays or positive rays), forming a bundle of particles that are influenced by the electric field, deflecting them towards the negatively charged plate, as well as being affected by the magnetic field.

The main properties of these rays (protons):

- They are attracted towards the negative pole, confirming that they are positively charged.
- They have mass and velocity, indicating that they are material particles and are usually heavier than electrons. Their mass depends on the type of gas present in the discharge tube.
- The e/m ratio for these particles was found to be much smaller, by thousands of times, compared to that calculated for negative particles, indicating that the mass of the proton is much larger than the mass of the electron.
- They are influenced by both electric and magnetic fields.

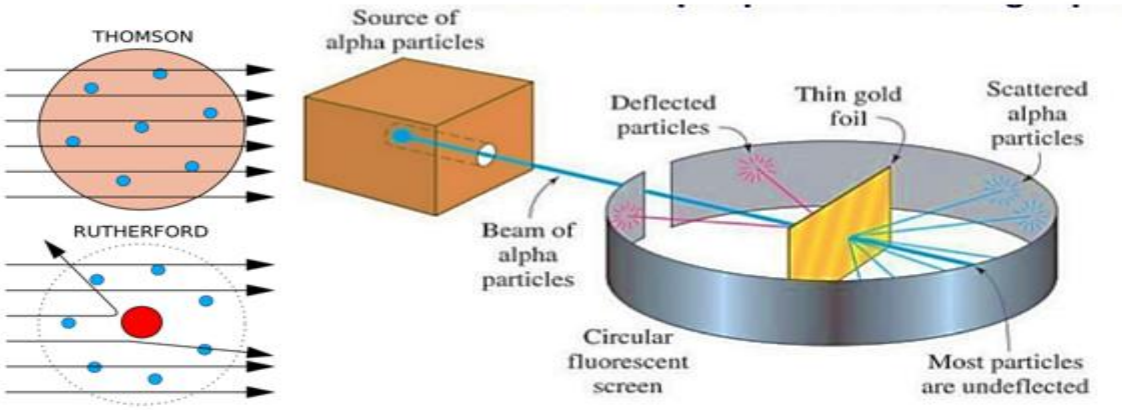
2- Proving the existence of the nucleus (The Rutherford Experiment 1909)

The Rutherford Experiment, also known as the Gold Foil Experiment or Geiger-Marsden Experiment (named after the scientists who conducted it under Rutherford's supervision in 1909), is an experiment that involved directing alpha particles (helium nuclei, He^{++}) at a thin gold foil (approximately 10^{-3} mm thick). It was expected that alpha particles, due to their high speed, would mostly pass through the gold foil with only slight deflections in their paths, caused by their positive charge interacting with the positive charge of the atoms.

The observations made were as follows:

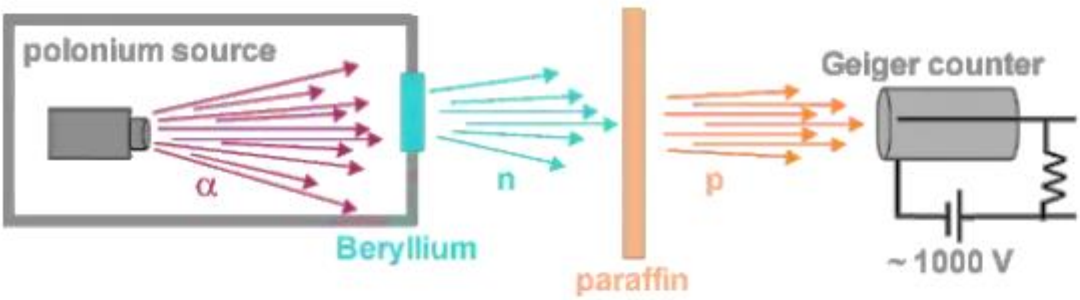
• Most of the alpha particles penetrated the gold foil, indicating that most of the atomic volume is empty space.

• Another portion rebounded from the foil, suggesting a direct collision with particles occupying a small fraction of the atomic volume and carrying a positive charge, presumed to be the atomic nucleus. • Some of them were deflected from their path, indicating that they passed near the positively charged nucleus and experienced repulsion.



3- Proof of the existence of neutrons (Chadwick Experiment 1932)

The prevailing belief was that the mass of an atom is roughly equal to the mass of the protons present in its nucleus.



However, mass spectrometry proved that the mass of the atom is about twice the mass of the nucleus obtained through calculations ($m_{nucleus} = Z.m_{proton}$). Therefore, it was assumed that there are neutral particles in the nucleus in addition to the positively charged protons.

The scientist Chadwick was able to experimentally prove the existence of neutrons by conducting an experiment in which he bombarded a thin slice of beryllium (^{9}Be) with alpha particles. This resulted in the generation of particles with high penetrating power. When these particles were directed onto a paraffin wax target, they caused it to emit high-speed protons.

Chadwick was able to demonstrate that these particles emitted from beryllium were uncharged (not affected by electric or magnetic fields) and had a mass approximately equal to that of a proton. He named these particles "neutrons.

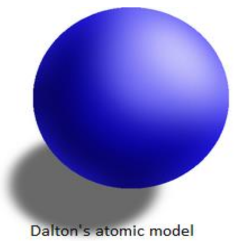
It's worth noting that the use of paraffin wax in the Chadwick experiment was to slow down the released neutrons. Neutrons were collided with hydrogen nuclei (protons) of equal mass, causing the neutrons to slow down and the protons to be released (elastic collision between two bodies of equal mass with one at rest).

A "Geiger counter" is an instrument used to detect "ionizing" radiations.

2- Proposed Atomic Models

Dalton's Atomic Model

In 1803, Dalton proposed an atomic model in which he assumed that matter is composed of tiny, indivisible particles called atoms. These atoms were envisioned as small, homogeneous, and indivisible spheres, resembling small glass marbles.



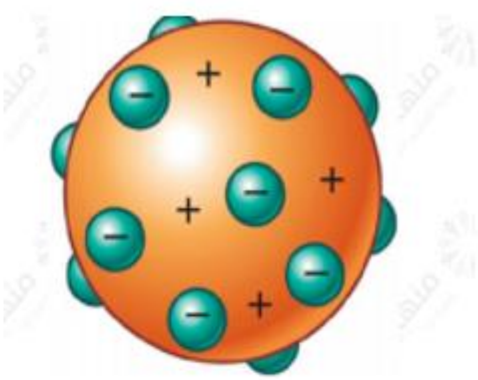
Dalton's atomic model

Drawbacks of Dalton's Model:

- Dalton's atomic model stated that atoms are indivisible, but it was later discovered that this is not true, and atoms are composed of smaller units known as electrons, protons, and neutrons.
- According to Dalton's theory, atoms of the same chemical element are identical in shape, mass, and chemical properties. However, it turned out that this assumption is also not true, as scientists discovered that atoms of the same chemical element can have different densities, weights, and properties, which are referred to as isotopes

Thomson's Model

Since atoms are electrically neutral, Thomson assumed that if atoms contain negatively charged electrons, they must also contain positively charged particles to make the atom neutral. Therefore, Thomson modified Dalton's atomic model. Instead of considering atoms as solid spheres, Thomson proposed that atoms are spheres filled with positive charges, with negatively charged electrons scattered within them.



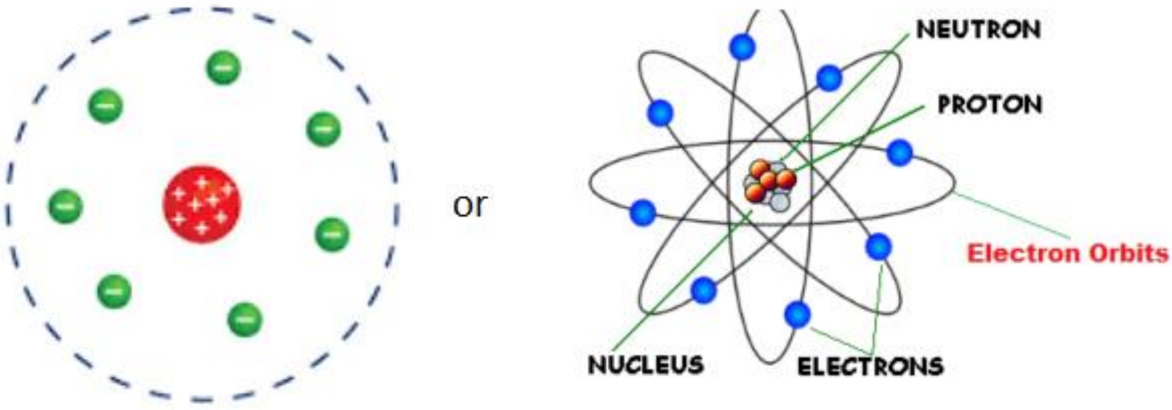
or

Drawbacks of Thomson's Model

Thomson's model had several shortcomings, including the idea that electrons are embedded in the atom, implying that they are stationary, which was incorrect. Additionally, Thomson's model failed to explain the scattering of alpha particles by thin metal foils.

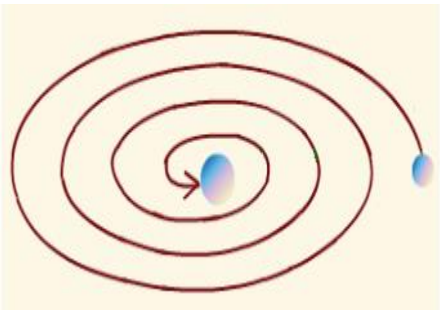
Rutherford's Model

Rutherford proposed a model of the atom after his famous experiment, in which he demonstrated that the positive charge and most of the mass of the atom are concentrated in a small central part called the nucleus. Electrons, which are negatively charged, orbit the nucleus like planets around the sun.



Shortcomings of Rutherford's Theory:

According to the physics laws at the time, any object in motion should lose energy. Since electrons were considered to be in constant circular motion around the nucleus in Rutherford's model, this motion would result in the continuous loss of energy. As the electron lost energy, its speed would decrease, and it would gradually spiral inward toward the nucleus until it eventually merged with it. This prediction did not align with observed behavior.



This classical physics concept did not match the actual behavior of atoms. For example, the hydrogen atom is stable because its electrons remain in stable orbits and do not continuously lose energy or spiral into the nucleus. Electrons only emit radiation when excited, not in their natural state.

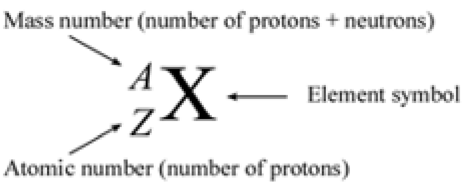
3- Summary of the experiments

Atoms are made up of extremely small subatomic particles called protons, neutrons, and electrons. **Protons** are positively charged particles with a relative mass of 1.673×10^{-27}Kg, which form part of the core **nucleus** of an atom. The other part of the atomic nucleus is made up of **neutrons,** electrically neutral particles with a relative mass almost identical to a proton (1.675×10^{-27}Kg). **Electrons** are extremely small (9.109×10^{-31}Kg) negatively charged particles that form an electron cloud, which orbits the nucleus.

Particle	Symbol	Mass (kg)	Relative Mass (pro-ton =1)	Relative Charge
Proton	p^+	1.673×10^{-27}	1	+1
Neutron	n^0	1.675×10^{-27}	1	0
Electron	e^-	9.109×10^{-31}	0.00055	-1

Elements are represented by a chemical symbol, with the atomic number and mass number sometimes affixed as indicated below. The mass number is the sum of the numbers of neutrons and protons in the nucleus.

$$^{Z+N}_{Z}X$$



-The charge of the nucleus is: $Q=+Ze$

-The charge of electrons is: $Q'=-Ze$

The mass of the atom is:

$$m = Z.m_p + (A-Z)m_n + Z.m_e$$

The mass of an electron is so small compared to the mass of a proton and a neutron, that it is generally neglected when calculating the atomic mass of an element.

$$m = Z.m_p + (A-Z)m_n \Leftarrow (m_e \lll m_p)$$

4-Isotopes

Are different atoms of the same element that have the same atomic number but different mass numbers. For example, oxygen isotopes $^{18}_{8}O$; $^{16}_{8}O$; $^{15}_{8}O$.

The chemical behavior of isotopes is similar, primarily due to their identical atomic numbers, which represent the number of electrons. It's well-known that chemical reactions involve electrons and are largely influenced by electron behavior, rather than the nucleus.

However, the physical properties of isotopes of the same element are inherently different, such as boiling points, melting points, and densities. This difference is attributed to variations in their nuclear masses.

Isotopic abundance

Is the percentage of occurrence of isotopes of a single element in nature.

Example: Natural uranium consists of three primary isotopes with the following abundances:

- $^{238}_{92}U$ in ratio (99.2745%)
- $^{235}_{92}U$ in ratio (0.72%)
- $^{234}_{92}U$ in ratio (0.0055%)

To calculate the average atomic mass of the isotopes, which is listed in the periodic table, we use the following relationship:

$$\bar{M}_x = \sum \frac{Y\% \times M_x}{100}$$

Where:

X: Chemical element

Y%: Isotopic abundance

$\bar{M}_x$: Average atomic mass of the element

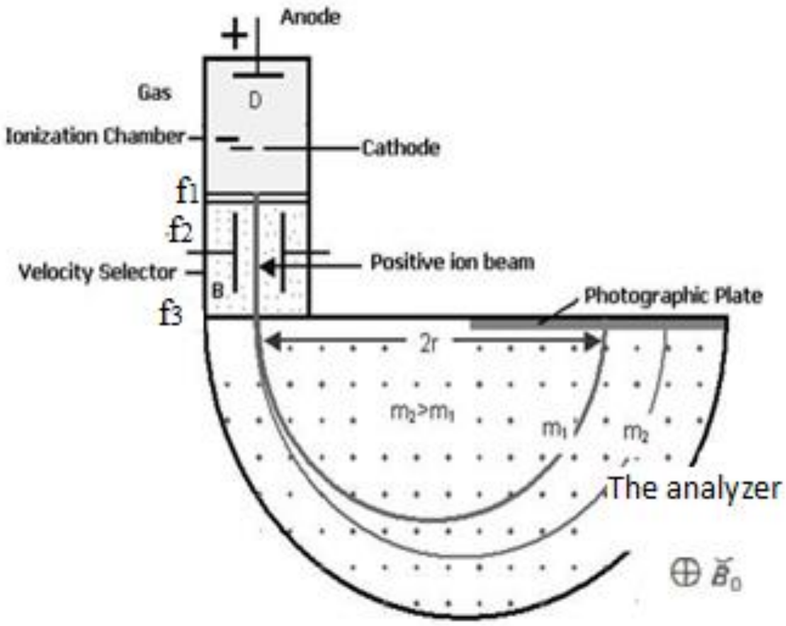
Determining atomic masses

To determine atomic masses, we measure the q/m ratio, where q stands for the atomic charge and m represents its mass. The instruments used for mass measurement are called mass spectrometers, with one of the most renowned ones being the "Bainbridge Mass Spectrometer," also referred to as the "Velocity Filter" (it's an analytical technique for identifying the elemental composition of a substance or molecule).

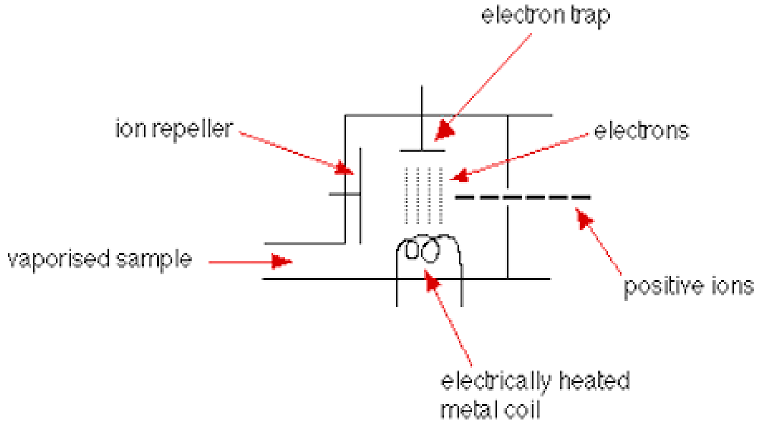
The Bainbridge Mass Spectrometer, 1933

The Bainbridge Mass Spectrometer, model, consists of the following elements:

1. Deflection Source
2. Velocity Selector
3. Analyzer and Detector



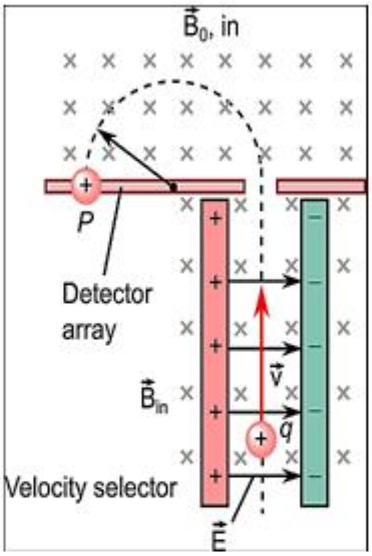
The ionization chamber

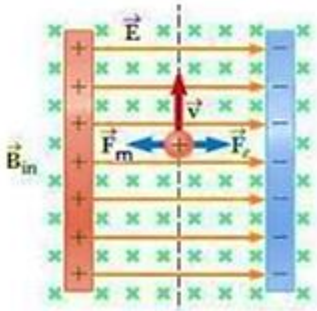


We send a stream of gas into the ionization chamber, where gas particles collide with a stream of high-speed electrons emitted from a heated wire. As a result, the gas particles ionize into positively charged ions.

The velocity Selector

Is composed of a capacitor and a magnet. When ions enter it, they are influenced by the electric field $\vec{E}$, causing a change in their path. By applying a magnetic field $\vec{B}$ perpendicular to it and of the same intensity, the deviation is corrected, allowing the ions to pass through hole f_3 to the analyzer.





$$F_e = F_m$$

$$q.E = q.\vartheta.B$$

$$\Rightarrow \vartheta = \frac{E}{B} \dots \dots 1$$

The analyzer

Ions enter through opening f_3 and are subjected to a constant magnetic field $\overrightarrow{B_0}$, which is perpendicular to the motion of the ions. As a result, the ions deviate from their straight path to trace a circular arc with a radius of R. The magnetic force equals the centripetal force.

$$F_c = F_m$$

$$\frac{m.\vartheta^2}{R} = q.\vartheta.B_0$$

$$\Rightarrow \frac{q}{m} = \frac{\vartheta}{R.B_0} \ldots\ldots 2$$

From (1) and (2), we find:

$$\frac{q}{m} = \frac{E}{R.B.B_0}$$

q: The charge of the ion (Coulombs)

m: The mass of the ion (Kilograms)

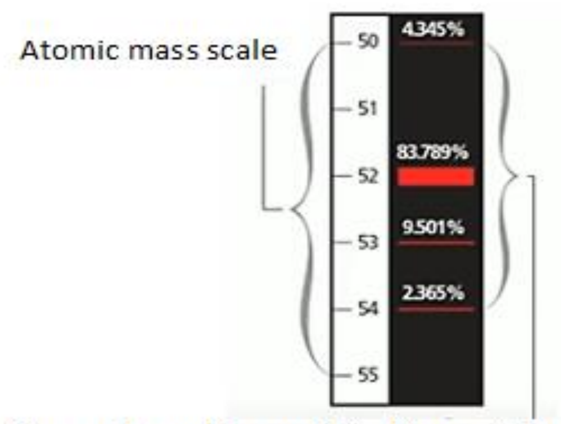
R: The radius of the ion's circular path (meters)

E: The electric field intensity (Volts per meter)

B_0, B: The magnetic field intensity (Tesla)

The detector

It is a photographic plate that displays spots where ions make contact. Isotopes collide with the detector at different positions, depending on their mass. The collision positions are influenced by their mass, with lighter ions producing smaller circular path radii. The abundance of each isotope is determined by the width of the mark it leaves on the detector.



The marks on the sensitive film and the
: percentage of isotope presence

Note:

In theory, with knowledge of E, R, B, B_0 from the device, and q, it is possible to calculate the mass (m) of the ion easily. However, it has been found experimentally that measuring m with acceptable precision using this method is challenging. Therefore, it is preferable to measure mass ratios by introducing the ion of interest alongside a reference ion like carbon ^{12}C.

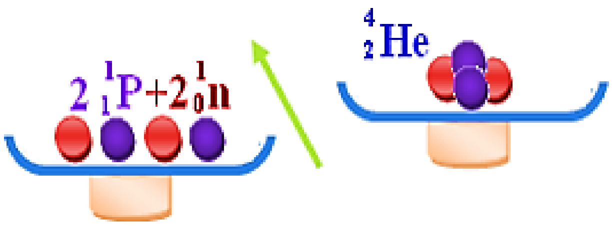
$$\begin{cases} R_1 = \dfrac{m_1 . \vartheta}{q . B_0} \\ R_c = \dfrac{m_c . \vartheta}{q . B_0} \end{cases}$$

$$\Rightarrow \frac{R_1}{R_c} = \frac{m_1}{m_c}$$

$$\Rightarrow m_1 = \frac{R_1 . m_c}{R_c}$$

4- Nuclear Binding Energy

Experimentally, it is observed that the mass of an atomic nucleus is always smaller than the sum of the masses of its constituent nuclei.



For example, the mass of the helium nucleus :

$$m({}^{4}_{2}He) = 6,6447 . 10^{-27} kg$$

As for the sum of the masses of its constituent nuclei, it equals :

$$2m_p + 2m_n = 2 \times 1,6726 . 10^{-27} + 2 \times 1,6750 . 10^{-27} = 6,6952 . 10^{-27} kg$$

Therefore, we conclude that the mass of the helium nucleus is smaller than the sum of the masses of its constituent nuclei.

The difference between them is called the mass deficiency (Δm) of the nucleus, and it is a positive quantity. Its expression is:

$$\Delta m = [Z m_P + (A - Z) m_N] - M({}^{A}_{Z}X)$$

where:

Δm = mass defect (amu)

m_p = mass of a proton

m_N = mass of a neutron

M = mass of nuclide

Z = atomic number (number of protons)

A = mass number (number of nucleons)

Indeed, studies have shown that the mass deficiency is converted into energy. This energy is what keeps the nucleus cohesive and stable, and it is known as nuclear binding energy.

Nuclear Binding Energy (Per Nucleus)

It is the energy required to separate the constituent nucleons (protons and neutrons) of a nucleus and leave them in a state of rest.

$$E_\ell = \Delta m.c^2$$

Where:

- Δm : is the mass loss in kilograms.
- E_ℓ: is the binding energy in joules.
- C: is the speed of light, c= $2.99792.10^8 \approx 3.10^8 m/s$

Conversely, when a nucleus forms from separate nucleons, energy is released, and this energy is represented by E_ℓ.

Example

what is the binding energy of the nucleus of iron $^{56}_{26}Fe$, given that:

m_p = 1.0072 uma ; m_N = 1.0086uma ; M_{exp} = 55.9375uma

- Energy is typically measured in the International System of Units (SI) using the Joule (J). However, in nuclear physics, it is often preferred to use the electron-volt (eV) and its multiples, where:

$$1MeV = 10^6 eV \quad و \quad 1eV = 1,602177.10^{-19} J$$

Nuclear Stability

We characterize nuclear stability by calculating the average binding energy, which is the energy required to remove one nucleon from the nucleus. It is represented as B:

$$\xi = \frac{E_l}{A}$$

Where:

ξ: The average binding energy per nucleon.

A: Mass number.

E_ℓ : Nuclear binding energy.

The greater the binding energy for a nucleus, the more stable the nucleus is.