Chapter II :

Main components of matter

- **Introduction :**

Dalton's Atomic Theory :

The English chemist **John Dalton** suggested that all matter is made up of atoms, which were indivisible and indestructible. He also stated that all the atoms of an element were exactly the same, but the atoms of different elements differ in size and mass.

Chemical reactions, according to Dalton's atomic theory, involve a rearrangement of atoms to form products. According to the postulates proposed by Dalton, the atomic structure comprises atoms, the smallest particle responsible for the chemical reactions to occur.

The following are the postulates of his theory:

- Every matter is made up of atoms.
- Atoms are indivisible.
- Specific elements have only one type of atom in them.
- Each atom has its own constant mass that varies from element to element.
- Atoms undergo rearrangement during a chemical reaction.
- Atoms can neither be created nor destroyed but can be transformed from one form to another.

Dalton's atomic theory successfully explained the Laws of chemical reactions, namely, the Law of conservation of mass, the Law of constant properties, the Law of multiple proportions and the Law of reciprocal proportions.

Demerits of Dalton's Atomic Theory :

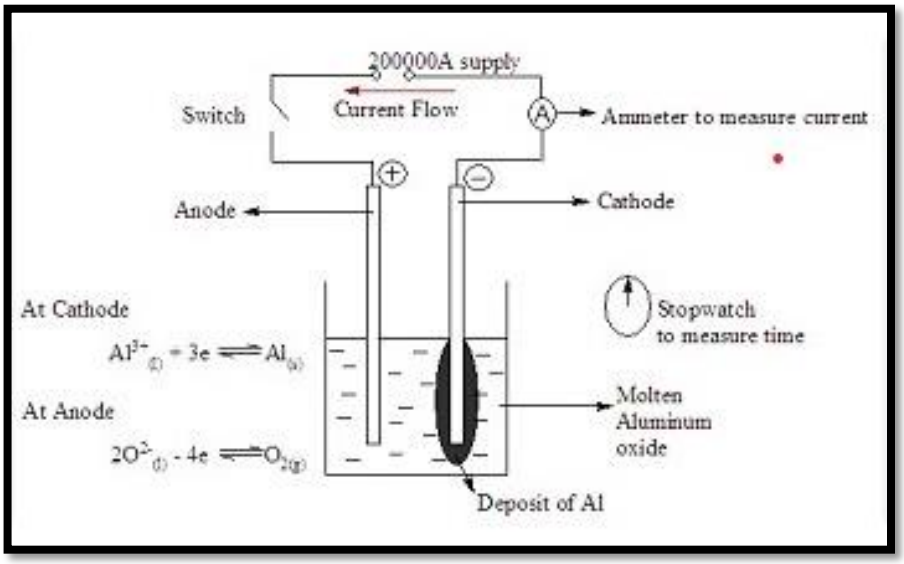
- The theory was unable to explain the existence of isotopes.
- Nothing about the structure of the atom was appropriately explained.
- Later, scientists discovered particles inside the atom that proved the atoms are divisible.

The discovery of particles inside atoms led to a better understanding of chemical species; these particles inside the atoms are called subatomic particles. The discovery of various subatomic particles is as follows:

1- Discovering Electron :

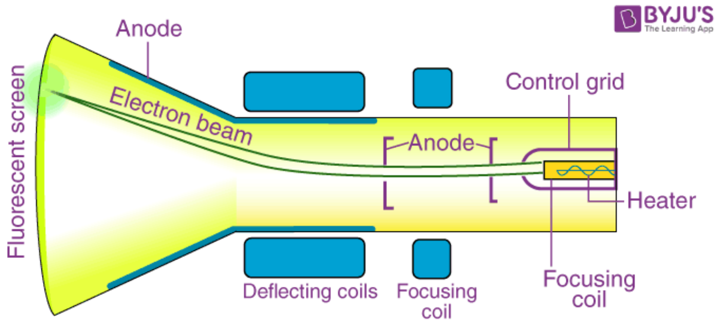
1-1- Electrical nature of the matter :

In the course of proving that electricities produced by various means are identical, Faraday discovered the two laws of electrolysis: the amount of chemical change or decomposition is exactly proportional to the quantity of electricity that passes in solution, and the amounts of different substances deposited or dissolved by the same quantity of electricity are proportional to their chemical equivalent weights. In 1833 he and the classicist William Whewell worked out a new nomenclature for electrochemical phenomena based on Greek words, which is more or less still in use today—ion, electrode, and so on.



1-2- Cathode Ray Experiment :

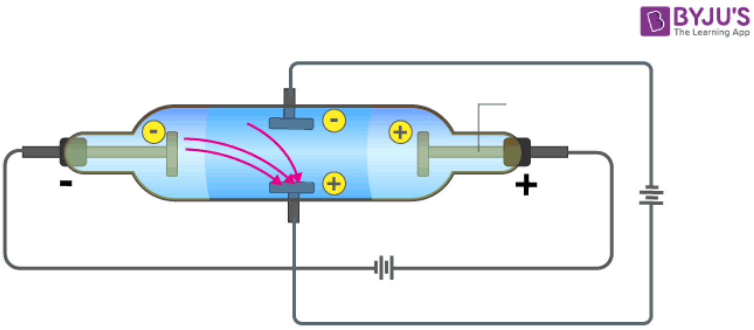
The first ideas about electrons came from experiments with cathode-ray tubes. A forerunner of neon signs, fluorescent lights, and TV picture tubes, a typical cathode-ray tube is a partially evacuated glass tube with a piece of metal sealed in each end. The pieces of metal are called electrodes; the one given a negative charge is called the cathode, and the one given a positive charge is called the anode.



Cathode Ray Tube Experiment

Observations:

- If 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 cathode rays will cause gases and fluorescent materials to glow, and will heat metal objects in its path to red heat, These rays were called “Cathode Rays”.
- When an external electric field is applied, the cathode rays get deflected towards the positive electrode, but in the absence of an electric field, they travel in a straight line.
- When rotor Blades are placed in the path of the cathode rays, they seem to rotate. This proves that the cathode rays are made up of particles of a certain mass so that they have some energy.

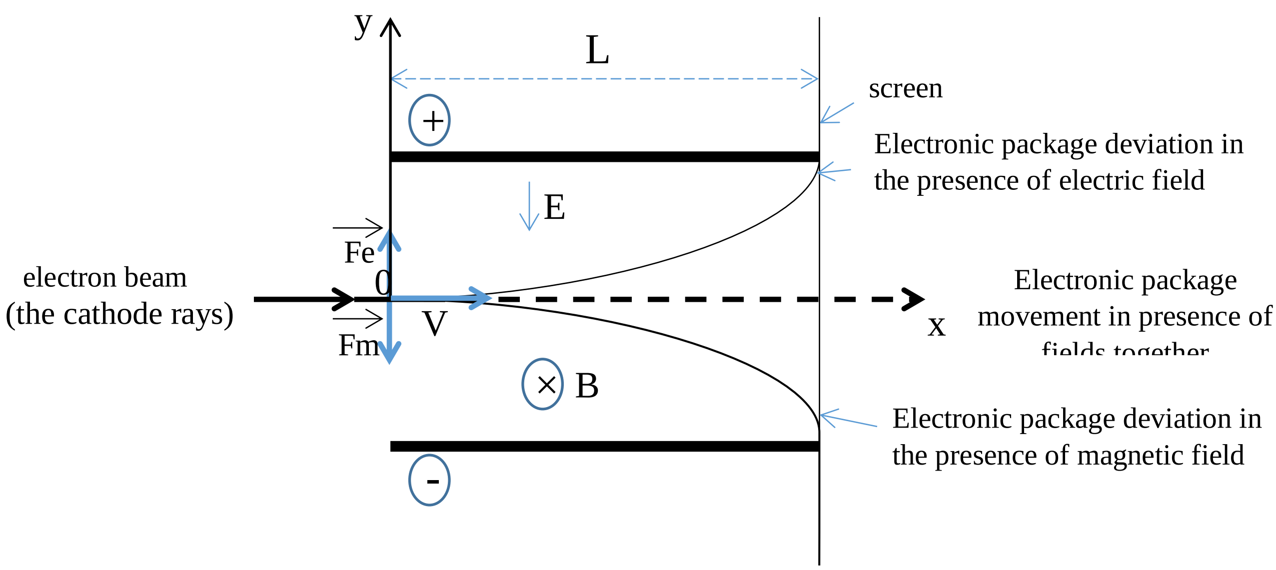


The direction of Cathode rays in the absence of an Electric Field

- With all this evidence, Thompson concluded that cathode rays are made of negatively charged particles called “electrons”.

I-3- J.J. Thomson experiment (determination of the e/m ratio) :

Thomson applied both electrical and magnetic fields to the cathode ray at the same time. Knowing the strength of the fields applied and measuring the deflection of the particle stream in the tube, he was first able to measure the velocity of the particles in the stream. Then, by measuring the deflection of the ray while varying the two fields, Thomson was able to measure the mass-to-charge ratio of the particles in the stream and he found something astonishing. The negative particles had a mass-to-charge ratio that was over 1,000 times lower than that of a hydrogen atom, suggesting that the particles were incredibly tiny – much smaller than the smallest atom known. This fact allowed Thomson to definitively say that the atom was not the fundamental building block of matter and that smaller (subatomic) particles existed. Thomson originally called these particles corpuscles, but later they became known as electrons.



- The electron pack veers towards the positive dyke due to electrical force

$\overrightarrow{Fe} = e\vec{E}$, This force is vertical on the direction of electrons speed and as is known from the basic relationship of your dynamic:

$$\sum \vec{F} = m\vec{\gamma}$$

Drop motion on the OX axis:

The force applied to the electron in the direction of the **ox** axis is non-existent:

$$F_x = 0 \Rightarrow \gamma_x = 0 \Rightarrow \frac{d^2x}{dt^2} = 0 \Rightarrow \frac{dx}{dt} = cte$$

$$\frac{dx}{dt} = v \Rightarrow dx = vdt$$

By integration we find:

$$x = vt + c$$

- At the moment of time $t = 0, \quad x = 0$, So:

$$x = vt \ldots \ldots \ldots . . (1)$$

It is the time equation of the movement of the electron towards the **OX** axis and indicates that the movement towards ox is regular.

Drop motion on the OY axis:

$$\|\vec{F}\| = m\gamma_y = eE \Rightarrow \gamma_y = \frac{eE}{m} \Rightarrow \frac{dv}{dt} = \frac{eE}{m}$$

- The first integration of acceleration γ_y gives the v_y speed:

$$\Rightarrow v = \frac{eE}{m}t + c$$

At the moment of time $t = 0, \quad y = 0$, So:

$$\Rightarrow v_y = \frac{eEt}{m} = \frac{dy}{dt}$$

- The second integration gives the distance y:

$$\int dy = \int \frac{eEt}{m} dt \Rightarrow y = \frac{1}{2}\frac{eE}{m}t^2 \ldots \ldots \ldots \ldots \ldots (2)$$

This is if we consider the principle of y distance to be the entry point of the electron in the electric field t=0⇒y=0

Equation (**2**) is the time equation of electron movement towards the oy axis, and it is a regular accelerated movement.

- **Movement equation :**

From equation (**1**) we conclude that $t = \frac{x}{v}$ with compensation in equation (**2**) we find:

$$y = \frac{1}{2}\frac{eE}{m}\frac{x^2}{v^2} \dots\dots\dots\dots\dots\dots..(3)$$

It is a parabole equation that gives y deviation to any distance x the electron travels under the electric field.

At the end of the electric field $x = L$, with compensation in equation 3 we find:

$$y = \frac{1}{2}\frac{eE}{m}\frac{L^2}{v^2} \dots\dots\dots\dots\dots.(4)$$

- From this relationship we note:

1- The more intense the electric field the more the deviation.

2- The longer my condensed plate gets the more deviant it gets.

3- The faster the electron the lower the deviation.

4- The deviation is expressly proportional to the charge and inversely to the mass.

This can be summarized as follows:

- The measured deviation fits into the constant related to the device (E, L) that we can easily measure.
- Speed can be measured by applying a suitable vertical magnetic field to the electric field so that:

If the intensity of both forces is equal $F_e = F_m$ the following can be written:

$$\mathrm{q.v.B} = \mathrm{q.E} \Rightarrow v = \frac{E}{B}$$

With compensation in equation (**4**) we find:

$$\frac{e}{m} = \frac{2y}{L^2}\frac{E}{B^2} \dots\dots\dots\dots\dots.(5)$$

And since E = U/d:

d : Distance between the two plates

U : Latency difference

$$\frac{e}{m} = \frac{2y}{L^2}\frac{U}{d}\frac{1}{B^2}$$

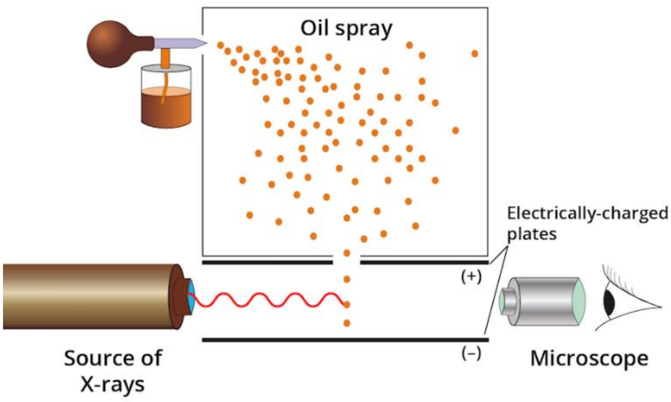
By conducting a numerical application of all values obtained experimentally we calculate the e/m ratio value where we find:

$$\boldsymbol{\frac{e}{m} = 1,759.10^{11}\ coulomb/Kg}$$

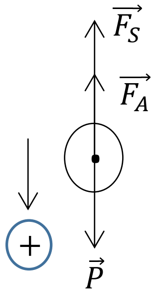
An important part of Thomson's experimentation was his use of twenty different metals for cathodes and of several gases to conduct the discharge. Every combination of metals and gases yielded the same charge-to-mass ratio for the cathode rays. This led to the belief that electrons are common to all of the metals used in the experiments, and probably to all atoms in general.

I-4- The Millikan Experiment :

Millikan measured the charge of electrons. The experiment consisted of measuring the electrical charge carried by tiny drops of oil that are suspended in an electrical field. By means of an atomizer, oil droplets were sprayed into the test chamber. As the droplets settled slowly through the air, high-energy X rays were passed through the chamber to charge the droplets negatively (the X rays caused air molecules to give up electrons to the oil). By using a beam of light and a small telescope, Millikan could study the motion of a single droplet. When the electrical charge on the plates was increased enough to balance the effect of gravity, a droplet could be suspended motionless. At this point, the gravitational force would equal the electrical force. Measurements made in the motionless state, when inserted into equations for the forces acting on the droplet, enabled Millikan to calculate the charge carried by the droplet.



- Until we understand exactly what's happening, let's follow the movement of one drop of oil.
- The drop drops rapidly with the effect of terrestrial gravity, but also quickly ceases to fall because air resistance increases by increasing speed, reaching its limit speed, becoming suspended (obviously the limit speed will vary from drop to drop depending on its mass).



There are three main forces affecting the drop of oil, the force of gravity downwards (the force of gravity), the force of Stokes (the force of air resistance), and the driving force of Archimedes (often is neglected).

- Since the oil drop has a marginal speed, the acceleration is non-existent:

$$F_S + F_A = P$$

$$F_S = 6\pi r\tau v, \quad F_A = \frac{4}{3}\pi r^3 \rho_{air} g, \quad P = mg = \frac{4}{3}\pi r^3 \rho_h g$$

F_S :The power of Stokes

F_A: The driving force of Archimedes.

ρ_h :Volume mass of oil.

m :Oil droplet mass.

g: gravitational acceleration

ρ_{air} :Volumetric mass of air.

v :Air speed.

τ :Air viscosity coefficient.

r :Radius of the droplet.

$$F_S + F_A = P \Rightarrow 6\pi r\tau v_0 + \frac{4}{3}\pi r^3 \rho_{air} g = \frac{4}{3}\pi r^3 \rho_h g$$

$$\Rightarrow \boldsymbol{v_0 = \frac{2r^2 g(\rho_h - \rho_{air})}{9\tau}}$$

Oil particles are of course equivalent, so Millikan exposed the room to X-rays.

• X-ray of air lost some of its electrons, which motivated the neutral particles to acquire them, becoming negative of the charge.

• After charging the room and having voltage teams, an electric field will be generated that affects the oil drops, the direction of the electric field from the top to the bottom, and because the oil drops are negative, the "opposite" of the field: from the bottom to the top.

In this case there are four main forces affecting the drop of oil which is the force of gravity, the electrical force to the top, the force of Stokes, the thrust of archimides.

$$F_e = qE.$$

F_e: electric force

q: electric charge

E :electric field

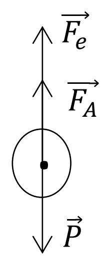
$$F_e + F_A = P + F_S \Rightarrow qE + \frac{4}{3}\pi r^3 \rho_{air} g = \frac{4}{3}\pi r^3 \rho_h g + 6\pi r \tau v$$

$$\Rightarrow q = \frac{\frac{4}{3}\pi r^3 g(\rho_h - \rho_{air}) + 6\pi r \tau v}{E}$$

- The drop of oil will move upwards until the gravitational force is equal to the force the electric field affects, and here it will stop momentarily and remain stationary. In this case three forces affect the drop of oil (the force of Stokes here is non-existent) which is, the force of gravity, the electric force upwards and the thrust of archimides.

$$F_e + F_A = P \Rightarrow qE' + \frac{4}{3}\pi r^3 \rho_{air} g = \frac{4}{3}\pi r^3 \rho_h g$$

$$\Rightarrow q = \frac{\frac{4}{3}\pi r^3 g(\rho_h - \rho_{air})}{E'}$$



Some oil droplets are larger than others, which means they have the ability to capture more than one electron, which means they need a "stronger" electric field to stop them moving.

Thus, Millikan changed the amount of electric field to stop the oil particles, calculating the charge in each case based on the previous law.

Also, Millikan changed the power of X-ray ionizing air, which means changing the number of electrons he would catch.

Millikan and his team studied the movement, thousands of shipped oil drops, repeating all the steps from calculating the mass of each drop and calculating its boundary speed, and then changing the field and calculating its intensity to arrive at the amount of charge carried by each drop of oil.

Millikan found different amounts of negative charge on different drops, but the charge measured each time was always a whole-number multiple of a very small basic unit of charge. The largest common divisor of all charges measured by this experiment was 1.60 x 10^{-19} coulomb (the coulomb is a charge unit). Millikan assumed this to be the fundamental charge, which is the charge on the electron.

- **mass of the electron:**

An electron is a micro-minute with a negative charge, and Millikan defined its charge as a racial charge and equal to:

$$|e| = 1{,}602.10^{-19}C$$

With a good estimate of the charge on an electron and the ratio of charge-to-mass as determined by Thomson

$$\frac{e}{m} = 1{,}759.10^{11}\ coulomb/Kg$$

m is the mass of the electron and e charge it, we can calculate the mass of the electron as follows:

$$e\Big/\frac{e}{m} = m = \frac{1{,}602.10^{-19}C}{1{,}759.10^{8}C/g} = 9{,}108.10^{-28}g$$

$$\boldsymbol{m_e = 9{,}108.10^{-28}g = 9{,}1.10^{-31}Kg}$$

II- Discovery of the Proton

- **Rutherford Discovering Proton**

By 1918, scientists had begun to isolate the nuclei of atoms and the smallest nuclei they had found was the hydrogen nucleus. Hydrogen nuclei were observed by shooting alpha particles at a tube full of hydrogen and occasionally the alpha particle would knock a hydrogen nucleus out of the atom. Those nuclei hit detectors that gave of twinkles of light when they were hit.
The way that Rutherford discovered the proton was in the following indirect way. He was shooting alpha particles through air, which is mostly nitrogen, with some oxygen and traces of other gasses. When he did this, he occasionally would see what looked like twinkles that occurred when the nuclei of hydrogen atoms hit his detector. The twinkles of the light were very clear.

- **Rutherford's Experiments**

Now, it could have been that what he was seeing was the alpha particles hitting hydrogen in the air, or maybe water molecules, which we know are made of two hydrogen atoms and one oxygen. So, he had to figure that all out.
He did so by repeating the experiment, this time with vials of pure nitrogen gas, pure oxygen gas, etc. What he found out was that it was the nitrogen that was the source of the hydrogen nuclei signal. He didn't see it very much at all with oxygen.

- **Findings from the Experiment**

So what he had done was use alpha particles to knock protons out of nitrogen nuclei, but he didn't know that of course. All he knew was that he could make hydrogen nuclei by hitting nitrogen nuclei and, of course, that's a bit puzzling. It took him a long time to figure it out.
Depending on how one reads the history in detail, it took between a year and six years for him to work through his thoughts. He saw the first creation of hydrogen nuclei in 1918 and it took until 1919 or perhaps until 1925 before he and others had put the entire picture together.
They realized that nitrogen nuclei must contain hydrogen nuclei and that they could be knocked out. Rutherford then coined the term proton, from the Greek word *protos*, which means *first.* The number of known constituents of the atom was now two. But there was a problem.

Properties of Proton

Some of the basic properties of Protons can be given as

- The mass of the proton is 1.6726×10^{-27}kg
- It is around 1836 time the mass of the electron

- Charge of Proton +1
- The radius of the proton is approximately 1.11×10^{-15}m

- **Problem of the 'Missing' Mass**

Scientists had figured out the mass of the proton and the electron and they knew from the fact that the atom was electrically neutral that they had to come in equal quantities. The problem was that the masses didn't add up.

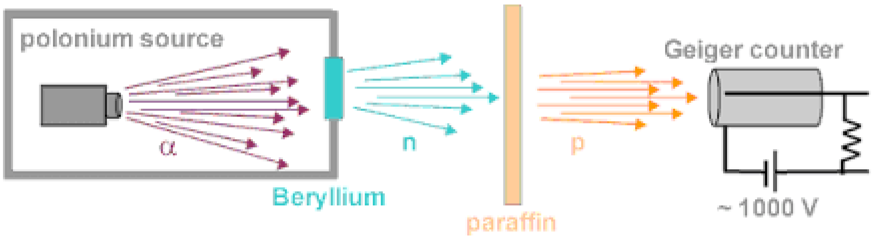
There were atoms with a known number of electrons, and therefore protons, and the mass of the atom was too big. There had to be another source of mass in the atoms and that source had to be electrically neutral. This is what we call the neutron these days, but they didn't know that back then.

III- Discovery of the Neutron :

Since the time of **Rutherford**, it had been known that the atomic mass number A of nuclei is a bit more than twice the atomic number Z for most atoms and that essentially all the mass of the atom is concentrated in the relatively tiny nucleus. In 1920 Rutherford proposed the existence of the third neutral particle in an atom. But up to 1930 proton-electron hypothesis was accepted. An experimental breakthrough came in 1930 with the observation by the German nuclear physicist Herbert Becker and Walther Bothe that bombardment of beryllium with alpha particles from a radioactive source produced neutral radiation which was penetrating but non-ionizing. They observed that the penetrating radiation was unaffected by electric fields and hence, they assumed it to be gamma radiation. In the year 1932, Frederic Joliot-Curie and Irene Joliot-Curie demonstrated that these rays have the potential to eject protons when it strikes paraffin or any H-containing compounds. The experiment proved that the assumption that the rays to be gamma rays was wrong. Because a photon that does not have mass cannot be capable to release a particle 1836 times heavier than an electron (protons). Therefore, it was concluded that the ejected rays cannot be photons.

- **Chadwick's Experiment:**

Neutron was discovered by Sir James Chadwick in 1932. He performed the same experiment performed by Frederic Joliot-Curie and Irene Joliot-Curie and used different bombardment targets other than paraffin.



He fired alpha radiation at the beryllium sheet from a polonium source. This led to the production of an uncharged, penetrating radiation. These radiations were made incident on paraffin wax, having relatively high hydrogen content. The range of the liberated protons was measured and the interaction between the uncharged radiation and the atoms of several gases was studied by Chadwick. The particle ejected was found to have a mass equal to that at proton and no charge. He called these particles as neutrons.

$$\mathbf{{}_4Be^9 + {}_2\alpha^4 \longrightarrow [{}_6C^{13}] \longrightarrow {}_6C^{12} + {}_0n^1}$$

Properties of Neutrons

Some of the basic properties of Neutron can be given as

- The mass of the electron is 1.6750×10^{-27} kg
- Charge of Neutron 0
- The radius of the neutron is approximately 1.11×10^{-15}m

Completing the Basic Structure of an Atom

With the discovery of the neutron, our understanding of basic structure of the atom was complete. Protons and neutrons are held together tightly in a small ball at the center of the atom called the nucleus. Surrounding the nucleus is a much larger volume—a trillion times larger in fact—where the electrons can be found.

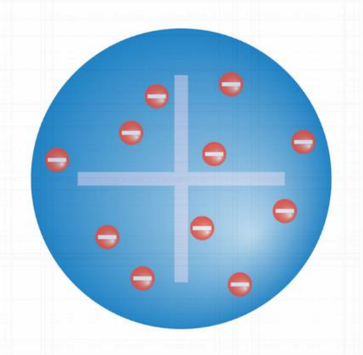
With that model, scientists can work out the details of chemistry and nuclear physics. This very impressive achievement was all wrapped up by the early 1930s. Surely, scientists have made advances since then.

IV- Different models of atom structure :

IV-1- Thomson Atom Model (The ‘Plum Pudding’ model):

With the electron now discovered, Thomson went on to propose an entirely new model of the atom that was known as "The Plum Pudding Model." The model was so called since it mimicked the British desert of the same name that had dried fruit (primarily raisins not plums), dispersed in a body of suet and eggs that made a dough.

In his model Thomson proposed that the negatively charged electrons (analogous to the raisins) were randomly spread out among what he called "a sphere of uniform positive electrification" (analogous with the dough or body of the pudding).



Thomson's "plum pudding model" of the atom, showing a positively-charged sphere containing many negatively-charged electrons in a random arrangement.

- **Limitations of Thomson’s Atomic Structure:**

Thomson’s atomic model does not clearly explain the stability of an atom. Also, further discoveries of other subatomic particles couldn’t be placed inside his atomic model.

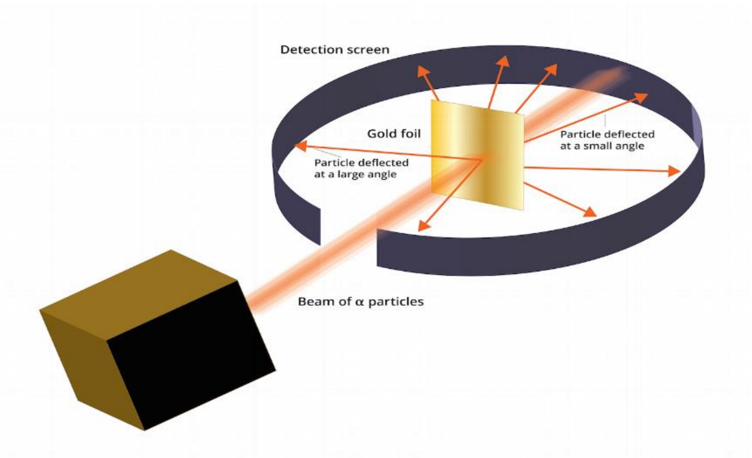
IV-2- Rutherford Atomic Theory

Rutherford, a student of J. J. Thomson, modified the atomic structure, where he did the next experiment.

Alpha Ray Scattering Experiment

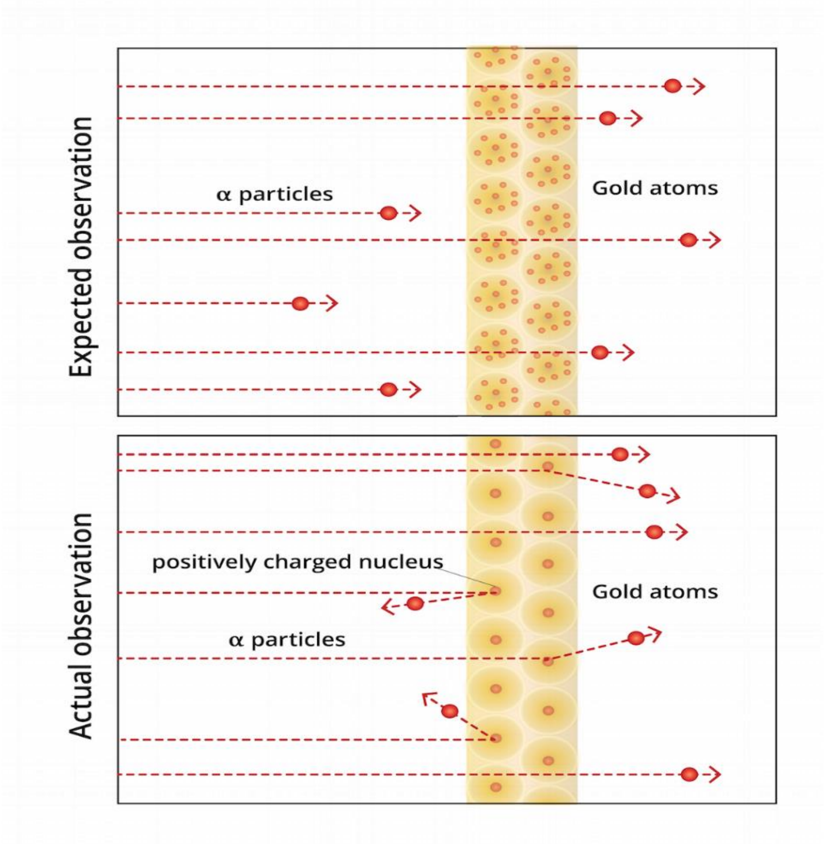
Construction:

- A very thin gold foil of 1000 atoms thick is taken.
- Alpha rays (doubly charged Helium He^{2+}) were made to bombard the gold foil.
- Zn S screen is placed behind the gold foil.



Observations:

- Most of the rays just went through the gold foil, making scintillations (bright spots) in the ZnS screen.
- A few rays got reflected after hitting the gold foil.
- One in 1000 rays got reflected by an angle of 180° (retraced path) 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 with some other positive charge inside the atom.
- 1/1000th of the rays got strongly deflected because of a very strong positive charge in the centre of the atom. He called this strong positive charge “nucleus”.
- He said most of the charge and mass of the atom resides in the nucleus.

Rutherford’s Structure of Atom

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

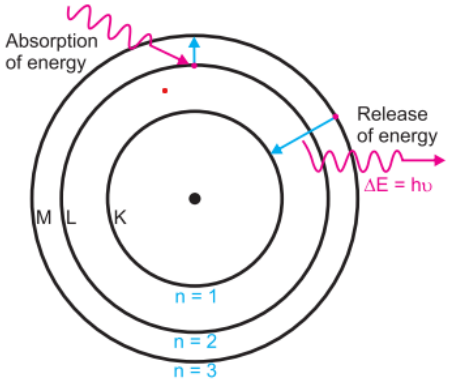
- The nucleus is at the centre of an atom, where most of the charge and mass is concentrated.
- The atomic structure is spherical.
- Electrons revolve around the nucleus in a circular orbit, similar to the way planets orbit the sun.

Limitations of the Rutherford Atomic Model

If electrons have to revolve around the nucleus, they will spend energy and that too against the strong force of attraction from the nucleus, a lot of energy will be spent by the electrons, and eventually, they will lose all their energy and will fall into the nucleus so the stability of atom is not explained.

IV-3- Bohr's Atomic Theory :

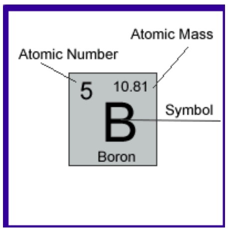
Keeping in view the defects in Rutherford's Atomic Model, Neil Bohr presented another model of atom in 1913. The Quantum Theory of Max Planck was used as foundation for this model. According to Bohr's model, revolving electron in an atom does not absorb or emit energy continuously. The energy of a revolving electron is 'quantized' as it revolves only in orbits of fixed energy, called 'energy levels' by him.



The Bohr's atomic model

V- Symbolic writing of atom :

- **Atomic Number:** Number of protons in nucleus
- **Mass Number (Atomic Mass):** Number of protons + neutrons (Units are g/mol)

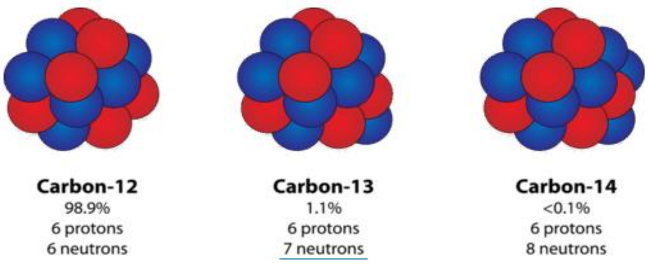


VI-Isotopes:

a) Definition :

Isotopes can be defined as the variants of chemical elements that possess the same number of protons and electrons, but a different number of neutrons. In other words, isotopes are variants of elements that differ in their nucleon (The total number of protons and neutrons) numbers due to a difference in the total number of neutrons in their respective nuclei.

Example:



- **Average mass of isotopes :**

The mass of an element shown in a periodic table or listed in a table of atomic masses is a weighted, average mass of all the isotopes present in a naturally occurring sample of that element. This is equal to the sum of each individual isotope's mass multiplied by its fractional abundance divided by 100.

$$\textbf{Average mass} = \sum_i (\textbf{fractional abundance} \times \textbf{isotopic mass}) / 100$$

For example, the element boron is composed of two isotopes : About 19.9% of all boron atoms are ^{10}B with a mass of 10.0129 amu, and the remaining 80.1% are ^{11}B with a mass of 11.0093amu The average atomic mass for boron is calculated to be :

$$\textbf{boron average mass} = (0.199\times10.0129) + (0.801\times11.0093) / 100 = 10.81\textbf{amu}$$

It is important to understand that no single boron atom weighs exactly 10.8 amu is the average mass of a large number of naturally occurring boron atoms. The individual boron atoms weigh either approximately 10 or 11 amu.

b) Isotope separation and identification of atomic blocks by mass spectrometry (Bainbridge):

The chemical properties of isotopes are identical and cannot be separated by classical chemical methods of this based on physical methods. These methods depend mainly on the different isotopes in the mass. Mass spectrometry is a device that allows:

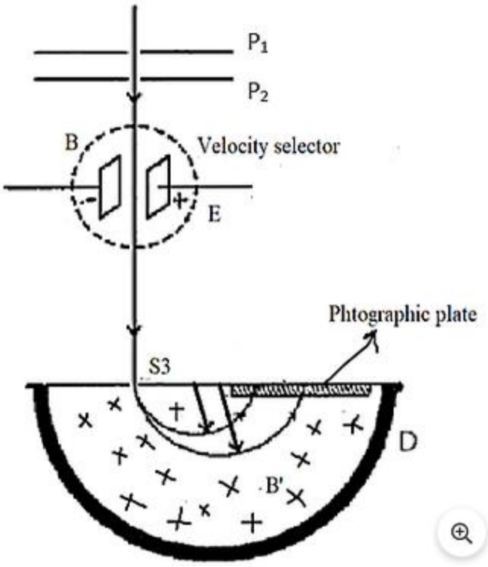
- Separation of various strands by the application of an electric field and a magnetic field
- Measuring the e/m ratio of a given ion
- Determination of isotope mixture ingredients in percentages

Among the types of spectometer:

- **Bainbridge Spectrometer :**

A device manifests its role in sorting ions with the same electric charge and different blocks using an electric field and a magnetic field, enabling the measurement of ion blocks with the same charge.

There are two types of fields used in Bainbridge mass spectrometer, electric and magnetic fields. Only ions having particular velocity are allowed to enter through the velocity selector. The force due to the magnetic field does not work, it only changes the direction of the moving particle.



An instrument called Bainbridge mass spectrometer is used for the determination of atomic masses. If one or more electrons are removed from the atom then it has a net positive charge and it becomes a positive ion. Using two narrow slits ***S*1** and ***S*2**, a beam of positive ions produced in a discharge tube is collimated into a fine beam. This fine beam then enters into a velocity selector. Only the ions of a particular velocity are allowed to come out from the velocity selector. The velocity selector consists of two plane parallel plates ***P*1** and ***P*2**, which produces a uniform electric field, it also has an electromagnet which produces a uniform magnetic field. These electric and magnetic fields are at right angles to each other and to the direction of the beam.

The electric field and magnetic field are so adjusted in such a way that the deflection produced by one field is nullified by the other so that the ions do not suffer any deflection within the velocity selector. Let E be the electric field intensity, B be the magnetic induction respectively and **q** be the charge on the positive ion.
The force exerted by the electric field is

$$F_E = qE$$

The force exerted by the magnetic field is :

$$F_B = qvB$$

Where v is the velocity of positive ions.

We have :

$$F_E = F_B$$

$$qE = qvB, \qquad v = E/B$$

The ions having this velocity are allowed to pass through the velocity selector. Then they pass through a slit $S3$ and then they enter in evacuated chamber D. The positive ions having the same velocity are again influenced by another strong uniform magnetic field of induction at right angles to the plane of the paper acting inwards. These ions then move in a circular path of radius R and strike the photographic plate. The centripetal force is provided by this magnetic field. Therefore,

$$mv^2/R = qvB', \qquad m = qB'R/v$$

Using $v = E/B$, **D** = **2R** we get :

$$m = BB'qR/E \Rightarrow \frac{q}{m} = \frac{2E}{B.B_0.D}$$

Ions having different masses trace different semi-circular paths and produce dark lines on the plate. The distance between the opening of the chamber and the position of the dark line is equal to the diameter **2*R*** from which radius **R** can be calculated.

By knowing ***B*,*B′*,*E*** and ***R***, the mass of the positive ions and isotopic masses can be calculated.

Note: In this method, we can distinguish the ions based on their mass as ions having different masses trace semicircles of different radii.

VI-Nuclear Binding Energy :

As a simple example of the energy associated with the strong nuclear force, consider the helium atom composed of two protons, two neutrons, and two electrons. The total mass of these six subatomic particles may be calculated as:

$$(2\times1.0073\text{amu})\text{protons}+(2\times1.0087\text{amu})\text{neutrons}+(2\times0.00055\text{amu})\text{electrons}=4.0331\text{amu}$$

However, mass spectrometric measurements reveal that the mass of an ${}_2^4He$ atom is 4.0026 amu, less than the combined masses of its six constituent subatomic particles. This difference between the calculated and experimentally measured masses is known as the mass defect of the atom. In the case of helium, the mass defect indicates a "loss" in mass of 4.0331 amu – 4.0026 amu = 0.0305 amu. The loss in mass accompanying the formation of an atom from protons, neutrons, and electrons is due to the conversion of that mass into energy that is evolved as the atom forms. The nuclear binding energy is the energy produced when the atoms' nucleons are bound together ; this is also the energy needed to break a nucleus into its constituent protons and neutrons. In comparison to chemical bond energies, nuclear binding energies are *vastly* greater, as we will learn in this section. Consequently, the energy changes associated with nuclear reactions are vastly greater than are those for chemical reactions.

The conversion between mass and energy is most identifiably represented by the mass-energy equivalence equation as stated by ***Albert Einstein***:

$$\Delta E = \Delta m C^2$$

where ΔE is energy, Δm is mass of the matter being converted, and C is the speed of light in a vacuum. This equation can be used to find the amount of energy that results when matter is converted into energy. Using this mass-energy equivalence equation, the nuclear binding energy of a nucleus may be calculated from its **mass defect**, as demonstrated in Example. A variety of units are commonly used for nuclear binding energies, including electron volts (eV), with 1 eV equaling the amount of energy necessary to the move the charge of an electron across an electric potential difference of 1 volt, making $1eV=1.602\times10^{-19}J$

.

Example: Calculation of Nuclear Binding Energy

Determine the binding energy for the nuclide ${}_2^4He$ in:

a. joules per mole of nuclei
b. joules per nucleus
c. MeV per nucleus

Solution :

The mass defect for a He
nucleus is 0.0305 amu, as shown previously. Determine the binding energy in joules per nuclide using the mass-energy equivalence equation. To accommodate the requested energy units, the mass defect must be expressed in kilograms (recall that 1 J = 1 kg m^2/s^2).

a) First, express the mass defect in g/mol. This is easily done considering the *numerical equivalence* of atomic mass (amu) and molar mass (g/mol) that results from the definitions of the amu and mole units (refer to the previous discussion in the chapter on atoms, molecules, and ions if needed). The mass defect is therefore 0.0305 g/mol. To accommodate the units of the other terms in the mass-energy equation, the mass must be expressed in kg, since 1 J = 1 kg m^2/s^2. Converting grams into kilograms yields a mass defect of 3.05×10–5kg/mol
. Substituting this quantity into the mass-energy equivalence equation yields:

b) The binding energy for a single nucleus is computed from the molar binding energy using Avogadro's number:

$$\Delta E = 2.74\times10^{12}\ Jmol^{-1}\times\ \frac{1mol}{6.022\times10^{23}nuclei} = 4.55\times10^{-12}J = 4.55J$$

c) that 1eV=1.602×10–19J

. Using the binding energy computed in part (b):

$$\Delta E = 4.55\times10^{-12}J\times\frac{1eV}{1.602\times10^{-19}J} = 2.84\times10^{7}eV = 28.4\ MeV$$

- ## Nuclear Stability

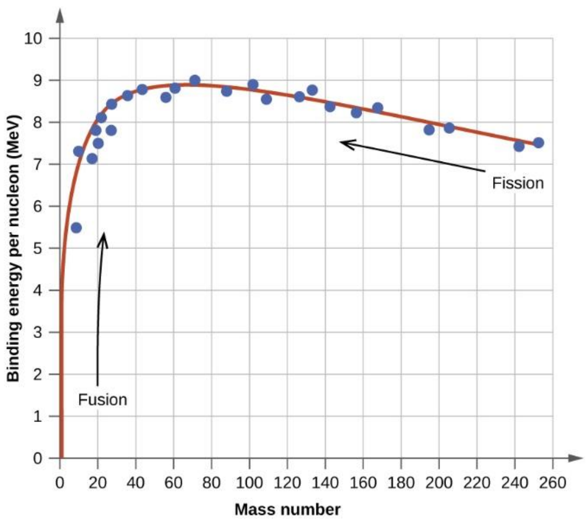
A nucleus is stable if it cannot be transformed into another configuration without adding energy from the outside. Of the thousands of nuclides that exist, about 250 are stable. A plot of the number of neutrons versus the number of protons for stable nuclei reveals that the stable isotopes fall into a narrow band.

represents nuclei that have a 1:1 ratio of protons to neutrons (n:p ratio). Note that the lighter stable nuclei, in general, have equal numbers of protons and neutrons. For example, nitrogen-14 has seven protons and seven neutrons. Heavier stable nuclei, however, have increasingly more neutrons than protons. For example: iron-56 has 30 neutrons and 26 protons, an n:p ratio of 1.15, whereas the stable nuclide lead-207 has 125 neutrons and 82 protons, an n:p ratio equal to 1.52. This is because larger nuclei have more proton-proton repulsions, and require larger numbers of neutrons to provide compensating strong forces to overcome these electrostatic repulsions and hold the nucleus together.

The relative stability of a nucleus is correlated with its binding energy per nucleon, the total binding energy for the nucleus divided by the number or nucleons in the nucleus. For instance, the binding energy for a ${}_{2}^{4}He$ nucleus is therefore:

$$28.4\ \text{MeV} / 4\text{nucleons} = 7.10\ \text{MeV/nucleon}$$

The binding energy per nucleon of a nuclide on the curve shown in Figure :



The binding energy per nucleon is largest for nuclides with mass number of approximately 56.

- **Summary**

An atomic nucleus consists of protons and neutrons, collectively called nucleons. Although protons repel each other, the nucleus is held tightly together by a short-range, but very strong, force called the strong nuclear force. A nucleus has less mass than the total mass of its constituent nucleons. This “missing” mass is the mass defect, which has been converted into the binding energy that holds the nucleus together according to Einstein’s mass-energy equivalence equation, $E = mc^2$. Of the many nuclides that exist, only a small number are stable. Nuclides with even numbers of protons or neutrons, or those with magic numbers of nucleons, are especially likely to be stable. These stable nuclides occupy a narrow band of stability on a graph of number of protons versus number of neutrons. The binding energy per nucleon is largest for the elements with mass numbers near 56; these are the most stable nuclei.