Atoms

- An **atomic model** is a scientific representation of the structure of an atom. Scientists proposed different atomic models to explain how electrons and protons are arranged inside an atom.
  The major atomic models are:

  i -Thomson's Atomic Model

  ii-Rutherford's Nuclear Model

  iii-Bohr's Atomic Model

  iv-Quantum Mechanical Model (Modern Atomic Model)

• Thomson's Atomic Model (Plum Pudding Model)

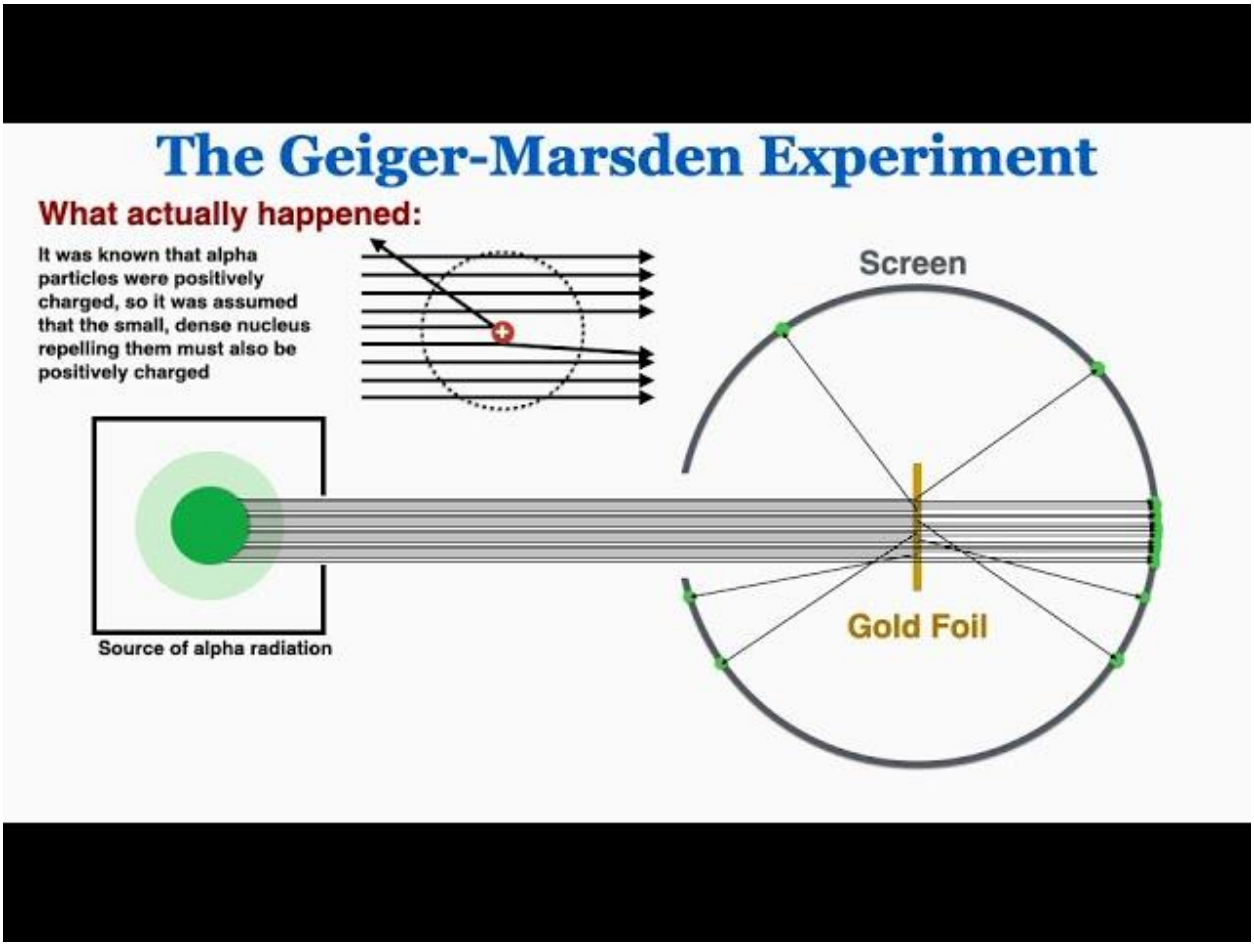
Proposed by: J. J. Thomson (1904)
According to Thomson:

An atom is a uniformly positively charged sphere. Negatively charged electrons are embedded in this sphere. The total positive charge is equal to the total negative charge. Therefore, an atom is electrically neutral.

- **Limitations:**
- It could not explain Rutherford's alpha-particle scattering experiment.
- It failed to explain the actual distribution of positive charge inside an atom.

Rutherford's Nuclear Model

- **Proposed by:** Ernest Rutherford (1911)
- Rutherford proposed this model on the basis of the alpha-particle scattering experiment.



Main observations of the experiment:

- Most alpha particles passed straight through the gold foil.
- Some alpha particles were deflected through small angles.
- Very few alpha particles were deflected through large angles or bounced back.

Conclusions:

- Most of the space inside an atom is empty.
- Almost the entire mass and positive charge of an atom are concentrated in a very small region called the nucleus.
- Electrons revolve around the nucleus.
- The nucleus is extremely small compared to the size of the atom.

Limitations:

- According to classical electromagnetic theory, revolving electrons should continuously radiate energy.

- Consequently, electrons should lose energy and fall into the nucleus.
- It could not explain the stability of atoms.
- It could not explain the line spectrum of hydrogen.

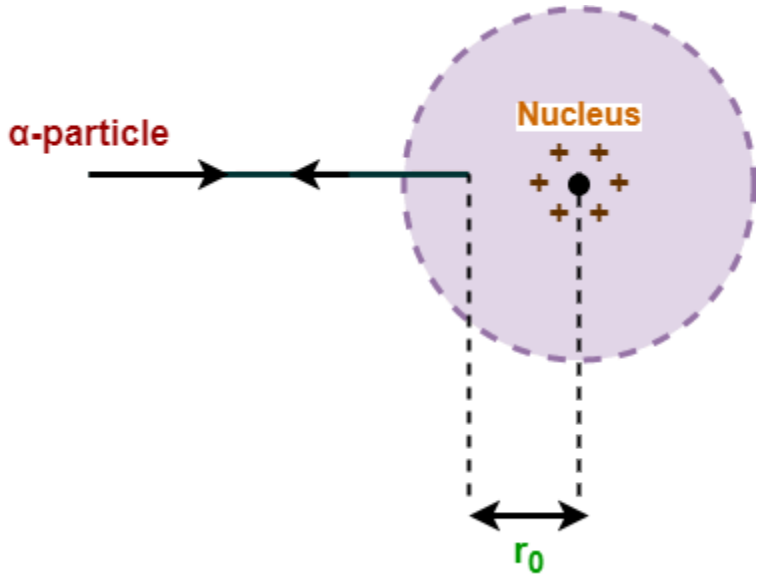
Distance of Closest Approach and Impact Parameter

- **Rutherford's Alpha-Particle Scattering Experiment**
- These two terms are very important for understanding Rutherford's alpha-particle scattering experiment.

Distance of Closest Approach

Definition:

The minimum distance between the positively charged nucleus and an alpha particle during a head-on approach is called the **distance of closest approach**.



Explanation-

- An alpha particle carries positive charge +2e.
- The nucleus also carries positive charge +Ze.
- Therefore, a repulsive electrostatic force acts between them.
- As the alpha particle approaches the nucleus, its kinetic energy gradually converts into electrostatic potential energy.
- At the closest point, its velocity momentarily becomes zero.

After that, the alpha particle moves back due to electrostatic repulsion

- Expression for distance of closest approach-

  According to the law of conservation of energy:

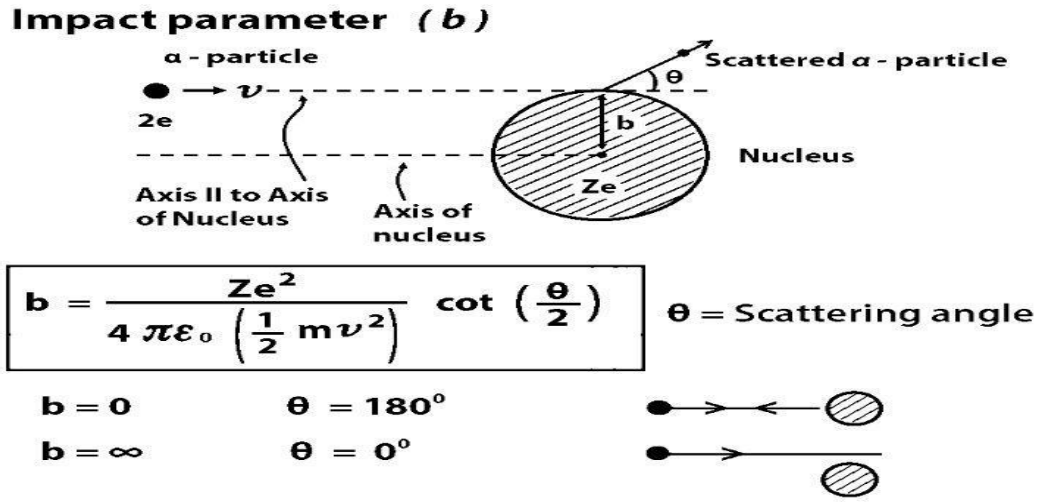
  Initial kinetic energy = Electrostatic potential energy at the closest point.

  $\frac{1}{2}mv^2 = \frac{2ze2}{4\pi\epsilon 0 r0}$= k (kinetic energy)

  Therefore r_0 = $\frac{2ze2}{4\pi\epsilon 0 K}$

  In Rutherford's classic gold-foil experiment, its value is on the order of **10^{-14} m (tens of femtometers)**

- **Important points:**
- It occurs in a head-on collision.
- At this point, the alpha particle has zero kinetic energy momentarily.
- The distance depends on the kinetic energy and nuclear charge.
- Impact Parameter
- **Definition:** The perpendicular distance between the initial direction of motion of an alpha particle and the centre of the nucleus is called the **impact parameter**.



  It is generally represented by b = $\frac{ze2cot\theta/2}{4\pi\epsilon 0 K}$ where K is the kinetic energy

- **Explanation:**

- When an alpha particle moves towards a nucleus, it may not travel directly towards its centre.Its initial path may have some perpendicular distance from the nuclear centre.This perpendicular distance is called the impact parameter.Due to electrostatic repulsion, the alpha particle changes its direction of motion.
- **Important points:**

  Impact parameter is a geometrical quantity.

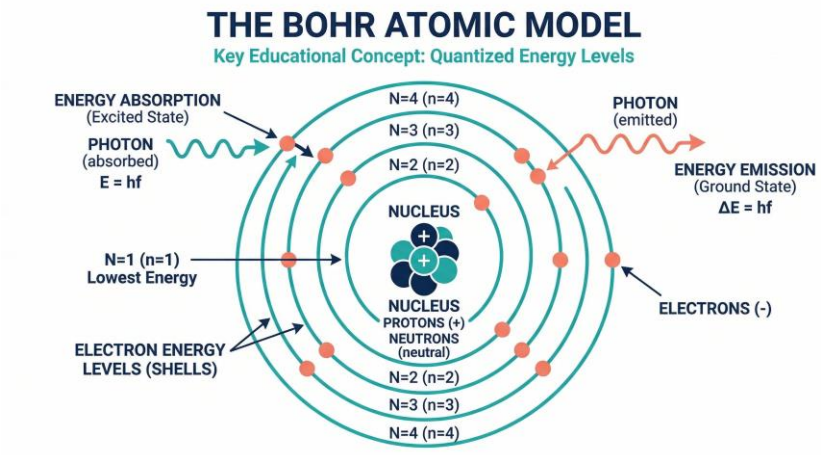
  It is measured before the alpha particle approaches the nucleus.

  For a head-on collision, impact parameter is zero.

  A larger impact parameter generally produces a smaller scattering angle.

• Bohr's Atomic Model

- **Proposed by:** Niels Bohr (1913)
  Bohr modified Rutherford's model and introduced the concept of fixed energy levels.



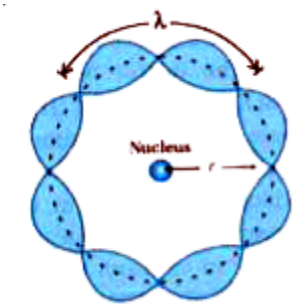
- **Main postulates:**
  - **Nuclear concept-**

An atom consist of a small and massive central core , called nucleus around which electrons revolve. The required centripetal force provided by electrostatic force of attraction, it may be given as $mv^2/r = kze^2/r$

Quantum condition-

Of all possible circular orbits allowed by the classical theory , the electrons are permitted to circulate only on those orbit in which the angular momentum of an electron is an integral multiple of ($h/2\pi$) so we can write $L = mvr = nh/2\pi$.

- **Proof of Bohr's quantization condition of angular momentum -**



Let r is the radius in which electron is moving with velocity v , having mass 'm',
If there are n waves formed in the given radius r then , $2\pi r = n\lambda$,
According to de-Broglie, $\lambda = h/mv$
Therefore $2\pi r = nh/mv$,
$mvr = nh/2\pi$ which is the second postulates.

stationary orbits-

While moving in the permissible orbits , an electron does not radiate energy. These non radiating orbits are called stationary orbits

- **Frequency condition-**

An electron can emit or absorb radiation in the form of photons only when electron jumps from one orbit to another orbit . i.e. $E_2 - E_1 = hv = hc/\lambda$
In short we can say , according to Bohr's

Electrons revolve around the nucleus only in certain permitted circular orbits called stationary orbits. Electrons do not radiate energy while revolving in these permitted orbits. Each permitted orbit has a definite energy.

- Energy is emitted or absorbed when an electron jumps from one energy level to another.

Limitations of Bohr's postulates :-

- It successfully explains the hydrogen spectrum but fails to explain spectra of multi-electron atoms.
- It cannot explain the fine structure of spectral lines.
- It cannot explain the Zeeman and Stark effects completely.
- Bohr's theory of the hydrogen atom derives quantized orbital radii and energy levels by combining classical circular mechanics with angular momentum quantization.

Force Balance and Velocity

For an electron (mass m, charge -e orbiting a hydrogen nucleus (charge +e, Z=1 in a circle of radius r at speed v, electrostatic attraction supplies the centripetal force:

$Ke(ze)/r^2 = mv^2/r$

$Kze^2/r^2 = mv^2/r$
So, $r= kze^2/mv^2$ ……..(i)
But according to Bohr's second postulates,
$Mvr=nh/2\pi$
Therefore $r= nh/2\pi mv$ …….(ii)
From eq. (I) and (ii).
$Kze^2/mv^2 = nh/2\pi mv$,
Or, $v= 2\pi kze^2/nh$……..(iii) [$v \alpha 1/n$]
Putting the value of v from eq. (iii) in eq. (i) we get $r= n^2h^2/(4\pi^2mkze^2)$
[$r \alpha n^2$] ($r_1 : r_2 : r_3$ …..= 1:4:9:……..)

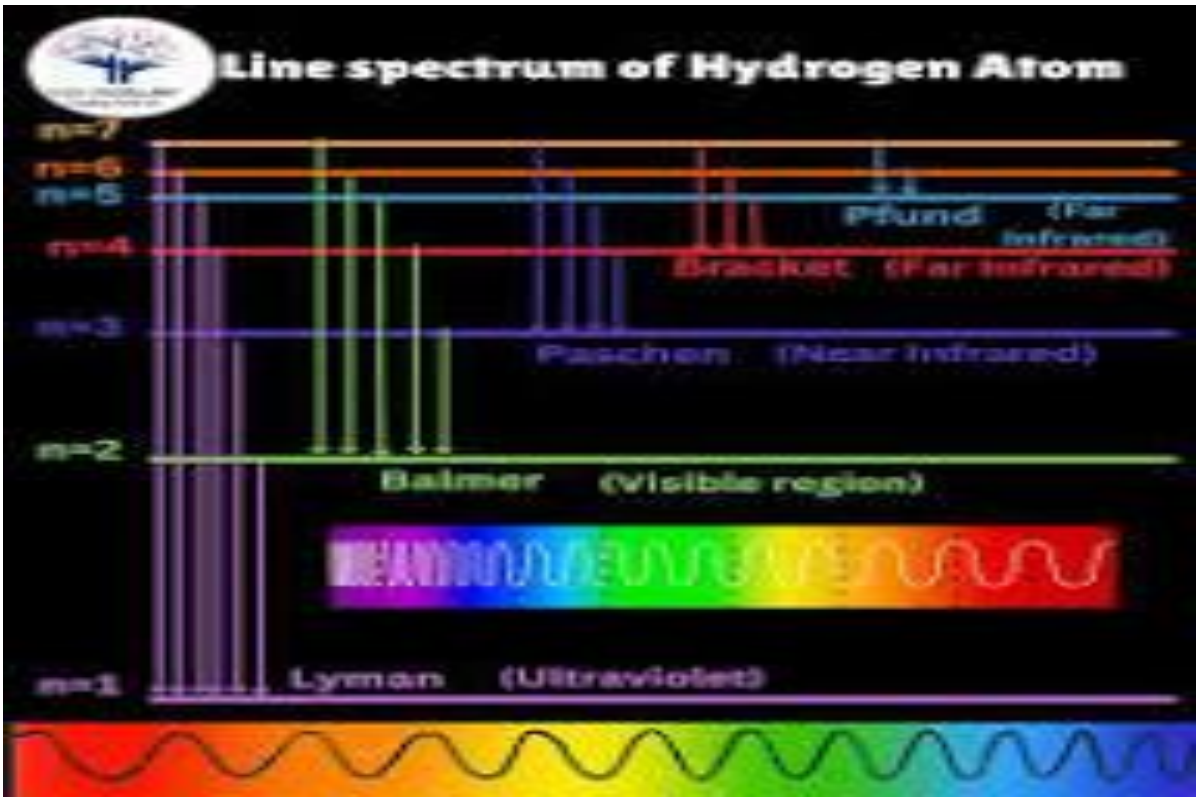
- ## Total Energy Derivation

Kinetic Energy (K):
$K = ½ (mv^2) = ½ m (2\pi kze^2/nh)^2 = (2\pi^2mk^2z^2e^4) /n^2h^2$
Potential Energy (U) :
$U= (-KQ_1Q_2 /r)$ putting the value of r from eq. (iv) we get
Total Energy-
Total energy E = K.E. + potential energy (U)
$E = [2\pi^2mk^2z^2e^4) /n^2h^2] + [- 4\pi^2mk^2z^2e^4) /n^2h^2] =$ **$- 2\pi^2mk^2z^2e^4 /n^2h^2$.i.e K.E. = - E = -U/2.**
Total energy for hydrogen atom
$E_n =(-13.6/n^2)$ eV .
$E_1= -13.6/1^2 = -13.6$eV
$E_2 = -13.6/2^2 = -3.4$eV
$E_3 = -13.6 /3^2 = -1.51$ eV
$E_4 = -13.6/4^2 = -0.85$ eV
$E_5 =-13.6 /5^2 = -0.54$ eV.

- **Spectral series of hydrogen atom-**
As we know total energy $E_n = - 2\pi^2mk^2z^2e^4 /n^2h^2$
Or, $E_{n1} = - 2\pi^2mk^2z^2e^4 /n_1{}^2h^2$
$E_{n2} = - 2\pi^2mk^2z^2e^4 /n_2{}^2h^2$
When electron jumps from n_1 to n_2 then
$E_{n2}- E_{n1} = (- 2\pi^2mk^2z^2e^4 /n_2{}^2h^2) – (- 2\pi^2mk^2z^2e^4 /n_1{}^2h^2)$. ……..(i)

And $E_{n2}- E_{n1} = hv$ (v = frequency of the emitted light)……..(ii)
From (i) and (ii)
We have; $hv = (- 2\pi^2mk^2z^2e^4 /n_2{}^2h^2) – (- 2\pi^2mk^2z^2e^4 /n_1{}^2h^2)$.
$hv = 2\pi^2mk^2z^2e^4 / h^2 (1/n_1{}^2 – 1/n_2{}^2)$;
or, $v = 2\pi^2mk^2z^2e^4 / h^3 (1/n_1{}^2 – 1/n_2{}^2)$;

or, $v = c/\lambda = \pi^2mk^2z^2e^4 / h^3 (1/n_1{}^2 – 1/n_2{}^2)$;
$1/\lambda = \pi^2mk^2z^2e^4 / ch^3 (1/n_1{}^2 – 1/n_2{}^2)$;
$1/\lambda = R (1/n_1{}^2 – 1/n_2{}^2)$ where, R = Rydberg constant ; $R = \pi^2mk^2z^2e^4 / ch^3$.

- 

- 1. Lyman series - when an electron jumps from any higher energy level to first level (or ground state) we get lymann series . this series lies in the ultraviolet region and is given by **$1/\lambda = R (1/1^2 – 1/n_2^2)$**
- **2. Balmer series -** - when an electron jumps from any higher energy level to second level (or first exited state) we get balmer series . this series lies in the visible region and is given by **$1/\lambda = R (1/2^2 – 1/n_2^2)$**

- **3.** - **Paschen series**- when an electron jumps from any higher energy level to third level (or second exited state) we get paschen series . this series lies in the infra red region and is given by **$1/\lambda = R (1/3^2 – 1/n_2^2)$**

- **4. Bracket series -** when an electron jumps from any higher energy level to fourth level (or third exited state) we get bracket series . this series lies in the infra red region and is given by **$1/\lambda = R (1/4^2 – 1/n_2^2)$.**

- **5 .Pfund series -** when an electron jumps from any higher energy level to fifth level (or fourth exited state) we get pfund series . this series lies in the infra red region and is given by **$1/\lambda = R (1/5^2 – 1/n_2^2)$.**

- **Excitation energy** –

For an atom it is the energy required by its electron to jump from the ground state to any one of the excited state.
First excitation energy of the electron is given as E_2-E_1 = -3.4 – (-13.6) =10.2 eV.

- **Excitation potential**-
  It is the accelerating potential which gives to a bombarding electron , sufficient energy to excite the target atom by rising one of its electron fro one inner to outer orbit.
  First excitation potential of hydrogen = -3.4-(-13.6)=10.2 V

- **Ionisation energy** –
  it is the energy required to knock an electron completely out of the atom .
  Ionisation energy of hydrogen atom =E_∞-E_1 = 0- (-13.6) = 13.6 eV .

- **Ionisation potential** –
  It is the accelerating potential which gives to a bombarding electron , sufficient energy to excite the target atom by rising one of its electron from one inner to complete out of the atom.
  Ionisation potential for hydrogen atom,= 0- (-13.6) = 13.6 eV