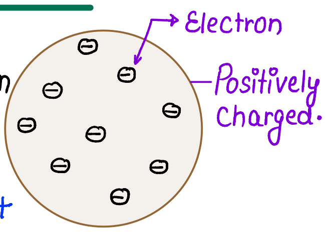
ATOMS

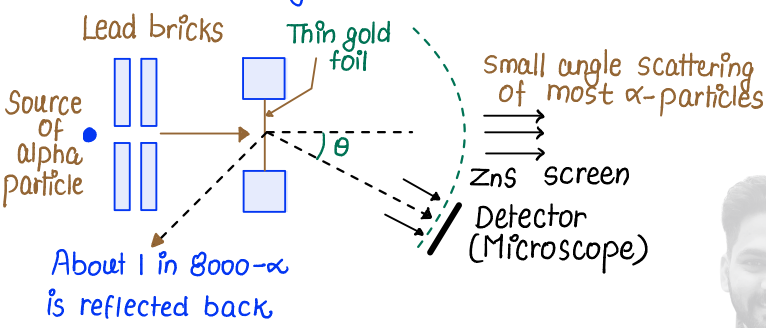
Thomson's model of atom

→ It was proposed by J.J Thomson in 1898. According to this model, the positive charge of atom is uniformly distributed throughout the volume of the atom and the negatively charged electrons are embedded in it like seeds in watermelon.



Rutherford's α-scattering Experiment

Rutherford and his two associates, Geiger and Marsden, studies the scattering of the α-particles from a thin gold foil in order to investigate the structure of the atom.

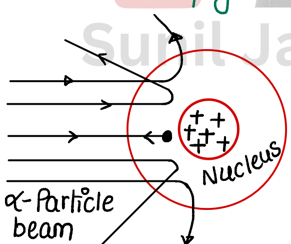


Observations: Most of the alpha particles went straight, some of α-particles deflected at acute angle and small fraction of α-particles deflect on obtuse angle and very-very small fraction α-particles suffers deflection of 180°. (This shows that size of Nucleus is very small, nearly 1/8000 times the size of the atom.

Conclusion: (a) Most of the space in atom is empty.

(b) The +ve charge is present at the centre of atom.

(c) The whole mass of atom is concentrated at the centre of atom in a tiny core i.e called atomic Nucleus.



(d) Negatively charged $e^{\ominus}$ are present out side of the nucleus of atom.

(e) Coulomb's Law holds for atomic distances.

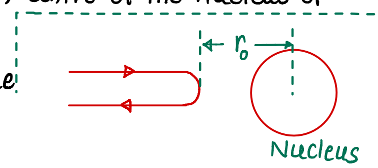
Rutherford's Nuclear Model of atom

According to this model the entire positive charge and most of the mass of the atom is concentrated in a small volume known as the nucleus with electrons revolving around it just as planets revolve around the sun.

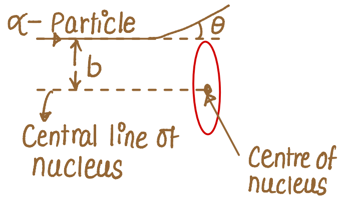
Distance of closest approach: (r_0)

The minimum distance between centre of the nucleus of atom to that point where alpha particle just deflected back due to repulsion.

i.e $\boxed{K.E = P.E}$ $\quad r_0 = \dfrac{2Ze^2}{4\pi\varepsilon_0 K_E}$

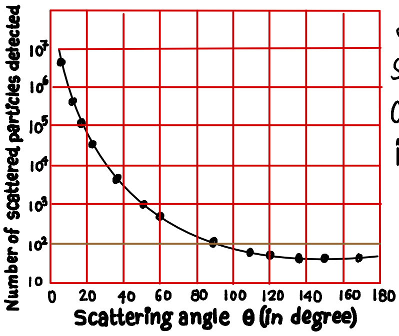


Impact Parameter: It is defined as the perpendicular distance of the initial velocity vector of α-particle from the centre of the nucleus, when the particle is far away from the nucleus of the atom.



$$b = \frac{Ze^2}{4\pi\varepsilon_0 K_E}\cot\left(\frac{\theta}{2}\right)$$

θ = Scattering Angle.



The number of α-particles scattered per unit area N(θ) at scattering angle θ varies inversely as

$$N(\theta) \propto \frac{1}{\sin^4 \theta/2}$$

Bohr model of the Hydrogen Atom

- An electron can revolve around the nucleus only in certain allowed circular orbits of definite energy and in these orbits it does not radiate. These orbits are known as **Stationary Orbits.**
- Angular momentum of the electron in a stationary orbit is an integral multiple of $h/2\pi$.

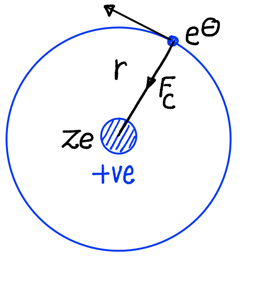
  i.e $L = \dfrac{nh}{2\pi}$ or $mvr = \dfrac{nh}{2\pi}$ $\quad h = 6.6\times10^{-34}\,Js$ (Planck's Constant)
- The emission of radiation takes place when an electron makes a transition from a higher to a lower orbit. The frequency of the radiation is given by $\boxed{h\nu = E_i - E_f}$ $\{E_i > E_f\}$

  E_i & E_f are the energies of the initial & Final states

▶ Bohr's Formulae

$$F_C = F_E \Rightarrow \frac{mv^2}{r} = \frac{Ze^2}{4\pi\varepsilon_0 r^2}$$

$$mv^2 = \frac{Ze^2}{4\pi\varepsilon_0 r} \quad \text{—(1)}$$



Acc to Bohr $\boxed{mvr = \dfrac{nh}{2\pi}} \Rightarrow v = \dfrac{nh}{2\pi mr}$

using value of v in eq (1)

$$\cancel{m}\frac{n^2h^2}{4\pi^2 m^{\cancel{2}} r^{\cancel{2}}} = \frac{Ze^2}{4\pi\varepsilon_0 \cancel{r}} \Rightarrow \boxed{r_n = \frac{n^2h^2\varepsilon_0}{\pi m Z e^2}}$$

Radius of first orbit of hydrogen atom is called Bohr radius. $(n=1, Z=1)$ $\quad r_1 = \dfrac{h^2\varepsilon_0}{\pi m e^2} \Rightarrow r_1 = 0.53\,A^\circ$

So $r_n = \dfrac{0.53\, n^2}{Z} A^\circ$

Velocity of electron

$$\frac{mv^2}{r} = \frac{Ze^2}{4\pi\varepsilon_0 r^2}$$

$$mv^2 r = \frac{Ze^2}{4\pi\varepsilon_0} \quad \text{—(1)}$$

Acc to Bohr

$$mvr = nh/2\pi \quad \text{—(2)}$$

Dividing eq (1) & (2)

$$\frac{mv^2 r}{mvr} = \frac{Ze^2}{4\pi\varepsilon_0}\,\frac{2\pi}{nh}$$

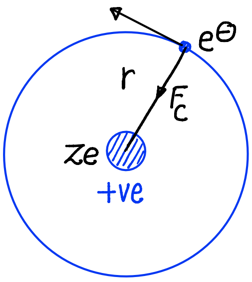
$$\boxed{V = \frac{Ze^2}{2nh\varepsilon_0}} \quad \boxed{V \propto \frac{1}{n}}$$

n = principle quantum Number.

Energy of Electron

Energy of Electron: The electron revolves around the nucleus in the circular orbits which possess electrostatic potential energy and kinetic energy.

$E = U + K.E$ —(1) $\boxed{U = -\frac{Ze^2}{4\pi\varepsilon_0 r}}$



Also $\frac{mv^2}{r} = \frac{Ze^2}{4\pi\varepsilon_0 r^2} \Rightarrow mv^2 = \frac{Ze^2}{4\pi\varepsilon_0 r}$

$\frac{1}{2}mv^2 = \frac{Ze^2}{8\pi\varepsilon_0 r}$ so $\boxed{K.E = \frac{Ze^2}{8\pi\varepsilon_0 r}}$

using value of U & K.E in equation

$E = -\frac{Ze^2}{8\pi\varepsilon_0 r}$ we know $r = \frac{\varepsilon_0 n^2 h^2}{\pi m Z e^2}$ $\rightarrow$ $\boxed{E = -\frac{mZ^2e^4}{8\varepsilon_0^2 n^2 h^2}}$

for H-atom $Z=1$

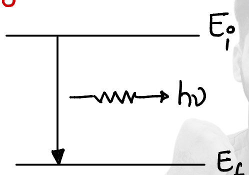
$E = \frac{-13.6\,eV}{n^2}$ $E = -\frac{13.6\,eV\,Z^2}{n^2}$

Expression for frequency and wave Number

Expression for frequency and wave Number: When electrons of higher orbit jump into lower orbit, then radiations emitted whose frequency is given by

$h\nu = E_i - E_f$

$h\nu = -\frac{mZ^2e^4}{8\varepsilon_0^2 n_i^2 h^2} + \frac{mZ^2e^4}{8\varepsilon_0^2 n_f^2 h^2}$



$h\nu = \frac{mZ^2e^4}{8\varepsilon_0^2 h^2}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right] \Rightarrow \nu = \frac{mZ^2e^4}{8\varepsilon_0^2 h^3}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$

Wave Number ($\bar{\nu}$) = The reciprocal of wavelength is called wave number

$\bar{\nu} = \frac{1}{\lambda} \Rightarrow \bar{\nu} = \frac{\nu}{c} \Rightarrow \frac{mZ^2e^4}{8\varepsilon_0^2 h^3 c}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$

$R = \frac{me^4}{8\varepsilon_0^2 h^3 c} = 1.097 \times 10^7\,m^{-1}$ = Rydberg Constant

$\bar{\nu} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$

Spectral Series of Hydrogen atom

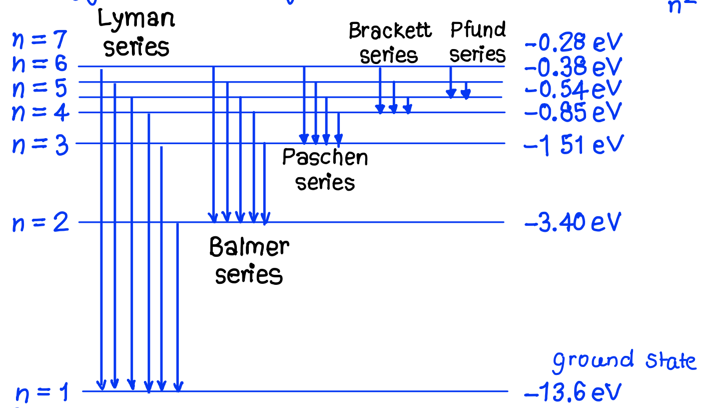
When the electron in a H-atom jumps from Higher energy level to lower energy level, the difference of energies of the two energy levels is emitted as radiation of particular wave length, known as spectral line.

For H-atom $Z=1$ $\frac{1}{\lambda} = R\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$

	n_f	n_i	
Lyman Series	1	2,3,4,5 --- ∞	U.V region
Balmer Series	2	3,4,5 --- ∞	visible region
Paschen Series	3	4,5,6 --- ∞	Low infrared region
Brackett Series	4	5,6,7 --- ∞	" " "
Pfund Series	5	6,7, ---- ∞	High infrared region.

Energy Level Diagram

$E = -\frac{13.6\,eV}{n^2}$



v.imp **No. of Spectral line**

$\boxed{N = \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}}$ $n_2 \rightarrow n_1$

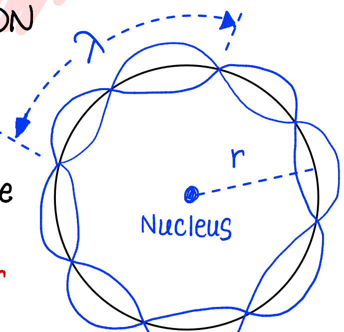
Ionization: The process of knocking an electron out of the atom is called ionization

Ionization energy: The energy required, to knock an $e^{\ominus}$ completely out of the atom.

Ionization energy $= \frac{13.6}{n^2}$ eV

DE-BROGLIE'S EXPLANATION OF BOHR'S SECOND POSTULATE OF QUANTIZATION

→ Acc to de Broglie, a stationary orbit is that orbit which contains an integral number of de-Broglie standing waves associated with the revolving electron.



For an $e^{\ominus}$ revolving in n^{th} circular orbit of radius r_n.

Total distance covered = Circumference of the orbit $= 2\pi r_n$

∴ for the Permissible Orbit $\Rightarrow 2\pi r_n = n\lambda$

Acc to de-Broglie $\lambda = \frac{h}{mv_n}$

v_n is the speed of $e^{\ominus}$ revolving in n^{th} Orbit.

so $2\pi r_n = \frac{nh}{mv_n} \Rightarrow mv_n r_n = \frac{nh}{2\pi}$

▶ Limitations of Rutherford's Atomic Model

- Failed to explain the stability of $e^{\ominus}$ in a circular path.
- cannot explain atomic line spectrum.

▶ Limitations of Bohr's Model

- The Bohr model is applicable to hydrogenic atoms.
- The model is unable to explain the relative intensities of the frequencies in the spectrum