UNIT — 8 (Marks — 6)

Atoms and Nuclei

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Thomsons Model of an atom ↦ In 1898 J. J. Thomson proposed that an atom is sphere with positive charg distributed uniformly over its entire volume and −vely charged electrons embedded in it. He compared the atom with a watermelon with +ve charge as that of red portion of the watermelon and electrons embedded like seeds in the watermelon.

Thomson atomic model is also k/a watermelon model or plum pudding model of atom.

Drawbacks of J. J. Thomsons atomic Model ↦

Thomson's model remained popular till 1911 but later on it was discarded because

(i) It could not explain the origin of spectral lines in case of hydrogen and other atoms.

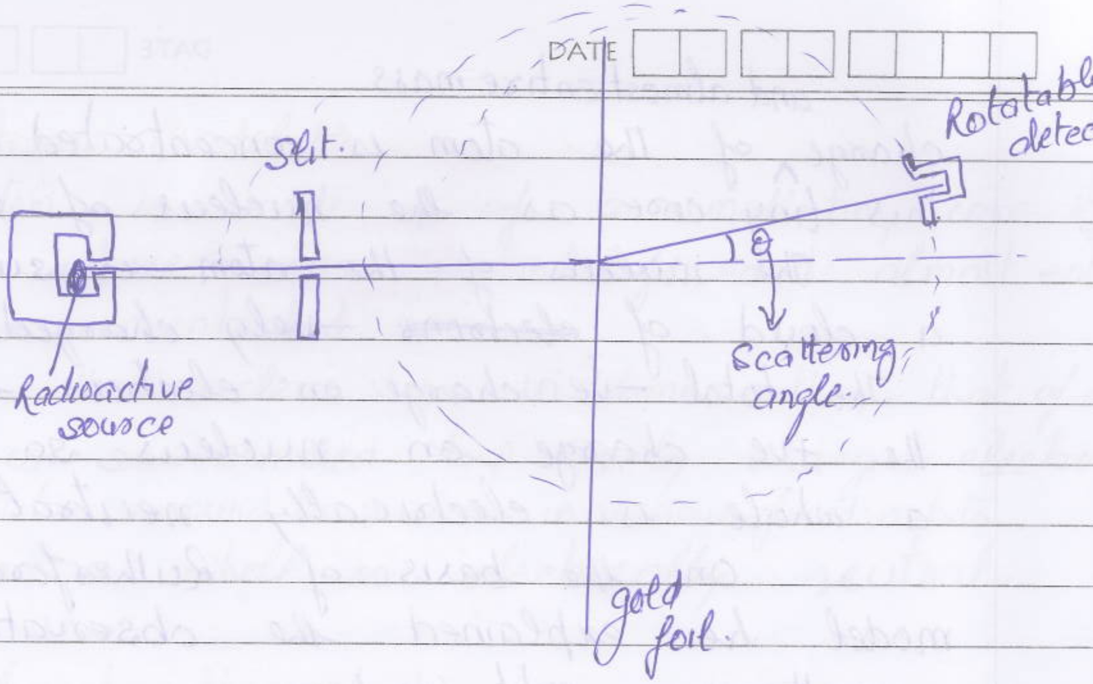
2. It could not explain the large angle scattering of α-particles in Rutherfords scattering.

Rutherfords α-particle scattering experiment ↦

α-particle ↦ An α-particle is a helium atom whose both the electrons have been removed.

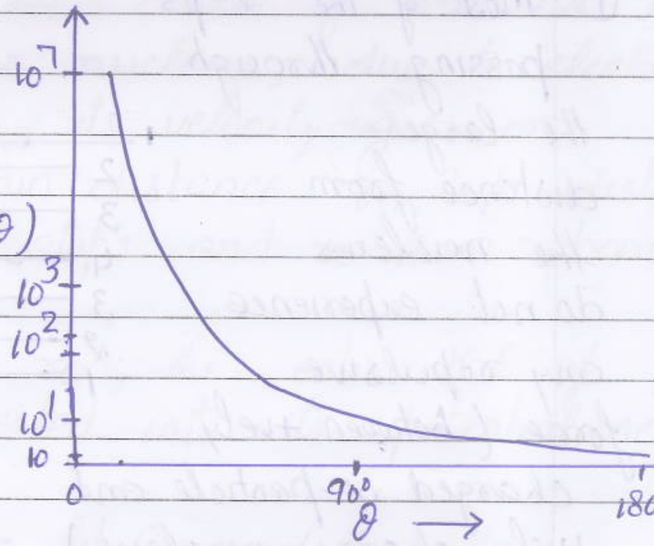
It has charge equal to +2e and mass nearly four times the mass of proton.

In Rutherfords α-particle scattering experiment, a beam of α-particles is allowed to fall on a thin gold foil of thickness 2×10^{-7} m. The number of α-particles scattered at different angles are counted with the help of a rotatable detector. A graph was plotted b/w the scattering angle θ and



~~no.~~ number of α-particles scattered at that angle θ $N(\theta)$.

The observations made by Rutherford are as below



① Most of the α-ps, pass through the gold foil without significant deviation.

2. A few α-ps deviate through a scattering angle more than 90°

3. Only one out of 8000, α-p rebounces back ie get deviated through an angle of 180°.

On the basis of above observations, Rutherford gave a new model of an atom k/a Rutherford atomic model.

In this model he suggested that atom is made up of small tiny core where whole +ve

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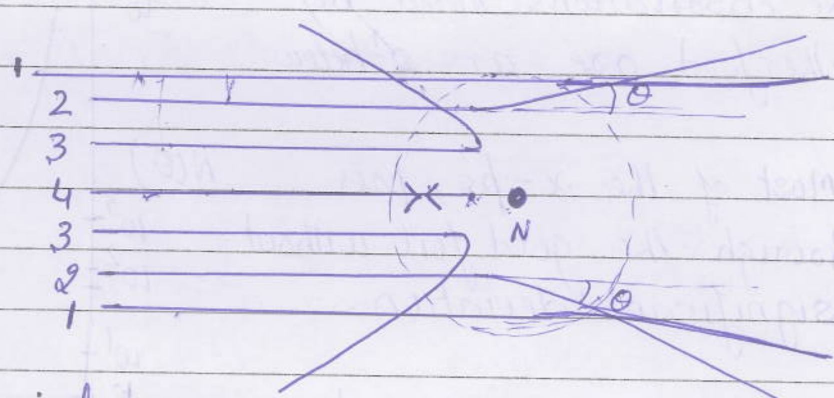
charge and almost entire mass of the atom is concentrated and he named this tiny core as the nucleus of the atom.

The nucleus of the atom is surrounded by a cloud of ~~electrons~~ -vely charged electrons.

The total -ve charge on electrons is equal to the +ve charge on nucleus so that atom as a whole is electrically neutral.

On the basis of Rutherfords atomic model he explained the observations of scattering expt as :-

① Most of the α-ps passing through the large distance from the nucleus do not experience any repulsive force (between +vely charged α-particle and +vely charge nucleus) and hence pass through the gold foil without any deviation.



2. α-particles passing through the periphery of the nucleus get deviated through small angle.

3. α-particles passing very close to the nucleus get strongly repelled and hence deviate through large angles.

4. A rare α-particle having head on collision with the tiny nucleus get deviated through 180° ie rebounces back.

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Rutherfords Atomic Model ↦

① An atom is made up of a small tiny core K/ nucleus where entire +ve charge and almost enti mass is concentrated.

2. Size of the nucleus is 10^{-5} times than that of a

3. Nucleus is surrounded by -vely charged electron These e^-s revolve arround the nucleus in various fixed orbits.

4. Atom as a whole is electrically neutral.

Distance of closest approach ↦

Consider an α-particle of mass m and velocity v is directed towards the centre of nucleus. A it approaches towards the nucleus, due to electro static force of repulsion, its velocity goes on decreasing and at certain distance r_0 k/a dista of closest approach, it stops and ~~re~~ then reboun back.

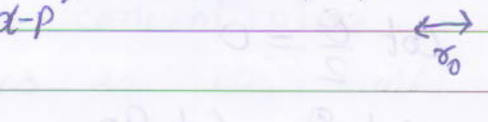
At this distance r_0, the entire K.E of the α-particle get converted into its potential ene

ie

Initial K.E of α-particle = Pot. Energy of the α-particle at the distance of closest approach —①

Where $K.E = \frac{1}{2} m v^2$ —②

and $P.E$ = charge × potential



$$P.E = 2e \times \frac{1}{4\pi\epsilon_0} \frac{Ze}{r_0} \quad —③$$

where $2e$ = charge on α-particle

and Ze = charge on the nucleus of atom.

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Put ② and ③ in eqn no-① we get

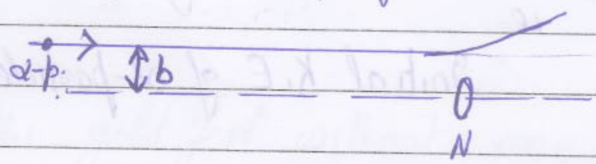
$$\frac{1}{2}mv^2 = qe \times \frac{1}{4\pi\epsilon_0}\frac{Ze}{r_0}$$

$$\therefore\ r_0 = \frac{1}{4\pi\epsilon_0}\frac{qZe^2}{\frac{1}{2}mv^2}$$

ie $$\boxed{r_0 = \frac{1}{4\pi\epsilon_0}\frac{qZe^2}{K.E}}$$

The distance of closest approach gives an approximation about the radius of the nucleus. As r_0 is just more than the radius of the nucleus.

Impact parameter ↦ The impact parameter is defined as the perpendicular distance of the velocity vector of the α-particle from central axis of the nucleus, when it is far away from the nucleus.



$$\boxed{b = \frac{1}{4\pi\epsilon_0}\frac{Ze^2 \cot \theta/2}{K.E.}}$$

for $b = 0$

$$\cot\frac{\theta}{2} = 0$$

$$\Rightarrow \cot\frac{\theta}{2} = \cot 90$$

$$\Rightarrow \frac{\theta}{2} = 90$$

$\Rightarrow \theta = 180^\circ$ ie scattering angle is 180°.

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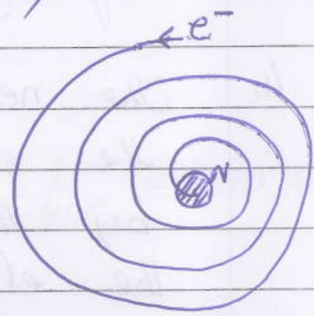
Limitations of Rutherfords atomic model ↦

(1) In Rutherford atomic model e^-'s revolving around the nucleus experience centripital acceleration, and we know accelerated charged particle radiates energy in form of em waves. Thus radius of the orbit should go on decreasing continuosly and ultim it should fall into the nucleus.
But atom is otherwise stable.
Thus Rutherford model cannot explain the stability of an atom



(2) Since revolving electron is continuosly emitting energy. Thus spectrum of the atom should be continuos but spectrum hydrogen like atom is discrete line spectrum
Thus Rutherfords atomic model fails to exp spectrum the hydrogen like atoms.

Postulates of Bohr's Theory of Hydrogen atom ↦

Bohr proposed an atomic model to explai the spectra emitted by Hydrogen atom. Bohr planetary model of an atom is based on following postulates.

(1) An atom consists of a small tiny core ~~othe~~ ~~or~~ k/a nucleus where almost entire mass whole +ve charge is concentrated.

(2) Size of the nucleus is 10^{-5} times the si of the atom

3. Electrons revolve around the nucleus only certain permitted orbits for which angular

momentum of the electron is integral multiple of $\frac{h}{2\pi}$, ie for permitted orbits angular momentum of the electrons is

$$mvr = n\frac{h}{2\pi}$$

for first orbit $n=1$, for second $n=2$ and so on

4. The necessary centripital force required by the electron for its rotation is provided by the electrostatic force of attraction b/w the electron and nucleus.

5. While revolving in these permitted orbits (stationary orbits) electron does not radiate energy.

6. Whenever an electron jumps from a higher orbit to lower orbit it emits energy equals to the difference of energies ~~betweel~~ between the two orbits

ie $h\nu = E_2 - E_1$

where E_2 is the energy of higher orbit.
and E_1 is the energy of lower orbit.

Bohr's Theory of Hydrogen atom ↦ Consider an electron of mass m and charge $-e$ is revolving around the nucleus of hydrogen like atom having $+Ze$ (for hydrogen $Z=1$) in n^{th} circular orbit of radius r_n.

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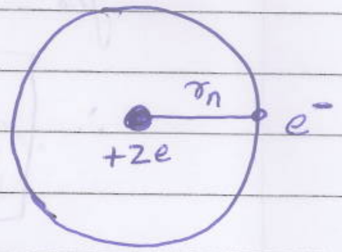
The necessary centripetal force required by the electron is provided by the electro static force of attraction b/w the electron and nucleus.



ie
$$\frac{mv^2}{r_n} = \frac{1}{4\pi\epsilon_0}\frac{(Ze)e}{r_n^2}$$

or
$$\frac{mv^2}{r_n} = K\frac{Ze^2}{r_n^2}$$ where $K = \frac{1}{4\pi\epsilon_0} = 9\times10^9 Nm$

$$\Rightarrow r_n = K\frac{Ze^2}{mv^2} \quad \text{—— (1)}$$

Also from Bohr's postulate of stationary orbits

$$mvr_n = \frac{nh}{2\pi}$$

$$\Rightarrow v = \frac{nh}{2\pi m r_n} \quad \text{—— (2)}$$

Substitute (2) in eqn no (1) we get.

$$r_n = K\frac{Ze^2}{m\left(\frac{nh}{2\pi m r_n}\right)^2}$$

$$\Rightarrow r_n = \frac{KZe^2 \cdot 4\pi^2 m^2 r_n^2}{mn^2h^2}$$

$$\therefore \boxed{r_n = \frac{n^2 h^2}{4\pi^2 KZe^2 m}} \quad \text{—— (3)}$$

for hydrogen atom $Z=1$

$$\therefore \boxed{r_n = \frac{n^2h^2}{4\pi^2 mke^2}}$$

for first orbit of hydrogen atom $n=1$

$$\therefore r_1 = \frac{h^2}{4\pi^2 mke^2} = a_0 = \text{Bohr's radius}.$$

on substituting the standard values.

$$a_0 = .53\,A^0.$$

for second orbit $n=2$.

$$\therefore r_2 = 4a_0$$

$$\text{llly} \quad r_3 = 9a_0 \quad \text{and so on.}$$

Thus $r_1 : r_2 : r_3 : \text{---} :: 1 : 4 : 9 : \text{---}$

Also from eqn no. 3

$$\boxed{r_n \propto n^2}$$

<u>Speed of electron</u> ↦

Substitute the value of r_n from eqn no 3 in eqn no. 2 we get

$$v = \frac{nh}{2\pi m\left(\frac{n^2h^2}{4\pi^2 mkZe^2}\right)}$$

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$$\therefore v = \frac{4\pi^2 m k Z e^2}{2\pi m n h}$$

$$\Rightarrow \boxed{v = \frac{2\pi k Z e^2}{nh}} \quad \text{———— (4)}$$

ie $v \propto \frac{1}{n}$

ie higher is the orbit, lesser will be th[e] velocity of electron.

for hydrogen atom $Z = 1$ and for first orbit $n = 1$

$$\therefore v_1 = \frac{2\pi k e^2}{h}$$

IIly $v_2 = \frac{v_1}{2}$ $\quad$ ($\because n = 2$

and $v_3 = \frac{v_1}{3}$ and so on.

<u>Time period of revolution</u> $\mapsto$ we know that time period of revolution of the electron is

$$T = \frac{\text{Circumference of the orbit}}{\text{Orbital velocity}}$$

ie $$T = \frac{2\pi r_n}{v}$$

ie $$T = 2\pi \frac{n^2 h^2}{4\pi^2 m k Z e^2} \times \frac{nh}{2\pi k Z e^2}$$ { from and

$$\Rightarrow \boxed{T = \frac{n^3 h^3}{4\pi^2 m k^2 Z^2 e^4}}$$

$$\Rightarrow T \propto n^3$$

ie $T_1 : T_2 : T_3 : \ldots :: 1 : 8 : 27 ! \ldots$

Energy of the electron ↦ while revolving around the nucleus, an electron possess K.E. due to its motion and potential energy due to potential of the nucleus. Thus total energy possessed by the electron in n^{th} orbit is

$$E_n = E_{K.E} + E_{P.E.} \quad \text{——— (5)}$$

where $E_{K.E} = \frac{1}{2} m v^2$

$$= \frac{1}{2} m \left(\frac{2\pi k Z e^2}{nh} \right)^2 \quad \{\text{using —(4)}\}$$

ie $$E_{K.E} = \frac{1}{2} m \cdot \frac{4\pi^2 k^2 Z^2 e^4}{n^2 h^2}$$

$$\Rightarrow \quad E_{K.E} = \frac{2\pi^2 m k^2 Z^2 e^4}{n^2 h^2} \quad \text{——— (6)}$$

and $E_{P.E} = \text{Charge} \times \text{Potential}$

$$= -e \cdot \frac{1}{4\pi\epsilon_0} \frac{Ze}{r_n}$$

$$= -e \times \frac{kZe}{r_n}$$

$$= -\frac{kZe^2}{\frac{n^2 h^2}{4\pi^2 m k Z e^2}}$$

$$\Rightarrow E_{P.E} = -\frac{4\pi^2 m k^2 Z^2 e^4}{n^2 h^2} \quad \text{--- (7)}$$

Thus using (5), (6) and (7).

$$E_n = \frac{2\pi^2 m k^2 Z^2 e^4}{n^2 h^2} - \frac{4\pi^2 m k^2 Z^2 e^4}{n^2 h^2}$$

$$\Rightarrow \boxed{E_n = -\frac{2\pi^2 m k^2 Z^2 e^4}{n^2 h^2}}$$

for hydrogen atom $Z = 1$

$$\therefore E_n = \frac{-2\pi^2 m k^2 e^4}{n^2 h^2} \quad \text{ie} \quad \boxed{E \propto -\frac{1}{n^2}}$$

ie higher orbit ~~higher energy~~

Spectral series of hydrogen atom ↦ When ever an jumps from any higher orbit n_2 to lower orbit n_1, it will radiate energy equals to

$$h\nu = E_{n_2} - E_{n_1}$$

$$= -\frac{2\pi^2 m k^2 e^4}{h^2 n_2^2} + \frac{2\pi^2 m k^2 e^4}{h^2 n_1^2}$$

$$h\nu = \frac{2\pi^2 m k^2 e^4}{h^2}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\therefore \nu = \frac{2\pi^2 m k^2 e^4}{h^3}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\Rightarrow \frac{c}{\lambda} = \frac{2\pi^2 m k^2 e^4}{h^3}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\frac{1}{\lambda} = \frac{2\pi^2 m k^2 e^4}{ch^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\Rightarrow \boxed{\bar{\nu} = \frac{2\pi^2 m k^2 e^4}{ch^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]}$$

where $\bar{\nu} = \frac{1}{\lambda}$ = wave number.

$$\Rightarrow \bar{\nu} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

where $R = \frac{2\pi^2 m k^2 e^4}{ch^3} = 1.0973 \times 10^7 \, m^{-1}$

= Rydbergs const.

<u>Lyman Series</u> $\mapsto$ when an electron jumps from any higher orbit to first orbit, we get Lyman Series

ie for Lyman series

$n_1 = 1$ and $n_2 = 2, 3, 4, ---$.

$$\therefore \bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{1^2} - \frac{1}{n_2^2}\right]$$

wavelength of the radiation emitted lies in ultraviolet region.

<u>Balmer Series</u> $\mapsto$

When ever an electron jumps from any higher ~~a~~ orbit to second orbit, we get Balmer Series.

for Balmer spectral lines

$n_1 = 2$ and $n_2 = 3, 4, 5 \ldots$

$$\therefore \bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{n_2^2}\right]$$

wavelength of the Balmer series lies in visible region.

Paschen Series ↦ When an electron jumps from any orbit to third orbit, we get Paschen series.
ie for Paschen Series

$n_1 = 3$ and $n_2 = 4, 5, 6 \ldots$

$$\therefore \bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{3^2} - \frac{1}{n_2^2}\right]$$

These Spectral lines lie in infrared region.

Bracket Series ↦ When an electron jumps from a higher orbit to fourth orbit, we get Bracket series.

ie for Bracket series

$n_1 = 4$ and $n_2 = 5, 6, 7 \ldots$

$$\therefore \bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{4^2} - \frac{1}{n_2^2}\right]$$

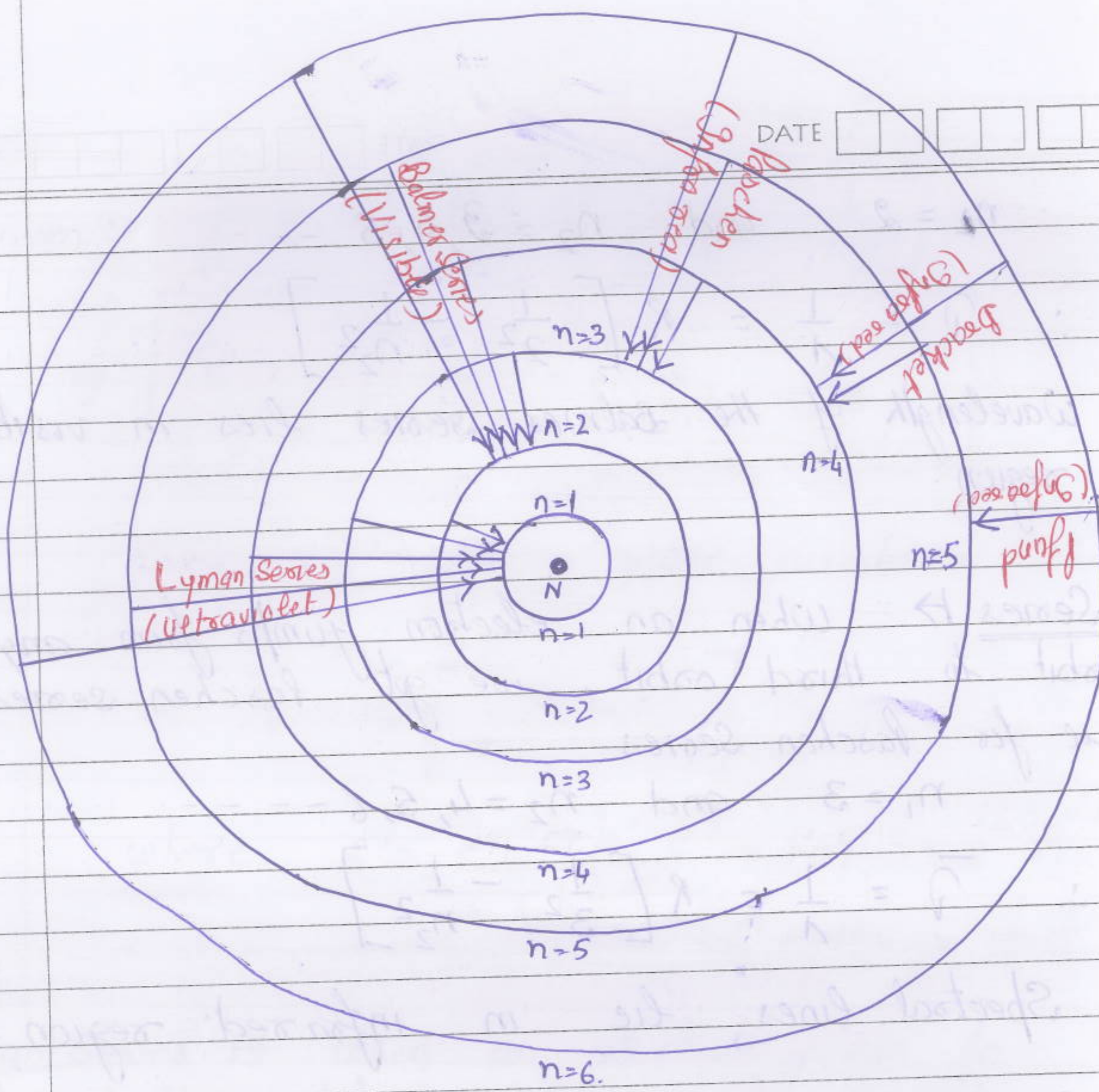
These spectral lines also lie in infra red region

Pfund Series ↦ When an electron jumps from a higher orbit to 5^{th} orbit, we get Pfund series
ie for Pfund series

$n_1 = 5$ and $n_2 = 6, 7, 8 \ldots$

$$\therefore \bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{5^2} - \frac{1}{n_2^2}\right]$$

Spectral lines in infra red region



Energy Level diagram for Hydrogen atom ↦

It is diagram in which energies of different stationary orbits of hydrogen atom are drawn according to some suitable energy scale. The various spectral series emitted are also indicated in the energy level diagram.

We know that energy of the electron in n^{th} energy orbit is given as

$$E_n = \frac{-2\pi^2 m k^2 e^4}{n^2 h^2}$$

on substituting the standard values

$$E_n = -\frac{13.6}{n^2} \; eV$$

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for first orbit $n=1$ $\therefore E_1 = -13.6$ eV. (Lowest Energy and Ground State)

" 2^{nd} " $n=2$ $\therefore E_2 = \frac{-13.6}{4} = -3.4$ eV.

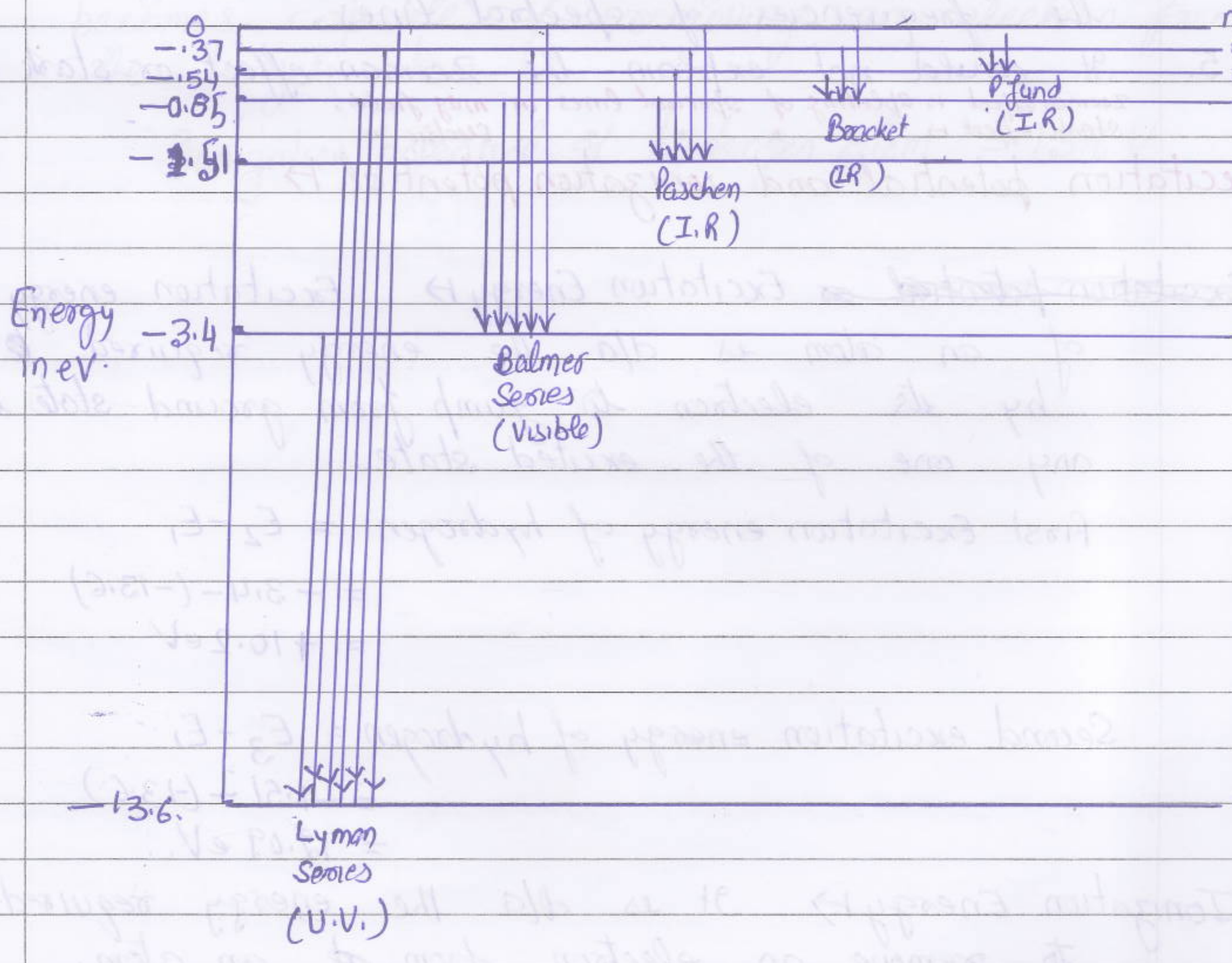
for 3^{rd} " $n=3$ $\therefore E_3 = -\frac{13.6}{9} = -1.51$ eV.

for 4^{th} " $n=4$ $\therefore E_4 = -\frac{13.6}{16} = -0.85$ eV

" 5^{th} " $n=5$ $\therefore E_5 = \frac{-13.6}{25} = -0.54$ eV.

for 6^{th} " $n=6$ $\therefore E_6 = \frac{-13.6}{36} = -0.37$ eV.

for ∞^{th} orbit $n=\infty$ $\therefore E_\infty = 0$ eV.



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Limitations of Bohr's Theory of a atom ↦

Bohr's theory successfully explained the spectrum of hydrogen atom but ~~it fails~~ yet it had following shortcomings-

1. It fails to explain the spectrum of multi electron atoms.
2. It fails to explain the fine structure of even hydrogen atom.
3. It ~~can~~ could not explain that why the orbits can be elliptical.
4. Bohr's theory does not tell anything about the intensity of spectral lines. It perdicts only about the frequencies of spectral lines.
5. It could not explain the Zeeman effect or stark effect.

Zeeman effect ↦ splitting of spectral lines in mag field.
Stark effect ↦ " " " " " Electric "

Excitation potential and ionization potential ↦

~~Excitation potential~~ ~~or~~ Excitation Energy ↦ Excitation energy of an atom is d/a the energy required by its electron to jump from ground state to any one of the excited state.

First Excitation energy of hydrogen $= E_2 - E_1$

$= -3.4 - (-13.6)$

$= +10.2\,eV$

Second excitation energy of hydrogen $= E_3 - E_1$

$= -1.51 - (-13.6)$

$= 12.09\,eV.$

Ionization Energy ↦ It is d/a the energy required to remove an electron from ~~at~~ an atom.

Ionization Energy of Hydrogen $= E_\infty - E_1$
$= 0 - (-13.6) = 13.6\ eV.$

<u>Excitation potential</u> ↦ It is the accelerating potential given to the bombarding electron so that it becomes capable of sending the electron of target atom from ground state to excited atom.

first excitation potential of hydrogen atom = 10.2 V
Second " " " " " = 12.09 V.

<u>Ionization potential</u> ↦ It is the accelerating potential given to the bomarding electron so that it becomes capable of removing an electron from the target atom.

Ionization potential of hydrogen atom = 13.6 V.