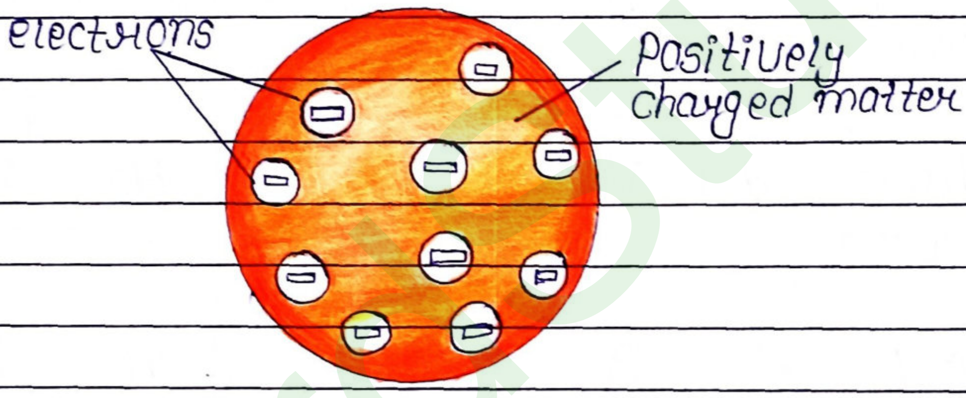
Date / /
Page

Atoms

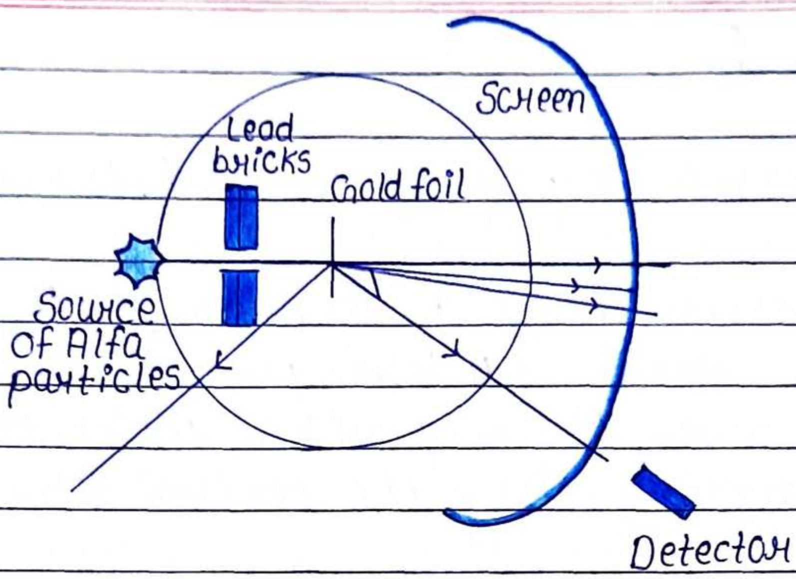
* Atom :- A atom is the smallest particle which can take part in all physical and chemical reactions.

Thomson Model (Plum Pudding) Of Atom :-



According to this model an atom consist of Nucleus in which there are protons and neutron and electron are embedded inside the nucleus like seeds in water-melon.
But this theory was wrong because electrons are not embedded inside nucleus.

* Rutherford α- Particle Scattering Experiment :-

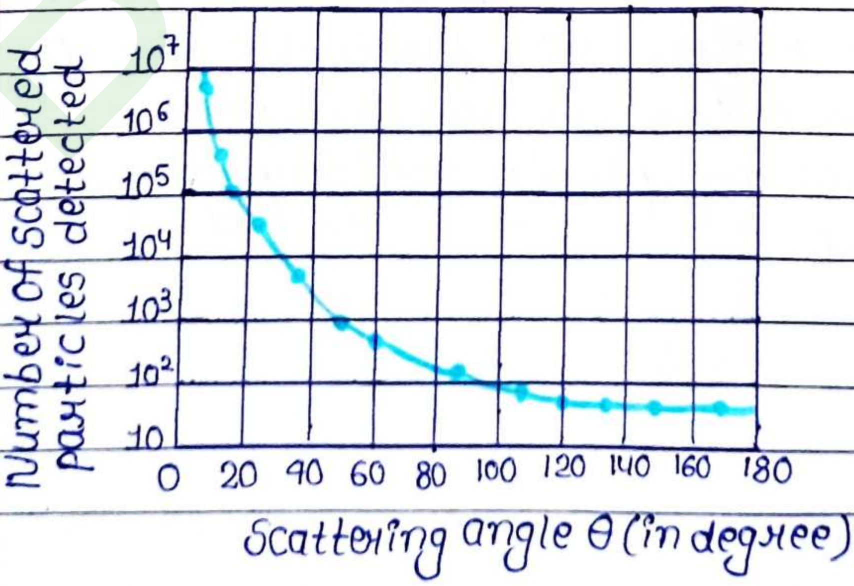


Following are the observations made through this experiment :

1) Most of α-particles were seen to pass through the fold foil without any appreciable deflection.

2) The various α particles, on passing through the gold foil, undergo different amounts of deflections. A large number of α particles caused fairly large deflections.

3) A very small number of α-particles (about 1 in 8000) practically retracted their paths or suffered deflection of nearly 180°.



Date __/__/__
Page ____

Conclusion :-

- There is some empty space in between the gold foil.
- There is some charge present on nucleus.
- There is a dense nucleus present at the centre because the α- particle which retraces its path moving along the central line.

* Rutherford's Atom Model :-

On the basis of the results of a scattering experiment, Rutherford suggested the following picture of an atom :

1. Atoms can be regarded as spheres of diameters 10^{10} m but whole of the positive charge and almost the entire mass of these atoms are concentrated in small central cores called nuclei having diameters of about 10^{14} m.

2. The nucleus is neighbored by electrons. In other words, the electrons are distributed over the remaining part of the atom leaving plenty of empty space in the atom.

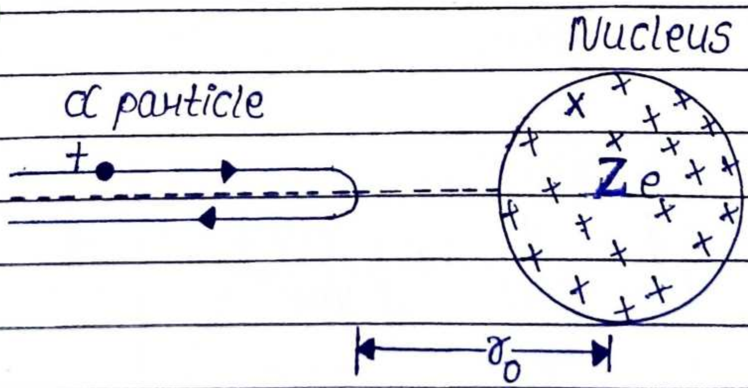
Date / /
Page

* Drawbacks of Rutherford's Atom Model :-

I. When the electrons revolve around the nucleus they get continuously accelerated towards the centre of the nucleus. According to Lorentz, an accelerated charged particle must radiate energy continuously. Thus, in the atom, a revolving electron must continously emit energy and hence the radius of its path must go on decreasing and finally, it must fall into the nucleus. However, electrons revolve around the nucleus without falling into it. Clearly, Ratherford's atom model couldn't explain the stability of the atom.

II. Suppose if Rutherford's atom model is true, the electron could revolve in orbits of all possible radii and thus it should emit a continuous energy sepectrum. But, atoms like hydrogen possess a line spectrum.

* Distance of closest approach :-



Note:- α- particle retraces its path when kinetic energy of α- particle get converted into electric potential energy.

Kinetic energy = Potential energy

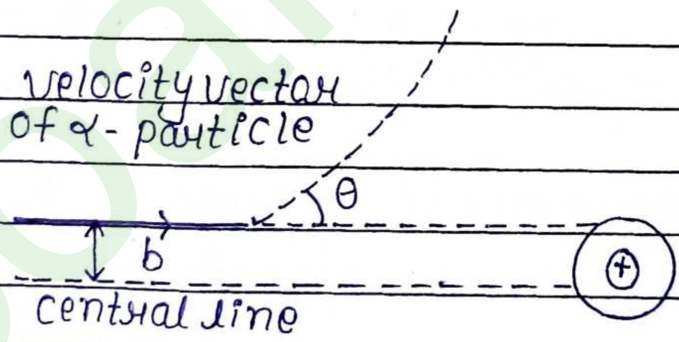
$$\frac{1}{2}mv^2 = \frac{Ze\ 2e}{4\pi\varepsilon_0 r_0}$$

$$r_0 = \frac{Ze\ 2e}{4\pi\varepsilon_0 \times \frac{1}{2}mv^2}$$

$$\boxed{r_0 = \frac{Ze\,2e}{4\pi\varepsilon_0 \times E}}$$

Impact Parameter :-

It is defined as the perpendicular distance of initial velocity vector of α- particle from the central line of the nucleus.



$$\boxed{b = \frac{1}{4\pi\varepsilon_0} \cdot \frac{Ze^2 \cot\frac{\theta}{2}}{\frac{1}{2}mv^2}}$$

Date / /
Page

* Alpha - particle Trajectory :-

- The path of an α- particle depends on the impact parameter, which is the distance from the particle's starting point to the centre of the nucleus.

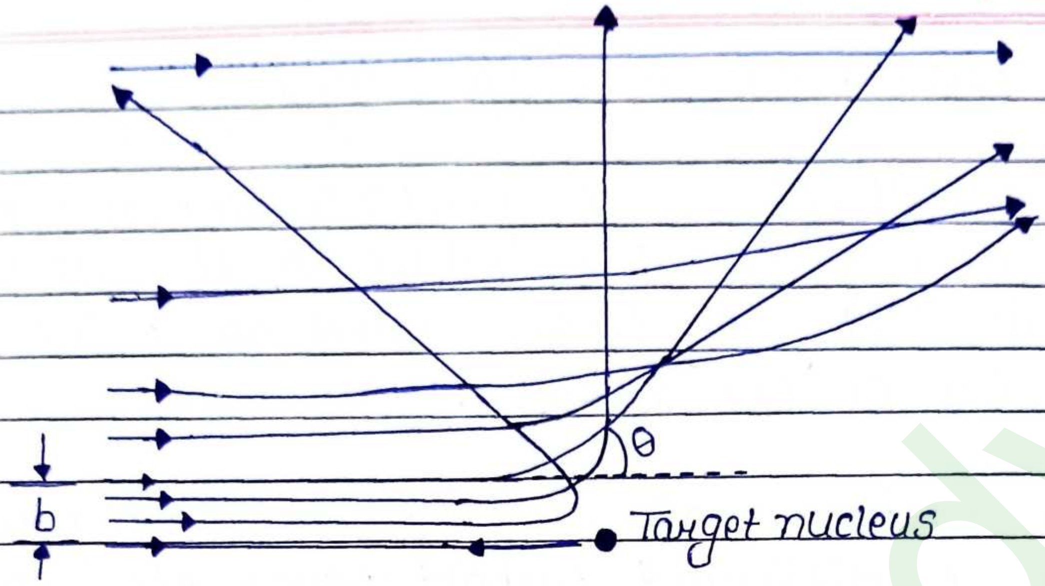
- In a beam of α- particles, each particle has a different impact parameter, causing them to scatter in various directions. They all have nearly the same energy.

- An α- particle that comes very close to the nucleus (small impact parameter) scatters a lot. If it hits the nucleus head-on (very small impact parameter), it bounces back ($\theta \approx 180°$).

- If the impact parameter is large, the α-particle barely changes direction ($\theta \approx 0°$).

- The fact that few α- particles bounce back means that head-on collisions are rare, showing that most of the atom's mass and positive charge is concentrated in a small area. This helps determine the size of the nucleus.

Date / /
Page



* Electron Orbits (Rutherford Model) :-

1. Rutherford's Nuclear Model :

The atom is modelled as a small, dense, positively charged nucleus with electrons revolving around it in fixed orbits.

The force of attraction between the nucleus and the electrons provides the necessary centripetal force to keep the electrons in stable orbits.

2. Centripetal Force :

The electrostatic force (F_e) between the electron and the nucleus equals the centripetal force (F_c) :

$$F_e = F_c = \frac{1}{4\pi\varepsilon_0}\frac{e^2}{r^2} = \frac{mv^2}{r}$$

Date / /
Page

Where :
ε_0 = permittivity of free space
e = electron charge
r = orbit radius
m = electron mass
v = velocity of the electron

3. Relation between orbit Radius and Electron Velocity :

The radius r of the electron's orbit is related to its velocity v by the formula :

$$r = \frac{e^2}{4\pi\varepsilon_0 m v^2}$$

4. Kinetic Energy (K) :

The kinetic energy K of the electron is given by :

$$K = \frac{1}{2} m v^2 = \frac{e^2}{8\pi\varepsilon_0 r}$$

5. Electrostatic Potential Energy (U) :

The potential energy U of the electron due to the nucleus is :

$$U = -\frac{e^2}{4\pi\varepsilon_0 r}$$

The negative sign indicates that the force is attractive.

Date ___/___/___
Page ______

6. Total Energy (E):

The total energy E of the electron in the hydrogen atom is the sum of its kinetic and potential energy:

$$E = K + U = \frac{e^2}{8\pi\varepsilon_0 r} - \frac{e^2}{4\pi\varepsilon_0 r}$$

$$E = -\frac{e^2}{8\pi\varepsilon_0 r}$$

The total energy is negative, meaning the electron is bound to the nucleus.

* Bohr Model of Hydrogen atom :-

Bohr combined classical and early quantum concepts and give theory in the form of 3 postulates :

1) According to Bohr, every atom consist of a central core called nucleus, in which entire positive charge and almost entire mass of atom is concentrated. A suitable number of electron revolve around the nucleus in circular orbit.

2) According to Bohr, e^- can revolve in certain non-radiating orbit called stationary orbit for which the total angular momentum of revolving electron is integral multiple of $\boxed{\frac{h}{2\pi}}$.

Date / /
Page

revolving e^-'s angular momentum = $\frac{nh}{2\pi}$

also, angular momentum = mvr

$$\because \boxed{\frac{nh}{2\pi} = mvr}$$

where, n is positive integer or principle quantum number.

↳ The quantum condition limits the number of allowed orbits.
The electron while revolving in such orbits, shall not lose energy.

3) The emission or absorption of energy occur only when e^- jumps from one of its orbit to another orbit and it is given as –

$$\boxed{E_2 - E_1 = h\nu}$$

* Radius of the Bohr's orbit :–

According to coulomb's law, electrostatice force of attraction (F) between moving electron and nucleus is –

$$F = \frac{KZe^2}{r^2}$$

where : K = constant = $\frac{1}{4\pi\varepsilon_0}$ = $9 \times 10^9 Nm^2/c^2$

and the centripetal force $F = \frac{mv^2}{r}$

Date ___/___/___
Page ______

Hence $\frac{mv^2}{r} = \frac{kze^2}{mr}$

or $v^2 = \frac{kze^2}{mr}$ —— (1)

From the postulate of Bohr,

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

or $v^2 = \frac{n^2h^2}{4\pi^2 m^2 r^2}$ —— (2)

From equation (1) and (2)

$$\therefore \quad r = \frac{n^2h^2}{4\pi^2 mkze^2}$$

On putting value or e, h, m

$$r = 0.529 \times \frac{n^2}{z} A$$

where A denotes armstrong, $1A = 10^{-10}$ meters.

* Velocity of an electron in Bohr's orbit :-

The total energy of an electron is revolving in a particular orbit is —

$$mvr = \frac{nh}{2\pi} \qquad v = \frac{nh}{2\pi mr}$$

putting the value of r in above equation then,

$$v = \frac{nh \times 4\pi^2 mZe^2}{2\pi mn^2h^2}$$

$$V = \frac{2\pi Z e^2}{nh}$$

On putting the value of e and h

$$V = 2.188 \times 10^6 \times \frac{Z}{n} \text{ m/sec.}$$

* Energy of an electron in Bohr's orbit :—

The total energy of an electron revolving in a particular orbit is —

$$T.E. = K.E. + P.E.$$

The K.E. of an electron $= \frac{1}{2}mv^2$

and the P.E. of an electron $= \frac{-kZe^2}{r}$

Hence, $T.E. = \frac{1}{2}mv^2 - \frac{kZe^2}{r}$ —— (3)

But $\frac{mv^2}{r} = \frac{kZe^2}{r^2}$

or $mv^2 = \frac{kZe^2}{r}$

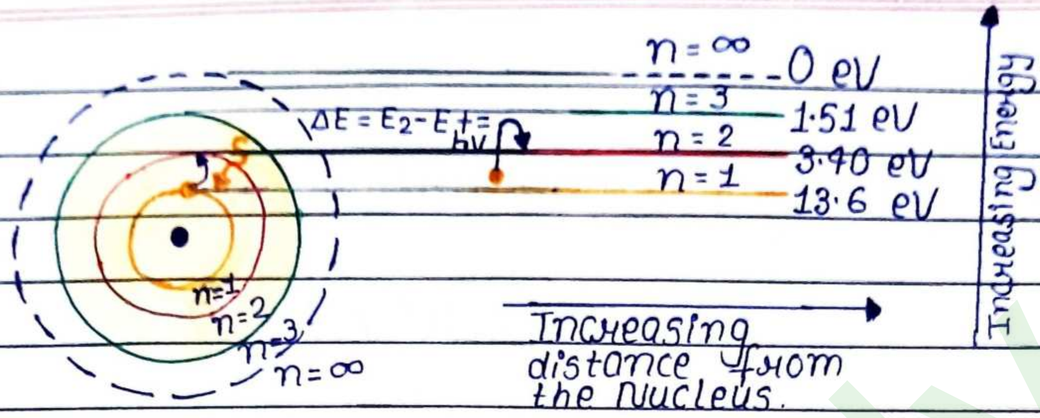
Substituting value of mv^2 in the eqn (3)

$$T.E = \frac{kZe^2}{2r} - \frac{kZe^2}{r} = \frac{kZe^2}{r}$$

So, $T.E = -\frac{kZe^2}{2r}$

Substituting value of 'r' in the equation of T.E.

$$E = -\frac{kZe^2}{2} \times \frac{4\pi^2 Z e^2 mk}{n^2h^2} = -\frac{2\pi^2 Z^2 e^4 mk^2}{n^2h^2}$$



Thus, the total energy of an electron in n^{th} orbit is given by

$$E = \frac{2\pi^2 Z^2 e^4 m k^2}{n^2 H^2}$$

$$= -13.6 \times \frac{Z^2}{n^2} \text{ eV/atom}$$

$$= -21.8 \times 10^{-19} \times \frac{Z^2}{n^2} \text{ J/atom}$$

$$= -313.6 \times \frac{Z^2}{n^2} \text{ Kcal/mole}$$

Relationship between P.E., K.E. & T.E. :-

$$P.E. = -\frac{KZe^2}{r}, \quad K.E. = \frac{1}{2}\frac{KZe^2}{r}, \quad T.E. = -\frac{1}{2}\frac{KZe^2}{r}$$

$$T.E. = \frac{P.E.}{2} = -K.E.$$

* The Line spectra of the Hydrogen Atom :-

↳ Hydrogen spectrum consist of bright lines in a dark background is known as hydrogen emission-spectrum.

Date / /

Page

↳ Hydrogen spectrum in which we get dark lines on the bright background is known as absorption spectrum.

Formula of wave number of radiation

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

where,

n_1 = final shell

n_2 = initial shell

R = Rydberg constant = $1.097 \times 10^{-7} m^{-1}$

or

$= 10^{-7} m^{-1}$.

1) For Lymon series :—

when an electron jumps to first orbit from any outer orbit.

$n_1 = 1$ and $n_2 = n$ $(2, 3 \ldots)$

$$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{(1)^2} - \frac{1}{n^2}\right]$$

This series is lies in UV region.

2) For Balmer Series :—

when an electron jumps to second orbit from any outer orbit.

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

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

Date ___/___/___
Page ______

$$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{(2)^2} - \frac{1}{n^2}\right]$$

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

This series lies in visible region.

3) For Paschen series :-

when an electron jumps to third orbit from any outer orbit.

$n_1 = 3$ and $n_2 = n$ (4, 5, 6 ----)

$$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{(3)^2} - \frac{1}{n^2}\right]$$

This series lies in infrared region.

4) For Brackett series :-

when an electron jumps to fourth orbit from any outer orbit.

$n_1 = 4$ and $n_2 = n$ (5, 6, 7 ----)

$$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{(4)^2} - \frac{1}{n^2}\right]$$

This series lies in infrared region.

5) Pfund Series :-

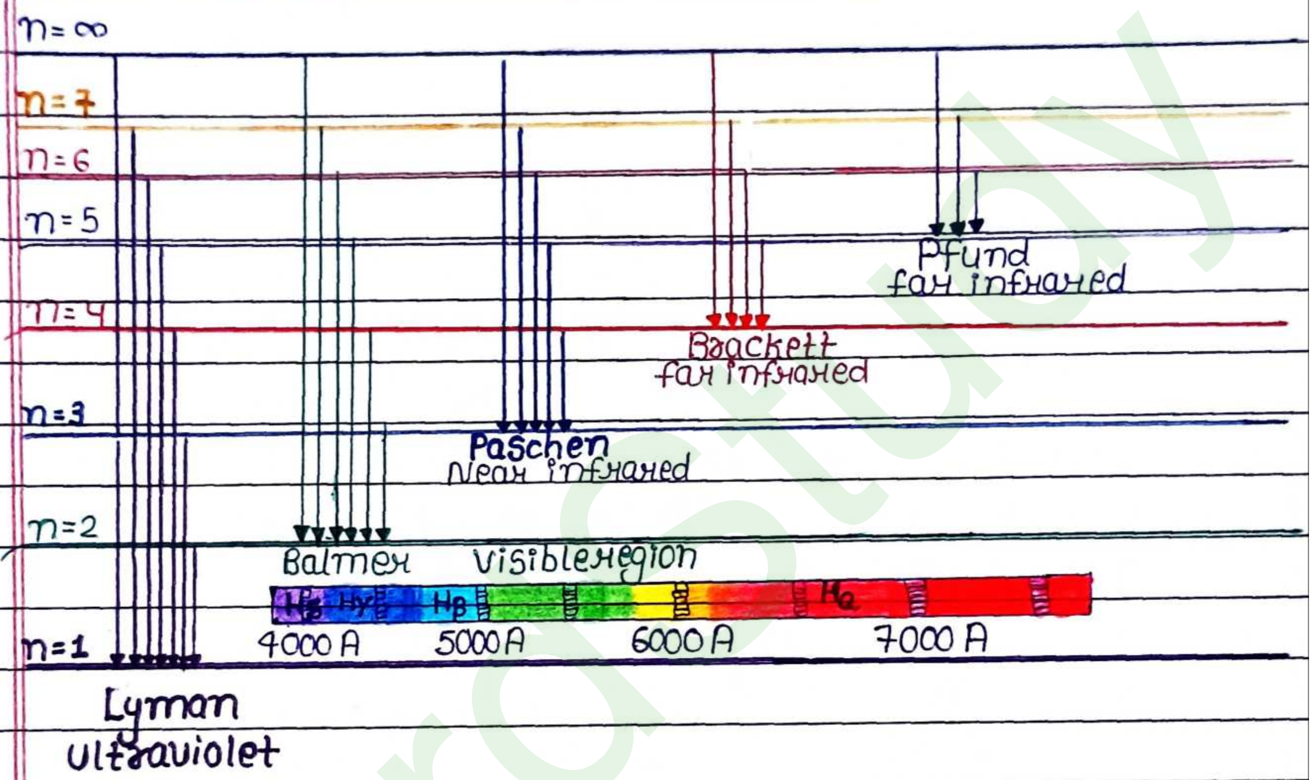
when an electron jumps to fifth orbit from outer orbit.

$n_1 = 5$ and $n_2 = n$ (6, 7 -----)

$$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{(5)^2} - \frac{1}{n^2}\right]$$

$$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{(5)^2} - \frac{1}{n^2}\right]$$

This series lies in infrared region.



* Limitation of Bohr's model :-

1) This model is applicable only to a simple atom like hydrogen having z = 1.

2) It does not explain the structure of spectral lines.

3) This model does not explain why orbital of e⁻ are taken as circular where as eliptical orbitales are also possible.

4) Bohr does say anything about relative intensities of spectral lines.