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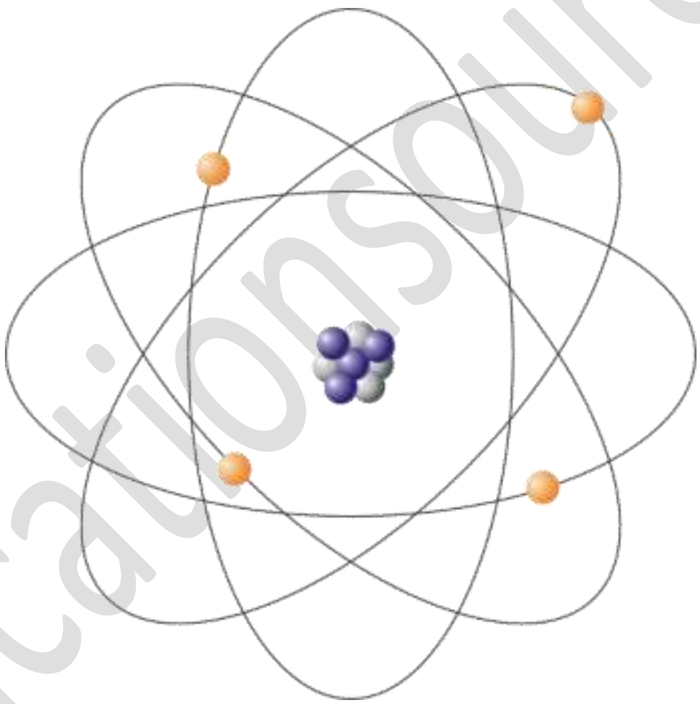
ATOMS

Chapter: - 12th



PHYSICS

CLASS: - 12TH

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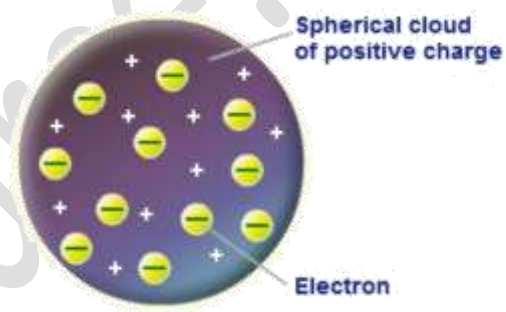
Chapter: - 12th
Atoms

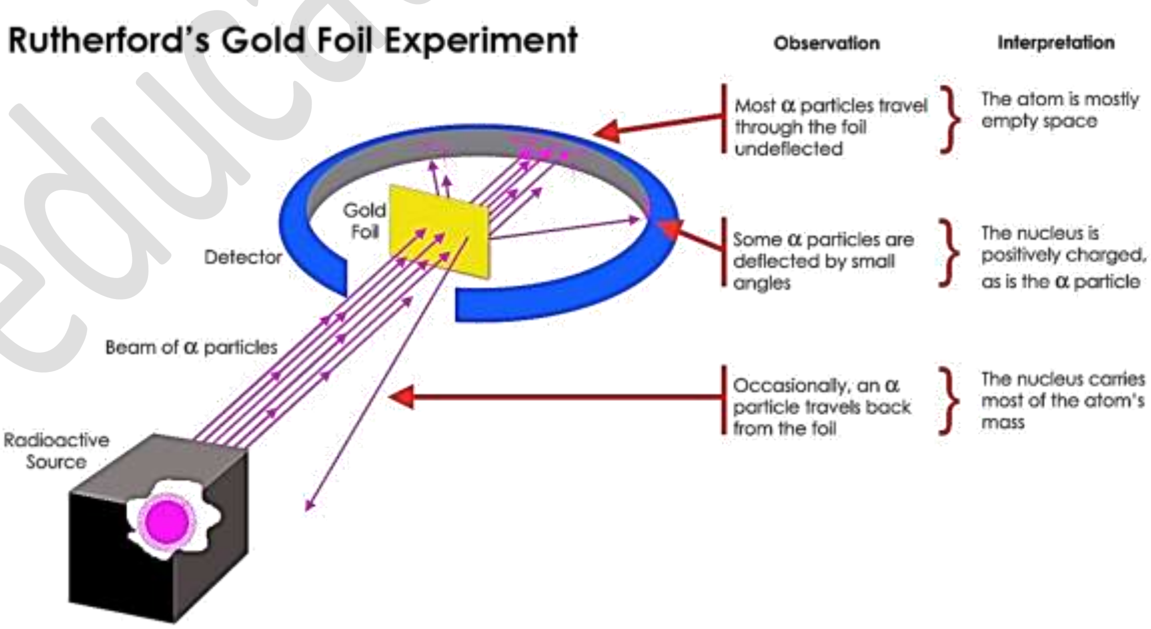
1) J.J. Thomson's Plum Pudding Model of the Atom:

- **Discovery:** J.J. Thomson discovered the electron, a negatively charged particle.
- **Basic Particles in Atoms:** Atoms consist of electrons, protons, and neutrons.
- **Thomson's Proposal:** Thomson proposed the first model of the atom based on his discovery.
- **Structure:** An atom is a sphere of positive charge with uniform density, approximately 10^{-10} m in diameter.
- **Electron Placement:** Negatively charged electrons are embedded in the positively charged sphere, resembling plums in a pudding.
- **Charge Neutrality:** The atom as a whole is neutral, with the positive and negative charges balanced.
- **Model Name:** This model is known as the **"plum pudding model"** or the **"positive sphere model."**



According to JJ Thomson's an atom is a sphere of positive charges of uniform density of about 10^{-10} m diameter in which negative charges **(i.e., electrons)** are embedded like plums **(i.e., fruits like dried grapes)** in the pudding. Thomson model of atom is also called **"plum pudding model".**

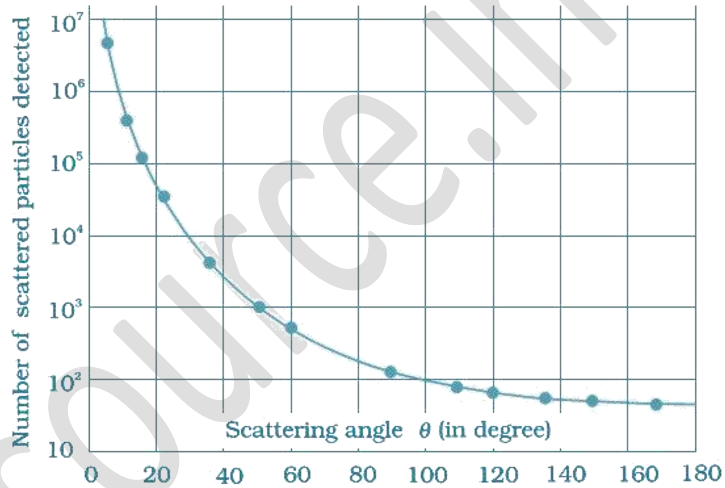
2) Alpha - particle scattering experiment: Rutherford's model of atom:



Experimental Setup:

- Source of alpha particles: Radium.
- Beam of alpha particles emitted from a radioactive substance, directed towards a gold foil.
- The gold foil is thin (10 nm) and placed inside an evacuated chamber.
- A circular screen coated with zinc sulfide (ZnS) is used to detect alpha particles.

Observations:



a) Most of the alpha particles pass through the gold foil undeflected.

b) Some alpha particles are deflected through small angles **(> 1°).**

c) A few alpha particles **(1 in 8000)** are deflected through large angles, and some even retrace their paths **(180° deflection).**

d) The number of alpha particles per unit solid angle, **N(θ),** varies with the scattering angle θ and follows the relationship:

$$N(\theta) = {}^{1}\!/_{Sin^4(\theta/2)}$$

e) The force between α- particle and nucleus is given by

$$F = \frac{1}{4\pi\varepsilon_o}\frac{(2e)(Ze)}{r^2}$$

r = distance between α- particle and nucleus.

Conclusion:

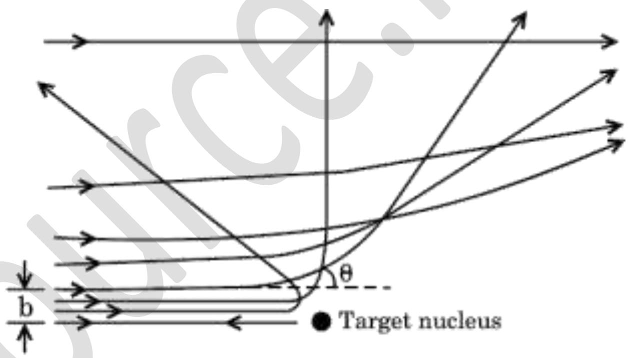
1. **Empty Space in Atom:**
   - Most alpha particles pass through undeflected, suggesting that most of the space in an atom is empty.
2. **Concentration of Mass and Charge:**
   - Alpha particles, being positively charged and having large mass, can only be deflected backward by a strong repulsive force.
   - Rutherford concludes that the positive charge in an atom is concentrated in a small, heavy region at the atom's center, called the **nucleus.**
3. **Nucleus:**
   - Rutherford proposes the existence of a nucleus where most of the atom's mass and positive charge are concentrated.

4. **Particle Interactions:**
   - Alpha particles traveling toward the nucleus are repelled and deflected.
   - Electrons, being lighter, do not significantly affect the path of alpha particles.

3) Alpha – Particle Trajectory and impact parameters: -

- **Impact Parameter (b):** The perpendicular distance of the velocity vector (u) of an incident alpha particle from the center of the nucleus when the alpha particle is not deflected.

Expression for Impact Parameter: $b = \frac{Ze^2 \cot(\theta/2)}{4\pi\varepsilon_0 E} = \frac{Ze^2 \cot(\theta/2)}{4\pi\varepsilon_0 (\frac{1}{2}mu^2)}$



- ***b*** is the impact parameter.
- ***Ze*** is the product of the atomic number and the elementary charge.
- ε_0 is the permittivity of free space.
- ***E*** is the kinetic energy of the alpha particle.
- ϑ is the scattering angle.

Observations:

1. **Large Impact Parameter (*b*):**
   - When ***b*** is large, **cot** $(\theta/2)$ is large, and ϑ (scattering angle) is small.
   - Alpha particles, traveling far from the nucleus, experience small deflections.

2. **Small Impact Parameter (�*b*):**
   - When ***b*** is small, **cot** $(\theta/2)$ is small, and ϑ is large.
   - Alpha particles, traveling close to the nucleus, experience significant deflections.

3. **Retracing Path:**
   - When ***b*=0,** the alpha particle is directed toward the center of the nucleus.

$$\frac{\theta}{2} = 90^o \quad or\ \theta = 180$$

   - meaning the alpha particle retraces its path when it travels directly toward the center of the nucleus.

Conclusion:

- The impact parameter plays a crucial role in determining the path adopted by alpha particles in the electrostatic field of the target nucleus during Rutherford's scattering experiment.
- Large impact parameters result in small scattering angles, indicating minor deflections, while small impact parameters lead to large scattering angles, indicating significant deflections or even retracing of paths.

4) Distance of closest approach: -

Distance of Closest Approach (r_0 or d): The minimum distance up to which an energetic alpha particle, traveling directly towards a nucleus, can approach before coming to rest and then retracing its path. This distance is denoted by ***r_0* or *d*** and estimates the size of the nucleus.

Assumptions made by Rutherford:

a) Elastic Collisions:
- Scattering occurs due to elastic collisions between the alpha particle and the nucleus.

b) Neglect of Nucleus Motion:
- The motion of the nucleus during impact is not taken into account because the nucleus is considered to be very heavy (e.g., a gold nucleus is 50 times heavier than an alpha particle).

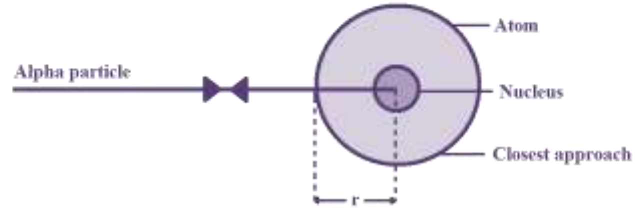
c) Point Charges:
- Both the alpha particle and the nucleus are treated as point charges.

d) Expression for Distance of Closest Approach:

This is the distance at which the kinetic energy (E) of a-particle is completely converted into electric potential energy of the system.

$$K.E. = potential\ energy$$

$$\boldsymbol{K.E. = electric\ potential + charge\ of\ \alpha - particle} \ldots\ldots\ldots \textbf{(1)}$$



Kinetic energy(E) = $\frac{1}{2}\boldsymbol{mu^2}$

$$\boldsymbol{electric\ potential = \frac{1}{4\pi\varepsilon_o}\frac{(Ze)}{r_o}}$$

Using in equation (1)

$$\boldsymbol{\frac{1}{2}mu^2 = \frac{1}{4\pi\varepsilon_o}\frac{(Ze)}{r_o} + 2e}$$

$$\frac{1}{2}mu^2 = \frac{1}{4\pi\varepsilon_o}\frac{(Ze)(2e)}{r_o}$$

$$r_o = \frac{2Ze^2}{4\pi\varepsilon_o\,(\frac{1}{2}mu^2)}$$

$$r_o = \frac{2Ze^2}{4\pi\varepsilon_o\,(E)}$$

- ***r_0*** is the distance of closest approach.
- ***e*** is the elementary charge.
- ***m*** is the mass of the alpha particle.
- ***u*** is the velocity of the alpha particle.
- ***Ze*** is the product of the atomic number and the elementary charge.

e) Explanation:

- As the alpha particle moves towards the nucleus, the repulsive force between the alpha particle and the nucleus increases.
- At the distance of closest approach, the alpha particle comes to rest, and its kinetic energy is completely converted into the electric potential energy of the system.
- The expression for r_0 relates the potential energy to the kinetic energy of the alpha particle at the point of closest approach.

Conclusion:

- The distance of closest approach is a crucial parameter in Rutherford's model, helping estimate the size of the nucleus based on the interaction between an alpha particle and the nucleus.
- It provides insights into the distribution of charge within the nucleus.

5) Summary of Rutherford's Nuclear Model of the Atom:

Nuclear Model:

a) Distribution of Mass and Charge:

- Almost all the mass of the atom and all the positive charges are concentrated in a very small region known as the atomic nucleus.
- The size of the nucleus is extremely small, with a diameter on the order of 10^{-15} meters, in stark contrast to the larger diameter of the entire atom (about 100 pm).

b) Electron Behavior:

- Negatively charged electrons revolve around the nucleus but are located far away from it.
- Most of the space in an atom is empty, indicating that the nucleus occupies a tiny fraction of the total volume.

c) Centripetal Force and Coulomb's Attraction:

- The centripetal force required for the electrons to revolve around the nucleus is provided by the Coulomb's force of attraction between the positively charged nucleus and the negatively charged electrons.

i.e., $F_c = F_e$

$$\frac{mv^2}{r} = \frac{1}{4\pi\varepsilon_o}\frac{(e^2)}{r^2}$$

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

Kinetic energy: - K = $\frac{1}{2}mv^2$ **(from equation 1)**

$$K = \frac{1}{2}\left(\frac{1}{4\pi\varepsilon_o}\frac{(e^2)}{r}\right) = \frac{1}{8\pi\varepsilon_o}\frac{(e^2)}{r}$$

Potential energy: - U = $-\frac{e^2}{4\pi\varepsilon_o r}$

Thus, total mechanical energy is

$$E = K + U = \frac{1}{8\pi\varepsilon_o}\frac{(e^2)}{r} + \left(-\frac{e^2}{4\pi\varepsilon_o r}\right)$$

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

d) Electrically Neutral Atom:

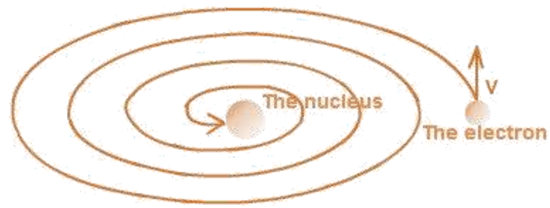
- The number of revolving electrons is equal to the number of positive charges in the nucleus, ensuring that the atom is electrically neutral.

e) Total Energy of Electrons:

- The total energy of electrons in the atom is negative.

6) Drawbacks of Rutherford's Model of the Atom:

a) Failure to Explain Stability:



- Rutherford's model proposed electrons orbiting the nucleus in circular paths.
- According to classical electrodynamics, an accelerated charged particle emits electromagnetic radiation and loses energy.
- Continuous loss of energy should cause the electron's orbit to gradually decrease, leading to it spiraling into the nucleus.
- This contradicts the observed stability of atoms, where electrons do not fall into the nucleus over time.

b) Inability to Justify Complex Spectra:

- Rutherford's model suggested that electrons could revolve around the nucleus in circular orbits of any possible radius.
- According to classical physics, this would result in atoms emitting a continuous energy spectrum.
- However, actual observations, especially in the case of hydrogen, revealed line spectra rather than continuous ones.
- Rutherford's model failed to explain the discrete spectral lines observed in the emission spectra of atoms.

7) BOHA'S Atom Model: -

Rutherford's model of atom failed to explain the stability of atom as well as spectrum of radiation emitted and absorbed by in atom. To explain the stability and the radiation spectrum concept of an atom, **Niels Henrik David Bohr** applied **Planck's quantum theory of radiation** to Rutherford's model. He used classical as well as quantum concepts to form his theory. He made the following assumptions: -

Postulates of Bohr's Atom Model: -

a) Electron Orbits:

- Electrons revolve around the nucleus in circular orbits without the emission of radiant energy.

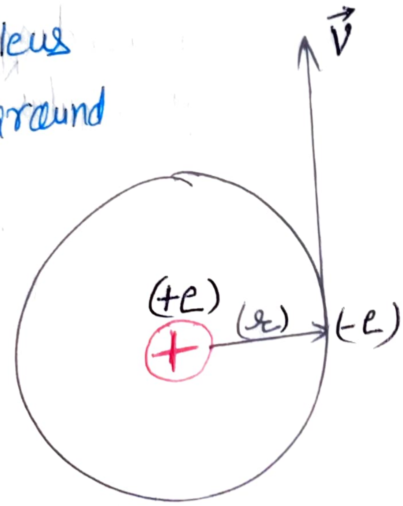
b) Quantization of Angular Momentum:

- Only certain orbits, with specific quantized values of angular momentum, are stable. The angular momentum $\mathbf{L} = mvr$ is quantized and is an integral multiple of $h/_{2\pi}$

  where m is the mass of the electron,
  v is its velocity,

⇒ Bohr's Theory of Hydrogen atom :-

Hydrogen atom Consists of a nucleus having charge (+e) and electron Revolve around it (-e).



The Coulomb force of attraction between nucleus and electron.

$$F_n = \frac{1}{4\pi\varepsilon_0} \frac{e \times e}{r_n^2}$$

$$\boxed{F_n = \frac{1}{4\pi\varepsilon_0} \frac{e^2}{r_n^2}} \quad \text{--- (I)}$$

This force provide necessary Centripital force for the electron

$$\text{ie } F_c = \frac{m v_n^2}{r} \quad \text{--- (II)}$$

equating (I) and (II)

$$F_n = F_c$$

$$\frac{m v_n^2}{r_n} = \frac{1}{4\pi\varepsilon_0} \frac{e^2}{r_n^2}$$

$$m v_n^2 = \frac{1}{4\pi\varepsilon_0} \frac{e^2}{r_n} \quad \text{--- (III)}$$

(A) Radius of an orbit of Hydrogen atom :-

According Bohr's Postulates of quantization of angular momentum.

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

$$\boxed{v_n = \frac{nh}{2\pi m r_n}} \quad \text{—— (IV)}$$

Using equation (IV) in (III)

$$m\left[\frac{nh}{2\pi m r_n}\right]^2 = \frac{e^2}{4\pi\varepsilon_0 r_n}$$

$$m\frac{n^2h^2}{4\pi^2 m^2 r_n^{\cancel{2}}} = \frac{e^2}{4\pi\varepsilon_0 (\cancel{r_n})}$$

$$\frac{\cancel{m}\, n^2h^2}{\cancel{4}\pi^{\cancel{2}} m^{\cancel{2}}} \times \frac{\cancel{4}\cancel{\pi}\varepsilon_0}{e^2} = r_n$$

$$\boxed{r_n = \frac{n^2h^2\varepsilon_0}{\pi m e^2}} \quad \text{—— (V)}$$

Since $\dfrac{h^2\varepsilon_0}{\pi m e^2} = \text{Constant } (k)$

$\therefore \boxed{r_n = kn^2}$ $\boxed{r_n \propto n^2}$

Thus Radius of an orbit is directly proportional to the square of the principal quantum number (n) of the orbit. [ie n = 1, 2, 3, 4 ----]

$r_1 : r_2 : r_3 : r_4 ---- = 1 : 4 : 9 : 16 : ---$

Bohr's Radius:- The Radius of the innermost orbit $(n=1)$ of an electron in hydrogen atom is called Bohr's Radius. It is denoted by (r_0).

$h = 6.625 \times 10^{-34}\ Js$, $\varepsilon_0 = 8.854 \times 10^{-12}\ m^{-2}c^2$

$m = 9.1 \times 10^{-31}\ kg$ and $e = 1.6 \times 10^{-19}\ c$; $n = 1$

$$r_0 = \frac{(1)^2 \times (6.625 \times 10^{-34})^2 (8.854 \times 10^{-12})}{(3.14)(9.1 \times 10^{-31})(1.6 \times 10^{-19})}$$

$$r_0 = 5.29 \times 10^{-11}\ m$$

$$r_0 = 0.529 \times 10^{-10}\ m$$

$$\boxed{r_0 = 0.529\ \mathring{A}}$$

(13) **Speed of an electron in an orbit of hydrogen atom:**

We know that formula for the velocity of an electron

$$V = n\frac{h}{2\pi m r}$$

Put the value of (r) from equ. (V)

$$v = \frac{nh}{2\pi m\left(\frac{n^2h^2\varepsilon_0}{\pi m e^2}\right)}$$

$$v = \frac{\cancel{n}\cancel{h}(\cancel{\pi}\cancel{m}e^2)}{2\cancel{\pi}\cancel{m}(n^{\cancel{2}}h^{\cancel{2}}\varepsilon_0)} = \left(\frac{e^2}{2nh\varepsilon_0}\right)$$

$$\boxed{v_n = \frac{e^2}{2h\varepsilon_0 n}} \text{ — (VI)}$$

$$\left[\text{Here } \frac{e^2}{2h\varepsilon_0} = k \text{ (Constant)}\right]$$

$$v_n = k\,\frac{1}{n}$$

$$\boxed{v_n \propto \frac{1}{n}}$$

Speed of an electron in an orbit is inversely proportional to the principal quantum number (n).

Relation between the speed of electron in first orbit of hydrogen and speed of light in vaccum:-

We know that

$$v_n = \frac{e^2}{2hn\varepsilon_0} \text{ — (VII)}$$

~~Speed light in vaccum~~

$$v_c = \left(\frac{e^2}{2\varepsilon_0 ch}\right)\frac{c}{n}$$

$$v_c = \alpha\left(\frac{c}{n}\right)$$

$$\alpha = \frac{e^2}{2\varepsilon_0 ch} = \frac{(1.6\times10^{-19})^2}{2(8.854\times10^{-12}\times3\times10^{8}\times6.625\times10^{-34}}$$

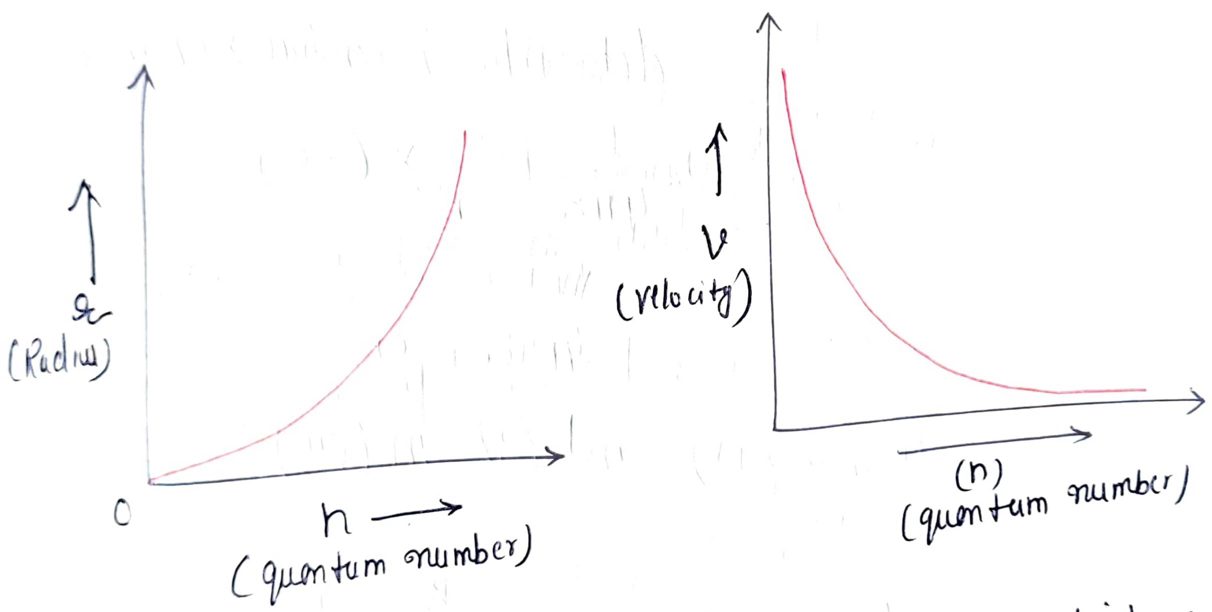
$$\alpha = \frac{1}{137}$$

$$V_n = \frac{c}{137(n)}$$

Here $\boxed{n=1}$

$$\boxed{V_1 = \frac{c}{137}}$$

Thus speed of electron in the first orbit of hydrogen is $\frac{1}{137}$ times the speed of light in vaccum.



(C) Energy of Electron in Stationary orbits :-

The total Energy of an electron in n^{th} orbit is the sum of its kinetic Energy and Potential energy in that orbit.

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

$$\text{Kinetic Energy (K.E.)} = \frac{1}{2} m v_n^2$$

$$= \frac{1}{2} \left(\frac{e^2}{4\pi \varepsilon_0 r_n} \right) \quad [\text{from (III)}]$$

$$\left[K.E. = \frac{1}{2} \frac{e^2}{4\pi \varepsilon_0 r_n} \right] \text{——(ix)}$$

$$\text{P.E (Potential Energy)} =$$

$$\text{electrostatic Potential} \times \text{Charge}$$

$$= \frac{1}{4\pi \varepsilon_0} \frac{e}{r_n} \times (-e)$$

$$= -\frac{1}{4\pi \varepsilon_0} \frac{e^2}{r_n} \text{——(x)}$$

Using (ix) and (x) in (VIII)

$$E_n = \frac{e^2}{8\pi \varepsilon_0 r_n} + \left(-\frac{1}{4\pi \varepsilon_0} \frac{e^2}{r_n} \right)$$

$$E_n = \frac{e^2}{4\pi \varepsilon_0 r_n} \left[\frac{1}{2} - 1 \right]$$

$$\boxed{E_n = -\frac{e^2}{8\pi \varepsilon_0 r_n}}$$

using $\left(r_n = \frac{n^2 h^2 \varepsilon_0}{\pi m e^2}\right)$

$$E_n = \frac{-e^2}{8\pi\varepsilon_0 \left(\frac{n^2 h^2 \varepsilon_0}{\pi m e^2}\right)}$$

$$E_n = \frac{-e^4 \cancel{\pi} m}{8\cancel{\pi}\varepsilon_0 (n^2 h^2 \varepsilon_0)}$$

$$\boxed{E_n = \frac{-me^4}{8h^2\varepsilon_0^2 n^2}}$$

$\varepsilon_0 = 8.854 \times 10^{-12}\ N^{-1} m^{-2} C^2$

$m = 9.1 \times 10^{-31}\ kg$

$e = 1.67 \times 10^{-19}\ C$

$h = 6.63 \times 10^{-34}\ Js$

$$E_n = \frac{-(9.1 \times 10^{-31})(1.67 \times 10^{-19})^4}{8(6.63 \times 10^{-34})^2 (8.854 \times 10^{-12})(n^2)}$$

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

Thus, total energy of eletron in a orbit is (-ve)
(-ve) energy tells that eletron and nucleus finally form an attractive system. (bounded system).

(D) Frequency (ν_n) of an electron in n^{th} orbit of Hydrogen atom :-

we know that
Relation between linear velocity and angular velocity

$$V_n = r_n \omega \qquad [\omega = 2\pi \nu_n]$$

velocity → $V_n = r_n (2\pi \nu_n)$ frequency

$$\nu_n = \frac{v_n}{r_n(2\pi)} \quad \text{—— (a)}$$

we know $r_n = \frac{n^2 h^2 \varepsilon_0}{\pi m e^2}$ and $v_n = \frac{e^2}{2h\varepsilon_0 n}$

Put in equation (a)

$$\nu_n = \frac{e^2}{2h\varepsilon_0 n} \times \frac{\cancel{\pi} m e^2}{n^2 h^2 \varepsilon_0} \times \frac{1}{2\cancel{\pi}}$$

$$\nu_n = \frac{m e^4}{4 h^3 \varepsilon_0^2 n^2}$$

Since $\frac{m e^4}{4 h^3 \varepsilon_0^2 n^2} = K$ (Constant)

$$\nu_n = K \frac{1}{n^3} \Rightarrow \boxed{\nu_n \propto \frac{1}{n^3}}$$

r is the radius of the orbit, and
h is Planck's constant.

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

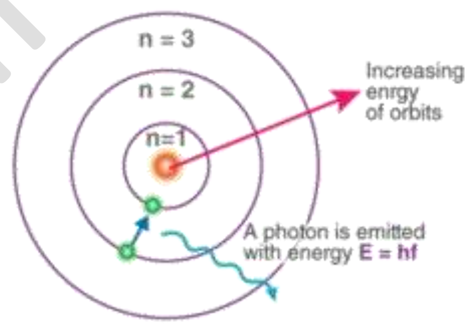
n is an integer called the principal quantum number, and it specifies the energy level or shell.

c) Radiation Absorption and Emission:

- Electrons in stable orbits neither emit nor absorb energy. They remain in these orbits without any radiation.

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

E_n is the energy of the electron in the nth orbit,
k is Coulomb's constant,
e is the charge of the electron,
Z is the atomic number,
h is Planck's constant, and
n is the principal quantum number.



d) Energy Changes in Transitions:

- When an electron transitions from a higher energy orbit **(n_2)** to a lower energy orbit **(n_1),** a photon is emitted or absorbed, and the frequency of the emitted or absorbed radiation is related to the energy difference between the two orbits.

$$hv = E_i + E_f$$

- ***ΔE*** is the energy change,
- ***h*** is Planck's constant,
- ***v*** is the frequency of the emitted or absorbed radiation.

e) Stationary Orbits:

- Electrons in stable orbits are in a "stationary" state, meaning they do not radiate energy. These orbits are also called "quantized" or "allowed" orbits.

f) Quantum of Action:

- The product of Planck's constant **(*h*)** and the frequency **(*v*)** is a multiple of **2*π*,** representing a quantum of action.

$$hv = 2\pi\hbar$$

- ***ħ*** is the reduced Planck's constant, equal to $h/_{2\pi}$.

Energy level of Hydrogen atom: -

The energy of the hydrogen atom in n^{th} state: -

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

a) When **n = 1** i.e., the electron is revolving in the innermost orbit

E_1 = -13.6 eV.

This is the lowest energy state of hydrogen atom and said to be in the **ground state.**

b) When **n = 2** i.e., the electron is revolving in the second orbit

E_2 = -3.4 eV.

This is the first excited state energy of hydrogen atom.

c) When **n = 3** i.e., the electron is revolving in the Third orbit

E_2 = -1.51 eV.

This is the second excited state energy of hydrogen atom.

d) When **n = ꝏ** i.e., the electron is revolving in the Third orbit

$$\boldsymbol{E_\varpi = \frac{-13.6}{\varpi^2} eV = 0}$$

The electron is completely free from the attraction.

The line spectra of Hydrogen atom :-

According to Bohr's Model

Energy is Radiated in the form of a photon when the eletron Returns from higher energy state to the lower state.

n_i = Initial energy state
n_f = final energy state.

$$h\nu = E_{n_i} - E_{n_f}$$

We know that $E_n = \dfrac{-me^4}{8h^2\varepsilon_0^2 n^2}$

$$h\nu = \frac{-me^4}{8h^2\varepsilon_0 n_i^2} - \left(\frac{-me^4}{8h^2\varepsilon_0^2 n_f^2}\right)$$

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

$$\nu = \frac{me^4}{8h^3\varepsilon_0}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right] \quad -(1)$$

$$C = \nu\lambda \Rightarrow \nu = \frac{C}{\lambda}$$

$$\frac{C}{\lambda} = \frac{me^4}{8h^3\varepsilon_0}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$$

$$\frac{1}{\lambda} = \frac{me^4}{8h^3 C\varepsilon_0}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$$

where $R = \frac{me^4}{8\varepsilon_0^2 Ch^3}$ [Rydberg Constant]

$$\bar{\nu} = \frac{1}{\lambda} \text{ (wave number)}$$

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

Various Spectral Series :-

when electron jumps from higher energy state to the lower energy state (orbit) in the hydrogen atom, Radiate energy of different wavelength and frequency. Known as spectral line.

(1) Lyman Series :-

when electron jump from any outer oRbit to first orbit form a ~~spectral~~ spectral series known as Lyman series.

$$\text{ie } \bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{1^2} - \frac{1}{n^2}\right] \qquad n = 2, 3, 4 \ldots$$

if $n = 2$ [Longest wavelength]

$$\frac{1}{\lambda} = R\left[\frac{1}{1} - \frac{1}{4}\right] = \frac{3R}{4}$$

$$\lambda_{max} = \frac{4}{3R} = \frac{4}{3\times(1.097\times10^7)} = 1.216\times10^{-7} m$$

$$\lambda_{max} = 1215.4 \text{ Å}$$

Shortest wavelength :-

When $(n = \infty)$

$$\frac{1}{\lambda_{min}} = R\left[1 - \frac{1}{\infty}\right]$$

$$\lambda_{min} = \frac{1}{R} = \frac{1}{1.097 \times 10^{-7}}$$

$$\boxed{\lambda_{min} = 911\,\mathring{A}}$$

This series lies in Ultra Voiolet Region.

(2) **Balmer series :-** When electron jump from any outer orbit to second orbit $(n_f = 2)$ form a spectral series known as Balmer series.

$$\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{2^2} - \frac{1}{n^2}\right] \quad [n = 3, 4, 5 \ldots]$$

Longest wavelength :- $\lambda_{max} = \frac{36}{5R} = 6559.2\,\mathring{A}$

Shortest wavelength :- $1/\lambda_{min} = \frac{R}{4}$

$$\lambda_{min} = 4/R = 3644\,\mathring{A}$$

This series lies in Visible Region.

③ **Paschen Series :-**

When the electron jump from any ~~out~~ outer orbit to the third orbit ($n_f = 3$) form a series known as Paschen series.

$$\bar{\nu} = \frac{1}{\lambda} = \frac{R}{\cancel{XR}}\left[\frac{1}{3^2} - \frac{1}{n^2}\right] \quad [n = 4, 5, 6 \ldots$$

(Longest wavelength) = $\frac{1}{\lambda} = \frac{R}{\cancel{XR}}\left[\frac{1}{9} - \frac{1}{16}\right]$

$$\lambda_{max} = \frac{144}{7R} = 18741\ \text{Å}$$

Shortest wavelength :- $\lambda_{min} = \frac{9}{R} = 8119\ \text{Å}$

This series lies in the Infra-Red

④ **Brackett series :-** The spectral lines emitted due to the transition of an electron from any outer orbit ($n = 5, 6, 7 \ldots$) to the fourth orbit ($n_f = 4$) form a spectral series known as Brackett series.

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

$$[n = 5, 6, 7 \ldots]$$

Longest wavelength :- $\frac{1}{\lambda_{max}} = R\left[\frac{1}{4^2} - \frac{1}{5^2}\right]$

$$\lambda_{max} = \frac{400}{9R} = 40589\,\mathring{A}$$

Shortest wavelength :-

$$\frac{1}{\lambda_{min}} = R\left[\frac{1}{4^2} - \frac{1}{\infty}\right] = \frac{R}{16}$$

$$\lambda_{min} = \frac{16}{R} = 14576\,\mathring{A}$$

This series lies in the Infra-Red Region.

(5) **Pfund Series :-** The spectral lines emitted due to the transition of an electron from any outer orbit ($n = 6, 7, 8 \ldots$) to the fifth orbit ($n_f = 5$) form a spectral series known as Pfund series.

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

$$(n = 6, 7, 8 \ldots)$$

Longest wavelength :-

$$\frac{1}{\lambda_{max}} = R\left[\frac{1}{5^2} - \frac{1}{6^2}\right]$$

$$= R\left[\frac{11}{900}\right]$$

$$\lambda_{max} = \frac{900}{11R} = 74536\,\mathring{A}$$

Shortest wavelength :-

$$\frac{1}{\lambda_{min}} = \frac{R}{25} = \frac{25}{R} = 22775\,\mathring{A}$$

This series lies in Infra-Red Region