

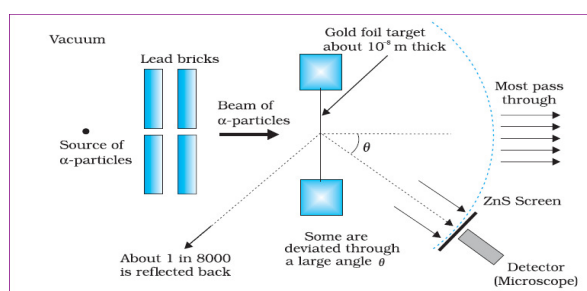
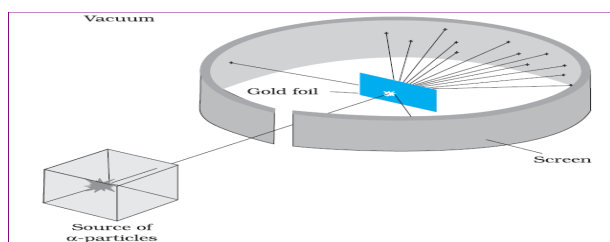
ATOMS

Thomson's model of atom

- The first model of atom was proposed by J. J. Thomson in 1898.
- According to this model, the positive charge of the atom is uniformly distributed throughout the volume of the atom and the negatively charged electrons are embedded in it like seeds in a watermelon.
- The total positive charge is balanced by the negative charges of electrons. As a result of this atom is electrically neutral
- This model was called **plum pudding model of the atom**.

ALPHA-PARTICLE SCATTERING EXPERIMENT

- This experiment was performed by H. Geiger and E. Marsden at the suggestion of Ernst Rutherford.

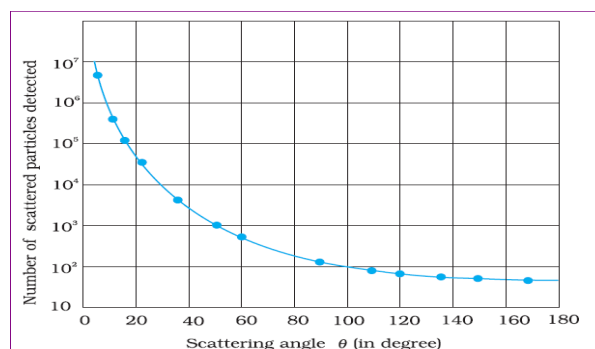


- Alpha-particles were collimated into a narrow beam by passing through lead bricks.
- The beam was allowed to fall on a thin foil of gold of thickness 2.1×10^{-7} m.
- The scattered alpha-particles were observed through a rotatable detector consisting of zinc sulphide screen and a microscope.
- The scattered alpha-particles on striking the screen produced brief light flashes or scintillations.

Variation of no. of α – particles scattered with scattering angle

- As the scattering angle increases, the number of α -particles scattered decreases.

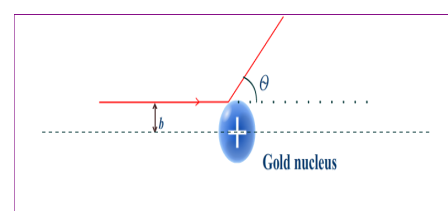
Graph of the total number of α -particles scattered at different angles, in a given interval of time



OBSERVATIONS

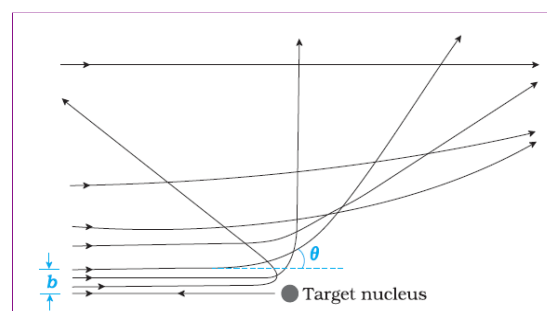
- Most of the α -particles were passed through the gold foil.
- Some were scattered through an angle and some of them in large angles.
- A very few were reflected back.

The impact parameter



- The perpendicular distance of the initial velocity vector of the α -particle from the centre of the nucleus.
- For small impact parameter- large scattering
- Large impact parameter- small scattering

Alpha-particle trajectory



- The trajectory traced by an α -particle depends on the impact parameter, b of collision.

Rutherford's planetary model of atom (nuclear model of the atom)

- An atom may be regarded as a sphere of diameter 10^{-10}m .
- Most of the mass and all the positive charges of the atom are concentrated in a very small central core (of diameter about 10^{-14}m) called the nucleus.
- The nucleus is surrounded by electrons.
- The electrons are spread over the remaining part of the atom, leaving plenty of empty space in the atom.
- As the atom is electrically neutral, the total positive charge on the nucleus is equal to the total negative charge of the electrons in the atom.
- Electrons are revolving round the nucleus in circular orbits.
- The necessary centripetal force for the revolution of electrons is provided by the electrostatic force of attraction between the electron and the nucleus.

Magnitude of repulsive force between α -particle and gold nucleus

- The magnitude of the force is

$$F = \frac{1}{4\pi\epsilon_0} \frac{(2e)(Ze)}{r^2}$$

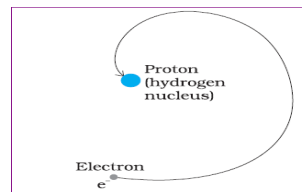
- where r is the distance between the α -particle and the nucleus

Advantages of Rutherford Model

- Large angle scattering of alpha particles through thin foils could be explained.
- The classification of elements in the periodic table on the basis of their atomic number, instead of atomic weight, was justified.

Limitations of Rutherford Model

- Rutherford's model fails to account for the stability of the atom.
- The energy of an accelerating electron should continuously decrease and the electron would spiral inward and eventually fall into the nucleus.



- Rutherford's model does not explain the line spectra of atoms.

Electron orbits

- The electrostatic force of attraction, F_e between the revolving electrons and the nucleus provides the centripetal force (F_c) to keep them in their orbits.
- For a dynamically stable orbit in a hydrogen atom

$$\frac{F_e}{r} = \frac{F_c}{r} \Rightarrow \frac{mv^2}{r} = \frac{1}{4\pi\epsilon_0} \frac{e^2}{r^2}$$

- Thus the relation between the orbit radius and the electron velocity is

$$r = \frac{e^2}{4\pi\epsilon_0 mv^2}$$

Total energy of an electron in an orbit

- The kinetic energy (K) and electrostatic potential energy (U) of the electron in hydrogen atom are

$$K = \frac{1}{2}mv^2 = \frac{e^2}{8\pi\epsilon_0 r} \quad \text{and} \quad U = -\frac{e^2}{4\pi\epsilon_0 r}$$

- Thus the total energy E of the electron in a hydrogen atom is

$$E = K + U = \frac{e^2}{8\pi\epsilon_0 r} - \frac{e^2}{4\pi\epsilon_0 r} = -\frac{e^2}{8\pi\epsilon_0 r}$$

- The total energy of the electron is negative.
- That is the electron is bound to the nucleus.

BOHR MODEL OF THE HYDROGEN ATOM

- Bohr combined classical and early quantum concepts and gave his theory in the form of three postulates:

Postulate I

- **Electrons in an atom can revolve in certain stable orbits without radiating energy.**
- According to this postulate, each atom has certain definite stable states in which it can exist, and each possible state has definite total energy.
- These are called the **stationary states** of the atom

Postulate II

- **The electron revolves around the nucleus only in those orbits for which the angular momentum is some integral multiple of $h/2\pi$ where h is the Planck's constant ($= 6.6 \times 10^{-34}$ J s).**
- Thus the angular momentum (L) of the orbiting electron is quantised.
- That is **$L = nh/2\pi$, where $n = 1, 2, 3, \dots$, is the principal quantum number.**

Postulate III



- **An electron might make a transition from one of its specified non-radiating orbits to another of lower energy.**
- A photon is emitted having energy equal to the energy difference between the initial and final states.
- The frequency of the emitted photon is then given by

$$h\nu = E_i - E_f$$

Radii of Bohr's Stationary orbits:

- The centripetal force for the revolution of electrons round the nucleus is provided by the electrostatic force of attraction between the nucleus and the electron.
- Thus

$$\frac{mv^2}{r} = \frac{1}{4\pi\epsilon_0} \cdot \frac{e^2}{r^2}$$

- Therefore

$$mv^2 = \frac{1}{4\pi\epsilon_0} \frac{e^2}{r}$$

- From Bohr's II postulate the angular momentum $L = mvr$, is given by

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

- Thus

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

- Therefore

$$m \left(\frac{nh}{2\pi mr} \right)^2 = \frac{1}{4\pi\epsilon_0} \frac{e^2}{r} \quad m \frac{n^2 h^2}{4\pi^2 m^2 r^2} = \frac{1}{4\pi\epsilon_0} \frac{e^2}{r}$$

- That is

$$\frac{n^2 h^2}{\pi m r} = \frac{e^2}{\epsilon_0}$$

- The radius is given by

$$r = \frac{\epsilon_0 n^2 h^2}{\pi m e^2}$$

- In general

$$r_n = \left(\frac{\epsilon_0 h^2}{\pi m e^2} \right) n^2$$

- Thus

$$r_n \propto n^2$$

- The radii of the stationary orbits are in the ratio, $1^2:2^2:3^2:\dots$ or $1:4:9:\dots$
- The stationary orbits are not equally spaced.

Bohr Radius

- The radius of the lowest orbit ($n=1$) is called Bohr radius.
- The Bohr radius is given by

$$a_0 = \frac{h^2 \epsilon_0}{\pi m e^2}$$

- Substituting the values we get

$$a_0 = 5.29 \times 10^{-11} \text{ m}$$

- Thus the radius of n^{th} orbit becomes:

$$r_n = a_0 n^2$$

Velocity of electrons in an orbit

- We have

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

- But

$$r = \frac{\epsilon_0 n^2 h^2}{\pi m e^2}$$

- Therefore

$$v = \frac{nh}{2\pi m \left(\frac{\epsilon_0 h^2}{\pi m e^2} \right) n^2}$$

- That is

$$v = \frac{e^2}{2\epsilon_0 nh}$$

- In general

$$v_n = \frac{e^2}{2\epsilon_0 nh}$$

Total energy of an orbiting electron

Kinetic energy

- For an orbiting electron, we have

$$\frac{mv^2}{r} = \frac{1}{4\pi\epsilon_0} \cdot \frac{e^2}{r^2}$$

- Or

$$mv^2 = \frac{1}{4\pi\epsilon_0} \frac{e^2}{r}$$

- Thus the kinetic energy is given by

$$K = \frac{1}{2}mv^2 = \frac{e^2}{8\pi\epsilon_0 r}$$

Potential energy

- The electrostatic potential energy of an orbital electron is given by

$$U = -\frac{e^2}{4\pi\epsilon_0 r}$$

Total energy

- Total energy is given by

$$E = K + U = \frac{e^2}{8\pi\epsilon_0 r} - \frac{e^2}{4\pi\epsilon_0 r}$$

- That is

$$E = \frac{-e^2}{8\pi\epsilon_0 r}$$

- But, we have

$$r = \frac{\epsilon_0 n^2 h^2}{\pi m e^2}$$

- Thus

$$E = \frac{-e^2}{8\pi\epsilon_0 \times \frac{\epsilon_0 h^2 n^2}{\pi m e^2}}$$

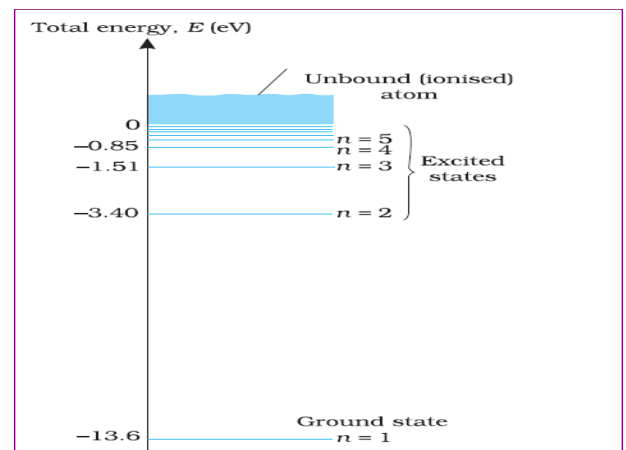
- In general

$$E_n = \frac{-me^4}{8\epsilon_0^2 h^2 n^2}$$

- Substituting the values we get

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

Energy level diagram of hydrogen atom



Excitation energy

- Excitation energy is the **energy required to excite an electron from its ground state to an excited state.**
- First excitation energy of hydrogen atom required to excite the electron from $n = 1$ to $n = 2$ orbit of hydrogen atom. That is $(-3.4) - (-13.6) = 10.2 \text{ eV}$.

Excitation potential

- Excitation potential of an excited state is the potential difference through which electron in an atom has to be accelerated so as to excite it from its ground state to the given excited state.
- The **first excitation potential of H atom** is 10.2V.

Ionization energy

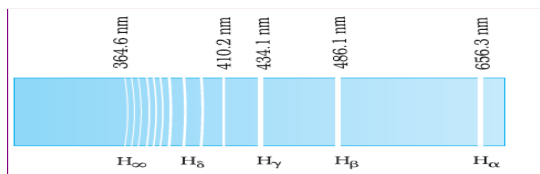
- Ionisation energy is the energy required to take an electron completely out of the atom.

- The ionization energy of hydrogen atom is 13.6 eV.

ATOMIC SPECTRA

Emission Spectrum

- When an atomic gas is excited at low pressure, the emitted radiation has a spectrum contains only certain wavelengths- emission spectrum.
- Emission line spectrum consists of **bright lines on a dark background**

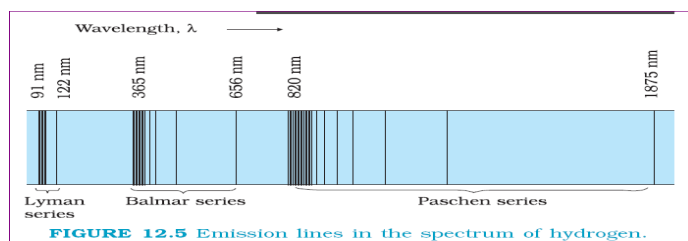


Absorption spectrum

- When white light is passed through certain gas, some dark lines are seen in the transmitted spectrum- absorption spectrum
- Absorption line spectrum has **dark lines on bright background**

Spectral series

- The spacing between lines within certain sets of the hydrogen spectrum decreases in a regular way. Each of these sets is called a **spectral series**.



- As the wavelength decreases, the lines appear closer together and are weaker in intensity.
- The first spectral series was observed by Johann Jakob Balmer in the visible region of the hydrogen spectrum and is known as **Balmer series**.

Balmer formula



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- The empirical formula for the observed wavelengths of Balmer series is

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

- Where λ is the wavelength, R is a constant called the Rydberg constant, and n may have integral values 3, 4, 5, etc.
- The value of R is $1.097 \times 10^7 \text{ m}^{-1}$.

Rydberg formula

- Using Bohr's postulate, the energy for a transition from n_1 to n_2 is

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

- That is

$$h\nu = \frac{-me^4}{8\epsilon_0^2 h^2 n_2^2} - \frac{-me^4}{8\epsilon_0^2 h^2 n_1^2}$$

$$= \frac{me^4}{8\epsilon_0^2 h^2} \left[\frac{1}{n_1^2} - \frac{1}{n_2^2} \right]$$

- Or

$$\nu = \frac{me^4}{8\epsilon_0^2 h^3} \left[\frac{1}{n_1^2} - \frac{1}{n_2^2} \right]$$

- Since $\nu = c/\lambda$, we get

$$\frac{c}{\lambda} = \frac{me^4}{8\epsilon_0^2 h^3} \left[\frac{1}{n_1^2} - \frac{1}{n_2^2} \right]$$

- Or

$$\frac{1}{\lambda} = \frac{me^4}{8\epsilon_0^2 h^3 c} \left[\frac{1}{n_1^2} - \frac{1}{n_2^2} \right]$$

- Where $1/\lambda$ is the **wave number**.
- Also

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

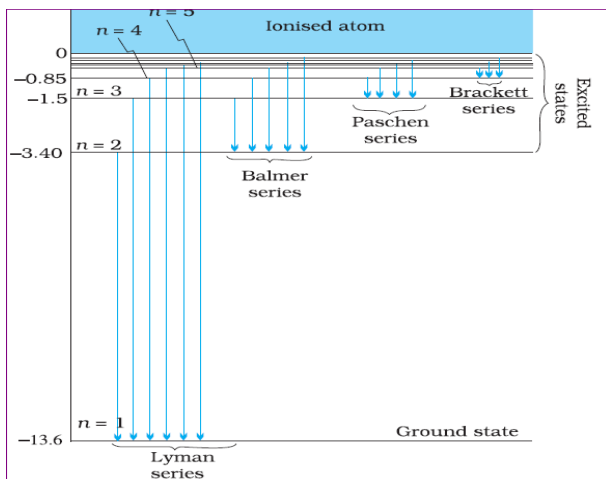
- This formula is called **Rydberg formula**.
- Where R – Rydberg constant

$$R = \frac{me^4}{8\epsilon_0^2 h^3 c}$$

- The value of R is $1.097 \times 10^7 \text{ m}^{-1}$.

Spectral lines of Hydrogen Spectrum

- The spectral series of hydrogen atom is shown in figure.



Lyman Series

- The spectral line of this series corresponds to the transition of an electron from some higher energy state to the innermost orbit ($n = 1$).
- For Lyman series, $n_1 = 1$ and $n_2 = 2, 3, 4, \dots$
- Thus

$$\frac{1}{\lambda} = R \left(\frac{1}{1^2} - \frac{1}{n^2} \right) \quad n = 2, 3, 4, \dots$$

- Lyman series lies in the **ultra violet** region.

Balmer Series

- The spectral lines of this series correspond to the transition of an electron from some higher state to an orbit having $n = 2$.
- For Balmer series $n_1 = 2$, $n_2 = 3, 4, 5, \dots$
- Thus

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

- The first four lines of the Balmer series lie in the **visible region** of the spectrum and the rest of the series goes to the **u.v. region**.

Paschen Series

- For Paschen series $n_1 = 3$, $n_2 = 4, 5, 6, \dots$
- Thus

$$\frac{1}{\lambda} = R \left(\frac{1}{3^2} - \frac{1}{n^2} \right) \quad n = 4, 5, 6, \dots$$

- Paschen series lies in the **infrared region**.

Brackett Series

- For Bracket series, $n_1 = 4$, $n_2 = 5, 6, \dots$
- Thus

$$\frac{1}{\lambda} = R \left(\frac{1}{4^2} - \frac{1}{n^2} \right) \quad n = 5, 6, 7, \dots$$

- Brackett series lies in the **infrared region**.

Pfund Series

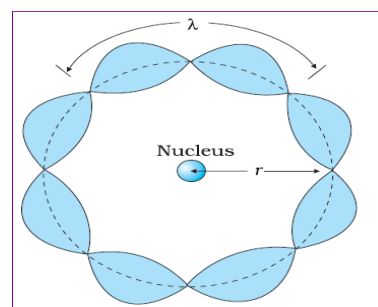
- For Pfund series $n_1 = 5$ and $n_2 = 6, 7, \dots$
- Thus

$$\frac{1}{\lambda} = R \left(\frac{1}{5^2} - \frac{1}{n^2} \right) \quad n = 6, 7, 8, \dots$$

- This series is in the **far infrared region** of the spectrum.

DE BROGLIE'S EXPLANATION OF BOHR'S SECOND POSTULATE OF QUANTISATION

- Louis de Broglie argued that the electron in its circular orbit, as proposed by Bohr, must be seen as a particle wave.
- In analogy to waves travelling on a string, particle waves too can lead to standing waves under resonant conditions.
- In a string, standing waves are formed when the total distance travelled by a wave down the string and back is one wavelength, two wavelengths, or any integral number of wavelengths.



- For an electron moving in n th circular orbit of radius r_n , the total distance is the circumference of the orbit, $2\pi r_n$.

$$2\pi r_n = n\lambda, \quad n = 1, 2, 3, \dots$$

- But we have $\lambda = h/p$,

$$2\pi r_n = n h / m v_n \quad \text{or} \quad m v_n r_n = n h / 2\pi$$

- This is the quantum condition proposed by Bohr for the angular momentum of the electron

Limitations of Bohr Model

- Bohr's theory is applicable only to single electron atoms.
- This theory gives no idea about relative intensities of spectral lines.
- Could not explain the fine structure of hydrogen spectrum.