BOHR MODEL OF THE HYDROGEN ATOM

In adequacy of Rutherford Model

The nucleus-electron system being charged objects, interact by Coulomb’s Law of force.

Electrons revolve in circular orbit experience a centripetal acceleration.

Drawback 1

According to **EM theory**, an **accelerated** electric charge must emit **EM radiations**. If the radiation is continuously emitted from the electron, its energy must **decrease** and electron should finally **fall into the nucleus**. That is atom is **unstable** as per the Rutherford model. But most of the atoms are **stable**

Drawback 2

According to the classical electromagnetic theory, the **frequency** of the electromagnetic waves emitted by the revolving electrons is **equal** to the **frequency of revolution**.

As the electrons spiral inwards, their **angular velocities** and hence their **frequencies** would change **continuously**, and so will the frequency of the light emitted. Thus, they would emit a **continuous spectrum**, in contradiction to the **line spectrum** actually observed.

Bohr’s first postulate

An electron in an atom could revolve in certain stable orbits without the emission of radiant energy.

That is an 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

Bohr’s second postulate

This postulate states that the electron revolves around the nucleus only in those orbits for which the **angular momentum** is some **integral** multiple of h/2π, where h is the Planck’s constant.

Thus the angular momentum (L)

of the orbiting electron is quantised

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

Bohr’s third postulate

It states that an electron might make a transition from one of its specified **non-radiating** orbits to another of lower energy. When it does so, 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

le $\mathbf{h\,v = E_i - E_f}$

where E i and E f are the energies of the initial and final states and $E_i > E_f$

Energy of an electron

The angular momentum L is given by **$L = mvr$**

According to the Bohr second postulates **$L_n = mv_n r_n$**

where n is an integer, **r_n** is the radius of n th possible orbit and **v_n** is the speed of moving electron in the **n^{th}** orbit.

n = 1, 2, 3 ..., which is called the principal **quantum number** of the orbit

The necessary centripetal force is provided by Coulomb’s attraction. For a Hydrogen atom

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

Ie

$$v^2 = \frac{e^2}{4\pi\epsilon_0 mr}$$

For n^{th} orbit

$$v_n = \frac{e}{\sqrt{4\pi\epsilon_0 m r_n}}$$

Using Bohr's second postulates

$$L_n = m v_n r_n = m \frac{e}{\sqrt{4\pi\epsilon_0 m r_n}} r_n = \frac{nh}{2\pi}$$

Squaring on both sides, we get

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

It is the radius of the n^{th} orbit.

If we put n =1, we get Bohr radius as

$$r_1 = a_0 = \frac{h^2 \epsilon_0}{\pi m e^2}$$

Energy of an orbiting electron

We have

$$E_n = -\frac{e^2}{8\pi\epsilon_0 r_n}$$

And

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

On substitution

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

Or

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

It is the total energy

Substituting the values, we get

$$E_n = -\frac{2.18\,x\,10^{-18}}{n^2} J = -\frac{13.6}{n^2} eV$$

Where **1 eV = 1.6 x 10 $^{-19}$** Joule

The **negative** sign of the total energy of an electron moving in an orbit means that the electron is **bound** with the **nucleus**

Ionisation

Energy will thus be required to remove the electron from the hydrogen atom to a distance infinitely far away from its nucleus

Energy Levels

The energy of an atom is the least (largest negative value) when its electron is revolving in an orbit closest to the nucleus i.e., the one for which $n = 1$.

That is E_1 is -13.6 eV

Ionisation energy

The minimum energy required to free the electron from the ground state of the hydrogen atom is 13.6 eV. It is called the ionisation energy of the hydrogen atom.

Excited State

At room temperature, most of the hydrogen atoms are in **ground** state. When a hydrogen atom **receives energy** by processes such as electron **collisions**, the atom may acquire sufficient energy to raise the electron to higher energy states. The atom is then said to be in an **excited state**.

First excited state

$E_2 – E_1$ = –3.40 eV – (–13.6) eV = **10.2 eV** amount of energy is required to excite an electron from ground state to first excited state

Second excited state

$E_3 - E_1 = 12.09$ eV, amount energy is required to excite the hydrogen atom from its ground state (n = 1) to second excited state (n = 3)

Energy Level Diagram