Atoms

1. Distance of closest approach. When an α-particle of mass m and velocity v moves directly towards a nucleus of atomic number Z, its distance of closest approach is given by

$$r_0 = \frac{2kZe^2}{E} = \frac{4kZe^2}{mv^2}$$

where $E = \frac{1}{2}mv^2$

and $k = \frac{1}{4\pi\varepsilon_0} = 9 \times 10^9 Nm^2C^{-2}$.

2. Bohr's atom model. This model is also called planetary model of an atom and is based on following postulates :

(i) Nuclear concept. An atom consists of a small massive central core called nucleus around which planetary electrons revolve. The centripetal force required for their revolution is provided by the electrostatic attraction between the electrons and the nucleus.

(ii) Quantum condition. Of all the possible circular orbits allowed by the classical theory, the electrons are permitted to circulate only in such orbits in which the angular momentum of an electron is an integral multiple of h/2π, h being Planck's constant,

$$L = mvr = \frac{nh}{2\pi}, n = 1, 2, 3, \ldots$$

where n is called principal quantum number.

(iii) Stationary orbits. While revolving in the permissible orbits, an electron does not radiate energy. These non-radiating orbits are called stationary orbits.

(iv) Frequency condition. An atom can emit or absorb radiation in the form of discrete energy photons only when an electron jumps from a higher to a lower orbit or from a lower to a higher orbit. If E_1 and E_2 are the energies associated with these permitted orbits, then the frequency v of the emitted or absorbed radiation is given by

$$h\nu = E_2 - E_1$$

3. Bohr's theory of hydrogen atom. An electron having charge - e revolves with speed v in a circular orbit of radius r round the nucleus having charge + e. For a circular orbit,

$$\frac{mvr^2}{r} = k\,Ze^2/r$$

From quantisation of angular momentum,

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

On solving the above two equations, we get

Radius of nth orbit,

$$r = \frac{n^2h^2}{4\pi mk\,Ze^2}$$

Speed of electron in nth orbit,

$$v = \frac{2\pi ke^2}{nh} = \alpha\frac{c}{n} = \frac{1}{137}.\frac{c}{n}.$$

where $\alpha = \frac{2\pi ke^2}{nh}$ is fine structure constant.

Total energy of an electron in nth orbit is

E_n = K.E. + P.E. = $\frac{k\,Ze^2}{2r} - \frac{kZe^2}{r} = -\frac{kZe^2}{2r}$ or $E_n = -\frac{2\pi\,mk^2\,Z^2e^4}{n^2h^2}$

$= -\frac{Z^2\,Rhc}{n^2} = -\frac{13.6\,Z^2}{n^2}\,eV$

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

is the Rydberg's constant.

4. Spectral series of hydrogen atom. Whenever an electron makes a transition from a higher energy level n_2 to a lower energy level n_1, the difference of energy appears in the form of a photon of frequency v given by

$$v = \frac{2\pi^2mk^2Z^2e^4}{h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

Wave number,

$$\bar{v} = \frac{1}{\lambda} = \frac{\mathrm{v}}{c} = \frac{2\pi^2mk^2Z^2e^4}{\mathrm{ch}^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

or $\bar{v} = \frac{1}{\lambda} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

Different spectral series for hydrogen atom are as follows:

(i) Lyman Series. Here n_2 = 2, 3, 4, ... and n_1 = 1. This series lies in the ultraviolet region.

$$\bar{v} = \frac{1}{\lambda} = R\left[\frac{1}{1^2} - \frac{1}{n_2^2}\right], n_2 = 2,3,4, \ldots \ldots$$