Bohr's atomic theory:

- Bohr recognized the relationship between the nature of the series of spectral lines and the arrangement of electrons in the atom.
- Bohr applied Planck's quantum theory to the electrons revolve around the nucleus. He retained the basic concept of Rutherford's model of atom that electrons revolve round the positively charged nucleus.
- Bohr proposed his theory to explain the structure of atom.

The important postulates of his theory are:

- Electrons revolve around the nucleus with definite velocities in concentric circular orbits. These orbits are called stationary orbits as the energy of the electron remains constant. As long as the electron revolves in the same circular orbit it neither radiates nor absorbs energy.
- The angular momentum of the electron is quantized. The electronic motion is restricted to those orbits where the angular momentum of an electron is an integral multiple of $h/2\pi$ or $mvr = nh/2\pi$. This is called Bohr's quantum condition or quantization of angular momentum.
- Energy of the electron changes only when it moves from one orbit to another orbit.
- Energy is absorbed when an electron jumps from a lower orbit to a higher outer orbit.
- If electron is in 1s orbital, it can only absorb but cannot emit energy.
- Energy is released when an electron jumps from higher orbit to a lower orbit.
- The released or absorbed energy is equal to the difference between the energies of the two orbits.
- If E_2 is the energy of the electron in the outer orbit (n_2) and E_1 is the energy of the electron in the inner orbit (n_1), then $E_2 – E_1 = \Delta E = h\upsilon$.
- Where n is called principal quantum number and it represents the main energy level.
- It takes all positive and integral values 1, 2, 3, 4… etc.
- With the help of these postulates Bohr derived the expression for the radius of the circular orbit, energy of the electron in a circular orbit and velocity of the electron in a circular orbit.
- Bohr's theory could satisfactorily explain the formation of different series of lines in hydrogen spectrum.

- The wavelengths and the frequencies of the lines determined experimentally are in excellent agreement with those calculated by using Bohr's equation.

Radius of orbit:

- Hydrogen atom contains one proton in the nucleus and one electron revolving around the nucleus in a circular orbit of radius r.
- The electron maintains the same circular motion in given orbit as centripetal and centrifugal forces are equal in magnitude and opposite in direction.
- Centripetal force = centrifugal force (columbic forces of attraction provides necessary centripetal force)

  i.e. $\frac{-e^2}{r^2} = -\frac{mv^2}{r}$; $\frac{e^2}{r} = mv^2$

  According to Bohr's quantum condition;

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

  $v^2 = \frac{n^2h^2}{4\pi^2m^2r^2}$; $\frac{e^2}{r} = \frac{m \cdot n^2h^2}{4\pi^2m^2r^2}$; $e^2 = \frac{n^2h^2}{4\pi^2mr}$; $r = \frac{n^2h^2}{4\pi^2me^2}$

  The radius of the nth orbit is given by

  $r_n = n^2\left(\frac{h^2}{4\pi^2me^2}\right)$= 0.529 x 10^{-8} n^2cm

  Where h = Planck's constant; m = mass of electron; e = charge of electron; r_n = radius of nth orbit

- The radius of the first orbit of hydrogen atom is called Bohr's radius which is denoted by r_0.

  r_0 = 0.529 x 10^{-8} cm = 0.529 A^0

Energy of electron:

- The total energy of the electron in a stationary orbit is equal to sum of its kinetic and potential energies.

  Total energy of electron E = K.E + P.E.

  K.E. is always positive and P.E is always negative.

  K.E. is half to that of P.E. in magnitude.

$$\frac{1}{2}mv^2 - \frac{e}{r} = \frac{1}{2}\frac{e^2}{r} - \frac{e^2}{r} (\because mv^2 = \frac{e^2}{r}) = -\frac{1}{2}\frac{e^2}{r}$$

Energy of electron for single electron species is $E_n = \frac{2\pi^2 e^4 m}{h^2} x \frac{z^2}{n^2}$.

By substituting the value of r.

$E_n = -\frac{1}{2}\frac{e^2 4\pi^2 me^2}{n^2h^2}$; $E_n = -\frac{2\pi^2 e^4 m}{n^2h^2}$;

$E_n = -\frac{k}{n^2}$ (k is constant ; k = $\frac{2\pi^2 e^4 m}{h^2}$);

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

(or) $-\frac{2.18x10^{-11}}{n^2}$ergs/atom

(or) $-\frac{2.18x10^{-18}}{n^2}$J/atom

(or) $-\frac{313.6}{n^2}$kcal/mole

$-\frac{1312}{n^2}$ kJ/mole

- The energy of electron is negative in the atom.
- As the value of n increases energy increases.
- When n is infinity the value of E is zero.
- When n value decreases the energy of electron also decreases.

Rydberg constant (R):

- When an electron jumps from outer energy level (n_2) to inner energy level (n_1), energy is released.

  i.e. $E_2 - E_1 = \Delta E = \Delta h\upsilon$

  E_2 = energy of electron in higher orbit (n_2)

  E_1 = energy of electron in lower orbit (n_1)

  $E_2 - E_1 = \frac{-2\pi^2 e^4 m}{n_2^2 h^2} + \frac{2\pi^2 e^4 m}{n_1^2 h^2}$;

  $\Delta E = \frac{2\pi^2 e^4 m}{h^2}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$;

  $h\upsilon = \frac{2\pi^2 e^4 m}{h^2}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$;

  $\upsilon = \frac{2\pi^2 e^4 m}{h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$;

  Wave number $\bar{\upsilon} = \frac{1}{\lambda} = \frac{v}{c}$

$\therefore \upsilon = c\bar{\upsilon}$; $c\bar{\upsilon} = \frac{2\pi^2 e^4 m}{h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$; $\bar{\upsilon} = \frac{2\pi^2 e^4 m}{ch^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$; $\bar{\upsilon} = \frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

R = $\frac{2\pi^2 e^4 m}{ch^3}$= 1,09,681 cm^{-1}.

- This value Rydberg constant (R) calculated by Bohr as above is in good agreement with experimental value.

Velocity of element in the nth orbit:

- As per Bohr's quantum condition, mvr = n$\frac{h}{2\pi}$; $V_n = \frac{nh}{2\pi mr}$;

  Substituting 'r'; $V_n = \frac{2\pi e^2}{nh}$ (for 'H' atoms for any other single electron species; $V_n = \frac{2\pi e^2}{h} x \frac{z}{n}$

- Sub situating the values of constants,

  $V_n = \frac{2.188}{10^{-8}} x \frac{z}{n}$ cm/sec.

- Number of revolutions per second, made by electron in circular orbit is = $\frac{velocity}{circumference} = \frac{v}{2\pi r}$