Spectra, Hydrogen Spectrum, Bohr's Model

1. Sun light or light from an incandescent bulb gives a continuous spectrum.
2. When a gas or a vapour of a metal is kept in a discharge tube and higher potential is applied a line spectrum is formed. It is the characteristic of an Element.
3 Line spectra can be used in chemical analysis to identify unknown Element.
4. The spectra obtained from the energy emitted by the excited atoms are called **emission spectra.** They consist of bright lines on the dark background.
5. When white light is passed through a gas and the transmitted of light is allowed to fall on a photographic plate, the spectrum obtained is called absorption spectrum. In this dark lines appear on bright background.
6. For same element dark lines in the absorption spectrum and the bright lines in the emission spectrum occur at same wavelength.
7. The hydrogen spectrum is the simplest of all the atomic spectra.

   It contains a number of groups of lines. They can be classified into five series namely Lyman, Balmer, Paschen, Brackett and Pfund series.

Name Of Series	**n_1**	**n_2**	**Region**
Lyman series	1	2, 3, 4, 5..	Ultraviolet
Balmer series	2	3, 4, 5, 6..	Visible
Paschen series	3	4, 5, 6. 7	Near Infrared
Brackett series	4	5, 6, 7 8..	Infrared
Pfund series	5	6, 7, 8. 9	Far Infrared
Humphry series	6	7, 8, 9. 10	Far Infrared

8. The wavelength or wave number of various lines in the visible region can be expressed by an equation. $\bar{\nu} = \frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

9. The value of R = 1, 09,678 is valid only for the lines in the hydrogen spectrum.

   For a spectral line of any single electron species like He^{+}, Li^{2+} the value of $R = R_H \times Z^2$

10. The first line in each series is called H_α line, the second line is called H_β line, and third line is H_γ

   If n_2 is taken as infinity, the line is called **limiting line** in the series.

   For Balmer series, wave length of limiting line is

$$\bar{\nu}_\infty = \frac{1}{\lambda_\infty} = R\left[\frac{1}{2^2} - \frac{1}{\infty^2}\right] = \frac{R}{4} = 27,419 cm^{-1}$$

12. Maximum number of lines produced when an electron jumps from nth level to ground level $= \frac{n(n-1)}{2}$.

13. Maximum number of lines produced when an electron jumps from n_2 level to n_1 level $= \sum n_2 - n_1$.

14. Postulates of Bohr's Atomic Model.

   a) The electrons in an atom revolve round the nucleus in definite circular orbits or shells or energy levels.

   b) These are called stationary orbits as the Energy of electron remains constant in each orbit.

   c).The angular momentum of electron in an orbit is equal to integral multiple of $\frac{h}{2\pi}$.

i.e., $mvr = n\frac{h}{2\pi}$ where m = mass of electron,v = velocity of electron, r = radius of orbit

n = 1, 2, 3, 4 ...... and h = Planck's constant

d). Angular momentum is the product of moment of inertia (I) and angular velocity (ω). Angular momentum = $I \times \omega$

Since $I = m_e r^2$ and $\omega = v/r$ where v is the linear velocity angular momentum = $mr^2 \times v/r = mvr$

e) When an electron drops from a higher orbit to a lower orbit, energy is released. When an electron jumps from a lower orbit to a higher orbit, energy is absorbed. The absorbed or evolved energy is equal to the difference in energies of two orbits, which is equal to quanta.

$\Delta E = E_2 - E_1 = h\nu$

Where E_2 = Energy of higher orbit and E_1 = Energy of lower orbit, ν = Frequency

f) The force of attraction between the nucleus and the electron $= \frac{-Ze^2}{r^2}$

g) The centrifugal force of the electron due to revolving round the nucleus $= \frac{-mV^2}{r}$

h) Expression for the radius of Bohr's orbit $r = \frac{n^2h^2}{4\pi^2 m Ze^2}$ or

$r = 0.529 \times 10^{-8} \times n^2 \text{ cm} = 0.529 \times n^2 A^0$

i) Radius of orbits in H atom like ions $r = \frac{0.529 \times n^2}{Z} A^0$

j) Kinetic energy of electron $=\frac{1}{2}mV^2=\frac{Ze^2}{2r}$ & Potential energy of electron $=\frac{-Ze^2}{r}$

k) Total energy of electron $=KE+PE=\frac{Ze^2}{2r}-\frac{Ze^2}{r}=\frac{-Ze^2}{2r}$

l) Expression for the energy of Bohr's orbit $E=\frac{-2\pi^2 mZ^2e^4}{n^2h^2}$

m) As n value increases, kinetic energy decreases potential energy and the total energy increase.

15. Energy of orbits in hydrogen atom

$$E=\frac{-2.179\times10^{-11}}{n^2}\text{ergs}=\frac{-2.179\times10^{-18}}{n^2}\text{Joules}=\frac{-13.6}{n^2}\text{eV}=\frac{-313.6}{n^2}\text{K.cal/mole}$$

$$=\frac{-1312}{n^2}\text{KJ/mole}$$

$1\,eV=1.602\times10^{-19}\,J$

16. For Hydrogen like ions. $E=\frac{-2.179\times10^{-11}}{n^2}\times Z^2$ ergs

17. For. H atom and H like ions, Ionisation potential $=\frac{-E_1\times Z^2}{n^2}$.

18. Rydberg's constant $R=\frac{2\pi^2mZ^2e^4}{h^3C}=109680\,cm^{-1}\times Z^2$

19. Difference of energy between two Bohr orbits of hydrogen atom $\Delta E=Rhc\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]$

20. Velocity of electron in hydrogen atom $V=\frac{2\pi Ze^2}{nh}=\frac{2.188\times10^8}{n}$ cm/sec

21. Merits of Bohr's Model

a. It successfully explains the hydrogen spectrum and spectra of ions having one electron.

b. The experimental values of the energies and radii of possible orbits in hydrogen atom are in good agreement with that calculated on the basis of Bohr's theory.

c. It explains the stability of Atom by introducing stationary orbits.

22. Demerits of Bohr's Model

a. It failed to explain the multi electron spectra.

b. It failed to explain Zeeman Effect, Stark effect. And Fine spectrum of hydrogen atom.

- The splitting of spectral lines in a magnetic field is called **Zeeman Effect.**
- The splitting of spectral lines in an electric field is called **Stark Effect.**

c. It is in contradiction with Heisenberg uncertainity principle.