CHAPTER-2

ATOMIC STRUCTURE

→ At first, John Dalton proposed that atom is the smallest particle of matter and it cannot be divided.

→ But, due to advance research it was found that atom can be divided into sub-atomic particles.

Sub Atomic Particles

Particle	Mass	Charge	Discovered by	e/m ratio
Electron (e^-)	9.1×10^{-31} kg 9.1×10^{-28} g	-1.6×10^{-19} C	J.J. Thomson	1.76×10^{11} C/kg
Proton (p)	1.66×10^{-27} kg 1.66×10^{-24} g	$+1.6 \times 10^{-19}$ C	Rutherford	9.58×10^{7} C/kg
Neutron (n)	1.67×10^{-27} kg 1.67×10^{-24} g	0	James Chadwick	0

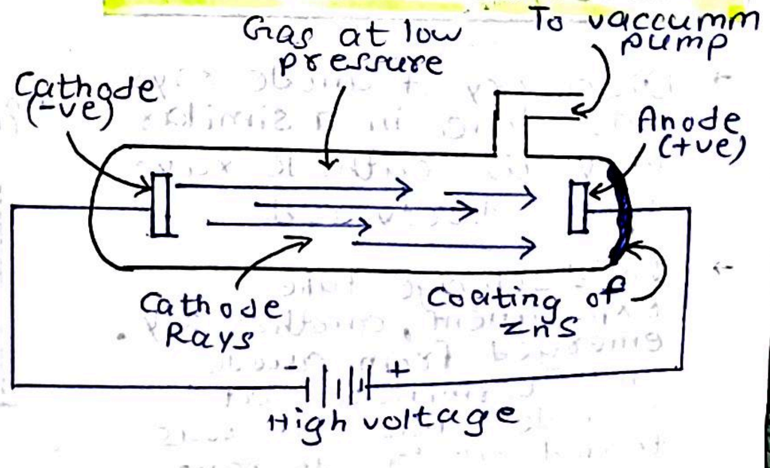
Discovery of Cathode Rays

→ Discovery of cathode ray is the result of Faraday's discharge tube experiment.

→ In his experiment, he passed current of very high voltage through a discharged glass tube and observed that a ray emerged out from cathode and terminates at anode.

→ He termed this ray as 'cathode ray' because it emerged from cathode (-ve electrode).

→ In this experiment, the voltage was very high (about 10000 V) and the pressure inside the tube was very low.



Observations of this experiment :-

→ Cathode rays travels in a straight path with a very high velocity.

→ Cathode rays produces shadow of the object placed in its path.

→ Cathode rays produced green glow when it striked on ZnS coating beyond anode.

→ They affect photographic plates.

→ They penitrate thin sheets of aluminium and other metals.

→ Cathode rays caused movement in paddle wheel placed between electrodes. This indicates that the cathode rays have some mass.

→ When electric and magnetic fields are applied, cathode rays deflected from their path. This shows that cathode rays have somecharge.

→ The direction of deflection shows that cathode rays have -ve charge on them.

→ The charge to mass ratio is same for all cathode rays irrespective of the gas used in the tube.

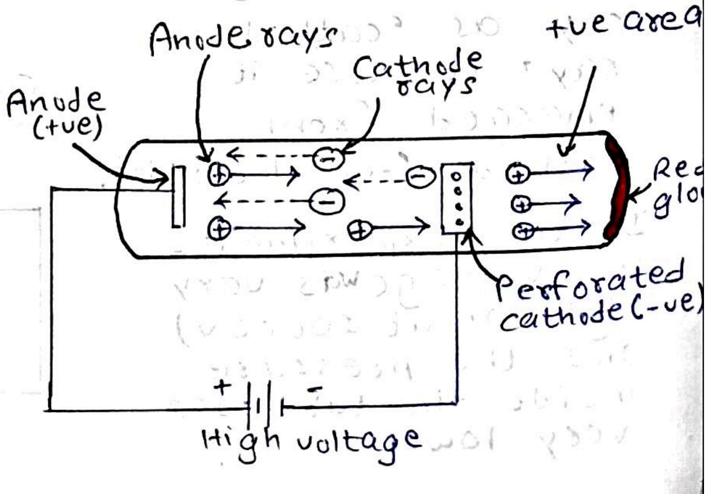
Conclusion made from this experiment:-

→ From this experiment, it was observed that ~~cha~~ cathode rays have some mass and they have -ve charge

→ Later, these rays are termed as 'electron', considering that they are the basic constituent of all the atoms.

Discovery of Anode Rays

→ Discovery of anode rays was done in a similar way as cathode rays were discovered.



→ In discharge tube experiment, another ray emerged from anode and terminates at cathode. This ray was termed as 'anode ray'.

Observations:-

- Anode rays have similar properties like cathode rays.
- Like cathode rays, anode rays travels in a straight line with high velocity. They produces shadow of objects placed in their path.
- Anode rays like cathode rays, caused movement in paddle wheel. This indicates that anode rays also have some mass.
- The direction of deflection of anode rays by electric and magnetic field shows that they have +ve charge.
- For different gases used in the discharge tube, the charge to mass ratio of anode ray is different.

Conclusion :-

- In this experiment, it was observed that, like cathode rays, anode rays have some mass and they also carries +ve charge on themselves.
- Later, anode rays are termed as 'proton', considering ~~they~~ like electrons, they are also a basic constituent of all the atoms.

Discovery of Neutron

- James Chadwick in 1932 bombarded alpha particles on thin sheet of Beryllium.
- He observed not only carbon but also a electrically neutral particle having mass slightly greater than that of proton was obtained.
- He termed this particle as 'neutron', considering like electrons and protons, neutron's are also a basic constituent of all the atoms.

$$\boxed{\text{Nuclear Reaction}} \rightarrow \; {}^{9}_{4}Be + \underbrace{\boxed{{}^{4}_{2}He}}_{\boxed{\alpha\text{-particle}}} \quad {}^{12}_{6}C + \underbrace{\boxed{{}^{1}_{0}n}}_{\boxed{\text{Neutron}}}$$

Dependency of e/m ratio of 'e⁻' & 'p'

→ Charge to mass ratio (e/m ratio) of electrons depends upon nothing. They are independent of gas used in discharged glass tube.

→ But, in case of protons, it depends upon the gas used in discharged glass tube.

Reason →

1) $H_2 \longrightarrow H_2^+ + e^-$
2) $O_2 \longrightarrow O_2^+ + e^-$
3) $NO_2 \longrightarrow NO_2^+ + e^-$
4) $He \longrightarrow He^+ + e^-$

In each case, every gas release same electron.

In each case, every +ve ion is different.

That's why, the e/m ratio of cathode rays is independent of nature of gas. While, e/m ratio of anode rays depends upon nature of gas.

Story after discovery of Sub Atomic Particles

→ Once, Dalton said that atoms can never be divided. But after the discovery of subatomic particles, he failed.

→ Thereafter, different atomic models were discovered.

→ They are as follows:-
(i) Thomson Atomic Model
(ii) Rutherford's model
(iii) Bohr's model
(iv) Quantum mechanical model

Thomson's Model of an Atom

→ This model is also known as 'plum pudding' model or 'watermelon model'.

After discovery of e^-, Thomson assumed that :-

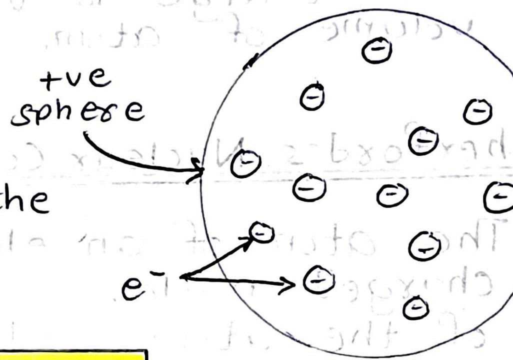
→ Atoms are spherical in shape.

→ Atoms consists of +vely charged sphere.

→ Electrons are embedded in the sphere to maintain electrical neutrality.

Drawback :-

→ Thomson doesn't explained the location of +ve charge.



Rutherford's α-Scattering Experiment

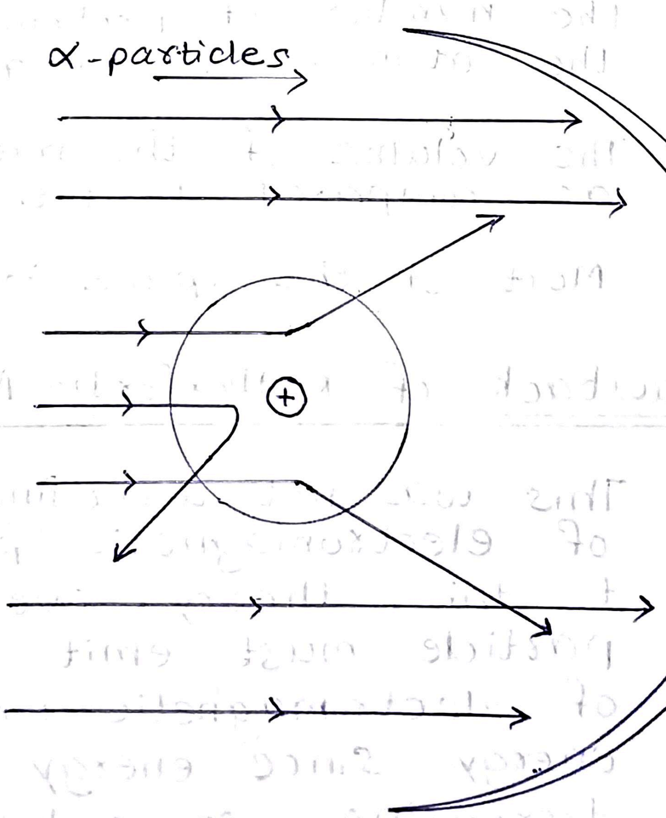
→ Rutherford bombarded α-particles on a thin sheet of gold which was surrounded by photographic plate.

Observations :-

→ Most of the alpha particles passed away without any deflection.

→ Few alpha particles were deflected by small angle while a very few were deflected by large angle.

→ Very very few alpha particles were deflected by an angle of 180° and retraces their path.



Conclusion :-

→ Most of the α-particles passed away undeflected, it means that most of the space inside an atom is empty.

→ Very few α-particles deflected by large angles, it means that there is the presence of concentrated +ve charge in an atom,

→ Very few α-particles deflected by the angle of 180°, it means that volume of concentrated +ve charge is very less as compared to the volume of atom.

Rutherford's Nuclear Concept of the Atom :-

→ The atom of an element consists of a small +vely charged nucleus which is situated at the centre of the atom and which carries almost the entire mass of the atom.

→ The electrons are distributed in the empty space of the atom around the nucleus in different concentric circular paths called orbit.

→ The number of electrons in orbits is equal to the number of protons in the nucleus. Hence, the atom is electrically neutral.

→ The volume of the nucleus is negligibly small as compared to the volume of the atom.

→ Most of the space in the atom is empty.

Drawback of Rutherford's Model:-

→ This was not according to the classical theory of electromagnetic proposed by Maxwell. According to this theory, every accelerated charged particle must emit radiation in the form of electromagnetic waves and loses its total energy. Since energy of electrons keep on decreasing, so radius of the circular orbit should also decrease and ultimately the electron should fall in nucleus.

→ It couldnot explain the line spectrum of H-atom.

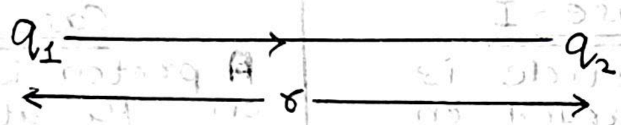
Properties of Charge

→ $Q = ne$ (Charge in quantized).

→ Charge are of two types:
(i) Positive Charge
(ii) Negative Charge

$$\boxed{\begin{aligned} e &= -1.6 \times 10^{-19}\,C \\ p &= +1.6 \times 10^{-19}\,C \end{aligned}}$$

→ Charge is a scalar quantity, and the force between the charge alawys acts along the line joining the charges.



The magnitude of the force between the two charges placed at a distance 'r' is given by

$$\boxed{\underset{\text{(Electrical force)}}{F_E} = \frac{1}{4\pi\varepsilon_0}\frac{q_1 q_2}{r^2}} \qquad \frac{1}{4\pi\varepsilon_0} = \underset{\text{(Constant)}}{K}$$

→ Potential energy between 2 charged particles :-

$$\boxed{P.E. = \frac{Kq_1q_2}{r}} \qquad \boxed{K = 9 \times 10^9}$$

→ If a charged particle q is placed on a surface of potential V then the potential energy of the charge is $\boxed{q \times V}$.

Eg: If a charge q is placed on a surface of potential 5V. Then find P.E.

$\Rightarrow$ P.E. $= q \times V = 1.6 \times 10^{-19} \times 5 = \boxed{8 \times 10^{-19}\,CV}$ Ans

Eg: If α-particle is placed on the surface of potential 10V. The find P.E.

$\Rightarrow$ α-particle has 2+ charge.

$\therefore$ P.E. $= qV = 2 \times 1.6 \times 10^{-19} \times 10 = \boxed{3.2 \times 10^{-18}\,CV}$ Ans

Closest Distance of Approach

→ If an α-particle is bombarded towards nucleus of an atom. The minimum distance between α-particle and nucleus is called closest distance of approach.

→ Closest distance of approach depends upon magnitude of repulsion.

→ If force of repulsion is more, then closest distance of approach is more. If force of repulsion is less, then closest distance of approach is less.

→ The force of repulsion depends upon how much charge the bombarded particle have.

Eg:

Case - I

α-particle is bombarded on Au atom.

Here, $\boxed{q_1 q_2 = 158}$

Case - II

A proton is bombarded on Au atom.

Here, $\boxed{q_1 q_2 = 79}$

Thus in case - I force of repulsion is more. Therefore, α-particle has more closest distance of approach.

Note: Force of repulsion jitna kam, particle nucleus ke utna najdik ja sakta hai. And distance of closest approach bhi utna hi kam hoga.

RBI: $\boxed{q_1 q_2 \text{ more, distance more.}}$

For α-particle

$$d_c = \frac{4KZe^2}{m_\alpha \cdot V_\alpha^2}$$

$K = 9 \times 10^9$

$Z =$ Atomic no.

$e =$ Charge of $1e^-$

$m_\alpha =$ Mass of α-particle

$m_\alpha =$ Velocity of α-particle

For any other charge particle

$$d_c = \frac{2K(Ze)q}{mv^2}$$

$q =$ charge on that particle

$m =$ mass of that particle

$v =$ velocity of that particle

Size of the Nucleus

→ Radius of atom $\approx 10^{-10}$ m or 10^{-8} cm

→ Radius of nucleus $\approx 10^{-15}$ m or 10^{-13} m

→ Ratio of volume of nucleus to the volume of atom :

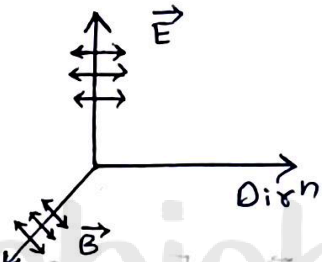
$$\Rightarrow \frac{\text{Vol}^m \text{ nucleus}}{\text{Vol}^m \text{ atom}} = \frac{\cancel{\frac{4\pi}{3}} r^3}{\cancel{\frac{4\pi}{3}} R^3} = \left(\frac{10^{-15}}{10^{-10}}\right)^3 = (10^{-5})^3$$

$$\Rightarrow \boxed{10^{-15}}$$

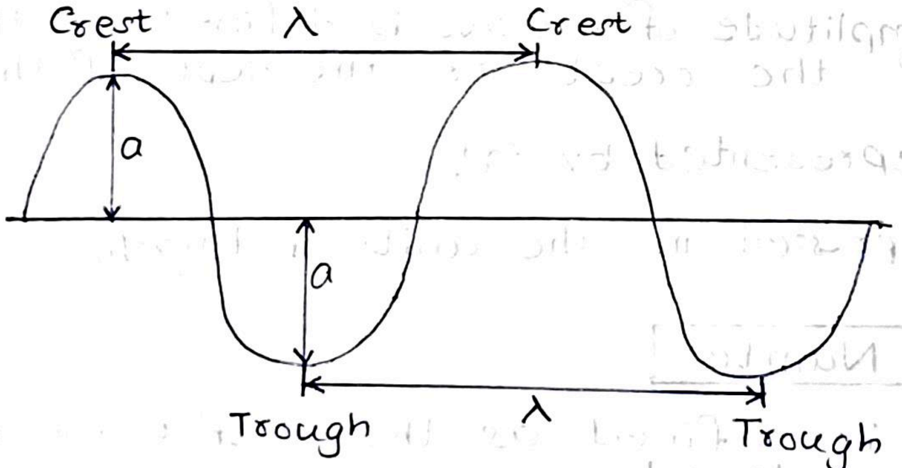
Wave Nature of Electromagnetic Radiation

→ EM waves
- Oscillating Magnetic Field
- Oscillating Electric Field
- Direction of propogation

↳ All of these are mutually perpendicular to each other.



Some important characteristics of a wave :-



1) Wavelength (λ)

→ Wavelength of a wave is defined as the distance between any two consecutive crests or troughs.

→ Unit : $1\,\text{Å} = 10^{-10}$ m , 1 pm $= 10^{-12}$ m , 1 nm $= 10^{-9}$ m

2) Frequency (ν)

→ Frequency of a wave is defined as the number of waves passing through a point in one second.

→ It is represented by 'ν'(nu).

→ Unit : Hertz (Hz) , sec^{-1}

$$\boxed{\nu = \frac{c}{\lambda}} \quad \boxed{\nu = \frac{1}{T}}$$

T= Time Period

3) Velocity

→ velocity of a wave is defined as the linear distance travelled by the wave in one second.

→ Represented by 'V'.

→ Unit: cm/s or m/s (ms^{-1}).

$$\boxed{V = \lambda \nu}$$

λ = Wavelength
ν = Frequency.

→ For EM wave: $V = C = 3 \times 10^8$ m/s

So, $\boxed{C = \lambda \nu}$

4) Amplitude

→ Amplitude of a wave is defined as the height of the crest or the depth of the trough.

→ Represented by 'a'.

→ Expressed in the units of length.

5) Wave Number

→ It is defined as the no. of waves present in 1 cm length.

→ Reciprocal of wavelength.

→ Represented by '$\bar{\nu}$' (nu bar).

→ Unit : cm^{-1}

$$\boxed{\bar{\nu} = 1/\lambda}$$

Order of wavelength in Electromagnetic Spectrum:

- Cosmic Rays
- Gama Rays
- Ultraviolet Rays
- Visible light
- Ifrared
- Micro waves
- Radio waves

Increasing wavelength ↓

Visible Light → Violet, Indigo, Blue, Green, Yellow, Orange, Red

$$\boxed{\begin{aligned}\lambda_{max} &= \text{Red}\\ \lambda_{min} &= \text{Violet}\end{aligned}}$$

Since, $\nu \propto \frac{1}{\lambda}$

$$\therefore \boxed{\begin{aligned}\nu_{max} &= \text{Violet}\\ \nu_{min} &= \text{Red}\end{aligned}}$$

Q. 1 The Vividh Bharti station of AIR, Delhi, broadcasts on a frequency of 1368 KHz. Calculate the wavelength of the em radiation emitted by transmitter.

$\Rightarrow$ $\nu = 1368 \text{ KHz} = 1368 \times 10^3 \text{ Hz}$

$c = \nu\lambda$

$\lambda = \frac{c}{\nu} = \frac{3 \times 10^8}{1368 \times 10^3} = \boxed{219.3\text{m}}$ Ans

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Q. 2 The wavelength of the visible spectrum extends from violet (400nm) to red 750(nm). Express these wavelengths ~~an~~ in frequencies. ($1\text{nm} = 10^{-9}$)

$\Rightarrow$ $\lambda_v = 400\text{nm} = 400 \times 10^{-9}\text{m}$

$\lambda_r = 750\text{ nm} = 750 \times 10^{-9}\text{m}$

$\Rightarrow$ $\nu_v = \frac{c}{\lambda_v} = \frac{3 \times 10^8}{400 \times 10^{-9}} = \boxed{7.5 \times 10^{14}\text{ Hz}}$ Ans

$$\Rightarrow \frac{c}{\lambda_r} = \nu_r = \frac{3\times10^8}{750\times10^{-9}} = \boxed{4\times10^{14}\ Hz}$$ Ans

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Q.3 Calculate (a) wave number and (b) frequency of yellow radiation having wavelength 5800 Å.

$\Rightarrow$ $\lambda = 5800$ Å $= 5800\times10^{-10}$ m

(a) $\bar{\nu} = \frac{1}{\lambda} = \frac{1}{5800}\times\cancel{10^{10}}\ 10^8 = \frac{100\times10^6}{58} = \boxed{1.7\times10^6\ m^{-1}}$ Ans

(b) $\nu = \frac{c}{\lambda} = \frac{3\times10^8}{5800\times10^{-10}} = \frac{3}{58}\times10^{16}$ Hz

Ans $\Rightarrow$ $\boxed{5.1\times10^{14}\ s^{-1}}$

Particle Nature of EM Radiation

→ Some of the experimental phenomenon such as diffraction and interference can be explained by the wave nature of electromagnetic radiation.

→ However, some phenomenon are only explained by particle nature of electromagnetic radiation.

Eg: (i) Black Body Radiation
(ii) Photoelectric Effect

Black Body Radiation

→ When solids are heated they emit a wide range of wavelengths.

→ The ideal body, which emits and absorbs all frequencies is called a black body.

→ The radiation emitted by such a body is called black body radiation.

→ The exact frequency distribution of the emitted radiation from a black body depends only on its temperature.

Example : • When an iron rod is heated, it changes its colour on increasing temperature.

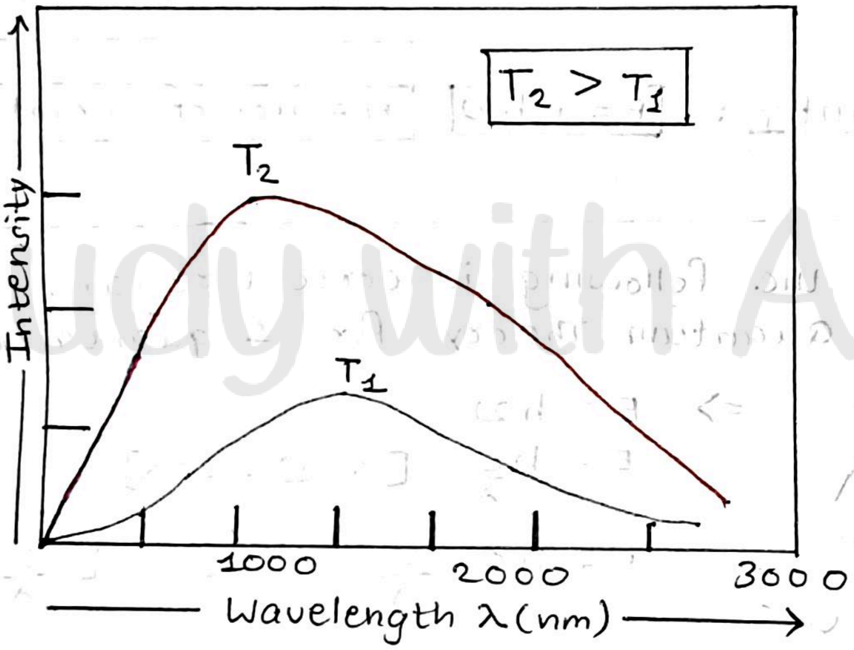
Red → Orange → Yellow → Bluish White

Temperature →

- The change in colour shows that it emits radiations of different frequencies.

- It also emits UV, IR, etc. but our eyes can only detect VIBGYOR.

Wavelength Intensity relationship graph :-



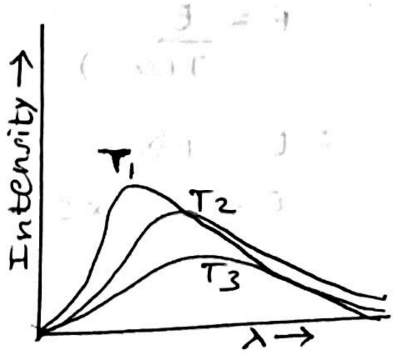
Ques Consider the graph of Black body radiation at 3 different temperatures T_1, T_2, T_3. Compare T_1, T_2 and T_3.

⇒ (a) $T_1 < T_2 < T_3$ ✗

(b) $T_1 = T_2 = T_3$ ✗

(c) $T_1 > T_3 > T_2$ ✗

(d) $T_1 > T_2 > T_3$ ✓



Plank's Quantum Theory

→ Proposed by Max Plank.

→ Quanta → Small Packets.

→ Transfer of energy is done in quanta (small packets).

→ For light, Quanta = Photon

According to Plank's Theory :-

$\boxed{h = 6.6 \times 10^{-34}}$

→ The energy of a quanta is directly proportional to the frequency.

→ For 1 quanta : $E \propto \nu$

$\boxed{E = h\nu}$

E = Energy
ν = Frequency
h = Planks constant

→ For n quanta : $\boxed{E = nh\nu}$

n = no. of quanta

Q.1 Which of the following is correct regarding Plank's Quantum Theory for 1 quanta (MCQ).

(a) $E = h\nu$ ✓

(b) $E = hc\nu$ ✗

(c) $E = \frac{hc}{\lambda}$ ✓

(d) $E = hc\bar{\nu}$ ✓

⇒ $E = h\nu$

$E = \frac{hc}{\lambda}$ $\left[\because \nu = 1/\lambda\right]$

$E = hc \times \frac{1}{\lambda} = hc\bar{\nu} = \left[\because \frac{1}{\lambda} = \bar{\nu}\right]$

Q.2 Consider 100W surya ~~blu~~ bulb emitting yellow light (λ = 500 nm). Find no. of photons emitted by the bulb in 1 sec.?

⇒ Power = Energy emitted per second.

$P = \frac{E}{T(sec.)}$

$\therefore E = P \times t$

$E = 100 \times 1 = 100J$

⇒ $E = nh\frac{c}{\lambda}$

$100 = n \times \frac{6.6 \times 10^{-34} \times 3 \times 10^{8}}{500 \times 10^{-9}}$

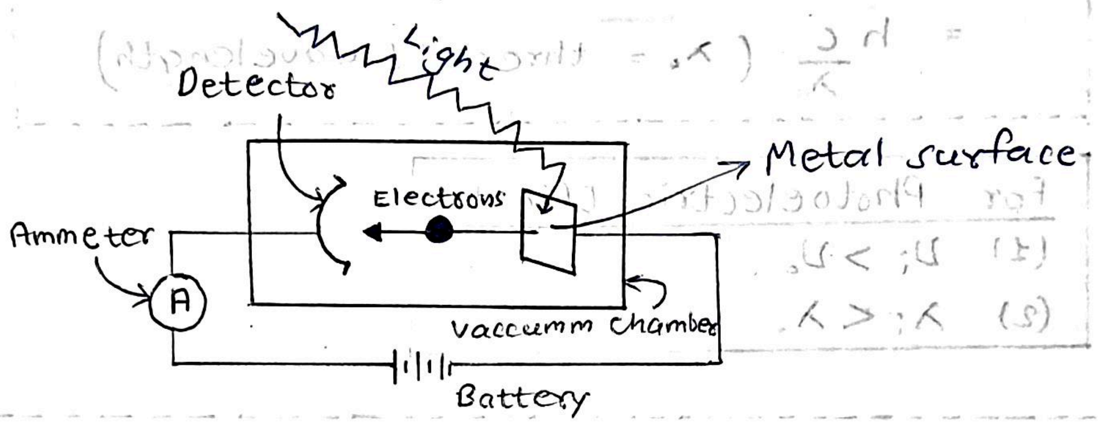
$n = \frac{5 \times 10^{4} \times 10^{-9}}{6.6 \times 3 \times 10^{-26}} = \frac{5}{20} \times 10^{20}$

Ans ⇒ $\boxed{2.5 \times 10^{20} \text{ photons}}$

Photo Electric Effect

→ When monochromatic light fall (incident) on metal surface and e^- is ejected from the metal surface, this phenomena is called photoelectric effect.

→ Monochromatic → single frequency wali light.



Results observed in this experiment were:-

(i) The electrons are ejected from the metal surface as soon as the beam of light strike on the surface.

(ii) The no. of electrons ejected is proportional to the intensity or brightness of light.

(iii) For each metal, there is a characteristic minimum frequency, ν_0 (threshold frequency) below which photoelectric effect is not observed. At a frequency $\nu_i > \nu_0$, ejected electrons come out with certain kinetic energy. The kinetic energies of these electrons increase with the increase of frequency of the light used.

Albert Einstein Equation :-

$$\boxed{E_i = \phi_0 + KE_{max}}$$

E_i = Energy of incident radiation
ϕ_0 = Work function of metal
KE_{max} = Maximum kinetic energy of photo e^-

$\Rightarrow \quad E_i = \phi_0 + (KE)_{max}$

$$\boxed{h\nu_i = \phi_0 + \frac{1}{2} m_e v_e^2}$$

ν_i = frequency of incident radiation
m_e = mass of e^-
v_e = velocity of e^-
$\phi_0 = h\nu_0$ (ν_0 = threshold)
$= h\frac{c}{\lambda_0}$ (λ_0 = threshold wavelength)

For Photoelectric Effect:
(1) $\nu_i > \nu_0$
(2) $\lambda_i < \lambda_0$

Threshold Energy (ϕ_0) : Minimum energy required to eject a photo electron from metal surface.

In case if, $E_i = \phi_0$, Then $KE = 0$.

Threshold Frequency (ν_0): The frequency corresponding to threshold energy.

$$\boxed{\nu_0 = \frac{\phi_0}{h}}$$

If $\nu_i < \nu_0$, then Photoelectric effect doesn't occur.

Threshold Wavelength (λ_0): The wavelength corresponding to ϕ_0.

$$\Rightarrow \phi_0 = \frac{hc}{\lambda_0}$$

$$\therefore \boxed{\lambda_0 = \frac{hc}{\phi_0}}$$

If $\lambda_i < \lambda_0$, then photo electric effect occur.

1 ev : Energy of electron accelerated under potential difference of 1 volt.

$$\Rightarrow KE = qv$$
$$= 1.6 \times 10^{-19} \times 1v$$
$$\therefore \boxed{1 ev = 1.6 \times 10^{-19} J}$$

Q.1 If ϕ_0 of metal is 4.5 ev. Which of the following photons will emit photo e^-?

$\Rightarrow$ (i) $E_i = 5.0$ ev

(ii) $E_i = 16 \times 10^{-19} J$

(iii) $E_i = 4 \times 10^{-18} J$

(iv) $E_i = 4.8$ ev

$\Rightarrow$ (i) $E_i > \phi_0$. $\boxed{\text{It emit photo } e^-}$

(ii) $E_i = 16 \times 10^{-19} J$

$$= \frac{16 \times 10^{-19}}{1.6 \times 10^{-19}} ev = 10 ev > 4.5 ev$$

$\Rightarrow \boxed{\text{It emit photo } e^-}$

(iii) $E_i = 4 \times 10^{-18}$

$$= \frac{4 \times 10^{-18}}{1.6 \times 10^{-19}} ev = \cancel{0.4}\ 25 ev > 4.5 ev$$

$\Rightarrow \boxed{\text{It emit photo } e^-}$

(iv) $E_i = 4.8 ev > 4.5 ev$. $\boxed{\text{It emit photo } e^-}$

Q.2 Fill in the blanks :-

S.No.	ν_i	E_i	ϕ_0	PEE	$(KE)_{max}$	$(Velocity)_{e^-}$
1	300 kHz	$19.8 \times 10^{-29} J$	5 ev	×	×	×
2	500 kHz	$33 \times 10^{-29} J$	5 ev	×	×	×
3	700 kHz	$46.2 \times 10^{-29} J$	5 ev	×	×	×
4	1000 kHz	$6.6 \times 10^{-28} J$	5 ev	×	×	×

(1) $\nu_i = 300 kHz$
$= 3 \times 10^5 Hz$

$E_i = h\nu_i$
$= 6.6 \times 10^{-34} \times 3 \times 10^5$
$= \boxed{19.8 \times 10^{-29} J}$

(2) $\nu_i = 500 kHz$
$= 5 \times 10^5 Hz$

$E_i = h\nu_i$
$= 6.6 \times 10^{-34} \times 5 \times 10^5$
$= \boxed{33 \times 10^{-29} J}$

(3) $\nu_i = 700\ KHz$

$= 7 \times 10^5\ Hz$

$E_i = h\nu_i$

$= 6.6 \times 10^{-34} \times 7 \times 10^5$

$= \boxed{46.2 \times 10^{-29}\ J}$

(4) $\nu_i = 1000\ KHz$

$= 10^6\ Hz$

$E_i = h\nu_i$

$= 6.6 \times 10^{-34} \times 10^6$

$= \boxed{6.6 \times 10^{-28}\ J}$

Various Forms of PEE equation :-

1) $\boxed{E_i = \phi_0 + (KE)_{max}}$

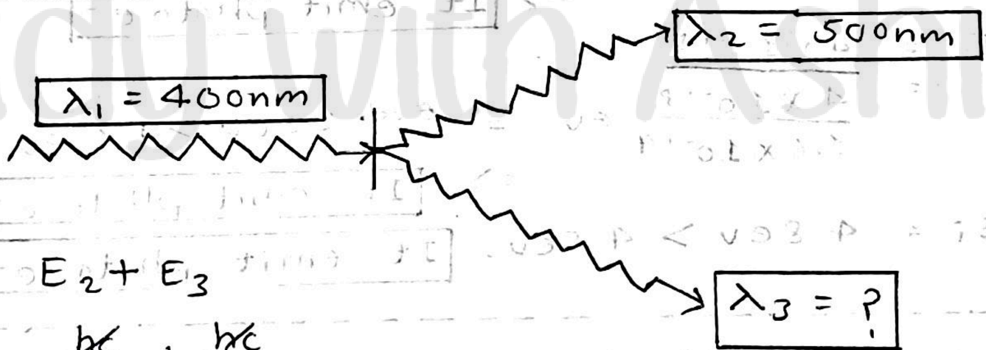
2) $\boxed{h\nu_i = h\nu_0 + (KE)_{max}}$

3) $\boxed{h\nu_i = \phi_0 + (KE)_{max}}$

4) $\boxed{h\nu_i = h\nu_0 + \frac{1}{2}mv^2}$

5) $\boxed{\frac{hc}{\lambda_i} = \frac{hc}{\lambda_0} + \frac{1}{2}mv^2}$

Q.3



$\Rightarrow E_1 = E_2 + E_3$

$$\frac{\cancel{hc}}{\lambda_1} = \frac{\cancel{hc}}{\lambda_2} + \frac{\cancel{hc}}{\lambda_3}$$

$$\frac{1}{\lambda_1} = \frac{1}{\lambda_2} + \frac{1}{\lambda_3}$$

$$\frac{1}{400} = \frac{1}{500} + \frac{1}{\lambda_3}$$

$$\frac{1}{\lambda_3} = \frac{1}{400} - \frac{1}{500}$$

$$\frac{1}{\lambda_3} = \frac{500 - 400}{400 \times 500} = \frac{100}{400 \times 500} = \frac{1}{2000}$$

$$\frac{1}{\lambda_3} = \frac{1}{2000} \Rightarrow \boxed{\lambda_3 = 2000 nm}$$

Important Points regarding Photoelectric Effect :-

1) No. of photo electrons depends upon intensity. And is independent of (E_i, ν_i or λ_i).

2) $(KE)_{max}$ of photo electrons depends upon (E_i, ν_i or λ_i) and is independent of intensity.

3) <u>Graph between ν_i and KE_{max} :-</u>

<u>Eqn</u> => $h\nu_i = h\nu_0 + (KE)_{max}$

=> $(KE)_{max} = h\nu_i - h\nu_0$

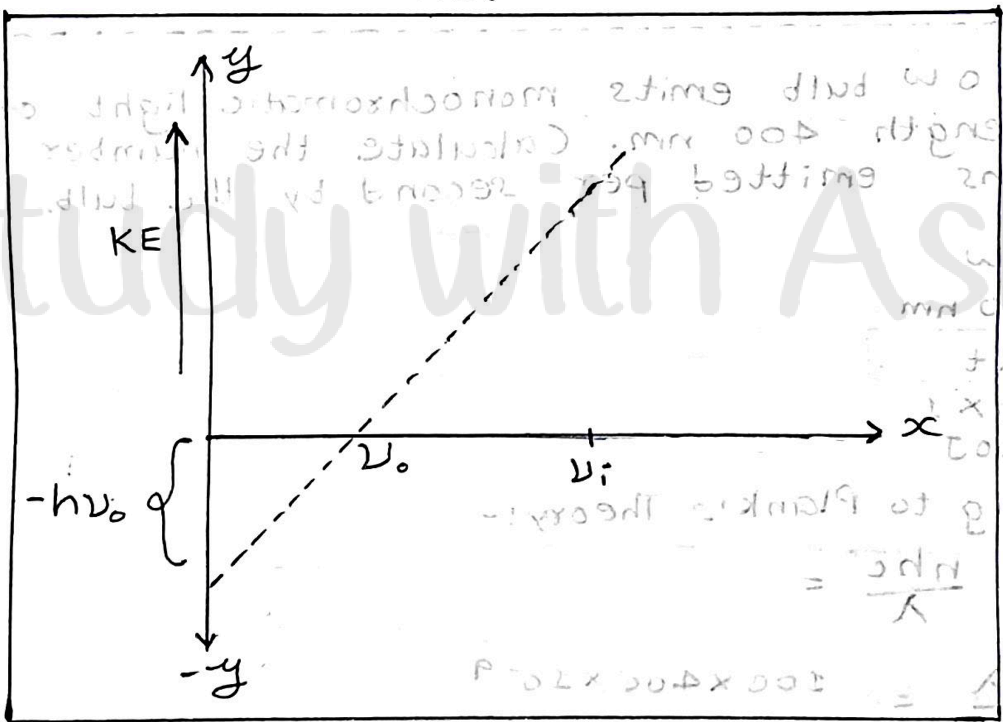
<u>Comparing with $y = mx + c$</u>

$y = (KE)_{max}$

m = slope = h (Plank's constant)

$x = \nu_i$

$c = -h\nu_0$



Q.1 The threshold frequency ν_0 for a metal is $6 \times 10^{14}\ s^{-1}$. Calculate the kinetic energy of an electron emitted when radiation of frequency $\nu_i = 1.1 \times 10^{15}\ s^{-1}$ hits the metal.

=> $\nu_i = 1.1 \times 10^{15}$ Hz $= 11 \times 10^{14}$ Hz

$\nu_0 = 6 \times 10^{14}$ Hz.

=> $h\nu_i = h\nu_0 + (KE)_{max}$

$(KE)_{max} = h\nu_i - h\nu_0$

$= h(\nu_i - \nu_0)$

=> $(KE)_{max} = h(11 \times 10^{14} - 6 \times 10^{14})$

$= 6.6 \times 10^{-34} \times 5 \times 10^{14}$

$= 33 \times 10^{-20}$ J

$= \boxed{3.3 \times 10^{-19}\ J}$ Ans

Q.2 When electromagnetic radiation of wavelength 310 nm fall on the surface of Sodium, electrons are emitted with KE = 1.5eV. Determine the work function (W_0) of sodium.

$\Rightarrow$ KE = 1.5 ev

$1 ev = 1.6 \times 10^{-19} J$

$KE = 1.5 \times 1.6 \times 10^{-19} J = 2.4 \times 10^{-19} J$

$\Rightarrow \lambda_i = 310 \text{ nm} = 310 \times 10^{-9} m$

$E_i = \frac{hc}{\lambda_i} = \frac{6.6 \times 10^{-34} \times 3 \times 10^{8}}{310 \times 10^{-9}} = 6.6 \times 10^{-19}$

$\Rightarrow E_i = \Phi_0 + (KE)_{max}$

$\Phi_0 = E_i - (KE)_{max}$

$= 6.6 \times 10^{-19} - 2.4 \times 10^{-19}$

$= \boxed{4.2 \times 10^{-19} J}$ Ans

Q.3 A 100W bulb emits monochromatic light of wavelength 400 nm. Calculate the number of photons emitted per second by the bulb.

$\Rightarrow P = 100W$

$\lambda = 400 \text{ nm}$

$E = P \times t$

$= 100 \times 1$

$\Rightarrow 100J$

$\Rightarrow$ According to Plank's Theory:-

$E = \frac{nhc}{\lambda}$

$\Rightarrow n = \frac{E\lambda}{hc} = \frac{100 \times 400 \times 10^{-9}}{19.8 \times 10^{-26}}$

$= \frac{4 \times 10^{21}}{19.8}$

$= 0.2 \times 10^{21}$

Ans $\Rightarrow \boxed{2 \times 10^{20} \text{ photons}}$

Q.4 When em radiation of $\lambda = 300nm$ falls on the surface of sodium, electrons are emitted with a KE of $1.68 \times 10^{5} J mol^{-1}$. What is the minimum energy needed to remove an electron from sodium? What is the maximum wavelength will cause a photoelectron to be emitted?

$\Rightarrow \lambda_i = 300\ nm$

$KE = 1.68 \times 10^5\ J/mol$

$KE\ of\ 1e^- = \frac{1.68 \times 10^5}{N_A}\ J$

$\phi_0 = ?$

$\Rightarrow E_i = \phi_0 + (KE)_{max}$

$$\frac{hc}{\lambda_i} = \phi_0 + \frac{1.68 \times 10^5}{6 \times 10^{23}}$$

$$\frac{6.6 \times 10^{-34} \times \cancel{3} \times 10^8}{\cancel{300}_{100} \times 10^{-9}} = \phi_0 + \frac{1.68 \times 10^5}{6 \times 10^{23}}$$

$$6.6 \times 10^{-19} - \frac{1.68}{6} \times 10^{-18} = \phi_0$$

$$\phi_0 = \left(6.6 - \frac{16.8}{6}\right) 10^{-19} = \boxed{3.8 \times 10^{-19}}\ \text{Ans}$$

$\Rightarrow \phi_0 = \frac{hc}{\lambda_0}$

$$\Rightarrow \lambda_0 = \frac{hc}{\phi_0} = \frac{6.6 \times 10^{-34} \times 3 \times 10^8}{3.8 \times 10^{-19}}$$

$$= \frac{6.6 \times 3 \times 10^{-7}}{3.8} = \boxed{5.2 \times 10^{-7}\ nm}\ \text{Ans}$$

Bohr's Atomic Model

→ The electrons revolve around the nucleus with high velocities in certain fixed orbits, called levels.

→ There are two forces acting on it. Electrostatic force of nucleus, which provides centripetal force.

→ It's magnitude is equal to ze^2/r^2, where z = atomic number, e = electronic charge, r = radius of orbit.

Important Postulates :-

1) Atom has a nucleus where all protons and neutrons are present. The size of nucleus is very small and it is present at the centre of the atom.

2) -vely charged electrons revolve around the nucleus in circular orbits. The attraction force between nucleus and electron is equal to the centrifugal force on electron.

Coulumb force

$$F_c = \frac{kq_1q_2}{r^2} = \frac{k(ze)(e)}{r^2} = \frac{kZe^2}{r^2}$$

In circular motion

$$F_c = \frac{mv^2}{r^2}$$

$$\therefore \boxed{\frac{kZe^2}{r^2} = \frac{mv^2}{r}} \quad \ldots (i)$$

3) Electrons can revolve only in those orbits in which angular momentum (mvr) of electron is integral multiple of h/2π.

$$h/2\pi, \frac{2h}{2\pi}, \frac{3h}{2\pi}, \frac{4h}{4\pi}, \ldots$$

$$\boxed{mvr = nh/2\pi} \quad \ldots (ii)$$

m = mass of e^-
v = velocity of e^-
r = distance b/w nucleus & e^-
$n \in N$

4) The orbit in which electron can revolve is known as stationary orbit because in these orbits energy of electron is always constant.

5) Each stationary orbit is associated with definite amount of energy. Therefore, these orbits are also called as energy levels and are numbered as (1, 2, 3, 4, 5, ...) or (K, L, M, N, O, ...) from nucleus onwards.

$$\boxed{\text{Max no. of } e^- \text{ in an orbit} = 2n^2}$$

Where, n = no. of orbit

Eg: Max no. of e^- in N shell.

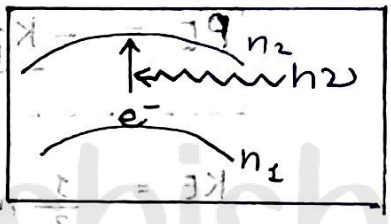
For N, $n = 4$

$e^-_{max} = 2 \times 4^2 = \boxed{32}$

6) The emission or absorption of energy in the form of photon can only occur when electron jumps from one stationary state to another.

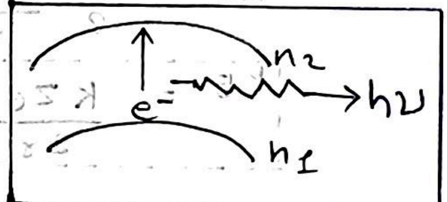
If electron jumps in higher orbit then it absorb energy.

$$\boxed{(E_{n_2} - E_{n_1}) = h\nu}$$



If electron jumps in lower orbit then it emit energy.

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



Finding values of r, v and E :-

1) For 'v' (velocity of electron in n^{th} shell)

$$\boxed{\frac{kze^2}{r^2} = \frac{mv^2}{r}} \quad \ldots (i)$$

$$\boxed{mvr = nh/2\pi} \quad \ldots (ii)$$

Eq^n - 1 :-

$$\frac{kze^2}{r^{\cancel{2}}} = \frac{mv^2}{\cancel{r}}$$

$$kze^2 = mv^2 r$$

$$kze^2 = (mvr)v$$

Putting value of $mvr = \frac{nh}{2\pi}$ from eq^n - 2 :-

$$kze^2 = \frac{nh}{2\pi} \cdot v$$

$$V_n = \frac{kze^2 \times 2\pi}{nh}$$

$$\boxed{V_n \propto \frac{Z}{n}}$$

$$V_n = \frac{2\pi ke^2}{h} \times \frac{Z}{n}$$

$$\boxed{V_n = 2.188 \times \frac{Z}{n} \times 10^6 \text{ m/s}}$$

2) $\boxed{\text{For '}r\text{' (distance b/w nucleus \& electron)}}$

$$\boxed{v = \frac{2\pi k e^2}{h} \times \frac{Z}{n}} \quad \ldots (i)$$

$$\boxed{mvr = \frac{nh}{2\pi}} \quad \ldots (ii)$$

From eqn 2 :-

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

$$r = \frac{nh}{2\pi \times mv}$$

Putting value of 'v' from eqn-1 :-

$$r = \frac{nh \times hn}{2\pi m \times 2\pi k e^2 \times Z}$$

$$r = \frac{n^2 h^2}{4\pi^2 k e^2 Z}$$

$$r = \frac{h^2}{4\pi^2 k e^2} \times \frac{n^2}{Z}$$

$$\boxed{r_n = 0.529 \frac{n^2}{Z} \text{ Å}}$$

$$\boxed{r_n \propto n^2/Z}$$

3) $\boxed{\text{For '}E_n\text{' (energy of electron in Bohr orbit}}$

$$\boxed{PE = \frac{-k q_1 q_2}{r} = \frac{-kZe^2}{r}}$$

$$KE = \frac{1}{2} mv^2$$

$$KE = \frac{1}{2} \frac{kZe^2}{r} \quad \left[\because \frac{mv^2}{\cancel{r}} = \frac{kZe^2}{r^2}\right]$$

$$\boxed{KE = \frac{kZe^2}{2r}}$$

$$TE = KE + PE$$

$$\boxed{KE = \frac{kZe^2}{2r}} \quad \ldots (i)$$

$$\boxed{PE = -\frac{kZe^2}{r}} \quad \ldots (ii)$$

$$TE = \frac{kZe^2}{2r} + \left(\frac{-2kZe^2}{2r}\right) = \frac{kZe^2 - 2(kZe^2)}{2r} = \frac{-kZe^2}{2r}$$

$$\boxed{E_n = -13.6 \cdot \frac{Z^2}{n^2} \text{ ev/atom}}$$

$$\boxed{E_n = -1312 \cdot \frac{Z^2}{n^2} \text{ kJ/mol}}$$

$$\boxed{E_n = -313.6 \cdot \frac{Z^2}{n^2} \text{ kcal/mol}}$$

Q.1 Find the energy of 1st and 2nd orbit of He^+ ion.

⇒ He^+ $(z = 2)$

$$E_1 = -13.6 \times \frac{2^2}{1^2} = -13.6 \times 4 = \boxed{-54.4\ eV}$$

$$E_2 = -13.6 \times \frac{2^2}{2^2} = \boxed{-13.6\ eV}$$

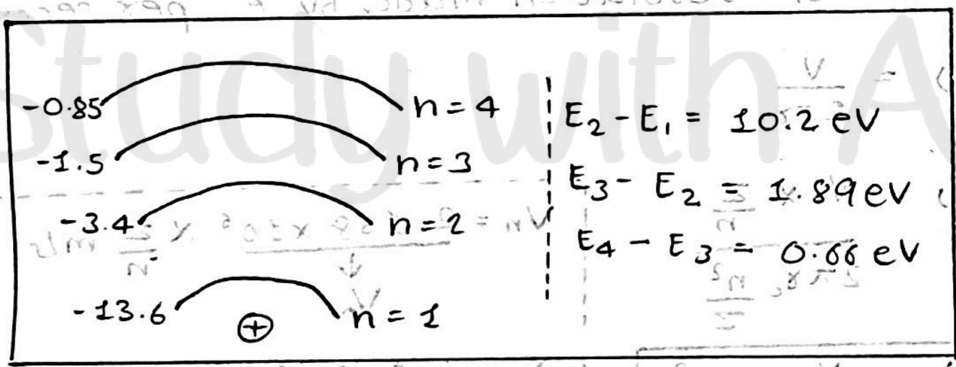
Q.2 Find the energy of first 4 orbits of H-atom

⇒ $E_1 = -13.6 \times \frac{1^2}{1^2} = -13.6\ eV$

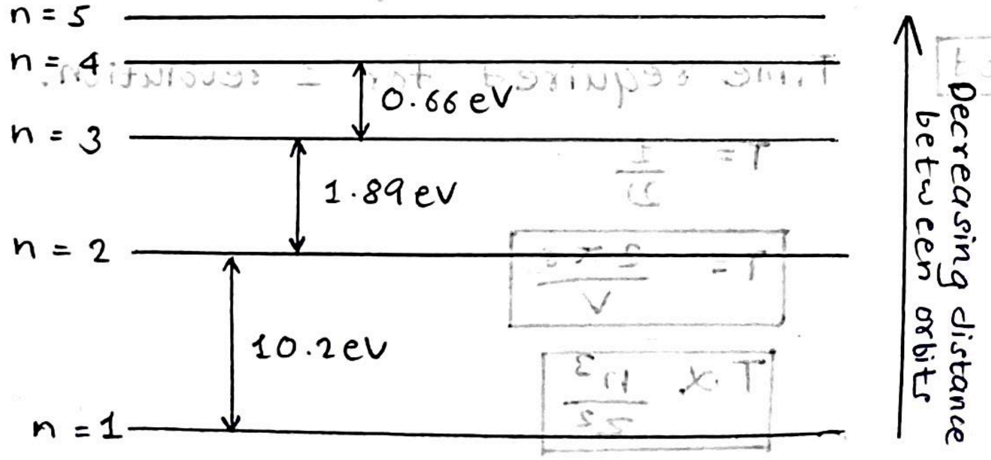
$$E_2 = -13.6 \times \frac{1^2}{2^2} = -13.6 \times \frac{1}{4} = \boxed{-3.4\ eV}$$

$$E_3 = -13.6 \times \frac{1^2}{3^2} = -13.6 \times \frac{1}{9} = \boxed{-1.5\ eV}$$

$$E_4 = -13.6 \times \frac{1^2}{4^2} = -13.6 \times \frac{1}{16} = \boxed{-0.85\ eV}$$



H - atom



Energy distance between two orbits :-

$$\boxed{E_{n_1} = -13.6 \times \frac{Z^2}{n_1^2}} \quad \ldots (i)$$

$$\boxed{E_{n_2} = -13.6 \times \frac{Z^2}{n_2^2}} \quad \ldots (ii)$$

$$\Delta E_n = E_{n_2} - E_{n_1} = -13.6 \times \frac{Z^2}{n_2^2} - \left(-13.6 \times \frac{Z^2}{n_1^2}\right)$$

$$\Delta E_n = 13.6 \times Z^2 \left[\frac{-1}{n_2^2} + \frac{1}{n_1^2}\right]$$

$$\boxed{\Delta E_n = 13.6\, Z^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]}$$

Frequency and Time period of Revolution :-

Frequency : Number of revolution made by e^- per second.

$$\nu = \frac{v}{2\pi r}$$

$$\nu = \frac{v_0 \times \frac{Z}{n}}{2\pi r_0 \frac{n^2}{Z}}$$

$$\boxed{\nu = \frac{v_0}{2\pi r_0} \times \frac{Z^2}{n^3}}$$

$$v_n = \underbrace{2.188 \times 10^6}_{v_0} \times \frac{Z}{n}\ m/s$$

$$r_n = \underbrace{0.529}_{r_0} \times \frac{n^2}{Z}\ \text{Å}$$

$$\boxed{\nu \propto \frac{Z^2}{n^3}}$$

Time period : Time required for 1 revolution.

$$T = \frac{1}{\nu}$$

$$\boxed{T = \frac{2\pi r}{v}}$$

$$\boxed{T \propto \frac{n^3}{Z^2}}$$

Q.1 Find the ratio of time period of an e^- for 2nd orbit of He^+ to 4th orbit of Li^{+2}.

$\Rightarrow$ He^+ $[Z_1 = 2]$
$[n_1 = 2]$

Li^{2+} $[Z_2 = 3]$
$[n_2 = 4]$

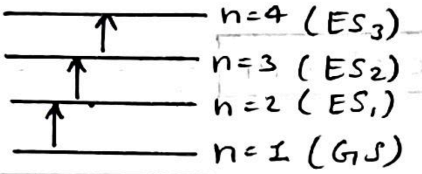
$$\frac{T_{He^+}}{T_{Li^{2+}}} = \frac{n_1^3}{Z_1^2} \times \frac{Z_2^2}{n_2^3}$$

$$= \frac{(2)^3 \times 3^2}{2^2 \times (4)^3} = \frac{9}{32} = \boxed{9:32}$$ **Ans**

Few Important Definitions (Single e^- species) :-

1) $\boxed{\text{Ground State}}$: $n = 1$ (First orbit).

2) $\boxed{\text{Excited state}}$: Orbit other than $n = 1$, i.e., $n = 2, 3, \ldots$



$\boxed{\text{no. of ES} = n - 1}$

3) $\boxed{\text{Ionisation Energy}}$: Energy required to move an e^- from ground state $(n=1)$ to $n = \infty$.

\# $n_1 = 1$, $n_2 = \infty$

$$\Delta E = IE = 13.6 \times Z^2 \left[\frac{1}{1^2} - \frac{1}{\infty^2}\right]$$

$$\therefore \boxed{IE = 13.6 \times Z^2 \text{ eV}} \quad [\because 1/\infty^2 = 0]$$

Eg: (i) H-atom $= 13.6 \times 1^2 \Rightarrow 13.6$ eV
(ii) He^+ $= 13.6 \times 2^2 = 13.6 \times 4 \Rightarrow 54.4$ eV
(iii) Li^{+2} $= 13.6 \times 3^2 = 13.6 \times 9 \Rightarrow 122.4$ eV

4) $\boxed{\text{Excitation Energy}}$: Energy required to move an e^- from GS $(n_1 = 1)$ to any excited state (n_2).

$$\Delta E_n = EE = 13.6 \times Z^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\boxed{EE = 13.6 \times Z^2 \left(1 - \frac{1}{n_2^2}\right) \text{ ev/atom}}$$

Eg: EE of He^+ ion if e^- transists from GS to 6th ES.

$\Rightarrow$ $Z = 2$, $n_1 = 1$, $n_2 = 7$

$\Delta E = 13.6 \times 2^2 \left(1 - \frac{1}{49}\right)$

$= 13.6 \times 4 \left(\frac{48}{49}\right) \approx \boxed{53.3 \text{ eV}}$ **Ans**

5) **Seperation or Binding Energy** :-

Energy required to move an e^- from n_1 (any excited state) to $n_2 = \infty$.

$$SE = \Delta E_n = 13.6 \times Z^2 \left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$SE = 13.6 \times Z^2 \left(\frac{1}{n_1^2} - 0\right) \quad \because \left(\frac{1}{\infty^2} = 0\right)$$

$$\therefore \boxed{SE = 13.6 \times \frac{Z^2}{n_1^2}}$$

Eg: SE for 3rd ES of He^+

=> $Z = 2$

$n_1 = 4$

$$\Rightarrow SE = 13.6 \times \frac{2^2}{4^2} = 13.6 \times \frac{4}{16} = \frac{13.6}{4} = \boxed{3.4 \ \frac{eV}{atom}}$$

Types of Spectrum

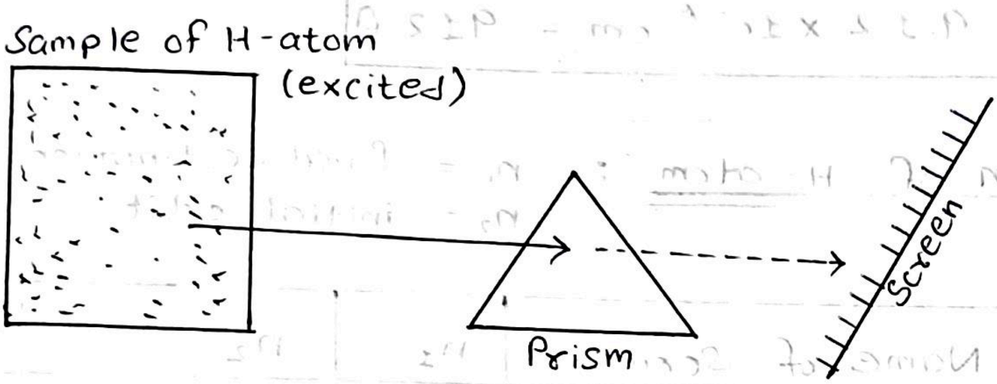
- Spectrum
  - Continuous Spectrum
    - 1 colour overlaps other.
    - Eg: Rainbow, white light from prism.
  - Non-continuous / Discrete Spectrum
    - Absorption
      - Line Spectrum
      - Band Spectrum
    - Emission
      - Line Spectrum
      - Band Spectrum

- **Absorption** : e^- jumps to higher orbit (absorbs light).
- **Emission** : e^- goes back to lower orbit (emits light).

In our syllabus : Non Continuous ⟶ Emission ⟶ Line Spectrum

Emission Spectrum of H-atom

- Spectrum of any atom is always line spectrum.
- Emission spectrum means transition of e^- from higher orbit to lower orbit.



- Suppose that there is a sample of H-atoms (excited state) (energy ↑). High energy means e^- is in any other orbit other than $n = 1$. If we stop giving energy to the electron, it jumpes in lower orbit and emits photon. And this photon can be observed by placing a prism. The photons are observed as line spectrum on the screen. So this is called line spectrum of H- atom.

Note 1

Jabtak electron ko energy mil rahi hai tabtak of excited state me rahega. Jab energy milna band hoga tab electron lower orbit me ~~de-exit~~ de-excite hoga aur photon emit karega.

Note 2

Spectrum kisi atom ka fingerprint hota hai.

Rydberg Formula :-

$$\Delta E_n = 13.6 \times z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \quad \cdots (1)$$

From Plank's Theorem :-

Energy of a photon $= h\nu = \frac{hc}{\lambda}$

$$\Delta E_n = \frac{hc}{\lambda} \quad \cdots (2)$$

From eq^n (1) and (2) :-

$$\frac{hc}{\lambda} = 13.6 \times z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

or

$$\frac{1}{\lambda} = \frac{13.6}{hc} \times z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\boxed{\frac{1}{\lambda} = R z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) = \bar{\nu}}$$

$\therefore$ **Rydberg Formula** $\Rightarrow$ $\boxed{\bar{\nu} = \frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)}$

R = Rydberg constant $= 109678\ cm^{-1} = 109700\ cm^{-1}$
$= 10970000\ m^{-1} = 1.1 \times 10^7\ m^{-1}$

$\boxed{\frac{1}{R} = 9.12 \times 10^{-6}\ cm = 912\ \mathring{A}}$

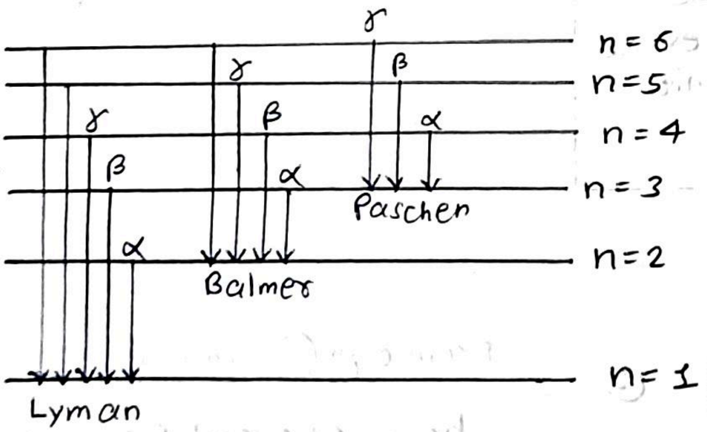
Spectrum of H-atom : n_1 = final destination
n_2 = initial orbit

No.	Name of Series	n_1	n_2	Type
1	Lyman Series	1	2, 3, 4, 5....	UV
2	Balmer series	2	3, 4, 5....	Visible
3	Paschen Series	3	4, 5, 6....	IR
4	Brakett Series	4	5, 6, 7....	IR
5	Pfund series	5	6, 7, 8....	IR
6	Humphery Series	6	7, 8, 9....	IR

Trick
लाल $\rightarrow$ Lyman
बाल $\rightarrow$ Balmer
पाल $\rightarrow$ Paschen
भारत $\rightarrow$ Brakett
प्रेमी $\rightarrow$ Pfund
हूं $\rightarrow$ Humphery

Terms :-

1) α line = nearest transition (1st line)
2) β line = 2nd line
3) γ line = 3rd line



Max. energy = min. λ = $\infty \rightarrow n_1$
Min. energy = λ_{max} = α line

1) **Lyman Series**

Transition ($n_2 \rightarrow n_1$)

$n_1 = 1$

$n_2 = 2, 3, 4, \ldots$

Ryd. Formula :

$$\frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

λ for α line :-

$n_1 = 1, \; n_2 = 2$

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

$$\frac{1}{\lambda} = R\left(\frac{4-1}{4}\right)$$

$$\frac{1}{\lambda} = \frac{3R}{4}$$

$$\therefore \lambda_\alpha = \frac{4}{3R}$$

λ for β line :-

$n_1 = 1, \; n_2 = 3$

$$\frac{1}{\lambda} = R \times 1^2\left(\frac{1}{1^2} - \frac{1}{3^2}\right)$$

$$\frac{1}{\lambda} = R \times \frac{8}{9}$$

$$\therefore \lambda_\beta = \frac{9}{8R}$$

λ for γ line :-

$n_1 = 1, \; n_2 = 4$

$$\frac{1}{\lambda} = R \times 1^2\left(\frac{1}{1^2} - \frac{1}{4^2}\right)$$

$$\frac{1}{\lambda} = R \times \frac{15}{16}$$

$$\therefore \lambda_\gamma = \frac{16}{15R}$$

λ maximum :-

$\lambda_{max} = \text{energy}_{min.} = \alpha\text{-line}$

$n_1 = 1, \; n_2 = 2$

$\therefore \lambda_{max} = \frac{4}{3R}$ (solved above)

λ minimum :-

$\lambda_{min.} = \text{energy}\,max. = \infty \rightarrow n_1$

$n_1 = 1, \; n_2 = \infty$

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

$$\frac{1}{\lambda} = R(1-0)$$

$$\therefore \lambda_{min} = \frac{1}{R}$$

2) **Balmer Series**

$n_1 = 2$

$n_2 = 3, 4, 5, \ldots$

Ryd. Formula :

$$\frac{1}{\lambda} = RZ^2\left(\frac{1}{2^2} - \frac{1}{n_2^2}\right)$$

λ for α line :-

$n_1 = 2, \; n_2 = 3$

$$\frac{1}{\lambda} = R \times 1^2\left(\frac{1}{4} - \frac{1}{9}\right)$$

$$\frac{1}{\lambda} = R\left(\frac{9-4}{36}\right)$$

$$\frac{1}{\lambda} = \frac{5R}{36}$$

$$\therefore \lambda_\alpha = \frac{36}{5R}$$

λ for β line :-

$n_1 = 2, \; n_2 = 4$

$$\frac{1}{\lambda} = R \times 1^2\left(\frac{1}{4} - \frac{1}{16}\right)$$

$$\frac{1}{\lambda} = R\left(\frac{4-1}{16}\right)$$

$$\frac{1}{\lambda} = R\left(\frac{3}{16}\right)$$

$$\therefore \lambda_\beta = \frac{16}{3R}$$

λ for γ line :-

$n_1 = 2, \; n_2 = 5$

$$\frac{1}{\lambda} = R \times 1^2\left(\frac{1}{4} - \frac{1}{25}\right)$$

$$\frac{1}{\lambda} = R\left(\frac{25-4}{100}\right)$$

$$\therefore \lambda_\gamma = 100/21R$$

λ minimum :-

$n_1 = 2, n_2 = \infty$

$$\frac{1}{\lambda} = R \times 1^2\left(\frac{1}{4} - \frac{1}{\infty}\right)$$

$$\frac{1}{\lambda} = \frac{R}{4}$$

$$\therefore \lambda_{min} = \frac{4}{R}$$

λ maximum :-

$n_1 = 2, n_2 = 3$

$$\frac{1}{\lambda} = R(1)^2\left(\frac{1}{4} - \frac{1}{9}\right)$$

$$\therefore \lambda_{max} = \frac{36}{5R}$$

3) **Paschen Series**

$n_1 = 3$

$n_2 = 4, 5, 6, \ldots$

Ryd. Formula :

$$\frac{1}{\lambda} = R(Z)^2\left(\frac{1}{3^2} - \frac{1}{n_2^2}\right)$$

4) **Brakett Series :-**

$n_1 = 4, n_2 = 5, 6, 7, \ldots$

Ryd. Formula :

$$\frac{1}{\lambda} = R(z)^2\left(\frac{1}{4^2} - \frac{1}{n_2^2}\right)$$

5) **Pfund Series :-**

$n_1 = 5$

$n_2 = 6, 7, 8, \ldots$

Ryd. Formula :

$$\frac{1}{\lambda} = R \times z^2\left(\frac{1}{5^2} - \frac{1}{n_2^2}\right)$$

6) **Humphery Series -**

$n_1 = 6$

$n_2 = 7, 8, 9, \ldots$

Ryd. Formula :

$$\frac{1}{\lambda} = R \times z^2\left(\frac{1}{6^2} - \frac{1}{n_2^2}\right)$$

Maximum spectral for any particular series $= (n_2 - n_1)$

1) Lyman → $(n_2 - 1)$
2) Balmer → $(n_2 - 2)$
3) Paschen → $(n_2 - 3)$
4) Brakett → $(n_2 - 4)$
5) Pfund → $(n_2 - 5)$
6) Humphery → $(n_2 - 6)$

Note

The no. of different λ found in any transition contributes different spectral line.

Maximum no. of spectral line for any transition (all lines are included)

$$= \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$$

Eg: For e⁻ transition from 5th excited state to 1st excited state. Find the total maximum lines.

$\Rightarrow n_1 = 2, n_2 = 6$

$$\Rightarrow \frac{(6-2)(6-2+1)}{2} = \frac{4 \times 5}{2} = 10 \text{ Ans}$$

Limitations of Bohr's Model

1) Bohr's theory doesn't explain the spectrum of multi e^- species.

2) Why the angular momentum of the revolving e^- is equal to $nh/2\pi$, has not been explained by Bohr's theory.

3) Bohr inter-related quantum theory of radiation and classical laws of physics without any theoretical explaination.

4) Bohr's theory does not explain the fine structure of the spectral lines. Fine structure of the spectral line is obtained when spectrum is viewed by spectroscope of more resolution power.

5) Bohr's theory does not explain the splitting of spectral lines in the presence of magnetic ~~effe~~ field (zemman's effect) or electric field (stark's effect).

de-Broglie Duality

→ Light has dual nature (particle, wave).

→ e^- also has dual nature.

- Particle = mass, momentum, kinetic energy
- wave = λ, ν, amplitude.

→ Every matter have wave associated to it.

Energy of particle => $E = mc^2$ ... ①

Energy of wave => $\frac{hc}{\lambda} = E_n$ ... ②

From eq^n ① and ② :-

$$\frac{hc}{\lambda} = mc^2$$

$$\therefore \boxed{\lambda = \frac{h}{mv} = \frac{h}{p}}$$

Here, λ = de Broglie wavelength
m = mass of particle
v = velocity of particle
p = ~~mome~~ momentum = mv

$$\boxed{\lambda = \frac{h}{mv}}$$
(For matter)
(Having mass)

$$\boxed{\lambda = \frac{hc}{E}}$$
(For photon)
(Massless)

→ de Broglie duality does not hold good for particle with greater mass. It explaines only lighter particles (e^-, p, α, n, etc.).

Ques. 1 Find the wavelength associated to a car with mass 400 kg and moving with velocity 180 km/h.

$\Rightarrow$ $m = 400\,kg$

$v = 180$ kmph

$h = 6.6 \times 10^{-34}$

$$\lambda = \frac{h}{mv} = \frac{6.6 \times 10^{-34}}{400 \times 50}$$

$$= \frac{6.6 \times 10^{-34}}{20000}$$

$\lambda \Rightarrow 3.3 \times 10^{-38}$ m <u>Ans</u>

(negligible)

Ques. 2 Find λ for e^- moving with velocity 2.188×10^6 m/s.

$\Rightarrow$ $m = 9.1 \times 10^{-31}$ kg

$v = 2.188 \times 10^6$ m/s

$$\Rightarrow \lambda = \frac{h}{mv} = \frac{6.6 \times 10^{-34}}{9.1 \times 10^{-31} \times 2.188 \times 10^6} = \frac{6.6}{9.1 \times 2.188} \times 10^{-9}$$

<u>Ans</u> $\Rightarrow 0.33 \times 10^{-9}$ m

<u>de-Broglie wave length in terms of kinetic energy :-</u>

$mv = p$

<u>Sq. both side :-</u>

$(mv)^2 = p^2$

$m^2v^2 = p^2$

$mv^2 = \frac{p^2}{m}$

<u>Multiplying by $\frac{1}{2}$ both side :-</u>

$\frac{1}{2}mv^2 = \frac{p^2}{2m}$

$KE = \frac{p^2}{2m}$ or $p^2 = 2m(KE)$

$\therefore p = \sqrt{2m(KE)}$

$\lambda = \frac{h}{mv} = \frac{h}{p}$

Hence, $\boxed{\lambda = \frac{h}{\sqrt{2m(KE)}}}$

<u>Note</u>: KE of any particle if placed b/w potential difference of V volt is $KE = qv$

q = charge of particle

V = voltage

$\therefore \boxed{\lambda = \frac{h}{\sqrt{2m(KE)}} = \frac{h}{\sqrt{2mqv}}}$

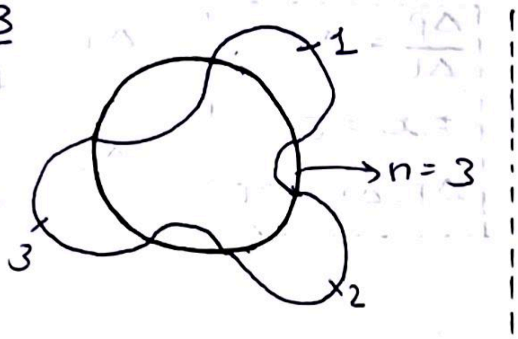
λ:

- $\frac{h}{p}$
- $\frac{h}{mv}$
- $\frac{h}{\sqrt{2mqv}}$
- $\frac{h}{\sqrt{2m(KE)}}$
- (for α) $\frac{0.101}{\sqrt{V}}$ Å
- $\frac{0.286}{\sqrt{V}}$ Å for (p^+)
- $\sqrt{150/V}$ Å (for e^-)

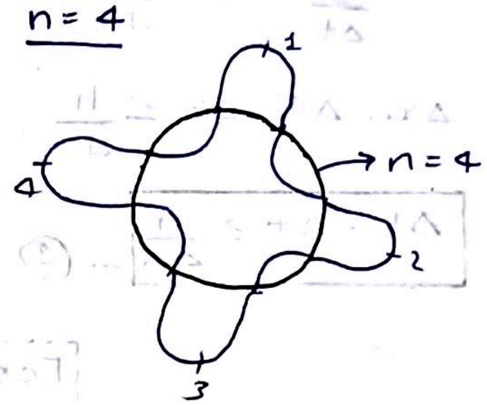
Explaination of angular momentum of e^- from de-Broglie's eqn:-

- e^- moves in an orbit in the form of a wave.
- The no. of wave in any orbit = no. of that orbit.

Eg: $n = 3$



$n = 4$



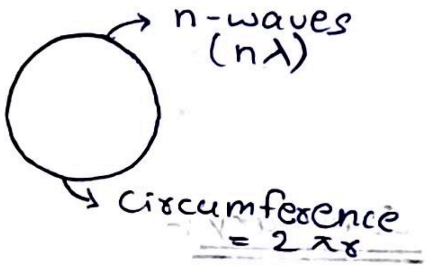
- **For n^{th} orbit** (Proof)

no. of waves = n

wavelength = λ

Total length of n wave = $n\lambda$

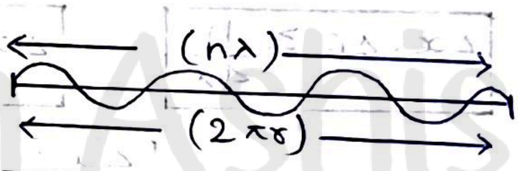
Circumference of circle = $2\pi r$



$\Rightarrow n\lambda = 2\pi r$

∵ de Broglie's $\lambda = \dfrac{h}{mv}$

$\Rightarrow \dfrac{nh}{mv} = 2\pi r$



$\therefore \boxed{mvr = \dfrac{nh}{2\pi}}$

Note: $\dfrac{h}{2\pi} = \hbar$

Heisenberg Uncertainity Principle (HUP)

- It is impossible to measure simultaneously the exact momentum and position of a particle as small as an electron.

$$\boxed{\Delta x \cdot \Delta p \geq \frac{h}{4\pi}} \quad \dots (1)$$

Δx = uncertainity in position

Δp = uncertainity in momentum

$$\boxed{\Delta x \cdot m \cdot \Delta v \geq \frac{h}{4\pi}} \quad \dots (2)$$

Δv = uncertainity in velocity

$$\Delta x . \Delta p \geq \frac{h}{4\pi}$$

Multiplying & dividing by Δt:-

$$\Delta x . \frac{\Delta p}{\Delta t} . \Delta t \geq \frac{h}{4\pi}$$

$$\Delta x . \Delta F . \Delta t \geq \frac{h}{4\pi}$$

$$\boxed{\Delta E . \Delta t \geq \frac{h}{4\pi}} \text{ --- } (3)$$

$\frac{\Delta p}{\Delta t} = \Delta F$

$F.x = E$

$\Delta F . \Delta x = \Delta E$

Δt = uncertainity in time

ΔF = uncertainity in force

ΔE = uncertainity in energy

For question solving

$$\Delta x . \Delta p = \frac{h}{4\pi}$$

$$h = 6.6 \times 10^{-34}$$

$$\pi = 22/7$$

Summary:-

$$\boxed{\Delta x . \Delta p \geq \frac{h}{4\pi}} \quad \boxed{\Delta x . m\Delta v \geq \frac{h}{4\pi}} \quad \boxed{\Delta E . \Delta t \geq \frac{h}{4\pi}}$$

Δx = uncertainity in position
Δp = uncertainity in momentum
Δv = uncertainity in velocity
ΔE = uncertainity in energy
Δt = uncertainity in time

Quantum Mechanical Model of an atom

→ Derived from Schrodinger Equation.

→ e^- is a wave. It follows an equation which is called 'e^- wave function' [(ψ) Psi].

→ ψ is function of [r, ϕ, θ]

$$\psi_{(r,\theta,\phi)} = \underbrace{f(r)}_{\text{radial part}} . \underbrace{f(\theta) . f(\phi)}_{\text{angular part}}$$

θ & ϕ = angle
r = distance

- Radial Part → f(r) depends upon 2 no. (n, l)
- Angular Part → f(θ) depends upon 2 no. (l, m)
  f(Φ) depends upon 1 no. (m)

n – principal quantum no.
l – azimuthal quantum no.
m – magnetic quantum no.

Shrondinger Wave Equation :-

$$\boxed{\frac{d^2\psi}{dx^2} + \frac{d^2\psi}{dy^2} + \frac{d^2\psi}{dz^2} + \frac{8\pi^2 m}{h^2}(E-V)\psi = 0}$$

→ E = Total energy
→ V = Potential energy

After solving this eqn we get value of 'ψ'

$$\boxed{\psi_{(r,\theta,\phi)} = \underset{n,l}{f(r)} \cdot \underset{l,m}{f(\theta)} \cdot \underset{m}{f(\phi)}}$$

Orbital : The 3-D space around a nucleus where e⁻ finding probability is greater than 95%.

Node : The region around nucleus where e⁻ finding probability is almost zero.

Radial Nodes ---- Nodes ---- Angular Nodes

Radial Nodes

Radial part = 0
f(r) = 0

Also known as 'spherical nodes'

No. of radial node = n - l - 1

Nodes

Total Nodes
= AN + RN
= $\cancel{l}$ + n - $\cancel{l}$ - 1

TN = n - 1
AN = l
RN = n - l - 1

Angular Nodes

Angular part = 0
f(θ) . f(Φ) = 0

Also known as 'angular nodes'

No. of angular node = l

Quantum Numbers

They are the set of 4 numbers which defines the position of e⁻ in an atom.

1) Principal Quantum no.
2) Azimuthal Quantum no.
3) Magnetic Quantum no.
4) Spin Quantum no.

e⁻ address

(n) Orbit / Shell
↓
(l) Subshells
↓
(m) Orbitals
↓
(s) Spin → 2e⁻

Principal Quantum No. (n)

→ It tells us about orbit in which e^- is present.

→ It also tells us about energy of an e^-.

→ n → Bohr Orbit. (1, 2, 3, 4, -----).
K, L, M, N, ------

→ Angular momentum of e^- in an orbit $= \frac{nh}{2\pi}$ or $n\hbar$.

→ It also tells us about size of orbit.

Maximum no. of e^- in any orbit $= 2n^2$

Maximum no. of orbital in any orbit $= \frac{2n^2}{2} = n^2$

n	Max. no. of e^-	Max. no. of orbitals
2	$2\times2^2 = 8$	$2^2 = 4$
3	$2\times3^2 = 18$	$3^2 = 9$
4	$2\times4^2 = 32$	$4^2 = 16$
5	$2\times5^2 = 50$	$5^2 = 25$
6	$2\times6^2 = 72$	$6^2 = 36$

Azimuthal Quantum No. (l)

→ Also known as subsidiary quantum no. or angular momentum quantum no.

→ It tells us about subshell.

→ value of l tells us about shape of orbital.

→ The value of l depends upon 'n'. [l = 0 to (n-1)]

l	Shape	Subshell
0	Spherical	s - subshell
1	Dumble	p - subshell
2	Double dumble	d - subshell
3	Complex	f - subshell
4	Complex	g - subshell
5	Complex	h - subshell

n	Possible value of l	Subshell
1	0	1s - subshell
2	0	2s - subshell
	1	2p - subshell
3	0	3s - subshell
	1	3p - subshell
	2	3d - subshell
4	0	4s - subshell
	1	4p - subshell
	2	4d - subshell
	3	4f - subshell

→ Angular momentum of an e^- in an orbital $= \sqrt{l(l+1)} = \frac{h}{2\pi}$

Orbit	Orbital Angular momentum
s	0
p	$\sqrt{2}\hbar$
d	$\sqrt{6}\hbar$
f	$\sqrt{12}\hbar$

Magnetic Quantum No. (m)

→ It tells us about orientation of orbitals.

→ Its value depends upon 'l'.

→ Possible values of m = ($-l$ to $+l$).

→ Total no. of possible value = $(2l+1)$

$$\boxed{\underbrace{-l, \dots, -2, -1}_{l},\ \underset{1}{0},\ \underbrace{+1, +2, \dots, l}_{l}}$$

→ Each value of 'm' represents 1 orbital.

→ Total no. of orbital in a subshell = $2l+1$

→ Total no. of e^- in a subshell = $2(2l+1)$

a) s - subshell

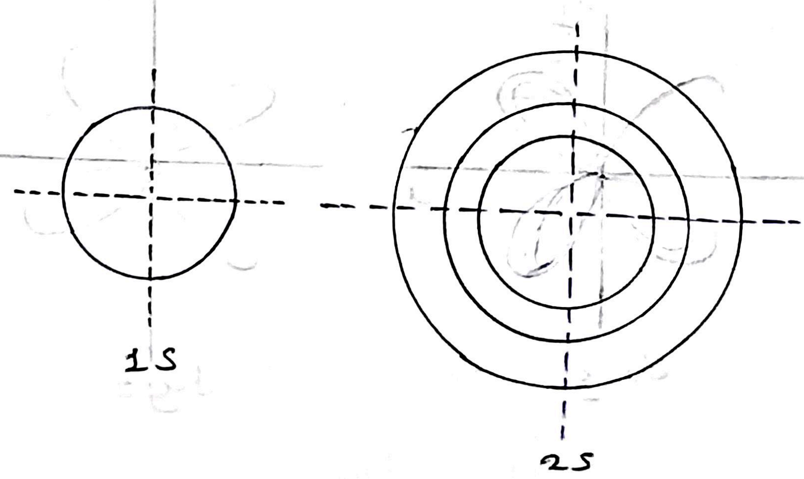
→ Shape = spherical

→ $l = 0$

→ $m = -l$ to $+l = 0$

→ Total possible value $= 2l+1$
$= 2\times0+1$
$= 1$

∴ An s-subshell contains 1 orbital.



b) p- subshell

→ $l = 1$

→ $m = (-1 \text{ to } +1)$

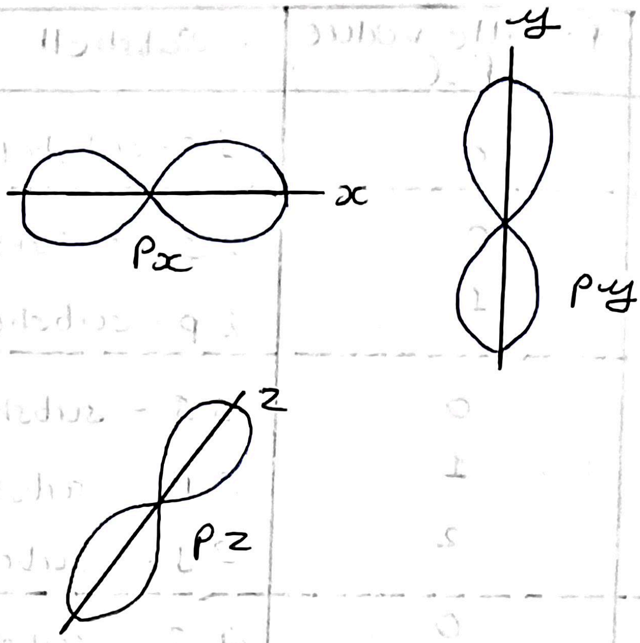
$= -1, 0, +1$

→ Total possible values $= (2l+1)$

$= 2 \times 1 + 1$

$= 3$

∴ p-subshell contains 3 orbitals.



c) d- subshell

→ $l = 2$

→ $m = (+2 \text{ to } +2)$

$= -2, -1, 0, +1, +2$

→ Total possible values $= (2l+1)$

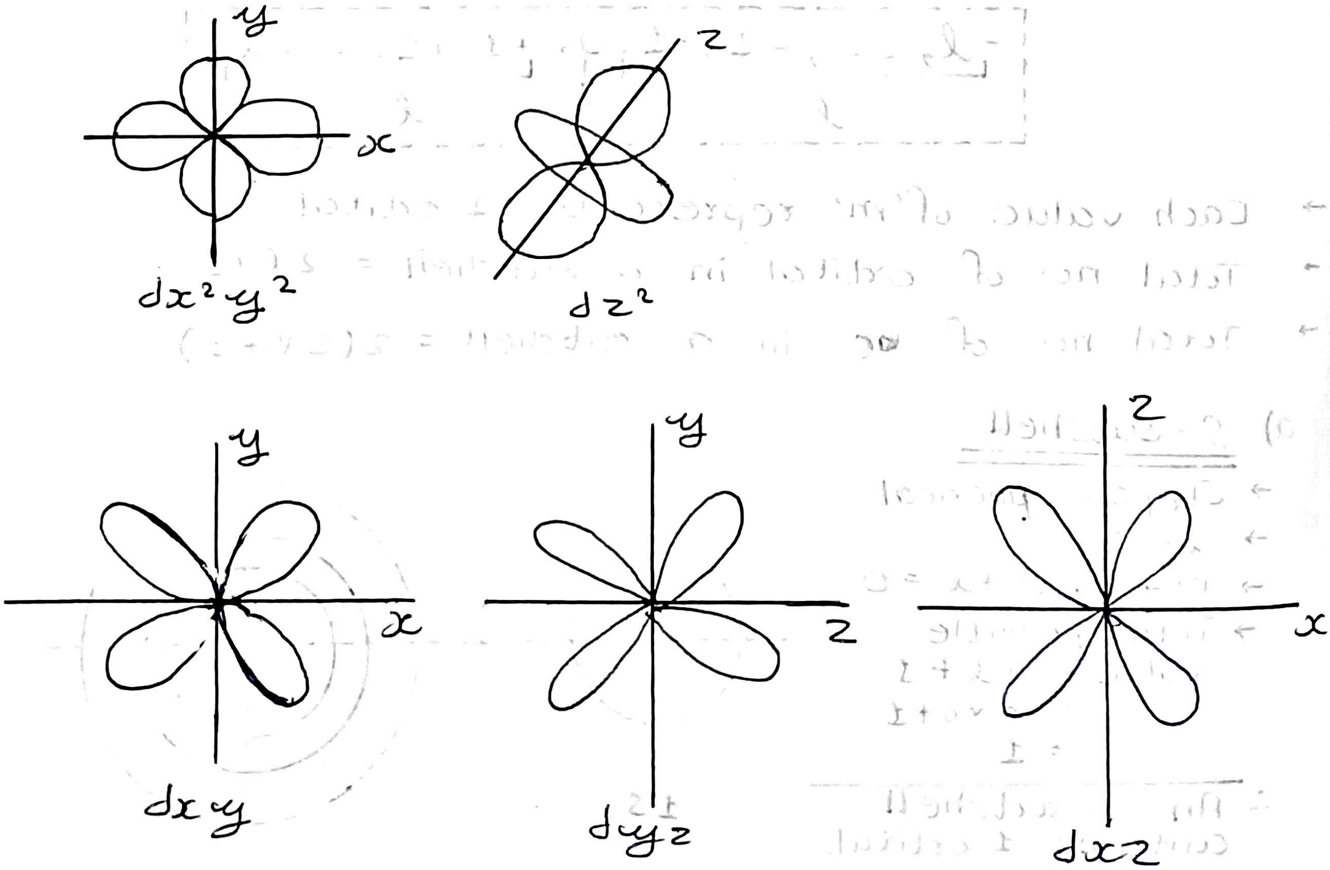
$= 2 \times 2 + 1$

$= 5$

∴ d- subshell contains 5 orbitals.

d-subshell

Axial	Interaxial
dz^2, dx^2y^2	dxy, dyz, dxz
Lobes along axis	Lobes b/w axis

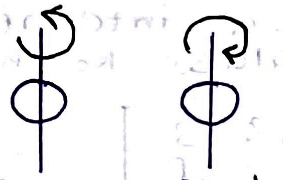


Spin Quantum No. (s)

→ Electronic spin

→ Clockwise = $+\frac{1}{2}$ or ↑ or 1

→ Anticlockwise = $-\frac{1}{2}$ or ↓ or ⇂



$+\frac{1}{2} \leftarrow \boxed{\uparrow\downarrow} \rightarrow -\frac{1}{2}$

→ 1 orbital

Note : Spin only magnetic moment

$\mu_s = \sqrt{n(n+2)}$ B.M. → unit (Bohr Magneton)

n = no. of unpaired electron

Questions from Q. no.

Type - 1 : Correct set of quantum no.

- n → 1, 2, 3, ...
- l → 0 to (n-1)
- m → -l to +l
- s → $+\frac{1}{2}$ or $-\frac{1}{2}$

Eg: Which of the following set is possible ?

	n	l	m	s	
1)	4	3	-2	-1/2	✓
2)	5	4	+5	+1/2	✗
3)	4	4	0	-1/2	✗
4)	6	2	0	-1/2	✓
5)	3	0	-1	+1/2	✓

Type - 2 : Maximum no. of e⁻

→ In an orbit

Max. no. of e⁻ = $2n^2$

Max. no. of orbital = n^2

→ In a subshell

Max. no. of e⁻ = $2(2l+1)$

Max. no. of orbital = $2l+1$

→ In an orbital

Max. no. of e⁻ = 2.

Eg: Find max. no. of e⁻

1) n = 4 (Orbit)
$2n^2 = 2 \times 4^2 = 32e^-$

2) n = 4, l = 2 (subshell)
$2(2l+1)$
$2(2\times2+1) = 10e^-$

3) n = 4, l = 2, m = -2 (1 orbital) = 2e⁻

4) n = 4, l = 2, m = -2, s = $+\frac{1}{2}$
max = 1 e⁻

5) Max. e^- in $(n=4, m=-1)$

=> 'l' is missing

=> ~~IA~~ In that case, intermediate Q. no. ki sabhi possible values ko manenge.

=> $n=4$, $\boxed{l = 0, 1, 2, 3 \quad (s, p, d, f)}$

s →	0
p →	-1, 0, +1
d →	-2, -1, 0, +1, +2
f →	-3, -2, -1, 0, +1, +2, +3

=> 3 orbitals

Ans => 6 e^- (1 orbital = $2e^-$)

---

6) $l = 3$

=> max. e^- in subshell $= 2(2l+1)$

$= 2(3\times2+1)$

$= 2\times7$

Ans $= 14\ e^-$

7) $n=3$, $s = +\frac{1}{2}$

=> l, m → all possible value

$n=3$ (orbit)

3s, 3p, 3d (subshell)

↓ ↓ ↓

$1 + 3 + 5 = 9$ (orbital)

Total = 9 orbital

$+\frac{1}{2}$ → $9e^-$, $-\frac{1}{2}$ → $9e^-$

Type - 3: No. of nodes

→ RN = $n - l - 1$ (Radial node)

Eg: Find no. of RN for

1) 2s	2) 3s	3) 3p	4) 3d	5) 6p
$n=2$	$n=3$	$n=3$	$n=3$	$n=6$
$l=0$	$l=0$	$l=1$	$l=2$	$l=1$
$n-l-1$	$n-l-1$	$n-l-1$	$n-l-1$	$n-l-1$
$2-0-1$	$3-0-1$	$3-1-1$	$3-2-1$	$6-1-1$
Ans => 1	Ans => 2	Ans => 1	Ans => 0	Ans => 4

→ AN = l (Angular node)

s → 0

p → 1

d → 2

f → 3

→ TN = $n - 1$ (Total nodes) Eg: Find TN for 6d

=> $n = 6$, $l = 2$ => TN $= 6-1 = 5$ Ans

Electronic Configuration

→ There are 3 rules for electronic configuration.

1) Aufbau's Principle
2) Hund's Rule
3) Pauli's Exclusion

Aufbau's Principle :-

→ Aufbau principle gives a sequence in which various subshell are filled up depending on the relative order of the energies of various subshells.

→ Principle : The subshell with min. energy is filled up first. When this subshell obtained max. quota of e^- then the next subshell with higher energy starts filling.

$$\boxed{\text{Energy} \propto (n+l)}$$

Eg: value of $(n+l)$ for -

1) $1s = 1+0 = 1$
2) $2s = 2+0 = 2$
3) $2p = 2+1 = 3$
4) $3s = 3+0 = 3$
5) $3p = 3+1 = 4$
6) $3d = 3+2 = 5$

$$\boxed{(n+l)\uparrow \ (\text{Energy})\uparrow}$$

Note: For 2p and 3s the value of $(n+l)$ is same

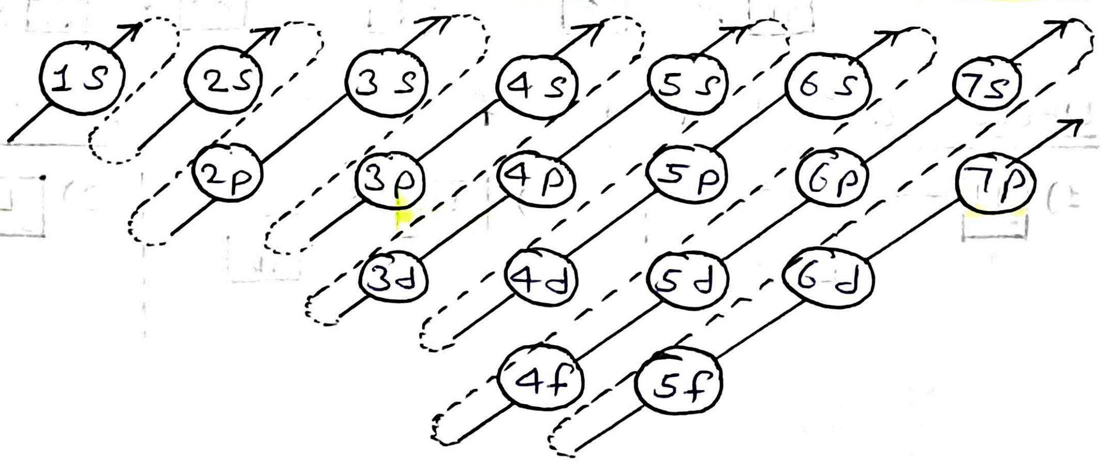
$2p = 2+1 = 3$

$3s = 3+0 = 3$

In that case, the subshell which has greater value of n, has more energy.

$\therefore \ 2p < 3s$

→ Sequence : 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < ....



→ **Important Note:** For single e^- species Energy $\propto$ 'n' only

Eg: H-atom

$1s < (2s = 2p) < (3s = 3p = 3d) < (4s = 4p = 4d = 4f)$

Hund's Rule of Maximum Spin Multiplicity :-

→ For degenerate orbitals e^- first filled up singly then paring takes place.

→ **Degenerate orbital:** Orbital with same energy. All orbitals of a subshells are degenerate.

$3p^{2}$ (2 = no. of e^-) → $\boxed{\uparrow}\boxed{\uparrow}\boxed{\ }$, $\boxed{\uparrow\downarrow}\boxed{\ }\boxed{\ }$, $\boxed{\uparrow}\boxed{\ }\boxed{\uparrow}$

Ek subshell me jabtak sabhi orbital me ek e^- na aa jaye hum pairing nahi karenge.

Eg: Which of the following is not according to Hund Rule?

1) $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$ (It follows)
2) $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\ }$ (Doesn't follow)
3) $\boxed{\downarrow}\boxed{\downarrow}\boxed{\downarrow}$ (It follow)
4) $\boxed{\downarrow}\boxed{\uparrow\downarrow}\boxed{\ }$ (Doesn't follow)

Pauli's Exclusion Principle :-

→ No two e^- in an atom can have same values of all 4 quantum no.

→ Orbital can accommodate maximum 2 e^- with opposite spin.

$\boxed{\uparrow\downarrow}$ ✓ $\boxed{\uparrow}$ ✓ $\boxed{\downarrow}$ ✓ $\boxed{\uparrow\uparrow}$ × $\boxed{\downarrow\downarrow}$ ×

Electronic Configuration from Z = 1 to Z = 30 :-

1) $_1H \rightarrow 1s^1$ $\boxed{\uparrow}$

2) $_2He \rightarrow 1s^2$ $\boxed{\uparrow\downarrow}$

3) $_3Li \rightarrow 1s^2\ 2s^1$ $\boxed{\uparrow\downarrow}\ \boxed{\uparrow}$

4) $_4Be$ → $1s^2\ 2s^2$

[↑↓] [↑↓]

5) $_5B$ → $1s^2\ 2s^2\ 2p^1$

[↑↓] [↑↓] [↑| |]

6) $_6C$ → $1s^2\ 2s^2\ 2p^2$

[↑↓] [↑↓] [↑|↑|]

7) $_7N$ → $1s^2\ 2s^2\ 2p^3$

[↑↓] [↑↓] [↑|↑|↑]

8) $_8O$ → $1s^2\ 2s^2\ 2p^4$

[↑↓] [↑↓] [↑↓|↑|↑]

9) $_9F$ → $1s^2\ 2s^2\ 2p^5$

[↑↓] [↑↓] [↑↓|↑↓|↑]

10) $_{10}Ne$ → $1s^2\ 2s^2\ 2p^6$

[↑↓] [↑↓] [↑↓|↑↓|↑↓]

11) $_{11}Na$ → $1s^2\ 2s^2\ 2p^6\ 3s^1$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑]

12) $_{12}Mg$ → $1s^2\ 2s^2\ 2p^6\ 3s^2$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓]

13) $_{13}Al$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓]

[↑| |]

14) $_{14}Si$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓]

[↑|↑|↑]

15) $_{15}P$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓]

[↑|↑|↑]

16) $_{16}S$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑↓|↑|↑]

17) $_{17}Cl$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑↓|↑↓|↑]

18) $_{18}Ar$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑↓|↑↓|↑↓]

19) $_{19}K$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓]

20) $_{20}Ca$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓]

21) $_{21}Sc$ → $\underbrace{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6}_{[Ar]}\ 4s^2\ 3d^1$

→ $[Ar]\ 4s^2\ 3d^1$

$\boxed{\uparrow\downarrow}\quad \boxed{\uparrow}\boxed{\ \ }\boxed{\ \ }\boxed{\ \ }\boxed{\ \ }$

22) $_{22}Ti$ → $\underbrace{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6}_{[Ar]}\ 4s^2\ 3d^2$

→ $[Ar]\ 4s^2\ 3d^2$

$\boxed{\uparrow\downarrow}\quad \boxed{\uparrow}\boxed{\uparrow}\boxed{\ \ }\boxed{\ \ }\boxed{\ \ }$

23) $_{23}V$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$

→ $[Ar]\ 4s^2\ 3d^3$

$\boxed{\uparrow\downarrow}\quad \boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\ \ }\boxed{\ \ }$

24) $_{24}Cr$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4$ (wrong)

(Exception)

→ $[Ar]\ 4s^2\ 3d^4$ ✗

$\boxed{\uparrow\downarrow}\quad \boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\ \ }$ ✗

→ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$ (correct)

→ $[Ar]\ 4s^1\ 3d^5$ ✓

$\boxed{\uparrow}\quad \boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$ ✓

25) $_{25}Mn$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$

→ $[Ar]\ 4s^2\ 3d^5$

$\boxed{\uparrow\downarrow}\quad \boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

26) $_{26}Fe$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

→ $[Ar]\ 4s^2\ 3d^6$

$\boxed{\uparrow\downarrow}\quad \boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

27) $_{27}Co$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$
→ $[Ar]\ 4s^2\ 3d^7$

↑↓		↑↓	↑↓	↑	↑	↑

28) $_{28}Ni$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$
→ $[Ar]\ 4s^2\ 3d^8$

↑↓		↑↓	↑↓	↑↓	↑	↑

29) $_{29}Cu$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$ (Wrong)
(Exception) → $[Ar]\ 4s^2\ 3d^9$ ✗

↑↓		↑↓	↑↓	↑↓	↑↓	↑

✗

→ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$ (correct)
→ $[Ar]\ 4s^1\ 3d^{10}$ ✓

↑		↑↓	↑↓	↑↓	↑↓	↑↓

✓

30) $_{30}Zn$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$
→ $[Ar]\ 4s^2\ 3d^{10}$

↑↓		↑↓	↑↓	↑↓	↑↓	↑↓

Explaination for exception to (n+ℓ) rule (Cr and Cu) :-

→ Half filled and fully filled orbitals are more stable.

→ There are 2 reasons (for extra stability of half filled orbital):

1) Symmetry

d^5

↑	↑	↑	↑	↑

Symmetrical

d^7

↑↓	↑↓	↑	↑	↑

Unsymmetrical

d^{10}

↑↓	↑↓	↑↓	↑↓	↑↓

Symmetrical

2) Exchange Energy — Same spin wale e^- aapas me exchange ho sakte hain.

No. of exchange ↑ Stability ↑

d^5

↑	↑	↑	↑	↑
4	3	2	1	0

Total exchange = 10

d^4

↑	↑	↑	↑	
3	2	1	0	

Total exchange = 6

d^{10}

4	3	2	1	0
↑↓	↑↓	↑↓	↑↓	↑↓
4	3	2	1	0

Total exchange = 20

Electronic Configuration for ions :-

$\boxed{\text{Cations}}$ → Remove e^- from outermost orbit.
→ Remove e^- from outermost subshell.
→ Remove e^- from 4s between 3d and 4s.

$\boxed{\text{Anions}}$ → Add electron in last subshell.

Eg: $[Na^+] \rightarrow 1s^2\ 2s^2\ 2p^6$
→ $[Na] = 1s^2\ 2s^2\ 2p^6\ 3s^1$

$[Mn^{2+}] \rightarrow [Ar]\ 3d^5$
→ $[Mn] = [Ar]\ 4s^2\ 3d^5$

$[O^{2-}] \rightarrow 1s^2\ 2s^2\ 2p^6$
→ $[_8O] = 1s^2\ 2s^2\ 2p^4$

Spin only magnetic moment $(\mu_s) = \sqrt{n(n+2)}$ B.M.
n = no. of unpaired e^-

Eg: ~~Find no. of unpaired e^-~~
Find μ_s of O-atom
⇒ $[_8O] \rightarrow 1s^2\ 2s^2\ 2p^4$

↑↓	↑↓	↑↓	↑	↑

⇒ unpaired e^- (n) = 2
⇒ $\mu_s = \sqrt{2(2+2)} = \sqrt{8}$ BM Ans

up.e⁻	μ_s
1 →	$\sqrt{1(1+2)} = \sqrt{3}$ BM
2 →	$\sqrt{2(2+2)} = \sqrt{8}$ BM
3 →	$\sqrt{3(3+2)} = \sqrt{15}$ BM
4 →	$\sqrt{4(4+2)} = \sqrt{24}$ BM

Note: If μ_s is given
Eg: $\mu_s = 4.84$ B M
Integer part — decimal part

Then integer part is no. of unpaired e^-.

μ_s	up.e⁻
1.73	1
2.85	2
4.84	4
5.87	5

Graphs

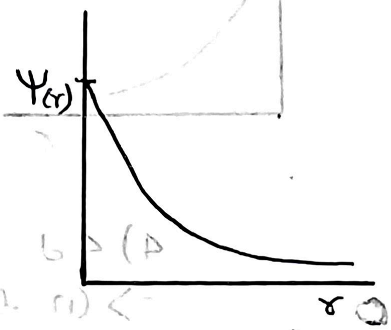
1) $\Psi(r)$ / $f(r)$ / $R(r)$ vs r

For s-subshell
$r = 0$; $\Psi(r) > 0$
$r = \infty$; $\Psi(r) = 0$

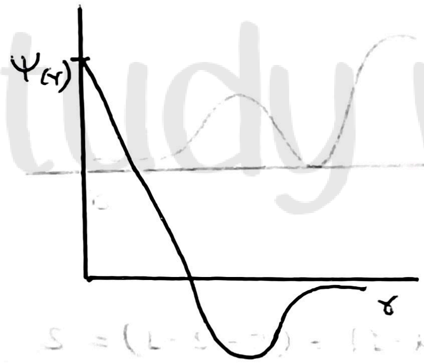
For p, d, f subshell
$r = 0$; $\Psi(r) = 0$
$r = \infty$; $\Psi(r) = 0$

__Note:__ The graph cuts x-axis $(n-l-1)$ times other than origin.

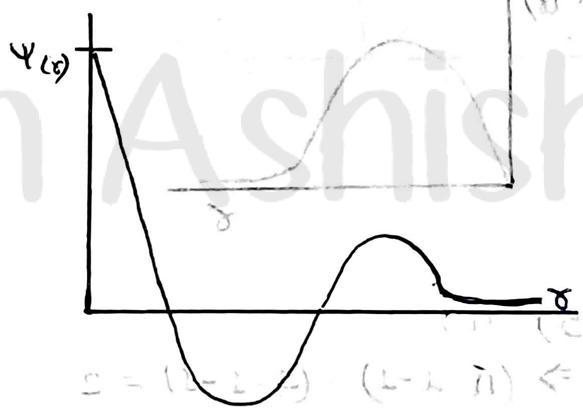
1) 1s
$\Rightarrow (n-l-1) = (1-0-1) = 0$



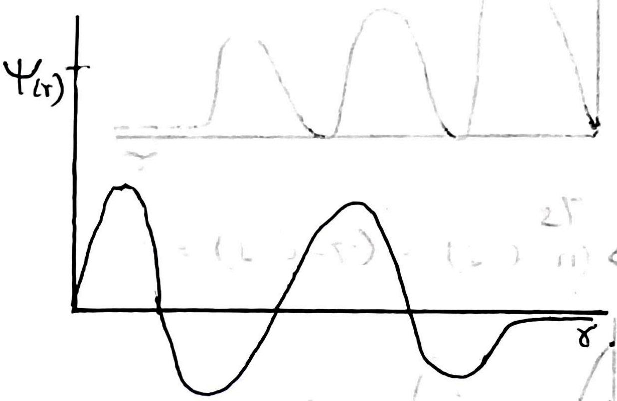
2) 2s
$\Rightarrow (n-l-1) = (2-0-1) = 1$



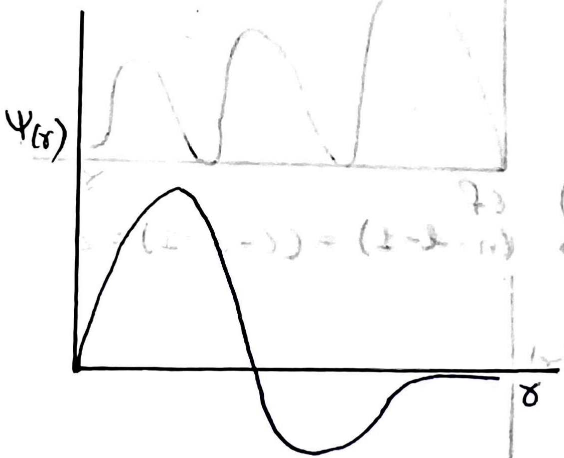
3) 3s
$\Rightarrow (n-l-1) = (3-0-1) = 2$



4) 6d
$\Rightarrow (n-l-1) = (6-2-1) = 3$



5) 3p
$\Rightarrow (n-l-1) = (3-1-1) = 1$



2) $\psi^2(r)$ / $R^2(r)$ / $f^2(r)$ vs r

For s-orbital

$r = 0$; $\psi^2(r) > 0$

$r = \infty$; $\psi^2(r) = 0$

For s, p, d

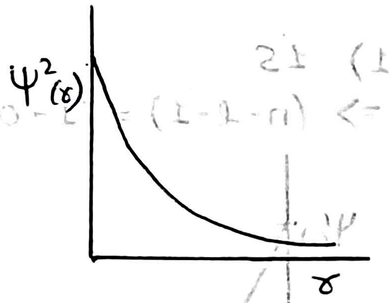
$r = 0$; $\psi^2(r) = 0$

$r = \infty$; $\psi^2(r) = 0$

Note: Due to square graph never be -ve.

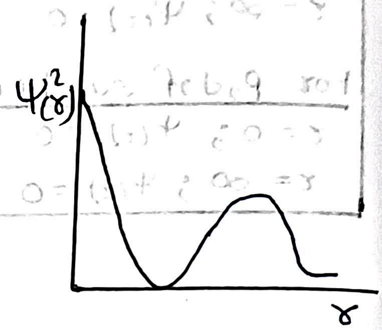
1) 1s

$\Rightarrow (n-l-1) = (1-0-1) = 0$



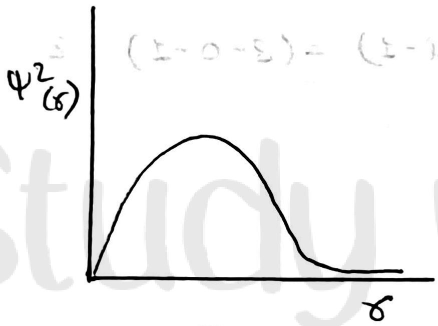
2) 2s

$\Rightarrow (n-l-1) = (2-0-1) = 1$



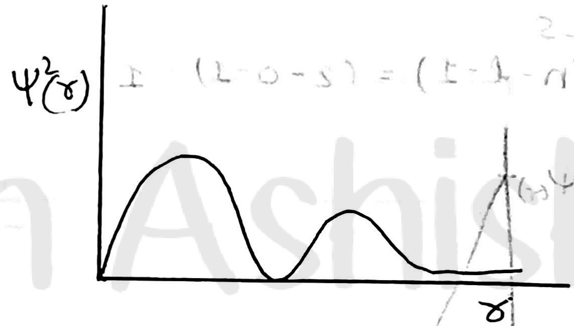
3) 2p

$\Rightarrow (n-l-1) = (2-1-\ ) =$



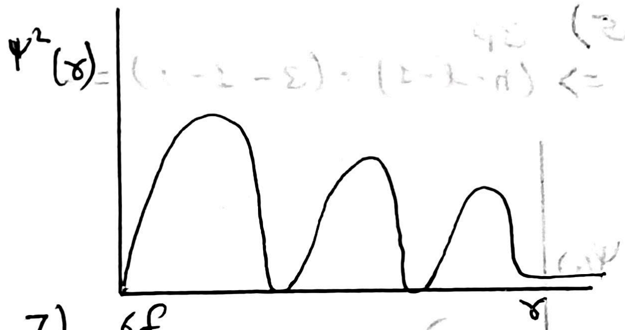
4) 4d

$\Rightarrow (n-l-1) = (4-2-1) = 1$



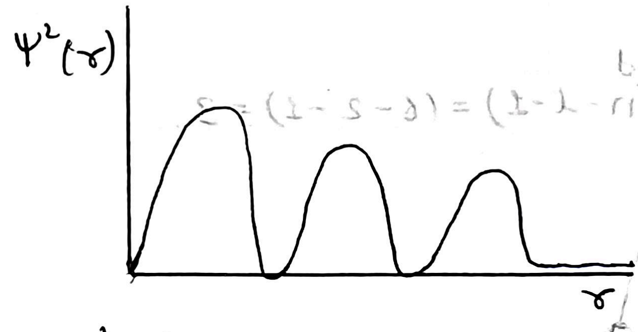
5) 4p

$\Rightarrow (n-l-1) = (4-1-1) = 2$



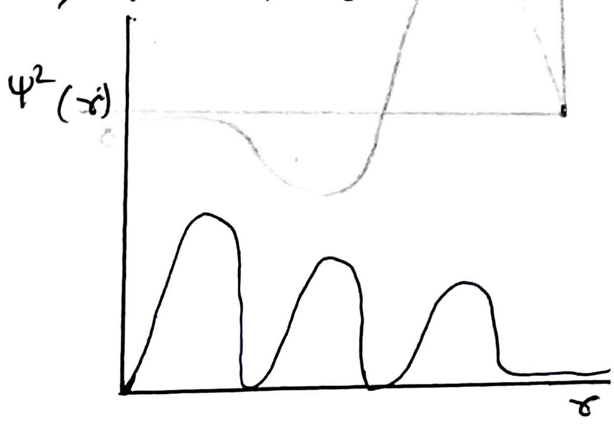
6) 5d

$\Rightarrow (n-l-1) = (5-2-1) = 2$



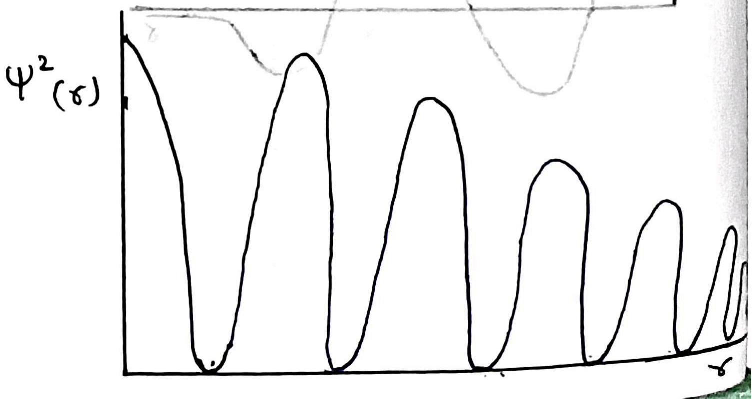
7) 6f

$\Rightarrow (n-l-1) = (6-3-1) = 2$



7) 7s

$\Rightarrow (n-l-1) = (7-0-1) = 6$



3) RPDF (Radial Probability Distibution Function)

$$\boxed{P(r) = 4\pi r^2 \psi^2(r) \cdot dr}$$

For s, p, d, f
$r = 0$; RPDF = 0
$r = \infty$; RPDF = 0

→ -ve nahi hoga

→ x-axis ko other than origin $(n-l-1)$

→ Peak badhta jayega.

1) 2S

=> $(n-l-1)$

$(2-0-1) = 1$

2) 5p

=> $(n-l-1)$

$(5-1-1) = 3$