# Atomic Structure

⇒ Atomic Terms :-

→ Nuclide :- Various species of atoms in general.

→ Nucleons :- Sub- atomic particles in the nucleus of an atom, i.e. protons and neutrons

→ Isotopes :- Atoms of an element with the same atomic no. but different mass number.

→ Mass Number (A) :- Sum of the no. of protons and neutrons i.e., the total no. of nucleons.

→ Atomic Number (Z) :- The no. of protons in the nucleus of an atom. This, when subtracted from A, gives the no. of neutrons

→ Isobars :- Atoms, having the same mass no. but different atomic numbers Eg. $_{15}P^{32}$ and $_{16}S^{32}$

→ Isotones :- Atoms having the same no. of neutrons but different no. of protons or mass number. Eg. $^{14}_{6}C$, $^{16}_{8}O$, $^{15}_{7}N$

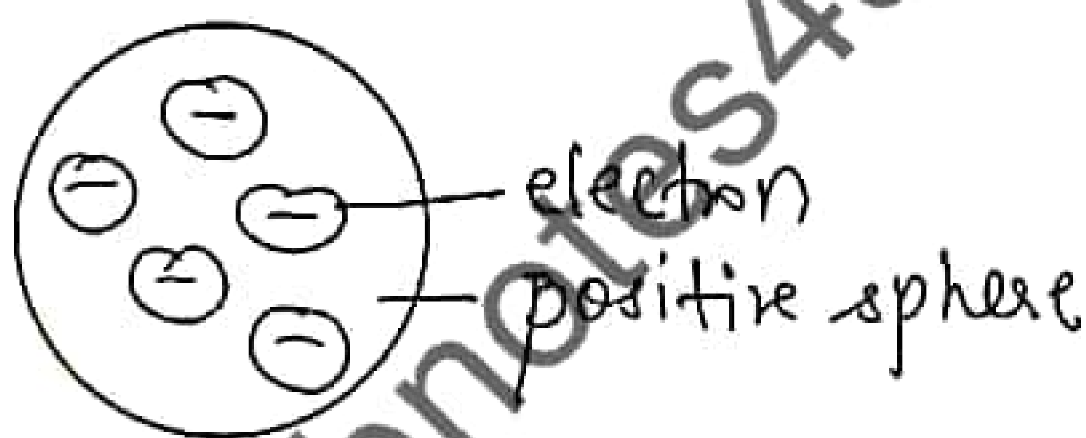
→ Isoelectronic species :- Atoms molecules or ions having the same no. of electrons Eg. $N_2$, $CO$, $CN^-$

→ Nuclear Isomers :- Atoms with the same atomic and mass numbers but different radioactive properties Eg. Uranium X (half life 1.4 min) and Uranium Z (half 6.7 hrs)

# Atomic Models

⇒ Thomson's Model :-

J.J. Thomson, in 1904, proposed that there was an equal and opposite positive charge enveloping the electrons in a matrix. This model is called the plum-pudding model after a type of Victorian dessert in which bits of plums were surrounded by matrix of pudding.




⇒ Rutherford's Model :-

$\alpha$-particles emitted by radioactive substance were shown to be dipositive Helium ions ($He^{++}$) having a mass of 4 units and 2 units of positive charge.

Rutherford allowed a narrow beam of $\alpha$-particles to fall on a very thin gold foil of thickness of the order of 0.0004 cm and determined the subsequent path of these particles with the help of a zinc sulphide fluorescent screen.

⇒ Observation :-

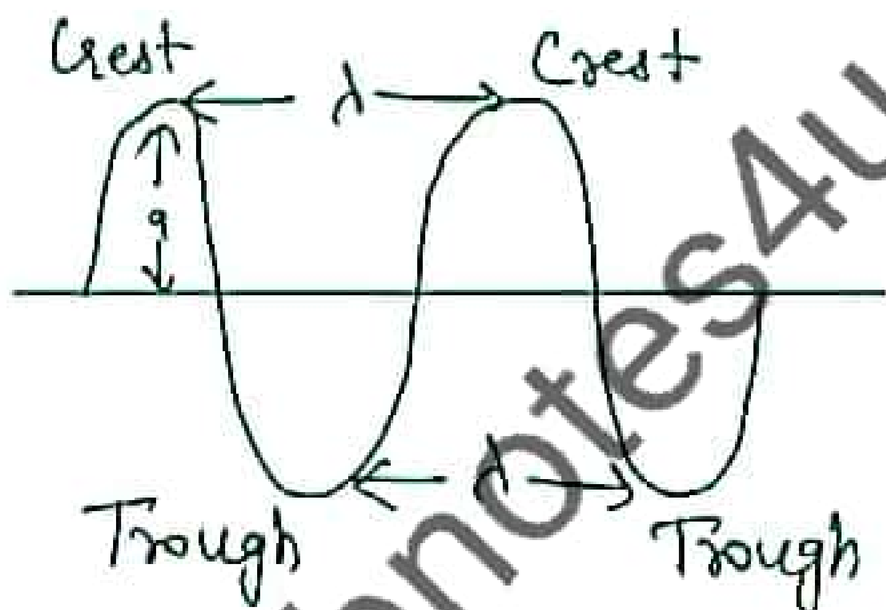
→ Majority of the α-particles pass straight through the gold strip with little or no deflection.
→ Some α- particles are deflected from their path and diverge
→ Very few α- particles are deflected backwards through angles greater than 90°.

⇒ Conclusions :-

→ The fact that most of the α-particles passed straight through the metal foil indicates the most part of the atom is empty
→ The fact that few α- particles are deflected at large angles indicates the presence of a heavy positively charge body i.e., for such large deflections to occur α- particles must have come closer to or collided with a massive positively charged body.
→ The fact that one in 20,000 have deflected at 180° backwards indicates that volume occupied by this heavy positively charged body is very small in Comparison to total volume of the atom.

# Characteristics of a Wave :-

A wave is a sort of disturbance which originates from some vibrating source and travels outward as a continuous sequence of alternating crests and troughs. Every wave has five important characteristics, namely, wavelength ($\lambda$), frequency ($\nu$), velocity (c), wave no. ($\bar{\nu}$) and amplitude (a)




⇒ Wavelength ($\lambda$) :- The distance between two neighbouring troughs or crests is known as wavelength. It is denoted by $\lambda$ and is expressed in cm, m, nanometers ($1\ nm = 10^{-9}\ m$) Armstrong ($1\ \mathring{A} = 10^{-10}\ m$)

⇒ Frequency ($\nu$) :- The frequency of a wave is the no. of times a wave passes through a given point in a medium in one second. It is denoted by $\nu$ (nu) and is expressed in cycles per second (cps) or hertz (Hz)
$1\ Hz = 1\ cps$.
The frequency of a wave is inversely proportional to its wave length ($\lambda$).

$$\nu \propto \frac{1}{\lambda} \quad \text{or} \quad \nu = \frac{c}{\lambda}$$

⇒ Velocity :- The distance travelled by the wave in one second is called its velocity. It is denoted by c and is expressed in cm $sec^{-1}$

$$c = \nu\lambda \quad or \quad \lambda = c/\nu$$

⇒ Wave number ($\bar{\nu}$) :- It is defined as number of wavelengths per cm. It is denoted by $\bar{\nu}$ and is expressed in $cm^{-1}$

$$\bar{\nu} = 1/\lambda \quad or \quad \bar{\nu} = \frac{\nu}{c}$$

⇒ Amplitude :- It is the height of the crest or depth of the trough of a wave and is denoted by (a). It determines the intensity or brightness of the beam of light

⇒ Electromagnetic Spectrum :- The arrangement of the various types of electromagnetic radiation in order of increasing or decreasing wavelengths or frequencies is known as Electromagnetic spectrum.

# # Atomic Spectrum :-

↙ Continuous Spectra

↘ line Spectra

⇒ Continuous Spectra:- When white light from any source such as sun or bulb is analyzed by passing through a prism, it splits up into seven different wide bands of colour from violet to red (like rainbow). These colour are so continuous that each of them merges into the next. Hence, the spectrum is called as continuous spectrum

⇒ Line Spectra:- When an electric discharge is passed through a gas at low pressure light is emitted. If this light is resolved by a spectroscope, it is found that some isolated coloured lines are obtained on a photographic plate separated from each other by dark spaces. This spectrum is called line spectrum.

# Planck's Quantum Theory :-

When a black body is heated, it emits thermal radiations of different wavelengths or frequency. To explain these radiations, Max Planck put forward a theory known as Planck's Quantum theory. The main points of Quantum theory are:-

⇒ Substances radiate or absorb energy discontinuously in the form of small packets or bundles of energy.

⇒ The smallest packet of energy is called Quantum. In case of light the quantum is known as photon

⇒ The energy of a quantum is directly proportional to the frequency of the radiation. $E \propto \nu$ (or) $E = h\nu$ where $\nu$ is the frequency of radiation and h is Planck's constant having the value $6.626 \times 10^{-27}$ erg-sec or $6.626 \times 10^{-34}$ J-sec.

⇒ A body can radiate or absorb energy in whole no. multiples of a quantum $h\nu, 2h\nu, 3h\nu$ ---- $nh\nu$. where 'n' is the positive integer.

⇒ Neils Bohr used this theory to explain the structure of atom.



## ⇒ Bohr's Atomic Model :-

Bohr developed a model for hydrogen atom and hydrogen like one-electron species (hydrogenic species). He applied quantum theory in considering the energy of an electron bound to the nucleus.

## # Important Postulates :-

⇒ An atom consists of a dense nucleus situated at the centre with the electron revolving around it in circular orbits without emitting any energy. The force of attraction between the nucleus and an electron is equal to the centrifugal force of the moving electron.

⇒ Of the finite no. of circular orbits around the nucleus, an electron can revolve only in those orbits whose angular momentum (mvr) is an integral multiple of factor $h/2\pi$

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

where, m = mass of the electron
v = velocity of the electron
n = orbit no. in which electron is present
r = radius of the orbit

⇒ As long as an electron is revolving in an orbit it neither loses nor gains energy. Hence, these orbits are called stationary states. Each stationary state is associated with a definite amount of energy and it is also known as energy levels. The greater the distance of the energy level from the nucleus, the more is the energy associated with it. The different energy levels are numbered as 1, 2, 3, 4 (from nucleus onwards) or K, L, M, N etc

⇒ Ordinarily an electron continues to move in a particular stationary state without losing energy. Such a stable state of the atom is called as Ground state or normal state

⇒ If energy is supplied to an electron, it may jump (excite) instantaneously from lower energy (say 1) to higher energy level (say 2, 3, 4 etc) by absorbing one or more quanta of energy. This new state of electron is called as excited state. The quantum of of energy absorbed is equal to the difference in energies of the two concerned levels. Since the excited state is less stable, atom will lose its energy and come back to the ground state.

Energy absorbed or released in an electron jump, ($\Delta E$) is given by $(\Delta E = E_2 - E_1 = h\nu)$

$\Rightarrow$ $\left[\text{Coulombic force} = \dfrac{kZe^2}{r^2}\right]$

$K = \dfrac{1}{4\pi\epsilon_0}$ (where $\epsilon_0$ is permittivity of free space)

$K = 9 \times 10^9 \ Nm^2 C^{-2}$

In C.G.S. units, value of $k = 1$ dyne $cm^2 (esu)^{-2}$

The centrifugal force acting on the electron is $\dfrac{mv^2}{r}$

Since the electrostatic force balances the centrifugal force, for the stable electron orbit.

$$\frac{mv^2}{r} = \frac{kZe^2}{r^2} \quad \text{——— (1)}$$

or $$v^2 = \frac{kZe^2}{mr} \quad \text{——— (2)}$$

According to Bohr's postulate of angular momentum quantization, we have

$$\left(mvr = \frac{nh}{2\pi}\right)$$

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

$$v^2 = \frac{n^2h^2}{4\pi^2m^2r^2} \quad \text{——— (3)}$$

Equating (2) and (3)

$$\frac{kZe^2}{mr} = \frac{n^2h^2}{4\pi^2m^2r^2}$$

Solving for $r$ we get $r = \frac{n^2h^2}{4\pi^2mkZe^2}$

where $n = 1, 2, 3 \text{ - - - - } \infty$

$r_n = 0.529 \times \frac{n^2}{Z}$ Å

Kinetic energy of the electron $= \frac{1}{2}mv^2$

Potential energy $= \frac{-kZe^2}{r}$

Total energy $= \frac{1}{2}mv^2 - \frac{kZe^2}{r}$ ——— (4)

From eq$^n$ (1), we know that

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

$$\therefore \frac{1}{2}mv^2 = \frac{kZe^2}{2r}$$

Substituting this in equation (4)

$$\text{Total energy } (E) = \frac{kZe^2}{2r} - \frac{kZe^2}{r} = \frac{-kZe^2}{2r}$$

Substituting for $r$, gives us

$$E = \frac{2\pi^2 m Z^2 e^4 k^2}{n^2 h^2} \quad \text{where } n = 1, 2, 3$$

$$E = -21.8 \times 10^{-12} \times \frac{Z^2}{n^2} \text{ erg per atom}$$

$$= -21.8 \times 10^{-19} \times \frac{Z^2}{n^2} \text{ J per atom}$$

$$= -13.6 \times \frac{Z^2}{n^2} \text{ eV per atom}$$

( $1\,eV = 3.83 \times 10^{-23}$ kcal
$1\,eV = 1.602 \times 10^{-12}$ erg
$1\,eV = 1.602 \times 10^{-19}$ J )

$$E = -313.6 \times \frac{Z^2}{n^2} \text{ kcal/mole} \quad (1\text{ cal} = 4.18\text{ J})$$

## ⇒ Calculation of Velocity :-

We know that

$$mvr = \frac{nh}{2\pi}; \quad v = \frac{nh}{2\pi m r}$$

by Substituting for $r$ we get

$$v = \frac{2\pi k Z e^2}{nh}$$

Where except $n$ and $Z$ all are constants

$$V = 2.18 \times 10^8 \frac{Z}{n} \text{ cm/sec.}$$

Eg:- The velocity of electron in the second orbit of $He^+$ will be:-

(A) $2.18 \times 10^6$ m/s (B) $1.09 \times 10^6$ m/s
(C) $4.36 \times 10^6$ m/s (D) None of these

sol$^n$ = $v_n = \frac{v_0 Z}{n}$ Hence, (A) is correct

## ⇒ Hydrogen Atom:-

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

where, $\bar{\nu}$ = wave number

$\lambda$ = wave length

R = Rydberg constant ($109678 \text{ cm}^{-1}$)

$n_1$ and $n_2$ have integral values as follows

Eg:- Find the wavelength of a spectral line produced when an electron in the H-atoms jumps from 4$^{th}$ level to 2$^{nd}$ level

sol$^n$:- $$\frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\frac{1}{\lambda} = 109678\left(\frac{1}{2^2} - \frac{1}{4^2}\right) \text{cm}^{-1} \Rightarrow \lambda = \frac{16}{3 \times 109678} \text{ cm}$$

$$= 4863 \text{ Å}$$

# # Quantum Numbers:-

Quantum numbers may be defined as a set of four numbers with the help of which we can get complete information about all the electrons in an atom. It tells us the address of the electron i.e. location, energy, the type of orbital occupied and orientation of that orbital.

⇒ **Principal Quantum Number (n):-** It tells the main shell in which the electron resides, the approximate distance of the electron from the nucleus and energy of that particular electron. It also tells the maximum no. of electrons that a shell can accomodate is $2n^2$, where n is the principal quantum no.

| Shell | K | L | M | N |
|---|---|---|---|---|
| Principal quantum no. (n) | 1 | 2 | 3 | 4 |
| Maximum no. of electrons | 2 | 8 | 18 | 32 |

⇒ **Azimuthal or Angular momentum quantum no. (l):-** This represents the no. of subshells present in the main shell. These subsidiary orbits within a shell will be denoted as 0, 1, 2, 3, 4 ---- or s, p, d, f ----. This tells the shape of the subshells. The orbital angular momentum of the electron is given as $\sqrt{l(l+1)}\,\frac{h}{2\pi}$ or

$\sqrt{l(l+1)}\,\hbar$ for a particular value of 'n' (where $\hbar = \frac{h}{2\pi}$)

For a given value of n values of possible l vary from 0 to n-1

⇒ The magnetic quantum number (m):- An electron due to its angular motion around the nucleus generates an electric field. This electric field is expected to produce a magnetic field. Under the influence of external magnetic field, the electrons of a subshell can orient themselves in certain preferred regions of space around the nucleus called orbitals. The magnetic quantum no. determines the no. of preffered orientations of the electron present in a subshell. The values allowed depends on the value of l, the angular momentum quantum no., m can assume all integral values between $-l$ to $+l$ including zero. Thus m can be $-1, 0, +1$ for $l=1$. Total values of m associated with a particular value of l is given by $2l+1$

⇒ The spin Quantum Number (s):- Just like earth which not only revolves around the sun but also spins about its own axis, an electron in an atom not only revolves around the nucleus but also spins about its own axis. Since an electron can spin either in clockwise direction or in anticlockwise direction, therefore, for any particular value of magnetic quantum no., spin quantum no. can have two values, i.e. $+1/2$ and $-1/2$ or these are represented by two arrows pointing in the opposite directions, i.e. ↑ and ↓. When an electron goes to a vacant orbital, it can have a clockwise or anticlockwise spin i.e. $+1/2$ or $-1/2$. This quantum no. helps to explain the magnetic properties of the substances.

**Eg:-** If the principal quantum no. n has a value of 3, what are permitted values of the quantum numbers 'l' and 'm'?

**Solⁿ:-** If Principal quantum no. (n) has the value of 3, subsidiary quantum no. (l) will also have three values i.e. 0, 1 and 2 and magnetic quantum number (m) will have $n^2$, i.e. $3^2$ i.e. 9 values in all and they are designated as under:-

| | | |
|---|---|---|
| $n = 3$ | $l = 0$ | $m = 0$ (s) |
| | $l = 1$ | $m = -1$ (p) |
| | | $m = 0$ |
| | | $m = +1$ |
| | $l = 2$ | $m = -2$ (d) |
| | | $m = -1$ |
| | | $m = 0$ |
| | | $m = +1$ |
| | | $m = +2$ |

Eg:- In which orbital the electron will reside if it has the following values of quantum numbers.

(i) $n = 3, l = 1, m = 0$

(ii) $n = 4, l = 0, m = 0$

Solⁿ:- (1) $l = 1$ for p subshell and $m = 0$ for $p_z$ orbital

Hence, electron will residue in $3p_z$ orbital.

(2) $l = 0$ for s subshell. Hence electron will residue in 4s orbital

Eg:- Write all the quantum numbers for the following orbitals

(i) $3d_{z^2}$ (ii) 3s (iii) $4p_x$

Solⁿ:- (i) $n = 3, l = 2, m = 0, s = \pm 1/2$

(ii) $n = 3, l = 0, m = 0, s = \pm 1/2$

(iii) $n = 4, l = 1, m = -1$ or $+1$, $s = \pm 1/2$

# Pauli's Exclusion Principle:-

According to this principle, an orbital can contain a maximum no. of two electrons and these two electrons must be of opposite spin.

Two electrons in an orbital can be represented by

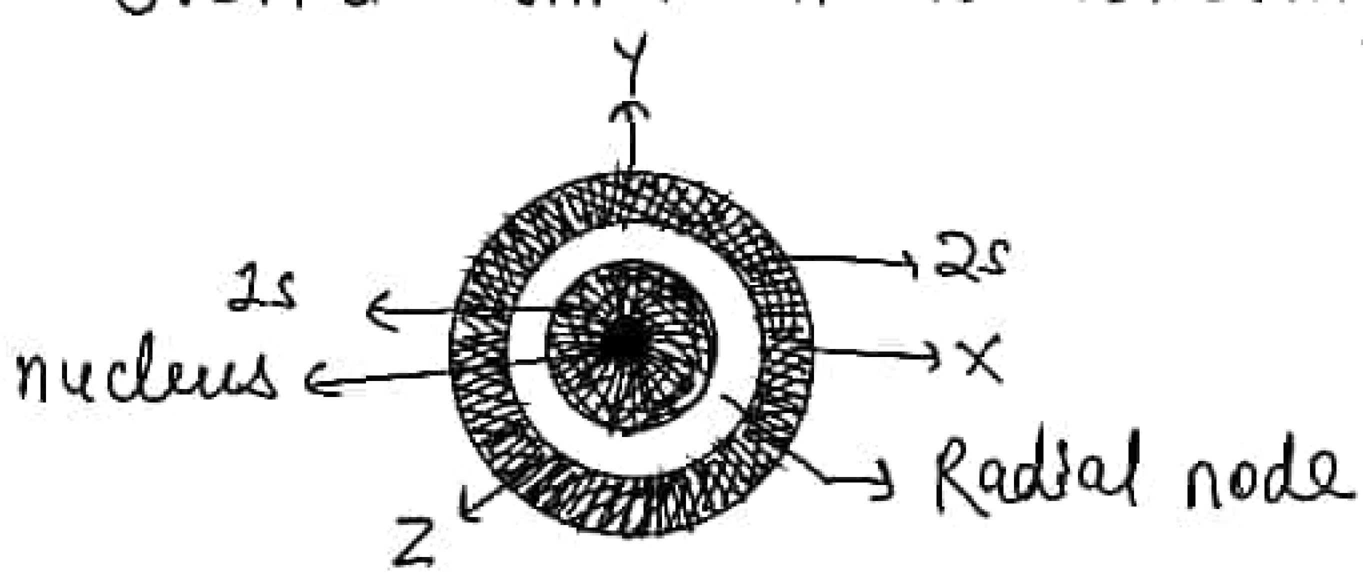
$\boxed{\uparrow\downarrow}$ or $\boxed{\downarrow\uparrow}$

# Shapes and Size of Orbitals :-

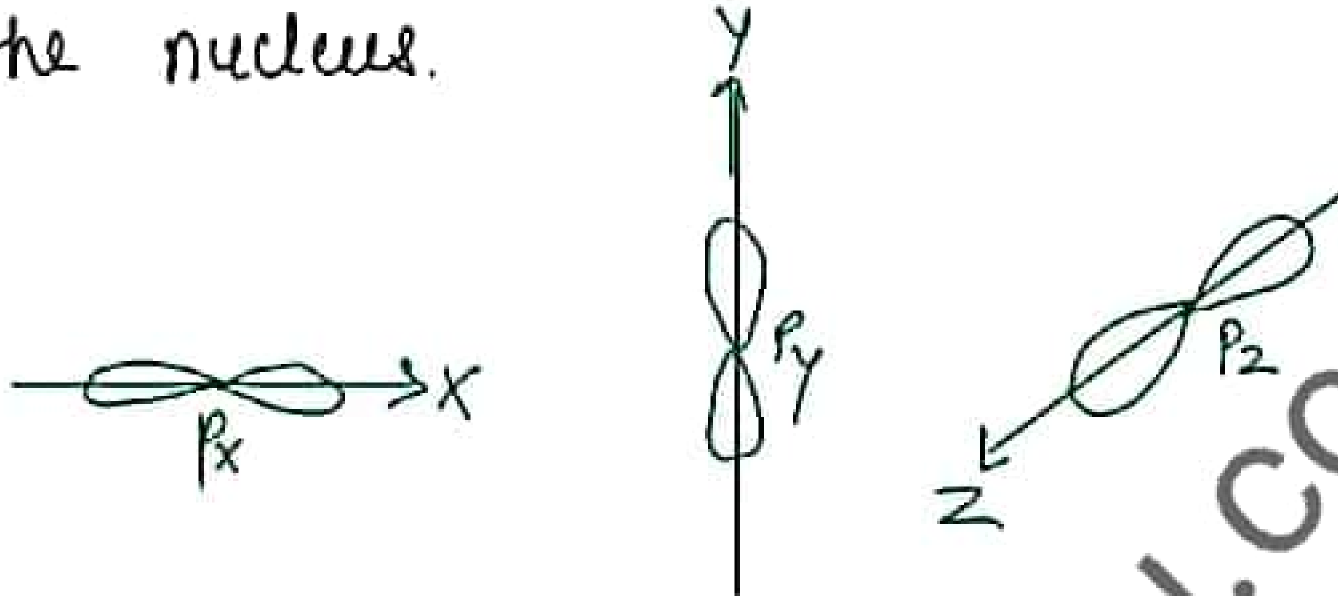
An Orbital is the region of space around the nucleus within which the probability of finding an electron of given energy is maximum (90-95%). The shape of this region (electron cloud) gives the shape of the orbital. It is basically determined by the azimuthal quantum no. l, while the orientation of orbital depends on the magnetic quantum no. (m).

1s 2s

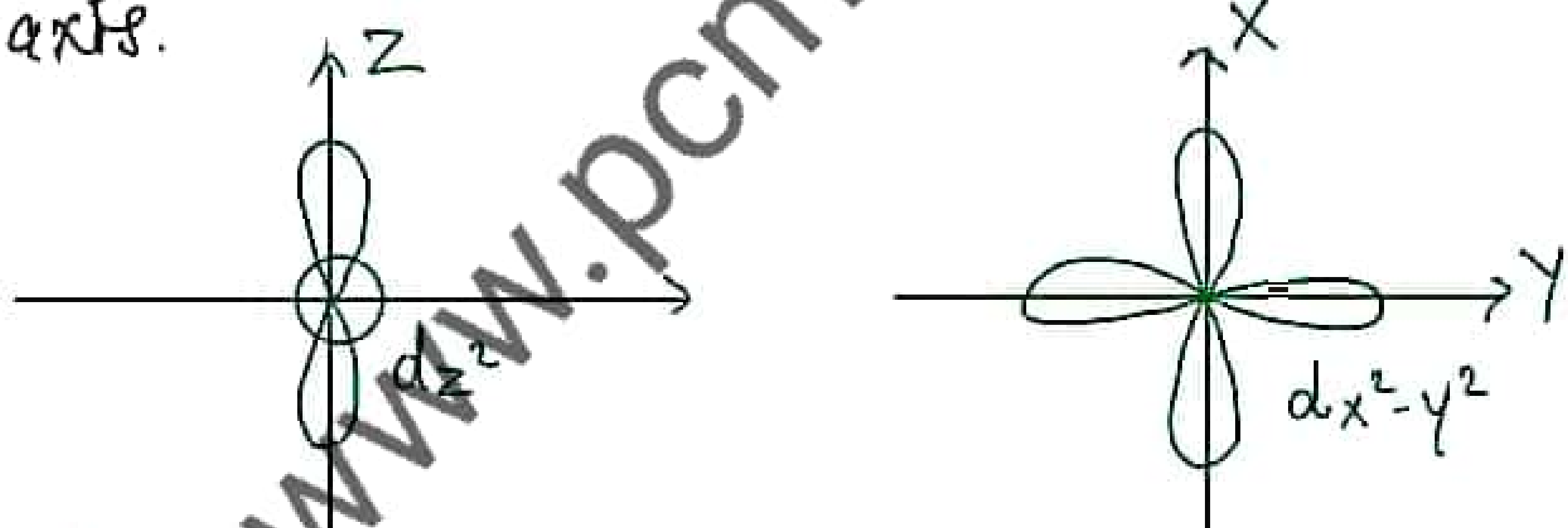
⇒ **S- orbitals :-** The probability of finding the electron is maximum near the nucleus and keep on decreasing as the distance from the nucleus increases. There is vacant space between two successive s-orbitals known as radial node. But there is no radial node for 1s orbital since it is starting from the nucleus.



⇒ **p-orbitals ($l = 1$):-** The probability of finding the p-electron is maximum in two lobes, on the opposite sides of the nucleus.

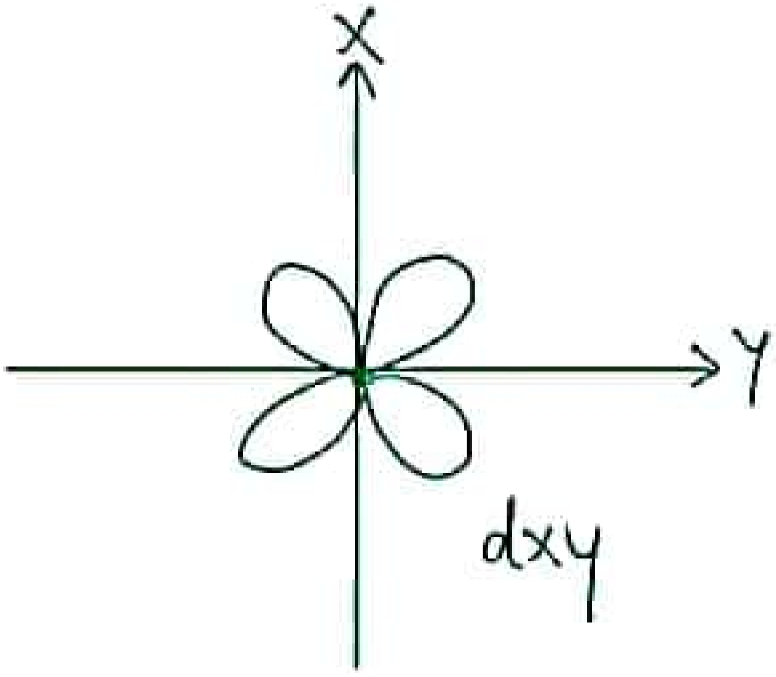


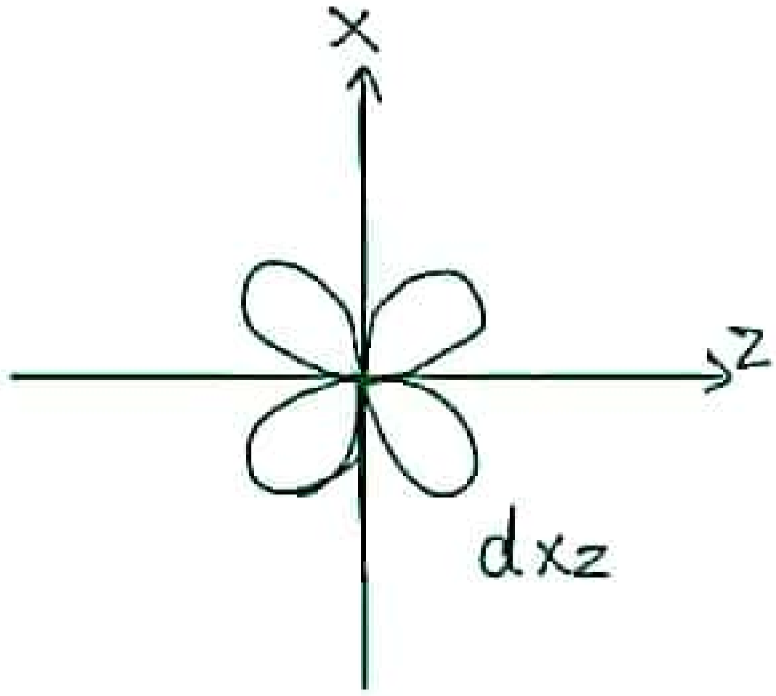

⇒ **d-orbitals ($l = 2$):-** For d-orbitals, $l = 2$, Hence $m = -2, -1, 0, +1, +2$. Thus there are 5d orbitals. They have relatively complex geometry. Out of the five orbitals, the three ($d_{xy}$, $d_{yz}$, $d_{zx}$) project in between the axis and the other two $d_{z^2}$ and $d_{x^2-y^2}$, lie along the axis.

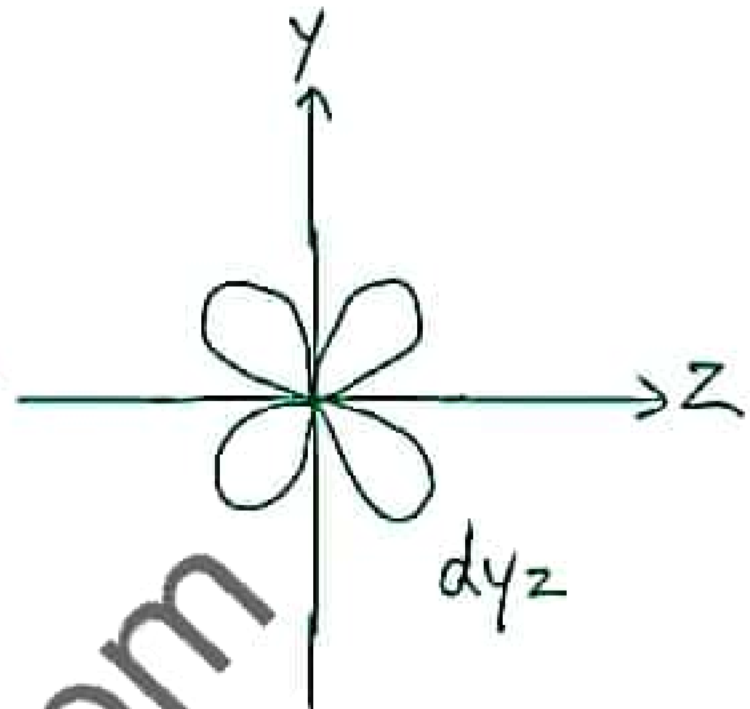



Dough-nut shape or Baby soother shape

Clover leaf shape




Eg:- An electron has a spin quantum no. $\frac{+1}{2}$ and a magnetic quantum no. $-1$. It cannot occupy a/an

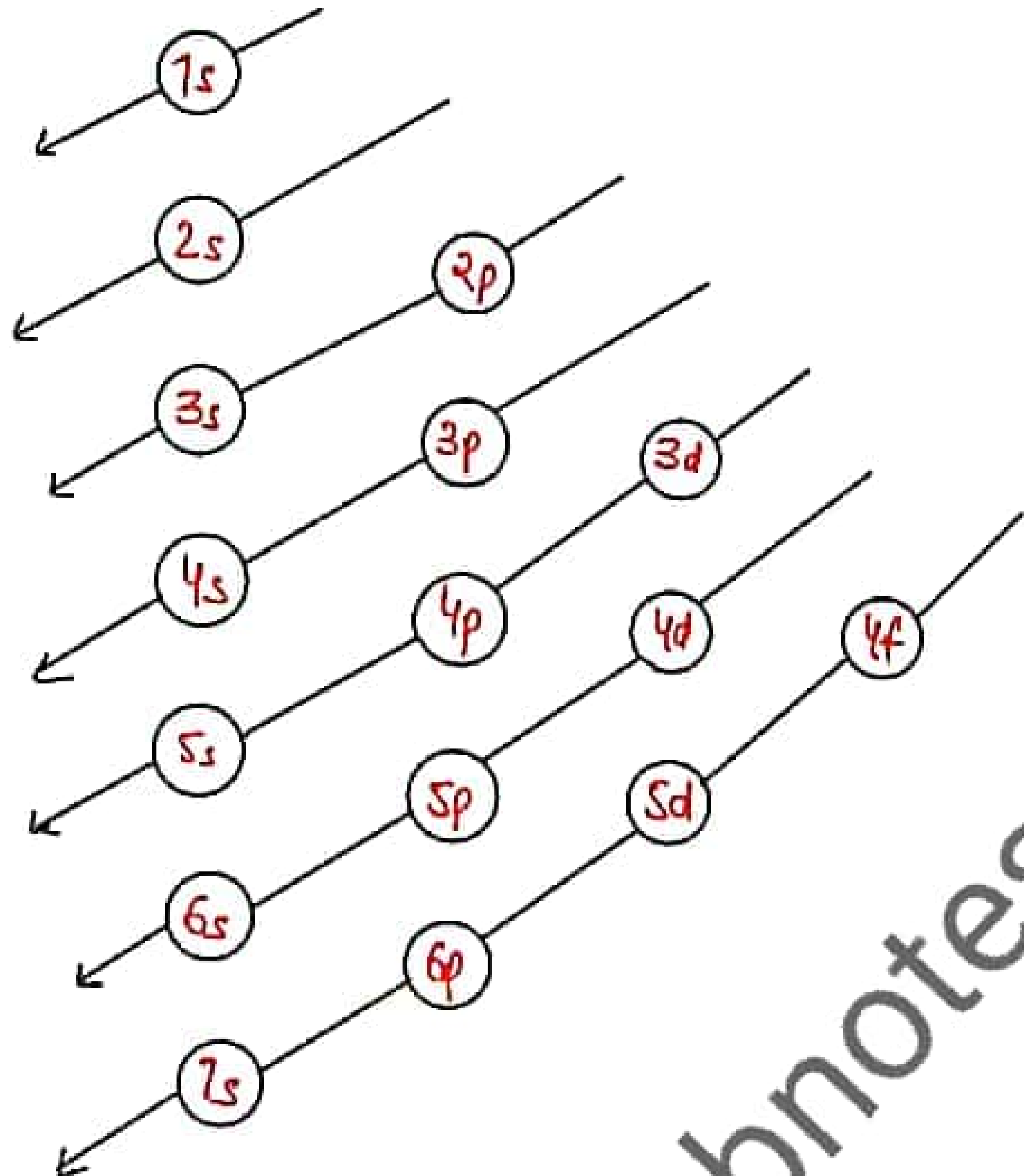
(a) d orbital (b) f orbital
(c) p orbital (d) s orbital

Sol$^n$:- For s orbital $m=0$. Hence, (d) is correct.

# Rules for filling of Electrons in Various Orbitals:-

Aufbau Principle:- This principle states that the electrons are added one by one to the various orbitals in order of their increasing energy starting with the orbital of lowest energy.

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 5f, 6d, 7p --------.






# Hund's Rule of Maximum Multiplicity:-

This rule deals with the filling of electrons in the equal energy (degenerate) orbitals of the same sub shell (p, d and f). According to this rule,

"Electron pairing in p, d and f orbitals cannot occur until each orbital of a given subshell contains one electron each or is singly occupied."

# Electronic Configuration of Elements:-

Electronic configuration is the distribution of electrons into different shells, subshells and Orbitals of an atom.

Keeping in view the above mentioned rules, electronic configuration of any orbital can be simply represented by the notation.

$$n\,l^{x}$$

- Principal quantum no. — $n$
- Symbol of subshell or orbitals (s, p, d, f) — $l$
- No. of electrons in the subshell — $x$

Eg:- Write the electronic configuration of Nitrogen (atomic no. = 7)

Sol$^n$:-

| $1s^2$ | $2s^2$ | $2p^3$ | | |
|---|---|---|---|---|
| ↑↓ | ↓↑ | ↑ | ↑ | ↑ |

## ⇒ Exceptional Configurations :-

**Stability of Half filled and Completely filled Orbitals:-**

Cu has 29 electrons. Its expected electronic configuration is
$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^9$

But a shift of one electron from lower energy 4s orbital to higher energy 3d orbital will make the distribution of electron symmetrical and hence will impart more stability.

Thus, the electronic configuration of Cu is
$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1\,3d^{10}$

Fully filled and half filled orbitals are more stable

Eg:- Which of the following metals has the highest value of exchange energy required for exchange stabilization!
(a) Mn (b) Cr
(c) Cu (d) Zn

Solⁿ:- Cu has full filled d-subshell. Hence, (C) is correct.

# Dual Character (Particle and Wave Character of matter and Radiation) :-

light has dual nature ( wave + particle)

## ⇒ de- Broglie Equation:-

The wavelength of the wave associated with any material particle was calculated by analogy with photon.

$E = h\nu$ —(1) (acc. to the Planck's quantum theory)

$E = mc^2$ — (2) (acc. to Einstein's equation)

$$h\nu = mc^2$$

but $\nu = c/\lambda$

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

(or) $\lambda = h/mc$

The above $eq^n$ is applicable to material particle, if the mass and velocity of photon is replaced by the mass and velocity of material particle. Thus, for any material particle like electron.

$\lambda = h/mv$ or $\lambda = h/p$ where $mv = p$ is the momentum of the particle.

# Relation Between Kinetic Energy and Wavelength:-

$$K.E\,(E) = \tfrac{1}{2}mv^2$$

$$v = \sqrt{\frac{2E}{m}} = \lambda = \frac{h}{mv}$$

$$\lambda = \frac{h}{\sqrt{2Em}}$$

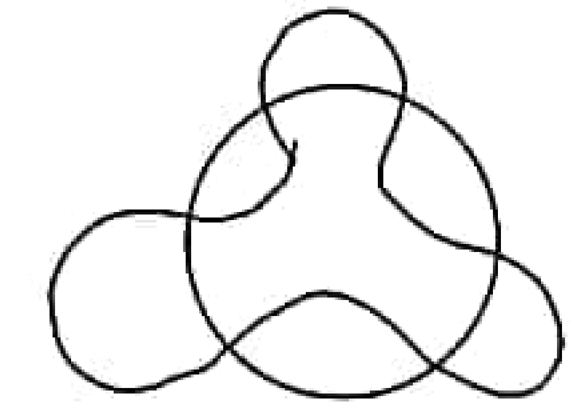
## ⇒ Derivation of Angular Momentum from de Broglie $Eq^n$:-

According to Bohr's model, the electron revolves around the nucleus in circular orbits. Acc. to de Broglie concept, the electron is not only a particle but has a wave character also. If the wave is completely in phase, the circumference of the orbit must be equal to an integral multiple of wave length ($\lambda$)

Therefore, $2\pi r = n\lambda$

But $\lambda = h/mv$

or $mvr = n\, h/2\pi$



No. of waves 'n' $= \frac{2\pi r}{\lambda} = \frac{2\pi r}{h/mv} = \frac{2\pi mvr}{h}$

where, $v$ and $r$ are the velocity of electron and radius of that particular Bohr orbit

Thus, the no. of revolutions per second is $\frac{v}{2\pi r}$

Eg:- An $e^-$, a proton and an alpha particle have K.E of 16 E, 4 E and E respectively. What is the qualitative order of their Broglie wavelengths?

(a) $\lambda e > \lambda p > \lambda a$  (b) $\lambda e > \lambda p = \lambda a$

(c) $\lambda p < \lambda e < \lambda a$  (d) $\lambda a < \lambda e = \lambda p$

Solⁿ = $\lambda = \frac{h}{\sqrt{2mK.E}}$ Hence, (B) is correct.

⇒ Distinction between the wave-particle nature of a photon and the particle - wave nature of a sub atomic particle:-

| Photon | Sub-Atomic Particle |
|---|---|
| ① Energy $= h\nu$ | Energy $= \frac{1}{2}mv^2$ |
| ② Wavelength $= \frac{c}{\nu}$ | Wavelength $= \frac{h}{mv}$ |

Eg:- What is the de-Broglie wavelength of electron having K.E. of 5eV?

sol$^n$ :- $K.E. = \frac{1}{2} mv^2$

$$v = \sqrt{\frac{2KE}{m}}$$

Now, $\lambda = \frac{h}{mv}$

$$= \frac{h}{\sqrt{2K.E\,m}} = \frac{6.62 \times 10^{-34}}{\sqrt{2 \times 5 \times 1.6 \times 10^{-19} \times 9.1 \times 10^{-31}}}$$

$$= 5.486 \times 10^{-10}\ m$$

⇒ **Heisenberg's Uncertainty Principle :-**

This Principle States:-

It is impossible to measure simultaneously the position and momentum of a small microscopic moving particle with absolute accuracy or certainty" i.e. if an attempt is made to measure any one of these two quantities with higher accuracy, the other becomes less accurate

$\Delta x . \Delta p \geq h/4\pi$    $\Delta x$ - change in position
$\Delta p$ - Change in momentum

Eg:- What is the uncertainity in the position of a ball of mass 10 g and which is moving with a velocity of 100 m/s with 0.002 % uncertainity?

Sol$^n$:- $\Delta x \times \Delta v = \frac{h}{4\pi m}$

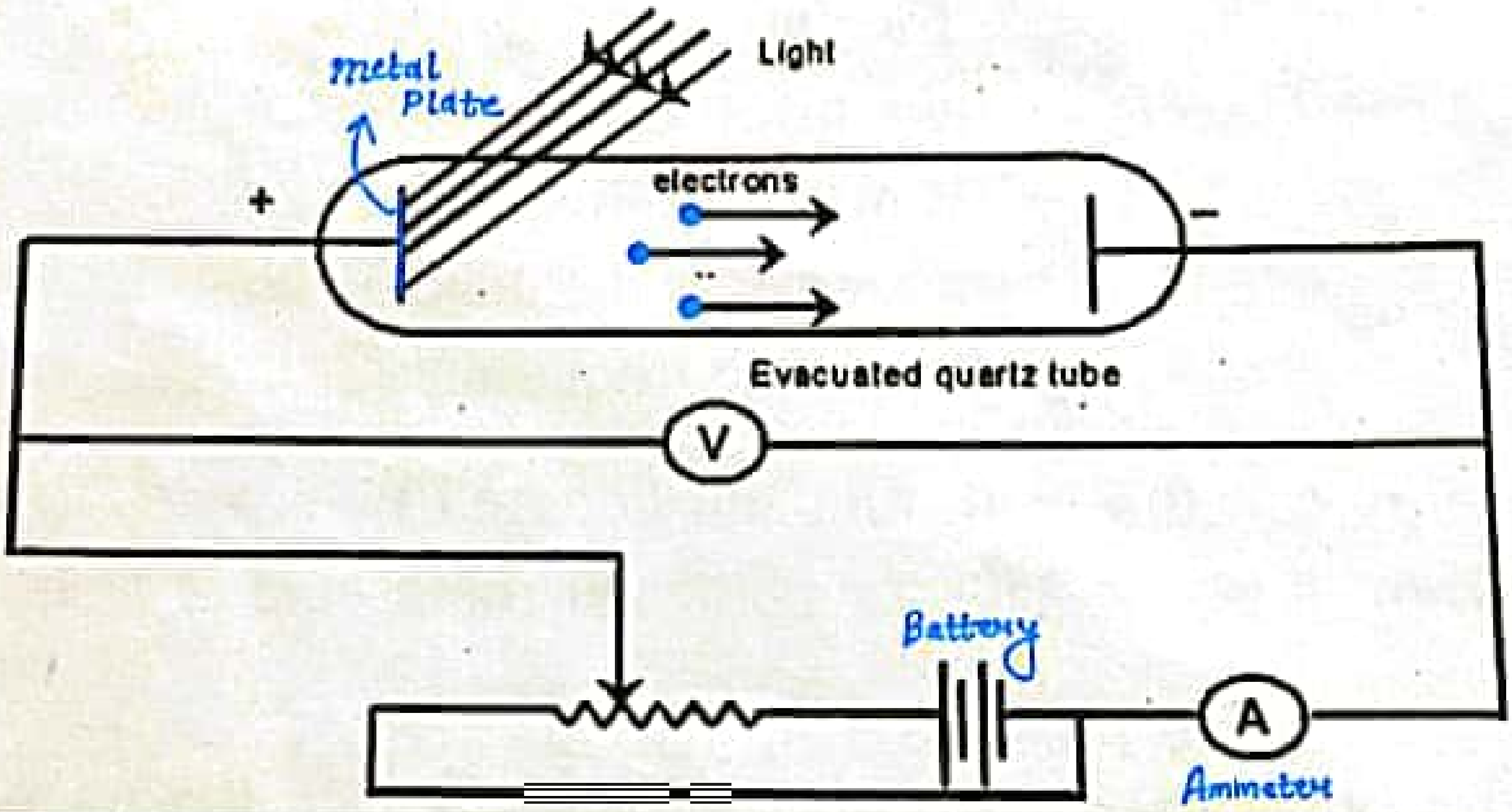
Now $\Delta v = \frac{100 \times 0.002}{100} = 0.002$ m/s

$$\Delta x = \frac{6.626 \times 10^{-34}}{4 \times 3.14 \times 9.1 \times 10^{-31} \times 0.002} = 0.0289 \text{ m}$$

# Photoelectric Effect :-

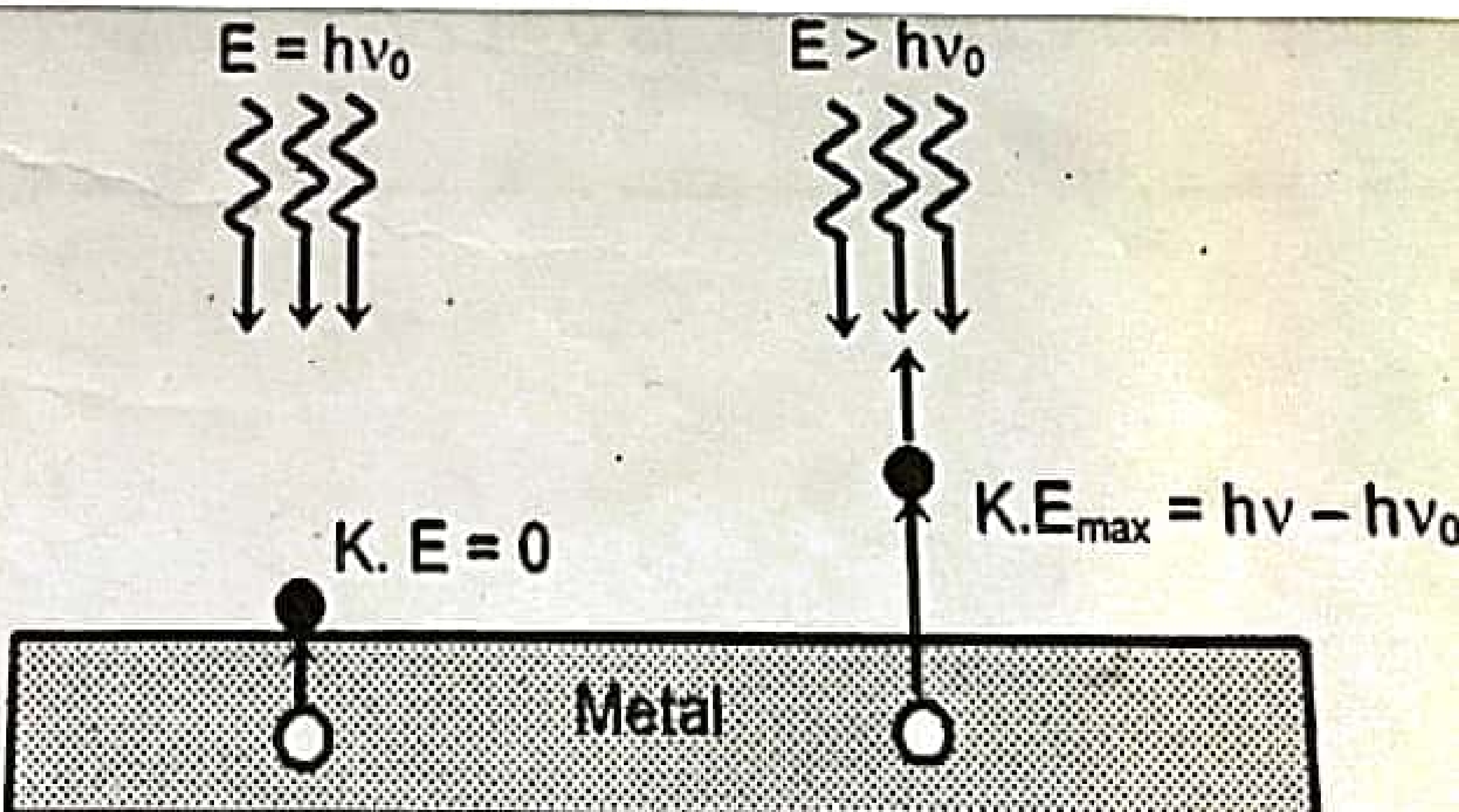
Sir J.J. Thomson, observed that when a light of certain frequency strikes the surface of a metal electrons are ejected from the metal. This phenomenon is known as Photoelectric effect and the ejected electrons are called Photoelectrons.

A few metals, which are having low ionization energy like Cesium, show this effect under the action of visible light but many more show it under the action of more energetic ultraviolet light




ns come out as soon as the light
energy) strikes the metal surface. Th
frequency, called the threshold (or cr
, which can just cause the ejection.
varies with the nature of the m
the frequency of the light, the me
electrons have. Blue light results in fa
light.





$$h\nu = h\nu_0 + K.E.$$

$$h\nu = h\nu_0 + \frac{1}{2}mv^2$$

$$\frac{1}{2}mv^2 = h\nu - h\nu_0$$

where, $\nu$ = frequency of the incident light
$\nu_0$ = threshold frequency

$h\nu_0$ is the threshold energy (or) the work function denoted by $\phi = h\nu_0$ (minimum energy of the photon to liberate electron. It is constant for particular metal and is also equal to the ionization potential of gaseous atoms.

→ The kinetic energy of the photoelectrons increases linearly with the frequency of incident light.


Eg:- Work function of sodium is 2.5eV. Predict whether the wavelength 6500 Å is suitable for a photoelectron ejection or not.

sol$^n$ :- Energy of incident light $= \frac{hc}{\lambda} = \frac{6.62 \times 10^{-34} \times 3 \times 10^{8}}{6500 \times 10^{-10}}$

$= 3.058 \times 10^{-19}$ J

$= 1.9$ eV

which is lower than work function. Hence no ejection will take place.