Ch-2-
Structure OF Atom

* Discovery of electron or cathode-ray discharge tube experiment.

(i) Construction :-

→ A partially evacuated tubes named as cathode ray discharge tubes were taken to study electrical discharge.

→ Cathode ray tube is made of glass containing 2 thin pieces of metal called electrodes through which electrical discharge could take place.

→ A flourescent screen was also put behind the anode.

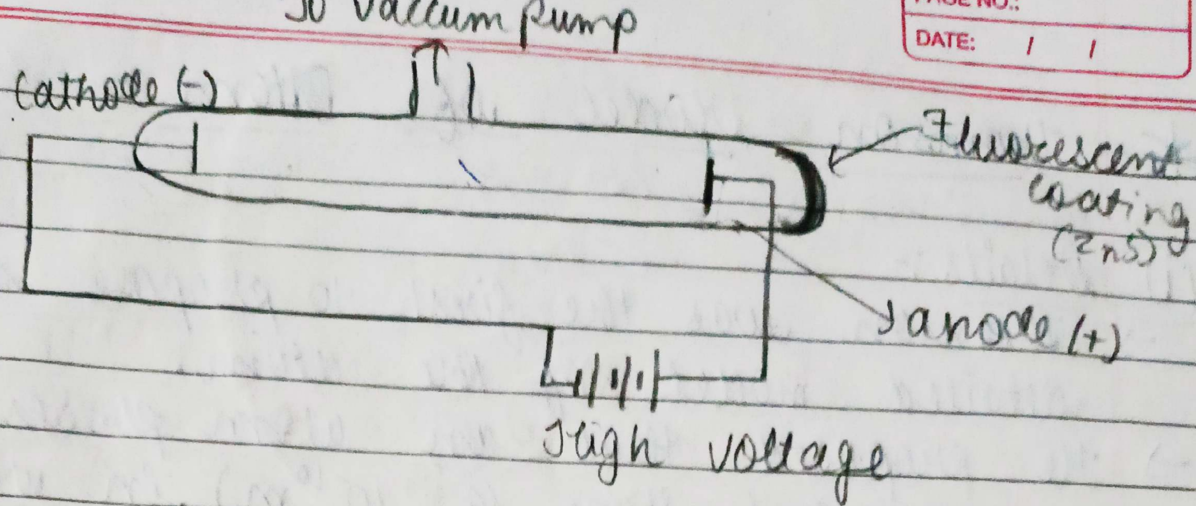
(ii) Working :-

→ The electrical discharge through the gases could be observed only at very low pressures and at very high voltages.

→ When sufficiently high voltage is applied, current starts flowing from cathode to anode. These were called cathode rays.

→ When these rays, passing through anode, strike the zinc sulphide coating, a bright spot is seen.

PAGE NO.:
DATE: / /



(iii) **Results :-**

→ The cathode rays start from cathode and move towards the anode.
→ These rays are not visible but their behaviour can be observed on fluorescent screen.
→ In absence of magnetic or electrical field, these rays travel in straight line.
→ Charachteristics of cathode rays don't depend upon the material of electrode & the nature of gas.

(iv) **Conclusion :-**

→ In the presence of magnetic or electrical field, cathode rays behave like negatively charged particles.
→ Therefore, cathode rays consist of −ve charged particles called electrons.

⇒ Thus, electrons are basic constituent of all atoms.

* Thomson Model of Atom :-

(i) Details :-

→ Thomson was the first to propose a detailed model of the atom.

→ He proposed that an atom possesses a spherical shape ($\approx 10^{-10}$ m) in which the positive charge is uniformly distributed.

→ Electrons are embedded in such a manner to give the most stable electrostatic arrangement.

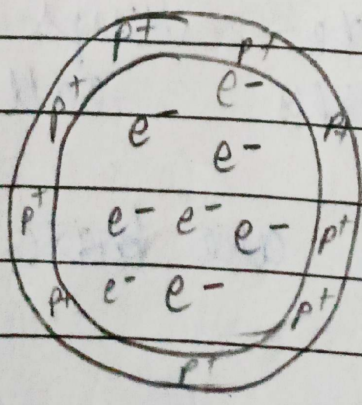
→ Different names are plum pudding, raisin pudding or watermelon.

(ii) Drawbacks (Limitations) :-

→ The main limitation is that the mass of the atoms is considered to be evenly spread over that atom.

→ It is a static model. It does not reflect the movement of electron.

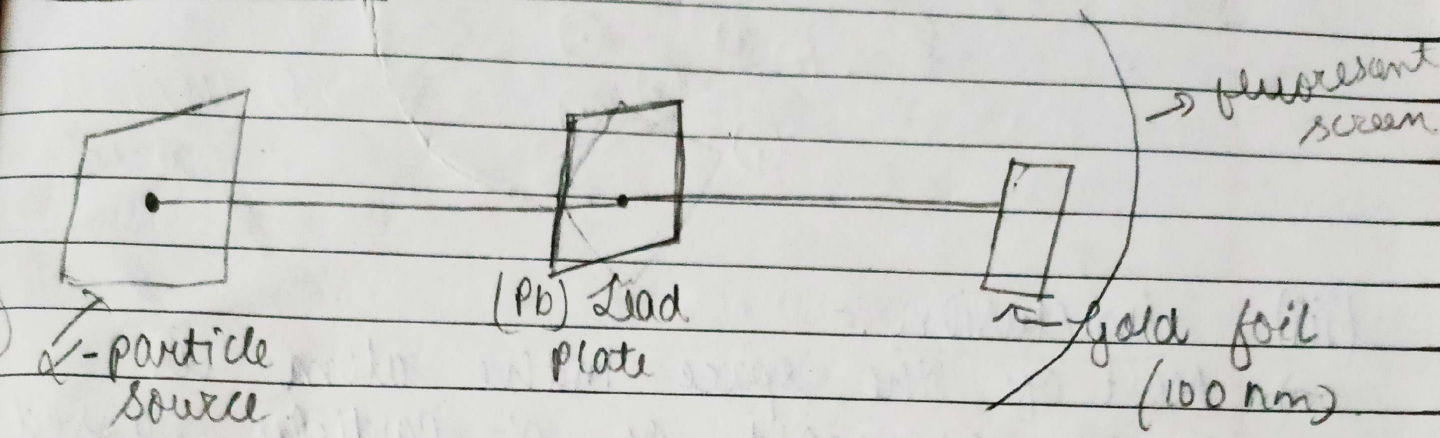
→ It couldn't explain the stability of an atom.



PAGE NO.:
DATE: / /

* Rutherford's Nuclear Model of Atom or α-particle scattering experiment

[i] Construction :-



→ He conducted a gold foil experiment where he bombarded a very thin gold foil with α-particles.

→ Thickness of gold foil is ~100 nm with α-particle in an evacuated chamber.

→ Around the foil, a circular fluorescent screen is coated with zinc sulphide.

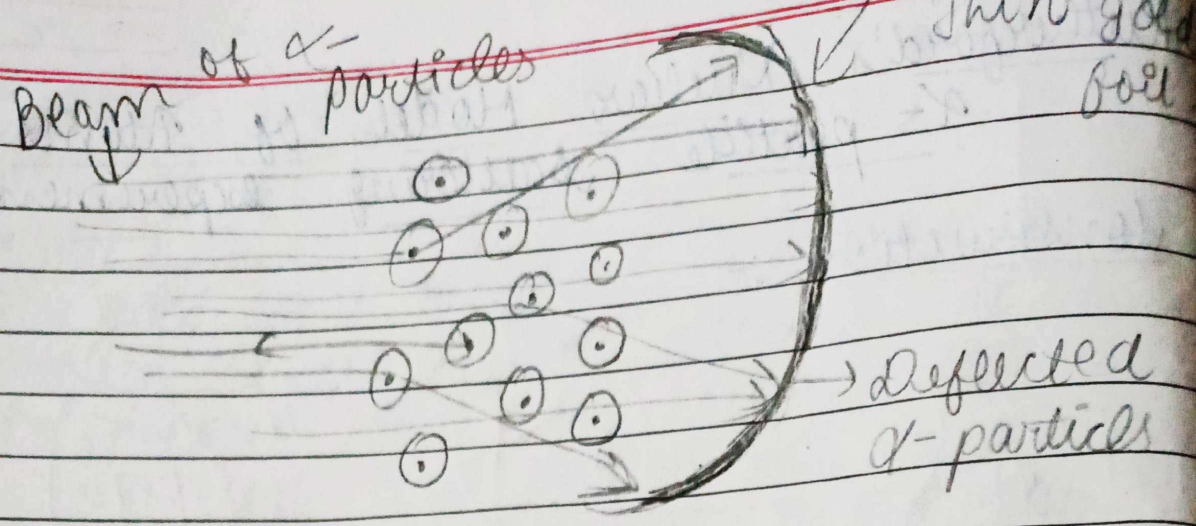
→ Whenever α-particles struck the screen, a tiny flash of light produced at that point.

[ii] Observation :-

→ most of the α-particles passed through the gold foil undeflected.

→ a small fraction of α-particles were deflected by small angles.

→ a very few α-particles (~ 1 in 20,000) bounce back.



(iii) Conclusion:-

→ Most of the space in the atom is empty as most of α-particles passed undeflected.

→ A few positively charged α-particles are deflected due to high concentration of positive charge in the centre.

→ The size of an atomic nucleus is negligible compared to the size of an atom. Atomic radius is about 10^{-10} m while that of nucleus is 10^{-15} m.

(iv) Rutherford proposed the nuclear model of atom, that was as follows:-

→ + charge & most mass of atom is concentrated in the center in an extremely small region. This was called nucleus by Rutherford.

→ Electrons move around the nucleus with a high speed in a circular path called an orbit.

DATE: / /

→ Electrons and nucleus are held together by electrostatic forces of attraction.

(v) Limitations :-

→ This model could not explain the motion of electrons in circular orbitals.

→ A charged particle (e^-) would release energy and should come closer to nucleus following a spiral path and eventually fall into the nucleus.

* Atomic Number (Z):-

→ No. of protons present.

→ Denoted by Z.

→ Neutral atom :- No. of e^- = No. of p^+.

→ For an ion :- No. of e^- = Z − (charge on ion).

* Mass Number (A) :-

→ Mass Number :- No. of neutrons + No. of protons

= No. of nucleons.

→ No. of neutrons = A − Z

• A is always a whole number.

• Representation of an Atom :-

→ $^{A}_{Z}X$

→ A → mass no.

Z → atomic no.

* Isotopes :-

→ Atoms of element which have the same atomic number but different mass number.

eg. :- $^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$ → radioactive

	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
p →	6	6	6
e →	6	6	6
n →	6	7	8

Same chemical property but different physical property

eg :- $^{1}_{1}H$, $^{2}_{1}H$, $^{3}_{1}H$

Protium (H), Deuterium (D), Tritium

	$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
p →	1	1	1
e →	1	1	1
n →	0	1	2

→ Neutron is not available in protium.

Note :-

1) Isotopes have the same no. of protons but differ in the no. of neutrons.

* Isotones :-

→ They are the atoms of different element having same no. of neutrons

eg. $^{3}_{1}H$ $^{4}_{2}H$

	$^{3}_{1}H$	$^{4}_{2}H$
p →	1	2
e →	1	2
n →	2	2

$^{39}_{19}K$ $^{40}_{20}Ca$

	$^{39}_{19}K$	$^{40}_{20}Ca$
p →	19	20
e →	19	20
n →	20	20

→ $A_1 - Z_1 = A_2 - Z_2$

* Isobars :-

→ Atoms of different element having same mass number but different atomic no.
→ They have different no. of electrons, protons & neutrons but sum of no. of neutron & protons remain same.

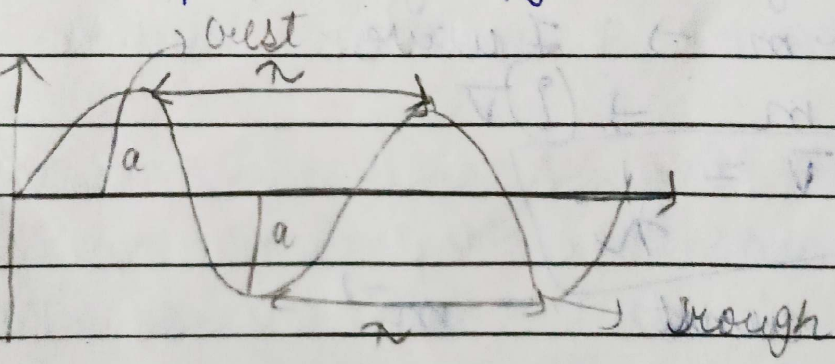
eg :-

	$^{3}_{1}H$	$^{3}_{2}He$
p =	1	2
e =	1	2
n =	2	1
p+n =	3	3

→ Don't have same chemical & physical property.

* electromagnetic waves :- (Radiant Energy)

→ These waves do not need any medium for propogation.
→ They have electric & magnetic fields & travel at right angle to these fields.
→ upper most point → crest
lower most point → trough.



(1) Frequency (ν) :- $\frac{\text{No. of waves}}{1 \text{ second}}$

→ SI unit → $\frac{1}{sec}$, s^{-1} → Hz

$\Rightarrow$ T → 1 wave
1 sec → wave? (ν)

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

② Time period (T) :-
→ Time taken by a wave.
→ Per wave → Time (second).
→ No. of waves (ν) = 1 sec.
1 = ? (T)

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

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

③ Wave length (λ) :- λ
→ distance covered by one oscillation (one crest + one trough).
→ SI unit → 'm'.

$$\lambda = \frac{c}{\nu}$$ → m/s = 3×10^8

m

④ Wave Number ($\bar{\nu}$) :-
→ No. of wavelength per unit time.
→ λ m → 1 wave
1 m → (?) $\bar{\nu}$

$$\bar{\nu} = \frac{1}{\lambda}$$

→ SI unit = m^{-1}

PAGE NO.:
DATE: / /

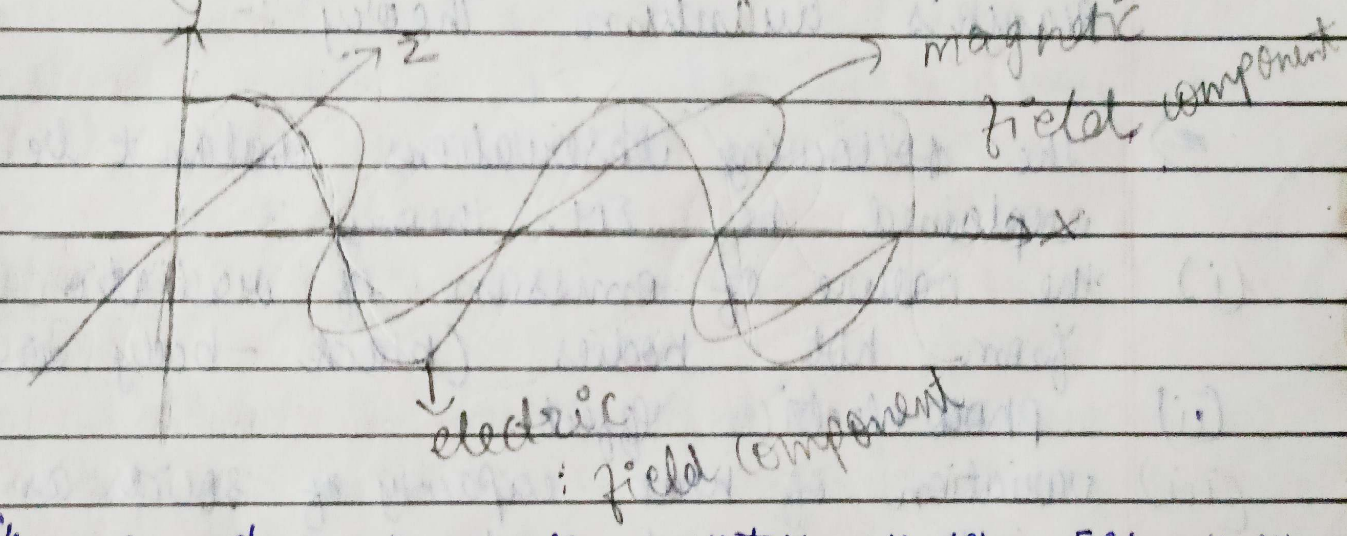
(James Maxwell).

* Wave Nature of EM Radiations:-

→ When electrically charged particle undergoes acceleration, alternating electrical and magnetic fields are produced & transmitted. These fields are transmitted in the form of waves called electromagnetic waves or EM radiations.

→ Properties of EM wave :-

(i) The oscillating electric & magnetic fields produced by oscillating charged particles are perpendicular to each other & also to the direction of propogation of wave.



(ii) Unlike sound waves or water waves, EM waves do not require medium & can move in vaccum.

(iii) There are many types of EM radiations, which differ from one another in wavelength (λ) (or frequency (ν)). These consititute which is called EM - spectrum.

Radio frequency : 10^6 Hz (broadcasting).
Microwave region : 10^{10} Hz (radar)
Infrared region : 10^{13} Hz (heating)

- Ultraviolet region : 10^{16} Hz (Component of Sun's radiation)
- Visible light : 10^{15} Hz (Component of Sun's radiation)

(iv) Different kinds of units are used to represent EM radiation.
→ These radiations → properties ⇒ frequency (ν), wavelength (λ) & no. of wave ($\bar{\nu}$).

* Particle nature of EM radiation: Planck's Quantum Theory :-

⇒ The following observations couldn't be explained by EM. theory :-
(i) the nature of emission of radiation from hot bodies (black-body radiation)
(ii) photoelectric effect.
(iii) variation of heat capacity of solids as a function of temperature.
(iv) Line spectra of atoms.

(1) Black - Body Radiation :-

⇒ ~~Black-body~~ An ideal body which emits and absorbs radiations of all frequencies uniformly is called a black body & the radiation emitted by such a body is called black body radiation.

DATE: / /

→ At a given temperature, intensity of radiation emitted increases with the increase of wavelength, reaches at maximum value at a given wavelength & then starts decreasing with further increase in wavelength.

→ Max Planck said that :-

(i) Atoms & molecules could emit or absorb energy in discrete quantities only, which are called quantum.

(ii) Energy of a quantum is proportional to its frequency as $E = h\nu$

h = Planck's constant ⇒ 6.626×10^{-34} Js.

(iii) Energy is always emitted in integral multiples of $h\nu$ as $2h\nu$, $3h\nu$, etc. In other words, energy is quantised.

* Photoelectric effect :-

→ When a light of certain frequency strikes the surface of a metal, e^- are ejected or given out from the metal surface.

→ This phenomenon of ejection of e^- from metal surface under the influence of incident radiation is called photoelectric effect.

→ The e^- ejected are called photo electrons.

→ According to wave theory, both the no. of e^- & their energies should depend upon the brightness (intensity) of light, but it is found that only the no. of e^- ejected depends upon the brightness (intensity) of light while their energies don't.

→ Einstein told that light consists of streams of particles called photons which move with the speed of light.

- **Observations :-**

(i) Only photons of light of certain minimum frequency called threshold frequency (ν_0) can cause the photoelectric effect. The value of ν_0 is different for every metal.

$$\text{Total energy} = \text{(threshold) work fn.} + \text{kinetic energy}$$

$$E = E_0 + K.E.$$

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

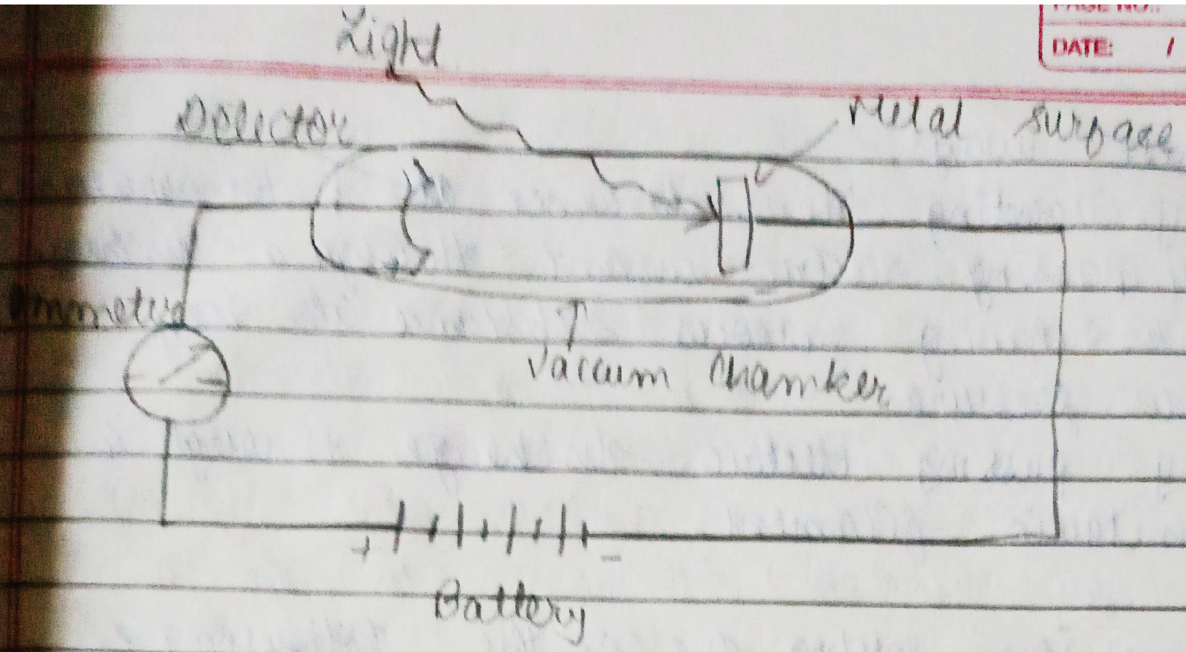
$$h(\nu - \nu_0) = \frac{1}{2} m_e v^2$$

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

(iii) $\frac{1}{\lambda}$ ∝ frequency change ↑ ∝ e^- velocity.

i.e. $\uparrow E \propto \nu\uparrow \propto \frac{1}{\lambda\uparrow}$

⇒ If the frequency of incident light (ν) is more than threshold frequency (ν_0), the excess energy ($h\nu - h\nu_0$) is imparted to e^- as kinetic energy.



* Dual behaviour of EM spectrum :-

→ It was concluded that light & EM radiations have dual nature i.e. particle & wave.
→ Radiation interacts with matter → particle like properties
Propogates → Wave like properties.

* Spectrum :-
→ A ray of white light spread out into a series of coloured bands is called spectrum.
→ Combination of different v or λ.
→ Arrangement of radiations in ↑ order of wavelength or ↓ order of v.

① Emission Spectrum :-
→ Atom, molecule & ions emit radiations after being excited i.e. after absorbing energy. The spectrum of such emitted radiations is called emission spectrum.

Normally a substance can be excited in any of

these ways:-
(a) By heating the substance at ↑ temperature
(b) By passing electric current through a discharge tube having gaseous substance at very low pressure.
(c) By passing electric discharge through a metallic filament.

• Emission spectra are of the following 2 types:-
(i) Continuous spectrum &
(ii) line spectrum or atomic spectrum.

(i) Continuous spectrum:-
→ When a ray of white light is analysed by passing through a prism, it is observed that it splits in 7 wide bands of colours like rainbow. They were so continuous that each of them merges into the next. ∴ This is called continuous spectrum.

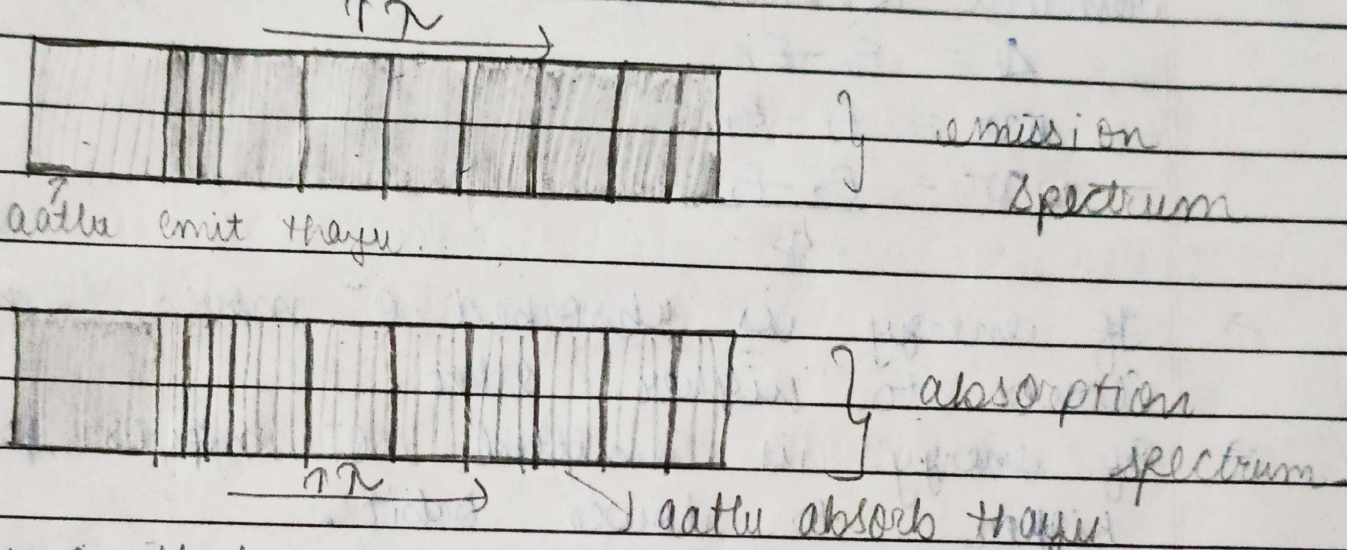
(ii) Line spectrum or atomic spectrum:-

→ When an atom of a substance is excited; it emits radiations. These radiations are analyzed with spectroscope & many fine bright lines of specific colours in a sequence are seen in the spectrum, which is not continuous i.e. dark zone in b/w any

PAGE NO.:
DATE: / /

② **Absorption spectrum :-**

→ An absorption spectrum is like the photographic negative of an emission spectrum. A continum of radiation is passed through a sample which absorbs the radiation of certain wavelengths. The missing λ (which correspond to radiations absorbed) leave dark spaces in bright continuous spectrum.



part emitted ✓ → ✗ not absorbed
part emitted ✗ → ✓ part absorbed

* **Bohr's Atomic Model :-**

• Introduction :-

→ Neil Bohr was the first to propose model for the hydrogen atom & describe arrangements of e^- in atom.

→ This model is applicable only for single e^- species like H, He^+, Li^{2+}, etc.

• Postulates :-

(i) Electron revolve around the nucleus in fixed circular path of definite energy called stationary orbits.

(ii) e^- can move from one stationary orbit to another either by absorbing or emitting energy. The amount of energy absorbed or emitted by e^- remains constant called transition of e^-.

$$\Delta E = E_2 - E_1$$
$$h\nu = E_2 - E_1$$
$$\nu = \frac{E_2 - E_1}{h}.$$

→ If energy is absorbed e^- moves from lower to higher orbit.

→ If energy is emitted e^- moves from higher to lower orbit.

(iii) e^- can revolve only in those orbits in which angular momentum of e^- is integral multiple of $\frac{h}{2\pi}$

i.e. $m_e v r = \frac{nh}{2\pi}$

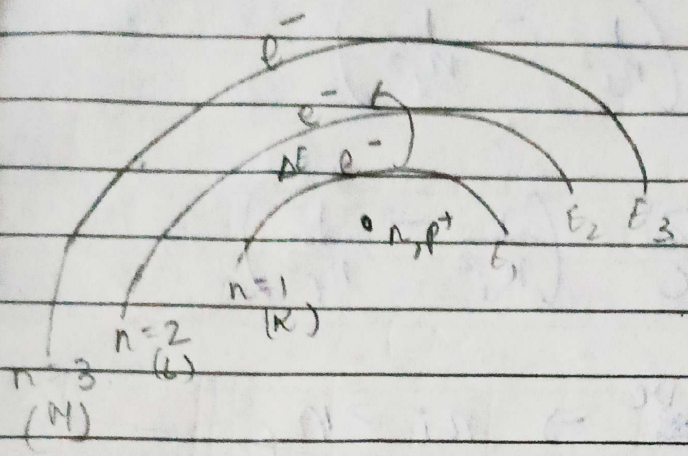
(iv) The stationary states for e^- are numbered $n = 1, 2, 3, \ldots$ These integral nos. are known as principal quantum nos.

(v) The radii of the stationary states are expressed as :-

PAGE NO.:
DATE: / /

$r_n = \frac{n^2 a_0}{Z}$ where, $a_0 = 52.9\ pm$
n = no. of orbit
Z = atomic no.

→ As 'n' increases the value of r will ↑.



(vi) energy of orbit (e⁻) is termed as below :-

$$E_n = \frac{-\pi^2 Z^2 m_e e^4}{n^2 h^2}$$

$$E_n = -R_H \frac{Z^2}{n^2}$$ where, $R_H = 2.18 \times 10^{-18}\ J$

→ The negative sign indicates e⁻ present in binding state i.e. e⁻ is hold by proton with energy.

(vii) e can move from one energy state to another either by absorption of energy or emission of energy. The energy difference is ΔE which can be given as:-

$$\Delta E = E_f - E_i$$

$$= \left(\frac{-R_H Z^2}{n_f^2}\right) - \left(\frac{-R_H Z^2}{n_i^2}\right)$$

$$\Delta E = R_H Z^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$h\nu = R_H Z^2 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\nu = \frac{R_H Z^2}{h} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\frac{c}{\lambda} = \frac{R_H Z^2}{h} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\bar{\nu} = \frac{1}{\lambda} = \frac{R_H Z^2}{hc} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

→ Here energy would be $-ve$ when → $n_i > n_f$
high → low
outer → inner
energy emission

→ Here energy is $+ve$ when → $n_i < n_f$
low → high
inner → outer
energy absorption

- **Limitations:-**

(i) It couldn't explain the atomic spectra of atoms containing more than one e^- i.e. multielectron atoms.

(ii) Failed to explain fine structure of atom (spectral lines).

(iii) No explanation of Zeeman effect (magnetic field) & Stark effect (e^- field).

PAGE NO.:
DATE: / /

(iv) Can't explain 3-D existence of atoms.

(v) Not in accordance with Heisenberg's uncertainity principle.

* Line spectrum of Hydrogen :- JEE

Series	n_i	n_f	Spectral region
Lyman	1 ← (Kα, Kβ)	2, 3, 4, 5, ...	ultraviolet
Balmer	2	3, 4, 5, 6 ...	Visible
Paschen	3	4, 5, 6, 7 ...	Infrared
Brackett	4	5, 6, 7, 8 ...	Infrared
Pfund	5	6, 7, 8, 9 ...	"
Humphrey	6	7, 8, 9, 10 ...	"

JEE 1)

		n_i		n_f
K_α	⇒	1	←	2
K_β	⇒	1	←	3
N_α	⇒	3	←	4
N_β	⇒	4	←	6

Outer → ΔE max $\propto \nu$ max $\propto \frac{1}{\lambda}$ (min).

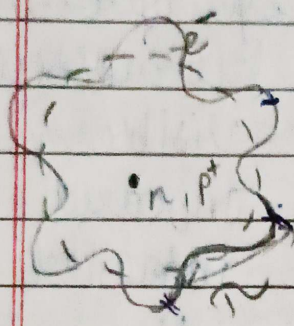
inner → ΔE min $\propto \nu$ min $\propto \frac{1}{\lambda}$ (max)

2) min / max ΔE / ν / λ

A) K_α	A) 2→1	A) Ly. 1st	ΔE min
B) K_β	B) 3→1	B) Ly. 2nd	
C) K_γ	C) 4→1	C) Ly. 3rd	
D) K_δ	D) 5→1	D) Ly. 4th	ΔE max

* **Dual behaviour of matter / De Broglie's Principle :-**

⇒ De Broglie proposed that matter like radiation should also have dual nature i.e. particle & wave. Which means that like photons, e^- also have momentum as well as wavelength.



$$n\lambda = 2\pi r \quad \text{— (1)}$$

$$\lambda = \frac{2\pi r}{n}$$

also,

$$\lambda \propto \frac{1}{p}$$

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

$$\lambda = \frac{h}{m_e v} \quad \text{— (2)}$$

(v → velocity, m_e → mass of electron)

⇒ From (1) & (2).

$$\frac{2\pi r}{n} = \frac{h}{m_e v} \Rightarrow m v r_n = \frac{nh}{2\pi}$$

⇒ $\lambda = \frac{h}{mv}$ can also be derived as :-

⇒ According to planck $E = h\nu$

" " Einstein $E = mc^2$

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

For a particle of mass 'm' moving with velocity

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

PAGE NO.:
DATE: / /

→ It is clear from the above eq. that the value of $\lambda \downarrow$ when $p\uparrow$ (either m or v or both). The wavelength of many fast-moving objects like aeroplane, ball, etc is very low because of their high mass.

* Heisenberg's Uncertainty Principle :-

→ It states that, "It is impossible to determine simultaneously, the exact position & exact momentum (or velocity) of an e^-.

→ Mathematically,

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

$$\Delta x \cdot m_e \Delta v \geq \frac{h}{4\pi}$$

$$\Delta x \cdot \Delta v \geq \frac{h}{4\pi m}$$

Δx = uncertainty in position
Δp = uncertainty in momentum
h = Planck's constant

● Significance of Uncertainty Principle :-

(i) It rules out existence of definite paths or trajectories of e^- & other similar particles.
(ii) The effect of this principle is significant only for microscopic objects & is negligible ~~of~~ for macroscopic particles.

→ If Δx small, Δv large.
Δx large, Δv small

* Quantum Mechanical Model of Atom :-

- Quantum Mechanics :- A theoretical science that deals with the study of motions of the microscopic objects that have both observable wave like & particle like properties.

→ For a system (such as an atom or a molecule) whose energy doesn't change with time, the Schrödinger eqⁿ is as :-

$$\hat{H}\Psi = E\Psi$$, where

$\hat{H}$ is a mathematical operator.
↳ Hamiltonian.

⇒ $\hat{H}$ = Kinetic energy operator ($\hat{T}$) + potential energy operator ($\hat{V}$)

⇒ $\hat{H} = \hat{T} + \hat{V}$.

* Important features of Quantum Mechanical Model of Atom:-

(i) The energy of e^- in atom is quantised i.e. can only have certain values.

(ii) The existence of quantized electronic energy level is direct result of wave like properties of e^- & are allowed solⁿ of Schrodinger eqⁿ.

(iii) Both the exact position & exact velocity of an e^- in an atom can't be determined simultaneously.

→ An atomic orbital has wave function ψ. There are many orbitals in an atom. e^- occupy an atomic orbital which has definite energy. An orbital → not more than 2 e^-. Orbitals filled in ↑ order of energy.

→ Probability of finding an e^- at a point in an atom ∝ $|\psi|^2$ (orbital wave fn.)

→ From the value of ψ^2 at different points within an atom, it is possible to predict the region around the nucleus where e^- most probably will be found.

Orbitals and Quantum Nos.

To obtain complete info. about an e^- in an atom 4 identification nos. are required & these nos. are called as quantum nos.

Principal quantum no. (n) ⟶ shell (orbit).
Azimuthal quantum no. (l) → sub shell (orbital)
Magnetic " " (m) → sub orbital
Spin " " (s) → spin of e^-

Principal quantum no. (n) :-
Gives the avg. distance of the e^- from nucleus.
Completely determines the energy of e^- in hydrogen atom or hydrogen like particles.
max. no. of e^- present in any shell is given by $2n^2$.

Principal quantum no. (n)	→	1	2	3	4
Letter designation	→	K	L	M	N
Max. no. of e^- $(2n^2)$	→	2	8	18	32

→ Orbit no. (n)

① $r_n = \frac{n^2 a_0}{Z}$

② $E_n = -R_H \frac{Z^2}{n^2}$

② Azimuthal / Subsidiary / Orbital angular / Secondary Quantum no. :-

→ This helps to explain the fine lines of the spectrum.

→ It gives info. about :-

(i) No. of subshells present in main shell.

(ii) angular momentum of e^- present in subshell

(iii) shape of subshell.

→ Denoted by l.

→ l = 0 1 2 3 4 5

notation = s (spherical), p (dumb bell), d (double dumb bell), f, g, h (complex)

→ For a given value of n, it can have any value ranging 0 to (n-1).

→ $l = 0$ to $(n-1)$

$l = 0$ → 1s

$l = 0, 1$ → 2s, 2p

$l = 0, 1, 2$ → 3s, 3p, 3d

$l = 0, 1, 2, 3$ → 4s, 4p, 4d, 4f.

DATE: / /

(3) Magnetic Quantum no. (m) :-

→ It determines the no. of preferred orientations of the e^- present in a subshell.
→ Denoted by m.
→ For a given value of l, it can have all the values ranging from -l to l. i.e.
-l . . . 0 . . . +l

s	l	m
s	0	0
p	1	-1, 0, 1 i.e. p_x, p_y, p_z
d	2	-2, -1, 0, 1, 2 i.e. d_{xy}, d_{yz}, d_{zx}, $d_{x^2-y^2}$, d_{z^2}
f	3	-3, -2, -1, 0, 1, 2, 3

→ For any sub-shell (l), 2l+1 values of m are possible:-

→ s → 1.
p → 2(1)+1 = 3 ⇒ p_x, p_y, p_z
d → 2(2)+1 = 5 → d_{xy}, d_{yz}, d_{zx}, $d_{x^2-y^2}$, d_{z^2}
f → 2(3)+1 = 7 → complicated

(4) Spin quantum no. (m_s) :-

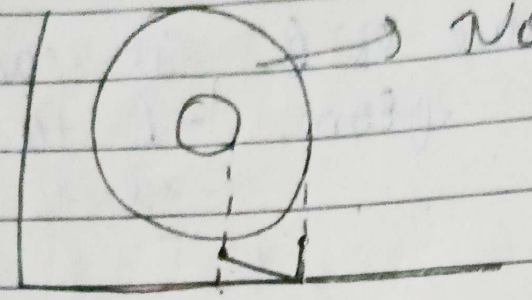
→ The spin could have 2 orientations given by m_s which can take values of $+\frac{1}{2}$ & $-\frac{1}{2}$
→ These are called 2 spin states of e^- represented by ↑ & ↓
→ 2 e^- having different m_s values are said to have opposite spins.
→ An orbital can't hold more than 2 e^- & these 2 e^- should have opp. spins.

* Shape of Orbitals :-

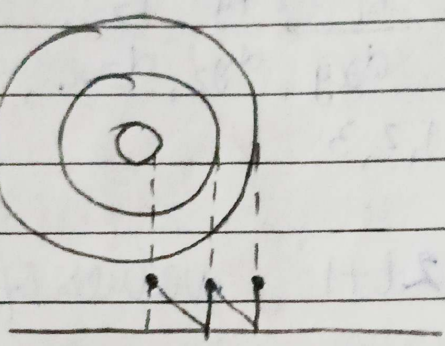
(1) Shape of S - orbital :- Spherical.

$n = 1, l = 0$

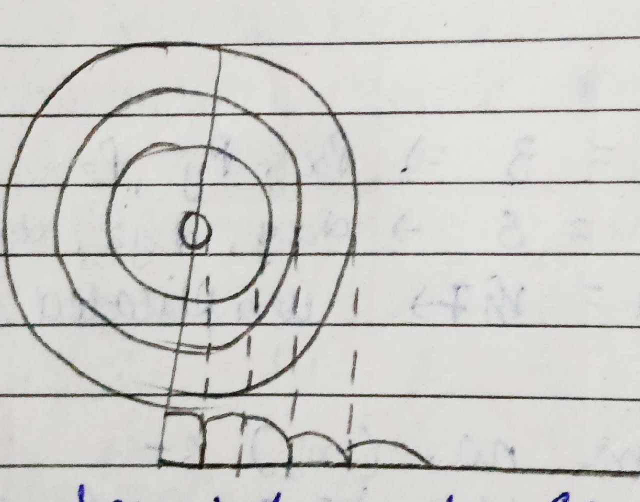
1s



2s $n = 2, l = 1$



3s.



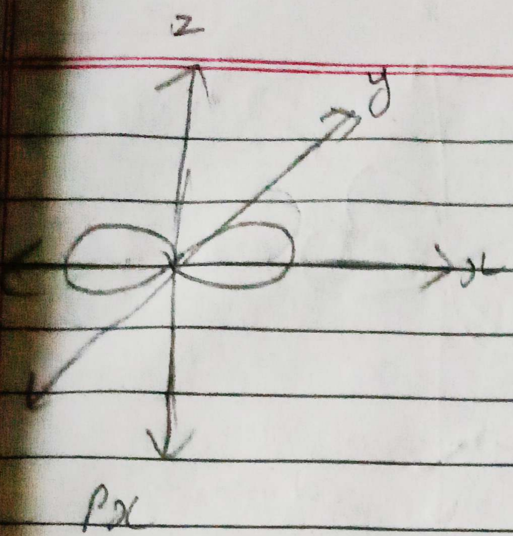
⇒ Nodes ∃ Spherical nodes Or radical node

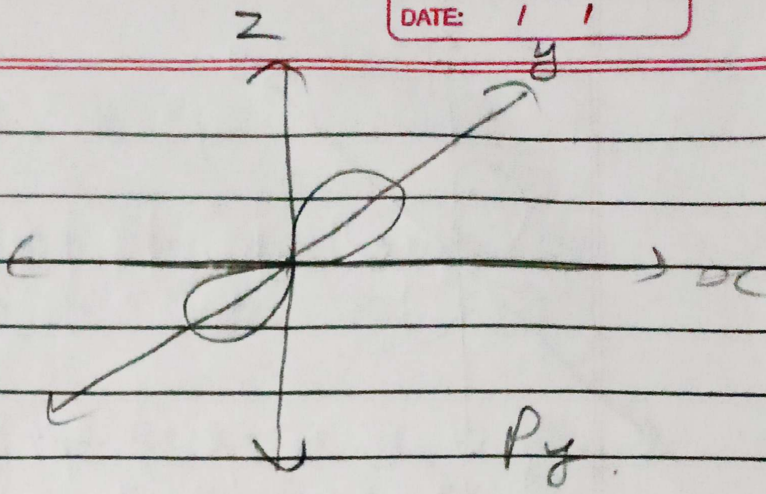
$= (n - l - 1)$ eg :- 4p

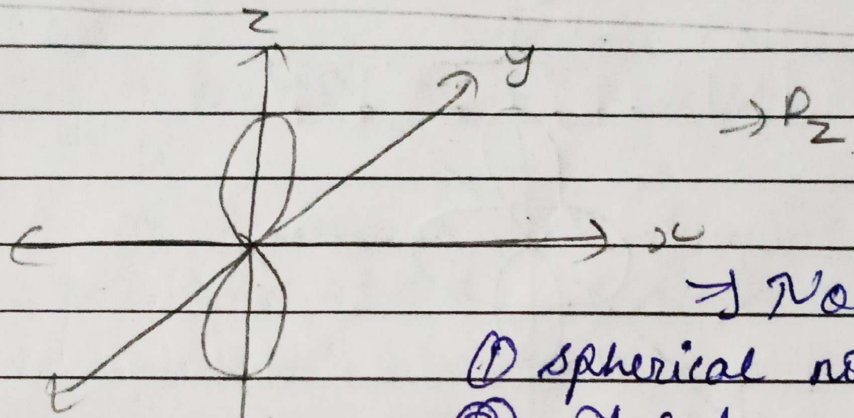
$= (4 - 1 - 1) = \boxed{2}$

* Shape of p - orbitals :- (Dumb bell)

p_x, p_y, p_z.







→ Nodes.

(1) spherical nodes = $(n-l-1)$

(2) Angular nodes = l.

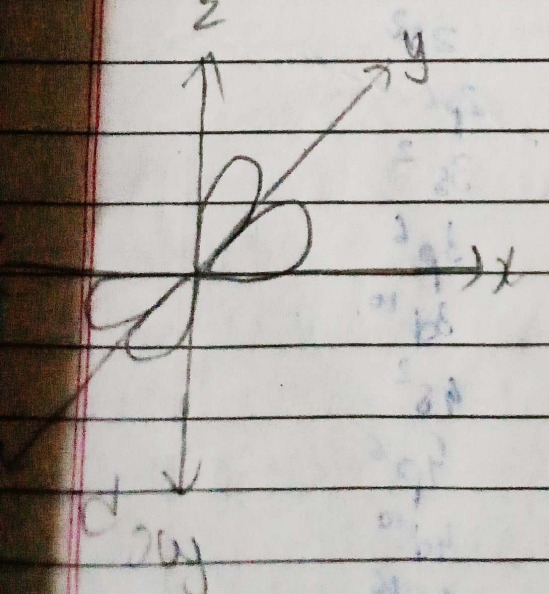
Total = $n - l - 1 - l$

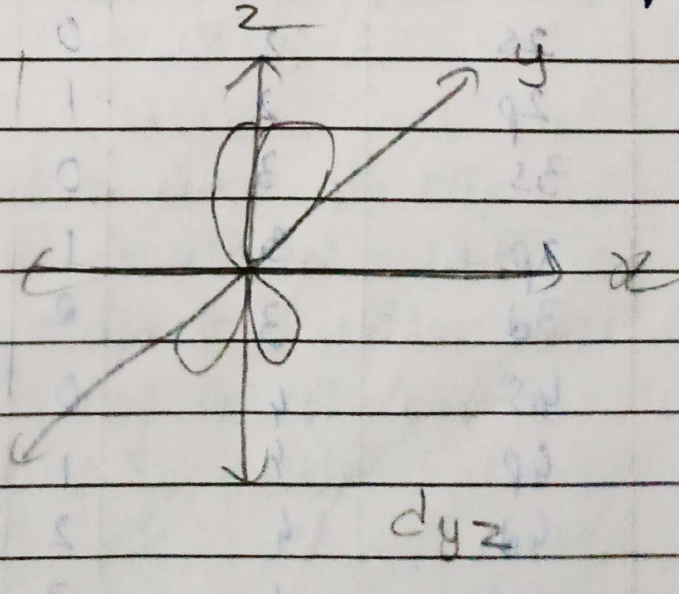
= $(n-1)$.

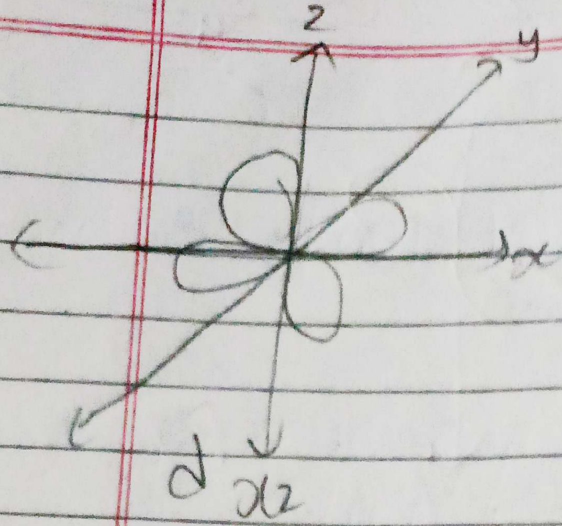
(8) Shape of d- orbitals:- double dumb bells.

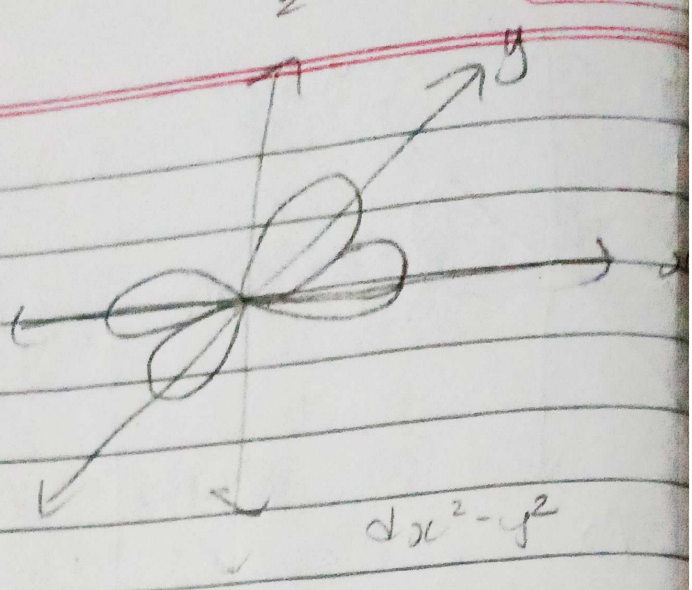
→ d_{xy}, d_{yz}, d_{zx}, $d_{x^2-y^2}$, d_{z^2}. except d_{z^2},

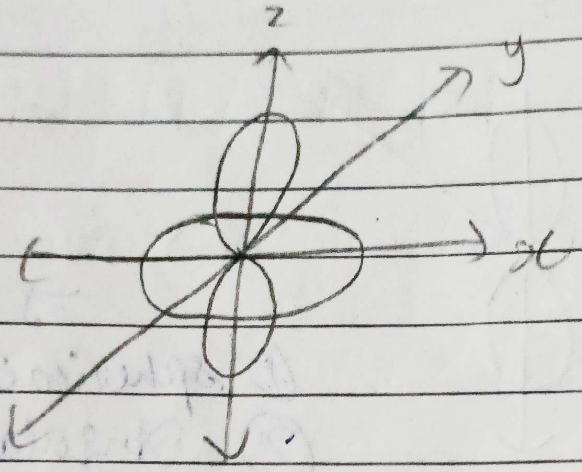
→ donut / peanut











* Energy of orbital :-

name of Orbital	value of n	l	n+l	Energy
1s	1	0	1	$1s^2$
2s	2	0	2	$2s^2$
2p	2	1	3	$2p^6$ (Px = 2, Py = 2, Pz = 2)
3s	3	0	3	$3s^2$
3p	3	1	4	$3p^6$
3d	3	2	5	$3d^{10}$
4s	4	0	4	$4s^2$
4p	4	1	5	$4p^6$
4d	4	2	6	$4d^{10}$
4f	4	3	7	$4f^{14}$

* Aufbau Principle :-

⇒ "In the ground state of the atoms, the orbitals are filled in order of their increasing energies."

• In other words, e^- first occupy the lowest energy orbital available to them & enter into higher energy orbitals only after the lower energy orbitals are fulfilled.

⇒ Order is as follows :-

⇒ 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p

* Pauli Exclusion Principle :-

⇒ "No 2 e^- in an atom can have the same set of 4 quantum nos."

⇒ Also, it can be stated as :
Only 2 e^- may exist in the same orbital & these e^- must have opposite spins.

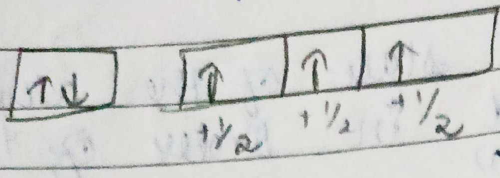
→ Maximum e^- = $2n^2$

* Hund's rule of maximum Multiplicity :-

→ Pairing of e^- in the orbitals belonging to the same subshell (p, d, or f) does not take place until each orbital belonging to that subshell has got one e^- each i.e., it is singly occupied.

eg :- $_6C$ ⇒ $1s^2$ $2s^2$ $2p^2$

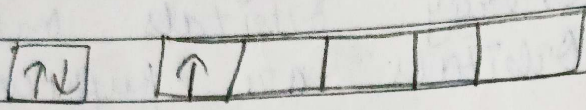
$_{15}P$ → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$



Total spin = +1½

$_{21}Sc$:- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$

[Ar] $4s^2\ 3d^1$



* Electronic Configuration of Atoms :-

→ Distribution of e^- into orbitals of an atom is called its electronic configuration. The electronic configuration can be represented in 2 ways :-

(i) s^a, p^b, d^c ... notation

(ii) Orbital diagram.

* Stability of completely filled & half filled subshells :-

→ In some cases like Cu, Cr, ~~...~~, ~~...~~ ~~...~~ where the 2 subshells differ slightly in their energies (4s & 3d), an e^- shifts from a subshell of lower energy (4s) to a subshell of higher energy (3d), resulting in all orbitals of subshell of higher energy getting either completely or half filled.

eg :- Cr → $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4$ ✗

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$ ✓

PAGE NO.:
DATE: / /

Cu :- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$ ×
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$ ✓.

→ The completely filled & half filled subshells are stable due to :-

(i) Symmetrical distribution of e^- :-
→ The completely filled or half filled subshells have symmetrical distribution of e^- in them & ∴ more stable.

(ii) Exchange energy :-
→ The stabilizing effect arises whenever 2 or more e^- with same spin are present in the degenerate orbitals of a subshell. These e^- tend to exchange their positions & the energy released due to their exchange is called exchange energy. The no. of exchanges that can take place is max. when the subshell is either half filled or completely filled.
→ As a result the exchange energy is max. & so is the stability.

* Degenerate orbitals :-
→ Orbitals of same subshell having equal energy level.