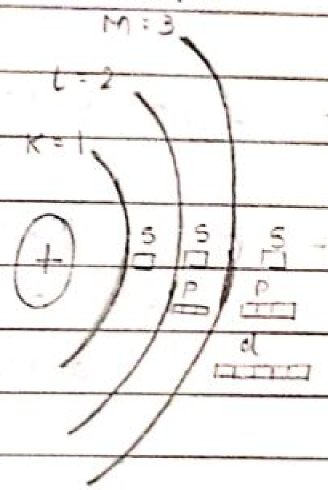
Quantum Numbers:

- shows address of electron in an atom.
- they are of four types:
  (a) Principal Q.N (n)
  (b) Azimuthal Q.N (l)
  (c) Magnetic Q.N (m)
  (d) Spin Q.N (s)



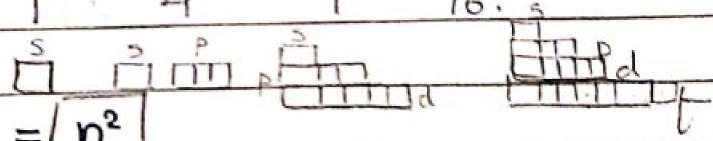
- Orbit (shell) ⟶ Sub-shell ⟶ Orbital ⟶ Spin

eg: 11^{th} NEET bench student

(a) Principal Quantum Number:

- represented by 'n'.
- represents 'shell (orbit)'.
- number of subshells = $\boxed{n}$ (1-s)(2-sp)(3-spd)(4-spdf)
- number of electrons = $\boxed{2n^2}$
- number of shell = 1 2 3 4
  K L M N

  orbitals ⟶ 1 4 9 16.



- number of orbitals = $\boxed{n^2}$

  K L M N

  electrons ⟶ 2 8 18 32

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(b) Azimuthal Quantam Number:

- represented by 'l'
- represents 'subshell'.
- azimuthal number is always 1 less than principal and it can never be equal or greater than it.

$$\boxed{l = 0 \ldots\ldots n-1} \quad \boxed{l < n \quad l \neq n}$$

- $l = 0 \quad 1 \quad 2 \quad 3$

subshell → s p d f

- number of orbitals = $\boxed{2l+1}$

subshell →	s	p	d	f
l →	0	1	2	3
orbitals →	1	3	5	7
electron →	2	6	10	14

- number of electrons = $\boxed{2(2l+1)}$

* Orbital Angular Momentum

$$\boxed{\sqrt{l(l+1)}\ \frac{h}{2\pi}}$$

here; $\frac{h}{2\pi}$ represent by $\hbar$

$\therefore \boxed{\sqrt{l(l+1)}\ \hbar}$

Q:38 find orbital ang. momentum for 's' and 'p' subshell.

$\Rightarrow$ $s = 0$ $p = 1$

for s $\Rightarrow$ $\sqrt{0(0+1)}\ \frac{h}{2\pi}$

$= \boxed{0}$

for p $\Rightarrow$ $\sqrt{1(1+1)}\ \hbar$

$= \boxed{\sqrt{2}\ \hbar}$

(c) Magnetic Quantum Number :

- represented by 'm'.
- represents 'orientation of subshell.'
- has one less, equal and one more than azimuthal.

i.e. [$m = -l \quad 0 \quad +l$.]

- s $\longrightarrow$ $m = 0$

$l = 0$

- p $\longrightarrow$ $m = -1 \quad 0 \quad +1$

$l = 0$.

- d $\longrightarrow$ $m = -2 \quad -1 \quad 0 \quad +1 \quad +2$

$l = 2$

- f $\longrightarrow$ $m = -3 \quad -2 \quad -1 \quad 0 \quad +1 \quad +2 \quad +3$

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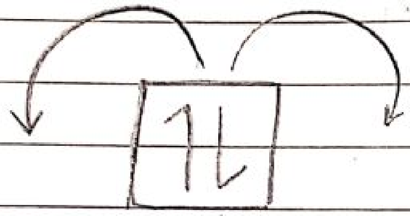
(d) Spin Quantam Number :

- represented by 's'
- represents 'spin of electron'. (in orbitals)
- $s = +\frac{1}{2}$ (clockwise) $-\frac{1}{2}$ (anti-clockwise).



- for single electron in orbitals
  spin may be $+\frac{1}{2}$ or may be $-\frac{1}{2}$.

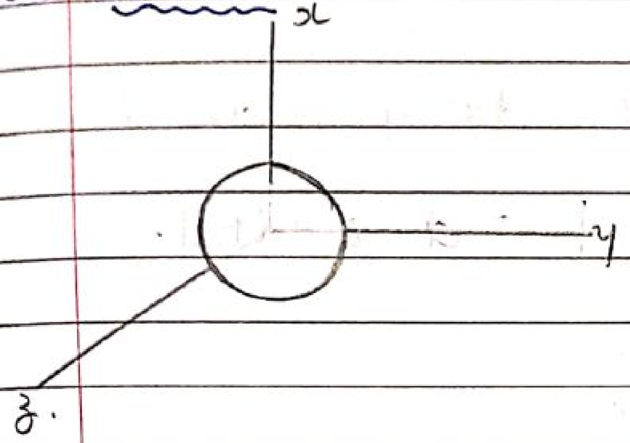
- for double electron in orbitals
  spin may be $+\frac{1}{2}$ and may be $-\frac{1}{2}$.

* Spin Magnetic Momentum

$$\boxed{\sqrt{s(s+1)}\,\hbar}$$

Shape of orbitals:

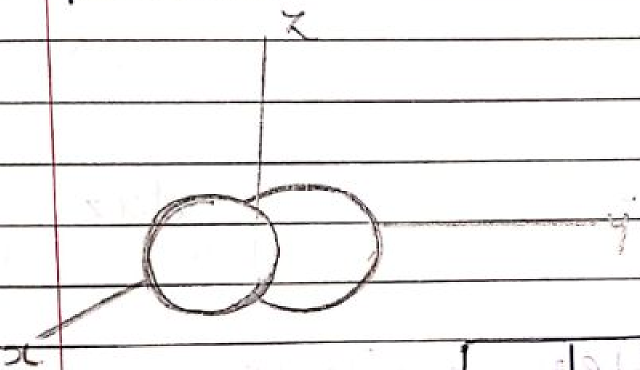
(a) s-orbital



→ Spherical shape.

$\boxed{}$ = s-orbital

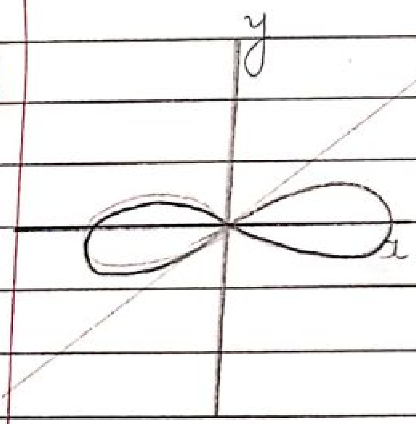
(b) p-orbital



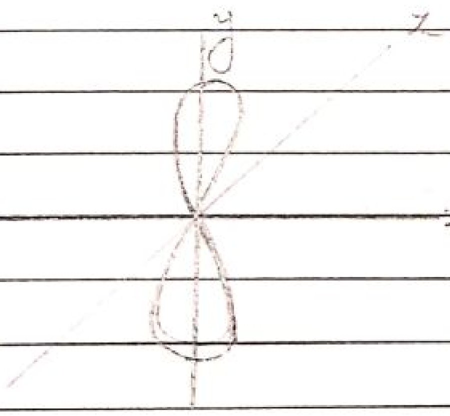
→ Dumb-bell shape
→ 03 number of degenerated p-orbitals (P_x, P_y, P_z).

Px	Py	Pz

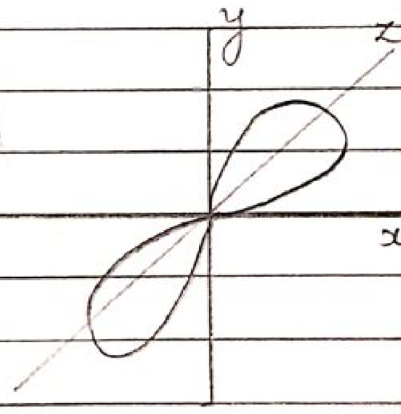
= p-orbital



px



py

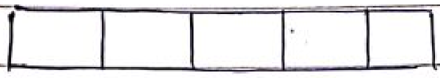


pz

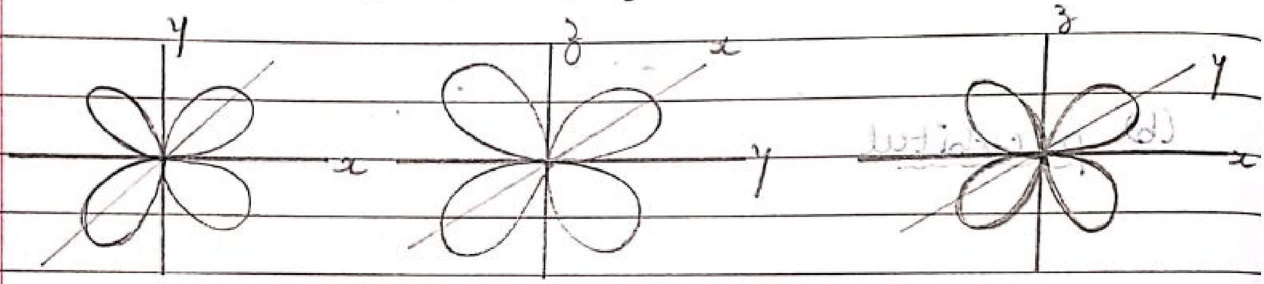
(c) d-orbital:

- Double dumb-bell shape.
- 05 number of degenerated orbitals.

$(d_{xy},\ d_{yz},\ d_{xz},\ d_{x^2-y^2},\ d_{z^2})$

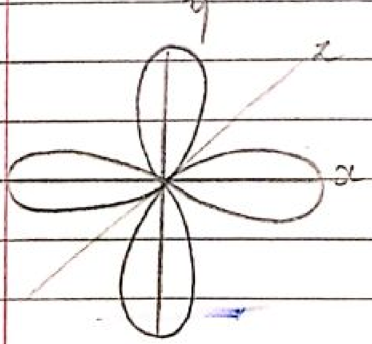
 = d-orbital.

d_{xy} d_{yz} d_{z^2} $d_{x^2-y^2}$ d_{zx}



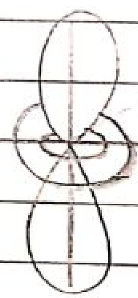
d_{xy} d_{yz} d_{xz}

Both dumb-bells available between axis.



$d_{x^2-y^2}$

(both dumb-bells on x-axis & y-axis)



d_{z^2}

(both dumb-bells on z-axis)

Nodes:

- Nodes are the points in orbitals where possibility of 'availability of electron' is negligible.

- There are 2 types of nodes:

NODES.

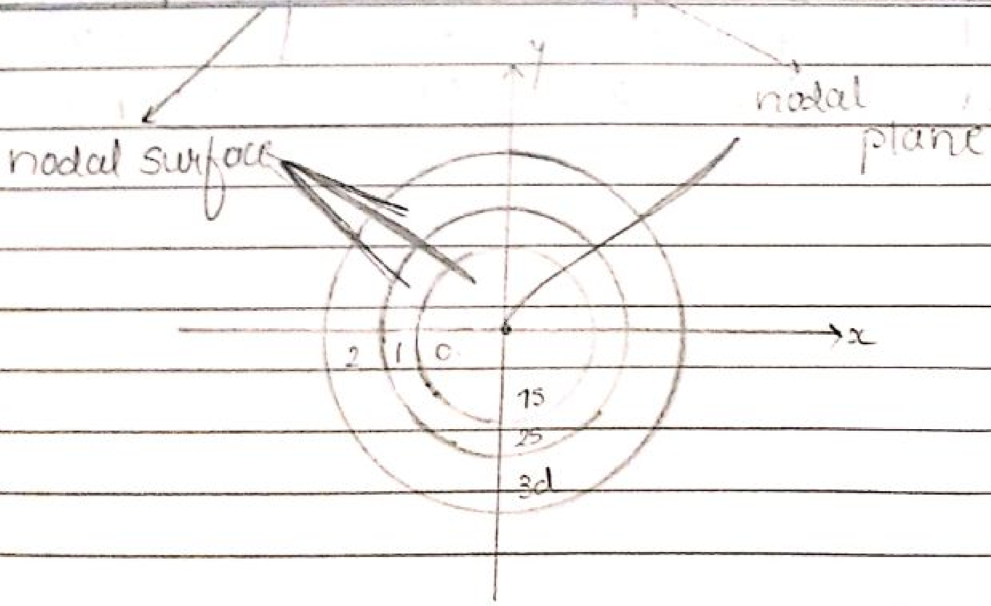
Radial Node (nodal surface) → $n-l-1$

Angular Node (nodal plane) → $= l$

(A) in S-orbital

$\frac{n-\cancel{l}-1+\cancel{l}}{n-1}$

S orbital	Radial node $(n-l-1)$	Angular node (l)	Total Node $(n-1)$
1s	$1-0-1=0$	0	0
2s	$2-0-1=1$	0	01
3s	$3-0-1=2$	0	02



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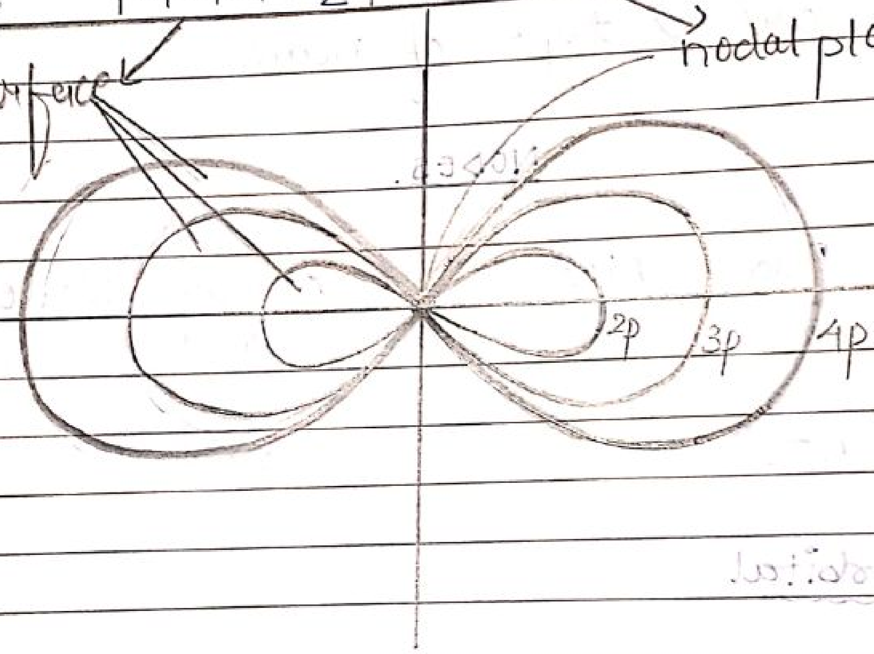
(B) in p-orbital

P-orbital	Radial node (n-l-1)	Angular node (l)	Total node (n-1)
2p	2-1-1 = 0	01	2-1 = 1
3p	3-1-1 = 1	01	3-1 = 2
4p	4-1-1 = 2	01	4-1 = 3

nodal surface

nodal plane



(C) in d-orbital

d-orbital	Radial node (n-l-1)	Angular node (l)	Total node (n-1)
3d	3-2-1 = 0	02	3-1 = 2
4d	4-2-1 = 1	02	4-1 = 3
5d	5-2-1 = 2	02	5-1 = 4

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