Chemistry

- Inorganic
- Organic
- Physical

* Chemistry is a branch of science that deals with the study of matter around us.

Inorganic

Quantum Numbers & Electronic configuration

Matter → Anything which has mass & occupies some space

- Matter
  - Physical Classification
    - Solid
    - Liquid
    - Gas
  - Chemical Classification
    - Pure
      - Element
      - Compound
    - Mixture
      - Homogeneous
      - Heterogeneous

Element → Made up of only one type of atom
eg. Helium (He), Sodium (Na), Chlorine gas (Cl_2)

Compound → Made up of two or more than two types of atoms
eg H_2O, $NaCl$, D_2O (Heavy Water)

→ Compounds follow fixed mass proportion

Homogeneous Mixture → It is a mixture of two or more types of substance having a uniform composition throughout

eg. Sugar + H_2O
$NaCl + H_2O$
Alchol + Water
Alloys - Brass [Cu + Zn)
Bronze (Cu + Sn)

Heterogenous Mixture ⇒ It is a mixture of two or more types of substances having non-uniform composition.
eg Oil + water
Sand + water
Smoke, milk, Blood

Q Classify :-

Element	Compound	Homogenous	Heterogenous
Gold	Glucose	Pure Air	Air with Dust
Silver	$C_{12}H_{22}O_{11}$	$NaCl + H_2O$	Ink + Water
		Bronze	

Atom → ① Smallest unit of matter
② It may or not exist in free state in nature
eg Na, Cl_2
③ It is electrically neutral
④ Made up of 3 fundamental particles - electrons
- protons
- neutrons

1. Electron $\boxed{^{\;0}_{-1}e}$	-1.602×10^{-19} C or -4.8×10^{-10} e.s.u.	9.11×10^{-28} g	Negative	Outside the Nucleus
2. Proton $\boxed{^{1}_{1}p}$	$+1.602 \times 10^{-19}$ C or $+4.8 \times 10^{-10}$ e.s.u.	1.6725×10^{-24} g or $\frac{1}{1837}$, mass of e	Positive	Inside the Nucleus
3. Neutron $\boxed{^{1}_{0}n}$	0	1.675×10^{-24} g	Neutral	Inside the nuc

e.s.u - electrostatic unit of charge

$\boxed{1.67 \times 10^{-24} \text{ gram} = 1 \text{ amu}}$

Representation of a atom

$$^{A}_{Z}X \quad \text{or} \quad _{Z}X^{A}$$

A - Atomic mass number

Z - Atomic number

no. of p = Z

no. of e = p = Z (In a absence of charge)

no. of e = Z - charge (for cation)

Z + charge (for anion)

atomic mass (A) = n + p

eg. $^{16}_{8}O$

O - Symbol of oxygen atom

16 - A - mass

8 - Z - atomic no

p = Z = 8

e = p = 8

n = A - Z = 16 - 8

= 8

Ex - 1

CH_4 $p = 6 + 4 = 10$
$e = 6 + 4 = 10$
$n = (12 - 6) + 4(1 - 1) = 6 + 0 = 6$

H_2O $p = 2 + 8 = 10$
$e = 2 + 8 = 10$
$n = 0 + 8 = 8$

H_2O_2 $p = 2 + 16 = 18$
$e = 2 + 16 = 18$
$n = 0 + \cancel{32} + 16 = 16$

D_2O $p = 2 + 8 = 10$
$e = 2 + 8 = 10$
$n = 2 + 8 = 10$

NH_3 $p = 7 + 3 = 10$
$e = 7 + 3 = 10$
$n = 7 + 0 = 7$

CO_3^{2-} $p = 6 + 24 = 30$
$e = 6 + 24 + 2 = 32$
$n = 6 + 24 = 30$

SO_4^{2-} $p = 16 + 32 = 48$
$e = 16 + 32 + 2 = 50$
$n = 16 + 32 = 48$

O_2 $p = 16$
$e = 16$
$n = 16$

$+1$ or 1^+ → monovalent cation
$+2$ or 2^+ → Divalent cation
$+3$ or 3^+ → Trivalent cation
-3 or 3^- → Trivalent anion
-2 or 2^- → Divalent anion
-1 or 1^- → monovalent anion

1 - mono
2 - Di
3 - Tri
4 - tetra
5 - penta
6 - Hexa
7 - Hepta
8 - Octa
9 - nona
10 - deca

Molecule → A molecule is formed when two or more than two atoms are chemically combined [either same or different) by a covalent bond.

Atomicity → Total number of atoms present in a molecule is called its atomicity.

eg. Ne → Monoatomic (atomicity -1)
O_2 → Diatomic (atomicity -2)
O_3 → Triatomic (atomicity -3)
P_4 → Tetratomic (atomicity -4)
S_8 → Octatomic (atomicity -8)

Isoterms

1. **Isotopes** - Atoms having same atomic number (z) of same element but different mass number (A).

eg. 1 ^1_1H → Protium (H) ^2_1H → Deuterium (D) ^3_1H → Tritium (T)

eg 2 $^{35}_{17}Cl$ → 75% $^{37}_{17}Cl$

Average A = 35.5

eg. 3 $^{12}_6C$ → 98.9 $^{13}_6C$ $^{14}_6C$

* Different number of neutrons.

2. **Isobars** - Atoms having same mass number but different atomic number. Atoms of different elements.

eg 1. $^{14}_{6}C$ $^{14}_{7}N$

eg 2. $^{40}_{18}Ar$ $^{40}_{20}Ca$

3. Isotones: Species having same number of neutrons.

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

$n = 3-1$ | $n = 4-2$

$= 2$ | $n = 2$

eg 2. $^{23}_{11}Na$ | $^{24}_{12}Mg$

$n = 23-11$ | $n = 24-12$

$= 12$ | $= 12$

4. Isodiaphers - Atoms with same value of $(n-p)$ or $(A-2Z)$

eg. 1 $^{19}_{9}F$ | $^{23}_{11}Na$

$A-2Z$ | $A-2Z$

$19-18$ | $23-22$

1 | 1

$^{238}_{92}U$ | $^{234}_{90}Th$

$A-2Z = 54$ | $A-2Z = 54$

5. <u>Iso-Electronic</u> – Species having same number of total electrons.

↗ atom / ion / molecule

eg 1. $^{20}_{10}Ne$ | $^{23}_{11}Na^{+}$

$e = 10$ | $e = 11 - 1 = 10$

eg 2. $_{6}C^{4-}$ $_{7}N^{3-}$ $_{8}O^{2-}$ $_{9}F^{1-}$ $_{10}Ne$ $_{11}Na^{+}$

$_{12}Mg^{2+}$ $_{13}Al^{3+}$ [All have $10e^-$]

eg 3. BO_3^{3-} CO_3^{2-} NO_3^- [All have $32e^-$]

eg 4. SO_4^{2-} PO_4^{3-} ClO_4^- (All have $50e^-$)

eg 5. H_2S HCl Ar (All have $18e^-$)

★ All iso-topes are iso-electronic.

6. <u>Isosters</u> - Species having same number of atoms as well as same no. of electrons.

eg. N_2O | CO_2

atoms = 3 | atoms 3

$e^- = 22$ | $e^- = 22$

★ All Isosters are Iso electronic but not all Iso-electronics are Isosters.

H.W.

Ex. O-1 (Q1-Q4)
O-2 (Q1)
ES-2 (Q6, 7)
JM (Q1, 6, 7, 9, 11)

O-1
Q1. B)
Q2 C)
Q3. C)
Q4 B)

O-2
Q1. a) b) d)

S-2
Q6-C)
Q7-b)

JEE-Mains

Q1. 1)
Q6 1)
Q7 4)
Q9 4)
Q11. 2)

Quantum Numbers

Atom – ~~Principal Quantum Number (n)~~

↓

Principal Quantum Number (n) – Shell/Orbit – ~~Azimuthal Quantum Number (l)~~

↓

Azimuthal Quantum Number (l) – Subshell – ~~Magnetic Quantum Number (m or m_l)~~

↓

Magnetic Quantum Number (m or m_l) – Orbital – ~~Spin Quantum Number (s, m_s)~~

↓

Spin Quantum Number (s or m_s) – Present/Absent

☆ First 3 Quantum Numbers gives information about the orbitals

☆ First 4 Quantum Numbers gives information about the electrons

B

Orbitals - It is defined as 3D space occupied by the electrons around the nucleus where the probability of finding the electron is maximum.

1. **Principle Quantum Number(n)** - Principal Quantum Number represent Shell/orbit/ energy level in which electron is present and its distance from the nucleus. (size of shell/orbit)

☆ It is also called as Primary Quantum number.

☆ Permissible values of n - 0 to ∞

1	2	3	4	5	6	... ∞
↓	↓	↓	↓	↓	↓	
K	L	M	N	O	P	

☆ Maximum number of possible shells in an atom is infinite.

n ↓ Attraction by Nucleus ↑ Stability ↑ Energy ↓

☆ n decides the energy of a electron

☆ as n increases, energy also increases

☆ Principal quantum number also gives the idea about Total number of orbitals & total no. of electrons in a Shell.

☆ In a ~~electron~~ orbital, There can be maximum $2e^-$

☆ Total no. of orbitals in a shell can be given by n^2.

☆ Total no. of electrons in a shell are $2n^2$

Shell no.	Max no. of orbitals	Max no. of electrons
$n = 1$ (K)	$= (1)^2$ $= 1$	$= 2(1)^2$ $= 2$
$n = 2$ (L)	$= (2)^2$ $= 4$	$= 2(2)^2$ $= 8$
$n = 3$ (M)	$= (3)^2$ $= 9$	$= 2(3)^2$ $= 18$
$n = 4$ (N)	$= (4)^2$ $= 16$	$= 2(4)^2$ $= 32$
$n = 5$ (O)	$= (5)^2$ $= 25$	$= 2(5)^2$ $= 50$

2. Azimuthal Quantum Number (l) / secondary Q.No. / subsidiary Q.No. –

1. The value of $l \Rightarrow 0$ to $(n-1)$ in Integral Step
2. This Quantum Number gives the information about subshell
   a) ~~name~~ name of subshell
   b) energy of subshell ($l\uparrow$, $e\uparrow$)
   c) shape of subshell

Value of l	Name of Subshell	Shape of subshell
$l = 0$ (s)	Sharp	Spherical 
$l = 1$ (p)	Principal	Dumbbell 
$l = 2$ (d)	Diffused	Double Dumbbell 
$l = 3$ (f)	fundamental	complex (not in JEE)
$l = 4$ (g)	–	–

☆ The number of subshells in n^{th} shell = n.

n	l (0 to $(n-1)$)	Subshells
$n=1$	$l=0$ (s)	1s (1 subshell)
$n=2$	$l=0$ (s) 1 (p)	2s 2p (2 subshells)
$n=3$	$l=0$ (s) 1 (p) 2 (d)	3s 3p 3d (3 subshells)

☆ For existence of a subshell, $(n>l)$

2s - exist

3p - exist

2d - not exist

4g - not exist

☆ For a particular energy level, the energy of subshell is in following order -

$$[ns < np < nd < nf]$$

eg. $4s < 4p < 4d < 4f$

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3. <u>Magnetic Quantum Number</u> (m or m_l)

1. It gives the information about possible orientation (arrangement) of an ~~ato~~ orbital in space.

2. Values of m depends upon the value of l and ranges from $-l$ to $+l$ (including 0) in integral steps.

O-1 (15, 5, 7)
O-2 (2, 5)

JEE-M (5, 8, 13, 14, 15, 10)
JEE-A (1, 2)

Answers

O-1

Q5. B)

Q7. C)

Q15. D)

O-2

Q2. C)

Q5. B) D)

J-M

Q5. 1)

Q8. 2)

Q10. 2)

Q13 A)

Q14 3)

Q15 C)

J-A

Q1. 98

Q2. 6

n	l	m	no. of orbitals	no. of e^-
$n=1$	$l=0(s)$	$m=0$	$1=(1)^2$	$1\times2=2$
$n=2$	$l=1(p)$	$m=-1,0,+1$	$(2)^2=4$	$4\times2=8$
$n=3$	$l=2(d)$	$m=-2,-1,0,1,2$	$(3)^2=9$	$9\times2=18$
$n=4$	$l=3(f)$	$m=-3,-2,-1,0,1,2,3$	$(4)^2=16$	$16\times2=32$

☆ For any given subshell, maximum number of orbitals are $(2l+1)$ & max no. of e^- are $2(2l+1)$

eg. $l=0(s)$

~~m=0 1 orbi~~ $2\times0+1=1$ orbital

$l=1(p)$

$2\times1+1=3$ orbitals

l	m	representation of orbital
$l=0(s)$	$m=0$	1 — s
$l=1(p)$	$m=-1,0,1$	3 — p_x p_y p_z
$l=2(d)$	$m=-2,-1,0,1,2$	5 — d_{xy} d_{yz} d_{xz} $d_{x^2y^2}$ d_{z^2}

Q. In an orbital has $m = -1$ Then it cannot be present in which subshell

✓a) s   b) p   c) d   d) f

4. ~~Magnetic~~ Spin Quantum Number (s, m_s)

1. It is defined to explain the rotation or movement of electron around its own axis.

	Clockwise	Anticlockwise
spin up	↑ or ↿	↓ or ⇂
spin down	↓ or ⇂	↑ or ↿
	$+\frac{1}{2}$	$-\frac{1}{2}$
	$-\frac{1}{2}$	$+\frac{1}{2}$

↿↿ ⇒ Parallel Spin

↿⇂ or ⇂↿ ⇒ Opposite spin

Summary of Quantum Numbers

1. $n \rightarrow$ Shell $\rightarrow 1 - \infty$
2. $l \rightarrow$ Subshell $\rightarrow 0$ to $n-1$
3. $m_l \rightarrow$ Orbital $\rightarrow -l$ to $+l$ (including zero)
4. $m_s \rightarrow$ spin of $e^- \rightarrow -\frac{1}{2}, +\frac{1}{2}$

Q1. What is the highest value of l for shell no. 3? 2

$n = 3$
$l = 0$ to $(n-1)$
$l = 0, 1, 2$

Q2. What is the highest value of m for P subshell? +1

$l = 1$
$m = -1, 0, +1$

Q3. What is the minimum value of m for d-subshell? -2

$l = 2$
$m = -2, -1, 0, 1, 2$

Q4. Max no. of e^- in S-subshell

$l = 0$
$e^- = (2l+1)2$
$= (0+1)2$
$\boxed{= 2}$

Q5. Max no. of e^- in P-orbitals?

$l = 1$
$e^- = (2l+1)2$
$\boxed{= 6}$

Q6. Select correct set of Quantum Nos.

	n	l	m	s
(A)	3	3	−2	−1/2
(B)	2	1	−2	+1/2
✓(C)	3	2	1	+1/2
(D)	4	2	−2	+1

Q7. How many max e^- can be described by Q. Nos. $n = 5$, $l = 2$ in a atom.

$l = 2$
$m = -2, -1, 0, 1, 2$
no of e^- = $2(2\times2+1)$
$\boxed{= 10}$

Scanned with CamScanner

Rule for Filling electrons

1. **Pauli's exclusion Rule** - According to this rule, no two electrons in an atom can have all four Quantum Numbers same, at least 1 Quantum Number should be different.

$\boxed{\uparrow\ \uparrow}$ 2s

$n=2$	$n=2$
$l=0$	$l=0$
$m=0$	$m=0$
$s=+\frac{1}{2}$	$s=+\frac{1}{2}$

not Possible

$\boxed{\uparrow\downarrow}$ 2s

$n=2$	$n=2$
$l=0$	$l=0$
$m=0$	$m=0$
$s=\frac{1}{2}$	$s=-\frac{1}{2}$

Possible

$\boxed{\uparrow\downarrow\uparrow}$ 2s

$n=2$	$n=2$	$n=2$
$l=0$	$l=0$	$l=0$
$m=0$	$m=0$	$m=0$
$s=\frac{1}{2}$	$s=-\frac{1}{2}$	$s=\frac{1}{2}$

Same
↓
not Possible

⟹ This is the reason, one orbital has only 2 e^-

2. **Afbau's Rule** :- (only for multi electronic species)

→ Not for single electronic species like (H, He^+, Li^{2+}, Be^{3+})

→ These rules will tell us about the filling of electron in different subshells.

→ According to this rule, electrons are filled in lower energy ~~level~~ subshell first and once they are filled, then filling of next higher energy subshell occurs.

$(n+l)$ rule → Higher the value of $(n+l)$ for a subshell, higher will be its energy

Ex-1

4s	3d
$n=4$	$n=3$
$l=0$	$l=2$
$n+l=4$	$n+l=5$

$3d > 4s$ (energy)

Ex-2

5P	4f
$n=5$	$n=4$
$l=1$	$l=3$
$n+l=6$	$n+l=7$

$4f > 5P$ (energy)

Ex-3

4P	5s
$n=4$	$n=5$
$l=1$	$l=0$
$l+n=5$	$l+n=5$

$5s > 4P$ (energy)

* If two subshells having same value of $(n+l)$, then subshell with higher value of n has higher energy.

Q energy order?

1. 3s	2P	3d	4f
$n=3$	$n=2$	$n=3$	$n=4$
$l=0$	$l=1$	$l=2$	$l=4$
$n+l=3$	$n+l=3$	$n+l=5$	$n+l=7$

$\boxed{4f > 3d > 3s > 2P}$

* for calculating no. of e^- in n^{th} period

$ns < (n-3)g < (n-2)f < (n-1)d < (nP)$

ii) 4f 5d 6s 4p

7 7 6 5

$\boxed{5d > 4f > 6s > 4p}$

iii) 1s < 2s < 3s < 3p < 3d < 4s < 4p < 4d < 4f

< 5s < 5p < 5d < 5f

5f > 5d > 4f > 5p > 4d > 5s > 4p > 4s > 3p > 3s > 2p > 2s > 1s

n/l	0	1	2	3	4
1	1s				
2	2s	2p			
3	3s	3p	3d		
4	4s	4p	4d	4f	
5	5s	5p	5d	5f	5g
6	6s	6p	6d	6f	
7	7s	7p			

Q

Q1. $4d_{xy}$ $4f$ $3d_{yz}$ $3d_{z^2}$ $5p_y$

6   7   5   5   ~~4~~ 6

$4f > 5p_y > 4d_{xy} > \underbrace{3d_{yz} = 3d_{z^2}}_{\text{degenerate orbitals}}$

<u>degenerate orbitals</u> – All the orbitals of a particular ~~set~~ subshell have same energy.

→ The orbitals having same value of n & l but different value of m have same energy in the absence of a electric & magnetic field.

→ These orbitals having same energy of a particular subshell are known as ~~same~~ degenerate orbitals.

<u>Degeneracy</u> :- Total no. of orbitals with same energy.

Q. find degeneracy

i) p subshell - 3

ii) d subshell - 5

iii) f subshell - 7

iv) s subshell - Not-Defined (D = 0)

✲ For single electron species, Or hydrogen like species, The energy depends only on n.

H. W. 09-4-24

O-1 (Q5, 7, 15)
O-2 (Q2, Q5)
~~Q~~ JM (Q5, 8, 10, 13, 14, 15)
JA (Q1, 2)

Answers:-

O-1
~~~~ (Q5 - B)
(Q 7 C)
(Q15. D)

O-2
(Q-2 C)
(Q5. B, D

Q JM
(Q5 1)
(Q 8 2)
(Q10 2)
(Q13 A)
(Q 14 3)
(Q15 C)

JEE-Advance

Q1. 9
Q 2. 6
~~~~

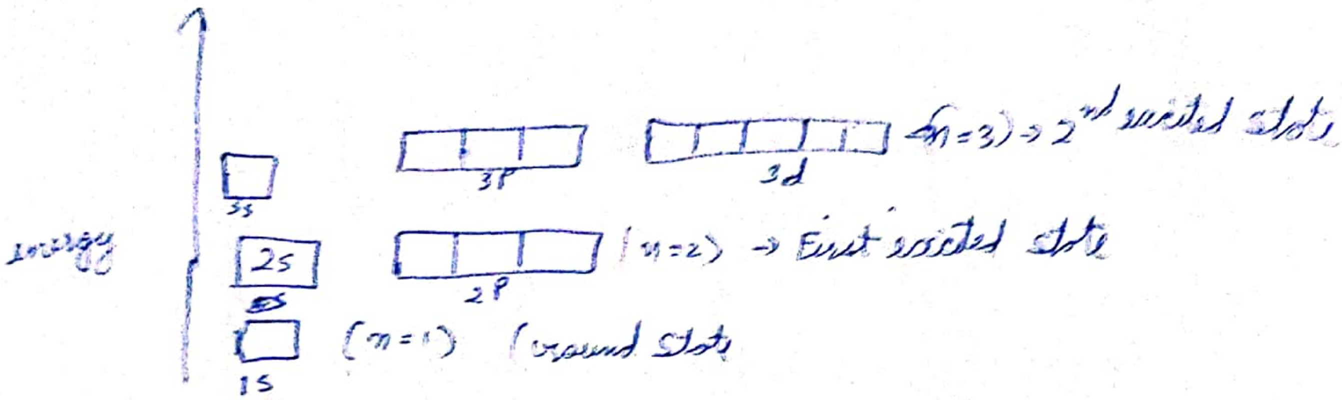
Single Electronic Species (H-like atom species)

- no. of $\bar{e} = 1$

Ex ($H, He^{+}, Li^{2+}, Be^{3+} \ldots$)

→ Energy will depend only on value of n

Energy Order- $1s < (2s = 2p) < (3s = 3p = 3d) < \ldots$



energy level diagram for single $\bar{e}$ species

eg. degeneracy of $n=2$ for H-atom $= 4$

degeneracy of 2nd excited state of H-atom $= 9$

Q1. what is the degeneracy of 3rd excited state of Li^{2+} ion. - $n = 4$

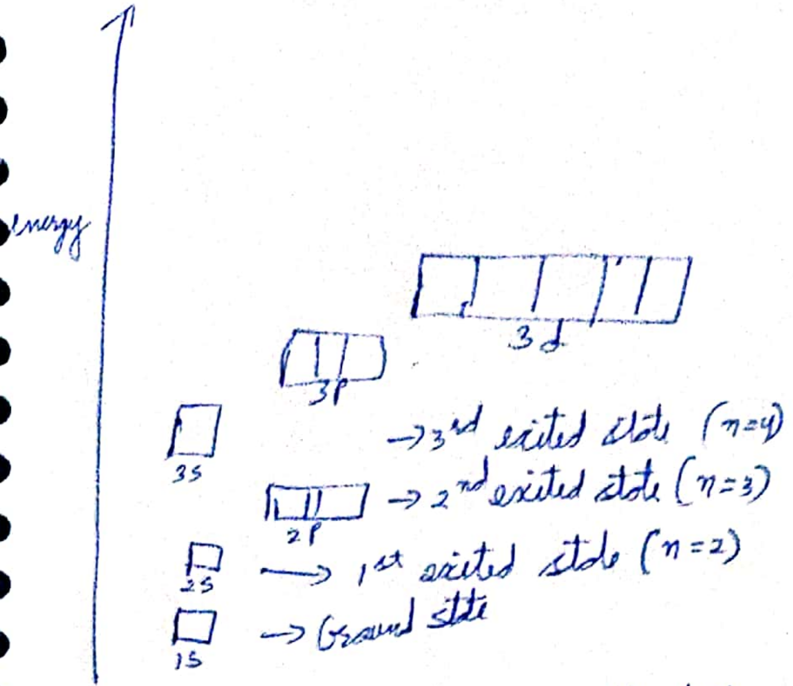
degeneracy $= n^2$
$= (4)^2$
$= 16$

Multi-Electronic species

no. of $\bar{e} > 1$

eg. He, H, Li, Be ....

energy of $\bar{e}$ depends on value of n & l



energy diagram for multi electronic species

Q what is the degeneracy of 2nd excited state of helium -3

Q what is the degeneracy of 3rd exited state of Lithum- I

Hunds Rule of maximum multiplicity

→ Applies to filling of e⁻ into degenerate orbitals

→ According to this rule electrons are filled into diff orbitals of a subshell so as to give maximum no. of unpaired e⁻ with parallel spin

→ Paring of e⁻ do not take place in orbitals until each orbital of the same subshell is filled with 1e⁻ of same spin.

OR

→ According to this rule, e⁻s are filled into diff orbitals of a subshell so as to give maximum spin multiplicity.

[↑ | ↑ | ↑ | ↑ | ↑] ✓

[↑ | ↑ | ↓ | ↑ | ↑] ✓

[↑ | ↓ | ↑ | ↑ | ↑] ✗

[↑↓ | ↑ | ↑ | ↑ | ↑] ✓

[↑↓ | ↑ | ↑ | ↑ |] ✗

[↑↓ | ↑ | ↑ | ↑ | ↑↓] ✓

Total Spin (S_T) :-

[↑ | ↑ | ↑]

$$|S_T| = \left|+\tfrac{1}{2} + \tfrac{1}{2} + \tfrac{1}{2}\right| = \frac{3}{2}$$

[↑↓ | ↑ | ↑]

$$|S_T| = \left|+\tfrac{1}{2} - \tfrac{1}{2} + \tfrac{1}{2} + \tfrac{1}{2}\right| = 1$$

Spin Multiplicity (S_m)

$$\boxed{S_m = 2|S_T| + 1}$$

[↑ | ↑ | ↑]

$|S_T| = \frac{3}{2}$

$S_m = 2 \times \frac{3}{2} + 1$

$S_m = 4$

[↑↓ | ↑ | ↑]

$|S_T| = 1$

$S_m = 2 \times 1 + 1$

$S_m = 3$

Q1. The orbital diagram in which violated Afbau's principal is

i) [↑↓] [↑↓ | ↑ |]

ii) [↑] [↑↓ | ↑ | ↑] ✓

iii) [↑↓] [↑ | ↑ | ↑]

iv) [↑↓] [↑↓ | ↑ | ↓]

Q2. The orbital diagram in which Hund's Rule is violated is

i) [↑ | ↑ | ↑ | | ↑]

ii) [↑ | ↑ | | | ↓] ✓

iii) [↑↓ | ↑↑ |] ✓

iv) [↑ | ↑ | ↑ | ↑↓ | ↑]

Q3. Match

Configuration	Violated Rule
~~i) [↑↓ | ↑↓]~~	
A i) [↑↓] [↑↓] [↑ | ↑ |]	P) Pauli's Rule
B ii) [↑↓] [↑↑] [↑ | ↓ |]	Q) No Rule violated
C iii) [↑] [↑↓] [↑ | ↓ | ↑]	R) Afbau's Rule
D iv) [↑↑] [↑] [↑ | ↓ | ↑]	S) Hund's Rule

A) Q
B) P S
C) R S
D) P R S

Electronic Configuration of a atom.

→ The destribution of electrons in various shells, subshells and orbitals in an atom is called ~~ort~~ its electronic configuration.

eg. $1s^2$ → no. of electrons

$n = 1$

$l = s$
$= 0$

$s \rightarrow 2e^-$
$p \rightarrow 6e^-$
$d \rightarrow 10e^-$
$f \rightarrow 14e^-$

eg.

$_1H \rightarrow 1s^1$

$_2He \rightarrow 1s^2$

$_3Li \rightarrow 1s^2 2s^1$

$_4Be \rightarrow 1s^2 2s^2$

$_5B \rightarrow 1s^2 2s^2 2p^1$

$_6C \rightarrow 1s^2 2s^2 2p^2$

$_7N \rightarrow 1s^2 2s^2 2p^3$

$_8O \rightarrow 1s^2 2s^2 2p^4$

$_9F \rightarrow 1s^2 2s^2 2p^5$

$_{10}Ne \rightarrow 1s^2 2s^2 2p^6$

Orbital notation method

eg

$_6C \rightarrow \underset{1s}{\boxed{\uparrow\downarrow}}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow\ \uparrow\ \ }$ } orbital diagram method

$_6C \rightarrow [\overset{1s^2}{He}]\ 2s^2 2p^2$ } condensed form

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$_{11}Na \rightarrow [Ne]\ 3s^1$

$_{12}Mg \rightarrow [Ne]\ 3s^2$

$_{13}Al \rightarrow [Ne]\ 3s^2\ 3p^1$

$_{14}Si \rightarrow [Ne]\ 3s^2\ 3p^2$

$_{15}P \rightarrow [Ne]\ 3s^2\ 3p^3$

$_{16}S \rightarrow [Ne]\ 3s^2\ 3p^4$

$_{17}Cl \rightarrow [Ne]\ 3s^2\ 3p^5$

$_{18}Ar \rightarrow [Ne]\ 3s^2\ 3p^6$

$_{19}K \rightarrow [Ar]\ 4s^1$

$_{20}Ca \rightarrow [Ar]\ 4s^2$

$_{21}Sc \rightarrow [Ar]\ 4s^2\ 3d^1$

$_{22}Ti \rightarrow [Ar]\ 4s^2\ 3d^2$

$_{23}V \rightarrow [Ar]\ 4s^2\ 3d^3$

$_{24}Cr \rightarrow \cancel{[Ar]\ 4s^2\ 3d^4}$ $[Ar]\ 4s^1\ 3d^5$

$_{25}Mn \rightarrow [Ar]\ 4s^2\ 3d^5$

$_{26}Fe \rightarrow [Ar]\ 4s^2\ 3d^6$

$_{27}Co \rightarrow [Ar]\ 4s^2\ 3d^7$

$_{28}Ni \rightarrow [Ar]\ 4s^2\ 3d^8$

$_{29}Cu \rightarrow \cancel{[Ar]\ 4s^2\ 3d^9}$ $[Ar]\ 4s^1\ 3d^{10}$

$_{30}Zn \rightarrow [Ar]\ 4s^2\ 3d^{10}$

Inert Gas	Atomic no.
He	2
Ne	10
Ar	15
Kr	36
Xe	54
Rn	86
Uuo	118

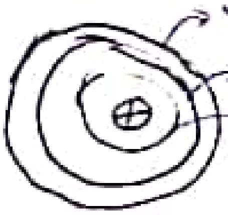
Exception

$_{24}Cr$:- $[Ar]\ 4s^1\ 3d^5$

$_{29}Cu$:- $[Ar]\ 4s^1\ 3d^{10}$

Valence Shell(n)

* It is the outermost shell which is occupied by a electron in the electronic configuration
* The highest value of 'n' is valence shell.



$Na - 1s^2\ 2s^2\ 2p^6\ 3s^1$

Q1. write down the all four Quantum No. for last e^- of Fe (iron)?

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

$v e^- \rightarrow n = 4$, $l = 0$, $m = 0$, $s = +\frac{1}{2} / -\frac{1}{2}$

last $e^- \rightarrow n = 3$
$l = 2$
$m = -2, -1, 0, 1, 2$
$s = -\frac{1}{2} / +\frac{1}{2}$

Q2. Write write down the all four Q. No. for 7^{th} e^- of Cl (atom)

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

$v e^- \rightarrow n = 3$, $l = 1$, $m = +1/-1/0$, $s = +\frac{1}{2} / -\frac{1}{2}$

$n = 2$
$l = 1$
$m = -1, 0, +1$
$s = \pm\frac{1}{2}$

Q3. find ~~minimum~~ maximum no. of e^- in sulfur for which $m = 0$?

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

for $m = 0$, $l > 0$ ~~any~~ & 2 in every subshell.

$\boxed{10}$

Q4. find total no. of e^- in manganese for which

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

a) value of $n = 3$
$= 6 + 5 + 2$
$\boxed{= 13}$ ✓

b) value of $l = 1$
$= 6 + 6$
$\boxed{= 12}$ ✓

c) value of $|m| = 1$
~~$= 2 + 2 + 2$~~
$= 4 + 4 + 4 = 12$
$\boxed{= 12}$ ✓

d)

d) value of $(n + l) = 4$
$= 6 + 2$
$\boxed{= 8}$ ✓

e) value of $(l + m) = 0$
$= 2 + 2 + 2 + 2 + 2 + 2 + 1$
$\boxed{= 11} = \boxed{13}$ ✓

f) value of $(l + m) = 2$
$+ 2 + 2 + 1$
$\boxed{}\ \boxed{= 5}$ ✓

g) $n \times l \times m = 0$
$= 8 + 5 = 13$
$\boxed{= 13}$ ✓

Q find out minimum no. of e^- in Magnese for $s = -\frac{1}{2}$

$10 + 2 \rightarrow 12 - \boxed{15}$

min no. of e^- = 10

max no of e^- = 15

Electronic configoration of ions

* Step-1 – first write down the electronic configuration of neutral atom.

Step-2 removal of electron takes place from outermost shell i.e n^{th} shell.

within a shell, removal of electron takes place from those subshells which have more value of l (Azimuthal Quantum number)

eg $_{11}Na^+$

$Na \rightarrow 1s^2 \; 2s^2 \; 2p^6 \; 3s^1 \rightarrow$ outermost

$Na^+ \rightarrow 1s^2 \; 2s^2 \; 2p^6$

$Mn \rightarrow 1s^2 \; 2s^2 \; 2p^6 \; 3s^2 \; 3p^6 \; 4s^2 \; 3d^5$

$Mn^{2+} \rightarrow 1s^2 \; 2s^2 \; 2p^6 \; 3s^2 \; 3p^6 \; 3d^5$

$Mn^{4+} \rightarrow 1s^2 \; 2s^2 \; 2p^6 \; 3s^2 \; 3p^6 \; 3d^3$

$Mn^{+7} \rightarrow 1s^2 \; 2s^2 \; 2p^6 \; 3s^2 \; 3p^6$

$N \rightarrow 1s^2 \; 2s^2 \; 2p^3$

$N^{3-} \rightarrow 1s^2 \; 2s^2 \; 2p^6$

Extra stability of half filled & fully filled subshell.

1. The electronic configuration of most of the atom follows Aufbau's rule.

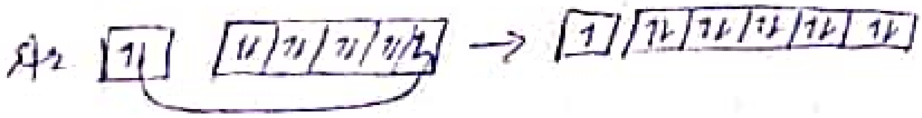
Exception of Aufbau's Rule:-

1. $_{24}Cr \rightarrow$ [Ar] $4s^2\ 3d^4 \rightarrow$ expected

   [Ar] $4s^1\ 3d^5 \rightarrow$ reality

2. $_{29}Cu \rightarrow$ [Ar] $4s^2\ 3d^9 \rightarrow$ expected

   [Ar] $4s^1\ 3d^{10} \rightarrow$ reality



2. Here the two subshell (4s & 3d) differ slightly in their energy. (4s<3d), An electron shifts from a subshell of lower energy (4s) to a subshell of higher energy (3d). Provided such a shift result in all orbitals of a subshell of higher energy getting either completely filled or half filled

Half filled subshell	Fully filled subshell
s - [↑]	[↑↓]
p - [↑][↑][↑]	[↑↓][↑↓][↑↓]
d - [↑][↑][↑][↑][↑]	[↑↓][↑↓][↑↓][↑↓][↑↓]
f - [↑][↑][↑][↑][↑][↑][↑]	[↑↓][↑↓][↑↓][↑↓][↑↓][↑↓][↑↓]

Note:- 1. Half filled & fully filled e^- configuration is more stable than partially filled e^- configuration.

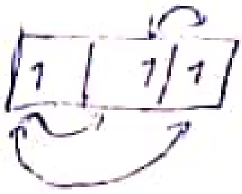
2. fully filled e^- configuration is more stable than half filled e^- configuration.

3. It has been found that there is an extra stability associated with this electronic configuration. The stabilisation is due to following 2 factors

Explanation 1- Symmetry leads to stability (Nature's law)

Exp 2 - Exchange energy - The electron present in degenerate orbitals having parallel spin can exchange their positions and energy released due to this exchange is known as exchange energy.

no. of exchange = 1



no. of exchange = 3

no. of exchange = 4

$$\boxed{\text{no. of exchange} = \frac{n(n-1)}{2} + \frac{n'(n'-1)}{2}}$$

n = no. of e^- in up spin
n' = no. of e^- in down spin

No. of exchange ↑ → energy ↓ → stability ↑

Excited energy

Excited energy – The minimum amount of energy required to excite an electron from ground state (lower energy level) of an atom to any excited state (higher energy) is called excited energy.

eg.

C - [↑↓] [↑↓] [↑|↑|] → Ground State ($1s^2\ 2s^2\ 2p^2$)

C^* : [↑↓] [↑] [↑|↑|↑] → excited state ($1s^2\ 2s^1\ 2p^3$)

Exceptional Electronic Configurations

S.no.	Element	Z	Config
1.	Cr	24	[Ar] $4s^1 3d^5$
2.	Cu	29	[Ar] $4s^1 3d^{10}$
3.	Nb	41	[Kr] $5s^1 4d^4$
4	Mo	42	[Kr] $5s^1 4d^5$
5.	Ru	44	[Kr] $5s^1 4d^7$
6.	Rh	45	[Kr] $5s^1 4d^8$
7	Pd	46	[Kr] $4d^{10}$
8	Ag	47	[Kr] $5s^1 4d^{10}$
9.	La	57	[Xe] $6s^2 5d^1$
10	Pt	78	[Xe] $6s^1 4f^{14} 5d^9$
11.	Au	79	[Xe] $6s^1 4f^{14} 5d^{10}$
12.	Ac	89	[Rn] $7s^2 6d^1$
13.	Th	90	[Rn] $7s^2 6d^2$

HW 13-04-24 34-Questions

O-1 (Q 6, 8, 9, 10, 11, 13, 16, 17)

O-2 (Q 3, 4, 6, 7, 8, 9, 10, 11, 12, 13, 14, 16)

S-1 (Q1, 2, 3, 5, 6)

S-2 (Q1, 2, 3, 9)

JM (Q2, 4, 12, 16)

JA (Q3)

JA Q3 - 3

JM Q2 - 1)
Q4 - 3)
Q 12 - 3)
Q 16 - A)

S-2
Q 1 - C
Q 2 - B
Q 3 - C
Q 9 - C

S-1
Q 1-4
Q 2-9
Q 3- 8
Q 5, 6 (Radut)
Q6 - 2. (no. of orbitals)

O-2
Q 3 - A, B
Q 4 - C, D
Q 6 - B, D
Q 7 - A, B, D
Q 8 - B, D
Q 9 - A, B, D
Q 10 - C, D
Q 11 - A, B, D
Q 12 - C, D
Q 13 - C, D, B
Q 14 - B, C
Q 16 - A, B

O-1
Q 6 - A
Q 8 - B
Q 9 - C
Q 10 - C
Q 11 - B
Q 12 - B
Q 13 - B C
Q 16 - C D
Q 17 - A

Q. Find maximum no. of e^- in 5^{th} period.

$n = 5$

$5s$, $(5-3)g$, $(5-2)f$, $(5-1)d$, $5p$

$5s$, ~~$2g$~~, ~~$3f$~~, $4d$, $5p$

$5s \rightarrow 2$, $4d \rightarrow 10$, $5p \rightarrow 6$

$10+6+2 \rightarrow 18$ elements

Magnetic Properties

Paramagnetic Diamagnetic

Spin Magnetic moment (μ):-

μ(net)

→ The magnetic moment due to spinning of e^- is called spin magnetic moment.

→ $\boxed{\mu = \sqrt{(n)(n+2)}}$

n: no. of unpaired electrons

→ unit - B.M. (Bohr Magneton)

eg. $_{12}Mg$ - $1s^2\ 2s^2\ 2p^6\ 3s^2$

$\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}\ \boxed{\uparrow\downarrow}$

no of unpaired electron = n = 0

Eg- $_7N \rightarrow 1s^2\ 2s^2\ 2p^3$

$\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow|\uparrow|\uparrow}$

$n = 3$

$\mu = \sqrt{3(3+2)}$

$\mu = \sqrt{15}$

$\mu = 3.87$

no. of unpaired e^-	$\sqrt{n(n+2)}$	μ
0	0	0
1	$\sqrt{3}$	1.73
2	$\sqrt{8} = 2\sqrt{2}$	2.84
3	$\sqrt{15}$	3.87
4	$\sqrt{24}$	4.91

value of μ before decimal (1.73) = no. of unpaired e^-

Q If Co^{x+} has magnetic moment 3.87 B.M. find x.

$4s^2\ 3d^7$

$\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow|\uparrow|\uparrow}$

$+1 \rightarrow \boxed{\uparrow}\ \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow|\uparrow|\uparrow} \rightarrow n = 4 \qquad x = 1$

$+2 \rightarrow \boxed{\uparrow|\uparrow\downarrow|\uparrow|\uparrow|\uparrow} \rightarrow n = 3 \qquad x = 2$

$+6 \rightarrow \boxed{\uparrow|\uparrow|\uparrow|\ |\ } \rightarrow n = 3 \qquad x = 6$

$\boxed{x = 2, 6}$

Magnetic

Paramagnetic

1. weakly attracted by external magnetic field.
2. $\mu \neq 0$
3. At least 1 unpaired e^- will be present

eg. N, Cl, C

Diamagnetic

1. weakly repelled by external magnetic field.
2. $\mu = 0$
3. no unpaired e^- present

eg. Ne, Mg, ~~Be~~, Na^+, F^-, N^{3-}

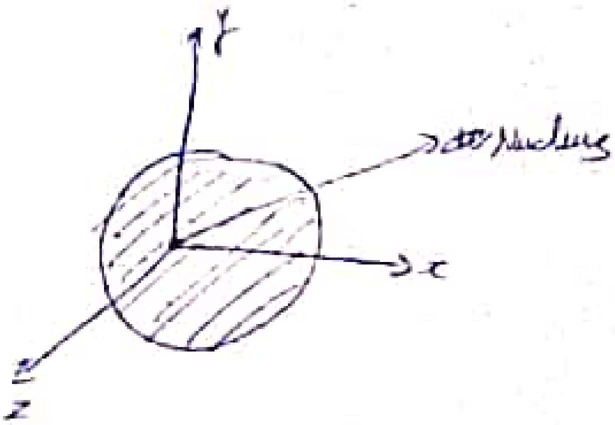
Shape of orbitals & their orientation:-

1. Nodal Plane - It is defined as the plane passing through the nucleus where probability of finding an electron is zero

\# no. of nodal planes = value of l

s - orbital	$l = 0$	N.P. = 0
p - orbital	$l = 1$	N.P. = 1
D - orbital	$l = 2$	N.P. = 2

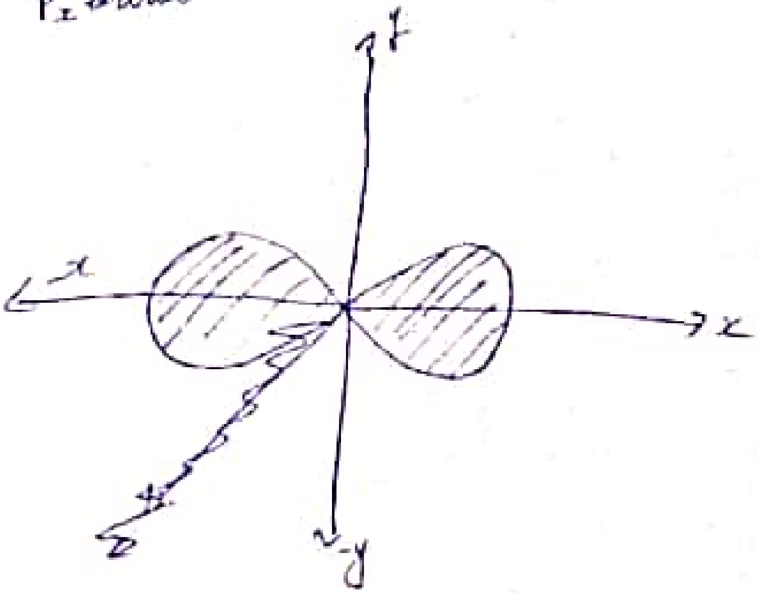
1. s-orbital (spherical)



N.P = 0

2. P-orbital (Dumbell)

P_x orbital

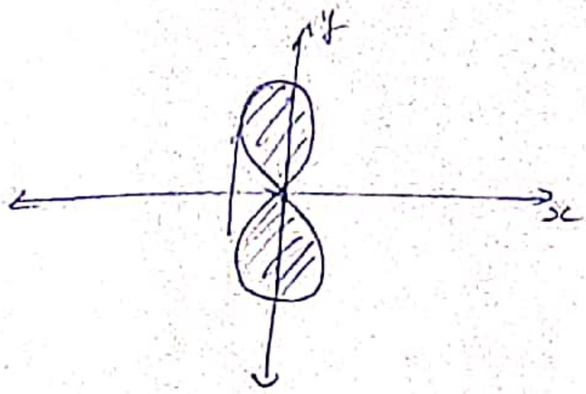


N.P = 1

Nodal plane - YZ axis or ZY

Scanned with CamScanner

P_y orbital -



Nodal Plane - XZ or ZX

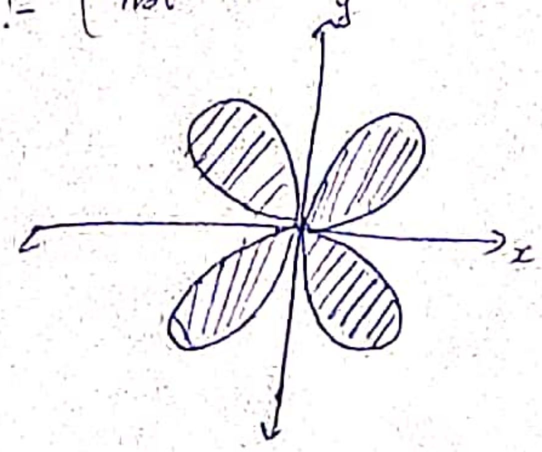
These orbitals are axial orbitals because e^- density lies of axises.

P_z orbital:-

Nodal plane - XY or YX

3. d-orbital (double-dumbell) (d_{xy}, d_{zx}, d_{yz}, $d_{x^2-y^2}$, d_{z^2})

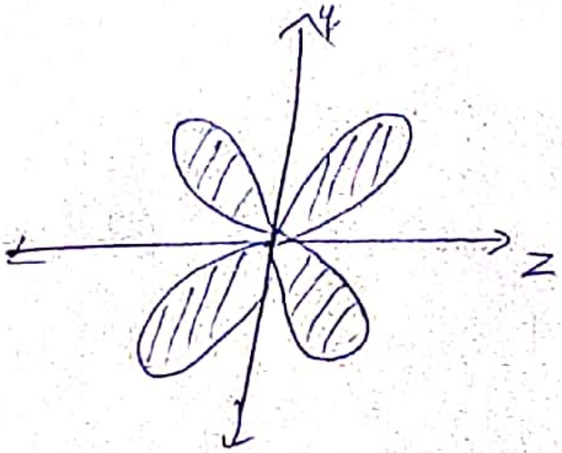
d_{xy} :- (not axial orbital)



Nodal Plane - XZ and YZ

-2

d_{yz} (not axial orbital



Nodal Planes :- XY, XZ

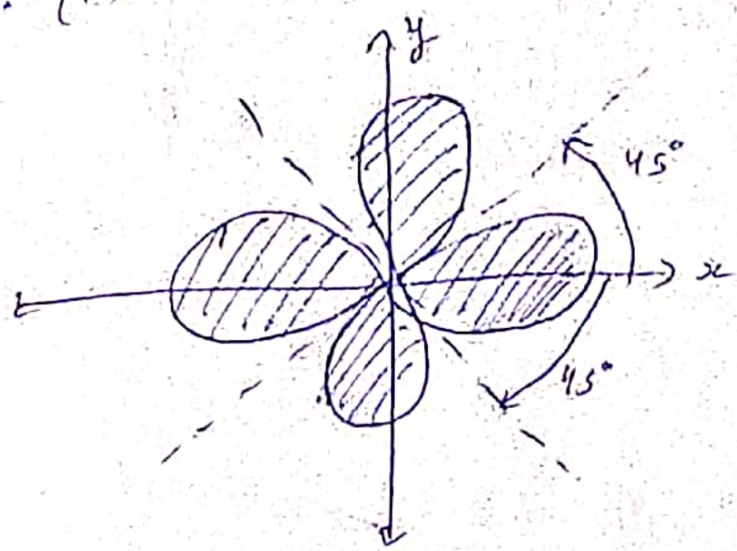
- 2

d_{xz} :- (not axial orbital)

Nodal Planes :- YZ, XY

- 2

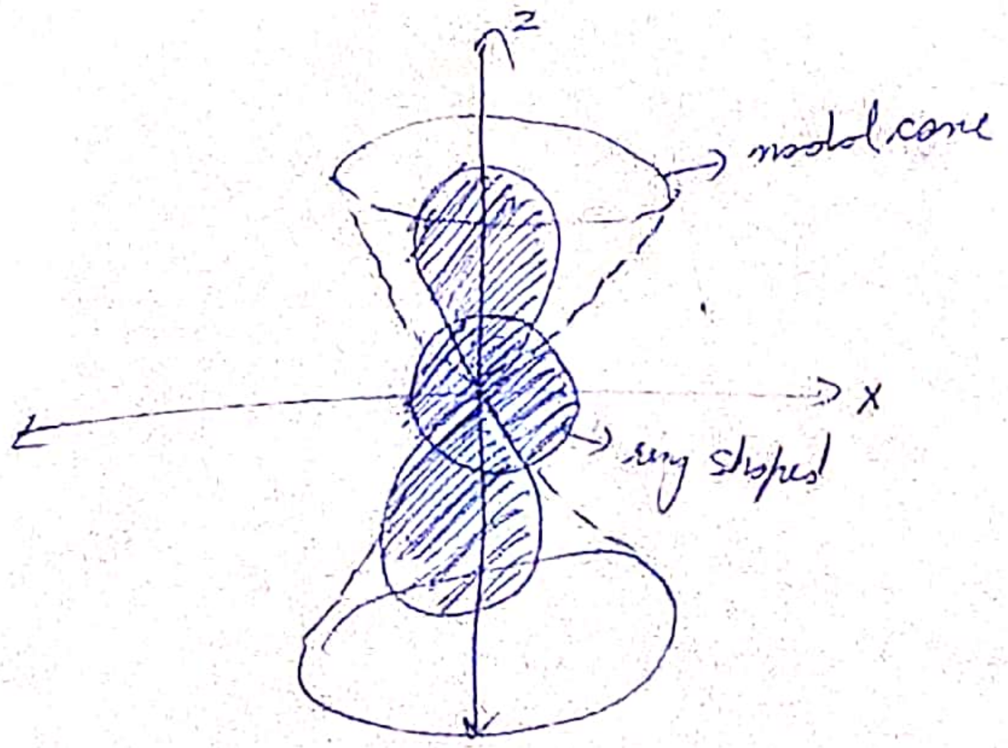
$d_{x^2-y^2}$:- (Axial orbital)



Nodal Planes :- 2

- Two nodal plane inclined at 45° from both X & Y axis

d_{z^2} :- (Axial orbital)



Nodal Planes :- 0

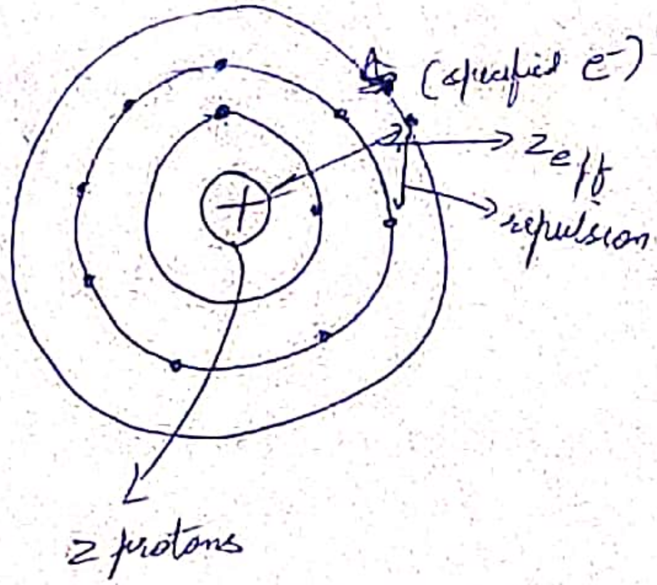
- No nodal planes
- two nodal cones exist

Screening constant (σ) & effective nuclear charge (Z^*/Zeff)

σ → sigma

(Effective Nuclear charge) –

→ It is the force of attraction between nucleus and any specified electron.



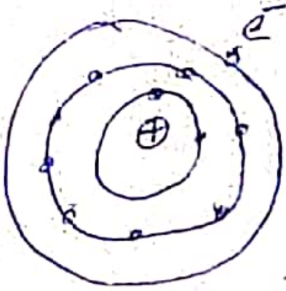
for single e^- system

$\bar{e}$ specified

force of attraction on specifies e^- of Z no. of protons

$Z_{eff} = Z$

for multi e^- system



$\bar{e} \Rightarrow$ force of attraction on specifies $\bar{e}$ of Z no. of protons

$$\underset{\substack{\swarrow \\ \text{net attraction force}}}{Z_{eff}} = \underset{\substack{\downarrow \\ \text{Total attraction}}}{Z} - \sigma \rightarrow \text{Repulsion}$$

Z_{eff} = effective Nuclear charge

Z = atomic no. / no. of proton

σ = slater constant / shielding constant / screening constant

Note:-
→ Due to the presence of inner electron, the attractive force of nucleus on specified electron is reduced.

→ The inner e^- present in an atom apply a repulsive force on specified e^-. This repulsive force of inner shell e^- is called schielding effect or screening effect.

Calculation of slater constant (σ)

Rule 1 - write electronic configo of an atom and arrange ec in following manner.
↓
electronic config.

(1s), (2s, 2p), (3s, 3p), (3d), (4s, 4p), (4d), (4f), (5s)...

Rule 2 - Decide the group of e^- for which σ is to be calculated.

2 Types of groups

~~ns/np~~

ns/np

1. All the e^- present right hand side of the specified e^- contribute nothing to sigma.
2. The remaining e^- in same group (ns/np) contribute 0.35 ~~each~~ each to sigma.
3. The e^- present in (n-1) shell contribute 0.85 each to sigma.
4. The e^- present in (n-2) (n-3) ..... contribute ~~1~~ 1 each to sigma.

nd/nf

1. e^- present right hand side of Nd or Nf grp contribute nothing to sigma.
2. e^- present in same group (nd/nf) contribute 0.35 each to sigma.
3. All e^- present in LHS of (nd/nf) grp contribute 1 each to σ.

Z_{eff} for (4s)e^- of Fe: 3.75
(3d)e^- of Fe: 6.25

It is easier to remove an electron from 4s as compared to 3d because Z_{eff} on 3d is more.

Q1. Calculate value of σ & Z_{eff} for valence shell e^- of Be.

$Be - (1s^2)(2s^2)$ 2s → specified e^- / test e^-

$\sigma = 0.35 + 2(0.85)$

$\sigma = 0.35 + 1.7$

$\sigma = \boxed{2.05}$

$Z_{eff} = Z - \sigma$

$Z_{eff} = 4 - 2.05$

$\boxed{Z_{eff} = 1.95}$

~~Q2. Sol:-~~

Q2. Calculate value of Z_{eff} & σ for 3s e^- of Na

$Na \rightarrow (1s)^2(2s^2 2p^6)(3s^1)$

$\sigma = (8 \times 0.85)(1 \times 2) + (0 \times 0.35)$

$\sigma = 6.80 + 2$

$\sigma = 8.8$

$Z_{eff} = Z - \sigma$

$= 11 - 8.8$

$\boxed{= 2.2}$

Q3. Calculate value of Z_{eff} & σ for 3d e^- of iron, also calculate σ & Z_{eff} for 4s e^-

$$Ec - (1s)^2 \, (2s, 2p)^8 \, (3s, 3p)^8 \, (3d)^6 \, (4s)^2$$

$$3d \ \sigma = (5 \times 0.35)\,(1 \times 18)$$

$$= 18 + 1.75$$

$$\boxed{= 19.75}$$

$$3d \ Z_{eff} = Z - 19.75$$

$$= 26 - 19.75$$

$$\boxed{= 6.25}$$

$$\cancel{4s \ \sigma = (1 \times 0.35)\,(6 \times 0.85)\,(1 \times 18)}$$

$$\cancel{= 18 + 0.35 + 5.10}$$

$$\cancel{\boxed{= 23.45}}$$

$$\cancel{4s \ Z_{eff} = 26 - 23.45}$$

$$\cancel{\boxed{= 2.65}}$$

$$4s \ \sigma = 0.35 + (14 \times 0.85) + (10 \times 1)$$

$$= 10 + 0.35 + 11.9$$

$$\boxed{= 22.25}$$

$$4s \ Z_{eff} = 26 - 22.25$$

$$\boxed{= 3.75}$$

$$\begin{array}{r} 2 \\ 85 \\ \times 14 \\ \hline 340 \\ 850 \\ \hline 11.90 \end{array}$$

$_3Li$ - $(1s)^2 (2s)^1$

$a = 0.85 \times 2$
$= 1.7$

$Z_{eff} = Z - a$
$= 3 - 1.7$
$\boxed{= 1.3}$

$_4Be$ - $(1s)^2 (2s)^2$

$Z_{eff} = 4 - (1 \times 0.35 + 2 \times 0.85)$
$= 4 - 2.05$
$\boxed{= 1.95}$

$_5B$ - $(1s)^2 (2s, 2p)^3$

$Z_{eff} = 5 - (2 \times 0.85 + 2 \times 0.35)$
$= 5 - 2.4$
$\boxed{= 2.6}$

$_6C$ - $(1s)^2 (2s, 2p)^4$

$Z_{eff} = 6 - (2 \times 0.85 + 3 \times 0.35)$
$= 6 - 2.75$
$\boxed{= 3.25}$

$_7N$ - $(1s)^2 (2s, 2p)^5$

$Z_{eff} = 7 - (2 \times 0.85 + 4 \times 0.35)$
$= 7 - 3.1$
$\boxed{= 3.9}$

$_8O - (1s)^2 \; (2s, 2p)^6$

$$Z_{eff} = 8 - (2 \times 0.85 + 5 \times 0.35)$$

$$\cancel{= 8 - (1.7 + 2.10)}$$

$$\cancel{= 8 - 3.8}$$

$$= 8 - (1.7 + 1.75)$$

$$= 8 - 3.45$$

$$\boxed{= 4.55}$$

$_9F - (1s)^2 \; (2s, 2p)^7$

$$Z_{eff} = 9 - (2 \times 0.85 + 6 \times 0.35)$$

$$= 9 - 3.8$$

$$\boxed{= 5.2}$$

$_{10}Ne - (1s)^2 \; (2s, 2p)^8$

$$Z_{eff} = \cancel{18}\; 10 - (2 \times 0.85 + 7 \times 0.35)$$

$$= 10 - 4.15$$

$$\boxed{= 5.85}$$

$_{11}Na - (1s)^2 \; (2s, 2p)^8 \; (3s)^1$

$$Z_{eff} = 11 - (2 \times 1 + 8 \times 0.85 + 0 \times 0.35)$$

$$= 11 - (2 + 6.80 + 0)$$

$$= 11 - 8.8$$

$$\boxed{= 2.2}$$

Q Calculate the Z_{eff} of $3d\ e^-$ of Cr^{3+} ?

$_{24}Cr^{3+}$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \cancel{4s^1}\ 5d^{5}$

$(1s)^2\ (2s, 2p)^8\ (3s, 3p)^8\ (5d)^3$

$\sigma = (7 \times 0.35) + (8 \times 0.85) + 2$

$= 2.45 + 3.60 + 2$

$= 6.05 + 2$

$= 18.70$

$Z_{eff} = \cancel{24 - 5.05} = \cancel{18.95}$ $\quad 24 - 18.7$ $\boxed{5.3}$

---

for ns, np e^-

$\sigma = (w \times 0) + (x-1) \times 0.35 + (y \times 0.85) + z \times 1$

for nd or nf e^-

$\sigma = (w \times 0)(x-1) \times 0.35 + y \times 1$

Note:- for $1s^2\ e^-$ configuration:- $\sigma = 0.3$

eg. He - $1s^2$
$\sigma = 0.3$
$Z_{eff} = 2 - 0.3$
$= 1.7$

Li^{+} - $1s^2$
$\sigma = 0.3$
$Z_{eff} = 3 - 0.3$
$= 2.7$

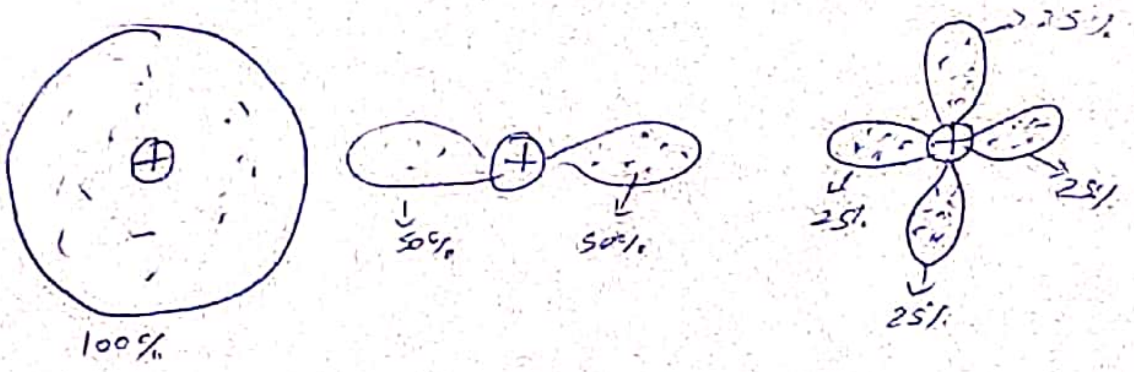
Ni - $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$
$\sigma = 0.3$
$Z_{eff} = 28 - 0.3$
$= 27.7$

Note :- For H atom , $Z_{eff} = 1$

Penetration Power - It is the ability of an orbital to attract e^-

Order - $ns > np > nd > nf$

(How easy can electron approach to nucleus)



Shielding Power / Screening power :-

Order - $ns > np > nd > nf$