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**Unit - 3**

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DATE: / / 20  
PAGE NO:

### Periodic properties

① Effective Nuclear charge  $\leftarrow$  As electrons in an atom are attracted by the nucleus, and at the same time repelled by other electrons in atom. Effective Nuclear charge is the average nuclear charge felt by individual electrons in an atom, taking into consideration the "shielding effect of inner shell electrons".

$$Z_{\text{eff}} = Z - \frac{n}{r} \rightarrow S \rightarrow \text{inner shell electrons.}$$

Proton in the nucleus

Example  $\leftarrow$  what is the effective nuclear charge felt by an electron in the n=3 shell of sulphur?

Ans  $\leftarrow$   $Z_{\text{eff}} = Z - S$

Sulphur At No = 16

No. of proton = 16

No. of electrons = 16 [2, 8, 6]

$$Z_{\text{eff}} = 16 - 10$$

$\Rightarrow$  for n=3

Example  $\leftarrow$  what is the effective nuclear charge felt by an electron in n=2 shell of chlorine?

Ans  $\leftarrow$   $Z_{\text{eff}} = Z - S$

Atomic no (Cl) = 17

No. of protons = 17

No. of electrons = 17 [2, 8, 7]

$$Z_{\text{eff}} = 17 - 2 = 15$$

## Slater's Rule for Effective Nuclear Charge

- ① Write the electronic configuration of the atom in the following form

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

- ② Identify the electrons of interest and ignore all electrons in higher groups (to the right in list from Rule 1). These do not shield electrons in lower groups.

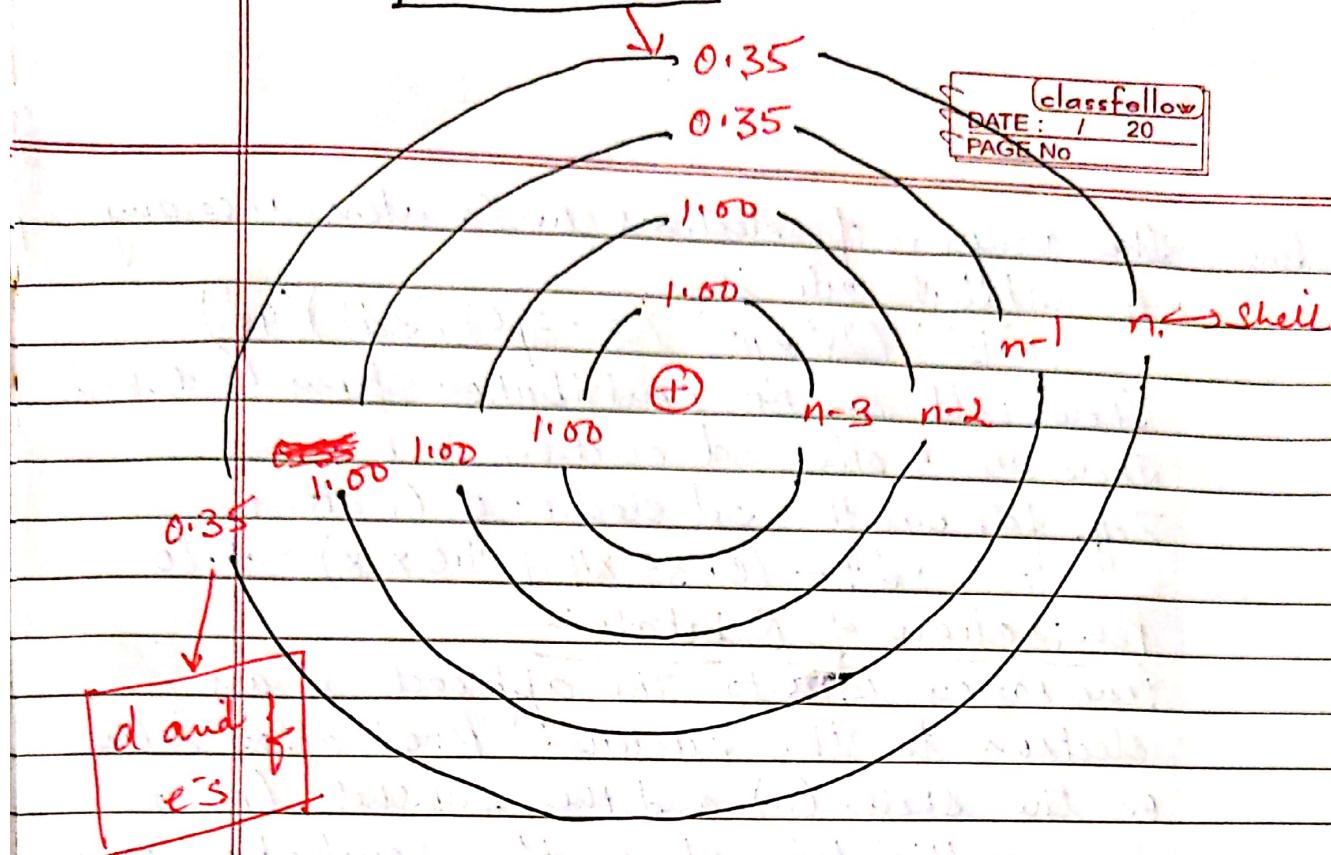
- ③ Slater rule is now broken into two cases

- ① The shielding experienced by an s or p electrons  
 → electrons within same group shield 0.35,  
 except the 1s which shield 0.30.  
 → electrons within the  $(n-1)$  group shield 0.85  
 → electrons within the  $(n-2)$  or lower groups shield 1.00

- ② The shielding experienced by nd or nf valence electrons

- electrons within same group shield 0.35  
 → electrons within the lower group shield 1.00  
 These rules are summarized in figures and Table

## 8 S and p electrons



### Contribution of Shielding

Group	Higher grp	Same grp	$(n-1) \text{ grp}$	$(n-2) \text{ grp.}$
1s	0	0.30	—	—
ns, np	0	0.35	0.85	1.00
nd, nf	0	0.35	1.00	1.00

Example What is the effective nuclear charge felt by 1s electron of helium atom?

Ans:- The electronic configuration of helium atom is  $1s^2$ . There is one more 1s electron, apart from the electron for which  $Z_{eff}$  is to be calculated.

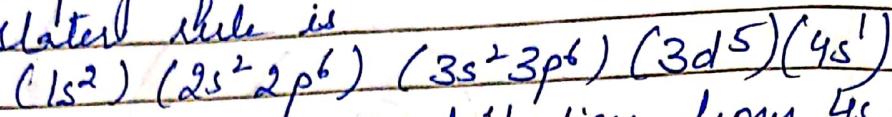
$$\text{The Screening constant } S = 1 \times 0.30 = 0.30$$

Hence  $Z_{eff}$  felt by 1s electron of He atom

$$= 2 - 0.30 = 1.70$$

Ques-2 What is the effective nuclear charge felt by a 3d electron of chromium atom ( $Z=24$ )

Ans) The grouping of electrons in Cr atom according to Slater's rule is



There will be no contribution from 4s electron.  
There are 4 other 3d electrons.

$Z_{\text{eff}}$  felt by the 3d electron of Cr atom  
 $= 24 - (0.35 \times 4 + 1.0 \times 18) = 4.60$

### Penetration of Orbitals :-

Penetration refers to the approach of an electron to the nucleus. Penetration relies on the shell ( $n$ ) and the subshell ( $m$ ).

In a multi-electron atom, the penetration power of an electron depends upon the value of both the shell and the sub-shell.

The penetration power of an electron within the same shell value ( $n$ ) follows the following trend in subshells:  $s > p > d > f$ .

For different value of shell ( $n$ ) and subshell ( $l$ ), penetration power of an electron is as follows.

$$1s > 2s > 2p > 3s > 3p > 4s > 3d > 4p > 5s > 4d > 5p \\ > 6s > 4f$$

The energy of an electron for each shell and sub-shell follows the following trend

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$$

Electrons having greater penetration have strong attraction to the nucleus, less screening and hence larger effective nuclear charge.

Electronic Configuration :- The distribution of electrons in different shell, Subshell

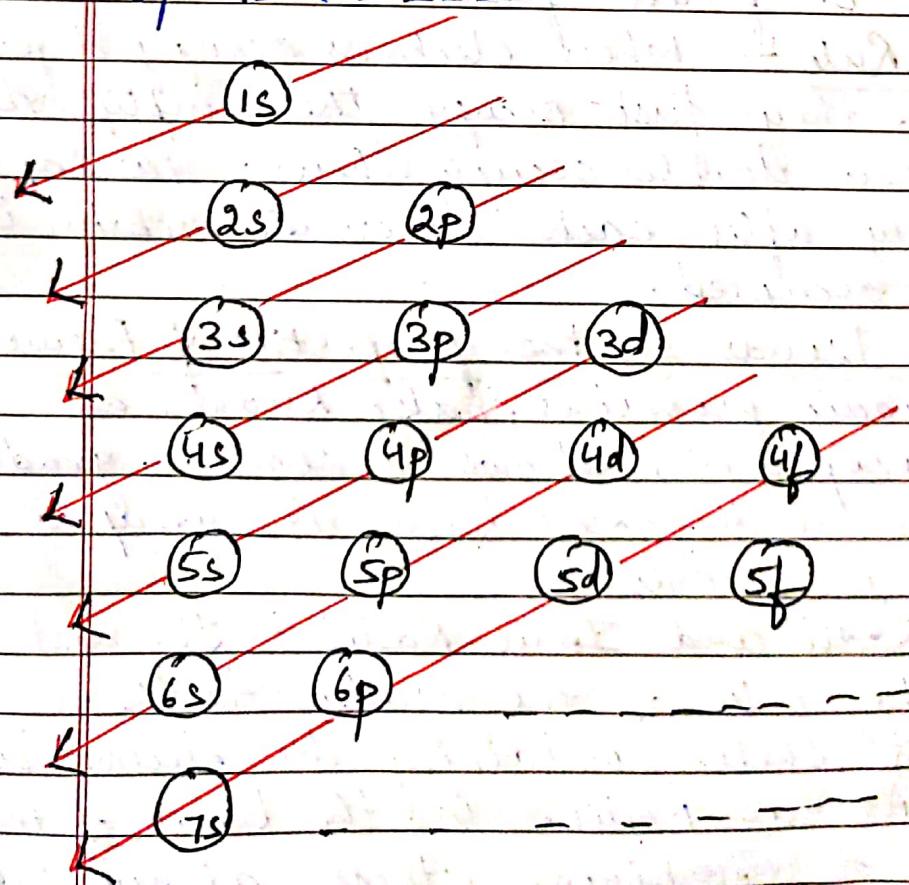
and orbital of an atom is called electronic configuration.

Rules followed for filling electrons:

(1) Pauli's exclusion Principle:— An orbital can have maximum two electrons, that too with opposite spin. No two electrons of an atom or molecule can have the equivalent four electronic quantum numbers.

(2) Aufbau's Principle:— Electrons are filled in various orbitals (subshells) in increasing order of their energies. Orbital with lower energy is filled first followed by orbital with higher energy.

The order of energy for various subshell/orbital.  
 1s 2s 2p 3s 3p 3d 4s 4p 4d 5s 4f 5p 5d 6s 6p 6f 7s ---



(n+l) rule + The orbital for which the sum of principal quantum no and azimuthal quantum no is less has lower energy, it will be filled first and that for which it is higher, has higher energy and will be filled later.

- ② For orbitals having same (n+l) value, the orbital with lower value of Principal Quantum Number is filled first.

Ques why does 4s fill prior to 3d?

Ans -

$$\begin{array}{ccccc}
 & & 4s & & 3d \\
 n = 4 & & & 3 & \\
 l = 0 & & & 2 & \\
 n+l = 4 & & & 5 &
 \end{array}$$

So. n+l value for 4s is lower to 3d, it will be filled first.

Hund's Rule - When electrons occupy degenerate orbitals, they first occupy the empty orbitals and then double occupy them. Pairing occurs only after each degenerate orbital is singly occupied.

Periodic Trends in the properties of Elements - The periodic properties, also known as atomic properties, shows a regular gradation (increase or decrease) both in group as well as in periods.

Atomic Radii and Ionic Radii - The distance from the centre of the nucleus to the outermost shell containing the electrons is called Atomic Radii. On the basis of the nature of combining atoms, atomic radii

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DATE: / / 20
PAGE No.

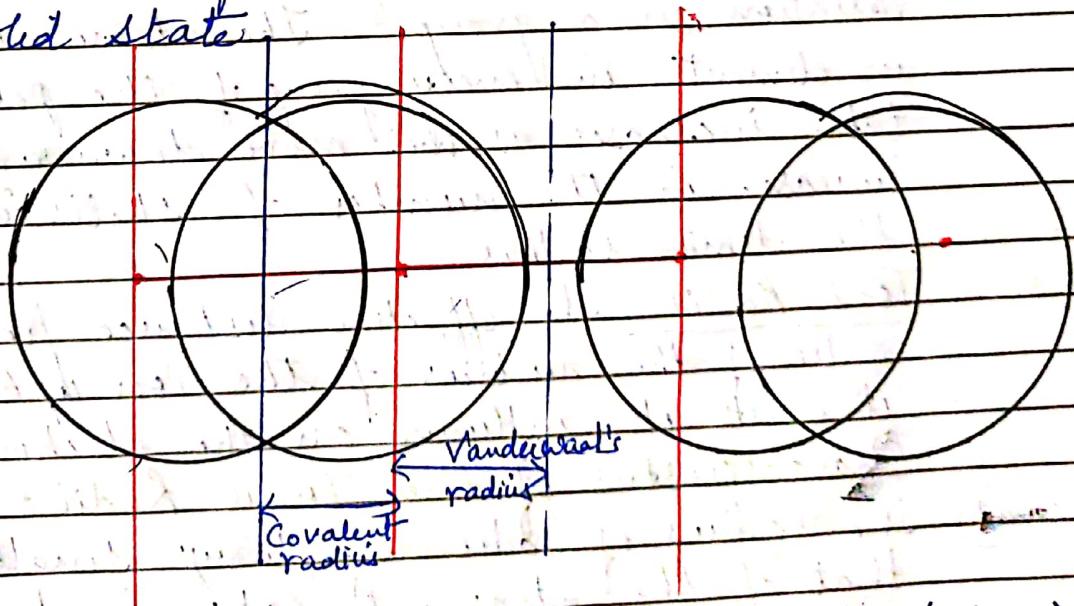
Can be classified into the following :-

- (1) Covalent radii
- (2) Vanderwaal's radii
- (3) Metallic radii
- (4) Covalent radii :- In homoatomic molecules containing same type of atoms, Covalent radius may be defined as :-

"One half of the distance between the centres of the nuclei of two adjacent covalently bonded atoms".

$$\text{Covalent Radius} = \frac{1}{2} (\text{Bond length})$$

- (2) Vanderwaal's Radii :- Half of the inter-nuclear distance between two similar adjacent atoms belonging to the two neighbouring molecules of the same substance in the solid state.



- (3) Metallic Radii :- Half of the distance between the centre of nuclei of two adjacent atoms in metallic crystals / Ionic compounds.

A comparison of three atomic radii

Shows that Vanderwaal's radius is maximum while covalent radius has the least value.

The order is

Vanderwaal's radius > Metallic radius >  
Covalent radius.

### Variation of Atomic Radii in the periodic

Table :-

Variation in a Period! - Along a period, the atomic radii of elements generally decreases from left to right. This is due to

to increase in the effective nuclear charge on the given element by which electron feels more attraction towards nucleus, and results in decrease in radius.

Example O F Ne

1.4 1.2 1.6

Variation in a group! - The atomic radii of an element in every group of the periodic table increases as we move downwards.

Increase in the number of shells increases the atomic radius.

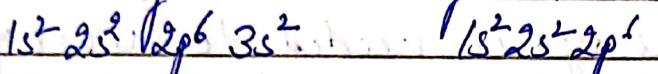
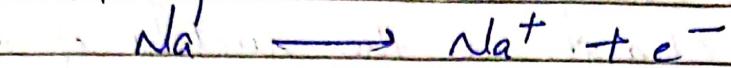
Ionic Radii - The effective distance from the centre of the nucleus of the ion upto which it exerts its influence on the electron cloud.

Characteristics of the Ionic Radii :-

- ① The radius of cation is always smaller than that of the atom. A cation is formed by loss of one or more electrons from the gaseous atom. In most of the cases, the electrons present in valence shell of the atom are completely removed so that the

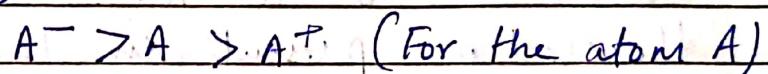
cations formed may have the configuration of nearest noble gas atom.

Example:



(2) The radius of the anion is always larger than that of the atom.

(3) In case of isoelectronic, more the number of protons, smaller will be the radius and more attraction between electrons & nucleus.



Ques:- (a)  $\text{Na}^+$  E = 10 P = 11

(b)  $\text{Al}^{3+}$  E = 10 P = 13

(c)  $\text{Mg}^{2+}$  E = 10 P = 12

Ans:- a > c > b

More the protons in case of b, so smallest is the radius.

Ionisation Energy - Minimum amount of energy required to remove the most loosely held outermost electron from an isolated neutral gaseous atom in its ground state.

Units -  $\text{KJ mol}^{-1}$

Factor affecting Ionization Energy -

(1) Atomic Size - With the increase in atomic size, the no. of shell increases. Therefore the force that binds the electron with nucleus decreases. The ionization energy

thus decrease with increase in the atomic size.

- (2) Nuclear Charge:- As the magnitude of positive charge on the nucleus of atom increases, its attraction with the electron also increases. This means greater energy is needed to remove the electron present in Valence shell of the gaseous atom. Therefore ionization energy increases with increase in the nuclear charge.
- (3) Screening / Shielding effect of inner shell electrons:- In a multi-electron atom, the electron present in the inner shells shield the electron present in Valence shell from the attraction of the nucleus. This is called Shielding effect. As a result, these electrons do not experience the complete charge of the nucleus. The actual charge felt by Valence electrons from the nucleus of the atom is called Effective Nuclear charge. Greater the screening effect, lesser will be the ionisation energy.

- (4) Penetration effect of the electrons:- More the penetration of electrons towards the nucleus of atom, more will be the ionisation energy.

Variation of Ionisation Energy in Periodic Table?

Along a period! - Ionisation energy is expected to increase on moving left to right, as nuclear charge increases and atomic size decreases.

Down a group! - The ionisation energy of the element decrease on moving from top to bottom in any group as there is increase in number of shells in moving from one element to another.

- (6) There is also increase in the magnitude of the screening effect due to gradual increase in number of inner electrons.

Electron Gain Enthalpy! - Energy released when the atom gains an electron in the gaseous state. It is denoted by (-ve) sign.

Factors affecting Electron gain enthalpy!

- (1) Atomic Size! - As the size of an atom increases, the distance between its nucleus and incoming electron also increases. So incoming electron experience lesser attractions towards the nucleus of the atom. So, electron gain enthalpy become less negative down the group.

- (2) Nuclear Charge! - With the increase in nuclear charge, forces of attraction between nucleus and incoming electron increases and so the value of electron gain enthalpy decreases. The electron gain enthalpy become more negative with increase in nuclear charge.

- (3) Symmetry of electronic Configuration!

The atoms with symmetrical configuration (half filled or fully filled) don't have any urge to take up extra electron. So electron gain enthalpy will be positive.

Electronegativity! - The tendency of an atom to attract shared pair of electron towards itself in a chemical bond is called electronegativity.

Factors affecting Electronegativity?

Effective nuclear charge - The value of electronegativity increases with increase in effective nuclear charge.

Atomic Radius - The value of electronegativity decrease with increase in atomic radius.

Oxidation States - The value of electronegativity increase with increase in oxidation state.

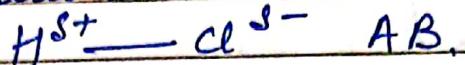
Hybridisation state of an atom in the molecule - The value of electronegativity increases with the increase in s-character in the hybridisation state of an atom.

The value of electronegativity also increase due to the closeness of  $\sigma$ -electrons to the nucleus.

Variation of electronegativity in Periodic Table.

The value of electronegativity increases along a period due to decrease in atomic size and increase in effective nuclear charge and the value of electronegativity decreases down a group to increase in atomic size of an element.

Uses of Electronegativity :- Finding bond between two atoms and also to determine its nature example if covalent bond is formed by two or more than two atoms, then due to electronegativity difference, polarization is calculated



$$X_A = E.N \text{ of } A$$

$$X_B = E.N \text{ of } B$$

If  $X_A - X_B = 0 \rightarrow$  Pure covalent

If  $(X_A - X_B) > 0 < 1.7 \rightarrow$  Polar covalent

$X_A - X_B = 1.7 \rightarrow 50\% \text{ Ionic}, 50\% \text{ covalent}$

$X_A - X_B \gg 1.7 \rightarrow \text{Mainly Ionic}$

### Difference between Electron affinity & Electronegativity

#### Electron Affinity

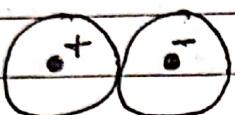
- ① It is the energy release to attract outside electron.
- ② It is absolute electron attracting tendency of the atom.
- ③ It is the property of an isolated atom.
- ④ It has specific unit KJ/mol.

#### Electronegativity

- ① It is tendency of an atom to attract shared pair of electrons.
- ② It is relative electron attracting tendency of an atom.
- ③ It is the property of bonded atoms.
- ④ It has no units.

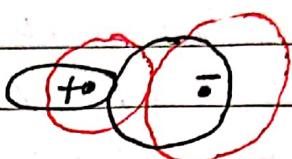
Polarizability :- Anions are larger in size than cations and therefore, their electron clouds are less tightly held. When two

oppositely charged ions approach each other closely, the positively charged cations attract the outermost electrons of the anions and repels its positively charged nucleus. This result in distortion or polarization of the anion as illustrated in figure

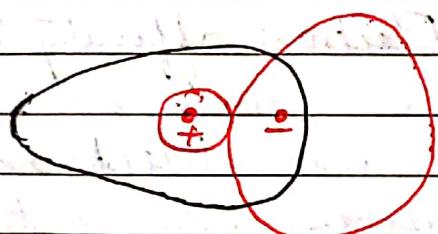


(a)

Idealised ion pair  
with no polarization



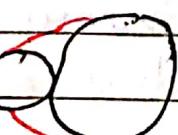
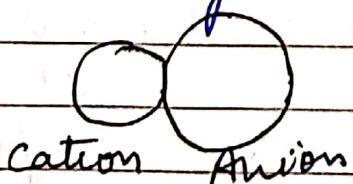
mutually polarized ion



Red lines  
represents  
hypothetical  
unpolarized ions

Polarization sufficient  
to form covalent bond

When a large and soft anions comes under the influence of small cations, the cation is able to polarize the anion. Due to compact charge cloud of the cation, polarization of cation is very small.



polarized Anion

The polarizability of the anion depends both on its size and charge. Anions of lower charge density are more polarizable. For anions having the same charge, the larger anion is polarized to a greater extent ( $\text{F}^- < \text{Cl}^- < \text{Br}^- < \text{I}^-$ ). For anions of similar size, the more polarizable anion has a greater negative charge ( $\text{S}^{2-} > \text{Cl}^-$ ). This is because a larger charge indicates a larger excess of electrons over protons. For polarization (covalent character) to be significant for a bond, the cation must be polarising and the anion must be polarizable.

Neutral non-polar species have spherically symmetrical arrangement of electrons in their electron cloud. While in the presence of electric field, their electron clouds are distorted. The ease of this distortion is defined as Polarizability of the atom / molecule.

#### Factors influencing Polarizability :

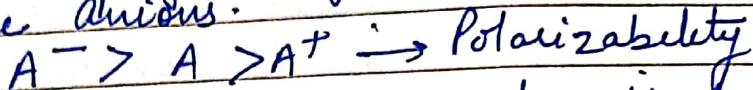
- (1) Greater the number of electrons, less effective nuclear charge, higher will be its polarizability (as nucleus have less control over outer electrons)
- (2) Greater the size of atom/ion, higher will be the polarizability.

#### Variation of Polarizability in the Periodic Table:

(1) Polarizability increases down a group, as atomic size increases and electron cloud could be easily distorted

⑤ Polarizability decreases across a period as size decrease and effective nuclear charge increases.

Cations are smaller than their parent atom, so they generally polarize the highly polarizable anions.



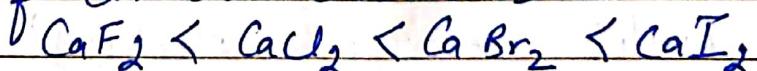
Fajan's Rule: ① A compound with large cation and smaller anion has ionic bond.

② A compound with high positive charge means smaller cation and larger anion are covalently bonded.

Ques: Write decreasing order of Melting Point and increasing order of covalent character in  $\text{CaF}_2$ ,  $\text{CaI}_2$ ,  $\text{CaCl}_2$ ,  $\text{CaBr}_2$ .

Ans: Melting point depends on Ionic character or attraction between two opposite ions, strong attraction will be having more ionic character, so decreasing order of Melting Point  $\text{CaF}_2 > \text{CaCl}_2 > \text{CaBr}_2 > \text{CaI}_2$

② According to Fajan's Rule covalent character increases with increasing size of anion and decreasing size of cation, so increasing order of covalent character



Quest: LiCl is more soluble in pyridine than in  $\text{H}_2\text{O}$ . Why?

Ans: According to Fajan's Rule, smaller is the cation greater is the covalent character, hence LiCl has covalent character thus it is more soluble in Organic solvents like

pyridine than  $H_2O$ .

Ques: Write decreasing order of covalent character in  $NaCl$ ,  $MgCl_2$  &  $AlCl_3$ .

Ans: As per Fajan's rule, cation having more positive charge is more covalent.

Order of Covalent character :  $AlCl_3 > MgCl_2 > NaCl$

Oxidation States :- It is defined as the number representing the electric charge on the central metal atoms of a complex ion.

Following are the rules to determine the Oxidation Number :-

- 1) Ion must be treated separately. Each ion has its own Oxidation number.
- 2) The Oxidation numbers of atoms in the pure neutral element are zero. Example  $H_2$ ,  $F_2$ ,  $O_2$ ,  $Py$ ,  $S_8$  have Oxidation number = 0
- 3) The most electronegative atom in a polyatomic species has an Oxidation number equal to number of Valence electron minus eight.
- 4) Hydrogen has Oxidation no. of +1, except when it is bonded to a metal.
- 5) Fluorine is most electronegative element and has 7 valence electrons. Therefore the Oxidation number of a compound containing fluorine is  $7-8 = -1$
- 6) The sum of Oxidation numbers of all the atoms is equal to the net charge of the species. For Example, Oxidation state of  $Ca^{2+}$  is +2.

Hard and Soft Acids and Bases :- Acids and Bases can also be classified on the basis of

their polarizability. Higher polarizability results in softer species and lower polarizability results in harder species.

A species having electron pairs of low polarizability is called hard base. They also have high electronegativity.  $F^-$  is the anion of most electronegative element. Smaller is the size of negative ions, lesser will be its polarizability and harder will be the base. Molecules and ions containing Oxygen or Nitrogen atoms are hard bases but relatively less hard than Fluorine.

For Example -  $H_2O$ ,  $CH_3OH$ ,  $NH_3$  and  $N_2NCH_3$

A species having electron pairs of high polarizability is called soft base. They have low electronegativity. As the size of anions increase, more will be its polarizability, softer will be the base.  $I^-$  is highly polarizable anion and is very soft base. Other molecules and polyatomic anions with donor atoms that belongs to rows 3 to 6 are also called soft bases.

For Example:  $CN^-$ ,  $SCN^-$ ,  $I^-$  etc

A species having acceptor atoms with low polarizability is called Hard Acids. Most of the metal atoms and ions are hard acids. Acids having smaller ionic radius and higher oxidation state are Harder. The low polarizability is the result of higher effective nuclear charge.

For Example -  $Li^+$ ,  $Na^+$ ,  $K^+$ ,  $Al^{3+}$ ,  $Cr^{3+}$  etc

A species having high polarizability is called soft Acid. Atoms having a larger size and low Oxidation states are soft Acid.

Example :  $\text{GaCl}_3$ ,  $\text{GaBr}_3$ ,  $\text{GaI}_3$ ,  $\text{Cu}^+$ ,  $\text{Ag}^+$ ,  $\text{Au}^{+}$  etc.

### Co-ordination Number and Geometries

The no. of atoms of the ligands that are directly bound to the central metal atom or ion by co-ordinate bonds is known as Co-ordination number of central metal atom or ion.

OR

The total no. of co-ordination bonds formed by the ligands in the complex.

$$\text{C.N} = \text{No. of ligands} \times \text{Denticity of ligands.}$$

ligand	Name	Denticity
1) $\text{NH}_3$	amine	monodentate
2) $-\text{NH}_2$	ethylene diammine (en)	bidentate
3) $-\text{NH}_2$	diethylenetriamine (dien)	Tridentate
4)	EDTA	Hexadentate

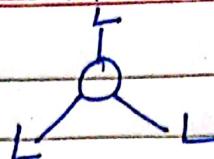
Co-ordination Number	Geometry	Structure	Example
2	Linear	$L - M - L$	$\text{Ag}^{+}$ in silver cyanide, $[\text{Ag}(\text{CN})_2]^-$ in $\text{KAg(CN)}_2$

C.N

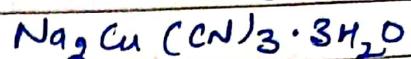
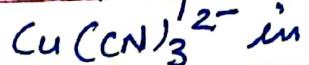
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Geometry  
Trigonal  
Planar

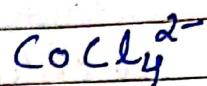
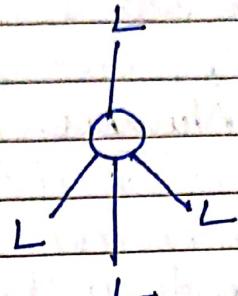
Structure



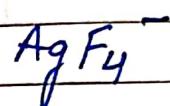
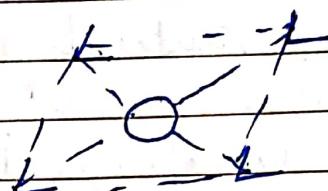
Example



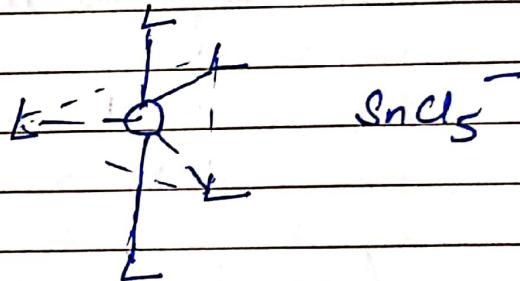
4. Tetrahedral  
Geometry



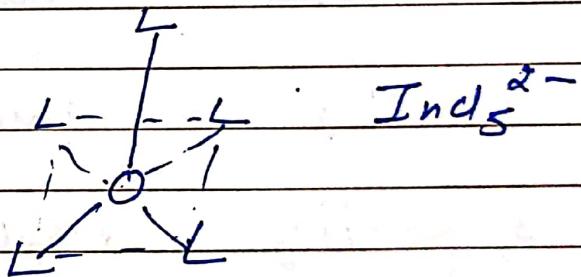
5. Square planar  
Geometry



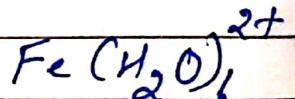
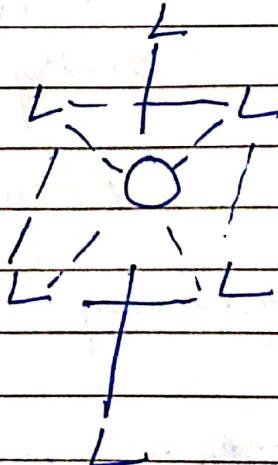
5. Trigonal bipyramidal  
Geometry



6. Square pyramidal  
Geometry



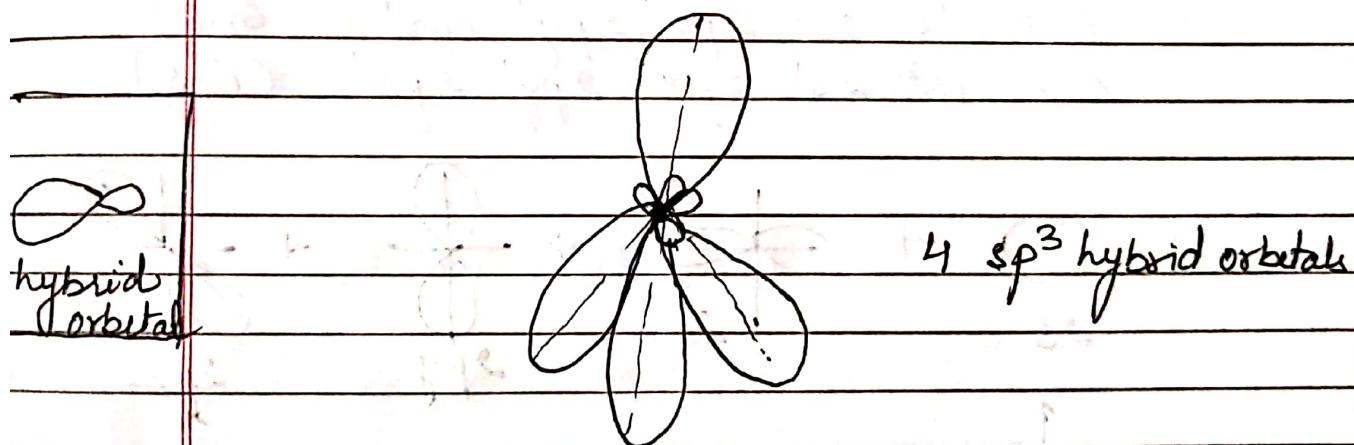
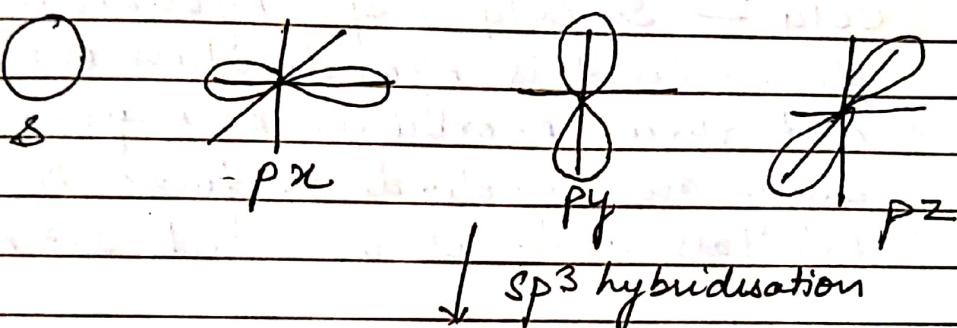
6. Octahedral  
Geometry



## Molecular Geometries ( $\text{H}_2\text{O}$ , $\text{NH}_3$ , $\text{PCl}_5$ , $\text{SF}_6$ , $\text{CCl}_4$ )

The phenomenon of mixing of atomic orbitals belonging to same shell of an atom having slightly different energies so that redistribution of energy takes place between them resulting in formation of equal no. of new orbitals of equal energies and identical shape called hybrid orbital called hybridisation.

Example :-



### Prediction of hybridisation :-

- ① Count the total Valence electrons.
- 2) If total number of Valence electron is less than 8, then divide by 2.
- 3) If total no. of Valence electron is greater than 8, then divide by 8.
- 4) If maximum remainder comes out to be more than 1, then again divide the remainder

by 2 ..

Result :- Quotient of 1 divisor + Quotient  
of 2 divisor + Remainder

Result

2

3

4

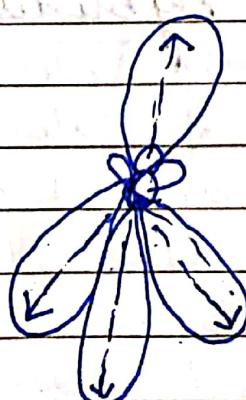
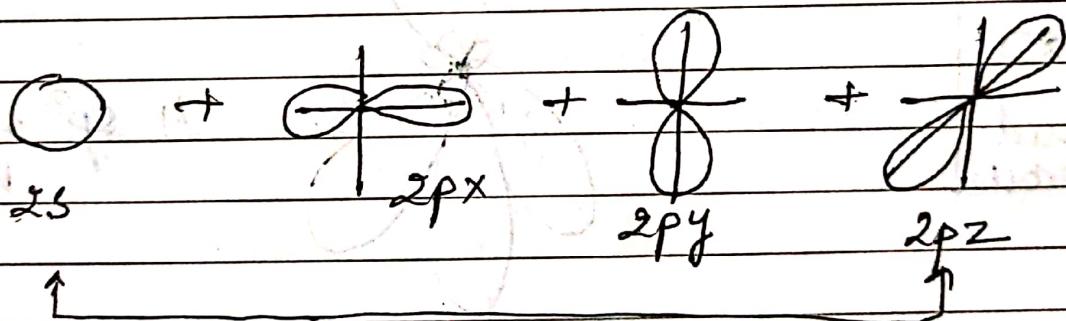
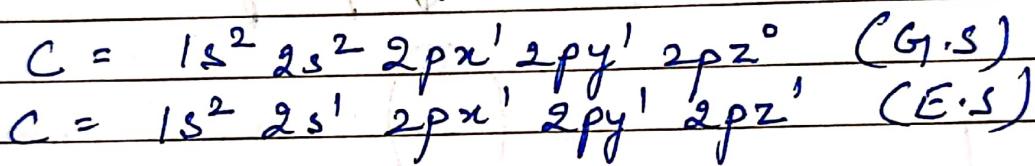
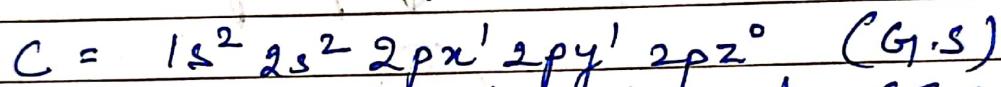
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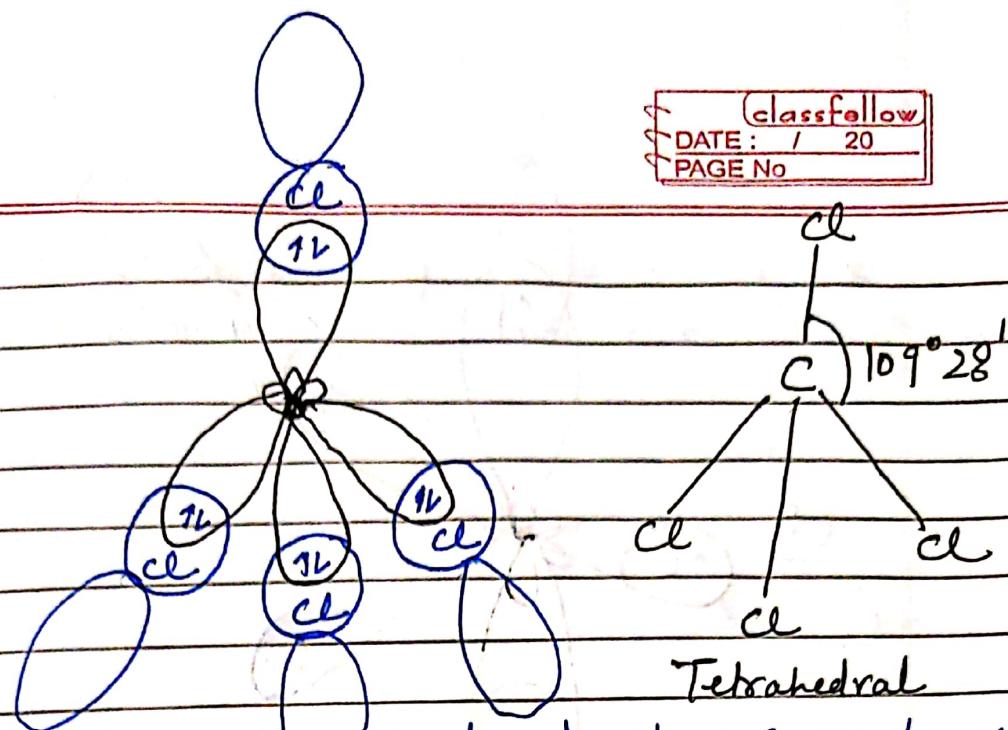
7

Hybridisation $sp$  $sp^2$  $sp^3$  $sp^3d$  $sp^3d^2$  $sp^3d^3$ 

CCl<sub>4</sub> :- In CCl<sub>4</sub>, sp<sup>3</sup> hybridisation takes place. In this hybridisation one s-orbital and three p-orbital belonging to same shell are involved to form 4 hybrid orbitals and it is called sp<sup>3</sup> hybridisation.



4 sp<sup>3</sup> hybrid orbitals  
of C-atom



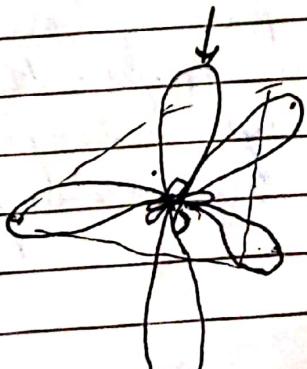
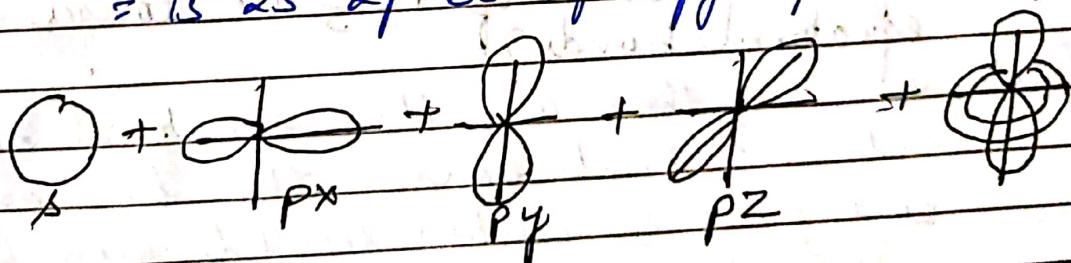
In  $\text{CCl}_4$ , the central atom C undergoes  $\text{sp}^3$  hybridisation to form 4 hybrid orbitals having one unpaired electron in each.

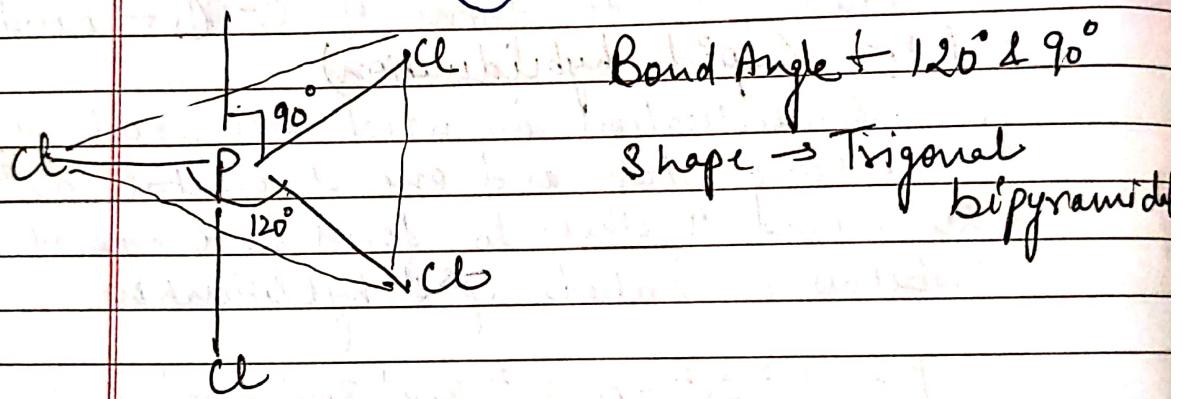
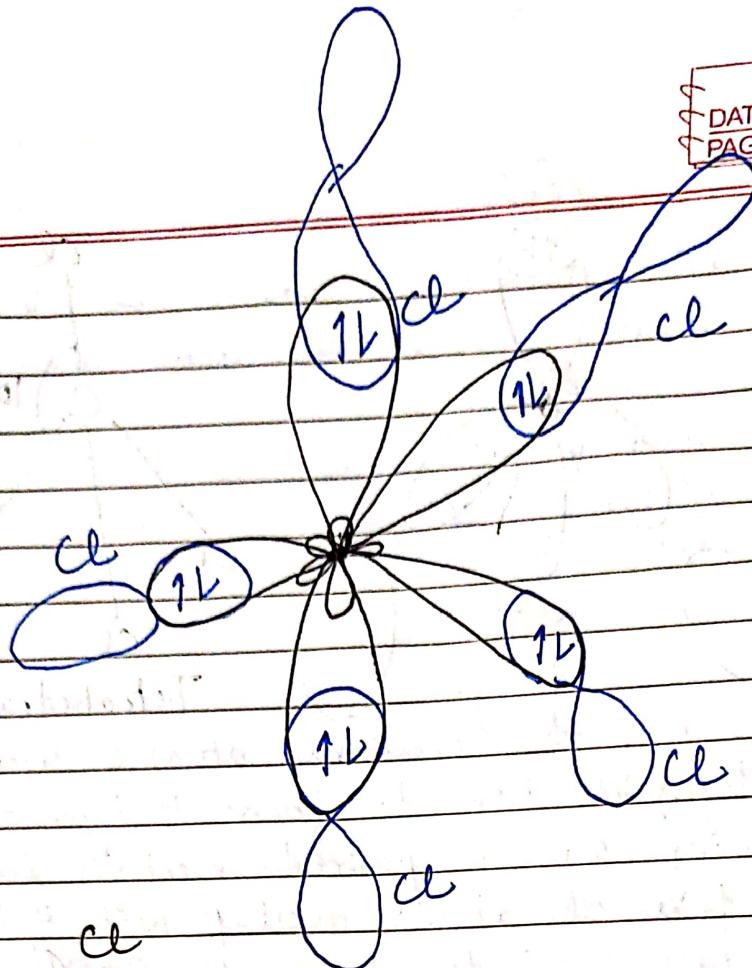
The four Cl atoms overlap with 4 hybrid orbitals of C to form 4  $\text{C}-\text{Cl}$  bond.

$\text{PCl}_5$  ( $\text{Csp}^3\text{d}$  hybridisation)

The hybridisation in which one s-orbital, three p-orbitals and one d-orbital ( $\text{d}_{z^2}$ ) are mixed together to form five hybrid orbitals is called  $\text{sp}^3\text{d}$ -hybridisation.

$$\begin{aligned} \text{P} &= 1s^2 2s^2 2p^6 3s^2 3p^6 3p_x^1 3p_y^1 3p_z^1 \text{ (G.S)} \\ &= 1s^2 2s^2 2p^6 3s^1 3p^6 3p_y^1 3p_z^1 3d_{z^2}^1 \text{ (E.S)} \end{aligned}$$





Ques: Why Axial bonds of  $\text{PCl}_5$  are longer than equatorial bonds?

Ans: In  $\text{PCl}_5$ , there are three equatorial bonds and two axial bonds. The two axial bonds pairs of electrons repel more by three equatorial bond pair of electrons. Hence axial bonds are longer than equatorial bonds.

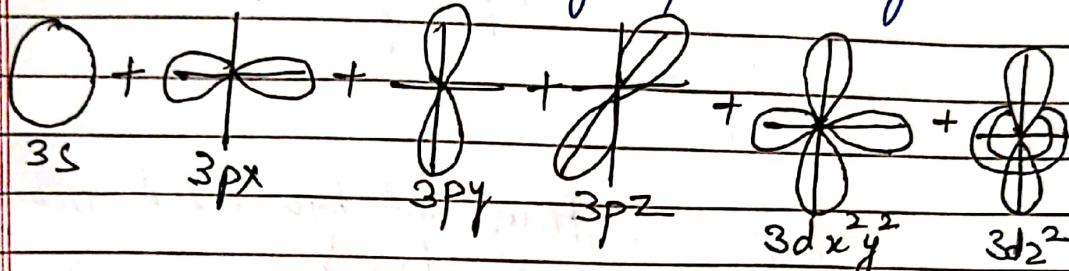
$$\text{Axial bond length} = 219 \text{ pm}$$

$$\text{Equatorial bond length} = 204 \text{ pm}$$

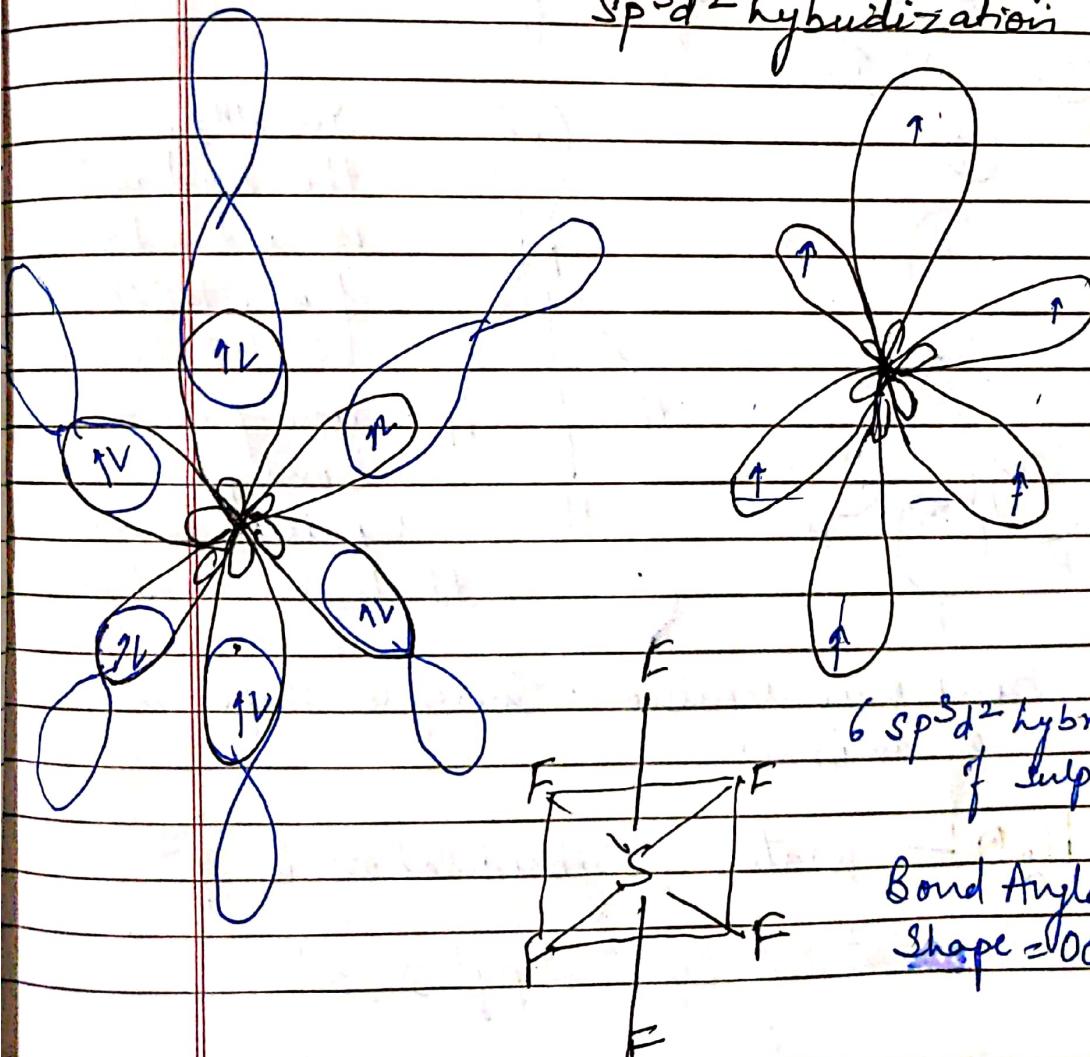
$sp^3d^2$  hybridisation :- The hybridisation in which one s-orbital and three p-orbitals and two d-orbitals ( $3d_{x^2-y^2}$ ),  $d_{z^2}$ , mixed together to form 6 hybrid orbitals is called  $sp^3d^2$  hybridisation.

Example : SF<sub>6</sub>

$$\begin{aligned} S &= 1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^1 3p_z^1 \text{ (G.S)} \\ &= 1s^2 2s^2 2p^6 3s^1 3p_x^1 3p_y^1 3p_z^1 3d_{x^2-y^2}^1 3d_{z^2}^1 \end{aligned}$$



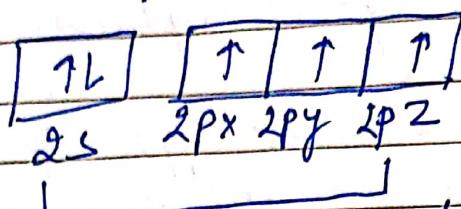
$sp^3d^2$  hybridization



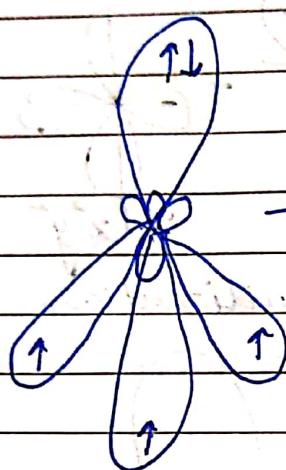
## Exceptional Hybridisation :-

Example:  $\text{NH}_3$  (Ammonia)  $\text{sp}^3$

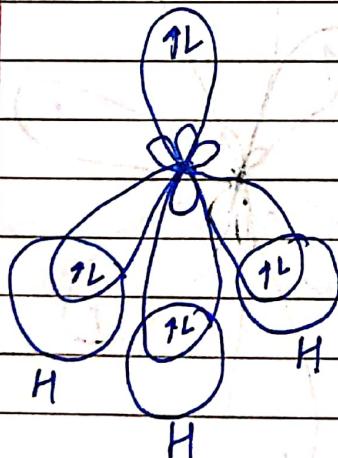
$$N = 1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$$



$\text{Sp}^3$  hybridisation



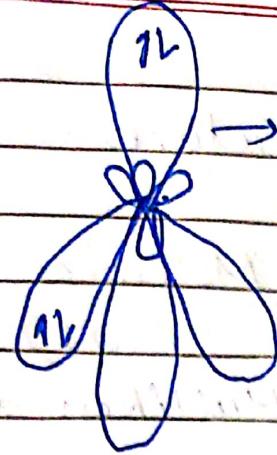
→ Four  $\text{sp}^3$  hybridised orbitals of Nitrogen.



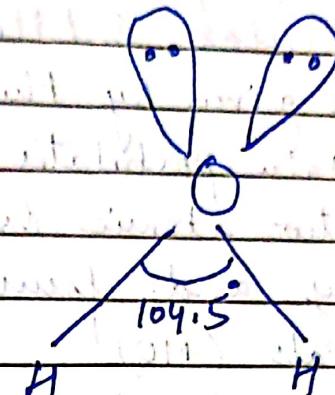
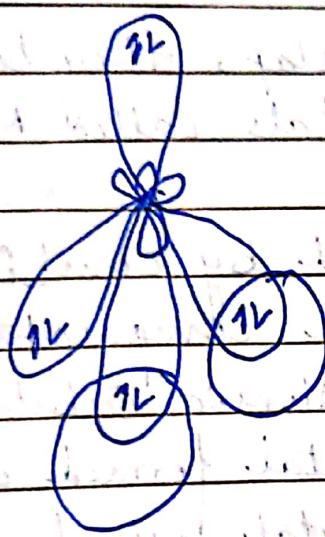
In  $\text{NH}_3$ , the bond angle is reduced from  $109^\circ 28'$  to  $107^\circ$  because lone-pair - bond pair repulsion is more than the bond-pair - bond pair repulsion. Therefore shape becomes pyramidal.

$\text{H}_2\text{O}$  - Water, Hybridisation is  $\text{sp}^3$

$$\text{O} = 1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$$



→ Four  $sp^3$  hybridised  
orbitals of Oxygen



In  $H_2O$ , the bond angle is reduced from  $109^\circ 28'$  to  $104.5^\circ$   
because  $dp-dp > lp-lp > bp-bp$   
and hence shape is inverted V-shape