

9. Coordination Compounds

Complex compounds in which the metal atoms are bound to a no. of anions or neutral molecules by sharing of electrons. In modern terminology such compounds are called coordination compounds.

Werner's theory of coordination compound :-

Alfred Werner (1866-1919), a Swiss chemist was the first to formulate his ideas about the structures of coordinate compounds.

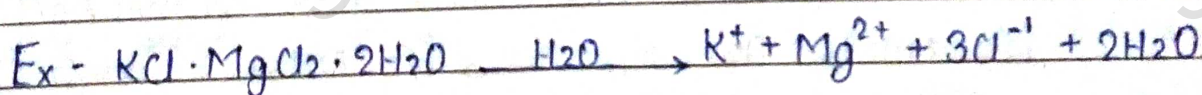
Addition compounds

Compounds which are formed by combination of two or more simpler compounds that can be packed in a definite ratio into a crystal.

Types of addition compounds

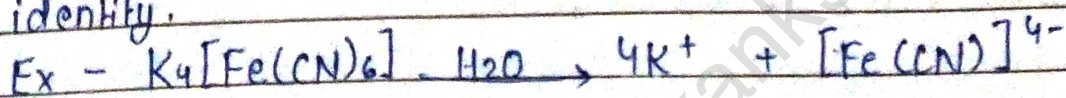
(i) Double salts

These are salts (addition) which exist only in solid state and dissociate into ions in aqueous solution completely and lose their identity.



(ii) Co-ordinate salts :

Addition salts which exist in solid state as well as when dissolved in water, they do not completely lose their identity.



Main postulates of Werner's theory

- (i) In coordination compounds metal show two types of linkages (valences) - primary and secondary.
- (ii) Primary valences are normally ionisable & satisfied by negative ions.
- (iii) The secondary valences are non ionisable & satisfied by neutral molecules or negative ions. Secondary valence is equal to coordination no and fixed for a metal.
- (iv) The ions / groups bound by secondary valence to the metal have characteristic spatial arrangements corresponding to different coordination numbers.

NOTE - In modern formulation, such spatial arrangements are called coordination polyhedra.

Primary valency - No of ion dissociable (shape not specific)

Secondary valency - No of non dissociable ion (shape specific)

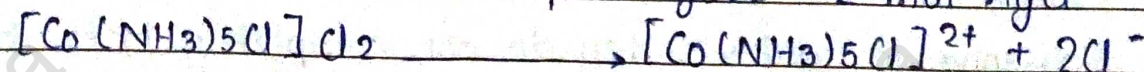
Werner's observation:

1 mol $\text{CoCl}_3 \cdot 6\text{NH}_3$ (Yellow) gave 3 mol AgCl



$$\therefore \text{P.V} = 3, \text{S.V} = 6$$

1 mol $\text{CoCl}_3 \cdot 5\text{NH}_3$ (Purple) gave 2 mol AgCl



$$\therefore \text{P.V} = 2, \text{S.V} = 6$$

1 mol $\text{CoCl}_3 \cdot 4\text{NH}_3$ (Green) gave 1 mol AgCl
 $[\text{Co}(\text{NH}_3)_4\text{Cl}_2]\text{Cl} \longrightarrow [\text{Co}(\text{NH}_3)_4\text{Cl}_2]^{+1} + \text{Cl}^{-1}$
 $\text{P.V} = 1$, $\text{S.V} = 6$

Some important terms :

Coordination entity / sphere - Constitutes of a central metal atom or ion bonded to a fixed no of ions or molecules.
Ex - $[\text{CoCl}_3(\text{NH}_3)_3]$, $[\text{Ni}(\text{CO})_4]$ etc.

Central atom / ion - The atom or ion to which a fixed no of ions or molecules are bonded are definite geometrical arrangement. Ex - In $[\text{Ni}(\text{CO})_4]$, Ni is central atom.

Coordination sphere - Central metal atom or ion and the molecules or ions bonded to it are enclosed in a square bracket and is collectively called coordination sphere []

Counter ion - The ionizable group written outside the brackets called counter ion. Ex - In $[\text{Co}(\text{NH}_3)_4\text{Cl}_2]\text{Cl}$ Cl^{-} is counter ion.

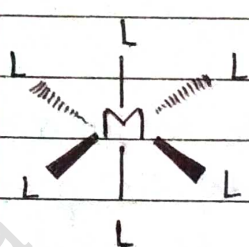
Ligand - The neutral molecules or ions bonded to the central atom or ion in the coordination entity are called ligand. $[\text{Ni}(\text{CO})_4]$, here $(\text{CO})_4$ is ligand

Coordination number (CN) - CN of a metal ion in a complex can be defined as the no of ligand donor atoms to which the metal is directly bonded. Ex - In $[\text{Ni}(\text{CO})_4]$ CN of Ni is 4.

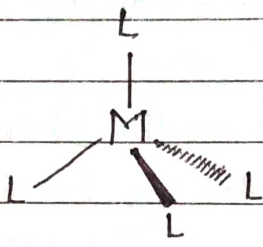
NOTE :- CN of central atom / ion determined only by no of sigma bonds formed by ligand with the central atom / ion.

Coordination polyhedron - The spatial arrangement of ligand atoms which are directly attached to the central atom / ion defines a coordination polyhedron about the central atom.

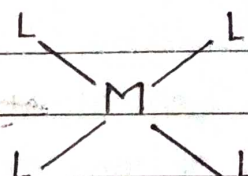
Some most common shape



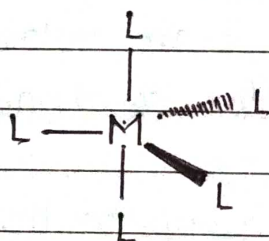
Octahedral



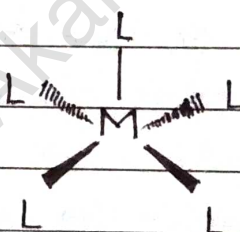
Tetrahedral



Square planar



Trigonal bipyramidal



Square pyramidal

Oxidation number of central atom - The oxidation number of central atom in a complex is defined as the charge it would carry if all the ligands are removed along with the electron pairs shared with the central atom.

Ex - In $[\text{Cu}(\text{CN})_4]^{3-}$, O.S. of Cu can be

$$x + 4(-1) = -3 \Rightarrow x = 1$$

And it is written as Cu(I)

Homopletic complexes - Complexes in which a metal is bound to only one kind of donor group. Ex - $[\text{Co}(\text{NH}_3)_6]^{3+}$

Heteropleptic complexes - Complexes in which a metal is bound to more than one kind of donor groups.
Ex - $[\text{Co}(\text{NH}_3)_4(\text{Cl}_2)]^+$

Classification of ligands

On the basis of charge :-

(i) Anionic ligands (neg charge)

F^- - Fluoro	SO_4^{2-} - Sulphato
Cl^- - Chloro	SCN^- - Thiocyanato
Br^- - Bromo	NCS^- - Isothiocyanato
I^- - Iodo	NO_2^- - Nitrito (N)
CN^- - Cyano	ONO^- - Nitrito (O)
OH^- - Hydroxo	N^{3-} - Nitrido
S^{2-} - Sulphido	NO_3^- - Nitrate

(ii) Neutral ligands (zero charge)

CO - Carbonyl	NH_3 - Ammine	H_2O - Aqua
CS (thiocarbonyl)	NO (Nitrosyl)	

(iii) Cationic ligands (Positive charge)

NO^+ (Nitrosonium)	NO_2^+ (Nitronium)	$\text{NH}_2 - \text{NH}_3^+$ - Hydrazinium
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On the basis of Denticity :-

Denticity - No of lone pair donated or no of coordinate bond formed.

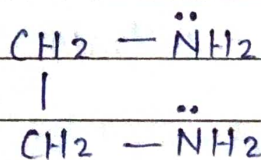
Monodentate or Unidentate - When a ligand is bound to a metal ion through a single donor atom the ligand is said to be unidentate.

Ex - F^- , Cl^- , Br^- , I^- , H^- , NO_2^- , $ONNO^-$, NO_3^- , OH^- , NH_2^- , SCN^- , NCS^- , S^{2-} , H_2O , NO^+ , SO , PH_3 , CO , CS , NO_2^+ , SO_4^{2-} , $NH_2^+-NH_3$

Some other :- C_5H_5N (Pyridine)
 CH_3NH_2 (Methylamine)
 CH_3COO^- (Acetate)

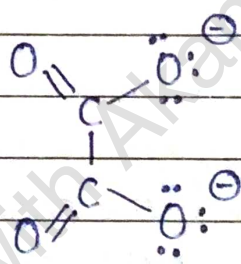
Didentate or Bidentate : When a ligand can bind through two donor atom, the ligand is said to be didentate.

Ex -



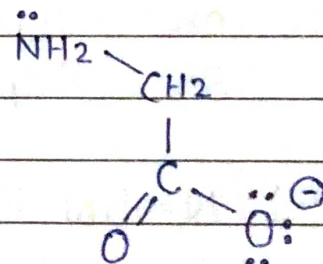
Ethylene diamine (en)

ON = 0



Oxalato (ox)

ON = -2



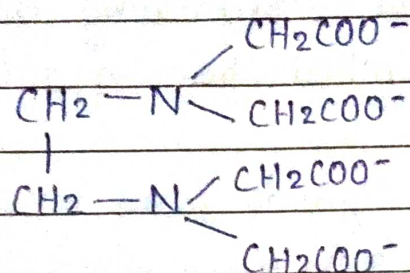
Glycinate (gly)

ON = -1

They can donate 2 lone pair at a time & form 2 coordinate bond.

Polydentate - When several donor atom are present in a single ligand & can donate more than 2 lone pair at a time, the ligand is said to be polydentate.

Ex -



Ethylenediamine tetraacetate (EDTA)

Oxidation No = -4

It is a hexa-dentate ligand

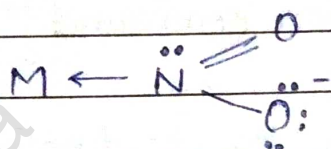
Ambidentate - Ligand which has two different donor atoms but either of two donates lone pair at a time.



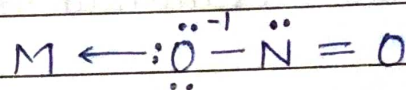
Thiocyanato - S



Thiocyanato - N



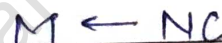
Nitrito - N



Nitrito - O



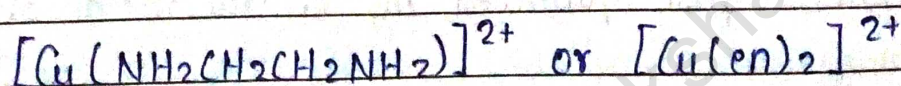
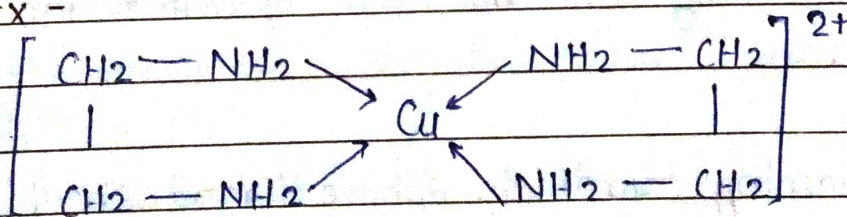
Cyano - C



Isocyano - N

Chelation - When a didentate or polydentate ligands uses its two or more donor atoms to bind to the same central metal atom or ion forming a ring structure is called chelation.

Ex -



Chelating agent - The ligand coordinating through two or more donor groups is called chelating agent or ligand.

Ex - (en)

Chelation effect - Chelate ligands form more stable complexes than similar ordinary complexes in which ligand act as monodentate. This is called chelation effect.

NOTE - Ligands are Lewis base
Central atom are Lewis acid.

Nomenclature of coordination compound

The formulas and names adopted for coordination entities are based on the recommendations of IUPAC (International Union of Pure and Applied Chemistry)

Rules :-

- 1) Positive part (Cationic part) should be named first, followed by negative part (anionic part).



Cationic part first named

- 2) We should only write the name of counter ion not their no.
- 3) While naming coordinate sphere ligand should be named first followed by central metal atom.
- 4) In case of heteroleptic complex ligand should be named according to their alphabetical order.

5) We should mention no. of ligands by adding prefix before the name of ligands.

di - 2 , tri - 3 , tetra - 4 , penta - 5 , hexa - 6

If there is (en) - ethylenediamine

(EDTA) - ethylenediaminetetraacetate

Then, instead of above prefixes, we would use:

bis - 2 , tris - 3 and tetrakis - 4

6) While naming anionic ligands should end with "o" or "ido".

Cl^- - chloro or chlorido , CN^- - cyano

While cationic ligands should be end with "ium".

NO^+ - Nitrosonium & NO_2^+ - Nitronium

While neutral ligands are named as such. Ex - NH_3 - ammine

7) Now central atom should be named follow by mentioning their oxidation no. in a bracket in roman numerals.

8) If the coordination sphere is anionic (-ve charged) then the name of central atom should be end with "ate".

Cr - chromium , Mn - Manganate , Fe - Ferrate , Co - Cobaltate , Ni - Nickelate , Cu - Cupperate , Pt - Platinatate , Pd - Palladate , Ag - Argentate , Au - Aurate , Hg - Murcurate , Al - Aluminate.

9) While naming only charge coordinate sphere i.e. $[\text{Mn}(\text{H}_2\text{O})_6]^{2+}$ etc, we write ion in the end.

Ex - $[\text{Ni}(\text{NH}_3)_6]\text{Cl}_2$ - Hexaamminenickel (II) chloride

$\text{K}_4[\text{Fe}(\text{CN})_6]$ - Potassiumhexacyanoferrate (II)

$\text{Ni}(\text{CO})_4$ - Tetracarbonylnickel (0)

$[\text{Mn}(\text{H}_2\text{O})_6]^{2+}$ - Hexaqua manganese (II) ion

$[\text{Co}(\text{en})_3]_2(\text{SO}_4)_3$ - Trisethylenediamminecobalt (III) sulphate

$[\text{Pt}(\text{NH}_3)_4][\text{PtCl}_4]$ - Tetraammineplatinum (II) tetrachloroplatinate (II)

$[\text{Pt}(\text{NH}_3)_2\text{Cl}(\text{NO}_2)]$ - Diamminechloronitrito-N platinum (II)

$[\text{Cr}(\text{H}_2\text{O})_4\text{Cl}_2]^+$ - Tetraquadiochlorochromium (III) ion

$\text{Hg}[\text{Co}(\text{SCN})_4]$ - Mercurytetra thiocyanatocobaltate (III)

Isomerism in Coordination Compound

Two or more compounds having same molecular formula and but different arrangement of atoms and hence differ in one or more physical or chemical properties are called isomers and the phenomenon is called isomerism.

Two types of isomerism are :

(i) Structural isomerism

(ii) Stereoisomerism

Structural isomer - This type of isomerism arises due to the difference in the structure of co-ordination compound. It is of 4 types :-

Linkage isomerism - The isomerism arises due to presence of an ambidentate ligands in a coordination compounds.
Ex - $[\text{Co}(\text{NH}_3)_5\text{NO}_2]\text{Cl}_2$ and $[\text{Co}(\text{NH}_3)_5\text{ONO}]\text{Cl}_2$
 $[\text{Mn}(\text{CO})_5\text{SCN}]$ and $[\text{Mn}(\text{CO})_5\text{NCS}]$

Coordination isomerism - Due to interchange of ligands between the cationic and anionic entities of different metal ions present in a complex.

Ex - $[\text{Co}(\text{NH}_3)_6][\text{Cr}(\text{CN})_6]$ and $[\text{Cr}(\text{NH}_3)_6][\text{Co}(\text{CN})_6]$

Ionisation isomerism - Compounds which give different ions in solution due to exchange of ions in coordination sphere & counter ions are called ionisation isomers.

Ex - $[\text{Co}(\text{NH}_3)_5\text{Br}]\text{SO}_4$ and $[\text{Co}(\text{NH}_3)_5\text{SO}_4]\text{Br}$
 $[\text{Pt}(\text{NH}_3)_4\text{Cl}_2]\text{Br}_2$ and $[\text{Pt}(\text{NH}_3)_4\text{Br}_2]\text{Cl}_2$

Solvate Isomerism or Hydrate Isomerism - This is a special kind of ionisation isomerism which arises due to differ in the no. of water molecules attached to the metal atom or ion as ligands.

Ex - $[\text{Cr}(\text{H}_2\text{O})_6]\text{Cl}_3$
 $[\text{Cr}(\text{H}_2\text{O})_5\text{Cl}]\text{Cl}_2 \cdot (\text{H}_2\text{O})$
 $[\text{Cr}(\text{H}_2\text{O})_4\text{Cl}_2]\text{Cl} \cdot 2(\text{H}_2\text{O})$

Stereoisomerism - Which have the same molecular formula and chemical bond but they have different spatial arrangement of ligands.

It is of 2 types :-

(i) Geometrical

(ii) Optical

This type of isomerism arises in heteroleptic complexes of coordination no 4 or 6 due to different possible geometric arrangements of the ligand.

In square planar

In octahedral

MA_2X_2

MA_4X_2

MA_3X_3

$MABX_3$

Cis

Trans

MA_2XY

Fac-isomer

Mer-isomer

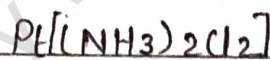
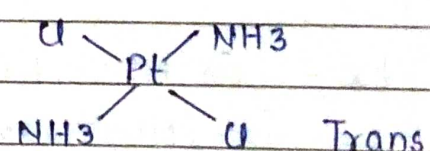
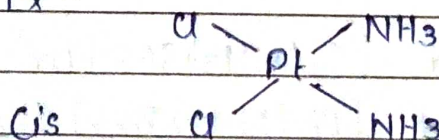
$MABXY$

Cis - Two identical ligands occupy adjacent position.

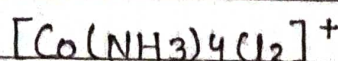
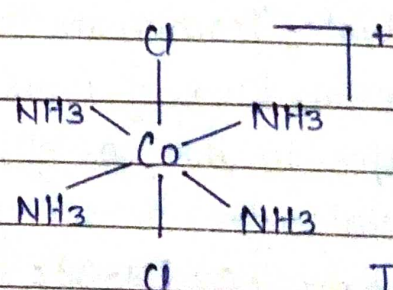
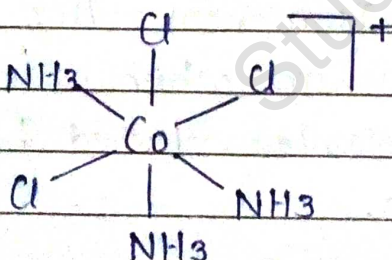
Trans - Two identical ligands occupy opposite position.

Ex -

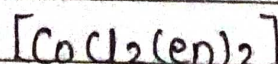
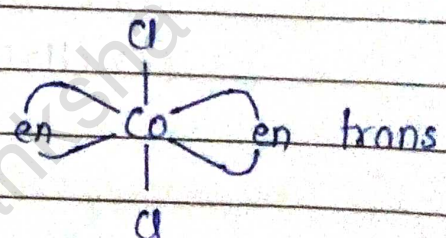
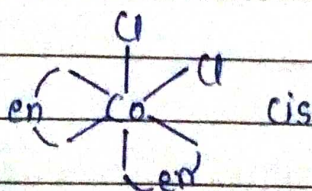
(i)



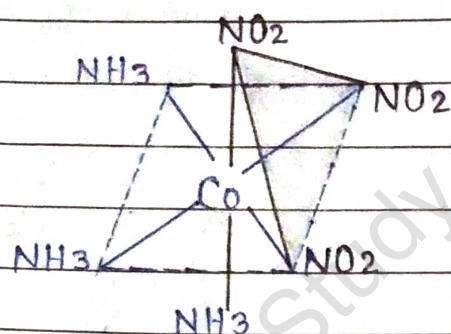
(ii)



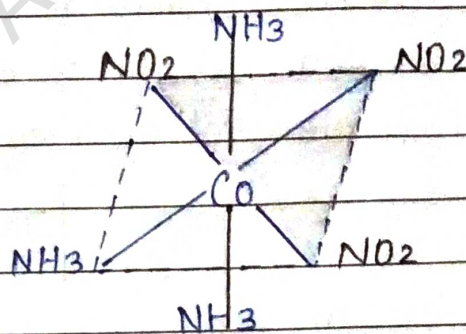
(iii)



Facial (fac) isomer - If three donor atoms of the same ligands occupy adjacent positions at the corners of an octahedral face, we have the facial (fac) isomer.



Fac - $[\text{Co}(\text{NO}_2)_3(\text{NH}_3)_3]$



Mer -

Meridional (mer) isomer - When the positions are around the meridian of the octahedron, we get the meridional (mer) isomer.

This is special type of geometrical isomerism occur in octahedral coordination entities of type $[\text{MA}_3\text{B}_3]$

Optical isomerism

The isomers which rotate the plane of polarised light (PPL) equally but in opposite direction are called optical isomerism or enantiomers or enantiomorphs.

Optical isomers are mirror images that cannot be superimposed on one another.

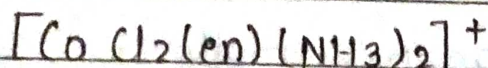
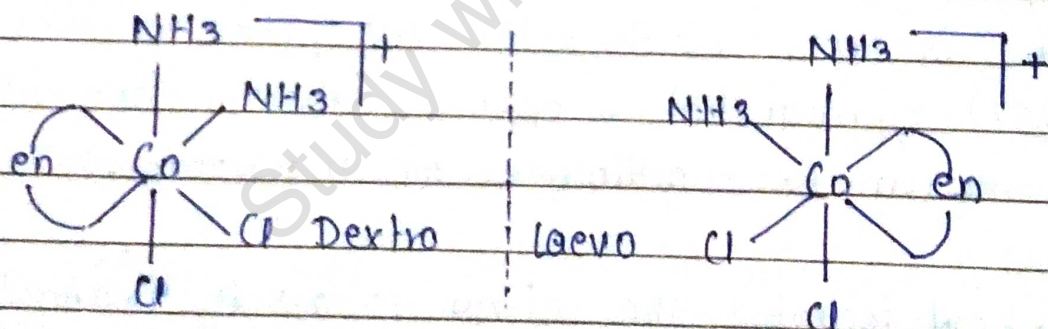
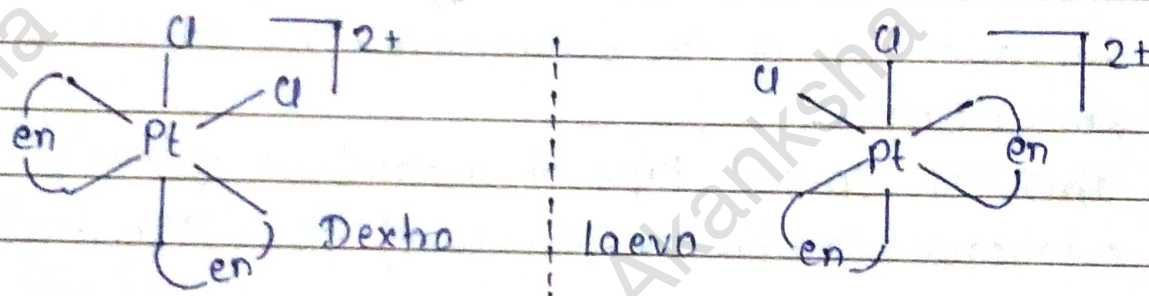
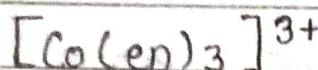
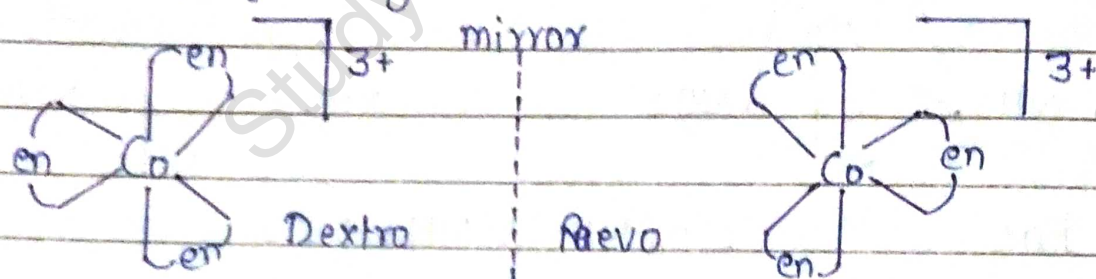
NOTE : The molecules or ions that cannot be superimposed are called chiral.

Dextro rotatory (d) - which rotate PPL towards Right.

Laevo rotatory (l) - which rotate PPL towards Left.

Racemic mixture - equilibrium mixture of d and l isomers gives a net zero rotation, called racemic mixture.

NOTE - Essential condition for a substance to show optical activity is that the substance should not have a plane of symmetry.



Bonding in coordination compounds

Warner's theory does not explain

- (i) why only certain elements possess the property of forming coordination compounds?
- (ii) Why bonds in coordination compound have directional properties?
- (iii) Why coordination compounds have characteristic magnetic and optical properties?

Valence bond theory (VBT)

- (i) According to this theory, the metal atom or ion under the influence of ligands can use its $(n-1)d$, ns , np , and nd orbitals for hybridisation to yield a set of equivalent orbitals of definite geometry such as octahedral, tetrahedral and square planar.
- (ii) These hybridised orbitals are allowed to overlap with ligand orbitals that can donate electron pairs for bonding.

Inner orbital complex

Involves inner d -orbitals
i.e. $(n-1)d$ -orbitals

Low spin complex

Have less or no unpaired electron.

Generally diamagnetic

Outer orbital complex

Involves outer d -orbitals

i.e. nd orbital for hybridisation

High spin complex

Have large no. of unpaired electron.

Generally Paramagnetic.

Coordination no	Type of hybridisation	Geometry
2	Sp	Linear
3	Sp ²	Trigonal planar
4	sp ³	Tetrahedral
4	dsp ²	Square planar
5	dsp ³	Trigonal Bipyramidal
5	sp ³ d	Square pyramidal
6	d ² sp ³	Octahedral
6	sp ³ d ²	Octahedral

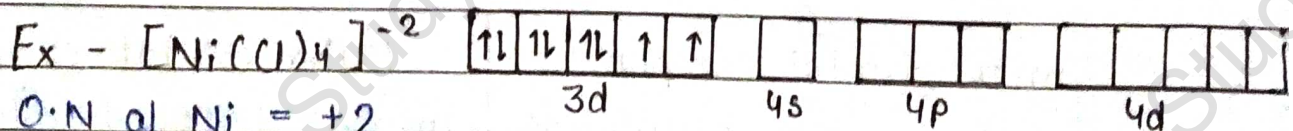
Strong ligands - such as CN^- , NH_3 & CO which force unpaired e^- to become pair.

Magnetic properties of coordination compound -

A coordination compound is paramagnetic in nature if it has unpaired electrons and diamagnetic if all electrons are paired.

Magnetic moment $\mu = \sqrt{n(n+2)}$ where n is no of unpaired electron.

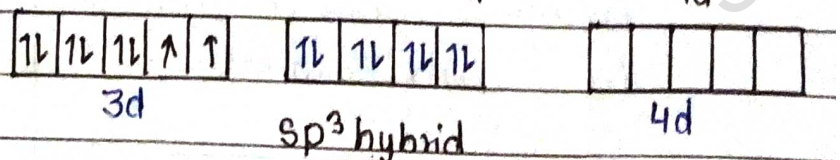
Unit is BM.



O.N of Ni = +2

Cl is weak ligand

$\text{Ni}^{2+} = [\text{Ar}]4s^03d^8$

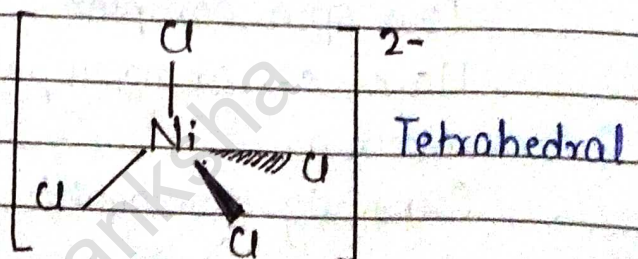


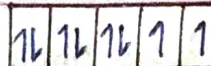
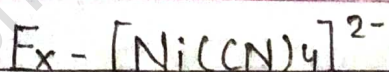
Hybridisation - sp³

Paramagnetic

High spin complex

Outer orbital complex





3d

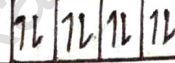
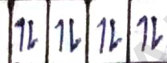
4s

4p

4d

O.N of Ni = +2

CN is strong ligand



3d

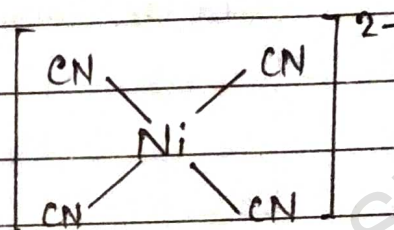
dsp² hybrid 4p

4d

Hybridisation - dsp²

Diamagnetic, inner orbital complex

Low spin complex



Square planar

TRICK :-

If $\mu = a.b$ (such as 5.9, 4.8 etc) then a is no of unpaired electron (ex 5 & 4 are no of unpaired e in above)

Limitations of VBT :

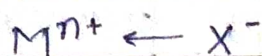
- (i) It involves a no of assumptions
- (ii) It does not give quantitative interpretation of magnetic data.
- (iii) It does not explain colour exhibited by coordination compound
- (iv) It does not give a quantitative interpretation of thermodynamic or kinetic stability of coordination compound.
- (v) It does not make exact predictions regarding the tetrahedral and square planar structure of 4 coordinate complexes.
- (vi) It does not distinguish between weak and strong ligands.

Crystal Field Theory (CFT)

Crystal field theory considers the metal ligands bond to be ionic arising from electrostatic interactions between the metal ion and the ligands are treated as point charge.

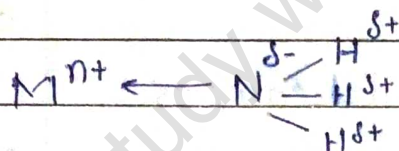
Case of anionic ligands - Such as Cl^- , Br^- , CN^-

The interaction with metal ions are ion-ion interaction.

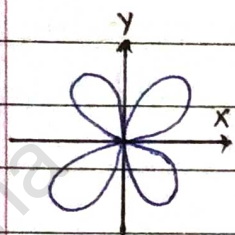


Case of neutral ligands - NH_3 , H_2O , CO etc

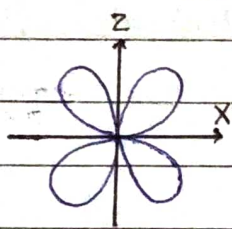
The interaction with metal ions are ion-dipole interaction.



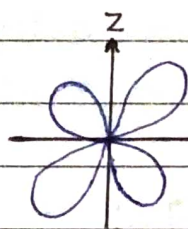
Now, in case of free metal ions, all the five d-orbitals have the same energy, called degenerate d orbitals.



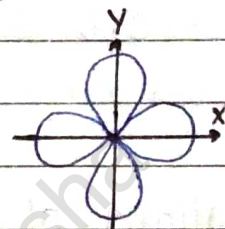
d_{xy}



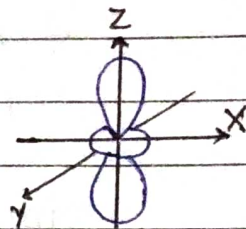
d_{xz}



d_{zy}



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



d_{z^2}

Non axial orbitals or
Triply degenerate orbitals, (t_{2g})

Axial orbitals or evenly
degenerated orbitals, (e_g)

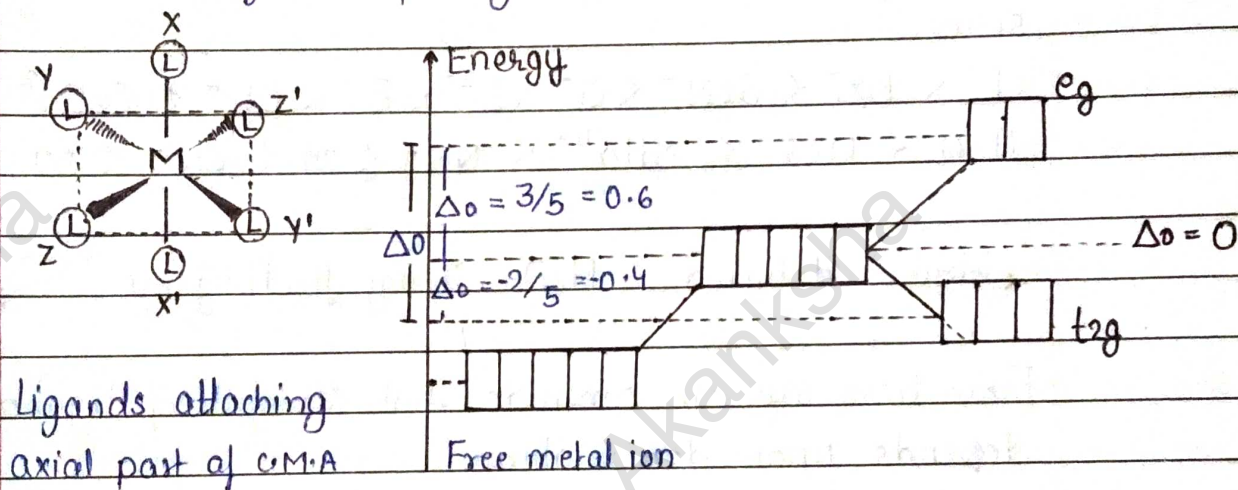
When ligands approach central metal atom, its d orbital loses its degeneracy, resulting in the splitting of d-orbitals (t_{2g} & e_g). However, splitting of d-orbitals depends on the nature of the crystal field.

Crystal field splitting energy - The energy difference between t_{2g} and e_g level is designated by Δ and is called crystal field splitting energy.

Splitting in octahedral compound

In an octahedral coordination entity coordination no is 6. 6 ligands approach central atom along the axis of dx^2-y^2 and dz^2 orbitals.

Thus degeneracy of d orbitals has been removed due to ligand electron-metal electron repulsion in octahedral complex to yield 3 orbitals of lower energy, t_{2g} set and 2 orbitals of higher energy of e_g set. because in octahedral ligands attack on axial part of CMA. Hence axial orbital energy becomes high on splitting.



Δ_0 = Crystal field splitting energy in octahedral field

$$\Delta_0 (\text{CFSE}) = [-0.4 \times (n_{t_{2g}}) + 0.6 \times (n_{e_g})] n = \text{no of } e^-$$

Splitting depends on type of ligand

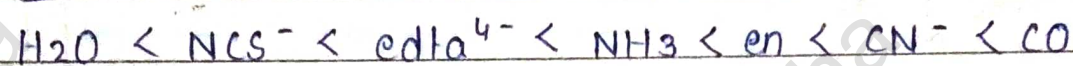
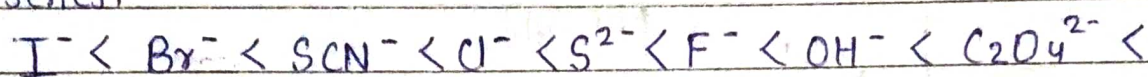
Strong ligand - Splitting more, CFSE (Δ) increases

Weak ligand - Splitting less, CFSE (Δ) decreases.

Pairing energy (P) - It represents the energy required for electron pairing in single orbital.

- (i) For weak field ligands $\Delta < P$ means electron prefer entering in higher energy orbital rather than pairing. It forms high spin complex.
- (ii) For strong field ligands $\Delta > P$ means electron prefer pairing rather than entering in higher energy level. It forms low spin complex.

Spectrochemical series - The arrangement of ligands in order of their CFSE values is known as spectrochemical series.



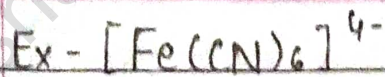
← Weak field ligand H_2O Strong field ligand →

From here we can conclude that strength of ligand depends upon donor atom.

C-donor > N-donor > O-donor > Halogen-donor

? Define crystal field splitting.

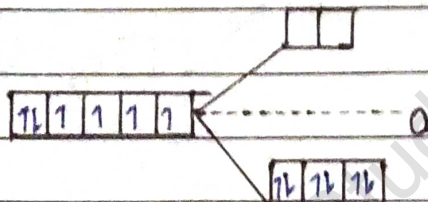
The conversion of five degenerated d-orbitals of the metal ion into different sets of orbitals having different energies in the presence of electrical field of ligand is called crystal field splitting.



O.N of Fe = +2



CN = Strong ligand

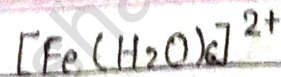


Colour - Colourless

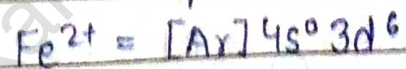
Diamagnetic, $\mu = 0 \text{ BM}$

$$\Delta_o = 0 \times 0.6 + 6 \times (-0.4)$$

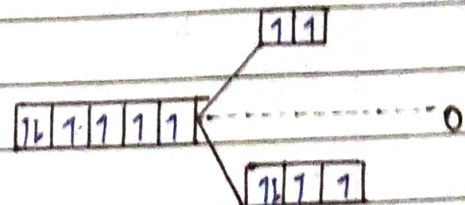
$$= -2.4$$



O.N of Fe = +2



H_2O = Weak ligand



Coloured

Paramagnetic, $\mu = \sqrt{24} \text{ BM}$

$$\Delta_o = 2 \times 0.6 + 4 \times (-0.4)$$

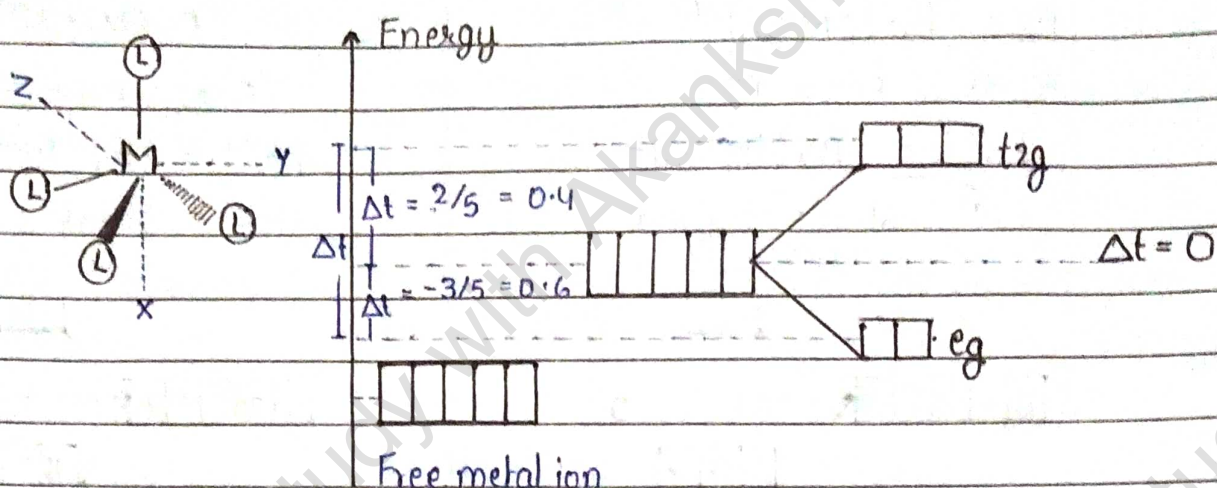
$$= 1.2 - 1.6 = -0.4$$

Crystal field splitting in tetrahedral coordination entities

In a tetrahedral coordination entity coordination no is 4. 4 ligands approach central atom from non axial position.

Thus degeneracy of d orbitals has been removed by ligand electron-metal electron repulsion to yield 3 orbitals (non axial) of high energy, t_2g and 2 orbitals of lower energy of e_g set, as ligands attack from non axial part on CMA. Hence non axial orbital (t_2g) energy becomes high on splitting.

In tetrahedral complexes, d-orbital splitting is inverted and smaller as compared to octahedral field splitting.



$$\Delta t (\text{CFSE}) = [+0.4 \times (n_{t2g}) - 0.6 \times (n_{eg})] n - \text{no of } e^-$$

For same metal, same ligand & metal-ligand distances

$$\Delta t = \frac{4}{9} \Delta_0$$

∴ Tetrahedral complexes lack symmetry, 'g' subscript is not used with energy levels.

CFSE $\ll$ P, hence pairing of e^- are not possible so these are high spin complexes. Low spin configuration are rarely observed.

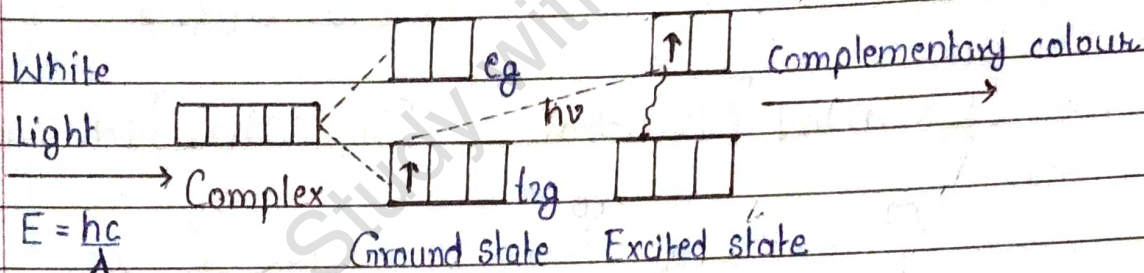
Colour in coordination compounds

When light of certain frequency falls on the complex, it absorbs light from visible range for transition of electrons from lower d-energy level to higher d energy level. This transition is called d-d transition of electron.

But in higher energy level electrons are unstable. So, they come back to a lower energy level, it releases energy which give the complementary colour of absorbed light to the compound.

In absence of ligand crystal field splitting does not occur and hence the substance is colourless.

Also, when d-orbital don't have any electron, crystal field splitting will not occur.



Ex - Anhydrous CuSO_4 is white but

$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ is blue in colour.

Removal of water from $[\text{Ti}(\text{H}_2\text{O})_6]\text{Cl}_3$ on heating renders it colourless.

Limitations of crystal Field theory :-

From the assumptions that the ligands are point charges it follows that anionic ligands should exert the greatest splitting effect. The anionic ligands actually are found at the low end of the spectrochemical series.

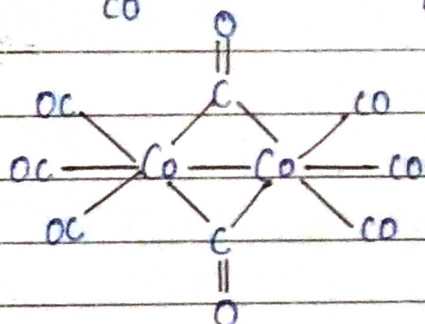
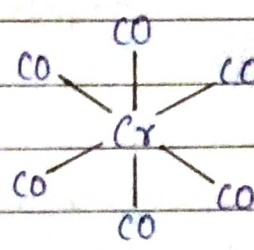
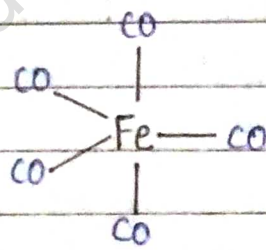
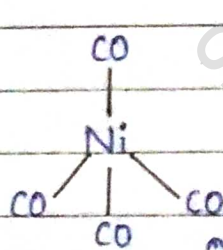
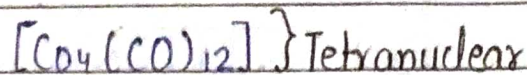
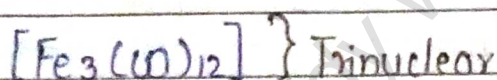
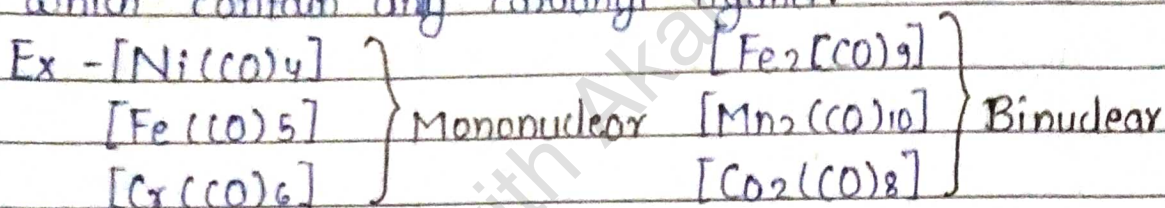
Further, it does not take into account the covalent character of bonding between the ligand and central atom.

Bonding in metal Carbonyls

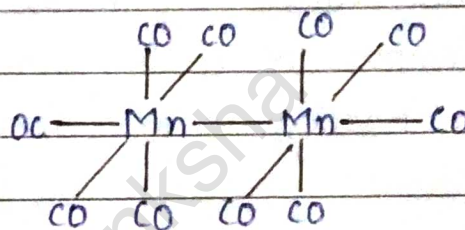
Compounds which have atleast one metal - carbon bond are called organometallic compounds.

Metal carbonyls are organometallic compounds in which carbonmonoxide (CO) acts as a ligand.

Homoleptic carbonyls - The organometallic compound which contain only carbonyl ligand.



Octacarbonylcobalt(0)



Decacarbonylmanganese(0)

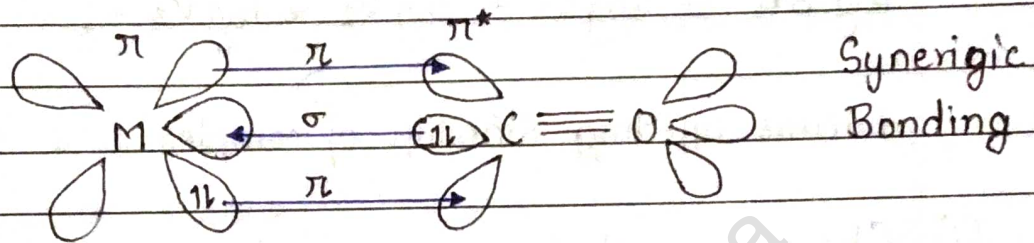
Metal carbon bonds in metal carbonyl have both σ and π character.

First overlap takes place between filled π_{2p} orbital of CO with an empty metal d-orbital resulting in σ bond formation.

The second overlap takes place between filled metal d-orbital with an empty antibonding π^*_{2p} orbital of CO resulting in additional π bond between the metal and same CO molecule. It is called back bonding

Synergic effect - The effect of σ bond formation strengthens the π bond and vice versa i.e. strengthens bond between CO and metal is called synergic effect (i.e. working together towards same goal)

Ligands such as CO, NO^+ , CN^- which are capable of accepting an appreciable amount of e^- density from the metal atom into empty π or π^* orbital of their own are called π acceptor or π acid ligands.



Stability of coordination compound

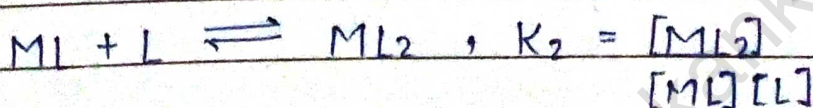
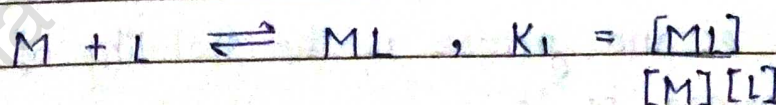
Most of the complexes are highly stable. The stability of a complex in solution means the degree of association between the metal and the ligands involved in the state of equilibrium.

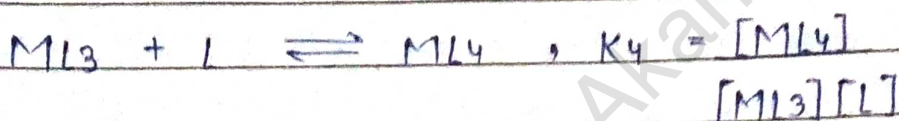
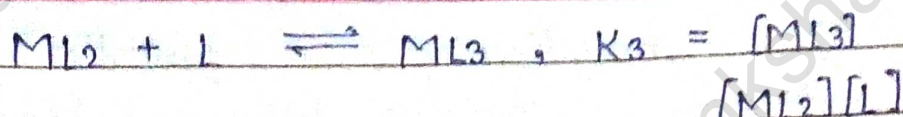
A coordination compound is formed by stepwise addition of ligands to a metal ion.

Let us consider a general reaction, $\text{M} + 4\text{L} \rightleftharpoons \text{ML}_4$

Ex - $\text{M} + 4\text{L} \rightleftharpoons \text{ML}_4$

Stability constant, $K_s = \frac{[\text{ML}_4]}{[\text{M}][\text{L}]^4}$ { K_s also denoted as β_4 }





K_1, K_2, K_3 and K_4 are stepwise stability constant & we can write $\beta_4 = K_1 \times K_2 \times K_3 \times K_4$

In general form - $\beta_n = K_1 \times K_2 \times K_3 \times \dots \times K_n$

or

$$\log \beta_n = \log K_1 + \log K_2 + \log K_3 + \dots + \log K_n$$

Factors affecting stability of complex

- (i) Charge on central metal ion - Greater the charge density on the central ion, greater is the stability of complexes.
- (ii) Size of metal ions - As the size of metal ion decreases stability of complex increases.
- (iii) Chelate effect - Chelate effect increases stability.
- (iv) Size & charge of ligand - For anionic ligand, higher the charge & smaller the size, the more stable complex is formed.
- (v) Basic strength of ligand - More basic a ligand, greater is the ease with which it can donate its lone pair of e^- and greater is the stability of complexes formed by it.