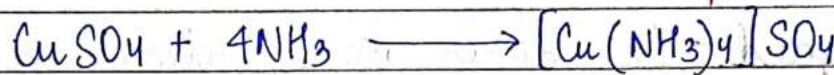
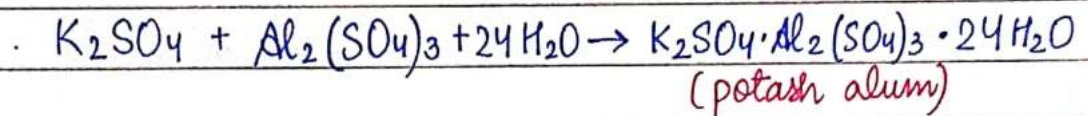
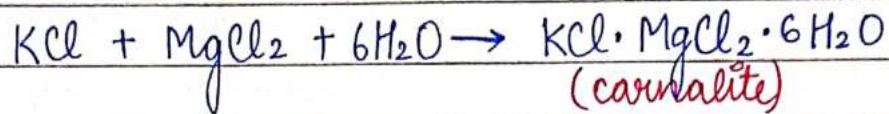


30 Aug '22

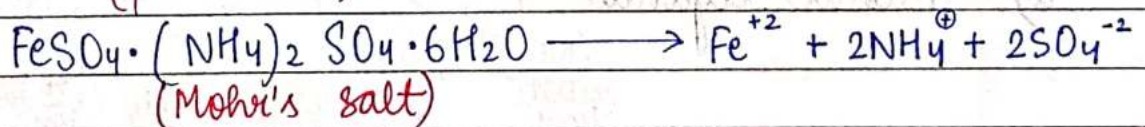
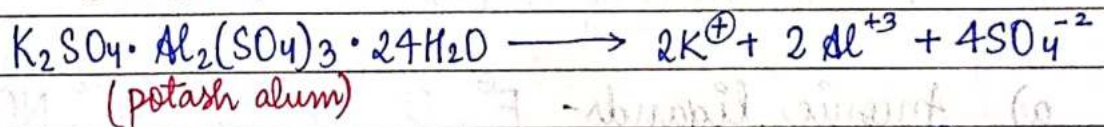
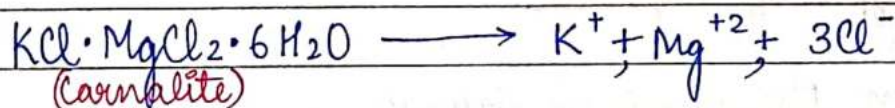
COORDINATION COMPOUNDS

Addition compounds - Those compounds which are made by mixing two or more stable compounds by crystallisation in a fixed ratio are called addition compounds.

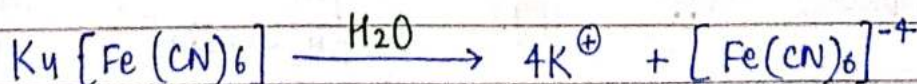
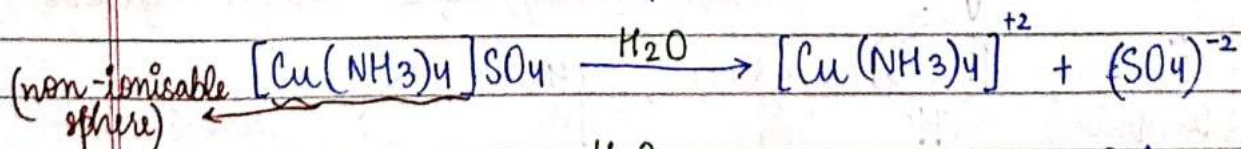


Addition compounds are of two types

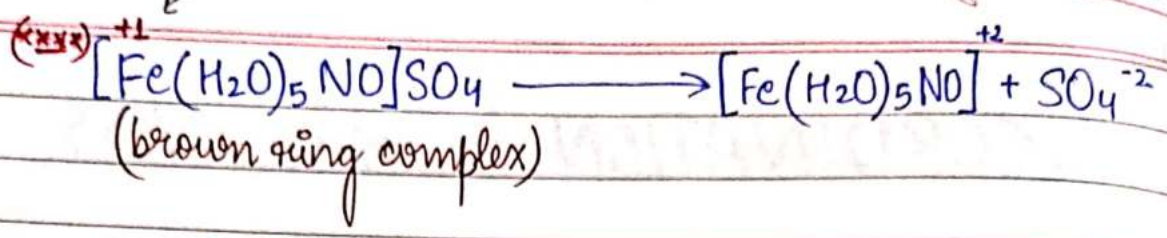
- 1) Double salt - Those compounds which lose their identity in solution are called double salts.



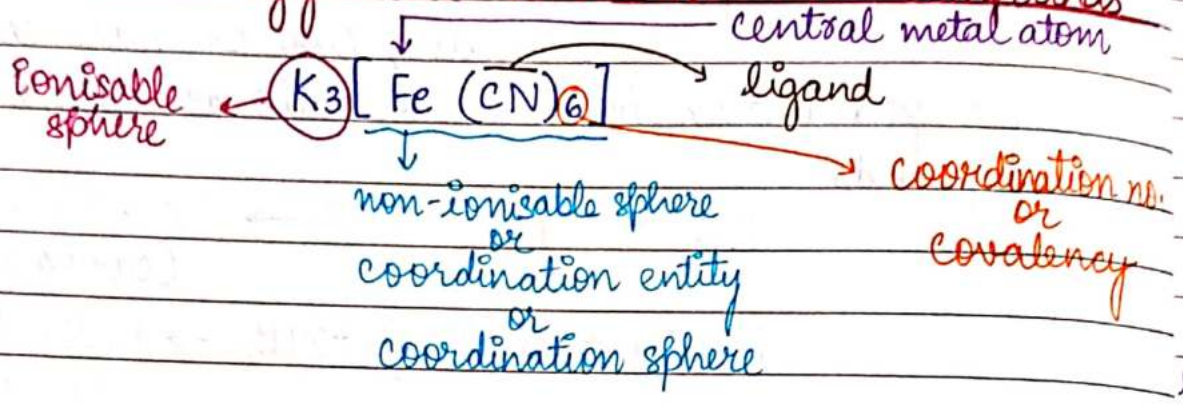
- 2) Complex or coordination compounds - Those compounds which retain their identity in solution are called complex or coordination compounds!



+1 oxidation state (very rare)



Some terminology used in coordination compounds



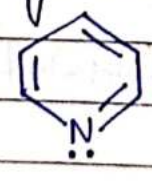
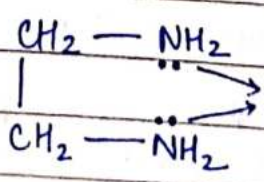
* **Ligand** - The donor atom, molecule or ion which donate its e⁻ pair to the central metal atom is called ligand.

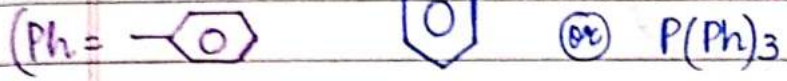
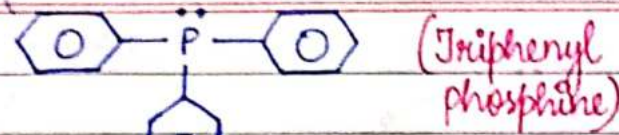
(**) Ligands act as lewis base.

Types of ligands

① On the basis of charge

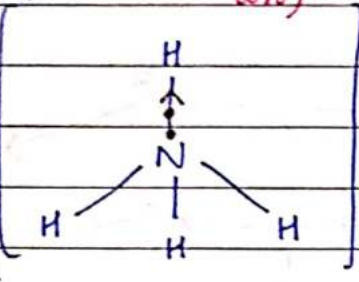
- a) **Anionic ligands** - $\text{F}^\ominus, \text{Cl}^\ominus, \text{Br}^\ominus, \text{I}^\ominus, \text{O}^{-2}, \text{NO}_3^{-1}, \text{S}^{-2},$
 $(\text{NO}_2)^{-1}$ (nitrite ion), $(\text{CN})^{-1}$ (cyanide), $(\text{NC})^{-1}$ (isocyanide), N^{-3} (nitride ion), $(\text{SCN})^{-1}$ (thiocyanate ion),
 $(\text{NCS})^{-1}$ (isothiocyanate ion), $(\text{N}_3)^{-1}$ (azide ion), $(\text{C}_2\text{O}_4)^{-2}$ (oxalate ion), $(\text{CO}_3)^{-2}$ (carbonate ion), $(\text{SO}_4)^{-2}$ (sulphate ion)

- b) **Neutral ligands** - (NH_3) (ammine), $:\text{CO}$ (carbonyl), $\text{H}_2\text{O}:$ (aqua), $:\text{NO}$ (nitrosyl), NH_2NH_2 (hydrazine),
 PH_3 (phosphine),  (pyridine or py),  (ethane-1,2-diamine or en)



c) Cationic ligands - NO_2^+ (nitronium ion), NO^+ (nitrosonium ion), $NH_2-NH_3^+$ (hydrazinium ion)

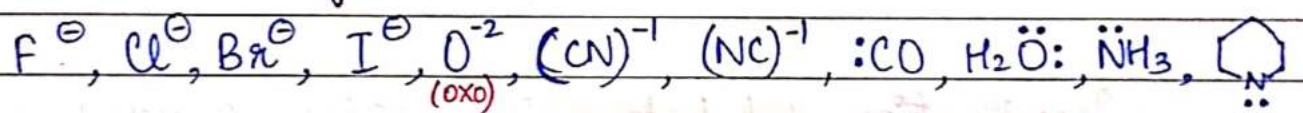
NH_4^+ or it cannot act as ligand because there are no e^- pair or lone pair to donate



(एक ligand एक साथ कितने coordinate bond बना सकता है वही उसकी denticity होती है)

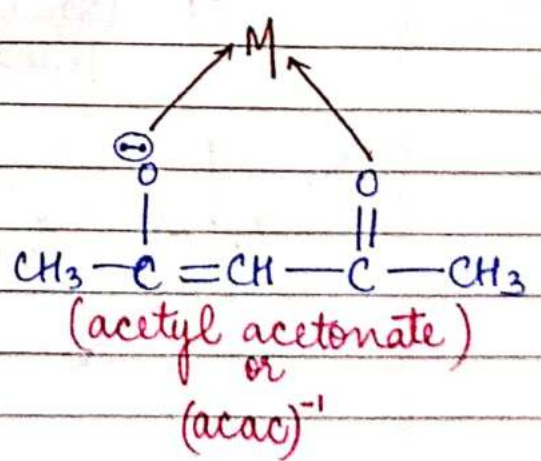
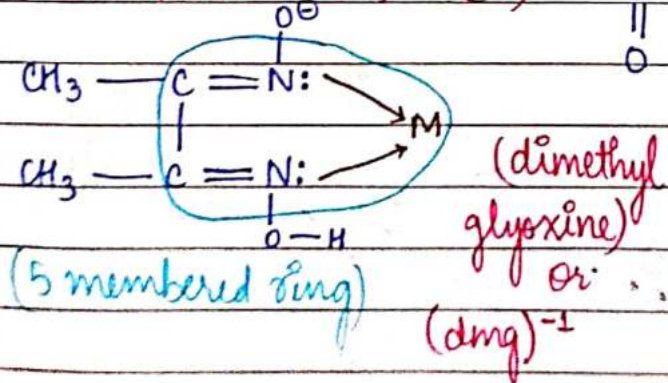
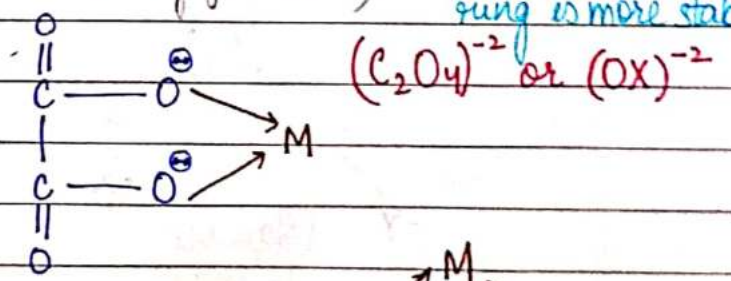
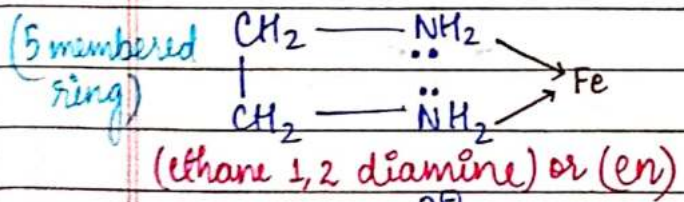
② Classification of ligand on the basis of denticity

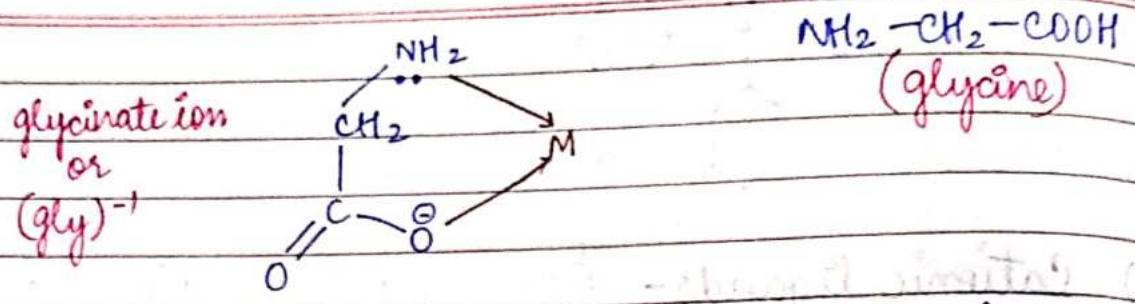
a) Monodentate ligand - एक बार में एक ही coordinate bond बनेगा



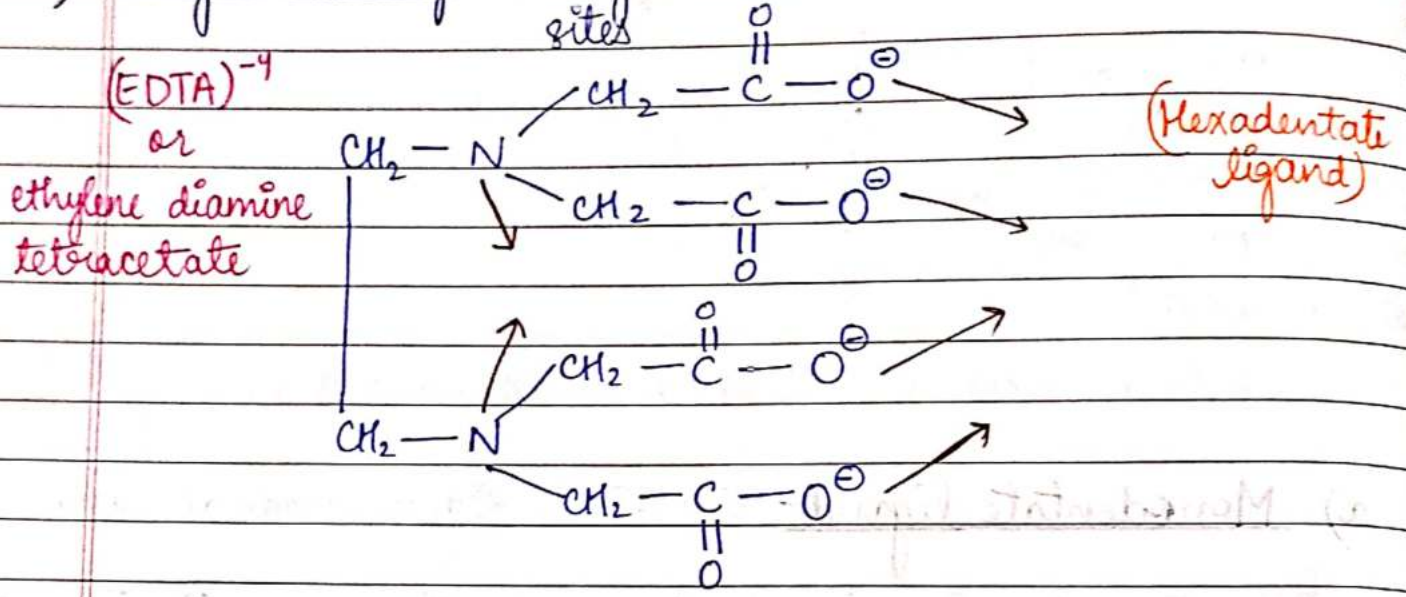
b) Bidentate ligand - ① One ligand forms two coordinate bond simultaneously.

② In case of bidentate ligand "chelation" occurs (5 membered or 6 membered ring formation) (5 membered or 6 membered ring is more stable)

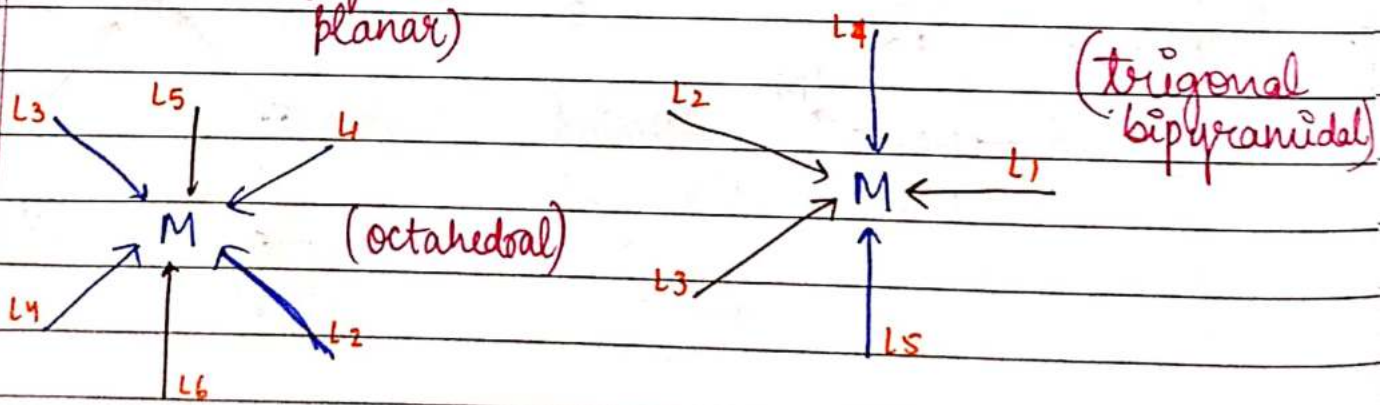
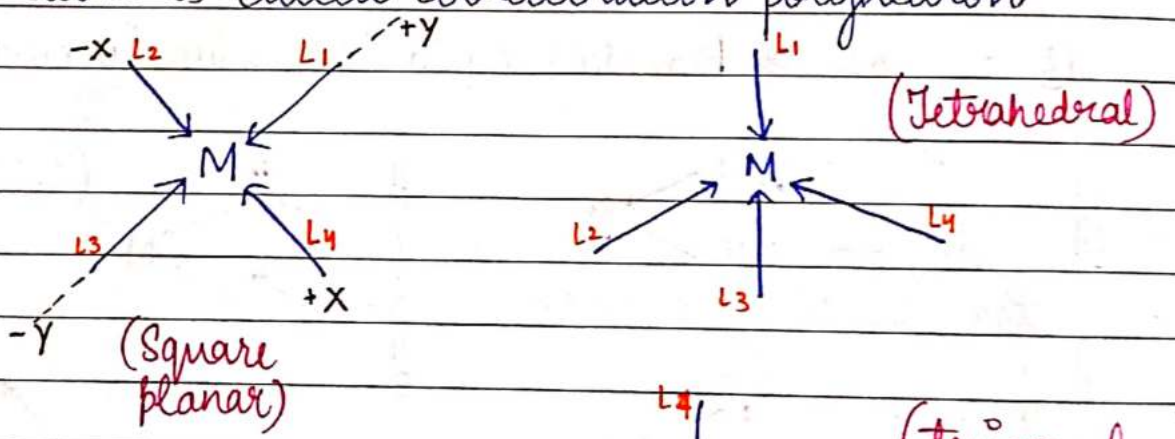




c) Polydentate ligand - Ligands which have several donor sites



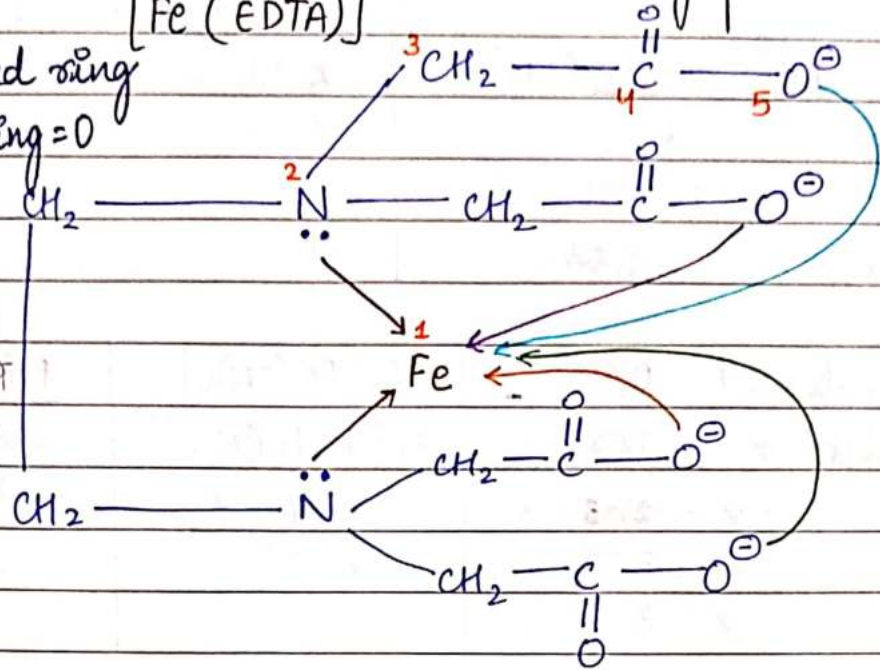
Coordination polyhedron - The spatial arrangement of ligands around the central metal atom is called coordination polyhedron.



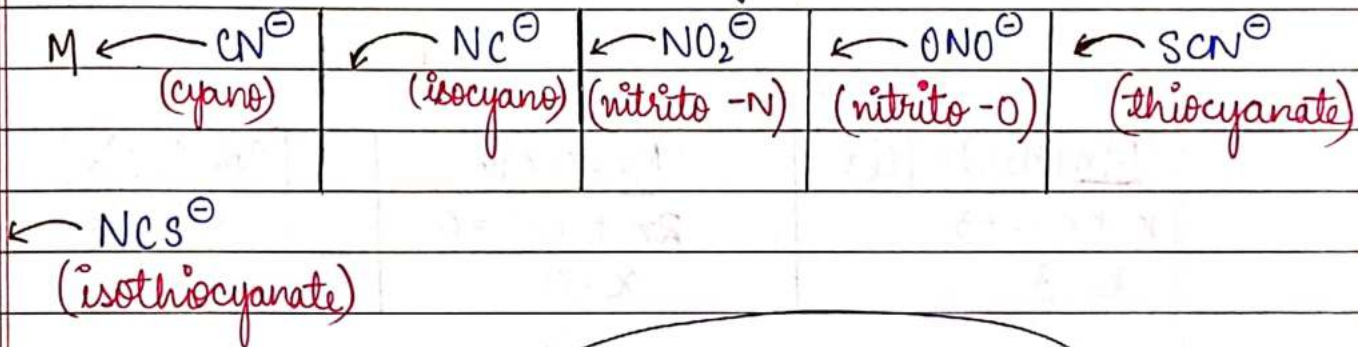
Q How many 5 member and 6 member ring possible in $[\text{Fe}(\text{EDTA})]^{-2}$

Ans "five" 5 membered ring
six membered ring = 0

Coordination no.
of Fe = 6 \rightarrow central
atom ने जितने
coordinate bond बनाए

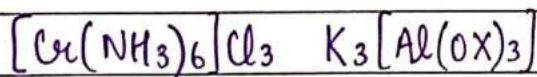
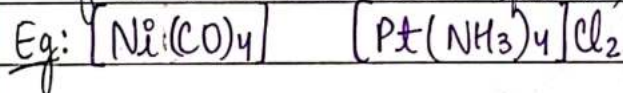


Ambidentate ligand - Those ligands which have two donor atom but only one can donate at a time -

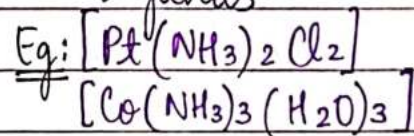


Homoleptic and heteroleptic complexes -

\hookrightarrow if in a coordination sphere, all ligands are same.



\hookrightarrow more than one type of ligands.

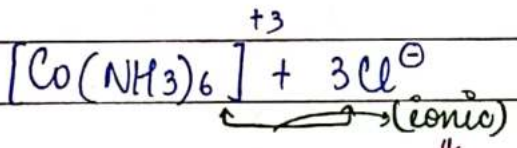
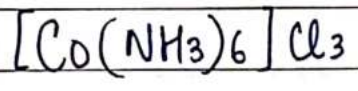


Calculation of charge on central metal atom

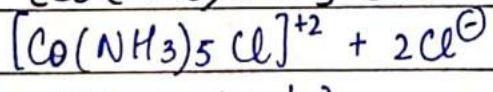
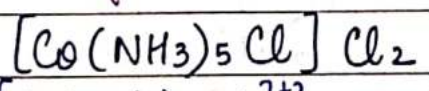
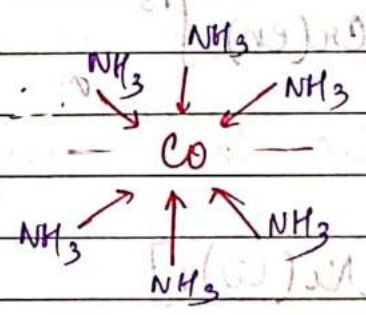
Werner's theory:-

According to the Werner's theory, coordination compounds has two types of valencies.

(1) Primary valency	(2) Secondary valency
→ Primary valency is equal to oxidation state of central metal atom.	→ Secondary valency is equal to coordination no.
→ Primary valencies are normally ionisable.	→ Secondary valencies are non ionisable.
→ Primary valency are satisfied by negative ion.	→ Secondary valencies are satisfied by negative, positive or neutral ligands.
→ Primary valencies are non directional.	→ Secondary valencies are directional, geometry of the compound is decided by secondary valency.



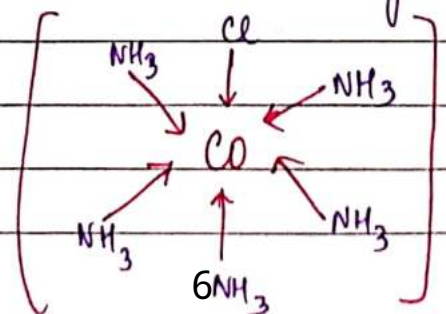
secondary valency = 6 (Ionic bond is non-directional)
 primary valency = 3



$x + 0 - 1 = +2$

$x = +3$

primary valency = 3
 secondary valency = 6



IUPAC Naming of Coordination Compounds

1) Cation is named first followed by the name of anion.

2) Name of anionic ligands ends with "o".

$\text{Cl}^\ominus = \text{chloro}$

$\text{C}_2\text{O}_4^{-2} = \text{oxalato}$

$\text{CN}^\ominus = \text{cyano}$

$(\text{CO}_3)^{-2} = \text{carbonato}$

$(\text{SO}_4)^{-2} = \text{sulphato}$

3) Ligands are named in alphabetical order.

4) If more than one ligand of similar type then prefix such as di, tri, tetra etc. are used. If ligand already contain a numerical prefix then bis, tris, tetrakis are used.

5) Name of some neutral ligands

$\text{NH}_3 = \text{amine}$

$\text{H}_2\text{O} = \text{aqua}$

$\text{CO} = \text{carbonyl}$

6) If coordination sphere has negative charge then the name of central metal atom ends with "ate".

$\text{Co} = \text{cobaltate}$

$\text{Zn} = \text{Zincate}$

$\text{Fe} = \text{Ferrate}$

$\text{Ag} = \text{Argentate}$

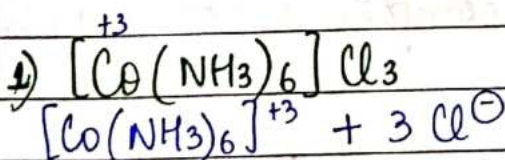
$\text{Cu} = \text{Cuprate}$

$\text{Au} = \text{Aurate}$

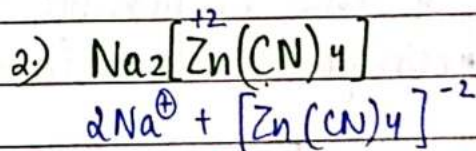
$\text{Pt} = \text{Platinate}$

$\text{Pb} = \text{Plumbate}$

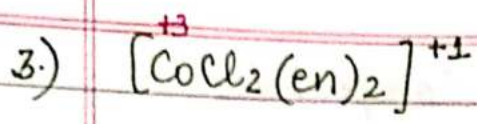
7) Oxidation state of central metal atom is written in roman numbers in paranthesis ()



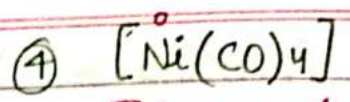
Hexammine Cobalt (III) chloride



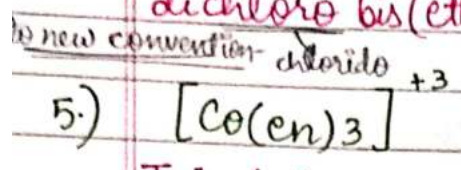
Sodium tetracyano zincate (II)



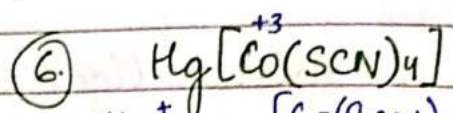
dichloro bis(ethane 1,2 diamine)
cobalt(III)



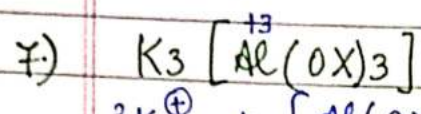
Tetracarbonyl nickel(0)



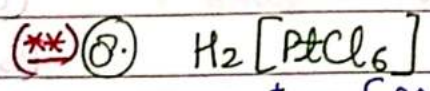
Tris(ethane 1,2 diamine)
cobalt(III)



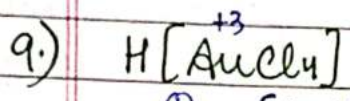
Mercury tetrathiocyanato Cobaltate(III)



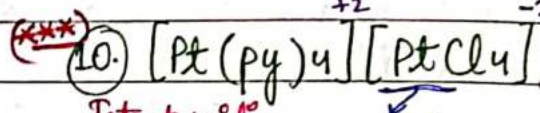
Potassium trioxalato Aluminate(III)



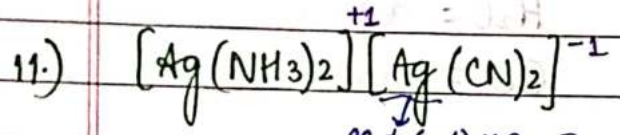
Hexachloro platinumic(IV) acid



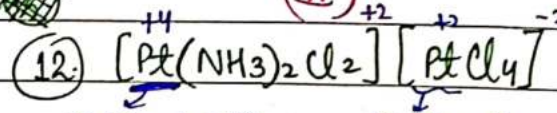
Tetrachloro Auric(III) acid



Tetrapyridine platinum(II) tetrachloro platinate(II)



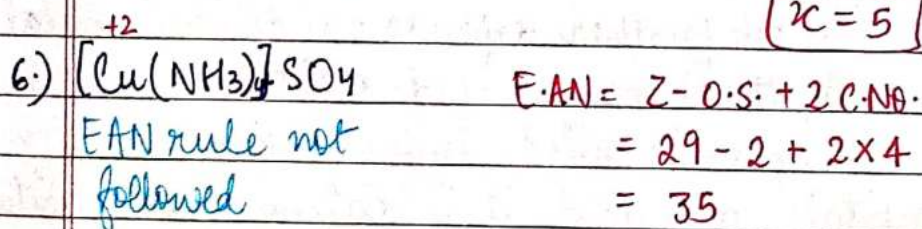
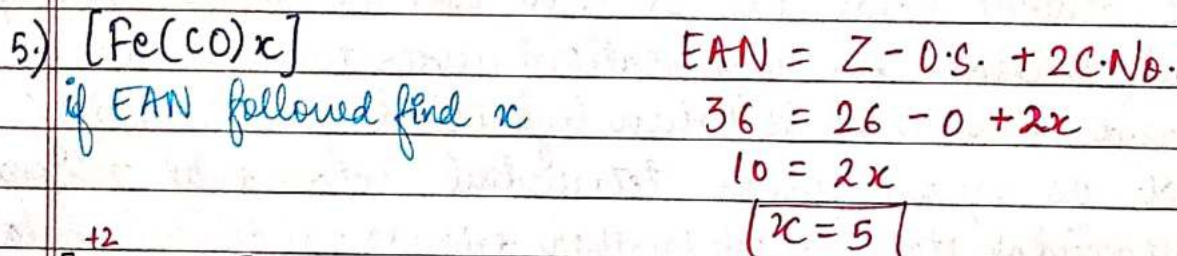
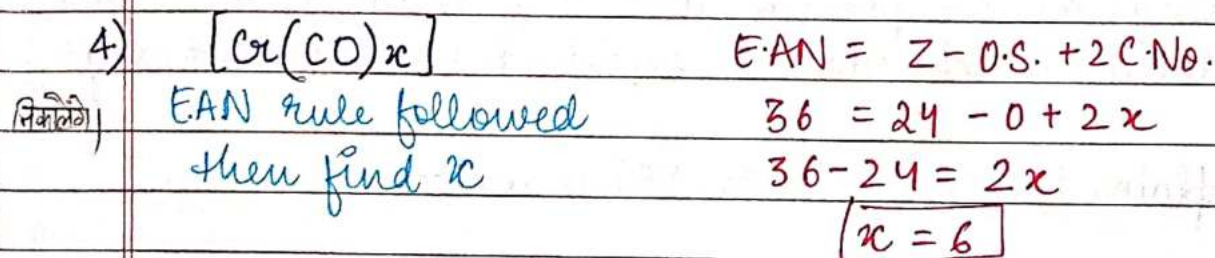
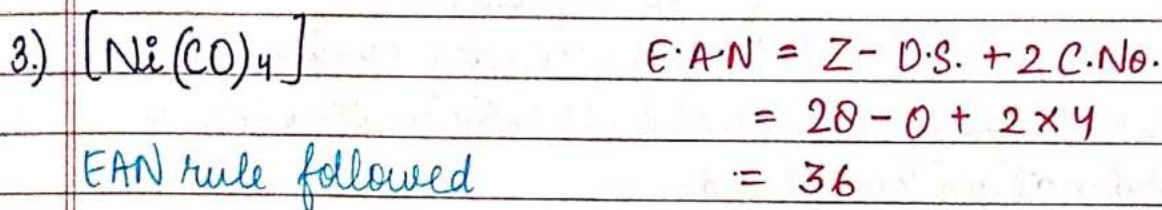
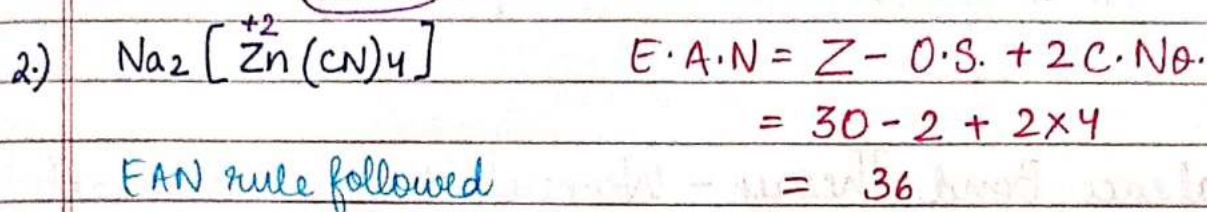
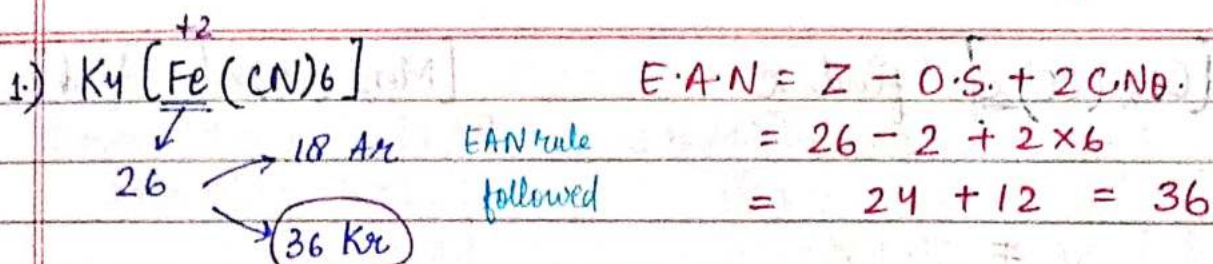
diammine silver(I) cyanargentate(I)



diamine dichloro platinum(IV) tetrachloro platinate(II)

EAN rule - According to this rule, EAN of central metal atom is equal to atomic no. of noble gas coming just after it. If this is possible then any complex or coordination compound considered stable.

$$E \cdot A \cdot N = Z - O.S + 2 C.No.$$

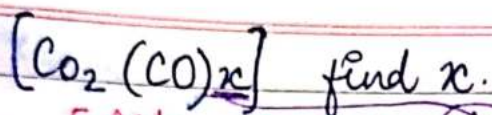


→ There exist many compounds where EAN rule is not followed but still they are stable means EAN rule can not decide stability.

(*) → In metal carbonyls, oxidation state of central metal atom is '0' and they generally follow EAN rule:

→ If in a metal carbonyl atomic no. of metal is odd, then

$$E.A.N = Z - O.S + C.No. + 1$$

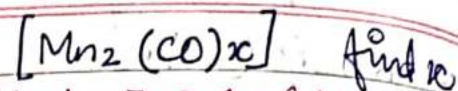


$$\text{EAN} = Z - O.S + (C.No.) + 1$$

$$36 = 27 - 0 + x + 1$$

$$36 = 28 + x$$

$$x = 8$$



$$\text{EAN} = Z - O.S + (C.No.) + 1$$

$$36 = 25 - 0 + x + 1$$

$$x = 10$$

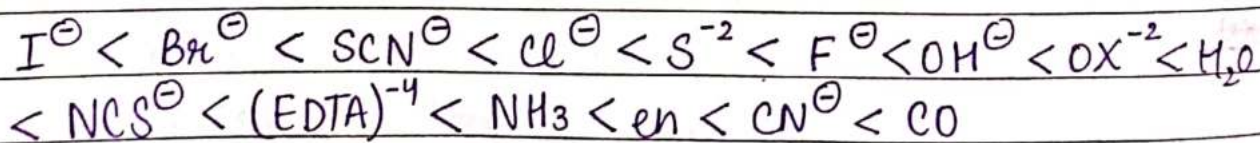
Valence Bond Theory - Werner's theory was unable to explain

- why only some elements form complex compounds.
- Werner's theory was unable to explain stability of coordination compounds.
- Coordination compounds also have magnetic and optical properties which can't be explained by Werner's theory.

→ To explain the above points VBT is used -

(**) → The central metal atom or ions uses its $(n-1)d$, ns , np & nd orbitals for hybridisation under the influence of ligand. These orbitals when hybridised gives geometry such as square planar, tetrahedral, octahedral, trigonal bipyramidal etc. which further depends on central metal atom and ligand. If the inner $(n-1)d$ orbitals are used ~~the~~ then the complex is called inner orbital complex. If outer nd orbitals are used it is called outer orbital complex.

Strength of ligand halogens - weak ligands



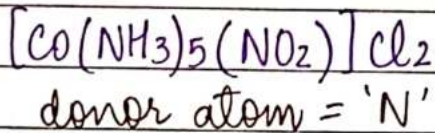
Isomerism1. Structural isomerism

- Linkage isomerism
- Coordination isomerism
- ionisation isomerism
- solvate or hydrate isomerism

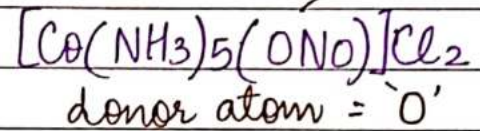
2. Stereoisomerism

- Geometrical isomerism
- Optical isomerism

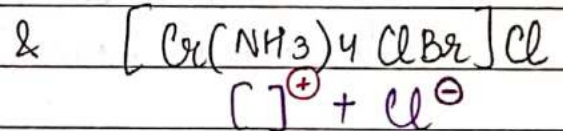
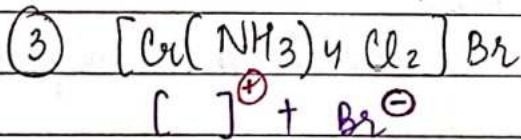
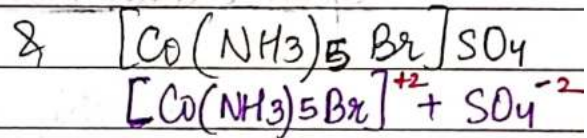
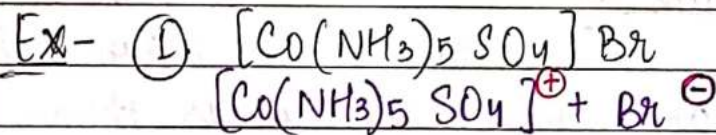
→ Linkage isomerism - This type of isomerism arises when a complex has ambidentate ligands. (o-nitro)



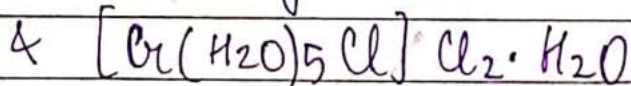
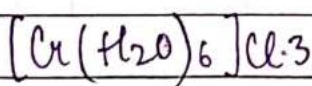
&



→ ~~isomerism~~ ionisation isomerism - This type of isomerism arises due to exchange of ligands with the coordination sphere.

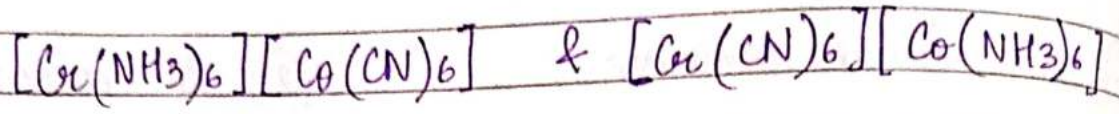


→ Solvate or hydrate isomerism - In this isomerism the isomers differ by No. of water molecules as ligands.



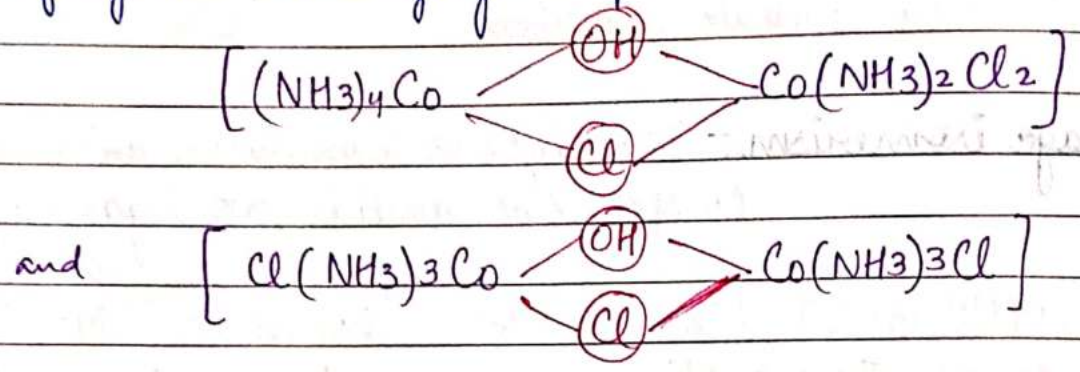
→ Coordination isomerism - This type of isomerism arises due to exchange of ligands with the coordination

Spheres. (cationic & anionic)



extra

Coordination-position isomerism - This isomerism arises due to different arrangement of ligands in bridging complexes.

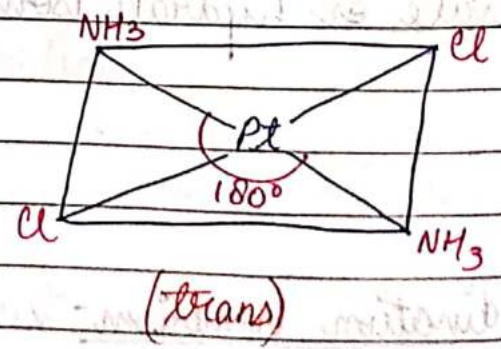
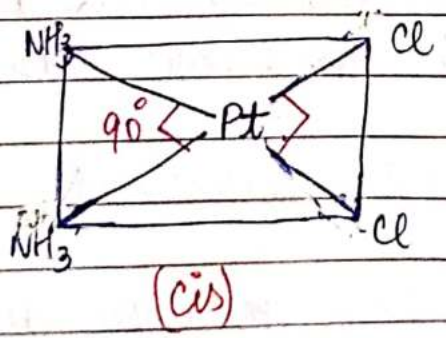


2.) Stereoisomerism - stereoisomers have same chemical formula and chemical bonds but different arrangement in space.

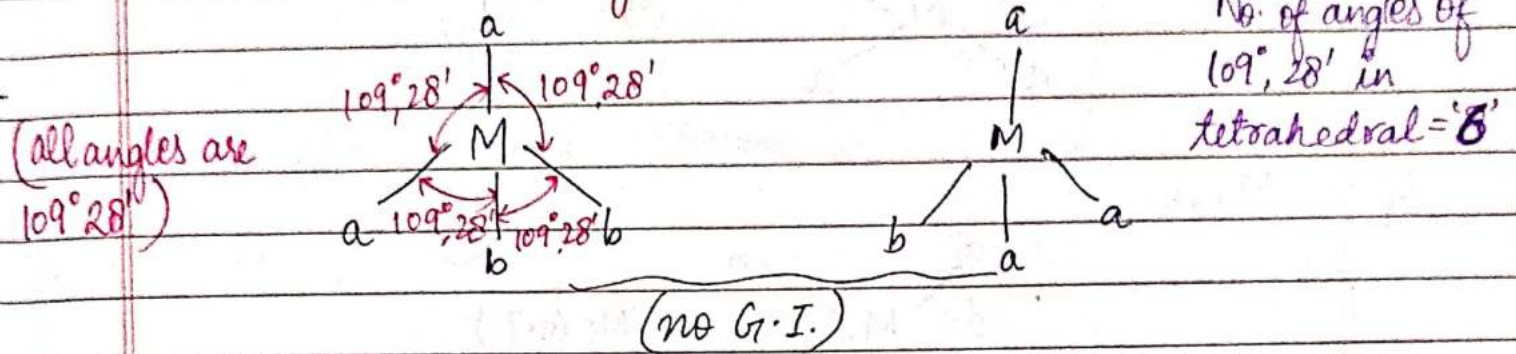
→ Geometrical isomerism - If due to exchange of ligands within the coordination sphere changes geometry or arrangement of ligands then geometrical isomerism arises.

(**) # Geometrical isomerism is only possible in heteroleptic complexes.

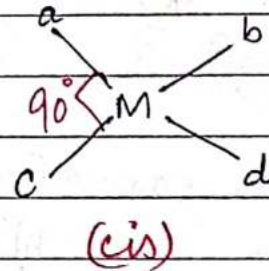
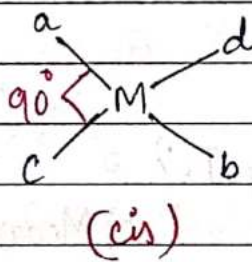
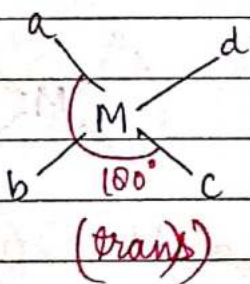
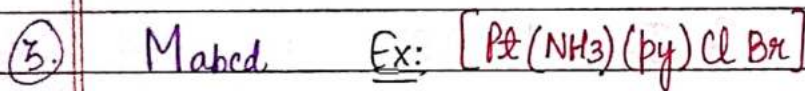
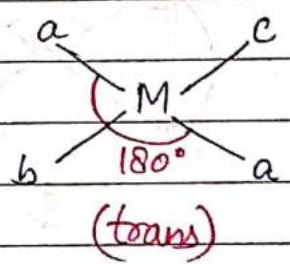
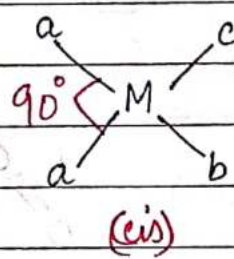
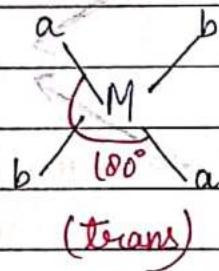
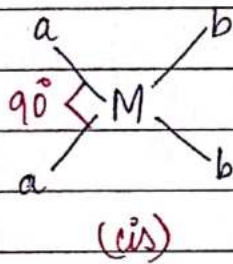
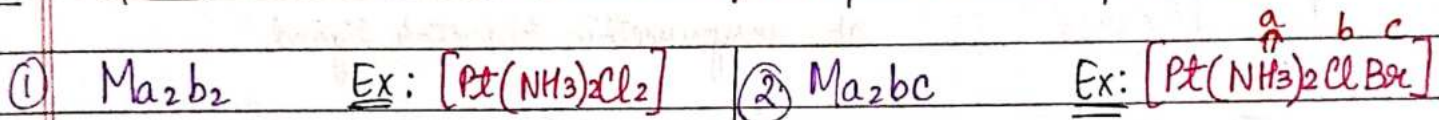
Ex: $[Pt(NH_3)_2Cl_2]$ Pt^2+ = square planar complex



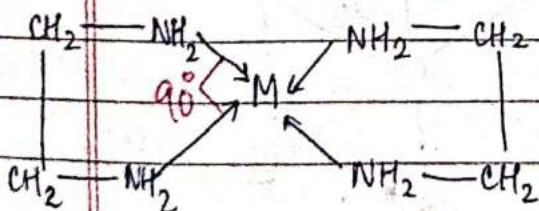
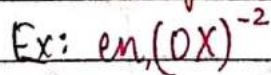
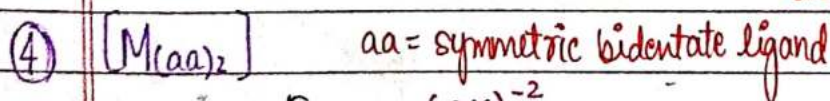
(***~~*)~~ Tetrahedral complex (sp^3) does not show geometrical isomerism because relative angles are same.



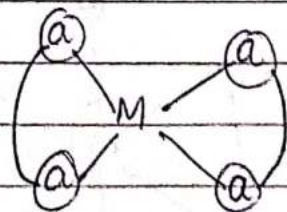
Case I Geometrical isomerism in square planar complexes -



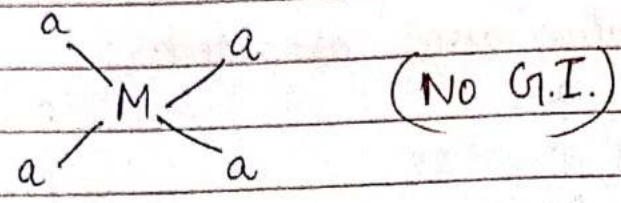
→ Total 3 geometrical isomer (2 cis & 1 trans with respect to any two ligands)



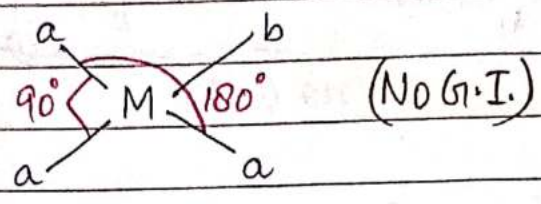
(No G.I.)



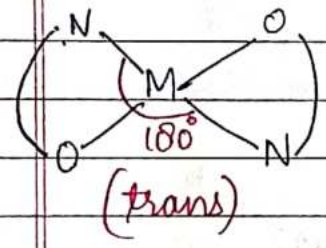
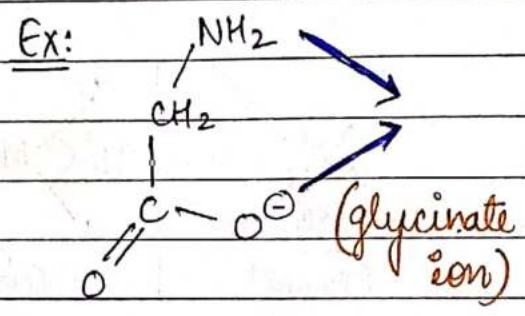
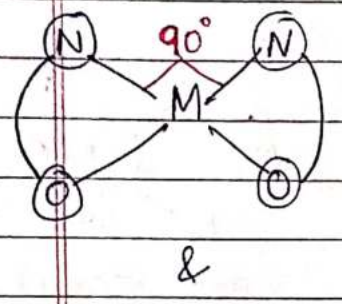
5. Ma_4



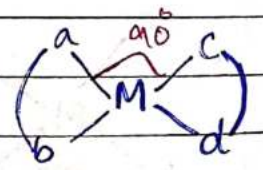
6. Ma_3b



7. $M(ab)_2$ $ab = \text{unsymmetric bidentate ligand}$



G.I. (✓)

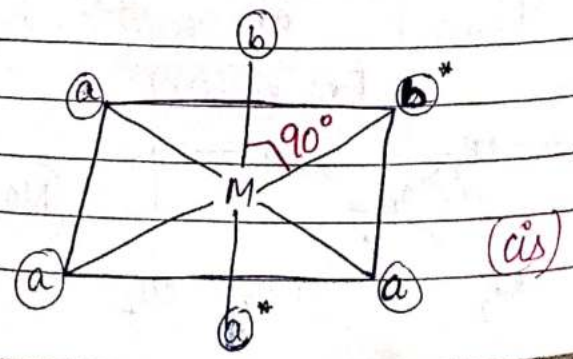
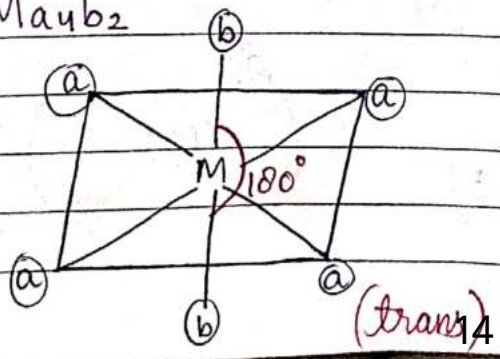


Q Which can show G.I.?

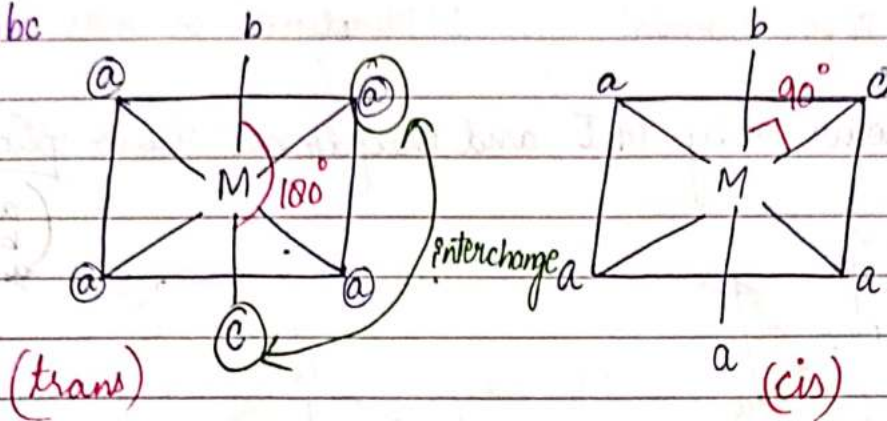
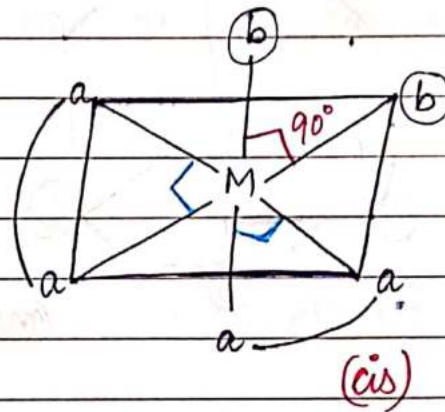
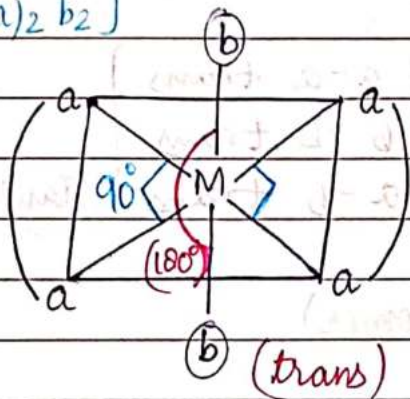
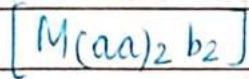
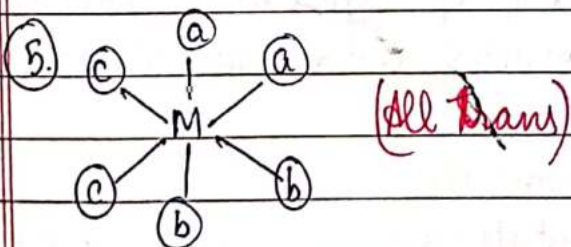
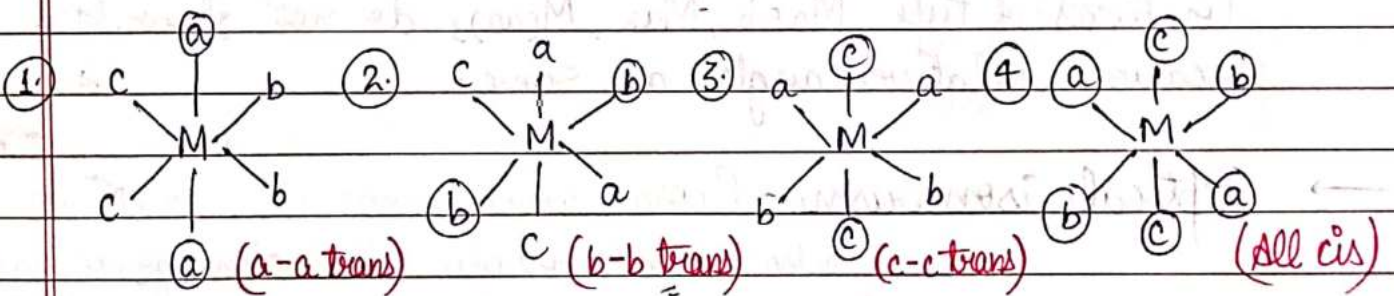
- a) Ma_4 b) Ma_3b c) $M(aa)_2$ ~~d) $M(ab)(cd)$~~

Case II Geometrical isomerism in Octahedral Complexes

1) Ma_4b_2



② Maybe

③ $[M(aa)_2 bc]$ & $[M(aa)_2 b_2]$ Ex: $[Co(en)_2 Cl Br]^+$ ④ $Ma_2b_2c_2$ (5 G.I.)

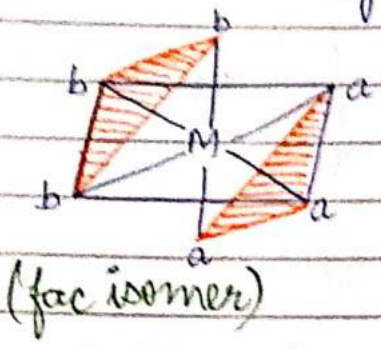
⑤

Ma_3b_3 → In this case cis and trans isomer are also formed. But it shows geometrical isomerism of two types.

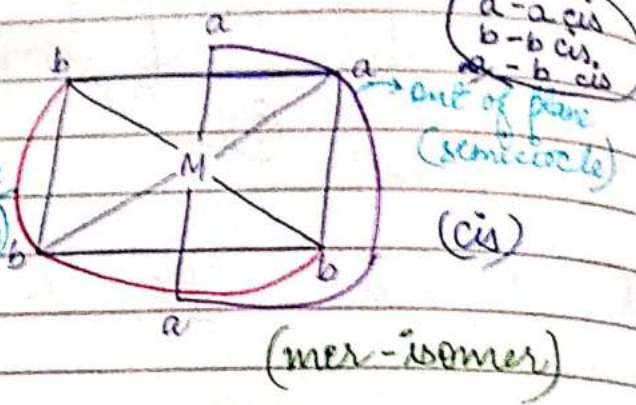
a) facial or fac isomer

b) Meridional or mer isomer

(XX) → Ma_3b_3 shows only G.I and they do not show optical isomerism

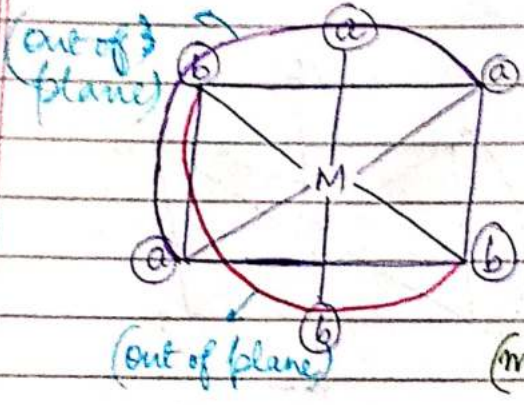


(in plane semicircle)



a-a cis
b-b cis
a-b cis

out of plane (semicircle)



a-a trans
b-b trans
a-b trans (trans)

(6) Ma_2b_2cd → 15 G.I.

Complexes of type Ma_2b_2 , Ma_2c , $M(ac)_3$ do not show G.I. because relative angles are same.

→ Optical isomerism - Coordination compounds which are non-superimposable mirror image of each other and do not have a plane of symmetry are called optical isomers or enantiomers and the phenomenon is called optical isomerism.

They are of (2) types → dextrorotatory (rotates right)
levorotatory (rotates left)

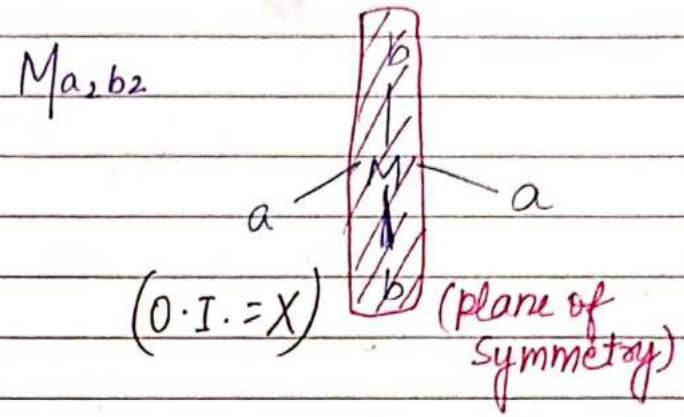
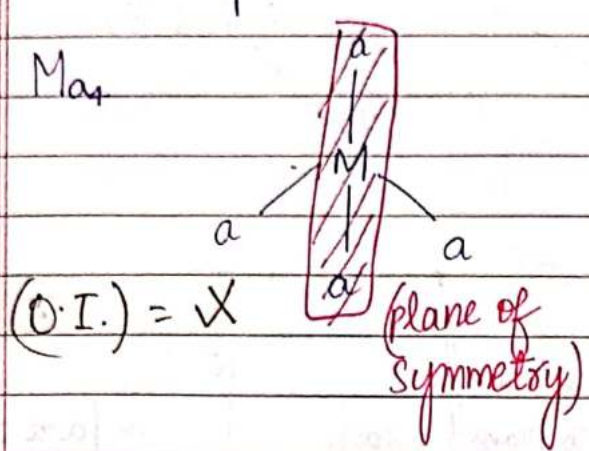
Optical isomers don't have plane of symmetry

For optical isomerism, compound must be non-planar. So, square planar complexes don't show optical isomerism.

(*) $[CrAs]$ is an intrinsic semiconductor as it is a compound

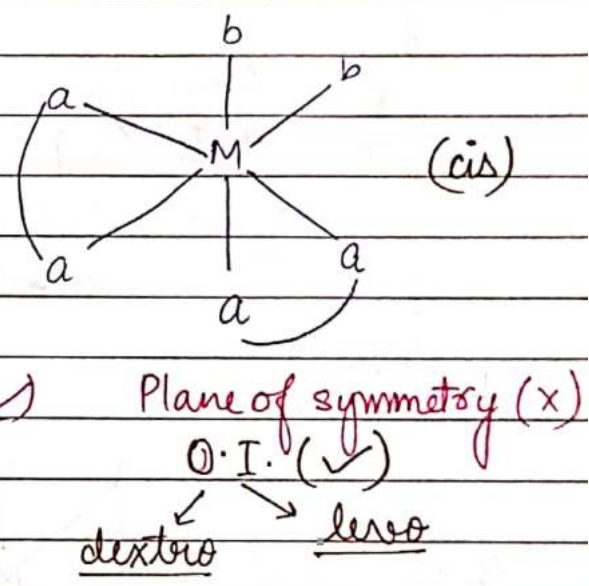
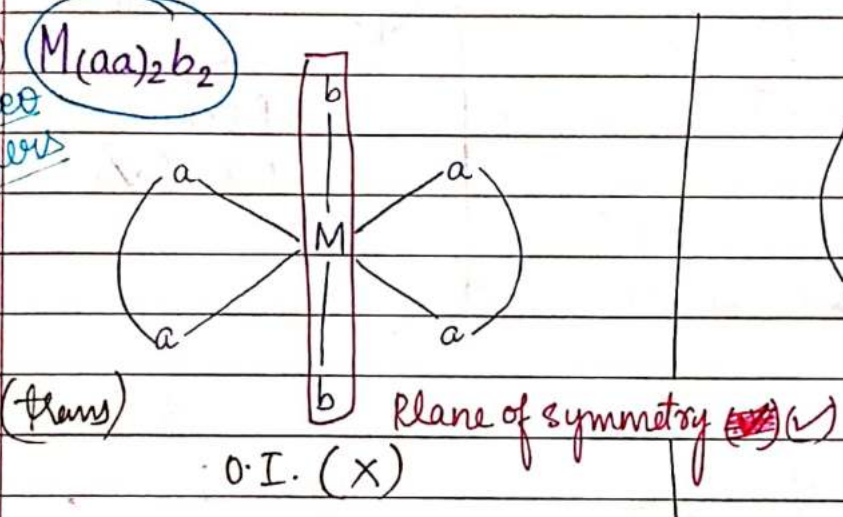
Date _____
Page _____

(*) \Rightarrow Tetrahedral complexes of type Ma_2cd , $M(cab)_2$, $M(cab)(cd)$ can show optical isomerism.

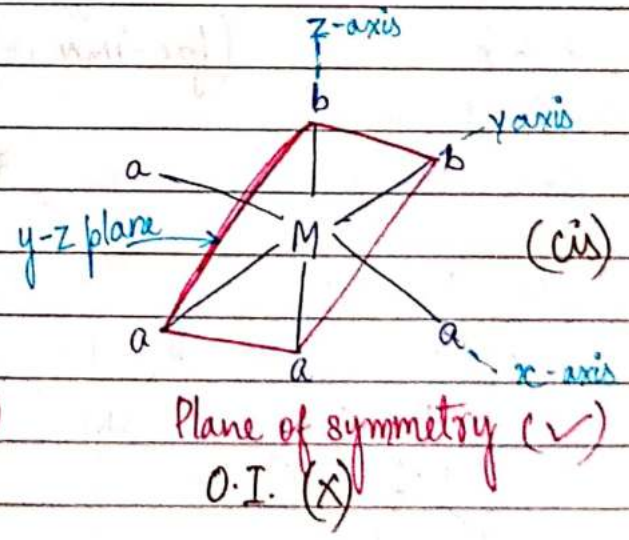
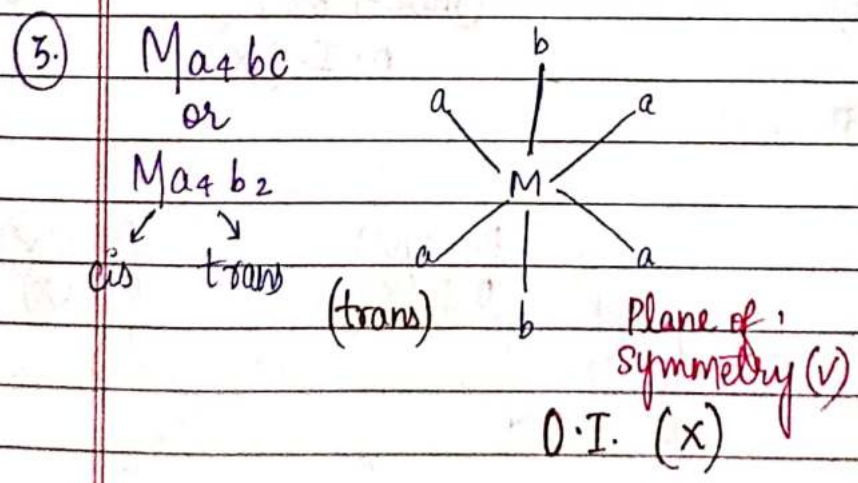


\rightarrow Optical isomerism in octahedral complexes

3 stereoisomers



(2) $M(aa)_2bc \rightarrow$ cis \leftrightarrow d \leftrightarrow e } 3 isomers
 \rightarrow trans
(Same as above)



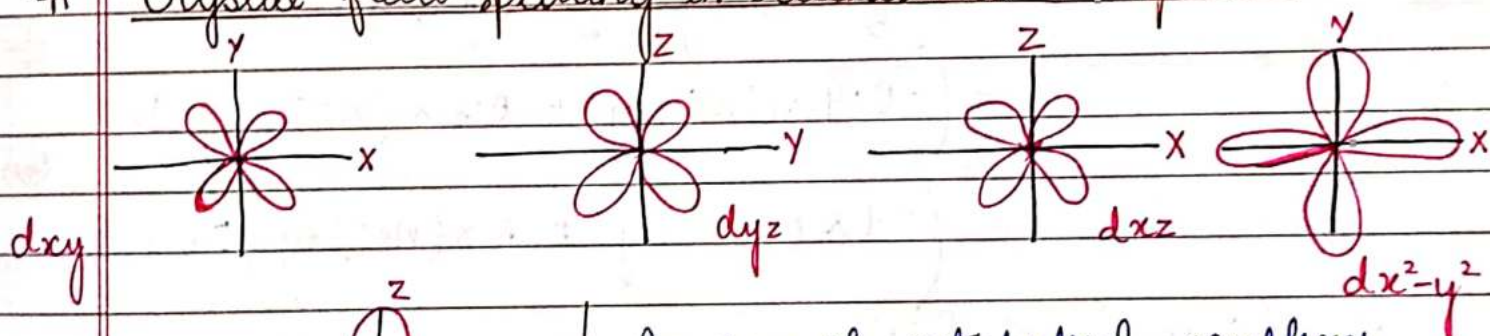
CRYSTAL FIELD THEORY

Crystal Field theory is an electrostatic model, according to CFT the interaction b/w metal and ligands is purely electrostatic.

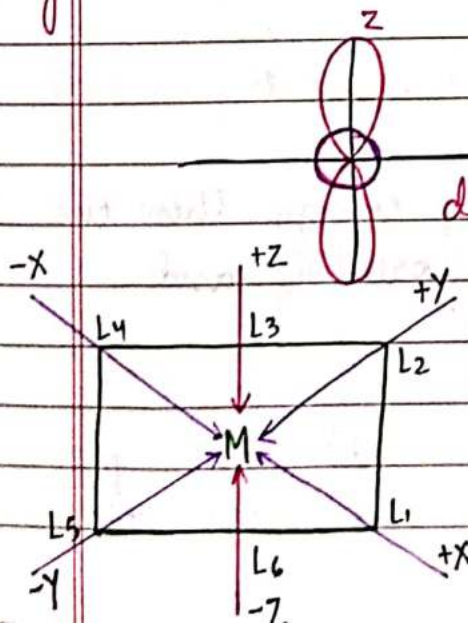
Main points -

- 1.) Anionic ligands are treated as negative point charges & Neutral ligands are dipoles with -ve end orientated towards central metal atom.
- 2.) There is no intermixing of atomic orbitals or hybridisation
- 3.) In the absence of ligand, the d orbitals of central metal atom are degenerate but in the presence of ligands these d-orbitals will not degenerate & they get splitted and this phenomenon is called crystal field splitting.

Crystal field splitting in octahedral complexes.



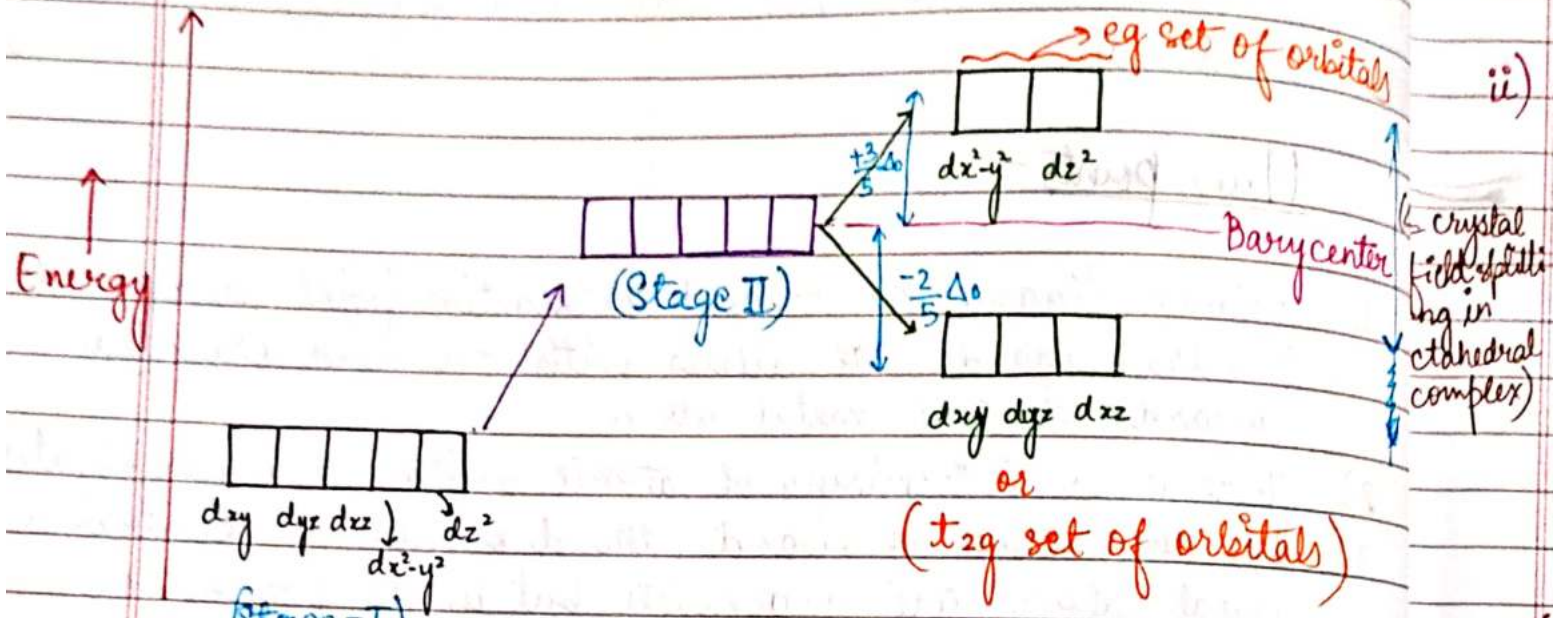
In case of octahedral complexes ligands approach the central metal atom from $(+x, -x)$, $(+y, -y)$, $(+z, -z)$



Ligands approach towards metal along the axis. So the orbitals which lie along the axis feels more repulsion. & their energy is raised & those orbitals which lie b/w the axis their energy is relatively low.

($d_{x^2-y^2}$, d_{z^2})

and this phenomenon is called crystal field splitting in octahedral complexes.



$$\begin{aligned}
 (\Delta_o)_{CFSE} &= \left(\frac{-2}{5} \Delta_o \times (n_{e^-})_{t_{2g}} + \frac{3}{5} \Delta_o \times (n_{e^-})_{e_g} \right) \\
 &= \left(\frac{-2}{5} (n_{e^-})_{t_{2g}} + \frac{3}{5} \times (n_{e^-})_{e_g} \right) \Delta_o \\
 &= \left(-0.4 \times (n_{e^-})_{t_{2g}} + 0.6 \times (n_{e^-})_{e_g} \right) \Delta_o \\
 (\Delta_o = 10 Dq) &= \left(-4 \times (n_{e^-})_{t_{2g}} + 6 \times (n_{e^-})_{e_g} \right) Dq
 \end{aligned}$$

i) $(\Delta_o)_{CFSE} < \text{Pairing energy (P)}$ → energy required to pair the e⁻

• if $(\Delta_o)_{CFSE}$ is less than the pairing energy then the 4th e⁻ enters one of the eg set of orbitals and configuration will become t_{2g}³ eg¹.

$d^3 = t_{2g}^3 e_g^0$	$d^6 = t_{2g}^4 e_g^2$	$d^9 = t_{2g}^6 e_g^3$
$d^4 = t_{2g}^3 e_g^1$	$d^7 = t_{2g}^5 e_g^2$	$d^{10} = t_{2g}^6 e_g^4$
$d^5 = t_{2g}^3 e_g^2$	$d^8 = t_{2g}^6 e_g^2$	

The above case $(\Delta_o)_{CFSE} < P$ is possible by weak field ligands which form high spin complexes.

ii) $(\Delta_o)_{CFSE} > \text{Pairing energy}$

= crystal field splitting in octahedral complex

In this case the 4th electron enters one of the t_{2g} set of orbitals and the configuration will become $t_{2g}^4 e_g^0$.

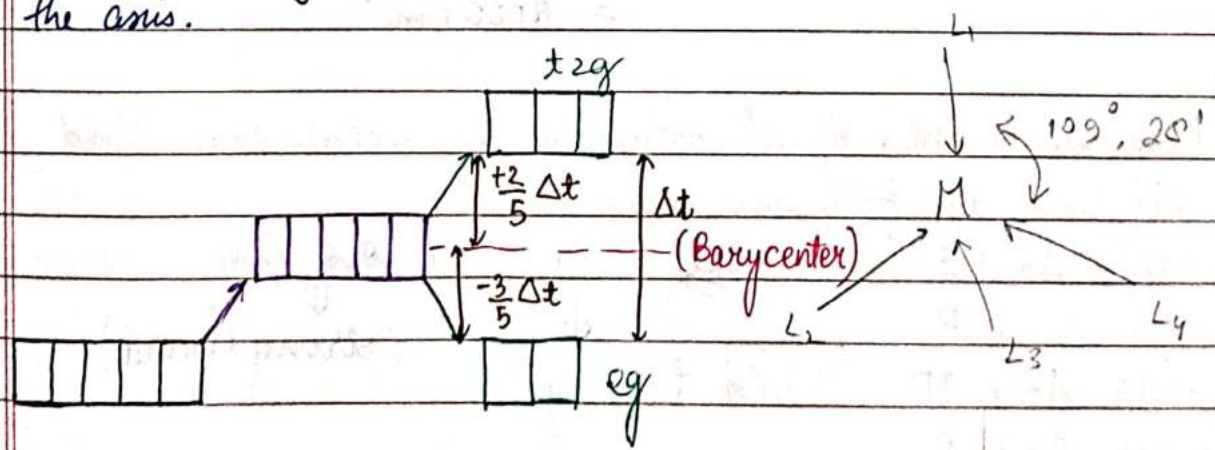
- $d^3 = t_{2g}^3 e_g^0$
- $d^4 = t_{2g}^4 e_g^0$
- $d^5 = t_{2g}^5 e_g^0$
- $d^6 = t_{2g}^6 e_g^0$
- $d^7 = t_{2g}^6 e_g^1$
- $d^8 = t_{2g}^6 e_g^2$
- $d^9 = t_{2g}^6 e_g^3$
- $d^{10} = t_{2g}^6 e_g^4$

(**) This case is possible by strong field ligands which form low spin complexes.

(***) $(\Delta_o)_{CFSE \text{ net}} = (-4 \times (n_{e^-})_{t_{2g}} + 6 \times (n_{e^-})_{e_g}) + P$ (Pairing energy)

Crystal field splitting in tetrahedral complexes

(*) In case of tetrahedral complexes the d orbital splitting is inverted as ligands approach the central metal atom between the axis.



$$(\Delta_t)_{CFSE} = \left(-\frac{3}{5} (n_{e^-})_{eg} + \frac{2}{5} \times (n_{e^-})_{t2g} \right) \Delta_t$$

$$= \left(-0.6 \times (n_{e^-})_{eg} + 0.4 \times (n_{e^-})_{t2g} \right) \Delta_t$$

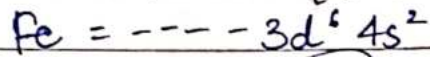
$$(\Delta_t = 10Dq) = (-6 \times (n_{e^-})_{eg} + 4 (n_{e^-})_{t2g}) Dq$$

$$(\Delta_t)_{CFSE} = \frac{4}{9} (\Delta_o)_{CFSE}$$

NEET
2019 Q

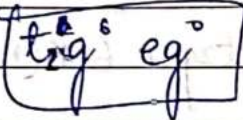
What is the correct electronic configuration of central metal atom in $K_4[Fe(CN)_6]$ acc. to CFT?

Solⁿ



Fe⁺² = --- (3d⁶) & ligand is strong field

so 4th e⁻ gets paired in t_{2g}



Q
Solⁿ

The CFSE for $[CoCl_6]^{-4}$ is 18000 cm^{-1} . Find CFSE for $[CoCl_4]^{-2}$?

$(\Delta_o)_{CFSE} = 18000 \text{ cm}^{-1}$ $(\Delta_t)_{CFSE} = ?$

$$(\Delta_t)_{CFSE} = \frac{4}{9} (\Delta_o)_{CFSE}$$

$$= 8000 \text{ cm}^{-1}$$

Q Low spin complex of d⁶ cation in an octahedral field will have the following energy:

a) $-12/5 \Delta_o + P$

b) $-2/5 \Delta_o + 2P$

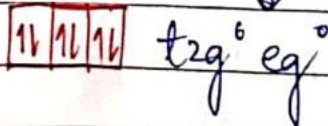
c) $-12/5 \Delta_o + 3P$

d) $-2/5 \Delta_o + P$

Solⁿ d⁶ (low spin)

↓

(strong ligand)



$$(\Delta_o)_{CFSE} = \left(-\frac{2}{5} \times (n_{e^-})_{t2g} + \frac{3}{5} (n_{e^-})_{eg} \right) \Delta_o + 3P$$

$$= \left(\frac{-2 \times 6}{5} \right) \Delta_0 + 3P \Rightarrow \frac{-12}{5} \Delta_0 + 3P$$

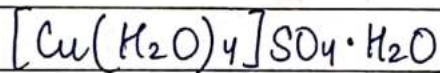
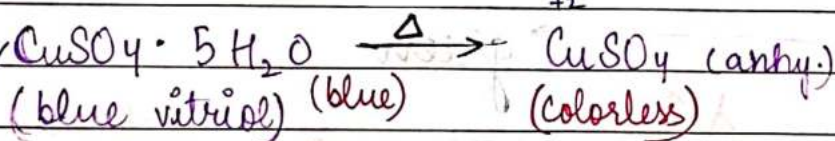
Q ^(Δ_0) CFSE for metal ion in d^7 configuration in the presence of strong ligand?

solⁿ d^7 (strong ligand) = $t_{2g}^6 e_g^1$

$$\begin{aligned} (\Delta_0)_{CFSE} &= \left(\frac{-2}{5} \times (ne^-) t_{2g} + \frac{3}{5} (ne^-) e_g \right) \\ &= \left(\frac{-2 \times 6}{5} + \frac{3 \times 1}{5} \right) \Delta_0 + 3P \\ &= \frac{-9}{5} \Delta_0 + 3P \Rightarrow -1.8 \Delta_0 + 3P \end{aligned}$$

Colour in coordination compounds

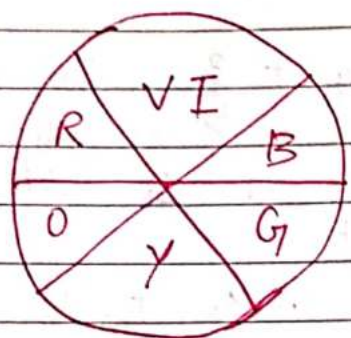
Color in coordination compounds can be explained on the basis of ~~CFSE~~ CFT. For color in coordination compounds, both unpaired electrons and ligands are necessary.



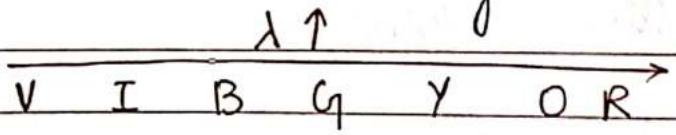
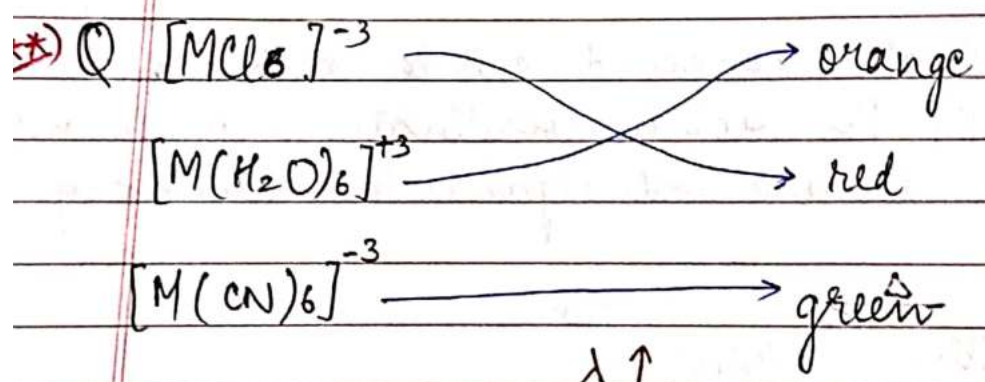
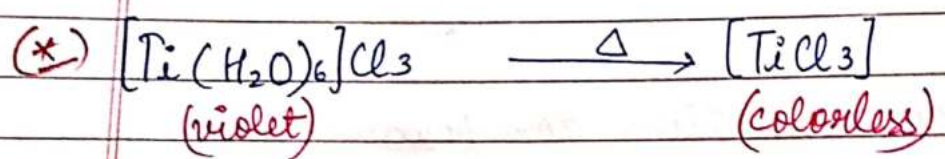
(***) For colour in coordination compound, the complex must absorb light from visible region: (4000 Å - 7500 Å)

→ The complementary colour is generated from the wavelength left over.

Coordination entity	Wavelength of light absorbed (nm)	Colour of light absorbed	Colour of coordination ion entity
$[\text{CoCl}(\text{NH}_3)_5]^{2+}$	535	yellow	Violet
$[\text{Co}(\text{NH}_3)_5(\text{H}_2\text{O})]^{3+}$	500	Blue green	Red
$[\text{Co}(\text{NH}_3)_6]^{3+}$	475	Blue	Yellow orange



(Complimentary colour cycle)



(strength of ligand ↑, λ ↓)

e) Color absorbed in $\text{CuSO}_4(\text{aq})$? → Blue

- Orange b) Blue c) Yellow d) Pink

Factors affecting CFSE -

- 1.) Strength of ligand, stronger the ligand more will be CFSE.
- 2.) $CFSE \propto$ oxidation state of central ~~metal~~ atom
- 3.) Nature of central metal atom

3d series	$_{21}Sc$	Ti	V	Cr	-----	-----	$_{30}Zn$
4d							\downarrow 40-50% \uparrow
4d series	$_{39}Y$	Zr	Nb	Mo	-----	-----	$_{48}Cd$
							\downarrow 50% increase
5d series	$_{57}La$	Hg	-----	-----	-----	-----	Hg

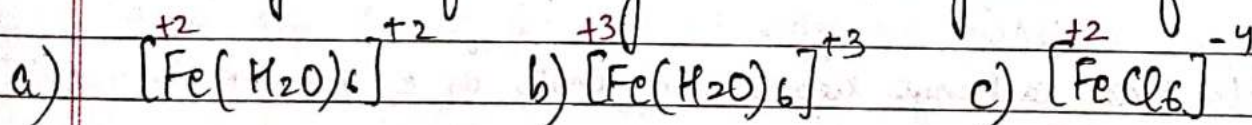
Order of CFSE $3d < 4d < 5d$

- 4.) Geometry of the complex.

$$(\Delta_t) CFSE = \frac{4}{9} (\Delta_o) CFSE$$

- 5.) If chelation occurs, the complex will be more stable
 \downarrow CFSE will be more:
 ring formation

Q Arrange the following in increasing order of CFSE.



Order- $c < a < b$
 \downarrow
 ligand is weak

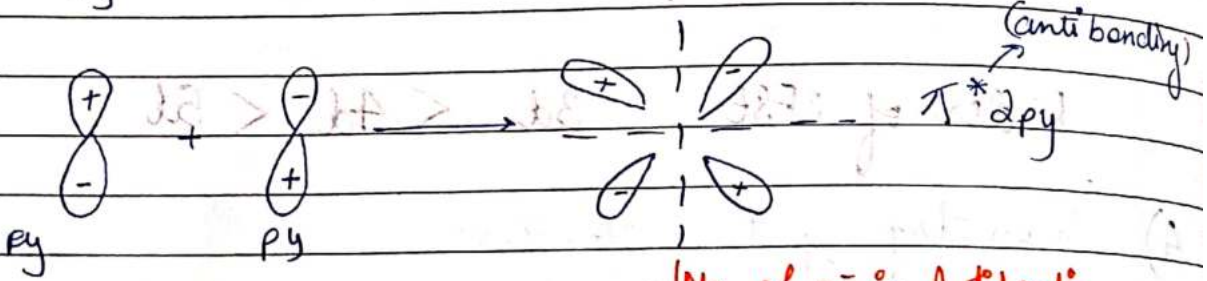
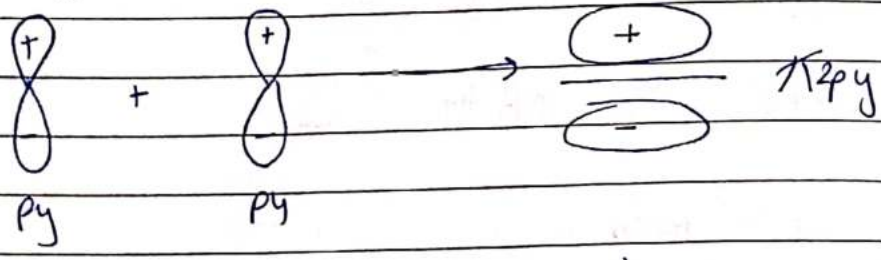
Limitations of CFT-

- 1.) Acc. to CFT, the anionic ligands should have greatest

splitting effect but the anionic ligands are actually found at the lower end of spectrochemical series.

2.) CFT was unable to explain covalent nature of bond b/w ligand and central metal atom.

Bonding in metal Carbonyls



(No. of e⁻ in antibonding molecular orbital)

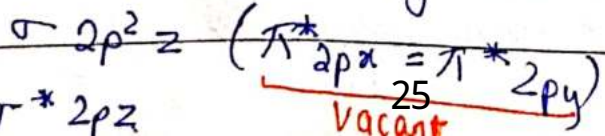
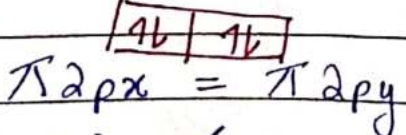
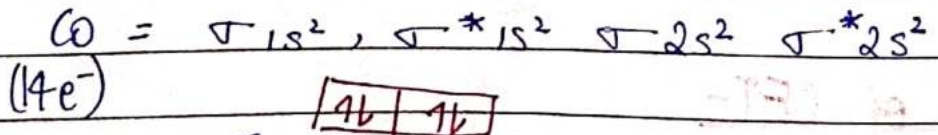
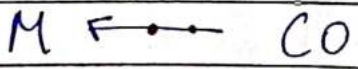
$$\text{Bond order} = \frac{1}{2} (\text{Nb} - \text{Na})$$

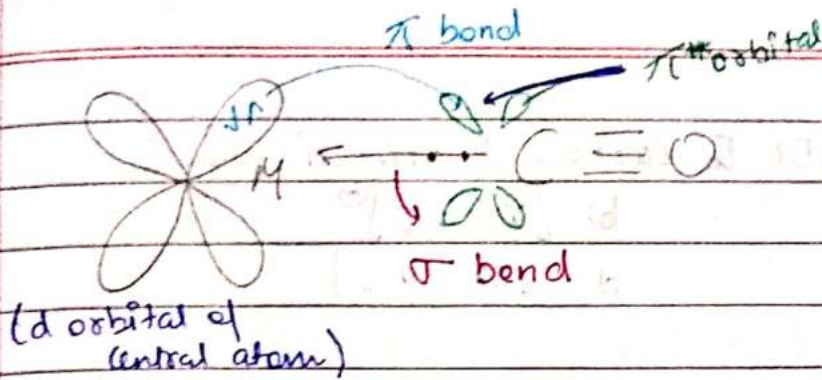
(it gives idea about 2 ↓)

no. of bonds b/w atoms) (No. of e⁻ in bonding molecular orbital)

(Na ↑, Bond order ↓ Bond length ↑)
Bond order ∝ 1 / Bond length

In metal carbonyls ligand CO donate its e⁻ pair to the central metal atom.





The metal-carbon σ bond is formed by donation of lone pair of electrons from CO but the metal-carbon π bond is formed by ~~donation~~ transfer of electrons from filled d orbital of metal to vacant π^* of CO, due to formation of additional π bond the bond length of metal-carbon bond decreases. And due to shifting of electrons from d orbital to π^* of CO, the bond order of CO decreases and bond length between C and O increases.

Synergic bonding

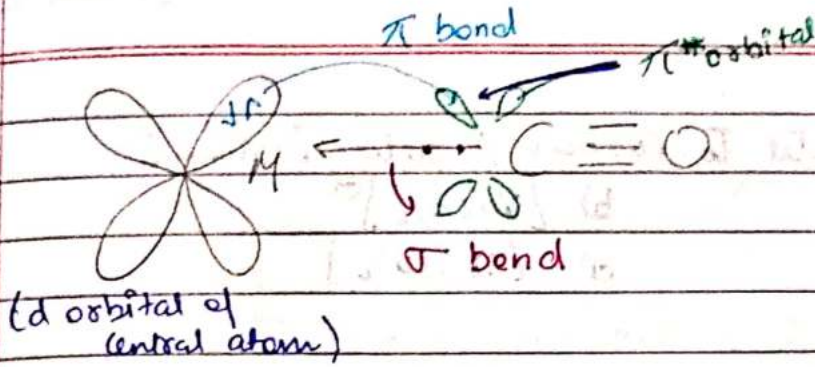
If a ligand donates its e^- pair to the central metal atom and accepts e^- pair from the filled d orbital then this type of bonding is called synergic bonding and the ligands are called π acid ligands.

eg CO, NO etc.

Ques Which has lowest C-O bond order or highest C-O bond length?

- a) $[\text{Fe}(\text{CO})_5]$ b) $[\text{Mn}(\text{CO})_6]^+$
 c) $[\text{Ni}(\text{CO})_4]$ d) $[\text{V}(\text{CO})_6]^-$

(central metal atom $4d$ -ve charge $\frac{5}{2}$ \uparrow \uparrow \uparrow \uparrow \uparrow
 \uparrow e^- CO in π^* \uparrow $\frac{1}{2}$ \uparrow \uparrow \uparrow \uparrow \uparrow , So bond order of CO \downarrow)



The metal-carbon σ bond is formed by donation of lone pair of electrons from CO but the metal-carbon π bond is formed by donation of transfer of electrons from filled d orbital of metal to vacant π^* of CO, due to formation of additional π bond the bond length of metal-carbon bond decreases. And due to shifting of electrons from d orbital to π^* of CO, the bond order of CO decreases and bond length between C and O increases.

Synergic bonding

If a ligand donates its e^- pair to the central metal atom and accepts e^- pair from the filled d orbital then this type of bonding is called synergic bonding and the ligands are called π acid ligands.

eg CO, NO etc.

Q. Which has lowest C-O bond order or highest C-O bond length?

- a) $[Fe(CO)_5]$ b) $[Mn(CO)_6]^+$
c) $[Ni(CO)_4]$ d) $[V(CO)_6]^-$

(central metal atom $4d$ -ve charge $\frac{1}{2}$ π π^* σ σ^*
 π e^- CO π π^* $\frac{1}{2}$ σ σ^* σ , So bond order of CO \downarrow)