

# NCERT SOLUTIONS

## CLASS-XII CHEMISTRY

### CHAPTER-9

### COORDINATION COMPOUNDS

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**Q1.** Using Werner's postulates describe the bonding present in coordination compounds.

Ans :

- ( a ) A metal shows two kinds of valencies viz primary valency and secondary valency. Negative ions satisfy primary valencies and secondary valencies are filled by both neutral ions and negative ions.
- ( b ) A metal ion has a fixed amount of secondary valencies about the central atom. These valencies also orient themselves in a particular direction in the space provided to the definite geometry of the coordination compound.
- ( c ) Secondary valencies cannot be ionized, while primary valencies can usually be ionized.

**Q2.**  $\text{FeSO}_4$  solution is mixed with  $(\text{NH}_4)_2\text{SO}_4$  solution in the molar ratio of 1:1 molar. It gives a positive test for  $\text{Fe}^{2+}$  ion. However, when  $\text{CuSO}_4$  solution is mixed with aqueous ammonia in the molar ratio of 1:4, it does not test positive for  $\text{Cu}^{2+}$  ion. Why?

Ans :

$\text{FeSO}_4$  solution when mixed with  $(\text{NH}_4)_2\text{SO}_4$  in 1 : 1 molar ratio produces a double salt  $\text{FeSO}_4(\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$ . This salt is responsible for giving the  $\text{Fe}^{2+}$ .  $\text{CuSO}_4$  mixed with aqueous ammonia in the ratio of 1:4 gives a complex salt. The complex salt does not ionize to give  $\text{Cu}^{2+}$ , hence failing the test.

**Q3.** Explain the following giving two examples for each of them :ligand, coordination entity, coordination polyhedron , coordination number,heteroleptic and homoleptic.

Ans :

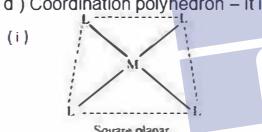
- ( a ) Ligands – they are neutral molecules or negative ions bound to a metal atom in coordination entity. Example-  $\text{Cl}^-$ ,  $\text{OH}^-$
- ( b ) Coordination entity – they are electrically charged radicals or species. They constitute of a central ion or atom surrounded by neutral molecules or ions. Example –  $[\text{Ni}(\text{CO})_4]$ ,  $[\text{COCl}_3(\text{NH}_3)_3]$

( c ) Coordination number– it is the number of bonds formed between ligands and central atom/ion.

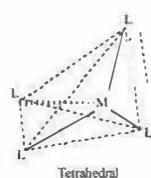
Example : ( i ) In  $\text{K}_2[\text{PtCl}_6]$ , 6 chloride ions are attached to Pt in the coordinate sphere. Thus, 6 is the coordination number of Pt.

( ii ) In  $[\text{Ni}(\text{NH}_3)_4]\text{Cl}_2$ , the coordination number of the central metal ion (Ni) is 4

( d ) Coordination polyhedron – it is the spatial positioning of ligands that are directly connected to the central atom in the coordination sphere. Example –



( ii )



( v ) Heteroleptic : they are complexes with their metal ion being bounded to more than one kind of donor group. Example –  $[\text{Co}(\text{NH}_3)_4\text{Cl}_2]^+$ ,  $[\text{Ni}(\text{CO})_4]$

( vi ) Homoleptic : they are complexes with their metal ion being bounded to only one type of donor. Example –  $[\text{PtCl}_4]^{2-}$ ,  $[\text{Co}(\text{NH}_3)_6]^{3+}$

**Q4.** Providing two examples in each case, explain the following terms :unidentate, ambidentate and didentate ligands?

Ans :

( i ) Unidentate ligands : these are ligands with one donor site. Example  $\text{Cl}^-$ ,  $\text{NH}_3$

( ii ) Ambidentate ligands : these are ligands that fasten themselves to the central metal ion/ atom via two different atoms.

Example  $\text{NO}_2^-$  or  $\text{ONO}^-$ ,  $\text{CN}^-$  or  $\text{NC}^-$

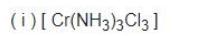
( iii ) Didentate – these are ligands with two donor sites.

Example – Ethane-1,2-diamine , Oxalate ion (  $\text{C}_2\text{O}_4^{2-}$  )

**Q5.** Present below are coordination entities, state the oxidation number of their metals :

- ( i )  $[\text{Cr}(\text{NH}_3)_3\text{Cl}_3]$
- ( ii )  $[\text{PtCl}_4]^{2-}$
- ( iii )  $K_3[\text{Fe}(\text{CN})_6]$
- ( iv )  $[\text{Co Br}_2(\text{en})_2]^+$
- ( v )  $[\text{Co}(\text{H}_2\text{O})(\text{CN})(\text{en})_2]^{2+}$

Ans :



$$\Rightarrow x + 3(0) + 3(-1) = 0$$

$$x - 3 = 0$$

$$x = 3$$



$$\Rightarrow x + 4(-1) = -2x$$

$$x = +2$$



$$\Rightarrow [Fe(CN)<sub>6</sub>]<sup>3-</sup>$$

$$\Rightarrow x + 6(-1) = -3$$

$$x = +3$$



$$\Rightarrow x + 2(-1) + 2(0) = +1$$

$$x - 2 = +1$$

$$x = +3$$

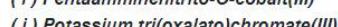
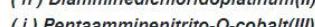
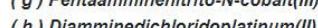
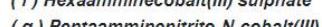
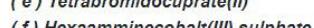
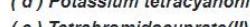
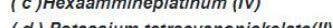
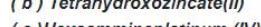
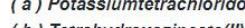
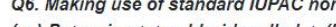


$$\Rightarrow x + 0 + (-1) + 2(0) = +2$$

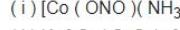
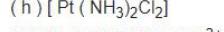
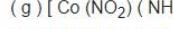
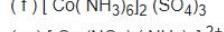
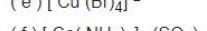
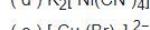
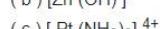
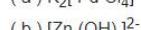
$$x - 1 = +2$$

$$x = +3$$

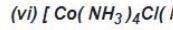
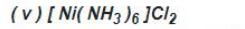
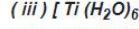
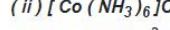
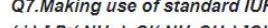
**Q6.** Making use of standard IUPAC norms give the formulae of the following compounds :



Ans :



**Q7.** Making use of standard IUPAC norms give the systematic names of the compounds given below :



Ans :

( i ) Diamminechlorido(methylamine) platinum(II) chloride

( ii ) Hexaamminecobalt(III) chloride

( iii ) Hexaquatitanium(III) ion

( iv ) Tetrachloridonickelate(II) ion

( v ) Hexaamminenickel(II) chloride

( vi ) Tetraamminichloridonitrito-N-Cobalt(III) chloride

( vii ) Hexaquamanganese(II) ion

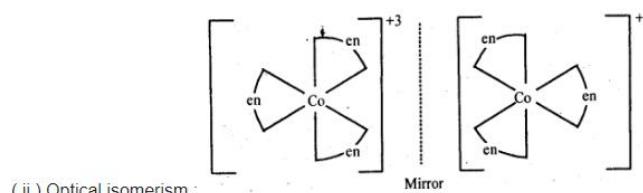
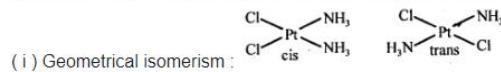
( viii ) Tetracarbonylnickel(0)

( ix ) Tris(ethane-1, 2-diammine) cobalt(III) ion

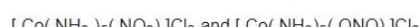
**Q8.** What are the various kinds of isomerism present in coordination compounds ? ( Provide an example for each case )

Ans :

The various types of **isomerism** present in coordination compounds are :



( iii ) Linkage isomerism : This is found in complexes that have ambidentate ligands. For e.g. :



( iv ) Coordination isomerism :

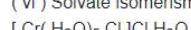
This kind of isomerism comes up when ligands are interchanged between anionic and cationic entities of different metal ions present in the complex.

Example –  $[\text{Cr}(\text{NH}_3)_6] [\text{Co}(\text{CN})_6]$

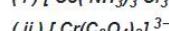
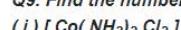
( v ) Ionisation isomerism :

This is the kind of isomerism where a counter ion takes the place of a ligand inside the coordination sphere.. For e.g.,  $[\text{Co}(\text{NH}_3)_5\text{Br}]\text{SO}_4$  and  $[\text{Co}(\text{NH}_3)_5\text{SO}_4]\text{Br}$

( vi ) Solvate isomerism :

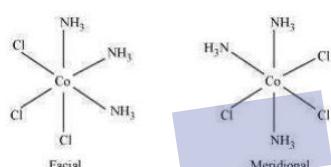


**Q9. Find the number of geometrical isomers possible in the coordination entities given below :**

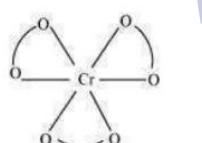


Ans :

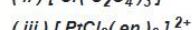
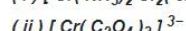
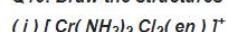
( i ) In  $[\text{Co}(\text{NH}_3)_3\text{Cl}_3]$  two isomers are possible.



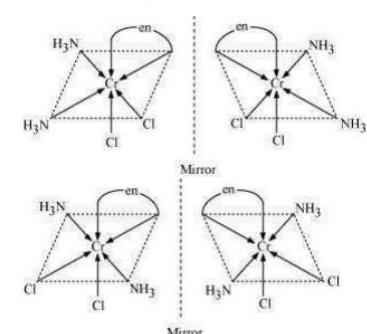
( ii ) In  $[\text{Cr}(\text{C}_2\text{O}_4)_3]^{3-}$  no geometric isomers are present because it is a bidentate ligand.

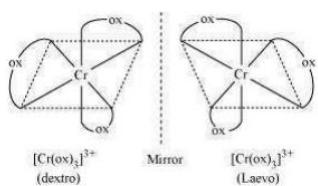
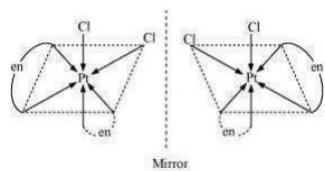


**Q10. Draw the structures of the optical isomers given below :**

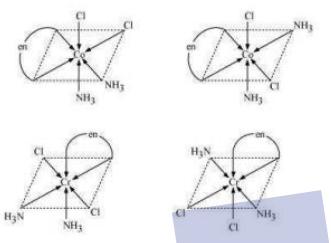
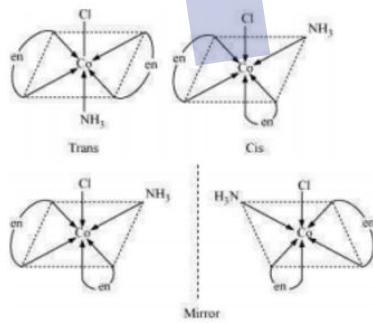
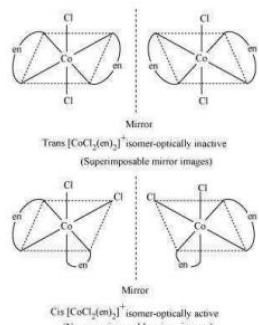


Ans :



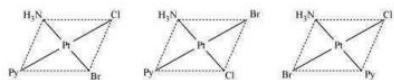
( ii )  $\text{Cr}(\text{ox})_3^{3+}$ ( iii )  $[\text{PtCl}_2(\text{en})_2]^{2+}$ **Q11.** Sketch all the isomers ( optical and geometrical) of:( i )  $[\text{Co}(\text{NH}_3)_2 \text{Cl}_2(\text{en})]^+$ ( ii )  $[\text{Co}(\text{NH}_3) \text{Cl}(\text{en})_2]^{2+}$ ( iii )  $[\text{CoCl}_2(\text{en})_2]^+$ 

Ans:

( i )  $[\text{Co}(\text{NH}_3)_2 \text{Cl}_2(\text{en})]^+$ ( ii )  $[\text{Co}(\text{NH}_3) \text{Cl}(\text{en})_2]^{2+}$ ( iii )  $[\text{CoCl}_2(\text{en})_2]^+$ **Q12.** What are the geometrical isomers of  $[\text{Pt}(\text{NH}_3)(\text{Br})(\text{Cl})(\text{py})]$ ? Also, what number of these isomers will show optical isomerism?

Ans:

[ Pt( NH<sub>3</sub>)( Br)(Cl)(py) ]



None of the above isomers will exhibit optical isomerism.

**Q13. Explain why :**

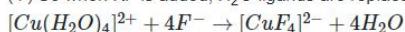
**Aqueous CuSO<sub>4</sub> solution, which is blue in color, yields :**

- (a) a green precipitate if aqueous KF is added to it, and
- (b) a bright green solution if aqueous KCl is added to it.

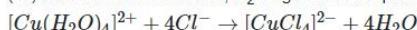
Ans:

The blue color of aqueous CuSO<sub>4</sub> solution is because of the presence of [ Cu(H<sub>2</sub>O)<sub>4</sub> ]<sup>2+</sup> ions.

(i) So when KF is added, H<sub>2</sub>O ligands are replaced by F<sup>-</sup> ligands which yield green colored [ CuF<sub>4</sub> ]<sup>2-</sup> ions.



(ii) So when KCl is added, H<sub>2</sub>O ligands are replaced by Cl<sup>-</sup> ligands which yield bright green colored [ CuCl<sub>4</sub> ]<sup>2-</sup> ions.



**Q14. Give the coordination entity produced when excess of aqueous KCN is put into an aqueous solution of CuSO<sub>4</sub>? When hydrogen sulphide gas is bubbled through this solution no precipitate of copper sulphide is formed. Why?**

Ans:



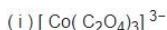
Therefore, the coordination entity obtained in the above process is K<sub>2</sub>[ Cu(CN)<sub>4</sub> ].

As the above coordination entity is highly stable it does not ionize to yield Cu<sup>2+</sup> ions. Thus, no precipitate is obtained when hydrogen Sulphide gas is bubbled through it.

**Q15. On the grounds of Valence bond theory, write about the nature of bonding in the given coordination entities :**

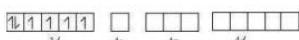
- (i) [ Co(C<sub>2</sub>O<sub>4</sub>)<sub>3</sub> ]<sup>3-</sup>
- (ii) [ CoF<sub>6</sub> ]<sup>3-</sup>
- (iii) [ Fe(CN)<sub>6</sub> ]<sup>4-</sup>
- (iv) [ FeF<sub>6</sub> ]<sup>3-</sup>

Ans:



In this complex, the oxidation state of cobalt is +3.

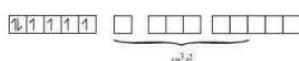
Orbitals of Co<sup>3+</sup> ion :



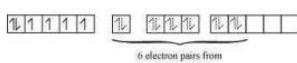
Oxalate is a weak field ligand. Thus, it will not cause the pairing of the 3d orbital electrons.

As there are 6 ligands, hybridization has to be either sp<sup>3</sup>d<sup>2</sup> or d<sup>2</sup>sp<sup>3</sup> hybridization.

sp<sup>3</sup>d<sup>2</sup> hybridization of Co<sup>3+</sup>:



The 6 electron pairs from the 3 oxalate ions (oxalate anion is a bidentate ligand) occupy these sp<sup>3</sup>d<sup>2</sup> orbitals :

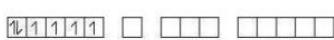


Thus, the complex shows octahedral geometry .

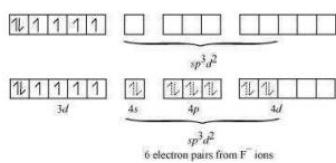


In this complex Cobalt has an oxidation state of +3.

Orbitals of Co<sup>3+</sup> ion:



As, fluoride ion is a weak field ligand it will not cause the 3d electrons to pair. Hence, the  $\text{Co}^{3+}$  ion will go through  $\text{sp}^3\text{d}^2$  hybridization.  
 $\text{sp}^3\text{d}^2$  hybridized orbitals of  $\text{Co}^{3+}$  ion are :



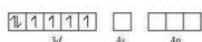
Thus, the complex has a geometric configuration of an octahedral and it is paramagnetic.

( iii )  $[\text{Fe}(\text{CN})_6]^{4-}$

In this coordination complex, the oxidation state of Fe is +3.

$\text{Fe}^{2+}$  : Electronic configuration is  $3\text{d}^6$

Orbitals of  $\text{Fe}^{2+}$  ion :



Since  $\text{CN}^-$  is a strong field ligand, it causes the unpaired 3d electrons to pair up:

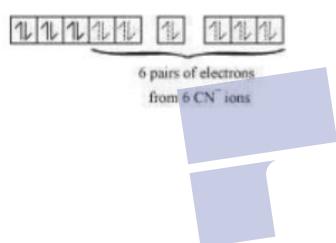


As there are six ligands around the central metal ion, the most practical hybridization is  $\text{d}^2\text{sp}^3$ ,  $\text{d}^2\text{sp}^3$  hybridized orbitals of  $\text{Fe}^{2+}$  are:



6 electron pairs from  $\text{CN}^-$  ions take the place of the six hybrid  $\text{d}^2\text{sp}^3$  orbitals.

Then,

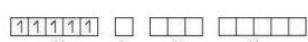


Thus, the geometry of the complex is octahedral and it is a diamagnetic complex ( since all the electrons are paired ).

( iv )  $[\text{FeF}_5]^{3-}$

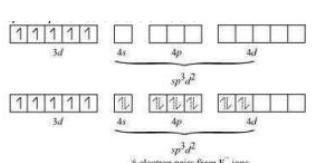
In this coordinate entity the oxidation state of iron is +3.

Orbitals of  $\text{Fe}^{3+}$  ion:



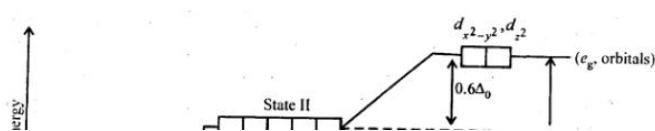
There are 6  $\text{F}^-$  ions. Hence, it will go through  $\text{d}^2\text{sp}^3$  or  $\text{sp}^3\text{d}^2$  hybridization.

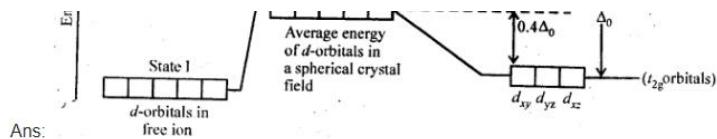
Since  $\text{F}^-$  is a weak field ligand, it does not cause the pairing of the electrons in the 3d orbital. Thus, the most practical hybridization is  $\text{sp}^3\text{d}^2$ .  $\text{sp}^3\text{d}^2$  hybridized orbitals of  $\text{Fe}$  are :



Thus, the geometry of this coordinate entity is octahedral.

Q16. Sketch a figure depicting the split of d orbitals within an octahedral crystal field.





Ans:

**Q17.** What do you understand by the term "spectrochemical series"? Differentiate between a strong field ligand and a weak field ligand.

Ans:

Series of common ligands in an ascending order of their crystal-field splitting energy (CFSE) is termed as Spectrochemical series. Strong field ligands have larger values of CFSE. Whereas, weak field ligands have smaller values of CFSE.

**Q18.** Define crystal field splitting energy. In what way is the d-orbital configuration depending upon the magnitude of  $\Delta_0$ ?

Ans:

Crystal-field splitting energy is the difference in the energy between the two levels (i.e.,  $t_{2g}$  and  $e_g$ ) that have split from a degenerated d orbital because of the presence of a ligand. It is symbolised as  $\Delta_0$ .

Once the orbitals split up, electrons start filling the vacant spaces. An electron each goes into the three  $t_{2g}$  orbitals, the fourth electron, however, can enter either of the two orbitals:

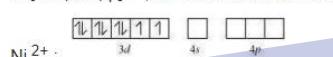
- (1) It can go to the  $e_g$  orbital (producing  $t_{2g}^3 e_g^1$  like electronic configuration), or
- (2) it can go to the  $t_{2g}$  orbitals (producing  $t_{2g}^4 e_g^0$  like electronic configuration).

This filling of the fourth electron is based on the energy level of  $\Delta_0$ . If a ligand has a  $\Delta_0$  value smaller than the pairing energy, then the fourth electron enters the  $e_g$  orbital. However, if the value of  $\Delta_0$  is greater than the value of pairing energy, the electron enters  $t_{2g}$  orbital.

**Q19.** Explain why  $[Ni(CN)_4]^{2-}$  is diamagnetic while  $[Cr(NH_3)_6]^{3+}$  is paramagnetic.

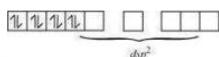
Ans:

In  $[Ni(CN)_4]^{2-}$ , Ni has an oxidation state of +2. Thus, it has  $d^8$  configuration.



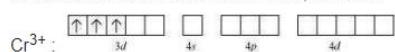
pdfelement

CN<sup>-</sup> being a strong field ligand causes the electrons in 3d orbitals to pair. This causes, Ni<sup>2+</sup> to undergo  $dsp^2$  hybridization.



Since all the electrons are paired, it is diamagnetic in nature.

Cr has an oxidation state of +3. Thus, it has a  $d^3$  configuration. As, NH<sub>3</sub> is not a strong field ligand it does not cause the electrons in the 3d orbital to pair.



It undergoes  $d^2sp^3$  hybridization and the 3d orbital electrons remain unpaired. Thus,  $[Ni(CN)_4]^{2-}$  is paramagnetic in nature.

**Q20.** Why is a  $[Ni(H_2O)_6]^{2+}$  solution green when a  $[Ni(CN)_4]^{2-}$  solution is colorless?

Ans:

$[Ni(H_2O)_6]^{2+}$  consists of Ni<sup>2+</sup> ion with  $3d^8$  electronic configuration. In this configuration, there are two unpaired electrons which cannot pair up because H<sub>2</sub>O is a weak ligand. Thus, the d-d transition absorbs the incoming light and it emits a green light. Thereby, giving a green color to the solution.

$[Ni(CN)_4]^{2-}$  consists of Ni<sup>2+</sup> ion with  $3d^8$  electronic configuration. But, CN<sup>-</sup> is present here, which is a strong ligand and in its presence, the unpaired electrons pair up. Thus, there is no d-d transition so no color.

**Q21.** Why do  $[Fe(H_2O)_6]^{2+}$  and  $[Fe(CN)_6]^{4-}$  have different colors in dilute solutions?

Ans:

$[Fe(H_2O)_6]^{2+}$  and  $[Fe(CN)_6]^{4-}$  have two different ligands H<sub>2</sub>O and CN<sup>-</sup>. CN<sup>-</sup> being a strong field ligand has a higher value of CFSE (crystal field splitting energy) than water. As a result, the d-d transitions absorb and give back different wavelengths of light. Thus, they have different colors in a solution.

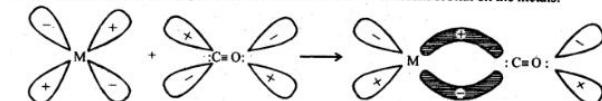
**Q22.** Write about the bonding present in metal carbonyls.

Ans:

In metal carbonyls, the metal-carbon bond contains both the  $\sigma$  and  $\pi$  bond characters. A  $\sigma$  bond forms when a lone pair of electrons is donated to the empty orbital of the metal by the carbonyl carbon. A  $\pi$  bond forms when a pair of electrons is donated to the empty antibonding  $\pi^*$  orbital by the filled d orbital of the metal. This in entirety

stabilizes and strengthens the metal-ligand bonding.

The above two types of bonding is represented as :



**Q23.** Given below are complex entities, provide the coordination number, oxidation state and  $d$ -orbital occupation of their central metal ion:

- (a)  $K_3[Co(C_2O_4)_3]$
- (b)  $cis-[Cr(en)_2Cl_2]Cl$
- (c)  $(NH_4)_2[CoF_4]$
- (d)  $[Mn(H_2O)_6]SO_4$

Ans:

(a)  $K_3[Co(C_2O_4)_3]$

Central metal ion: Co.

Coordination number = 6.

We know,

Oxidation state is :

$$x - 6 = -3$$

$$x = +3$$

The d orbital occupation :  $t_{2g}^6 e_g^0$ .

(b)  $cis-[Cr(en)_2Cl_2]Cl$

Central metal ion : Cr.

Coordination number = 6.

We know,

Oxidation state is :

$$x + 2(0) + 2(-1) = +1$$

$$x - 2 = -1$$

$$x = +3$$

The d orbital occupation :  $t_{2g}^3$ .

(c)  $(NH_4)_2[CoF_4]$

Central metal ion : Co.

Coordination number = 4.

We know,

Oxidation state is :

$$x - 4 = -2$$

$$x = +2$$

The d orbital occupation :  $e_g^4 t_{2g}^3$ .

(d)  $[Mn(H_2O)_6]SO_4$

Central metal ion : Mn.

Coordination number = 6.

We know,

Oxidation state is :

$$x + 0 = 2$$

$$x = +2$$

The d orbital occupation :  $t_{2g}^3 e_g^4$ .

**Q24.** State the IUPAC name for each of the complexes given below and provide their coordination number, oxidation state, and electronic configuration.

Also, indicate the stereochemistry and magnetic moment of these complexes :

- (a)  $K[Cr(H_2O)_2(C_2O_4)_2].3H_2O$
- (b)  $[Co(NH_3)_5Cl]Cl_2$
- (c)  $CrCl_3(py)_3$
- (d)  $Cs[FeCl_4]$
- (e)  $K_4[Mn(CN)_6]$

Ans:

(a) IUPAC name = Potassium diaquadioxalatochromate (III) trihydrate.

Coordination number = 6

Oxidation state of chromium :

$$x + 0 + 2(-2) = -1$$

$$\therefore x = +3$$

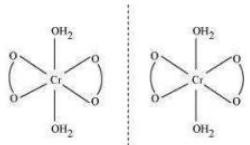
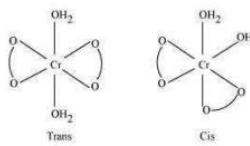
X = 3

Electronic configuration:  $3d^3 = t_{2g}^3$

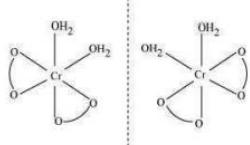
Shape : Octahedral

Stereochemistry :

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Trans is optically inactive



Cis is optically active

Magnetic moment,  $\mu = \sqrt{n(n+2)}$  [ n = unpaired electrons ]

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

$$= \sqrt{15} \approx 4BM$$

( b ) IUPAC name = Pentaamminechloridocobalt(III) chloride

Coordination number = 6

Oxidation state of Co :

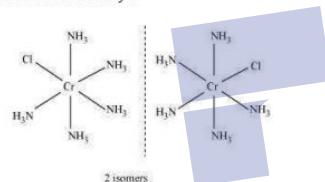
$$x + 0 - 1 = + 2$$

$$x = 3$$

Electronic configuration:  $3d^6 = t_{2g}^6$

Shape : Octahedral

Stereochemistry :



$$n = 0.$$

Thus, Magnetic moment = 0

( c ) IUPAC name = Trichloridotripyridinechromium (III)

Coordination number = 6

Oxidation state of Cr :

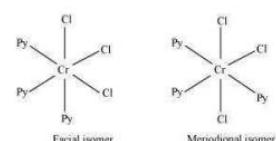
$$x - 3 + 0 = 0$$

$$x = 3$$

Electronic configuration:  $3d^3 = t_{2g}^3$

Shape : Octahedral

Stereochemistry :



$$n = 3$$

Magnetic moment,  $\mu = \sqrt{n(n+2)}$

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

$$= \sqrt{15} \approx 4BM$$

( d ) IUPAC name = Caesiumtetrachloroferrate (III)

Coordination number = 4

Oxidation state of Fe :

$$x - 4 = -1$$

$$x = 3$$

Electronic configuration:  $d^6 = e_g^2 t_{2g}^3$

Shape : Tetrahedral

Stereochemistry :- optically inactive

$$n = 5$$

Magnetic moment,  $\mu = \sqrt{n(n+1)} \approx 5BM$

Magnetic moment,  $\mu = \sqrt{n(n+4)}$

$$= \sqrt{5(5+2)}$$

$$= \sqrt{35} \approx 6 \text{ BM}$$

(e) IUPAC name = Potassium hexacyanomanganate(II)

Coordination number = 6

Oxidation state of Mn :

$$x - 6 = -4$$

$$x = +2$$

Electronic configuration:  $3d^5 = t_{2g}^5$

Shape : Octahedral

Stereochemistry :- optically inactive

$$n = 1$$

Magnetic moment,  $\mu = \sqrt{n(n+2)}$

$$= \sqrt{1(1+2)}$$

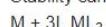
$$= \sqrt{3} = 1.732 \text{ BM}$$

**Q25. "Stability of a coordination compound in solution", what do you understand by it? What are the factors responsible for the stability of a complex?**

Ans:

Stability of a coordination compound in a solution is the degree/level of association among the species involved in a state of equilibrium.

Stability can also be written quantitatively in terms of formation constant or stability constant.



$$\text{Stability constant, } \beta = \frac{[ML_3]}{[M][L]^3}$$

Greater the value of  $\beta$ , stronger is the metal-ligand bond.

Factors responsible for the stability of a complex:

(1) Charge on the central metal ion – bigger the charge, more stable is the complex.

(2) Nature of ligand – chelating ligand produces a more stable complex.

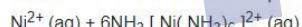
(3) The basic strength of ligand- more basic a ligand, more stable its complex.

**Q26. What is the chelate effect? Provide an example**

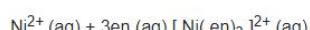
Ans:

When a polydentate or a didentate ligand fastens itself to a metal ion in such a way that it assumes the shape of a ring, the metal-ligand bond becomes more stable. These rings are called chelate rings.

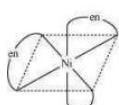
From here we can infer that complexes with chelate rings are more stable than complexes without the rings. This phenomenon is termed the chelate effect.



$$\log \beta = 7.99$$



$$\log \beta = 18.1 \text{ (more stable)}$$



**Q27. Write short notes on the role of coordination compounds in the following fields :**

(Also cite example in each case)

(a) biological system

(b) medicinal chemistry

(c) analytical chemistry

(d) metallurgy/extraction of metals

Ans:

(a) Role in biological systems:

In the body of animals, there are several very important coordination compounds e.g. hemoglobin is a coordination compound of iron.

In plants, the chlorophyll pigment is a coordination compound of magnesium.

(b) Role in medicinal chemistry:

So many coordinate compounds are used for curing purposes. For e.g., a coordination compound of platinum, cis-platin is used for checking the growth of tumors.

(c) Role in analytical chemistry:

Determination of hardness of water.

(d) Role in metallurgy or extraction:

During metal extraction from ores, complexes are formed. For e.g. gold combines with cyanide ions in an aqueous solution. Gold is then extracted from this complex using zinc.

**Q28. When you put the complex  $\text{Co}(\text{NH}_3)_6\text{Cl}_2$  in a solution, how many ions do you get?**

Ans:

The given complex  $[\text{Co}(\text{NH}_3)_6]\text{Cl}_2$  ionizes to give three ions, viz one  $[\text{Co}(\text{NH}_3)_6]^+$  and two  $\text{Cl}^-$  ions.

**Q29. From the ions given below which one has the largest value of magnetic moment ?**

- ( a )  $[\text{Cr}(\text{H}_2\text{O})_6]^{3+}$
- ( b )  $[\text{Fe}(\text{H}_2\text{O})_6]$
- ( c )  $[\text{Zn}(\text{H}_2\text{O})_6]^{2+}$

Ans:

- ( a )  $[\text{Cr}(\text{H}_2\text{O})_6]^{3+}$

number of unpaired electrons,  $n = 3$

$$\text{Magnetic moment, } \mu = \sqrt{3(3+2)} = \sqrt{3(3+2)}$$

$$= \sqrt{15} \approx 4 \text{ BM}$$

- ( b )  $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$

number of unpaired electrons,  $n = 4$

$$\text{Magnetic moment, } \mu = \sqrt{4(4+2)} = \sqrt{4(4+2)}$$

$$= \sqrt{24} \approx 5 \text{ BM}$$

- ( c )  $[\text{Zn}(\text{H}_2\text{O})_6]^2$

$n = 0$

Thus,  $[\text{Zn}(\text{H}_2\text{O})_6]^2$  has the highest magnetic moment value.

**Q30. What is the oxidation number of cobalt in  $\text{K}[\text{Co}(\text{CO})_4]$  ?**

Ans:



We know,

$$\Rightarrow x + 0 = -1 \quad [\text{Where } x \text{ is the oxidation number}]$$

$$x = -1$$

**Q31. Identify the most stable complex from among the following :**

- ( a )  $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$
- ( b )  $[\text{Fe}(\text{NH}_3)_6]^{3+}$
- ( c )  $[\text{Fe}(\text{C}_2\text{O}_4)_3]^{3-}$
- ( d )  $[\text{FeCl}_6]^{3-}$

Ans:

In all the cases Fe has an oxidation state of +3.  $(\text{C}_2\text{O}_4)_3$  is a didentate chelating ligand and it forms chelating rings. Thus, ( c ) is the most stable complex.

**Q32. Re arrange the following complexes in an ascending order of wavelengths they absorb from the visible spectrum :  $[\text{Ni}(\text{NO}_2)_6]^{4-}$ ,  $[\text{Ni}(\text{H}_2\text{O})_6]^{2+}$ ,  $[\text{Ni}(\text{NH}_3)_6]^{2+}$**

Ans:

All of the complexes have the same metal ion, so the energy absorption depends upon the CFSE values of the ligands. According to the spectro-chemical series, the CFSE values of the ligands are in the order of :  $\text{H}_2\text{O} < \text{NH}_3 < \text{NO}_2^-$

As,

$$E = hc / \lambda$$

$$\Rightarrow E \propto 1 / \lambda$$

Therefore, the values of the absorbed wavelength in an ascending order would be :

