

For B. Sc. II Year

Coordination Compounds

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Coordination compounds are series of compounds with at least one coordination bond. As we have already studied, difference between double salt and coordination compound is that those compounds which exist only in crystal state but dissociate in its constituents in solution state for eg.

Mohr's salt - $\text{FeSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$

Potash alum - $\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$

Carnallite - $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$

All the above mentioned salt are stable only in crystalline state but in solution dissociate into their constituent salts

Coordination compounds or complex compounds are those compounds which retain their identity even when dissolved in water or any other suitable solvent and their properties are completely different from the constituent salts. for eg.

1- $\{\text{Cu}(\text{NH}_3)_4\}\text{SO}_4$

2- $\text{K}_4\{\text{Fe}(\text{CN})_6\}$

Above mentioned (1) tetra ammine copper II sulphate exist both in crystalline and in solution state in the same form and its property is different from ammonia and sulphate ion. Similarly compound (2) also retain its identity in crystalline and solution state and it does not have property of potassium, iron and cyanide before we proceed further we should be familiar with some basic term of coordination compounds.

Coordination Complexes and Complex Ion- A coordination complex may be defined as "a compound that result from the combination of two or more stable chemical species and retain its identity in solid as well as liquid state." A water soluble complex compounds invariably contains a complex ion. Thus $\{\text{Cu}(\text{NH}_3)_4\}\text{SO}_4$ contain the complex ion $\{\text{Cu}(\text{NH}_3)_4\}^{2+}$ while $\text{K}[\text{Ag}(\text{CN})_2]$ contains the complex ion $[\text{Ag}(\text{CN})_2]^-$.

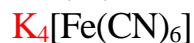
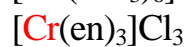
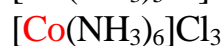
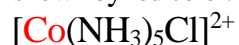
In coordination compounds there are two parts:

- Coordination sphere
- Ionisation sphere
- **Coordination sphere:** part of the coordination sphere possessing coordinate bond that is "The central ion together with molecule or ions coordinated to it constitute coordination sphere". In writing formulae of the complex compound coordination sphere is represented within square bracket [] for example $[\text{Fe}(\text{CN})_6]^{4-}$
- **Ionisation sphere:** part of the complex compound which ionizes when dissolved in a suitable solvent. During writing the formulae of the compound it is represented outside the square bracket shown in blue color. For example



- **Central metal ion**

In the complex ion an acceptor accepts a pair of electrons from the donor atoms. The acceptor is usually a metal / metal ion to which one (or) more of neutral molecules (or) anions are attached. The acceptor metal cation is referred to as central metal cation. Hence, central metal cation in a complex serves as a lewis acid. The cation to which one or more neutral molecules or ions are coordinated is called the central ion while the molecule or so attached are called ligands. Thus in the complex ion $[\text{Ni}(\text{NH}_3)_6]^{2+}$ the Ni^{2+} ion is the central ion and molecule of ammonia are the ligand. Central metal ions in some other complexes are shown by red color.



- **Coordination number**

The coordination number of a metal ion in a complex can be defined as the number of ligand donor atoms to which the metal is directly bonded. Numerically coordination number represents the total number of the chemical bonds formed between the central metal ion and the donor atoms of the ligands. For example in

$\text{K}_4[\text{Fe}(\text{CN})_6]$ the coordination number of Fe(II) is 6

and in $[\text{Cu}(\text{NH}_3)_4]\text{SO}_4$ the coordination number of Cu(II) is 4.

- **Charge on Complex Ion:** The charge carried by a complex ion is the algebraic sum of charges carried by the central metal ion and the ligands coordinated to it. Thus $[\text{Cu}(\text{NH}_3)_4]^{2+}$ carries a charge of +2 because Cu^{2+} carries a charge of +2 and ammonia is a neutral molecule, similarly $[\text{Fe}(\text{CN})_6]^{4-}$ carries a charge of 4- because Fe^{2+} carries a charge of +2 and six CN^- carries charge of 6-. A coordination compound some time may have no charge.

Example

1. $[\text{Cu}(\text{NH}_3)_4]^{2+}$ can be written as $[\text{Cu}^{2+}(\text{NH}_3)_4]^{2+}$ since NH_3 ligand is neutral. The sum of the charges on the metal ion and the ligands = +2.

This can be determined as shown below

Charge on the metal ion (Cu^{2+}) = +2

Charge on the ligand (NH_3) = 4 \times 0 = 0

Net charge on the complex ion = +2 + 0 = +2

2. Similarly for $[\text{Fe}(\text{CN})_6]^{4-}$ (or) $[\text{Fe}^{2+}(\text{CN})_6]^{4-}$

The sum of the charge on the metal ion and the ligand = -4.

Charge on the metal ion (Fe^{2+}) = +2

Charge on the ligand (CN^-) = 6 \times (-1) = -6

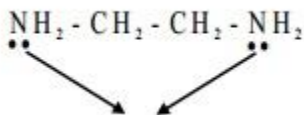
Net charge on the complex = +2 - 6 = -4

- **Ligand (Latin word meaning to bind)**

A ligand is an ion (or) a molecule capable of functioning as an electron donor. Therefore the neutral molecules or ions which are directly attached to the central metal ion are called as ligand (or) coordination groups. These coordination groups or ligands can donate a pair of electrons to the central metal ion (or) atom. Hence, in a complex compound ligands act as Lewis bases.

- **Types of ligands**

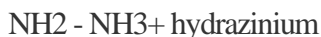
When a ligand is bound to a metal ion through a single donor atom, as with - Cl, H_2O or NH_3 , the ligand is said to be unidentate. Whenever a single coordinating group (or) ligand occupies two (or) more coordination position on the same central metal ions, a complex possessing a closed ring is formed. Such ligands are called polydentate ligands. When a single ligand has two coordinating positions, it is called bidentate ligand and when there are three coordinating positions available, it is called a tridentate ligand and so on. For example, ethylenediamine is a bidentate ligand because it has two amino groups each of which can donate a pair of electrons.



Name of the ligands

Positive ligands

The positive ligands are named with an ending -ium.



This ligand, though positive can bind through the uncharged nitrogen.

Neutral ligands

The neutral ligands are named as such without any special name. But water is written as 'aqua': Ammonia is written as ammine. Note that two m's to distinguish from organic amine CO-Carbonyl, NO-Nitrosyl, $\text{NH}_2 - \text{CH}_2 - \text{CH}_2 - \text{NH}_2$ -ethylenediamine (en), Pyridine $\text{C}_5\text{H}_5\text{N}$.

Negative Ligands

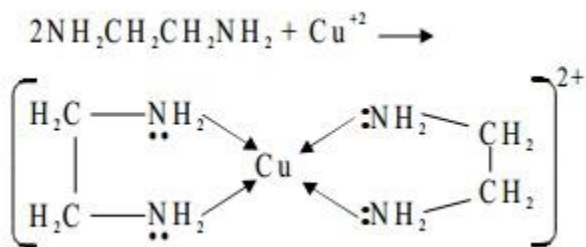
Negative ligands end in suffix 'O'.

Example

F--Fluoro, Cl--Chloro, $\text{C}_2\text{O}_4^{2-}$ --Oxalato, CN--Cyano, NO_2 --Nitro, Br--Bromo, SO_4^{2-} --Sulphato, CH_3COO^- --acetato CNS--thiocyanato, NCS--isothiocyanato, $\text{S}_2\text{O}_3^{2-}$ --thiosulphato.

Chelates

If a ligand is capable of forming more than one bond with the central metal atom (or) ion then the ring structures are produced which are known as metal chelates. Hence the ring forming groups are described as chelating agents (or) polydentate ligands.



- **Oxidation state**

This number denotes the charge, explaining the number of electrons it has lost to form the cation. It is oxidation number that denotes the charge, if the central metal atom would have if all the ligand in the complex were removed along with their electron pairs that were shared with the central atom. It is usually represented by Roman Numeral.

IUPAC NOMENCLATURE OF MONONUCLEAR COORDINATION COMPOUNDS

The rules are outlined below

1. In naming the entire complex, the name of the cation is given first and the anion second (just as for sodium chloride), no matter whether the cation or the anion is the complex species.
2. In the complex ion, the name of the ligand or ligands precedes that of the central metal atom (This procedure is reversed for writing formulae).
3. Ligand names generally end with 'O' if the ligand is negative ('chloro' for Cl^- , 'cyano' for CN^- , 'hydrido' for H^-) and unmodified if the ligand is neutral ('methylamine' for MeNH_2).
4. Special ligand names are 'aqua' for water, 'ammine' for ammonia, 'carbonyl' for CO , 'nitrosyl' for NO .
5. A Greek prefix (mono, di, tri, tetra, penta, hexa, etc.) indicates the number of each ligand (mono is usually omitted for a single ligand of a given type). If the name of the ligand itself contains the terms mono, di, tri, like triphenylphosphine, ligand name is enclosed in parentheses and its number is given with the alternate prefixes bis for 2, tris for 3 and tetrakis for 4.

For example, $[\text{Ni}(\text{PPh}_3)_2\text{Cl}_2]$ is named dichlorobis(triphenylphosphine) nickel(II).

6. A Roman numeral or a zero in parentheses is used to indicate the oxidation state of the central metal atom.
7. If the complex ion is negative, the name of the metal ends in 'ate' for example, ferrate, cuprate, nickelate, cobaltate etc.

8. If more than one ligand is present in the species, then the ligands are named in alphabetical order regardless of the number of each. For example, NH_3 (ammine) would be considered as 'a' ligand and come before Cl-(chloro).

Some additional notes

- i) Some metals in anions have special names

B Borate Au Aurate Ag Argentate Fe Ferrate
Pb Plumbate Sn Stannate Cu Cuprate Ni Nickelate

- ii) Use of brackets or enclosing marks.

Square brackets are used to enclose a complex ion or neutral coordination species.

Examples

$[\text{Co}(\text{en})_3]\text{Cl}_3$ - tris(ethylenediamine)cobalt(III) chloride

$[\text{Co}(\text{NH}_3)_3(\text{NO}_2)_3]$ triamminetrinitrocobalt (III)

$\text{K}_2[\text{CoCl}_4]$ potassiumtetrachlorocobaltate(II)

note that it is not necessary to enclose the halogens in brackets.

A systematic approach to drawing and naming coordination complexes

For the complex $\text{K}_3[\text{Cr}(\text{C}_2\text{O}_4)_3] \cdot 3\text{H}_2\text{O}$

Q1) What is the central metal ion?

A1) Central metal is Chromium

Q2) What is its oxidation state?

A2) O.S. is III

Q3) What is its electronic configuration?

A3) electronic configuration is d^3

Q4) What is its coordination number?

A4) C.N. is 6(3 bidentate ligands present)

Q5) What is the shape of the ion?

A5) structure is octahedral

Q6) Can the structure have isomers?

A6) Yes, optical isomers are possible

Q7) What is the IUPAC name of the complex?

A7) Potassiumtris(oxalato) chromate(III) trihydrate

IUPAC Nomenclature of mono nuclear coordination compounds

$[\text{Co}^{\text{III}}(\text{NH}_3)_5\text{Cl}]^{2+}$ pentaamminechlorocobalt(III) ion

$[\text{Co}^{\text{III}}(\text{NH}_3)_6]\text{Cl}_3$ - Hexaamminecobalt(III) chloride

$[\text{Cr}^{\text{III}}(\text{en})_3]\text{Cl}_3$ - Tris (ethylenediamine)chromium(III) chloride

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

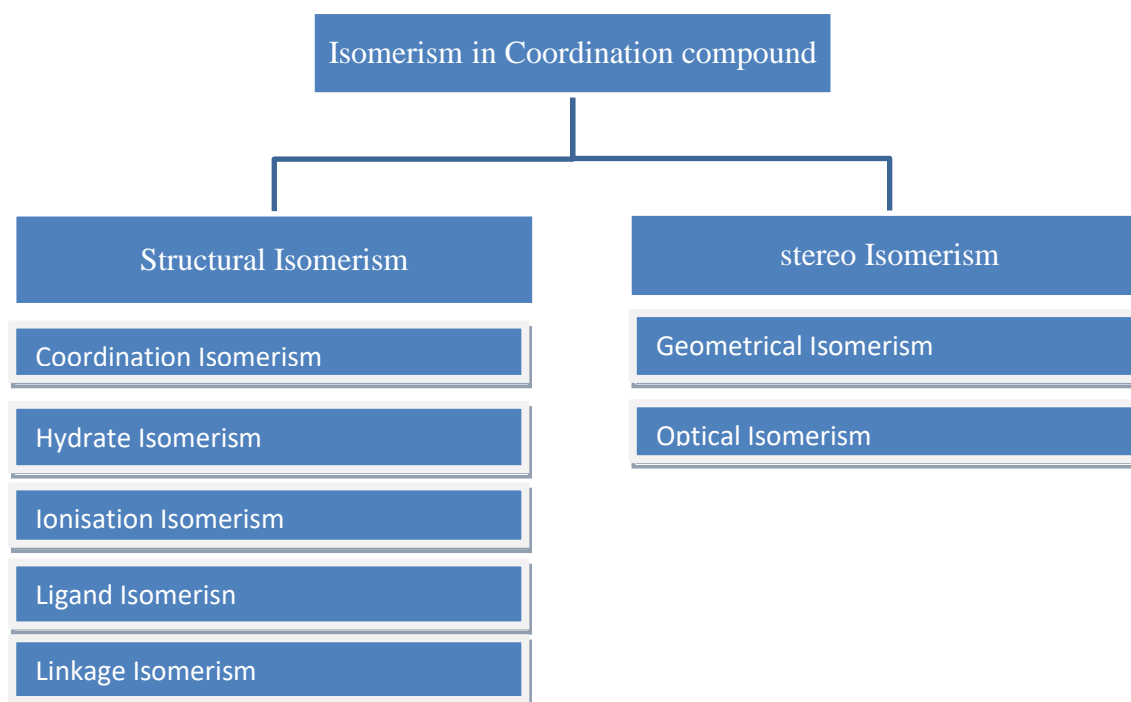
$[\text{Ni}^{\text{II}}(\text{CN})_4]^{2-}$ - Tetracyanonickelate(II) ion
 $[\text{Cu}^{\text{II}}(\text{NH}_3)_4]^{2+}$ - Tetraamminecopper(II) ion
 $[\text{Pt}^{\text{II}}\text{Cl}_2(\text{NH}_3)_2]$ - Diamminedichloroplatinum(II)

ISOMERISM

Compounds having the same chemical formula but different physical and chemical properties due to the different structural arrangements are called isomers. This phenomenon is known as isomerism.

Coordination compounds exhibit two major types of isomerism (A) structural isomerism and (B) stereoisomerism (space isomerism). Each of these is further classified as shown below.

-
- **Types of Isomerism:**
- Structural and
- Stereo isomerism



A) Structural isomerism a) Coordination isomerism c) Hydrate or Solvate isomerism e) Ligand Isomerism b) Ionisation isomerism d) Linkage isomerism B) Stereoisomerism a) Geometrical isomerism b) Optical isomerism

A) Structural isomerism

- a) Coordination isomerism
- c) Hydrate or Solvate isomerism
- e) Ligand Isomerism
- b) Ionisation isomerism
- d) Linkage isomerism

B) Stereoisomerism

- a) Geometrical isomerism
- b) Optical isomerism

1 A) Structural isomerism

a) Coordination isomerism

In a bimetallic complex, both complex cation and complex anion may be present. In such a case the distribution of ligands between the two coordination spheres can vary, giving rise to isomers called the coordination isomers. This phenomenon is called coordination isomerism. This isomerism is illustrated by the following pairs of complexes where the complex cation and anion contain different metal centres.

- | | |
|---|---|
| 1. $[\text{Co}^{\text{III}}(\text{NH}_3)_6]$ $[\text{Cr}(\text{CN})_6]$ and | $[\text{Cr}^{\text{III}}(\text{NH}_3)_6]$ $[\text{Co}^{\text{III}}(\text{CN})_6]$ |
| Hexamine hexacyano
cobalt(III) chromate(III) | Hexamine hexacyano
chromium (III) cobaltate (III) |
| 2. $[\text{Pt}^{\text{II}}(\text{NH}_3)_4]$ $[\text{CuCl}_4]$ and | $[\text{Cu}(\text{NH}_3)_4]$ $[\text{PtCl}_4]$ |
| Tetraammine Tetrachloro
platinum (II) cuparate (II) | Tetraammine Tetrachloro
copper (II) platinate (II) |

b) Ionisation isomerism

Coordination compounds having the same molecular formula but forming different ions in solution are called ionisation isomers. This property is known as ionisation isomerism.

An example of this type of isomerism is furnished by the red-violet,

$[\text{Co}(\text{NH}_3)_5\text{Br}]\text{SO}_4$	$[\text{Co}(\text{NH}_3)_5\text{SO}_4]\text{Br}$
pentaamminebromocobalt(III) sulphate	pentaamminesulphatocobalt (III) bromide

The red-violet isomer yields sulphate ion and the red isomer furnishes bromide ion in solution.

$[\text{Co}(\text{NH}_3)_4\text{Cl}_2]\text{NO}_2$ and	$[\text{Co}(\text{NH}_3)_4\text{NO}_2\text{Cl}]\text{Cl}$
Tetraamminedichlorocobalt(III) nitrite	Tetraamminechloronitrocobalt(III) chloride

$[\text{Co}(\text{NH}_3)_5\text{NO}_3]\text{SO}_4$ and	$[\text{Co}(\text{NH}_3)_5\text{SO}_4]\text{NO}_3$
pentaamminenitratocobalt(III) sulphate	pentaamminesulphatocobalt(III) nitrate

c) Hydrate isomerism or Solvate isomerism

The best known examples of this type of isomerism occurs for chromium chloride " $\text{CrCl}_3 \cdot 6\text{H}_2\text{O}$ " which may contain 4, 5, (or) 6 coordinated water molecules.

1. $[\text{Cr}(\text{H}_2\text{O})_4\text{Cl}_2]\text{Cl} \cdot 2\text{H}_2\text{O}$ - Bright green
Tetraaquadichlorochromium(III) chloride dihydrate
2. $[\text{Cr}(\text{H}_2\text{O})_5\text{Cl}]\text{Cl}_2 \cdot \text{H}_2\text{O}$ - grey-green
Pentaaquachlorochromium(III) chloride monohydrate
3. $[\text{Cr}(\text{H}_2\text{O})_6]\text{Cl}_3$ - Violet
Hexaaquachromium(III) chloride

These isomers have very different chemical properties and on reaction with AgNO₃ to test for Cl⁻ ions, would find 1, 2, and 3 Cl⁻ ions in solution respectively.

d. Linkage isomerism

Linkage isomerism occurs with ambidentate ligands. These ligands are capable of coordinating in more than one way. The best known cases involve the monodentate ligands SCN⁻/NCS⁻ and NO₂⁻/ONO⁻

For example

[Co(NH₃)₅ONO]Cl₂ the nitrito isomer - red colour
pentaamminenitritocobalt(III) chloride - O attached

[Co(NH₃)₅NO₂]Cl₂ the nitro isomer - yellow colour
pentaamminenitrocobalt(III) chloride - N attached

e) Ligand isomerism

Ligand isomerism arises from the presence of ligands which can adopt different isomeric forms. An example is provided by diaminopropane, which may have the amine groups in the terminal (1,3-) positions or in the 1,2-positions.

H₂N - CH₂ - CH₂ - CH₂ - NH₂

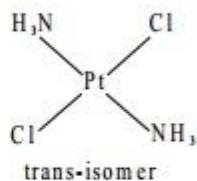
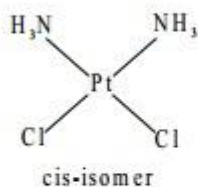
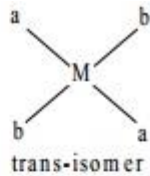
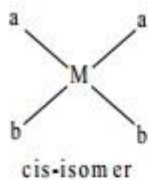
2 Stereoisomerism (space isomerism)

Consider two compounds containing the same ligands attached to the same central metal ion, but the arrangement of ligands in space about the central metal ion are different, then these two compounds are said to be stereoisomers and this phenomenon is known as stereoisomerism. There are two different types of stereoisomerism. a) Geometrical isomerism b) Optical isomerism.

a) Geometrical (or) cis-trans isomerism

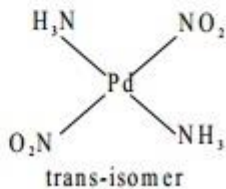
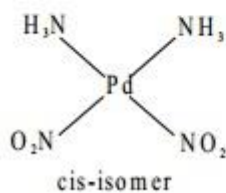
Geometric isomers are possible for both square planar and octahedral complexes, but not tetrahedral. In a cis-isomer two identical (or) similar groups are adjacent to each other whereas in a trans-isomer they are diametrically opposite to each other.

Square planar complexes of the type [Ma₂b₂]^{n±} where a and b are monodentate ligands, exist as cis and trans-isomers as shown below. Example of this type of complexes are [Pt (NH₃)₂ Cl₂] and [Pd(NH₃)₂ (NO₂)₂]. The cis- trans isomers of these compounds are represented as

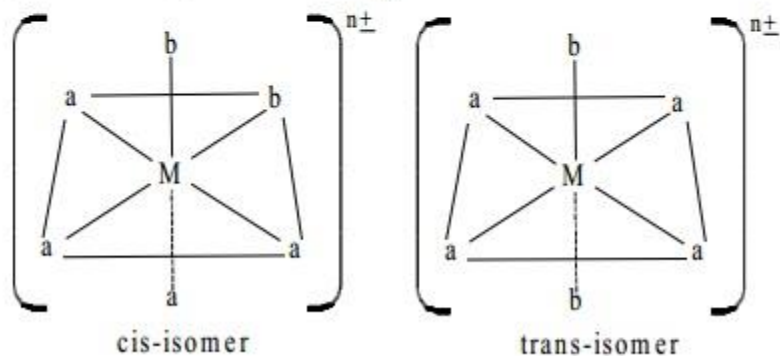
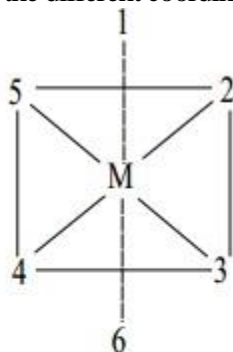


cis-diamminedichloroplatinum(II)

trans-diamminedichloroplatinum(II)



In the octahedral complex, the different coordination positions are numbered as shown below



Along the twelve edges of the octahedron, there are twelve cis positions.

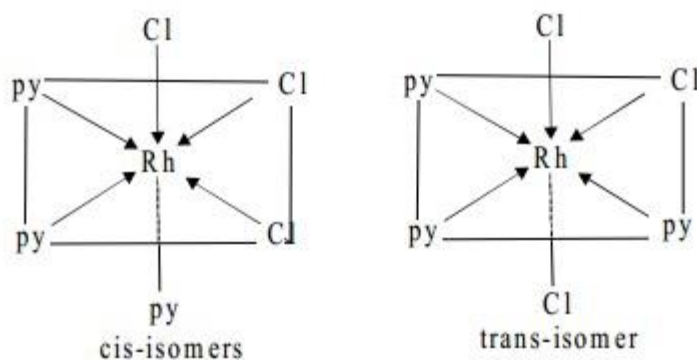
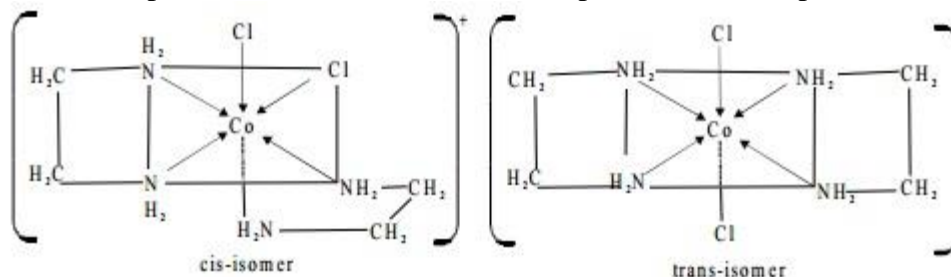
They are (1,2) (1,3) (1,4) (1,5) (2,6) (3,6) (4,6) (5,6) (3,4) (4,5) (2,5) and (2,3). In order to avoid confusion, generally it is assumed that the 1,2 positions are cis-positions. There are three trans positions; they are (1,6) (2,4) and (3,5). Normally it is taken that 1,6 positions are trans-positions in order to avoid confusion.

An octahedral complex of the type $[Ma_4b_2]$ where a and b are monodentate ligands, exists as two geometrical isomers:

A specific example for such Isomerism is $[Co(NH_3)_4 Cl_2]^+$ which exists as two geometrical isomers.

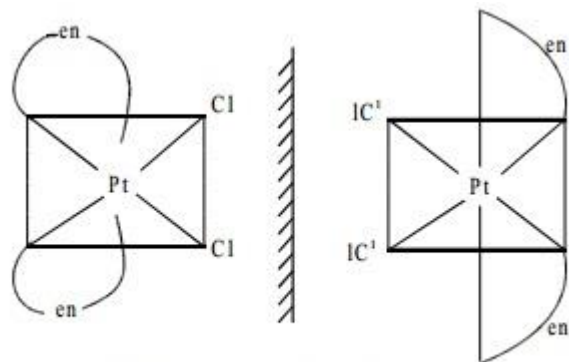
The octahedral complex are of the type $[M(AA)_2a_2]n^{\pm}$ where (AA) is a symmetrical bidentate ligand such as ethylenediamine $H_2N-CH_2-CH_2-NH_2$ and 'a' is a monodentate ligand. A specific example for this is $[Co(H_2N-CH_2-CH_2- NH_2)_2 Cl_2]^+$

The octahedral complex of the type, $[Ma_3b_3]$ in square planar complexes, where a and b are monodentate ligands also exist as geometrical isomers, For example, $[Rh(py)_3 Cl_3]$ exist as cis-(1,2,3 trichloro complex) and trans-(1,2,6-trichloro complex) isomers as represented below

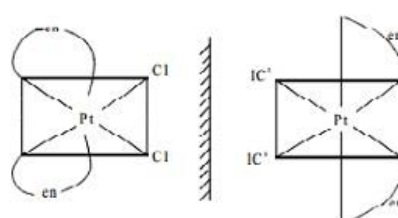
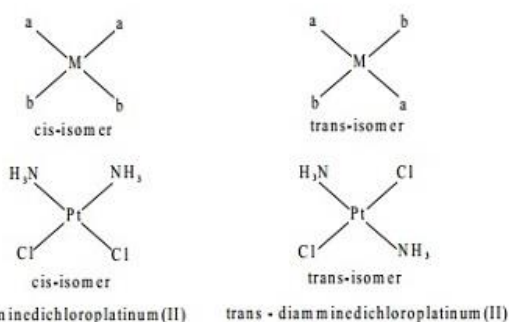


b) Optical Isomerism

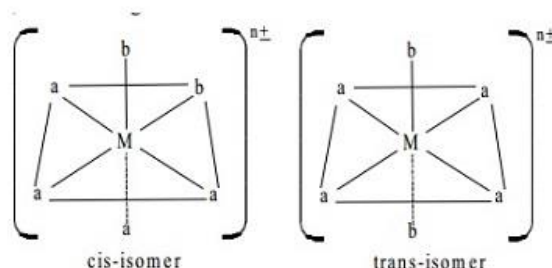
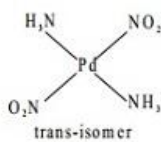
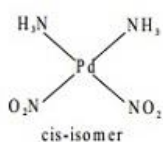
This is a phenomenon in which certain organic or inorganic compounds have the property of rotating plane polarised light. The compounds which exhibit this property are called optical isomers. The optical isomers of a compound have identical physical and chemical properties. The only distinguishing property is that the isomers rotate the plane of polarised light either to the left or right. In a coordination compound of type $[PtCl_2(en)_2]^{2+}$, two geometrical isomers are possible. They are cis and trans. Among these two isomers, cis isomer shows optical activity because the whole molecule is asymmetric.



Optical isomers of cis $(PtCl_2(en)_2)^{2+}$



Optical isomers of cis $(PtCl_2(en)_2)^{2+}$



Werners Theory of Coordination Compounds

One of the most important properties of metallic elements is their ability to act as Lewis acids that form complexes with a variety of Lewis bases. A metal complex consists of a central metal atom or ion that is bonded to one or more [ligands](#) (from the Latin ligare, meaning “to bind”), which are ions or molecules that contain one or more pairs of electrons that can be shared with the metal. Metal complexes can be neutral, such as $Co(NH_3)_3Cl_3$; positively charged, such as $[Nd(H_2O)_9]^{3+}$; or negatively charged, such as $[UF_8]^{4-}$. Electrically charged metal complexes are sometimes called *complex ions*. A *coordination compound* contains one or more metal complexes.

Werner, the son of a factory worker, was born in Alsace. He developed an interest in chemistry at an early age, and he did his first independent research experiments at age 18. While doing his military service in southern Germany, he attended a series of chemistry lectures, and he subsequently received his PhD at the University of Zurich in Switzerland, where he was appointed professor of chemistry at age 29. He won the Nobel Prize in Chemistry in 1913 for his

work on coordination compounds, which he performed as a graduate student and first presented at age 26. Apparently, Werner was so obsessed with solving the riddle of the structure of coordination compounds that his brain continued to work on the problem even while he was asleep. In 1891, when he was only 25, he woke up in the middle of the night and, in only a few hours, had laid the foundation for modern coordination chemistry.

Table 24.1.124.1.1: Werner's Data on Complexes of Ammonia with PtCl₄PtCl₄

Complex	Conductivity (ohm ⁻¹)	Number of Ions per Formula Unit	Number of Cl ⁻ Ions Precipitated by AgNO ₃
PtCl ₄ ·6NH ₃	523	5	4
PtCl ₄ ·5NH ₃	404	4	3
PtCl ₄ ·4NH ₃	299	3	2
PtCl ₄ ·3NH ₃	97	2	1
PtCl ₄ ·2NH ₃	0	0	0

The important postulates of Werner's Coordination Theory are-

(i) In Coordination Compounds the metal atom exhibit two types of valency, viz. Primary Valency, Secondary Valency.

The Primary Valency is ionisable and nondirectional whereas the secondary valency is nonionisable and directional. The primary valency corresponds to oxidation state and the secondary valency corresponds to the coordination number.

(ii) Every metal atom has a fixed number of secondary valency, i.e, it has a fixed coordination number.

(iii) The metal atom tends to satisfy both its primary and its secondary valencies.

Primary valencies are satisfied by negative ions whereas secondary valencies are satisfied by either negative or neutral or positive ligands.

(iv) The secondary valencies are always directed towards fixed position in space and this leads to definite geometry.

E.g., If a metal ion has six secondary valencies, these are arranged in octahedral manner around the central metal ion. If the metal ion has four secondary valency, these are arranged in either tetrahedral or square planar arrangement around the central metal ion. The secondary valencies thus determine the stereo chemistry of the compound.

Valence Bond Theory

It was proposed by G. N. Lewis Important postulates of valence bond theory are as follows:

(i) The suitable number of atomic orbitals of central metal ion (s, p, d) hybridise to provide empty hybrid orbitals.

(ii) These hybrid orbitals accept lone pair of electrons from the ligands and are directed towards the ligand positions according to the geometry of the complex.

(iii) When inner d-orbitals i.e. (n-1) d orbitals are used in hybridization, the complex is called – *inner orbital* or *low spin* or *hyperligated* complex.

(iv) A substance which do not contain any unpaired electron is not attracted by 2 magnet. It is said to be diamagnetic. On the other hand, a substance which contains one or more unpaired electrons in the electrons in the d-orbitals, is attracted by a magnetic field [exception O₂ and NO]. It is said to be paramagnetic.

Paramagnetism can be calculated by the expression, $\mu_s = \sqrt{n(n + 2)}$ where μ = magnetic moment.

s= spin only value and *n*= number of unpaired electrons.

Hence, if $n = 1$, $\mu_s = \sqrt{1(1 + 2)} = 1.73$ B.M, if $n = 3$ $\mu_s = \sqrt{3(3 + 2)} = 3.87$ B.M and so on

On the basis of value of magnetic moment, we can predict the number of unpaired electrons present in the complex. If we know the number of unpaired electrons in the metal complex, then it is possible to predict the geometry of the complex species.

(v) There are two types of ligands namely strong field and weak field ligands. A strong field ligand is capable of forcing the electrons of the metal atom/ion to pair up (if required). Pairing is done only to the extent which is required to cause the hybridization possible for that Co-ordination number. A weak field ligand is incapable of making the electrons of the metal atom/ ion to pair up.

Strong field ligands : CN⁻, CO, en, NH₃, H₂O, NO⁻, Py.

Weak field ligands : I⁻, Br⁻, Cl⁻, F⁻, NO₃⁻, OH⁻, C₂O₄²⁻, NH₃, H₂O

Geometry (shape) and magnetic nature of some of the complexes (Application of valence bond theory)

Atom/ion/ complex (1)	Configuration (2)	Oxidation state of metal (3)	Type of hybridi- zation (4)	Geometry shape (5)	No. of unpaired electrons (6)	Magnetic nature (7)
$\text{Ni}^{2+} (d^8)$		+2			2	Paramagnetic
$[\text{NiCl}_4]^{2-}$		+2	sp^3	Tetrahedral	2	Paramagnetic
$[\text{Ni}(\text{CN})_4]^{2+}$	<p>Rearrangement dsp^2</p>	+2	dsp^2	Square planar	0	Diamagnetic
Ni		0			2	Paramagnetic
$\text{Ni}(\text{CO})_4$	<p>Rearrangement sp^3</p>	0	sp^3	Tetrahedral	0	Diamagnetic
$[\text{Ni}(\text{NH}_3)_6]^{2+}$		+2	sp^3d^2 (Outer)	Octahedral	2	Paramagnetic
$\text{Mn}^{2+} (d^5)$		+2			5	Paramagnetic
$[\text{Mn}(\text{CN})_6]^{4-}$	<p>Rearrangement d^2sp^3</p>	+2	d^2sp^3 (Inner)	Octahedral	1	Paramagnetic
$[\text{MnCl}_4]^{2-}$		+2	sp^3	Tetrahedral	5	Paramagnetic
$\text{Cu}^{2+} (d^9)$		+2			1	Paramagnetic
$[\text{CuCl}_4]^{2-}$		+2	sp^3	Tetrahedral	1	Paramagnetic
$[\text{Cu}(\text{NH}_3)_4]^{2+}$	<p>One electron is shifted from 3d- to 4p-orbital</p>	+2	dsp^2	Square planar	1	Paramagnetic

(1)	(2)	(3)	(4)	(5)	(6)	(7)
$\text{Cr}^{3+} (d^3)$		+3			3	Paramagnetic
$[\text{Cr}(\text{NH}_3)_6]^{3+}$		+3	d^2sp^3 (Inner)	Octahedral	3	Paramagnetic
$[\text{Cr}(\text{H}_2\text{O})_6]^{3+}$		+3	sp^3d^2 (Outer)	Octahedral	3	Paramagnetic
$\text{Co}^{3+} (d^6)$		+3			4	Paramagnetic
$[\text{CoF}_6]^{3-}$		+3	sp^3d^2 (Outer)	Octahedral	4	Paramagnetic
$[\text{Co}(\text{NH}_3)_6]^{3+}$		+3	d^2sp^3 (Inner)	Octahedral	0	Diamagnetic
$\text{Co}^{2+} (d^7)$		+2			3	Paramagnetic
$[\text{Co}(\text{H}_2\text{O})_6]^{2+}$		+2	sp^3d^2 (Outer)	Octahedral	3	Paramagnetic
$\text{Fe}^{2+} (d^6)$		+2			4	Paramagnetic
$[\text{Fe}(\text{CN})_6]^{4-}$		+2	d^2sp^3 (Inner)	Octahedral	0	Diamagnetic
$[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$		+2	sp^3d^2 (Outer)	Octahedral	4	Paramagnetic
$[\text{Fe}(\text{NH}_3)_6]^{2+}$	Same	+2	sp^3d^2 (Outer)	Octahedral	4	Paramagnetic
$\text{Fe}^{3+} (d^5)$		+3			5	Paramagnetic
$[\text{Fe}(\text{CN})_6]^{3-}$		+3	d^2sp^3 (Inner)	Octahedral	1	Paramagnetic
Fe		0			4	Paramagnetic
$\text{Fe}(\text{CO})_5$		0	dsp^3 (Inner)	Trigonal bipyramidal	0	Diamagnetic

