Coordination complex

![Cisplatin, PtCl₂(NH₃)₂](image)

A platinum atom with four ligands

In chemistry, a **coordination complex** or **metal complex** consists of a central atom or ion, which is usually metallic and is called the **coordination centre**, and a surrounding array of bound molecules or ions, that are in turn known as **ligands** or complexing agents. Many metal-containing compounds, especially those of transition metals, are coordination complexes.

**Nomenclature and terminology**

Coordination complexes are so pervasive that the structure and reactions are described in many ways, sometimes confusingly. The atom within a ligand that is bonded to the central atom or ion is called the **donor atom**. In a typical complex, a metal ion is bound to several donor atoms, which can be the same or different. Polydentate (multiple bonded) ligands consist of several donor atoms, several of which are bound to the central atom or ion. These complexes are called chelate complexes, the formation of such complexes is called chelation, complexation, and coordination.

The central atom or ion, together with all ligands comprise the coordination sphere. The central atoms or ion and the donor atoms comprise the first coordination sphere.

**Coordination** refers to the "coordinate covalent bonds" (dipolar bonds) between the ligands and the central atom. Originally, a complex implied a reversible association of molecules, atoms, or ions through such weak chemical bonds. As applied to coordination chemistry, this meaning has evolved. Some metal complexes are formed virtually irreversibly and many are bound together by bonds that are quite strong.

The number of ligands attached to a metal ion is called the **coordination number**. The most common coordination numbers are 6, 4, 2. A hydrated ion is one kind of a
complex ion (or, simply, complex), a species formed between a central metal ion and one or more surrounding ligands, molecules or ions that contain at least one lone pair of electrons,

For example, Cobalt(II) hexahydrated ion or Hexaaquacobalt(II) ion \([\text{Co(H}_2\text{O)}_6]^{2+}\), is a hydrated-complex ion that consists of 6 ligands attached to a metal ion Co. The oxidation state and the coordination number reflects the number of bonds formed between the metal ion and the ligands in the complex ion.

**History**

Coordination complexes have been known since the beginning of modern chemistry. Early well-known coordination complexes include dyes such as Prussian blue. Their properties were first well understood in the late 1800s, following the 1869 work of Christian Wilhelm Blomstrand. Blomstrand developed what has come to be known as the complex ion chain theory. The theory claimed that the reason coordination complexes form is because in solution, ions would be bound via ammonia chains. He compared this effect to the way that various carbohydrate chains form.

![Alfred Werner](https://example.com/alfred-werner.jpg)

Following this theory, Danish scientist Sophus Mads Jorgensen made improvements to it. In his version of the theory, Jorgensen claimed that when a molecule dissociates in a solution there were two possible outcomes: the ions would bind via the ammonia chains Blomstrand had described or the ions would bind directly to the metal.

It was not until 1893 that the most widely accepted version of the theory today was published by Alfred Werner. Werner’s work included two important changes to the Blomstrand theory. The first was that Werner described the two different ion possibilities in terms of location in the coordination sphere. He claimed that if the ions were to form a chain this would occur outside of the coordination sphere while the ions that bound directly to the metal would do so within the coordination sphere.\[8\] In one of Werner’s most important discoveries however he disproved the majority of the chain theory. Werner was able to discover the spatial arrangements of the ligands that were involved in the formation of the complex hexacoordinate cobalt. His theory
allows one to understand the difference between coordinated and ionic in a compound, for example chloride in the cobaltammine chlorides and to explain many of the previously inexplicable isomers.

In 1914, Werner first resolved the coordination complex, called hexol, into optical isomers, overthrowing the theory that only carbon compounds could possess chirality.

**Structures**

The ions or molecules surrounding the central atom are called ligands. Ligands are generally bound to the central atom by a coordinate covalent bond (donating electrons from a lone electron pair into an empty metal orbital), and are said to be **coordinated** to the atom. There are also organic ligands such as alkenes whose pi bonds can coordinate to empty metal orbitals. An example is ethene in the complex known as Zeise's salt, $\text{K}^+\left[\text{PtCl}_3(\text{C}_2\text{H}_4)\right]^-$. 

**Geometry**

In coordination chemistry, a structure is first described by its coordination number, the number of ligands attached to the metal (more specifically, the number of donor atoms). Usually one can count the ligands attached, but sometimes even the counting can become ambiguous. Coordination numbers are normally between two and nine, but large numbers of ligands are not uncommon for the lanthanides and actinides. The number of bonds depends on the size, charge, and electron configuration of the metal ion and the ligands. Metal ions may have more than one coordination number.

Typically the chemistry of transition metal complexes is dominated by interactions between s and p molecular orbitals of the ligands and the d orbitals of the metal ions. The s, p, and d orbitals of the metal can accommodate 18 electrons. The maximum coordination number for a certain metal is thus related to the electronic configuration of the metal ion (to be more specific, the number of empty orbitals) and to the ratio of the size of the ligands and the metal ion. Large metals and small ligands lead to high coordination numbers, e.g. $[\text{Mo(CN)}_8]^{4-}$. Small metals with large ligands lead to low coordination numbers, e.g. $\text{Pt}[\text{P(CMe}_3)]_2$. Due to their large size, lanthanides, actinides, and early transition metals tend to have high coordination numbers.
Different ligand structural arrangements result from the coordination number. Most structures follow the points-on-a-sphere pattern (or, as if the central atom were in the middle of a polyhedron where the corners of that shape are the locations of the ligands), where orbital overlap (between ligand and metal orbitals) and ligand-ligand repulsions tend to lead to certain regular geometries. The most observed geometries are listed below, but there are many cases that deviate from a regular geometry, e.g. due to the use of ligands of different types (which results in irregular bond lengths; the coordination atoms do not follow a points-on-a-sphere pattern), due to the size of ligands, or due to electronic effects:

- Linear or bent for two-coordination
- Trigonal planar or trigonal pyramidal for three-coordination
- Tetrahedral or square planar for four-coordination
- Trigonal bipyramidal or square pyramidal for five-coordination
- Octahedral (orthogonal) or trigonal prismatic for six-coordination
- Pentagonal bipyramidal, capped octahedral or capped trigonal prismatic for seven-coordination
- Square antiprismatic or dodecahedral for eight-coordination
- Tri-capped trigonal prismatic (Triaugmented triangular prism) or capped square antiprismatic for nine-coordination.

Due to special electronic effects such as (second-order) Jahn–Teller stabilization, certain geometries are stabilized relative to the other possibilities, e.g. for some compounds the trigonal prismatic geometry is stabilized relative to octahedral structures for six-coordination.

**Isomerism**

The arrangement of the ligands is fixed for a given complex, but in some cases it is mutable by a reaction that forms another stable isomer.

There exist many kinds of isomerism in coordination complexes, just as in many other compounds.

**Stereoisomerism**

Stereoisomerism occurs with the same bonds in different orientations relative to one another. Stereoisomerism can be further classified into:

Cis–trans isomerism occurs in octahedral and square planar complexes (but not tetrahedral). When two ligands are adjacent they are said to be **cis**, when opposite each other, **trans**. When three identical ligands occupy one face of an octahedron, the isomer is said to be facial, or **fac**. In a **fac** isomer, any two identical ligands are
adjacent or *cis* to each other. If these three ligands and the metal ion are in one plane, the isomer is said to be meridional, or *mer*. A *mer* isomer can be considered as a combination of a *trans* and a *cis*, since it contains both trans and cis pairs of identical ligands.

\[
\text{cis-}[\text{CoCl}_2(\text{NH}_3)_4]^+ \\
\text{trans-}[\text{CoCl}_2(\text{NH}_3)_4]^+ \\
\text{fac-}[\text{CoCl}_3(\text{NH}_3)_3] \\
\text{mer-}[\text{CoCl}_3(\text{NH}_3)_3]
\]
Optical isomerism occurs when a molecule is not superimposable with its mirror image. It is so called because the two isomers are each optically active, that is, they rotate the plane of polarized light in opposite directions. The symbol Λ (lambda) is used as a prefix to describe the left-handed propeller twist formed by three bidentate ligands, as shown. Likewise, the symbol Δ (delta) is used as a prefix for the right-handed propeller twist.
Structural isomerism

Structural isomerism occurs when the bonds are themselves different. There are four types of structural isomerism: ionisation isomerism, solvate or hydrate isomerism, linkage isomerism and coordination isomerism.

1. **Ionisation isomerism** – the isomers give different ions in solution although they have the same composition. This type of isomerism occurs when the counter ion of the complex is also a potential ligand. For example, pentaamminebromidocobalt(III)sulfate \([\text{CoBr(NH}_3)_5\text{SO}_4]\) is red violet and in solution gives a precipitate with barium chloride, confirming the presence of sulfate ion, while pentaamminesulfatecobalt(III)bromide \([\text{CoSO}_4(\text{NH}_3)_5\text{Br}]\) is red and tests negative for sulfate ion in solution, but instead gives a precipitate of AgBr with silver nitrate.

2. **Solvate or hydrate isomerism** – the isomers have the same composition but differ with respect to the number of solvent ligand molecules as well as the counter ion in the crystal lattice. For example, \([\text{Cr(H}_2\text{O})_6]\text{Cl}_3\) is violet colored, \([\text{CrCl(H}_2\text{O})_5]\text{Cl}_2\cdot\text{H}_2\text{O}\) is blue-green, and \([\text{CrCl}_2(\text{H}_2\text{O})_4]\text{Cl}\cdot\text{2H}_2\text{O}\) is dark green.

3. **Linkage isomerism** occurs with ambidentate ligands that can bind in more than one place. For example, NO\(_2\) is an ambidentate ligand: It can bind to a metal at either the N atom or an O atom.

4. **Coordination isomerism** – this occurs when both positive and negative ions of a salt are complex ions and the two isomers differ in the distribution of ligands between the cation and the anion. For example, \([\text{Co(NH}_3)_6][\text{Cr(CN)}_6]\) and \([\text{Cr(NH}_3)_6][\text{Co(CN)}_6]\)

### Electronic properties

Many of the properties of transition metal complexes are dictated by their electronic structures. The electronic structure can be described by a relatively ionic model that ascribes formal charges to the metals and ligands. This approach is the essence of crystal field theory (CFT). Crystal field theory, introduced by Hans Bethe in 1929, gives a quantum mechanically based attempt at understanding complexes. But crystal field theory treats all interactions in a complex as ionic and assumes that the ligands can be approximated by negative point charges.

More sophisticated models embrace covalency, and this approach is described by ligand field theory (LFT) and Molecular orbital theory (MO). Ligand field theory, introduced in 1935 and built from molecular orbital theory, can handle a broader
range of complexes and can explain complexes in which the interactions are covalent. The chemical applications of group theory can aid in the understanding of crystal or ligand field theory, by allowing simple, symmetry based solutions to the formal equations.

Chemists tend to employ the simplest model required to predict the properties of interest; for this reason, CFT has been a favorite for the discussions when possible. MO and LF theories are more complicated, but provide a more realistic perspective.

The electronic configuration of the complexes gives them some important properties:

**Color of transition metal complexes**

![Image of copper(II)-tetraphenylporphyrin](image)

Synthesis of copper(II)-tetraphenylporphyrin, a metal complex, from tetraphenylporphyrin and copper(II) acetate monohydrate.

Transition metal complexes often have spectacular colors caused by electronic transitions by the absorption of light. For this reason they are often applied as pigments. Most transitions that are related to colored metal complexes are either d–d transitions or charge transfer bands. In a d–d transition, an electron in a d orbital on the metal is excited by a photon to another d orbital of higher energy. A charge transfer band entails promotion of an electron from a metal-based orbital into an empty ligand-based orbital (Metal-to-Ligand Charge Transfer or MLCT). The converse also occurs: excitation of an electron in a ligand-based orbital into an empty metal-based orbital (Ligand to Metal Charge Transfer or LMCT). These phenomena can be observed with the aid of electronic spectroscopy; also known as UV-Vis. For simple compounds with high symmetry, the d–d transitions can be assigned using Tanabe–Sugano diagrams. These assignments are gaining increased support with computational chemistry.
### Colours of Various Example Coordination Complexes

<table>
<thead>
<tr>
<th>Hydrated Ion</th>
<th>Fe$^{2+}$</th>
<th>Fe$^{3+}$</th>
<th>Co$^{2+}$</th>
<th>Cu$^{2+}$</th>
<th>Al$^{3+}$</th>
<th>Cr$^{3+}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>[Fe(H$_2$O)$_6$]$^{2+}$</td>
<td>Pale green Solution</td>
<td>Yellow/brown Solution</td>
<td>Pink Solution</td>
<td>Blue Solution</td>
<td>Colourless Solution</td>
<td>Green Solution</td>
</tr>
<tr>
<td>OH$^-$, dilute</td>
<td>[Fe(H$_2$O)$_4$(OH)$_2$]$^{3+}$</td>
<td>Dark green Precipitate</td>
<td>Brown Precipitate</td>
<td>Blue/green Precipitate</td>
<td>Blue Precipitate</td>
<td>White Precipitate</td>
</tr>
<tr>
<td>OH$^-$, concentrated</td>
<td>[Fe(H$_2$O)$_4$(OH)$_2$]$^{3+}$</td>
<td>Dark green Precipitate</td>
<td>Brown Precipitate</td>
<td>Blue/green Precipitate</td>
<td>Blue Precipitate</td>
<td>White Precipitate</td>
</tr>
<tr>
<td>NH$_3$, dilute</td>
<td>[Fe(H$_2$O)$_4$(NH$_3$)$_2$]$^{2+}$</td>
<td>Dark green Precipitate</td>
<td>Brown Precipitate</td>
<td>Straw coloured Solution</td>
<td>Deep blue Solution</td>
<td>White Precipitate</td>
</tr>
<tr>
<td>NH$_3$, concentrated</td>
<td>[Fe(H$_2$O)$_4$(NH$_3$)$_2$]$^{2+}$</td>
<td>Dark green Precipitate</td>
<td>Brown Precipitate</td>
<td>Straw coloured Solution</td>
<td>Deep blue Solution</td>
<td>White Precipitate</td>
</tr>
<tr>
<td>CO$_3^{2-}$</td>
<td>FeCO$_3$</td>
<td>[Fe(H$_2$O)$_3$(OH)$_2$]$^{2+}$</td>
<td>Brown Precipitate + bubbles</td>
<td>CoCO$_3$</td>
<td>Pink Precipitate</td>
<td>Blue/green Precipitate</td>
</tr>
<tr>
<td></td>
<td></td>
<td>CoCO$_3$</td>
<td>Pink Precipitate</td>
<td>CuCO$_3$</td>
<td>Blue/green Precipitate</td>
<td></td>
</tr>
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### Colors of Lanthanide complexes

Superficially lanthanide complexes are similar to those of the transition metals in that some are coloured. However for the common Ln$^{3+}$ ions (Ln = lanthanide) the colors are all pale, and hardly influenced by the nature of the ligand. The colors are due to 4f electron transitions. As the 4f orbitals in lanthanides are “buried” in the xenon core and shielded from the ligand by the 5s and 5p orbitals they are therefore not influenced by the ligands to any great extent leading to a much smaller crystal field splitting than in the transition metals. The absorption spectra of an Ln$^{3+}$ ion approximates to that of the free ion where the electronic states are described by spin-orbit coupling (also called L-S coupling or Russell-Saunders coupling). This contrasts to the transition metals where the ground state is split by the crystal field. Absorptions for Ln$^{3+}$ are weak as electric dipole transitions are parity forbidden (Laporte Rule forbidden) but can gain intensity due to the effect of a low-symmetry ligand field or mixing with higher electronic states (e.g. d orbitals). Also absorption bands are extremely sharp which contrasts with those observed for transition metals which
generally have broad bands. This can lead to extremely unusual effects, such as significant color changes under different forms of lighting.

**Magnetism**

Metal complexes that have unpaired electrons are magnetic. Considering only monometallic complexes, unpaired electrons arise because the complex has an odd number of electrons or because electron pairing is destabilized. Thus, monomeric Ti(III) species have one "d-electron" and must be (para)magnetic, regardless of the geometry or the nature of the ligands. Ti(II), with two d-electrons, forms some complexes that have two unpaired electrons and others with none. This effect is illustrated by the compounds TiX₂[(CH₃)₂PCH₂CH₂P(CH₃)₂]₂: when X = Cl, the complex is paramagnetic (high-spin configuration), whereas when X = CH₃, it is diamagnetic (low-spin configuration). It is important to realize that ligands provide an important means of adjusting the ground state properties.

In bi- and polymetallic complexes, in which the individual centers have an odd number of electrons or that are high-spin, the situation is more complicated. If there is interaction (either direct or through ligand) between the two (or more) metal centers, the electrons may couple (antiferromagnetic coupling, resulting in a diamagnetic compound), or they may enhance each other (ferromagnetic coupling). When there is no interaction, the two (or more) individual metal centers behave as if in two separate molecules.

**Reactivity**

Complexes show a variety of possible reactivities:

- **Electron transfers**

  A common reaction between coordination complexes involving ligands are inner and outer sphere electron transfers. They are two different mechanisms of electron transfer redox reactions, largely defined by the late Henry Taube. In an inner sphere reaction, a ligand with two lone electron pairs acts as a bridging ligand, a ligand to which both coordination centres can bond. Through this, electrons are transferred from one centre to another.

- **(Degenerate) ligand exchange**

  One important indicator of reactivity is the rate of degenerate exchange of ligands. For example, the rate of interchange of coordinate water in [M(H₂O)₆]ⁿ⁺ complexes varies over 20 orders of magnitude. Complexes where the ligands are released and rebound rapidly are classified as labile. Such labile complexes can be quite stable thermodynamically. Typical labile metal
complexes either have low-charge (Na\(^+\)), electrons in d-orbitals that are antibonding with respect to the ligands (Zn\(^{2+}\)), or lack covalency (Ln\(^{3+}\), where Ln is any lanthanide). The lability of a metal complex also depends on the high-spin vs. low-spin configurations when such is possible. Thus, high-spin Fe(II) and Co(III) form labile complexes, whereas low-spin analogues are inert. Cr(III) can exist only in the low-spin state (quartet), which is inert because of its high formal oxidation state, absence of electrons in orbitals that are M–L antibonding, plus some "ligand field stabilization" associated with the d\(^3\) configuration.

- Associative processes

Complexes that have unfilled or half-filled orbitals often show the capability to react with substrates. Most substrates have a singlet ground-state; that is, they have lone electron pairs (e.g., water, amines, ethers), so these substrates need an empty orbital to be able to react with a metal centre. Some substrates (e.g., molecular oxygen) have a triplet ground state, which results that metals with half-filled orbitals have a tendency to react with such substrates (it must be said that the dioxygen molecule also has lone pairs, so it is also capable to react as a 'normal' Lewis base).

If the ligands around the metal are carefully chosen, the metal can aid in (stoichiometric or catalytic) transformations of molecules or be used as a sensor.

**Classification**

Metal complexes, also known as coordination compounds, include all metal compounds, aside from metal vapors, plasmas, and alloys. The study of "coordination chemistry" is the study of "inorganic chemistry" of all alkali and alkaline earth metals, transition metals, lanthanides, actinides, and metalloids. Thus, coordination chemistry is the chemistry of the majority of the periodic table. Metals and metal ions exist, in the condensed phases at least, only surrounded by ligands.

The areas of coordination chemistry can be classified according to the nature of the ligands, in broad terms:

- Classical (or "Werner Complexes"): Ligands in classical coordination chemistry bind to metals, almost exclusively, via their "lone pairs" of electrons residing on the main group atoms of the ligand. Typical ligands are H\(_2\)O, NH\(_3\), Cl\(^-\), CN\(^-\), en. Some of the simplest members of such complexes are described in metal aquo complexes, metal ammine complexes,
Examples: [Co(EDTA)]\(^{-}\), [Co(NH\(_3\))\(_6\)]Cl\(_3\), [Fe(C\(_2\)O\(_4\))\(_3\)]K\(_3\)

- Organometallic Chemistry: Ligands are organic (alkenes, alkynes, alkyls) as well as "organic-like" ligands such as phosphines, hydride, and CO.

Example: (C\(_5\)H\(_5\))Fe(CO)\(_2\)CH\(_3\)

- Bioinorganic Chemistry: Ligands are those provided by nature, especially including the side chains of amino acids, and many cofactors such as porphyrins.

Example: hemoglobin contains heme, a porphyrin complex of iron
Example: chlorophyll contains a porphyrin complex of magnesium
Many natural ligands are "classical" especially including water.

- Cluster Chemistry: Ligands are all of the above also include other metals as ligands.

Example Ru\(_3\)(CO)\(_{12}\)

- In some cases there are combinations of different fields:

Example: [Fe\(_4\)S\(_4\)(Scysteinyl)\(_4\)]\(^{2-}\), in which a cluster is embedded in a biologically active species.

Mineralogy, materials science, and solid state chemistry – as they apply to metal ions – are subsets of coordination chemistry in the sense that the metals are surrounded by ligands. In many cases these ligands are oxides or sulfides, but the metals are coordinated nonetheless, and the principles and guidelines discussed below apply. In hydrates, at least some of the ligands are water molecules. It is true that the focus of mineralogy, materials science, and solid state chemistry differs from the usual focus of coordination or inorganic chemistry. The former are concerned primarily with polymeric structures, properties arising from a collective effects of many highly interconnected metals. In contrast, coordination chemistry focuses on reactivity and properties of complexes containing individual metal atoms or small ensembles of metal atoms.

Older classifications of isomerism

Traditional classifications of the kinds of isomer have become archaic with the advent of modern structural chemistry. In the older literature, one encounters:

- Ionisation isomerism describes the possible isomers arising from the exchange between the outer sphere and inner sphere. This classification relies on an archaic classification of the inner and outer sphere. In this classification, the
"outer sphere ligands," when ions in solution, may be switched with "inner sphere ligands" to produce an isomer.

- Solvation isomerism occurs when an inner sphere ligand is replaced by a solvent molecule. This classification is obsolete because it considers solvents as being distinct from other ligands. Some of the problems are discussed under water of crystallization.

**Naming complexes**

The basic procedure for naming a complex:

1. When naming a complex ion, the ligands are named before the metal ion.
2. Write the names of the ligands in alphabetical order. (Numerical prefixes do not affect the order.)
   - Multiple occurring monodentate ligands receive a prefix according to the number of occurrences: di-, tri-, tetra-, penta-, or hexa. Polyydentate ligands (e.g., ethylenediamine, oxalate) receive bis-, tris-, tetrakis-, etc.
   - Anions end in indo. This replaces the final 'e' when the anion ends with '-ate', e.g. sulfate becomes sulfato. It replaces 'ide': cyanide becomes cyanido.
   - Neutral ligands are given their usual name, with some exceptions: NH$_3$ becomes ammine; H$_2$O becomes aqua or aqo; CO becomes carbonyl; NO becomes nitrosyl.
3. Write the name of the central atom/ion. If the complex is an anion, the central atom's name will end in -ate, and its Latin name will be used if available (except for mercury).

<table>
<thead>
<tr>
<th>Ligands</th>
<th>Coding name</th>
<th>Old name</th>
</tr>
</thead>
<tbody>
<tr>
<td>H$_2$O</td>
<td>aqua</td>
<td>Aquo</td>
</tr>
<tr>
<td>NH$_3$</td>
<td>ammine</td>
<td>ammino</td>
</tr>
<tr>
<td>OH$^-$</td>
<td>hydroxo</td>
<td>hydroxy</td>
</tr>
<tr>
<td>Cl$^-$</td>
<td>chlorido</td>
<td>Chloro</td>
</tr>
<tr>
<td>F$^-$</td>
<td>fluorido</td>
<td>Fluoro</td>
</tr>
<tr>
<td>CN$^-$</td>
<td>cyanido</td>
<td>Cyano</td>
</tr>
</tbody>
</table>

1. If the central atom's oxidation state needs to be specified (when it is one of several possible, or zero), write it as a Roman numeral (or 0) in parentheses.
2. Name cation then anion as separate words (if applicable, as in last example)
Examples:

<table>
<thead>
<tr>
<th>Metal</th>
<th>changed to</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cobalt</td>
<td>cobaltate</td>
</tr>
<tr>
<td>aluminium</td>
<td>aluminate</td>
</tr>
<tr>
<td>chromium</td>
<td>chromate</td>
</tr>
<tr>
<td>vanadium</td>
<td>vanadate</td>
</tr>
<tr>
<td>Copper</td>
<td>cuprate</td>
</tr>
<tr>
<td>Iron</td>
<td>ferrate</td>
</tr>
</tbody>
</table>

\[
\text{[NiCl}_4\text{]}^{2-} \rightarrow \text{tetrachloridonickelate(II) ion}
\]

\[
\text{[CuCl}_5\text{NH}_3\text{]}^{3-} \rightarrow \text{amminepentachloridocuprate(II) ion}
\]

\[
\text{[Cd(CN)_2(en)_2]} \rightarrow \text{dicyanidobis(ethylenediamine)cadmium(II)}
\]

\[
\text{[CoCl(NH}_3\text{)}_5\text{]}^{\text{SO}_4} \rightarrow \text{pentaamminechloridocobalt(III) sulfate}
\]

The coordination number of ligands attached to more than one metal (bridging ligands) is indicated by a subscript to the Greek symbol μ placed before the ligand name. Thus the dimer of aluminium trichloride is described by \( \text{Al}_2\text{Cl}_4(\mu_2\text{-Cl})_2 \).

**Stability Constant**

The affinity of metal ions for ligands is described by stability constant. This constant, also referred to as the formation constant, is given the notation of \( K_f \) and can be calculated through the following method for simple cases:

\[
(X)\text{Metal}_{(aq)}+(Y)\text{Lewis Base}_{(aq)} = (Z)\text{Complex Ion}_{(aq)}
\]

\[
K_f = \frac{[\text{Complex Ion}]^Z}{[\text{Metal}]^X[\text{Lewis Base}]^Y}
\]

Formation constant vary widely. Large values indicate that the metal has high affinity for the ligand, provided the system is at equilibrium.

Sometimes the stability constant will be in a different form known as the constant of destability. This constant is expressed as the inverse of the constant of formation and is denoted as \( K_d = 1/K_f \). This constant represents the reverse reaction for the decomposition of a complex ion into its individual metal and ligand components. When comparing the values for \( K_d \), the larger the value is the more unstable the complex ion is.
As a result of these complex ions forming in solutions they also can play a key role in solubility of other compounds. When a complex ion is formed it can alter the concentrations of its components in the solution. For example:

$$\text{Ag}^{+} + 2\text{NH}_4\text{OH} = \text{Ag(NH}_3)_2^+ + \text{H}_2\text{O}$$

$$\text{AgCl} + \text{H}_2\text{O} = \text{Ag}^+ + \text{Cl}^- + \text{H}_2\text{O}$$

In these reactions which both occurred in the same reaction vessel, the solubility of the silver chloride would be increased as a result of the formation of the complex ion. The complex ion formation is favorable takes away a significant portion of the silver ions in solution, as a result the equilibrium for the formation of silver ions from silver chloride will shift to the right to make up for the deficit.

This new solubility can be calculated given the values of $K_f$ and $K_{sp}$ for the original reactions. The solubility is found essentially by combining the two separate equilibria into one combined equilibrium reaction and this combined reaction is the one that determines the new solubility. So $K_c$, the new solubility constant, is denoted by $K_c = K_{sp} * K_f$.

### Application of coordination compounds

Metals only exist in solution as coordination complexes, it follows then that this class of compounds are useful. Coordination compounds are found both in the natural world and artificially in industry. Some common complex ions include such substances as vitamin B12, hemoglobin, chlorophyll, and some dyes and pigments. One major use of coordination compounds is in homogeneous catalysis for the production of organic substances.

Coordination compounds have uses in both nature and in industry. Coordination compounds are vital to many living organisms. For example many enzymes are metal complexes, like carboxypeptidase, a hydrolytic enzyme important in digestion. This enzyme consists of a zinc ion surrounded by many amino acid residues. Another complex ion enzyme is catalase, which decomposes the cell waste hydrogen peroxide. This enzyme contains iron-porphyrin complexes, similar to that in hemoglobin. Chlorophyll contains a magnesium-porphyrin complexes, and vitamin B12 is a complex with cobalt and corrin.

Coordination compounds are also widely used in industry. The intense colors of many compounds render them of great use as dyes and pigments. Specifically Phthalocyanine complexes are an important class of dyes for fabrics. Nickel, cobalt, and copper can be extracted using hydrometallurgical processes.
involving complex ions. They are extracted from their ores as ammine complexes with aqueous ammonia. Metals can also be separated using the selective precipitation and solubility of complex ions, as explained in later sections. Cyanide complexes are often used in electroplating.

Coordination compounds can also be used to identify unknown substances in a solution. This analysis can be done by utilizing the selective precipitation of the complex ions, the formation of color complexes which can be measured spectrophotometrically, or the preparation of complexes, such as metal acetylacetonates, which can be separated with organic solvents.

A combination of titanium trichloride and triethylaluminum brings about the polymerization of organic compounds with carbon-carbon double bonds to form polymers of high molecular weight and ordered structures. Many of these polymers are of great commercial importance because they are used in common fibers, films, and plastics.

Other common uses of coordination compounds in industry include the following:

1. They are used in photography, i.e., AgBr forms a soluble complex with sodium thiosulfate in photography.
2. K[Ag(CN)₂] is used for electroplating of silver, and K[Au(CN)₂] is used for gold plating.
3. Some ligands oxidise Co²⁺ to Co³⁺ ion.
4. Ethylenediaminetetraacetic acid (EDTA) is used for estimation of Ca²⁺ and Mg²⁺ in hard water.
5. Silver and gold are extracted by treating zinc with their cyanide complexes.