Bonds in introductory chemistry are typically classified according to whether they are ionic or covalent in character. **Coordinate covalent bonds** are a third classification. In this type of bond, a lone pair of electrons from one chemical species is donated to an empty orbital on another chemical species to form the new bond. This is different from a covalent bond because both electrons come from one atom or molecule but are shared as in a typical covalent bond. Unlike an ionic bond, a coordinate covalent bond does not rely on formal electrostatic attraction between a cation and an anion to form. Ammonia-borane complexes are an example of this type of bonding (Figure 1).

![Figure 1](image1.png)

**Figure 1**: Ammonia-borane complexes contain a coordinate covalent bond where the lone pair of electrons on ammonia (NH$_3$) are donated to an empty $p$ orbital on borane (BH$_3$) to give NH$_3$BH$_3$.

These complexes are generally referred to as donor-acceptor complexes where ammonia would represent the donor and borane would represent the acceptor. Transition metal coordination complexes can be thought of in much the same way except that the transition metal is the acceptor and the molecules with lone pairs of electrons are referred to as ligands. Ligands can be neutral molecules (sometimes referred to as L type ligands), anionic (sometimes referred to as X type ligands), or in some cases even cationic. The ligand electrons are donated into $d$ orbitals on the metal to form the new coordinate covalent bonds (Figure 2).

![Figure 2](image2.png)

**Figure 2**: Formation of coordinate covalent bonds in transition metal complexes. M represents a transition metal.

Ligands can be **monodentate** or **multidentate**. Monodentate ligands like ammonia would coordinate to the metal via one coordination point while a multidentate ligand would coordinate to the metal via two or more points. Ethylene diamine (NH$_2$CH$_2$CH$_2$NH$_2$) is an example of a bidentate ligand because it coordinates to the metal through each of the two amine groups (-NH$_2$).

The number of coordination points in a complex is referred to as the **coordination number**. Coordination numbers of six and four are the most common and result in geometries that are octahedral (6), tetrahedral (4), or square planar (4) (Figure 3).

![Figure 3](image3.png)

**Figure 3**: Common geometries of coordination complexes with coordination numbers of 6 and 4.

Coordination numbers and geometries vary dramatically, but are dependent on the choice of metal and ligand (in some cases the geometry of the complex can even change depending on variations in variables like temperature). There are many trends that correlate coordination number and geometry with the type of metal and ligand, but explaining these trends in detail is beyond the scope of this primer.

Two common reactions encountered for transition metal coordination complexes are **aquation** and **ligand substitution** reactions. Dissolution of a transition metal salt in water can be thought of as an aquation reaction. In this type of reaction water ligands coordinate to the metal when the solid salt is placed in aqueous solution. For example, dissolution of NiCl$_2$ might be written in shorthand form as:

\[
\text{NiCl}_2(s) + 2\text{H}_2\text{O} \rightarrow \text{Ni}^{2+}(aq) + 2\text{Cl}^{-}(aq)
\] (eq. 1)
More information would be provided, however, if the equation for the reaction above were written as shown below.

\[ \text{NiCl}_2(s) + 6\text{H}_2\text{O}(l) \rightarrow [\text{Ni(H}_2\text{O)}_6]^{2+}(aq) + 2\text{Cl}^-(aq) \quad \text{(eq. 2)} \]

Here the water molecules that coordinate to the nickel (II) cation in aqueous solution are part of the **inner coordination sphere** meaning that they bond directly with the metal. The chloride anions are part of the **outer coordination sphere**. They do not coordinate directly to the metal and serve only to balance the positive charge on the hexaaqua nickel (II) complex.

A ligand substitution reaction occurs when a ligand in the inner coordination sphere of a metal is displaced by another ligand resulting in the formation of a new coordination complex. An example of such a reaction is shown below.

\[ [\text{Ni(H}_2\text{O)}_6]^{2+}(aq) + \text{NH}_3(aq) \rightarrow [\text{Ni(H}_2\text{O)}_5(\text{NH}_3)]^{2+}(aq) + \text{H}_2\text{O}(l) \quad \text{(eq.3)} \]

In the reaction above, a total of six successive ligand substitution reactions would need to occur to replace all six water molecules with amine ligands to form \([\text{Ni(NH}_3)_6]^{2+}\).

One of the readily apparent properties of many transition metal complexes is that they exhibit a wide range of colors. The color we perceive for a transition metal complex is the complementary color of the one on the color wheel that is being absorbed (Figure 4).

For example, if a molecule absorbs wavelengths of light that are blue in color, all other wavelengths of light will be transmitted but we will perceive that molecule as being orange.

Recall that when a molecule absorbs light it does so in discrete or *quantized* energies which correspond to its possible electronic transitions. In a coordination complex the energy of those electronic transitions are directly related to the ligands attached to the metal and to its geometry. If the ligands on a metal or its geometry are changed, the energies of the various electronic transitions in the molecule are altered and thus the wavelengths of light that are absorbed by the molecule also change. Carrying out a ligand substitution reaction changes the energies of the electronic transitions in the coordination complex and is therefore often accompanied by a change in color.

**Figure 4**: The color wheel. Perceived color of a coordination complex is complementary to the wavelength of light absorbed.