Oxidation – Reduction Reactions

These types of reactions are very common in our nature; extraction metals and non-metals from their ores, rusting of common metals, burning of fossil fuels, action of common household bleaches, etc

Oxidation-reduction reactions, or simply known as **redox** (derived from reduction and oxidation) **reactions** are electron-transfer reactions. There are two definitions of oxidation and reaction, old and new (current).

According to old definition, the oxidation is the process of combing with oxygen and reduction is the process of removing the oxygen. Some times, you hear the word "oxidized"; it means combining with oxygen.

The current definitions of oxidation and reduction are based on the transfer of electrons. *Oxidation is the process of losing the electrons while the reduction is the process of gaining the electrons*. Consider the example burning of magnesium ribbon (Mg). When magnesium bums, it combines with oxygen (O_2) from the air to form magnesium oxide (MgO) according to the following equation:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Magnesium oxide is an ionic compound containing Mg^{+2} and O^{-2} ions whereas Mg(s) and O_2(g) are elements with no charges. Then, how the Mg(s) with zero charge becomes Mg with +2 charges (Mg^{+2}) going from reactant side to product side? How the O_2 (g) with zero charge becomes -2 charge going from left side to right side? This is explained as follows. When Mg(s) becomes Mg^{+2}, it loses 2 electrons. Since there are 2 Mg on left side, a total of 4 electrons are lost according to the following equation.

$2Mg(s) \rightarrow 2Mg^{+2} + 4e^-$ oxidation process

This process is known as the oxidation or oxidation process. On the other hand, O_2 becomes O^{-2} by gaining 4 electrons according to

$O_2(g) + 4\ e^- \rightarrow 2O^{-2}$ reduction process

This process is known as reduction or reduction process. Note that each oxygen atom gains 2 electrons. Since there are 2 oxygen atoms, a total of 4 electrons are gained. The above each process is called **half-reaction,** i.e., **oxidation half-reaction and reduction half-reaction.** Overall reaction is the sum of both half-reactions:

$2Mg(s) \rightarrow 2Mg^{+2} + 4e^-$ oxidation process
$O_2(g) + 4\ e^- \rightarrow 2O^{-2}$ reduction process

$2\ Mg(s) + O_2(g) + 4\ e^- \rightarrow 2Mg^{+2} + 2O^{-2} + 4\ e^-$

When chemical reaction, especially, redox reaction takes place, we do not see the electrons as they appear and disappear during the course of the reaction. What we see is the reactants (starting material) and end products. Due to this, electrons appearing on both sides of the equation are canceled. After concealing, the equation is re-written as

$$2\ Mg(s) + O_2(g) \rightarrow 2Mg^{+2} + 2O^{-2} \rightarrow 2MgO(s) \qquad \text{oxidation-reduction reaction}$$

Two ions, positive (Mg^{+2}) and negative (O^{-2}) exist on product side and they combine immediately to form a compound magnesium oxide (MgO) due to their opposite charges (electrostatic attraction).

In any given oxidation-reduction reaction, there are two half-reactions – oxidation half-reaction and reduction half-reaction. The sum of these two half-reactions is the oxidation-reduction reaction.

In the above example, Mg loses the electrons because of the presence of O_2, i.e., O_2 makes the Mg to lose the electrons. Therefore, O_2 is the reason why Mg loses the electrons. Hence O_2 is known as the **oxidizing agent**. At the same time, O_2 gains electrons because Mg is willing to donate electrons to O_2. Thus Mg is the reason for O_2 to accept electrons. Hence Mg is known as the **reducing agent.** The term "reducing" means reducing the charge (decrease in charge going from reactant side to product side). Here Mg reduces the O_2 by supplying electrons. Then the oxidizing and reducing agents are defined as follows:

The oxidizing agent is the substance that is being reduced or the substance that accepts electrons. The reducing agent is the substance that is being oxidized or the substance that donates electrons.

Important: The definition of oxidizing and reducing agents apply only to substances present on the reactant side.

Example

In the following redox reaction, identify oxidation half-reaction, reduction half-reaction, oxidizing agent, and reducing agent.

$$3Mg(s) + N_2\ (g) \rightarrow Mg_3N_2$$

Answer

First, we need to assign oxidation numbers to each element on both sides of the equation. The Mg(s) and N_2(g) are elemental forms on the left side and hence carry no charges. On the product side, Mg carries +2 charges because it belongs to Group IIA. Based on this, we can assign -3 charges to N on the product side. Thus, Mg goes from zero charge to +2 charges by losing 2 electrons or a total of 6 electrons. On the other hand, N_2(g) becomes N_2^{-3} by gaining 6 electrons. The half-reactions are written as

$2\ Mg(s) \rightarrow 3Mg^{+2} + 6\ e^-$	oxidation half-reaction
$N_2(g) + 6\ e^- \rightarrow N_2^{-3}$	reduction half-reaction

The $N_2(g)$ is the oxidizing agent because it is the substance being reduced(accepts electrons) and Mg(s) is the reducing agent because it is the substance being oxidized(looses electrons).
