4. CORROSION BASICS

Corrosion is the deterioration of a metal through a reaction with its environment.

Typically, corrosion involves contact of the metal with moisture and air (oxygen). The deterioration that takes place during the corrosion process is the basic tendency of the metal to revert to its natural state prior to it being developed from its primary ore material. The most common example of corrosion is the rusting of steel into iron oxide, its primary ore as found in the earth.

The corrosion process of metals is an electrochemical process involving a transfer of electrons from the metal's surface to ions in the environment (the electrolyte).

Corrosion is often described in electrical terms as an electrical circuit consisting of four key components. These include an anode (where corrosion or chemical oxidation reactions are taking place), a cathode (where chemical reduction reactions are taking place), an electrolyte (solutions or conductive media providing the supply of chemicals needed to sustain the cathodic reactions at or near the metal surface), and a metallic path.

4.1 Corrosion Terms

The anode is where corrosion or chemical oxidation reactions are taking place. These are the locations on the metal surface where electrical current is being passed by chemical means from the metals surface to ions, elements, or compounds in the electrolyte. The metal is losing electrons and combining with other elements in the environments by means of an oxidation reaction.

The cathode is where chemical reduction reactions are taking place. These are the locations on the metal surface where electrical current is being passed by chemical means from ions, elements, or compounds in the electrolyte to the surface of the metal. This chemical reaction is a reduction reaction.

ANODE	**CATHODE**
An electrode where oxidation reactions (corrosion) occur.	An electrode where reduction occurs.

The anode and cathode in a corrosion process may be on two different metals connected together forming a bimetallic couple, or as they may be on the same piece of metal. An area on a metal surface (or with two dissimilar metals, each respective metal) either becomes an anode or a cathode depending on the electrical potential of one area relative to the other. The electrical potential difference is the electromotive force of the cell and is the voltage difference between the anode and cathode of the cell. In any electrochemical cell, the area that is more negative in potential will undergo corrosion if coupled to the more positive area. The corrosion process is initiated by differences in the natural potential between the two dissimilar metals (bimetallic couple); metallurgical variations in the state at different points on the surface a single metal; or localized variations in the environment such as variations in moisture content or oxygen concentration.

The electrolyte is the material in contact with both the anode and the cathode that allows ions to migrate between the two electrodes. This allows ionic current flow to occur between the anode and the cathode. The electrolyte includes the source of atoms, elements, or compounds required for ionic current flow to and from the metal electrodes. The electrolyte is the environment that the metal is in contact with including whatever salts and liquids are present (e.g., soil for a buried metal or concrete for reinforcing steel).

The metallic path completes the electrical circuit and allows electrons to flow from the anode to the cathode in the electrochemical cell.

ELECTROLYTE	METALLIC PATH
Soil and/or liquid or other conductive media adjacent to and in contact with the anode and the cathode that allows ions to migrate.	Any conductor that allows electrons to flow.

4.2 Electrochemical Equations

The corrosion process can be described through a series of anodic and cathodic reaction equations (electrochemical equations) depending on the metal and ionic species involved.

$$M_o \longrightarrow M+^{n+} \quad + \quad ne-$$
$$\text{(metal)} \longrightarrow \text{(ion)} \quad + \quad \text{(electron)}$$

Where M_o represents a metal atom such as iron in a metallic structure; the arrow indicates the direction that reaction is occurring; the symbol M+ represents a metal ion; and n+ along with ne- indicates the number of electrons involved in the chemical reaction.

A common example is that of iron corrosion:

At the anode, iron is oxidized to the ferrous state, releasing electrons.

$$Fe \longrightarrow Fe^{++} \quad + \quad 2e-$$
$$\text{(iron)} \longrightarrow \text{(iron ion)} \quad + \quad \text{(2 electrons)}$$

Iron gives up two electrons, which pass through the metal from an anodic site to a cathodic site where they are consumed in a cathodic (reduction) reaction. The cathodic reaction is in the form:

$$\underset{\text{(ion)}}{R^{+}} + \underset{\text{(electron)}}{e^{-}} \longrightarrow \underset{\text{(reduced species)}}{R^{o}}$$

Where R^{+} represents a positive ion in solution, e^{-} indicates the number of electrons gained, the arrow indicates the direction that reaction is occurring; and R^{o} is the reduced atom produced as a result of the chemical reaction.

Two common examples of cathodic reactions are the reduction of hydrogen and oxygen:

$$\underset{\text{(hydrogen ions in solution)}}{2H^{+}} + \underset{\text{(electron)}}{2e^{-}} \longrightarrow \underset{\text{(reduced species)}}{H_2}$$

$$\underset{\text{(oxygen in solution)}}{O_2} + \underset{\text{(water)}}{2H_2O} + \underset{\text{(electrons)}}{4e^{-}} \longrightarrow \underset{\text{(reduced species)}}{4OH^{-}}$$

4.3 Electromotive Force Series and Galvanic Series

From the previous discussion, there is a transfer of electrons that occurs during the corrosion process. The example shown was that of steel in the presence of hydrogen and oxygen (both present as dissolved species in water). Similar electrochemical reactions can occur regardless of the conductive contact medium (soil, water, concrete, etc.). The standard that has been established to provide a reference scale of electrical potential values for metals and non-metals is known as the electromotive force (EMF) series or REDOX potential series.

Another useful series is the galvanic series. The galvanic series is the electrical potential of a metal or alloy in a particular electrolytic solution (usually seawater). It is useful in determining how two or more metals will behave if electrically connected in an ionically conductive environment. The electrical potential of each metal is measured as before against a standard. A reference electrode is simply an electrode (electrical conductor) that has a stable and well know electrical potential. The high stability of the electrode potential is usually reached by employing a redox system with constant (buffered or saturated) concentrations of each participant of the redox reactions.

Examples of reference electrodes are the copper/copper sulfate ($Cu/CuSO_4$) reference electrode, saturated calomel electrode (SCE) reference, and the silver/silver chloride (Ag/AgCl) reference electrode.

In terms of corrosion, when looking at metals in an electrical potential series such as an electrical motive force series or a galvanic series, when two metals are coupled together, the more negative metal (the anodic metal) will corrode relative to the more cathodic (noble) metal.