CORROSION

Corrosion is fundamentally an electrochemical process. Examples of corrosion include the rusting of iron, the tarnishing of silver, and the formation of a green coating on copper and bronze. The detrimental effects of corrosion are evident in the substantial damage it causes to structures like buildings, bridges, ships, and various iron-made articles. In the process of corrosion, a metal undergoes oxidation through the loss of electrons to oxygen, resulting in the formation of metal oxide. The corrosion of iron, commonly known as rusting, occurs in the presence of both water and oxygen from the air.

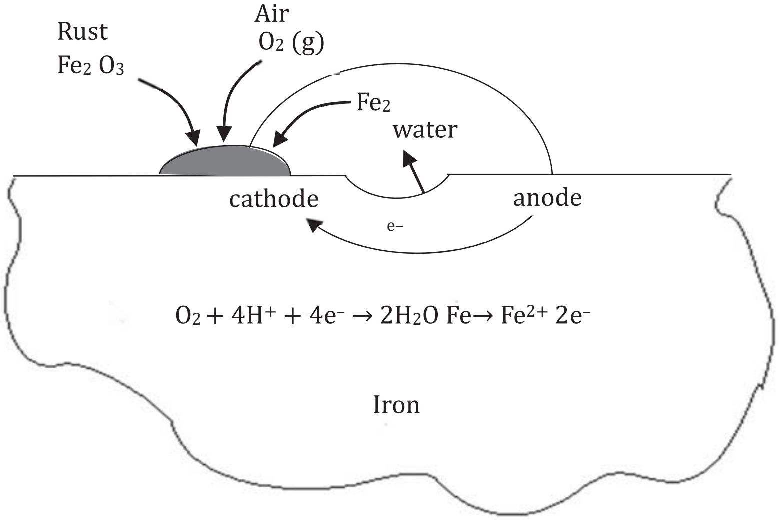

Fig.: The rusting of Iron (Iron in contact with water forms the anode and in contact with air forms the cathode. At the anode iron is oxidised to Fe and at the cathode oxygen is reduced to water

While the chemistry of corrosion is intricate, it is comprehended that oxidation transpires at a specific location on an iron object, with that particular spot acting as an anode.

Anode $Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^- ; ()$

Electrons liberated at the anode traverse through the metal, reaching another location where they participate in the reduction of oxygen in the presence of H^+ ions. These H^+ ions are believed to be derived from H_2CO_3, formed due to the dissolution of CO_2 in water. This specific location functions as a cathode.

Cathode $O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l);(E^0 = 1.23\ V)$

In this context, Fe^{2+} ions traverse through the water present on the surface of the iron object. If the water contains saline, it aids in facilitating the flow of current in the miniature cell formed, thereby intensifying the corrosion process. The comprehensive reaction of the miniature cell is the cumulative outcome of both the cathode and anode reactions, articulated as follows:

$$2Fe(s) + O_2(g) + 4H^+(aq) \longrightarrow 2Fe^{2+}(aq) + 2H_2O(l) ; E^0_{Cell} = 1.67\ V)$$

The Fe^{2+} ions undergo additional oxidation through atmospheric oxygen, resulting in the formation of Fe^{3+} (as Fe_2O_3), which manifests as rust in the configuration of hydrated iron (III) oxide denoted as Fe_2O_3. xH2O.

$$4Fe^{2+}(aq) + O_2(g) + 4H_2O(l) \longrightarrow 2Fe_2O_3(s) + 8H^+$$

H^+ ions produced in the above reaction help further in rusting.

The presence of impurities in iron accelerates the rusting process by establishing numerous small cells. In contrast, very pure iron exhibits a slower rate of rusting.

Prevention From Corrosion

To prevent the corrosion of metal, a protective coating (such as grease, paint, or metal coatings) is applied to its surface. In the case of iron, there are two methods for coating:

(i) Electrolysis involves coating with metals like Cr, Ni, and Cd.

(ii) Dipping iron objects in molten metal, a process known as Zn and Sn coating. When zinc coating is used to protect iron, it is referred to as Galvanization. In this process, zinc, being more reactive than iron, functions as an anode and undergoes oxidation.

It's crucial to note that even if the zinc coating is breached, iron does not rust, which differs from the scenario of tin coating over iron. In the case of tin coating, if the coating is damaged, iron is exposed, and due to its higher reactivity than both copper and tin, corrosion occurs more rapidly. However, tin protects copper similarly to how zinc protects iron. Instead of coating more reactive metals onto iron, these metals are often employed as sacrificial anodes.

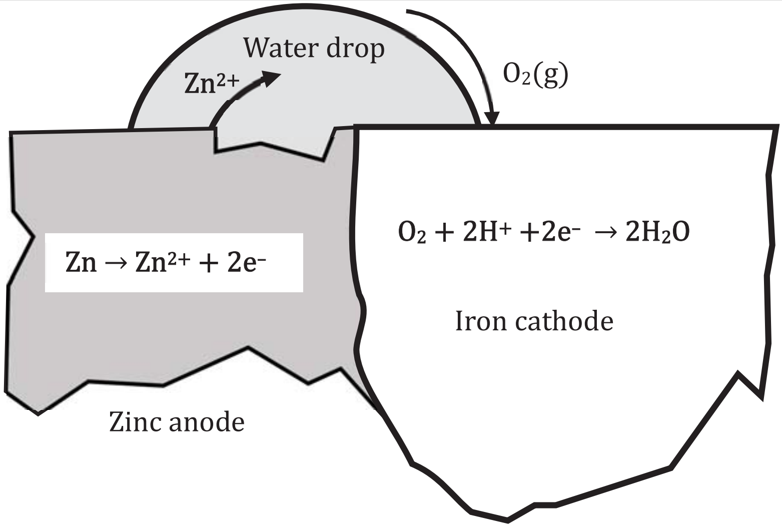

Fig.: Galvanized iron: Cathodic protection of iron in contact with zinc is provided by galvanization.

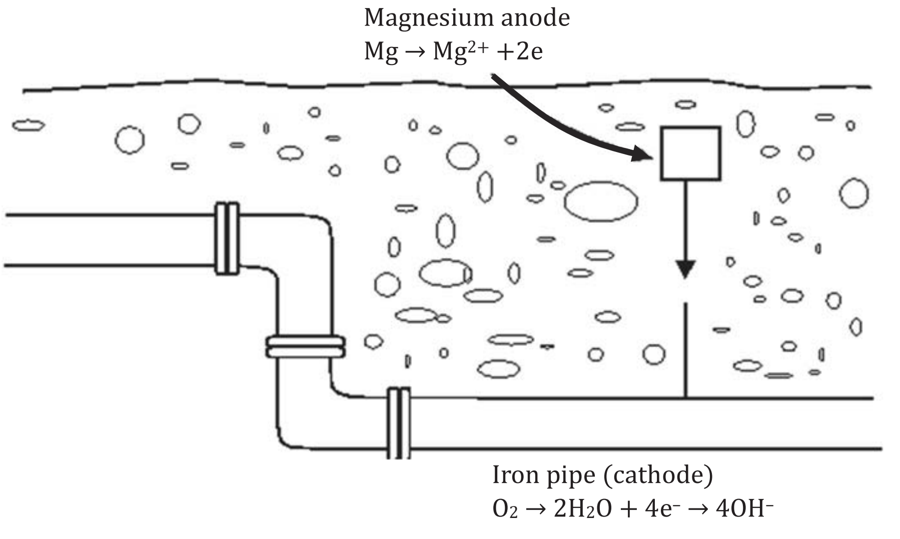

Fig.: Cathodic protection by connecting a block of magnesium to an iron pipe buried underground.

The technique employed to shield iron from corrosion is referred to as Cathodic Protection. In this approach, a plate composed of a reactive metal, such as Zn or Mg, is buried adjacent to the iron pipe or tank and linked to it through wires. In this configuration, iron assumes the role of the cathode, while the more reactive metal functions as the anode. The sacrificial oxidation of the reactive metal anode is crucial for safeguarding the iron. As these reactive metal plates undergo oxidation rapidly, they are periodically replaced, which is a straightforward process.