ELECTROCHEMISTRY

CORROSION

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Corrosion is basically an electrochemical phenomenon. The rusting of iron, the tarnishing of silver, development of a green coating on copper and bronze are some of the examples of corrosion. Corrosion causes enormous damage to building, bridges, ships and to many other articles made of iron.

In corrosion, a metal is oxidised by loss of electrons to oxygen and forms metal oxide. Corrosion of iron(which is commonly known as rusting) occurs in presence of water and oxygen (air).

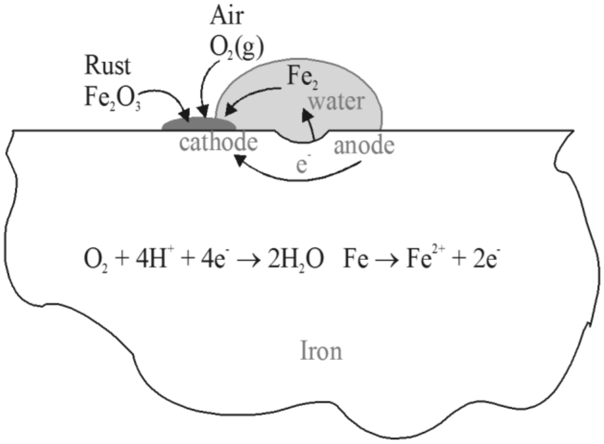

Fig. : The rusting of Iron (Iron in contact with water forms the anode and in contact with air forms the cathode. At the anode iron is oxidised to Fe^{2+},and at the cathode oxygen is reduced to water

Although the chemistry of corrosion is complex but it is understood that at one spot of an iron object oxidation occurs and that spot behaves as an anode.

Anode Fe(s) ¾® Fe^{2+} (aq) + $2e^-$; ()

Electrons released at anode move through the metal and go to another spot on the metal and reduce the oxygen in presence of H^+ ions (which is believed to be available from H_2CO_3 formed due to dissolution of CO_2 in water). This spot works as a cathode:

Cathode O_2(g) + $4H^+$ (aq) + $4e^-$ ¾® $2H_2O$ (l) ; (E^0 = 1.23 V)

Here Fe^{2+} ions move through water on the surface of the iron object. (If water present is saline, it will help more in carrying the current in the miniature cell thus formed and will enhance corrosion). The overall reaction of the miniature cell is the sum of the cathode and anode reactions as follows:

$$2Fe(s) + O_2(g) + 4H^+(aq) \text{ ¾® } 2Fe^{2+}(aq) + 2H_2O(l) \text{ ; } E^0_{Cell} = 1.67 \text{ V)}$$

The Fe^{2+} ions are further oxidised by atmospheric oxygen to Fe^{3+} (as Fe_2O_3) and comes out as rust in the form of a hydrated iron (III) oxide expressed as $Fe_2O_3 . xH_2O$.

$$4Fe^{2+}(aq) + O_2(g) + 4H_2O(l) \text{ ¾® } 2Fe_2O_3(s) + 8H^+$$

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

Impurities present in iron also enhance rusting by setting a number of miniature cells. Very pure iron does not rust quickly.

PREVENTION FROM CORROSION

Corrosion of a metal is prevented by applying protective coating (such as grease, paint or metal coatings) on its (metals) surface. In case of iron this coating is done in two ways:

(i) electrolysis (Cr, Ni and Cd coating)

(ii) dipping iron objects in a molten metal (Zn and Sn coating). Use of zinc coating to protect iron is called Galvanization. In galvanizing iron, zinc being more reactive than iron serves as an anode and is oxidised and).

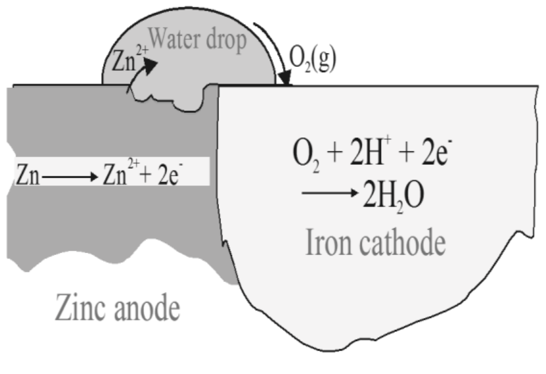

Fig. : Galvanised iron : Cathodic protection of iron in contact with zinc is provided by galvanization.

Here it is important to note that iron does not rust even after zinc coating is broken which is not true in case of tin coating over iron (Now if the coating is broken, iron is exposed and iron being more active than both copper and tin, is corroded. Here iron corrodes more rapidly than it does in the absence of tin. But tin protects copper in the same way as zinc protects iron). Instead of coating more reactive metals on iron, the use of such metals is made as sacrificial anode.

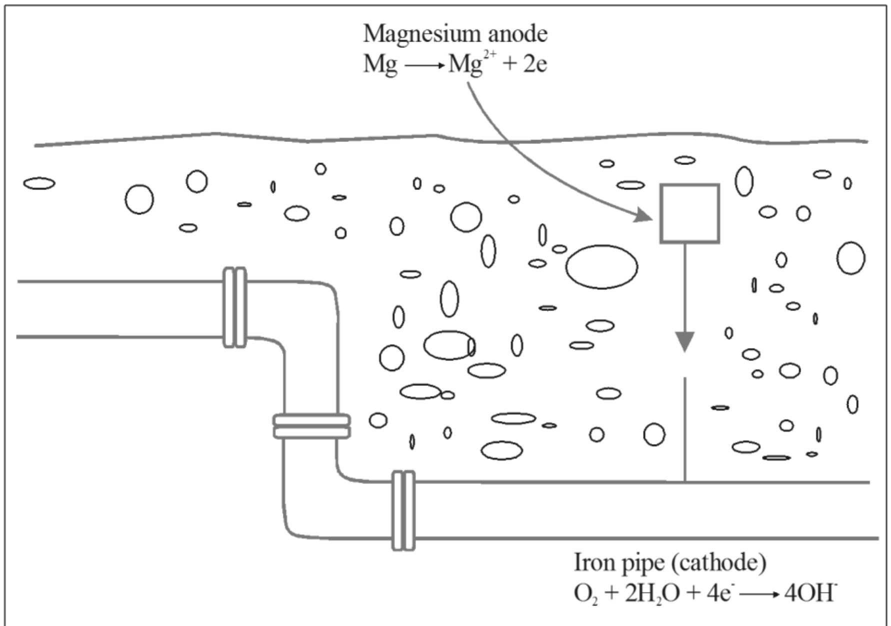

Fig. : Cathodic protection by connecting a block of magnesium to an iron pipe burried underground.

This method of preventing iron from corrosion is Called Cathodic Protection. In this method, a plate of reactive metal (Zn or Mg) is buried beside the iron pipe or tank and connected to it by wires.
Here iron becomes cathode and more reactive metal becomes anode. The reactive metal anode is sacrificed to protect the iron. Since these reactive metal plates are oxidised quickly, they are replaced from time to time which is easy to do.