SRI Y.N. COLLEGE (AUTONOMOUS) NARSAPUR, W.G.Dt.

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STUDY PROJECT

ON

CORROSION

BY

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CORROSION

As soon as the metals are extracted from their ores, the reverse process begins, i.e. nature tries to convert them back into the form in which they occur. This is due to the attack of the gases present in the atmosphere on the surface of the metal converting it into compounds such as oxides, sulphides, carbonates sulphates etc.

***Def. of Corrosion*:** **-** The process of slowly eating away of the metal due to attack of the atmospheric gases on the surface of the metal resulting into the formation of compounds such as oxides, Sulphides carbonates and sulphates is called **corrosion.**

Mechanism or theory of Corrosion

The most common example of corrosion is the **rusting of iron.** Rust is hydrated ferric oxide, $Fe_2 O_3$ x H_2O. Some other example includes tarnishing of silver, development of green coating on copper and bronze etc.

Mechanism or Theory of Corrosion: The theory of corrosion can be explained by taking the example of rusting of iron. The theory is called electrochemical theory because it explains the formation of rust on the basis of the formation of electrochemical cells on the surface of metal.

The formation of rust on the basis of this theory may be explained in the following steps:

(i) The water vapours on the surface of the metal dissolve CO_2 and O_2 from the air. Thus the surface of the metal is covered with the solution of CO_2 in water that is carbonic acid ($H_2 CO_3$)

$H_2O + CO_2$ -------> $H_2 CO_3$

This acts as an electrolytic solution of the cell. The carbonic acid and water dissociate to small extent as follows:

$$H_2CO_3 \text{ -------> } 2H^+ + CO^2{}_3{}^-$$

$$H_2O \text{ -------> } H^+ + OH^-$$

(ii) Iron in contact with the dissolved CO_2 and O_2 undergoes oxidation as follows:

Fe ------> Fe_2+ + 2e- (E_o = 0.44V)

Thus the sites where the above reaction takes place act as anodes. As a result of the above reaction, iron is converted into ferrous (Fe_2+) ions.

(iii) The electrons lost by iron are taken up by H+ ion present on the surface of the metal which was produced by the dissociation of H2CO3 and H2O. Thus H+ ions are converted into H atoms.

$H+ + e \text{--------}> H$ ---- 4

These H atoms either react with the dissolved oxygen or oxygen from the air to form water.

$4H + O_2 \text{-------}> 2H_2 O$ …..5

Multiplying eqn. (4) with 4 and adding to equation (5), the complete reduction reaction may be written as

$O_2 + 4H^+ + 4e- \text{-------------}> 2H_2O$ (Eored = 1.23 V)…..6

The dissolved oxygen may take up electrons directly to form OH- ion as follows:

$O_2 + 2H_2 O + e- \text{----------}> 4OH-$ …....7

The sites where the above reactions take place act as cathodes.

Multiplying equations (3) by (2) and adding to equation (6)

$2Fe(s) + O_2(g) + 4H+ (aq) \text{---}> 2Fe_2+ (aq) + 2H_2O(l)$; $E_0cell = 1.67V$

It may be mentioned here that if water is saline, it helps in the flow of current in the miniature cell and hence enhance the process of corrosion.

(iv) The ferrous ions formed react with the dissolved oxygen or oxygen from the air to form ferric oxide as follows:

$4Fe_2^+ + O_2 + 4H_2 O \text{---}> 2Fe_2 O_3 + 8H^+(aq)$

Ferric oxide then undergoes hydration to form rust as follows:

$Fe_2 O_3 + x H_2O \text{---}> Fe_2 O_3 . xH_2O$

Hydrated ferric oxide (Rust)

It may be noted that the rust is a non-sticking compound i.e., it does not stick to the surface, it peels off exposing fresh iron surface for further rusting.

Factors which promote corrosion

The factors which promote corrosion are:

(i) Reactivity of the metal. More active metals are readily corroded.

(ii) Presence of impurities: Presence of impurities in metals enhances the chances of corrosion. Pure metals do not corrode e.g. pure iron does not rust.

(iii) Presence of air and moisture. Air and moisture accelerate corrosion. Presence of gases like SO_2 and CO_2 in air catalyzes the process of corrosion. Iron when placed in vacuum does not rust.

(iv). Stains in metals. Corrosion (e.g. rusting of iron) takes place rapidly at bends, scratches, nicks and cuts in the metal.

(v) Presence of electrolytes. Electrolytes, if present, also increase the rate of corrosion. For example, iron rusts faster in saline water than in pure water.

Prevention of corrosion

Corrosion can be prevented by a number of ways. Some of these are explained below:

1. Barrier Protection: -

The metal surface is not allowed to come in contact with moisture, oxygen and carbon dioxide. This can be achieved by the following methods:

(i) The metal surface is coated with paint which keeps it out of contact with air; moisture etc. till the paint layer develops cracks.

(ii) By applying film of oil and grease on the surface of the iron tool and machinery, the rusting of iron can be prevented since it keeps the iron surface away from moisture, oxygen and carbon dioxide.

(iii) The iron surface is coated with non-corroding metals such as nickle, chromium aluminium, etc. (by electroplating) or tin, zinc, etc. (by dipping the iron article in the molten metal). This again shuts out the supply of oxygen and water to iron surface.

(iv). The iron surface is coated with phosphate or other chemicals which give a tough adherent insoluble film which does not allow air and moisture to come in contact with iron surface.

2. Sacrificial Protection: -

Sacrificial protection means covering the iron surface with a layer of metal which is more active (electropositive) than iron and thus prevents the iron from losing electrons. The more active metal loses electrons in preference to iron and converts itself into ionic state. With the passage of time, the more active metal gets consumed but so long as it present there, it will protect the iron from rusting and does not allow even the nearly exposed surface of iron to react. The metal which is most often used for covering iron with more active metal is zinc and the process is called **galvanization**. The layer of zinc on the surface of iron, when comes in contact with moisture, oxygen and carbon and carbon-dioxide in air, a protective invisible thin layer of basic zinc carbonate $ZnCO_3\ Zn(OH)_2$ is formed due to which the galvanized iron sheets lose their luster and also tends to protect it from further corrosion.

Iron can be coated with copper by electro-deposition from a solution of copper sulphate or with tin by dipping into molten metal. 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.

At time, zinc, magnesium and aluminium powders mixed with paints provide decorative protective coatings also.

3. Electrical Protection: -

Cathodic protection. The iron object to be protected from corrosion is connected to a more active metal either directly or through a wire. The iron object acts as cathode and the protecting metal acts as anode. The anode gradually used up due to the oxidation of the metal to its ions due to loss of electrons. Hydrogen ions collect at the iron cathode and prevent the rust formation. The iron object gets protection from rusting as long as some of the active metal is present. Metals widely used for protecting iron objects from rusting are magnesium, zinc and aluminium which are called sacrificial anodes.

Magnesium is often employed in the cathodic protection of iron pipes buried in the moist soil, canals, storage tanks etc. Pieces of magnesium are buried along the pipeline and connected to it by the wire.

4. Using anti-rust solutions:

These are alkaline phosphate and alkaline chromate solutions. The alkalinity prevents availability of hydrogen ions. In addition, phosphate tends to deposit an insoluble protective film of iron phosphate on the iron. These solutions are used in carradiators to prevent rusting of iron parts of the engine.