Dr. Uma K.
Assistant Professor
Dept. of Chemistry
S.I.T., Tumakuru

Unit- IV

CORROSION SCIENCE AND ENGINEERING

(From the Academic year 2022-223 onwards for ME stream)

Man has presented himself in a hostile environment in the early stages of human civilization by functioning as biochemical machines i.e., converting the solar power via electrochemical reactions. Of late, man has become increasingly a cyborg (a person with artificial parts and organs). The progress of civilization has been marked by an enhanced use of machine power. The machines are fabricated mainly from metals and these metals must be stable. In case the metals become unstable, the machines made from them undergo an undesired obsolescence. Therefore, the study of corrosion of metals has become more significant. The damage caused by corrosion costs several per cent of the gross national product of technologically advanced countries. One estimate puts the cost of metallic corrosion to the US economy in 1984 at ~80 billion dollar or 3% of the gross national product for that year.

Eg: Rusting of iron, tarnishing of silver and green putina formed on copper and brass.

Metallic corrosion- The gradual destruction or deterioration of a metal or its alloy caused by the direct chemical action or electrochemical reaction with its environment is called metallic corrosion.

Corrosion interrelated problems

Corrosion is an insidious process, which is often difficult to recognise until deterioration is well advanced. It may have the following interrelated problems.

1.   It causes damage to process plants, structural assemblies and other equipment resulting in the formation of a corrosion product.

2.   The process products get contaminated with the corrosion product.

3.   The contaminated product becomes no more useful resulting in loss of products.

4.   The contaminated product is dumped into the environment causing its contamination.

5.   To continue with the production, redesigning or over engineering is needed and consequent shut down for repair or replacement work.

6.   The customer will not wait and find some other manufacturer i.e. customer alienation.

Electrochemical theory of corrosion

According to electrochemical theory of corrosion, there will be formation of large number of minute galvanic cells on the surface of a metal. The cells arise due to the presence of heterogeneities at the interface of the metal and environment. The heterogeneities on a metal surface could arise in several ways as follows.

1.   ***Difference in concentration of air or oxygen on metal surface***- the part in contact with more air or oxygen becomes a cathode while the other part in contact with low air or oxygen content becomes an anode.

2.   ***Strain in the metal*** - Strained metal part serves as an anode while the unstrained part becomes a cathode.

| *Heterogeneity*	*Anode*	*Cathode*
Difference in concentration of air or oxygen on metal	Less aerated or lower concentration of air/oxygen	more aerated or higher concentration of air/oxygen
Strain in the metal	Strained metal part	unstrained metal part

Thus two unlike parts exist on the metal surface.

i)   *The part where the metal gets attacked* (or undergoes oxidation) - anode.

ii)   *The part where some substance from the environment gets reduced* - cathode.

The aqueous solution (derived from moisture) constitutes an electrolyte.

Anodic area:

At the anodic site, the metal undergoes oxidation-releasing electrons.

$$M \rightarrow M^{n+} + ne^- \ ; 2\ Fe \ \rightarrow \ 2\ Fe^{++} + 4e^- \ \underline{\hspace{3cm}} \ (1)$$
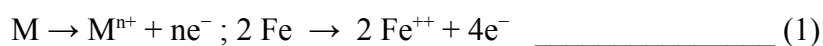
Cathodic area:

At the cathodic site, the following reduction reactions take place.
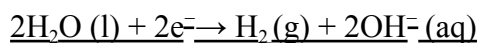
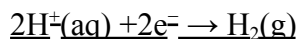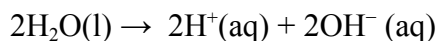a)   Reduction of H^+ ions to H_2 gas.

b)   Reduction of O_2 gas in presence of water to OH^- ions.

 i) **In acidic medium and in the absence of O_2**

  If the corrosion is occurring in acidic medium, the H^+ ions of the medium take up the electrons and get reduced to H_2 gas.

$$2\ H^+(aq) + 2\ e^- \rightarrow H_2(g)$$
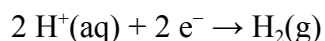
ii) **In acidic medium and in presence of O_2**

$$2\ H^+(aq) + \tfrac{1}{2}\ O_2 + 2\ e^- \rightarrow H_2O(g)$$
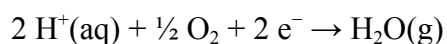
iii) **In neutral and deaerated (absence of O_2) medium**

   If the corrosion medium is neutral (or weakly alkaline) and deaerated, cathodic reaction takes place according to the equation,
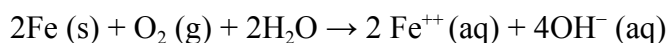
   $2H_2O(l) \rightarrow 2H^+(aq) + 2OH^- (aq)$

   $\underline{2H^+(aq) +2e^- \rightarrow H_2(g)}$

   $\underline{2H_2O (l) + 2e^- \rightarrow H_2(g) + 2OH^- (aq)}$

iv) **In neutral and aerated (presence of O_2) medium**

  If the medium is neutral and aerated, the reaction at cathode is,

   $O_2 (g) + 2H_2O (l) + 4e^- \rightarrow 4OH^- (aq)$ _____ (2)

 The cell reaction is,

  $2Fe (s) + O_2 (g) + 2H_2O \rightarrow 2 Fe^{++}(aq) + 4OH^- (aq)$
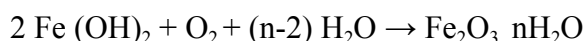
 The metal ions so formed at the anode combine with the OH^- ions to form the corresponding metallic hydroxide, a corrosion product.

  $2 Fe^{++} + 4OH^- (aq) \rightarrow 2 Fe (OH)_2$
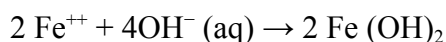
 The metallic hydroxide undergoes oxidation to ferric oxide, the final corrosion product.

  $2 Fe (OH)_2 + O_2 + (n-2) H_2O \rightarrow Fe_2O_3 . nH_2O$

Types of corrosion

1. Differential metal corrosion

When two different metals in contact are present in corroding environment, there exists a potential difference resulting in galvanic corrosion. One metal becomes anodic and undergoes oxidation, while the other becomes cathodic where reduction occurs. Under the aforesaid conditions, a small electric current, which is approximately proportional to the loss of metal, is caused. The rate of differential metal corrosion depends on the difference in the electrode potential. Higher the difference faster is the rate of corrosion

Fe	Sn

Fe	Cu

Anode   cathode   anode   cathode

-0.44 V   0.34 V   -0.44 V   -0.1 4 V

$E_{cell} = E^o_{Cu++|Cu} - E^o_{Fe++|Fe}$ $E_{cell} = E^o_{Sn++|Sn} - E^o_{Fe++|Fe}$

    $= 0.34 - (-0.44)$ $= -0.14 - (-0.44)$

    $= 0.78$ V $= 0.3$ V

The potential difference between iron and copper is more ($E^o{}_{Fe^{++}|Fe}$ = -0.44 V; $E^o{}_{Cu^{++}|Cu}$ = 0.34 V and potential difference is 0.78 V) than that between iron and tin ($E^o{}_{Fe^{++}|Fe}$ = -0.44 V; $E^o{}_{Sn^{++}|Sn}$ = -0.1 4 V and potential difference is 0.3 V). Hence, iron corrodes faster when in contact with copper than in contact with tin.
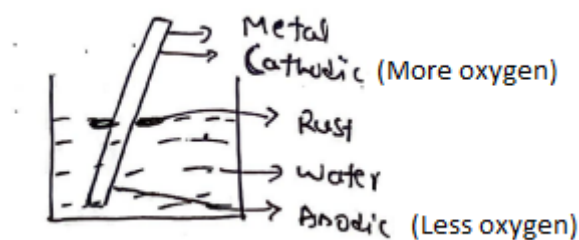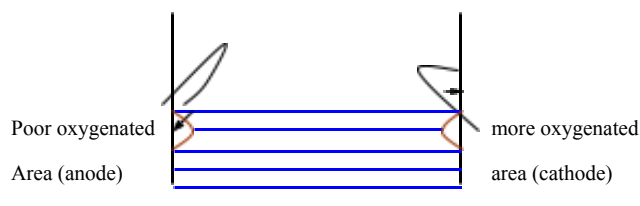
2. Differential aeration corrosion

The galvanic cell formation takes place due to difference in the oxygen concentration. The part of the metal that is associated with high concentration of oxygen acts as a cathode while the other part with a low oxygen concentration serves as an anode. A difference in the concentration of dissolved oxygen at the metal surface forms a concentration cell.

The pitting and waterline corrosion can be explained using the phenomenon of differential aeration.

i) **Pitting corrosion:** Pitting corrosion is usually observed when small particles of dust, sand or scale are deposited on a metal surface. The portion below the dust is less aerated (would become anode) compared to the exposed part of the metal, which is more aerated (would serve as cathode). Corrosion begins at the portion below dust with the formation of a pit in presence of an electrolyte or moisture. The same is depicted in figure. The rate of corrosion increases once a pit is formed and is due to small anodic area and large cathodic area.
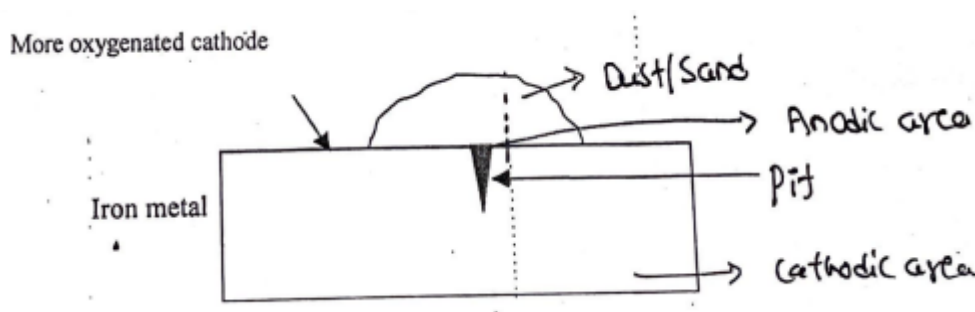

(ii) **Waterline corrosion:** Many investigators have reported that when a constant water level is maintained, corrosion often causes at or just below the water line. Evans showed that the metal in contact with water surface i.e., the creep (meniscus) is cathodic (the area above water level is more oxygenated) and the metal just below the water is anodic (the area below water level is less oxygenated). Therefore, corrosion takes place just beneath the level of water as shown in figure. The amount of creep of the water determines the rate of corrosion. The greater the creep, the more cathodic the water-line area and more is the rate of corrosion just below the water line.

Poor oxygenated Area (anode)

more oxygenated area (cathode)

Metal
Cathodic (More oxygen)
Rust
Water
Anodic (Less oxygen)

Waterline corrosion is commonly observed in steel tanks used to store water for a long time. This type of corrosion is also observed in ships floating in seawater for a long time.

3. Stress corrosion

The presence of both stress and a specific corrosion environment are necessary for the occurrence of stress corrosion. The stress may result from mechanical operations such as thermal processing (welding, quenching), fabrication (bending, deep drawing, cold working), assembly (press fitting, shrink fitting), *etc*.

Rust
Cathodic part
$H_2O + 1/2 O_2 + 2e^-$
Anodic part
$Fe \rightarrow Fe^{2+} + 2e^-$
$2OH^-$
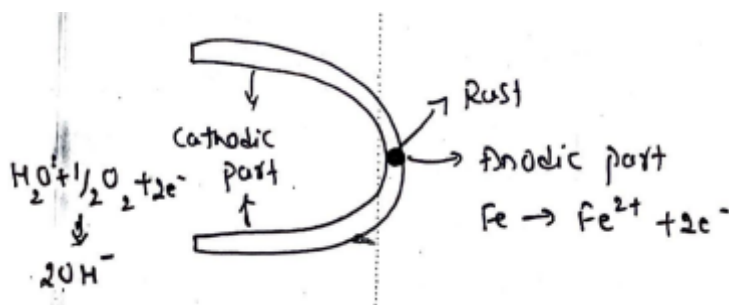
The precise mechanism of stress corrosion is thought to be related to hydrogen embrittlement and to localised increase of stress corrosion followed by stress relief *via* cracking. Cracking is caused when a static load is applied on a metal due to increase in tensile stresses.

Factors affecting the rate of corrosion

There are several factors, which influence the rate of metallic corrosion. These factors are classified into two categories, namely, *Primary factors and Secondary factors*.

Primary factors (which are related to the metal)	Secondary factors (which are related to the environment)
Nature of corrosion product	pH
Electrode potential	Temperature
Area effect- Relative proportion of anodic and cathodic areas	Polarization

Primary factors

1. Nature of corrosion product

If the corrosion product formed on the surface of a metal has stoichiometric composition, highly insoluble and has low ionic and electronic conductivities, it forms a thin protective layer (usually an oxide or hydroxide) on the surface of the metal and thereby decreases the rate of corrosion. Eg: Al, Cr, Ti. On the other hand, if the corrosion product formed has non-stoichiometric composition and has appreciable conductivity, it forms a thick porous layer on the surface of metal, which peels off and thereby exposes the fresh surface of metal for corrosion. Therefore, the rate of corrosion increases. Eg: Fe and Zn.

Justify: Iron corrodes faster than aluminium although it is placed below aluminium in the electrochemical series.

2. Electrode potential

It is the main factor, which determines the rate of corrosion. When two dissimilar metals exposed to atmosphere are in contact with each other, there results in galvanic corrosion. Higher the difference in electrode potentials faster is the rate of corrosion.

Eg. Iron corrodes faster when in contact with copper but slowly when in contact with tin.

Fe	Sn

Fe	Cu

Anode   cathode   anode   cathode

-0.44 V   0.34 V   -0.44 V   -0.1 4 V

$E_{cell} = E^o_{Cu++|Cu} - E^o_{Fe++|Fe}$ $E_{cell} = E^o_{Sn++|Sn} - E^o_{Fe++|Fe}$

$= 0.34 - (-0.44)$ $= -0.14 - (-0.44)$

$= 0.78 \text{ V}$ $= 0.3 \text{ V}$

The potential difference between iron and copper is more ($E^o_{Fe++|Fe} = -0.44$ V; $E^o_{Cu++|Cu} = 0.34$ V and potential difference is 0.78 V) than that between iron and tin ($E^o_{Fe++|Fe} = -0.44$ V; E^o

$_{Sn^{++}|Sn}$ = -0.1 4 V and potential difference is 0.3 V). Hence, iron corrodes faster when in contact with copper than in contact with tin.

Eg: Iron corrodes faster when in contact with copper but not when in contact with zinc.

Zn	Fe

Fe	Cu

Anode	cathode	anode	cathode
-0.44 V	0.34 V	-0.76 V	-0.44 V

$E_{cell} = E°_{Cu^{++}|Cu} - E°_{Fe^{++}|Fe}$ 　　　　$E_{cell} = E°_{Fe^{++}|Fe} - E°_{Zn^{++}|Zn}$

　　　$= 0.34 - (-0.44)$ 　　　　　　　　　　$= -0.44 - (-0.76)$

　　　$= 0.78 V$ 　　　　　　　　　　　　　$= 0.32 V$

When iron in contact with copper, iron acting as anode and copper becomes cathode. The rate of corrosion is more because more is the electrode potential value.

Eg: Nickel corrodes when in contact with lead but not when in contact with chromium.

Cr	Ni

Ni	Pb

Anode	cathode	anode	cathode
-0.74 V	-0.25 V	-0.25 V	-0.13 V

$E_{cell} = E°_{Ni^{++}|Ni} - E°_{Cr^{+++}|Cr}$ 　　　$E_{cell} = E°_{Pb^{++}|Pb} - E°_{Ni^{++}|Ni}$

　$= -0.25 - (-0.74)$ 　　　　　　　　　$= -0.13 - (-0.25)$

　$= 0.49 V$ 　　　　　　　　　　　　　$= 0.12 V$

Therefore, the use of different metals in contact with each other should be avoided to prevent the formation of galvanic corrosion. Bolt & nut or screw and washer should be made of the same metal or alloy.

**3. Area effect - Relative size of the anodic and cathodic areas**

If both the anodic and cathodic areas are minute and evenly distributed, a uniform corrosion results in. If the anode area is very small compared to large cathode area, intense corrosion will occur i.e., faster the rate of corrosion and vice versa. This is because, reduction occurs at a greater extent as the cathode area is large. As a result, there will be great demand for electrons, which has to be met by the dissolution of metal rapidly at the small anodic area.

Eg: Pinholes on tin coated iron are more prone to corrosion of iron compared to those on zinc-coated iron.

If a scratch or hole is formed in a tin layer, the corrosion of the underlying iron is rapid. This is because, tin coating is cathodic to iron and anodic oxidation of the small exposed area of iron occurs while oxygen reduction occurs over the whole tin surface (i.e. higher ratio of cathode to anode area). On the other hand, zinc is more electropositive (anodic) with respect to iron (cathodic) and the top layer is damaged, the iron does not corrode as zinc dissolves preferentially.
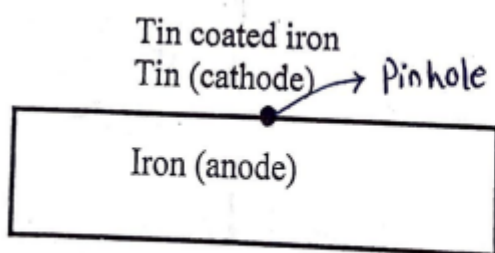
Smaller the anodic area, larger the cathodic area more is the rate of corrosion.
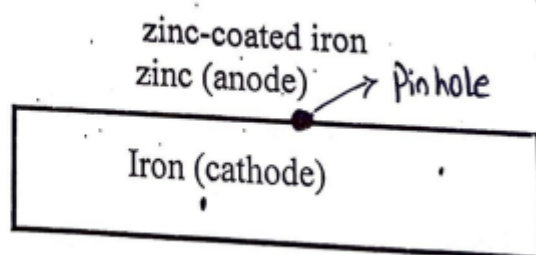
larger the anodic area, smaller the cathodic area less is the rate of corrosion.

Secondary factors

1. pH

In general, the rate of metallic corrosion is more in acidic solution (with lower pH value) than in a neutral or alkaline solution (with higher pH value).

i) If the pH of the medium is below 4, severe corrosion occurs even in the absence of air due to the continuous evolution of hydrogen on the cathode region. The increased rate of corrosion is attributed to the sum of both hydrogen evolution and oxygen depolarisation.

ii) The corrosion rate decreases above pH 10, because iron becomes increasingly passive in presence of alkali's and dissolved oxygen.

iii) The rate of corrosion is independent of pH in the range 4-10 and it depends only on how rapidly oxygen diffuses to the metal surface.

2. Temperature

Increase in temperature increases the rate of metallic corrosion. As the temperature is increased, the evolution (diffusion) of hydrogen gas at cathode becomes more rapid and the effect of polarisation decreases. The corrosion rate at a given oxygen concentration

approximately doubles for every 30 °C rise in temperature. When corrosion is attended by hydrogen evolution, the rate almost doubles for every 30 °C rise in temperature. The rate of corrosion of iron in HCl approximately doubles for every 10 °C rise in temperature.

Corrosion rate in terms of penetration

Corrosion penetration rate (CPR) = $\dfrac{kW}{\rho At}$

k = constant

W = metal lost

ρ = density of metal

A= area exposed for corrosion

t = duration of corrosion

Problem:

(1) An iron metal lost 2.4 $\times 10^{-3}$ g of weight when it is immersed in a deaerated acidic solution for 2 hrs. Calculate the corrosion penetration rate (CPR) for the exposed area of 4 cm^2. Given k = 534 and the density of metal = 5.37 g cm^{-3}. Express the CPR in terms of meter per second.

Solution: CPR = $\dfrac{kW}{\rho At}$

Given k = 534, W = 2.4 x 10^{-3} g, A = 4 cm^2, ρ= 5.37 g cm^{-3}

CPR = $\dfrac{2.4 \times 10^{-3} \text{ g} \times 534}{5.37 \text{ g cm}^{-3} \times 2 \text{ hrs} \times 4 \text{ cm}^2}$

CPR = 0.0298 cm / hrs

CPR in terms of meter per second is

CPR = $\dfrac{0.0298 \times 10^{-2} \text{ m}}{60 \times 60 \text{ second}}$

CPR = 8.27 x 10^{-8} meter/ second

Corrosion rate can also be calculated by the amount of metal undergoing corrosion in time, is given by

$m = \dfrac{iM}{nFt}$

M = atomic weight of metal

i = current

n = number of electrons

F = faraday constant

t = time

Problem:

(2) If the corrosion current in an iron specimen immersed in a deaerated acidic solution is 4.21 $\times 10^{-1}$ Am^{-2}, calculate the corrosion rate of iron in milligrams per square decimetre per day (mg $dm^{-2}d^{-1}$). Given atomic weight of iron =55.8 $gmol^{-1}$, faraday constant=96490 $cmol^{-1}$, t=1 second, density of iron=7.87 gcm^{-3}.

Solution:

i) Amount of iron m, undergoing corrosion in time 't' is given by

i=4.21 $\times 10^{-1}$, n = 2, F = 96490 $cmol^{-1}$, t= 1 sec, M= 55.8 g

$$m = \frac{i\,M}{nFt}$$

$$m = \frac{4.21 \times 10^{-1} Am^{-2} \times 55.8\ gmol^{-1}}{2 \times 96490\ cmol^{-1}}$$

$$m = \frac{4.21 \times 10^{-1} Am^{-2} \times 55.8 \times 10^{3}\ mgmol^{-1}}{2 \times 96490\ cmol^{-1}} \qquad \text{(since C=AS)}$$

$m = 0.122$ mg $m^{-2} S^{-1}$

ii) To calculate m in mg dm^{-3}

$$m = \frac{0.122\,mg}{m^2\,S^1}$$

$$m = \frac{0.122\,mg}{(10\ dm)^2 \times \dfrac{d}{(24 \times 3600)}}$$

$$m = \frac{0.122\,mg \times 24 \times 3600}{100\ dm^2 . d}$$

m= 105.408 mg $dm^{-2}d^{-1}$

1m= 10 dm

1d = 24 x 60 x 60 S

= (24x 3600)S

$$1S = \frac{d}{24 \times 3600}$$

Corrosion control methods

Corrosion can be completely avoided only under ideal conditions. Since it is impossible to attain such conditions, it can be minimized by using various corrosion control methods. They are:

a)   by protective coatings

b)   by cathodic protection

c)   by corrosion inhibitors

d)   by anodic protection

Protective coatings

●   Corrosion is prevented by the application of protective coating on the surface of metal; thereby the metal surface is isolated from the corrosive environment.

●   The coatings being chemically inert to the environment under specific conditions of temperature and pressure, forms a physical barrier between the coated surface and its environment.

●   Coatings are not only preventing corrosion but also decorate the surface of the metal.

●   Important types of protective coatings are:

    (i)   Inorganic coatings

    (ii)   Metal coatings and

    (iii)   Organic coatings

Inorganic coatings (Chemical conversion coatings): the surface of the metal is converted into a layer of stable compound by chemical or electrochemical reactions are called as inorganic coatings.

☐   These coatings are produced at the surface of the metal by chemical / electrochemical

reactions.

☐   These coatings are applied on the article for decorative effect and to increase the

corrosion resistance of the base metal.

☐   These serves as an excellent base coating for paints and enamels.

☐   Examples: Anodising and Phosphating

Anodising:

The process of forming surface oxide film by electrochemical oxidation is called Anodising.
Eg: Formation of a layer of Al_2O_3 *on Al.*
A number of other metals such as Mg, Ti, Ta, V, Zr, and Ni also form surface oxide films.

Anodising of aluminium

Aluminium article is first subjected to cleaning prior to anodising. The pre-treatment includes degreasing followed by electropolishing. A base metal Al is made as anode and lead (inert electrode) is made as cathode. The electrolyte used is 5-10% chromic acid or 10% sulphuric acid. The temperature of the bath maintained is 35 °c. Voltage is programmed to increase from 0-50 V, so as to maintain an anode current density of 10-20 mA cm^{-2}.

The electrode reactions occurring during anodising process are:

At anode: $2\ Al + 3\ H_2O \rightarrow Al_2O_3 + 6\ H^+ + 6e^-$

At cathode: $6\ H^+ + 6\ e^- \rightarrow 3\ H_2$

Over all reaction: $2\ Al + 3\ H_2O \rightarrow Al_2O_3 + 3\ H_2$

Applications: Anodized articles are used in automobiles, aircraft assemblies, window frames.

Phosphating:

Phosphate coating is obtained on iron and steel surface by converting the surface metal atoms into their phosphate by chemical reactions.
Eg: Formation of a layer of Iron phosphate on iron or steel.

The principal article is steel, although zinc and aluminium may also be phosphated.
Process: Phosphating is done by dipping process. A phosphating bath contains three essential components (a) phosphoric acid

(b) a primary metal phosphate and

(c) an accelerator

The pH of the bath is maintained in the range 1.8-3.2. Temperature of the bath is maintained at 35° C. The majority of metal phosphates are insoluble in water but are soluble in mineral acids. This constitutes the basis for a phosphate coating reaction. The reactions are slow; hence accelerators are used to speed up the reaction (coating rate) and reduce the crystal size. The commonly used accelerators include nitrites, nitrates, chlorates and peroxides (H_2O_2).

The chemical and electrochemical reactions occurring during phosphating involve the dissolution of the metal (M = Fe, Zn, Mn) followed by the deposition of an insoluble phosphate layer. i.e. The surface of iron or steel dissolves to form metal ions which combine with phosphate ions to form metal phosphate (Iron phosphate). Finally, the metal phosphate is deposited on the surface of metal iron or steel. Or the chemical reaction between the base metal and the phosphating mixture results in the formation of surface film consisting of crystalline Mn-Fe Phosphate, Zn-Fe Phosphate.

Applications: Phosphate coatings offer an improved corrosion resistance and render the surface, a good paint adhesion quality. The most important applications involve a mixed Fe and Zn phosphates on steel sheet as a standard pre-treatment before painting of automotive bodies and other steel fabrications such as refrigerators and washing machines.

Metal coatings: the deposition of a protective metal over the surface of a base metal is called metal coating.

☐   Metal coatings can be applied on the base metal by hot dipping process.

☐   This method is used for producing a coating of low melting metals such as Zn, Al, Sn etc., on iron / steel metals which have relatively high melting point.

☐   The process involves immersing of the base metal in a molten bath of coating metal

☐   The coating metal may be anodic or cathodic to the base metal.

☐   Examples: Galvanising and Tinning

There two types of metal coatings i) Anodic coatings
                                  ii) Cathodic coatings

(i) Anodic coating: Anodic coating is produced by coating a base metal with more active metal which is anodic to the base metal. Eg: Galvanising

(ii) Cathodic coating: Cathodic coating is produced by coating a base metal with less active metal which is cathodic to the base metal. Eg: Tinning

Galvanising	Tinning
▪ The process of coating **zinc** on iron or steel by hot dipping process is called galvanising. (M.P of Zn = 419°C) ▪ The article is washed with organic solvents to remove oil/grease. ▪ Then it is washed with dilute sulphuric acid to remove oxide film or scale/rust. ▪ Finally the article washed with water and dried, before coating. ▪ The article is dipped in a molten bath of Zn. The excess of coated metal is removed by passing through a pair of hot rollers to get a thin coating and cooled gradually. Applications: ▪ Galvanising is applied to nails, bolts, pipes, roofing sheets etc. ▪ Galvanised sheets cannot be used for preparing/storing food stuffs, since Zn dissolves in acidic medium and forms toxic compounds. ▪ If any crack is produced on the galvanised sheets, causes severe corrosion	▪ The process of coating **tin** on iron or steel by hot dipping process is called tinning. (M.P of Sn = 232°C). ▪ The metal surface is washed with organic solvents to remove oil/grease. ▪ Then it is washed with dilute sulphuric acid to remove oxide film or scale/rust. ▪ Finally the article washed with water and dried, before coating. ▪ The article is dipped in a molten bath of Sn. The excess of coated metal is removed by passing through a pair of hot rollers to get a thin coating and cooled gradually. Applications: ▪ Tinning is widely used for coating the steel sheets, Cu and brass sheets used for manufacturing containers for storing/packing food materials, oils and milk products cooking utensils, refrigeration equipments, etc. ▪ If any crack is produced on the tinned sheets, causes severe corrosion of the base metal.

on the coated Zn metal and the base metal is protected. ▪ Zn is chosen as a protective coating for iron/steel because of its natural resistance against corrosion in most atmospheric conditions, and Zn is electronegative to iron and can protect it sacrificially.	▪ Tin coatings form a useful preparation for protective painting in general applications.

Cathodic Protection

The principle is to force the metal to be protected, to behave as cathode**.**

Metals undergo corrosion due to electrochemical process. During corrosion anodic and cathodic areas are formed on the same metal surface and the anodic part corrodes. However, corrosion can be prevented by removing anodic area and converting entire metal surface into cathodic area. This process is known as cathodic protection.

 There are two types of cathodic protections namely,

1) Impressed voltage method (Impressed current cathodic protection)

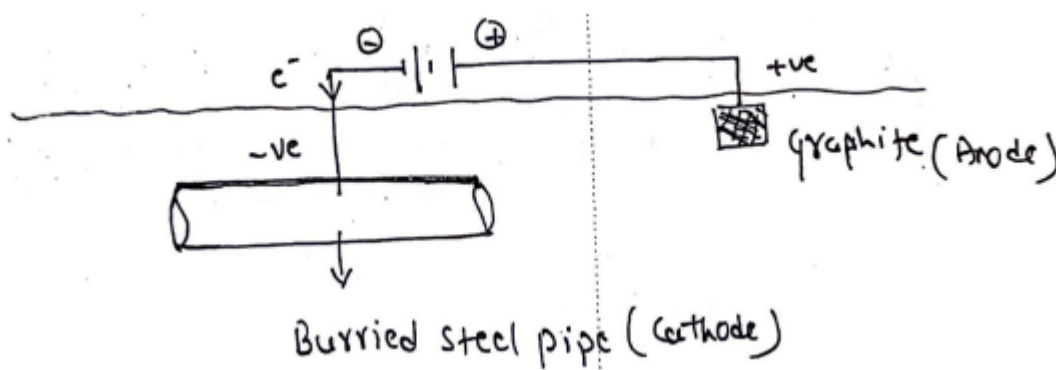2) Sacrificial anode method (Sacrificial anodic protection)

Impressed current cathodic protection	Sacrificial anodic protection
▪ The metallic structure to be protected is made as cathode by impressing the current. ▪ The current is applied in the opposite direction to nullify the corrosion current. ▪ The impressed current is obtained from a source like battery. ▪ An insoluble anode (eg: graphite, high silicon content iron, *etc*.) is buried in the	▪ The metallic structure to be protected is connected to a more anodic metal using a metallic wire. ▪ The more active metal gets corroded, while the parent structure is protected from corrosion. ▪ The more active metal so employed is called sacrificial anode.

soil, and connected to the structure to be protected. ▪ This method is suitable for large structures and for long term operations.	▪ The sacrificial anodes to be replaced by fresh ones as and when it is required. ▪ Commonly used sacrificial anodes are Mg, Zn, Al *etc*. ▪ This method is generally used for the protection of buried pipelines, ship hulls, water tanks, etc.

Impressed voltage method

In this method corrosion of metal can be prevented by applying direct current in the opposite direction to nullify the corrosion current. In practice the negative terminal of D.C. is connected to metal structure to be protected and positive terminal is connected to inert graphite. The metal structure becomes cathodic does not undergo corrosion and graphite anode being inert remains unaffected.

Eg: Underground buried pipe, steel tank etc.

Sacrificial anode method

In this the metal structure to be protected is connected to a more anodic metal like Mg, Zn and Al which acts as anodic. In practice active metal Mg or Zn is connected to metal structure to be protected. The more active metal Mg or Zn becomes anodic and corrodes but the metal structure becomes cathodic is protected. Since the anode metal sacrifices to protect the metal structure the method is known as Sacrificial anode method.

Eg: Ships floating on water surface, Underground buried pipe.
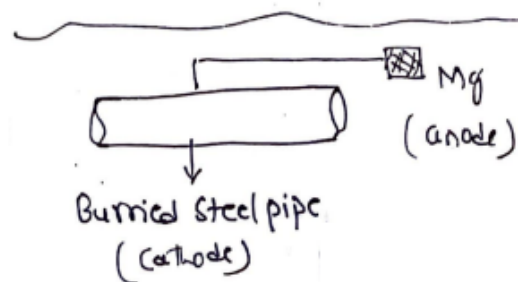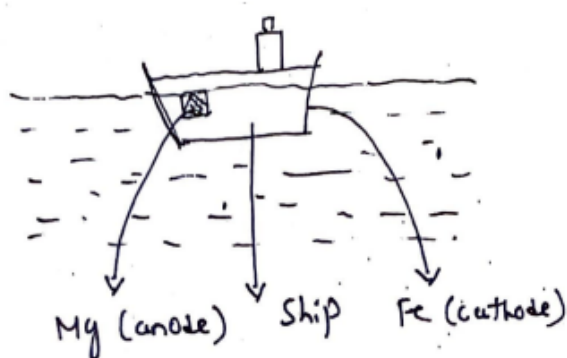
Questions:

1. Define metallic corrosion. Mention the corrosion interrelated problems.

2. What is metallic corrosion? Explain the corrosion of iron based on electrochemical theory.

3. Explain the effect of the following factors on the rate of metallic corrosion:

4. (i) pH (ii) Temperature (iii) Electrode potential

5. What is cathodic protection? Explain the cathodic protection given to metals exposed to corrosion environment.

6. Explain the type of corrosion occurring in each of the following cases.

   (i) The presence of small particles of dust on iron surface for a long time.

   (ii) Bolt and nut made from different metals are in contact with each other.

7. Explain:

   (i) Why is a metal corrodes below waterline?

   (ii) Why are the ships sailing in seawater corrode below the water line?

   (iii) Aluminium articles are self-protected against corrosion in air, while iron articles are not.

8.   (i) Illustrate the reactions involved in differential aeration corrosion with reference to iron metal.

   (ii) Pin holes on tin coated iron are more prone to the corrosion of iron than those on zinc coated iron. Account for this.

9. Explain the effect of the relative proportion of anodic and cathodic areas on the rate of metallic corrosion.

9.   What is anodizing? Explain anodising of aluminium.

10. Explain stress corrosion.

11. What is galvanising? How is galvanising done on iron articles?

12.   Give an account of phosphating employed in corrosion control.

13. What are corrosion inhibitors? How is metallic corrosion prevented using corrosion inhibitors?

14. If the corrosion current in an iron specimen in a deaerated acidic solution is 4.21×10^{-1} Am^{-2}, calculate the corrosion rate of iron in milligrams per square decimetre per day ($mg\ dm^{-2}d^{-1}$) and in milli meters per year ($mm\ y^{-1}$). Given atomic weight of iron = 55.8 g mol^{-1}, Faraday = 96490 C mol^{-1}, density of iron = 7.87 g cm^{-3}.