



# Corrosion in Materials

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**Corrosion process:** Corrosion, Cause of corrosion, Types of corrosion, Protection against corrosion.

CLO: Propose a solution to prevent corrosion.

**Corrosion** is the deterioration and loss of metal, or material by chemical or electrochemical reactions with its environments.

- ❑ It is a very natural process.
- ❑ Most of metallic elements (except noble metals) present in the most stable state having low energy, normally oxides (as metallic ores)

## Reason for corrosion

The tendency to decrease the energy of the system as a result of the corrosion reactions.

- ❑ Metals get oxidized
- ❑ Polymers react with oxygen and degrade
- ❑ Ceramic refractories may dissolved in contact with molten materials

**Example:** Formation of rust on the surface of iron, formation of green film on the surface of copper



The responsible factors for the corrosion of a metal are the **metal** itself, the **environmental chemicals**, **temperature** and the **design**.

**Electro-chemical reaction:** most common mechanism by which corrosion occur at or near room temperature as a result of a chemical reaction metals with water, o aqueous solutions of salts, acids or bases. It is named electro-chemical corrosion as a chemical reaction accompanied the passage of **electric-current takes place**.

Metal atoms characteristically lose or give up electrons in what is called an **oxidation** reaction.

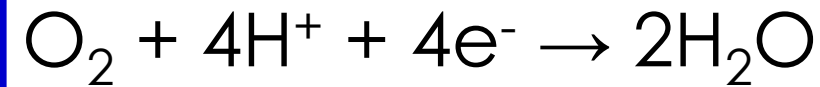


The electrons generated from each metal atom that is oxidized are transferred to and become a part of another chemical species. This is a **reduction** reaction.

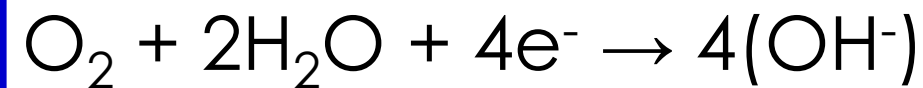
## Examples



Reduction of hydrogen ions in an acid solution



Reduction reaction in an acid solution containing dissolved oxygen



Reduction reaction in a neutral or basic solution containing dissolved oxygen

- ❑ When an electrode (e.g. Zn) is immersed in an aqueous solvent containing its own ions. Some of  $\text{Zn}^{2+}$  metal ions leave the electrode by giving their valence electrons to the metal electrode and go into the solvent. **The negative charge builds up in the electrode.**
- ❑ The **solvent becomes +ve** and the opposing electrical layers leads to a **dynamic equilibrium** wherein there is no further (net) dissolution of the electrode.
- ❑ If the metal-ion concentration in the solution is kept fixed, normally one mole of ions per litre (having unit activity) at a constant temperature ( $25^\circ\text{C}$ ) the electric potential of the metal now becomes the property of the metal. **It gives the tendency of a metal to go into solution i.e. to corrode, or to get plated.**
- ❑ The electrode potential is measured with the electrode in contact with a solution containing an unit concentration of ions of the same metal with the **standard hydrogen electrode** as the counter electrode (whose potential is taken to be zero).



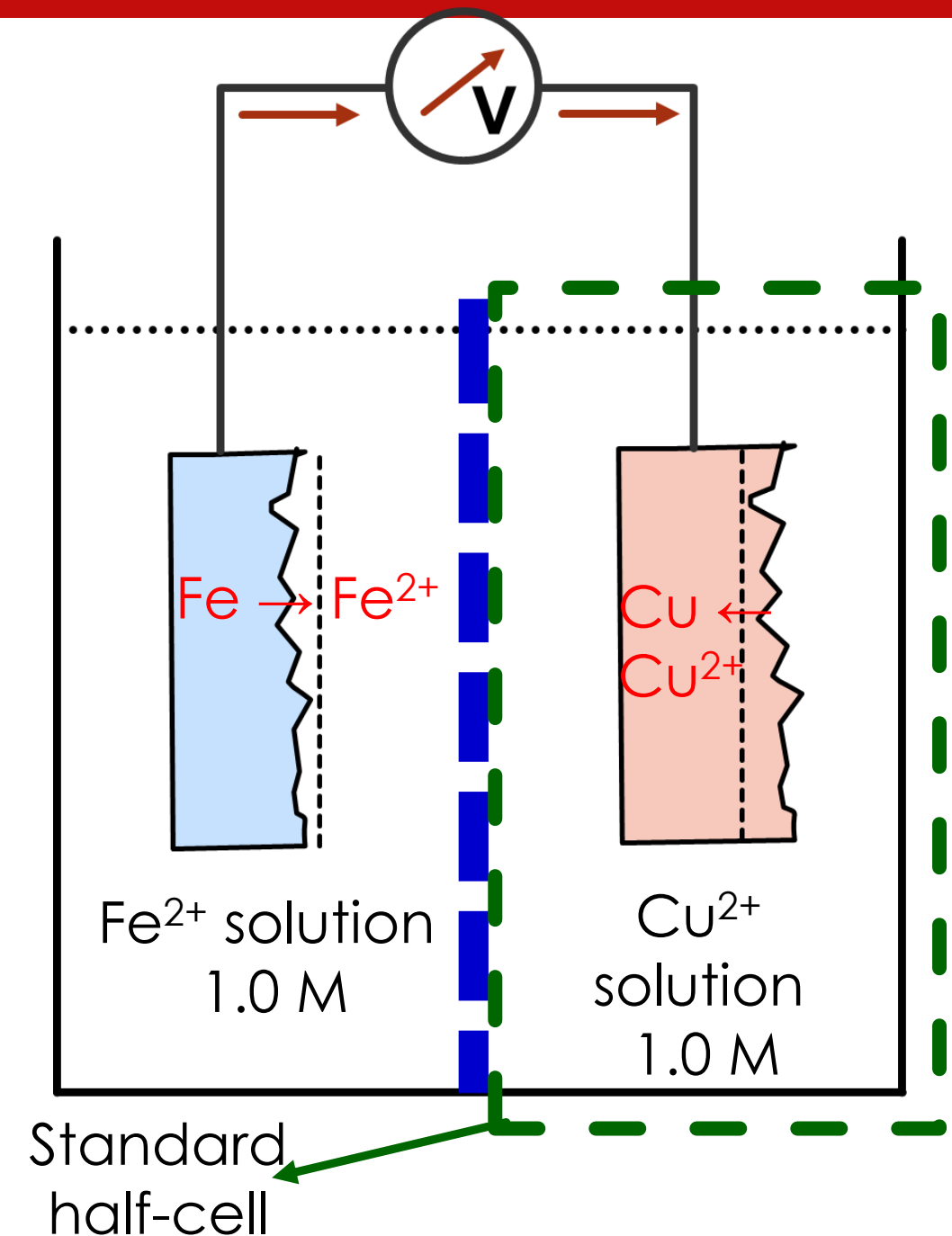
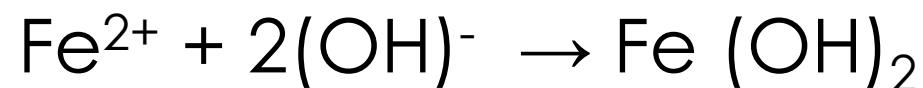
# Galvanic Cell

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Liberated electrons are move from anode to the higher potential area the cathode (copper side), where they consumed in the cathodic reaction:



The cations  $\text{Fe}^{2+}$  migrate through the electrolyte towards the cathode and the anions  $(\text{OH})^{-}$  migrate towards the anode. Both combine and produce an insoluble compound by the reaction:




Simple Galvanic Cell



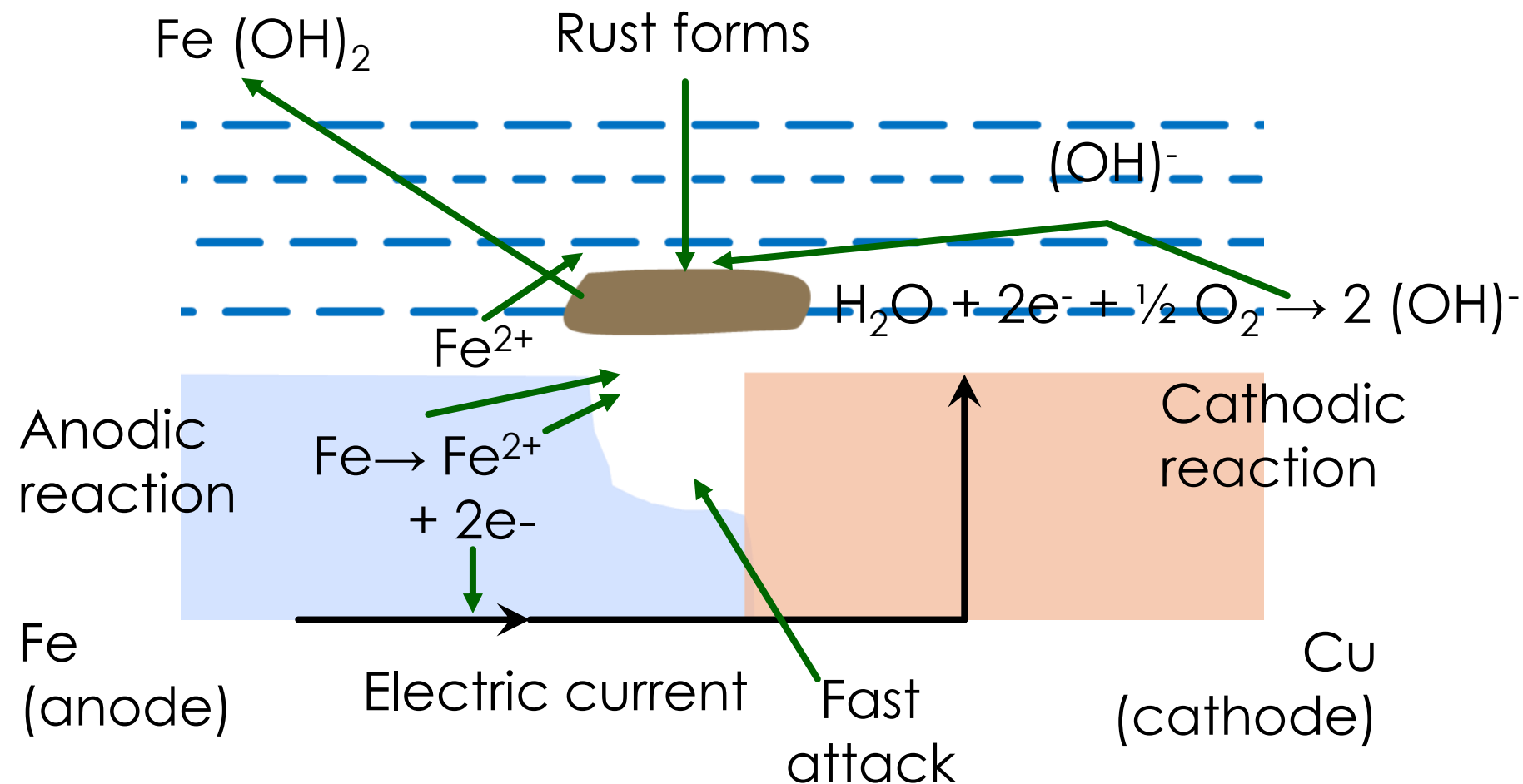
# Standard electrode potential of metals

Standard potential at 25°C

	System	Potential in V
<i>Noble end</i>	<i>Au / Au<sup>3+</sup></i>	<i>+1.5</i>
 Increasing propensity to dissolve	Ag / Ag <sup>+</sup>	+0.80
	Cu / Cu <sup>2+</sup>	+0.34
	H <sub>2</sub> / H <sup>+</sup>	0.0
	Pb / Pb <sup>2+</sup>	-0.13
	Ni / Ni <sup>2+</sup>	-0.25
	Fe / Fe <sup>2+</sup>	-0.44
	Cr / Cr <sup>3+</sup>	-0.74
	Zn / Zn <sup>2+</sup>	-0.76
	Al / Al <sup>3+</sup>	- 1.66
<i>Active end</i>	<i>Li / Li<sup>+</sup></i>	<i>-3.05</i>

Not all metallic materials oxidize to form ions with same degree of ease.

Two dissimilar metals Fe and Cu are placed in electrical contact in the presence of an electrolyte (e.g. salt solution having dissolved oxygen). The potential of each metal is related to its relative tendency to dissolved .



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Fe has lower potential acts as anode and dissolves (i.e. corrodes) by reaction,



Anodic reactions are oxidation reactions.

Cathodic reaction,  $\text{H}_2\text{O} + 2\text{e}^{-} + \frac{1}{2} \text{O}_2$  (dissolved oxygen)  $\rightarrow 2 (\text{OH})^{-}$   
Cathodic reactions are reduction reactions.

More anodic metal corrodes.



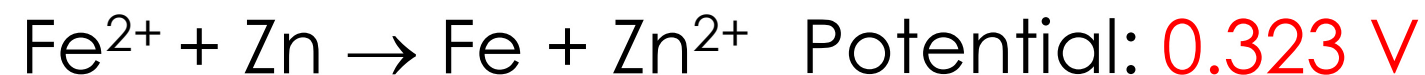
When externally connected, electrons generated from the oxidation of iron flow to the copper cell in order that  $\text{Cu}^{2+}$  be reduced.

In addition, there will be some net ion motion from each cell to the other across the membrane.

This is called a *galvanic couple* - two metals electrically connected in a liquid *electrolyte* wherein one metal becomes an anode and corrodes, while the other acts as a cathode.

An electric potential or voltage will exist between the two cell halves.

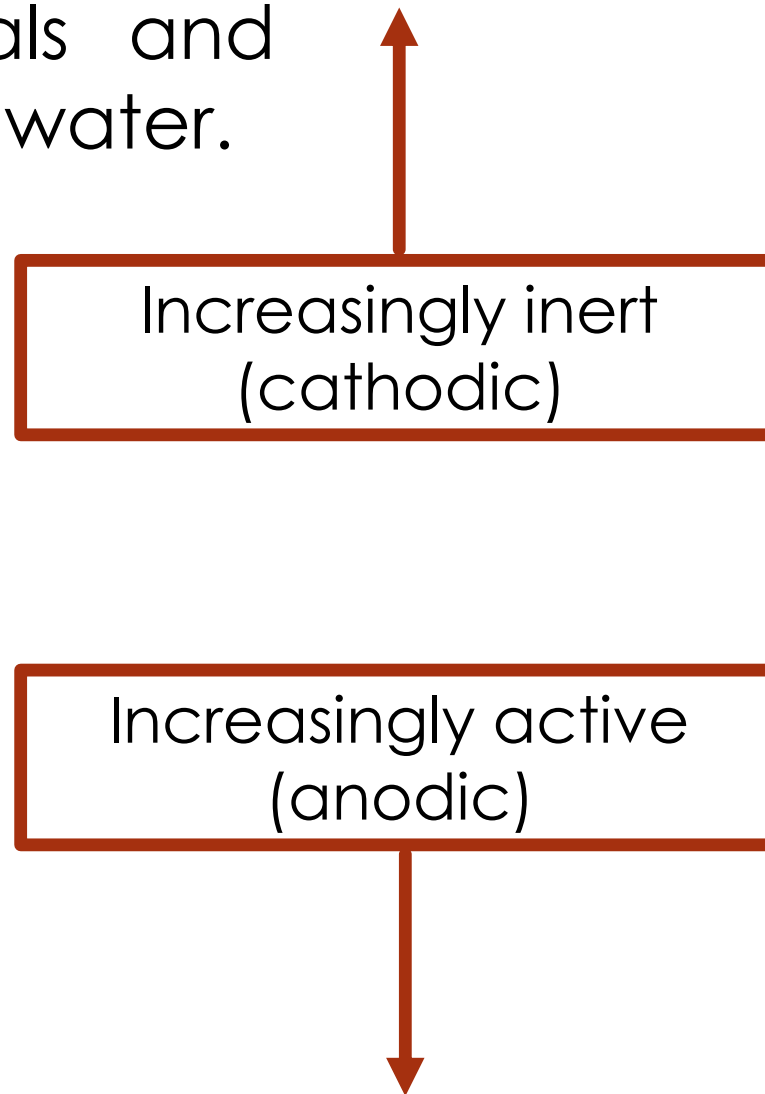
For example:



Various electrode pairs have different voltages !

# Galvanic Cell

Galvanic series: This represents the relative reactivity of a number of metals and commercial alloys in seawater.



Platinum  
Gold  
Graphite  
Titanium  
Silver  
316 Stainless steel (passive)  
304 Stainless steel (passive)  
Inconel (80Ni-13Cr-7Fe) (passive)  
Nickel (passive)  
Monel (70Ni-30Cu)  
Copper-nickel alloys  
Copper  
Brasses (Cu-Zn alloys)  
Inconel (active)  
Nickel (active)  
Tin  
Lead  
316 Stainless steel (active)  
304 Stainless steel (active)  
Cast iron  
Iron and steel  
Aluminium alloys  
Cadmium  
Commercially pure aluminium  
Zinc  
Magnesium and magnesium alloys

The corrosion rate, or the rate of material removal as a consequence of the chemical action, is an important corrosion parameter. It is the thickness loss of material per unit of time.

$$CPR = \frac{KW}{\rho At}$$

where  $W$  = weight loss after exposure time  $t$ ;  $\rho$  = density, and  $A$  = exposed specimen area,  $K$  is a constant.

There is an electric current associated with electrochemical corrosion reactions.

$$r = \frac{i}{nF}$$

where

$i$  = current density (the current per unit surface area of material corroding)  
 $n$  = number of electrons associated with the ionization of each metal atom  
 $F$  = 96,500 C/mol.



- Some normally active metals and alloys, under particular environmental conditions, lose their chemical reactivity and become extremely inert. E.g., chromium, iron, nickel, titanium, and many of their alloys.
- This passive behaviour results from the formation of a highly adherent and very thin oxide film on the metal surface, which serves as a protective barrier to further corrosion.
- e.g. Ashoka Pillar, Delhi



1. Corrosion is a loss due to electrochemical reactions in the environment.
2. Electrochemical reaction between two different metals gives rise to corrosion.
3. The corrosion in galvanic series is more when the potential difference is more.
4. Passivation is loss of chemical reactivity of the material.