Module – 3 – Corrosion

- Dry and wet corrosion
- detrimental effects to buildings, machines, devices & decorative art forms, emphasizing Differential aeration, Pitting, Galvanic and Stress corrosion cracking;
- Factors that enhance corrosion and choice of parameters to mitigate corrosion.

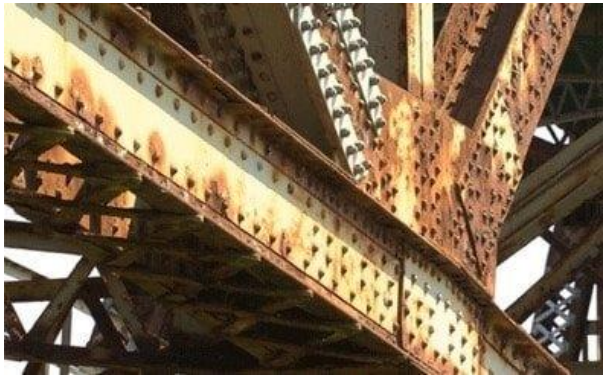
Reference – Corrosion Engineering by Mars G. Fontana

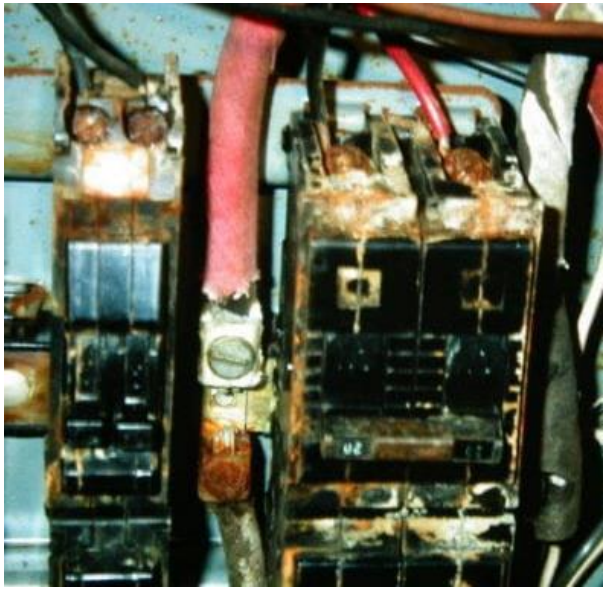
Corrosion

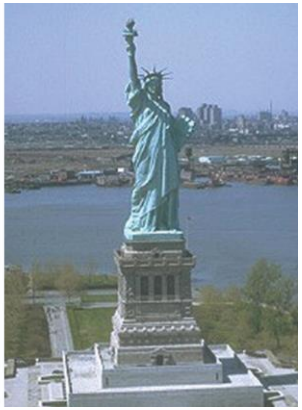
Corrosion is the deterioration of materials/components due to interaction with the environment.

Types of Interaction that leads to corrosion

- **Physical** — e.g.: Flow of liquid on a metal surface
- **Chemical** — e.g.: Iron at high temperature in air
- **Electrochemical** — e.g.: Iron in water







Electric Power Industry

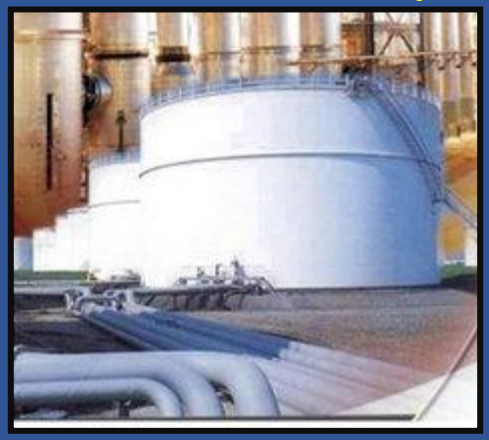
Transmission pipelines

Highway Bridges

Pulp & Paper

Defense

Petrochemical Industry



No industry is exempted

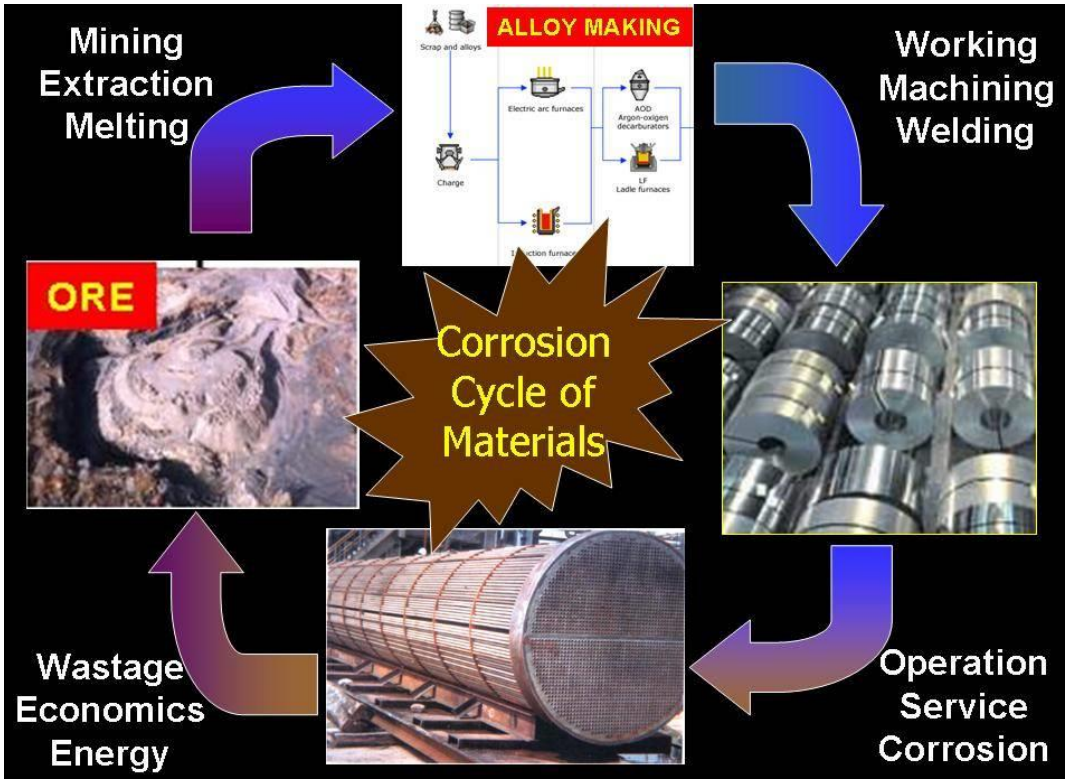
Nuclear Power Industry

Consequences of Corrosion

- Loss of aesthetic value
- Wall thinning in components, some times to unsafe levels
- Leaking of process fluids
- Loss of function, property or failure of components
- Economic loss
- Injury or loss of life
- Environmental damage

Why does corrosion takes place?

- Metals exist in nature as ores – stable compounds such as oxides, carbides, sulphides, carbonates, etc.
- Metals are in a higher energy state when compared to the ores – Energy has to be spent for extracting metals from ores
- Metals thus react spontaneously with the environment to revert back to the stable compounds



Types of Corrosion

Dry Corrosion / Chemical Corrosion	Wet Corrosion / Electrochemical Corrosion
Involves direct attack of atmospheric gases on metal in the absence of moisture/liquid phase	Involves electrochemical attack on metals in aqueous environments
Less prevalent	More prevalent
The corrosive media include vapours, gases, etc.	The corrosive media is conducting liquid (aqueous) phase
It is of various types: Oxidation corrosion, corrosion by other gases such as Cl_2, SO_2, H_2S, NO_x and liquid metal corrosion	It is of three types: differential metal corrosion, differential aeration corrosion and crevice corrosion

Dry Corrosion

- Dry corrosion occurs in the absence of a liquid phase or above the dew point of the environment. Vapors and gases are usually the corrodents.

- Dry corrosion is most often associated with high temperatures.

- e.g. Attack of Steel by furnace gases

Wet Corrosion

- Wet corrosion occurs when a liquid is present. This usually involves aqueous solutions or electrolytes and accounts for the greatest amount of corrosion by far.

- Types –
  - Differential metal corrosion
  - Differential aeration corrosion
  - Crevice corrosion

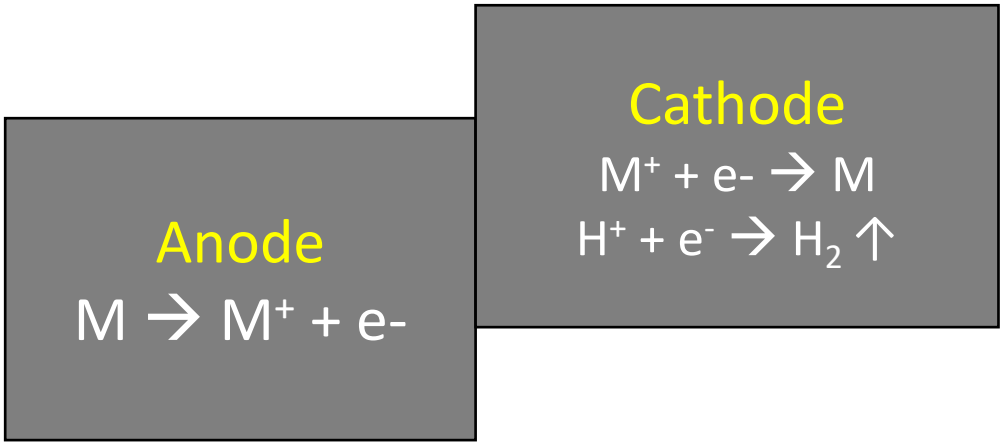
- e.g. Corrosion of Steel in water

Electrochemical Aspects of Corrosion

Electrochemical reaction

Oxidation takes place at anode – Metal dissolution

Reduction takes place at cathode – Gas evolution



Redox reactions

Metal atoms characteristically lose up electrons in what called an oxidation reaction.

$$M \rightarrow M^{n+} + ne^-$$

$$Fe \rightarrow Fe^{2+} + 2e^-$$

$$Al \rightarrow Al^{3+} + 3e^-$$

The site at which oxidation takes place is called the anode, oxidation is sometimes called an anodic reaction.

The electrons generated from each metal atom that is oxidized must be transferred to and become a part of another chemical species in what is termed a reduction reaction.

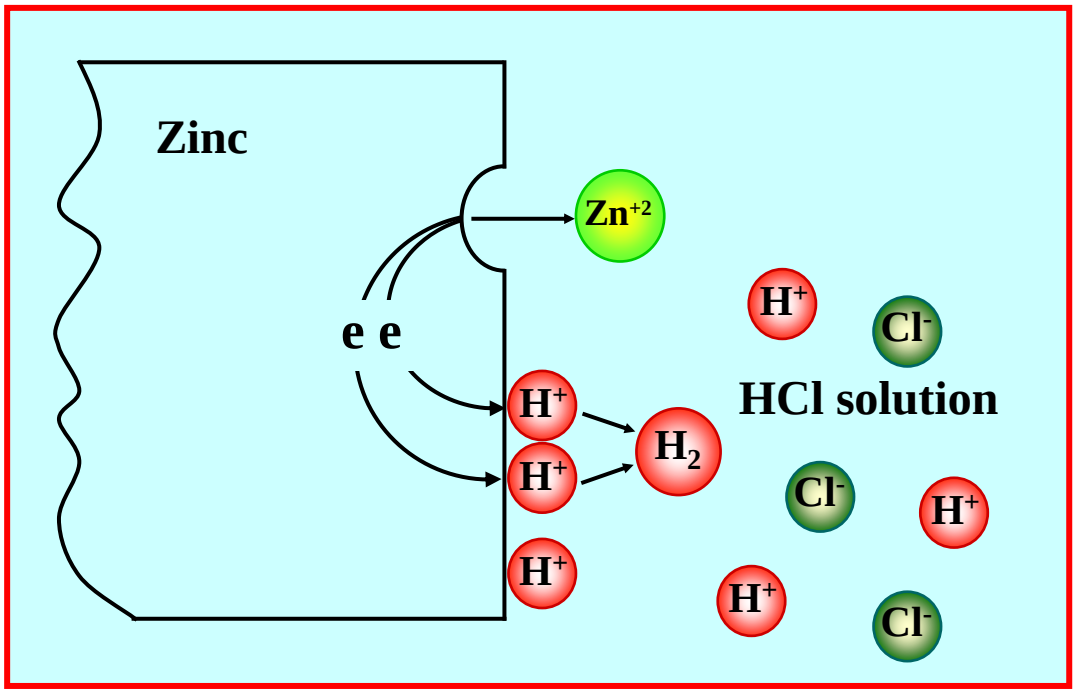
$$2H^+ + 2e^- \rightarrow H_2$$

Metal Corrosion – example

Oxidation (anodic reaction) $Zn \rightarrow Zn^{2+} + 2e^-$

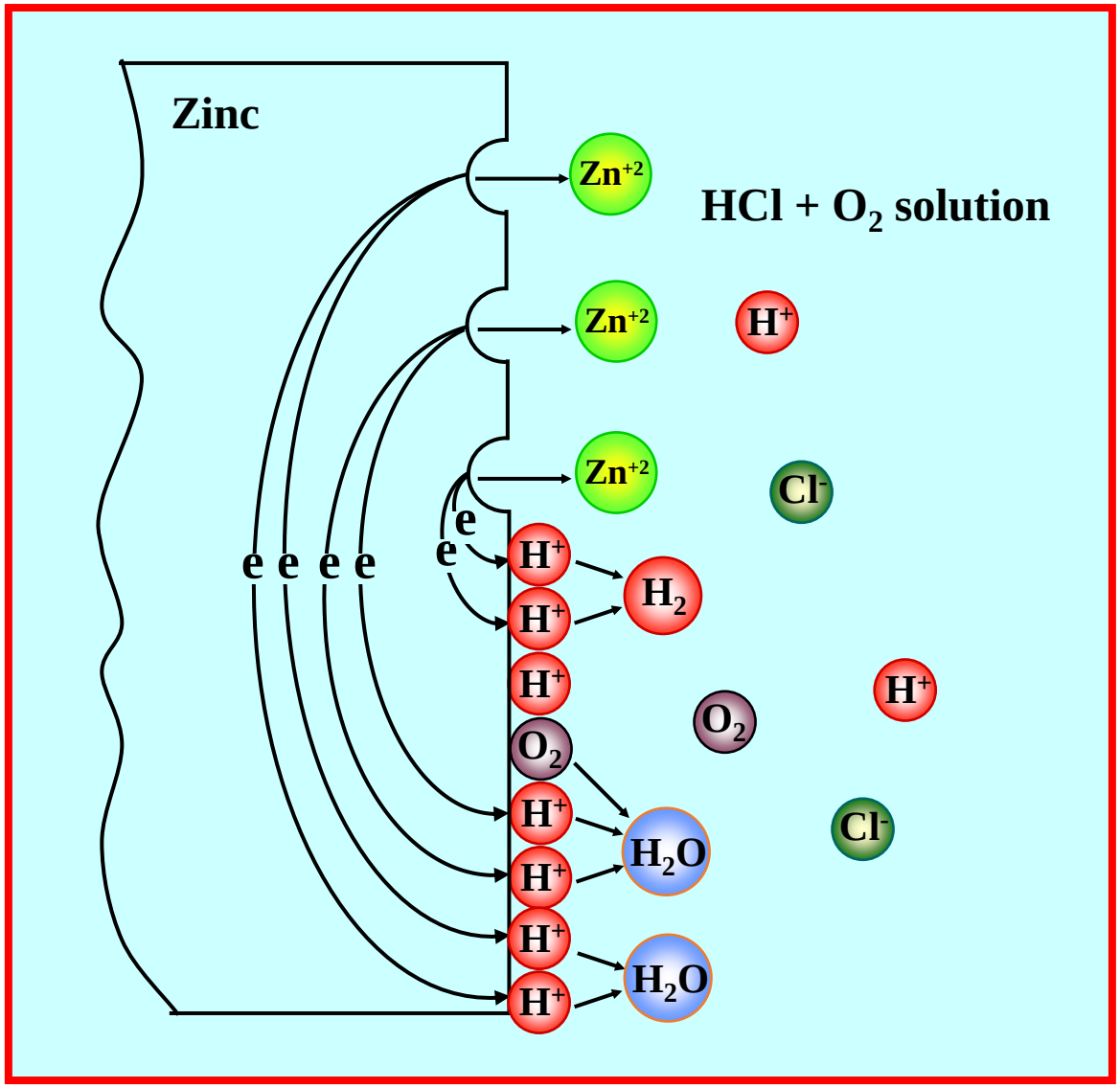
Reduction (cathodic reaction) $2H^+ + 2e^- \rightarrow H_2$

$$Zn + 2H^+ \rightarrow Zn^{2+} + H_2$$



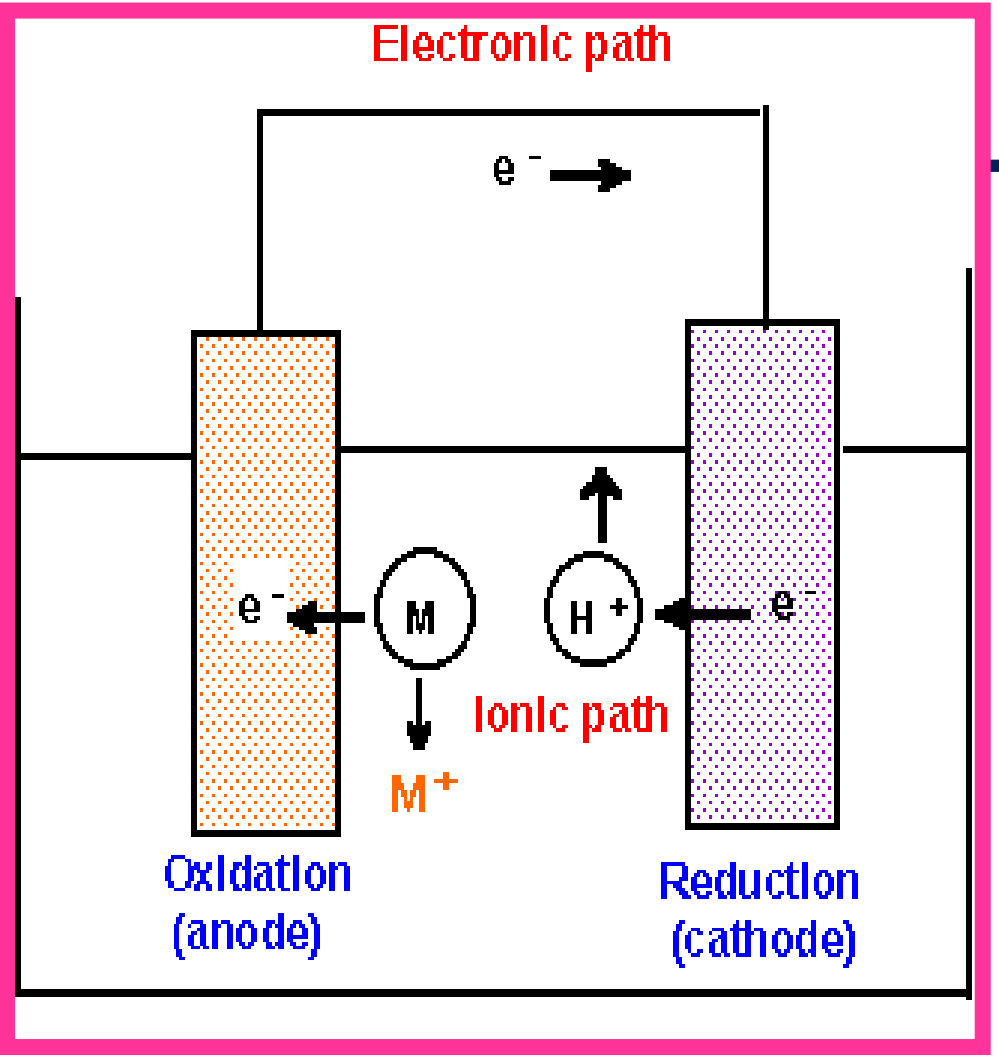
Electrochemical reaction in air free HCl

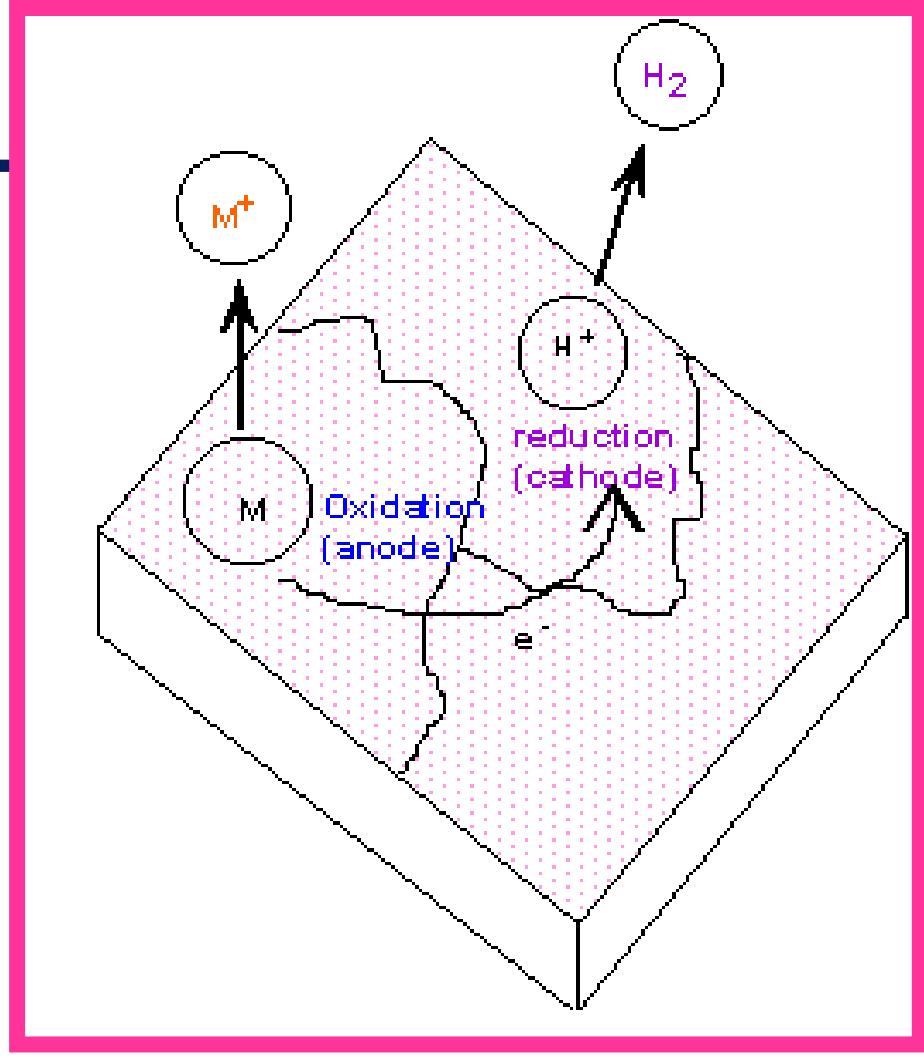
Metal Corrosion – example



Electrochemical reaction in aerated HCl

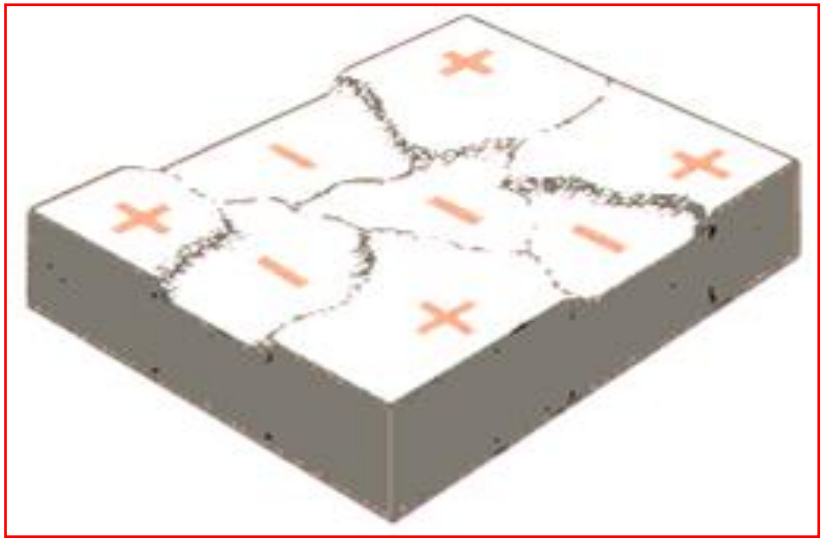
Anodic reaction same, Cathodic reaction will lead to formation of Water instead of H_2

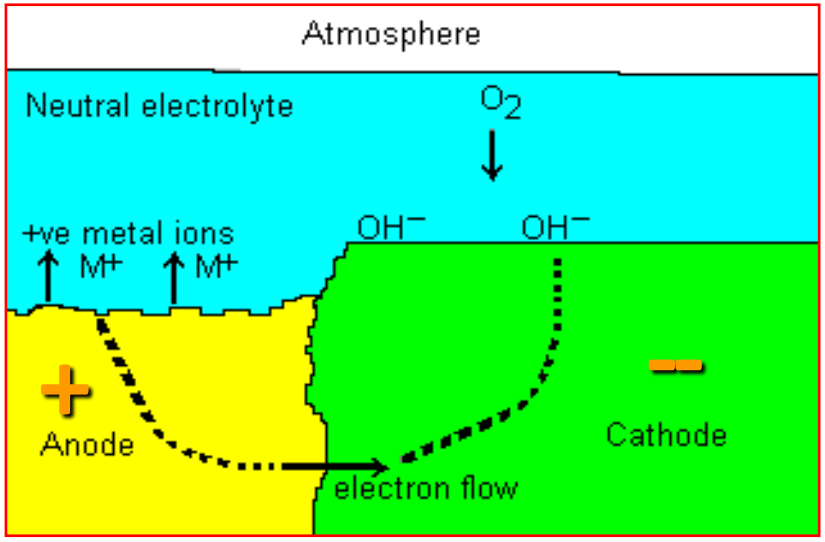


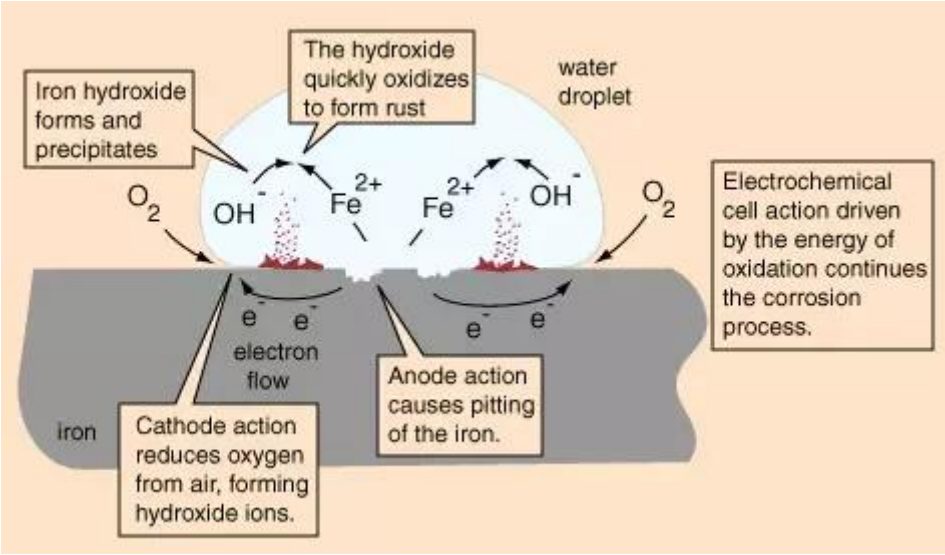


Corrosion occurs by metal dissolution (oxidation) at anode;
electron released thereby is consumed by reduction of ions at cathode;
Net charge is always near zero with anodic current ↔ cathodic current

Corrosion of metals and alloys

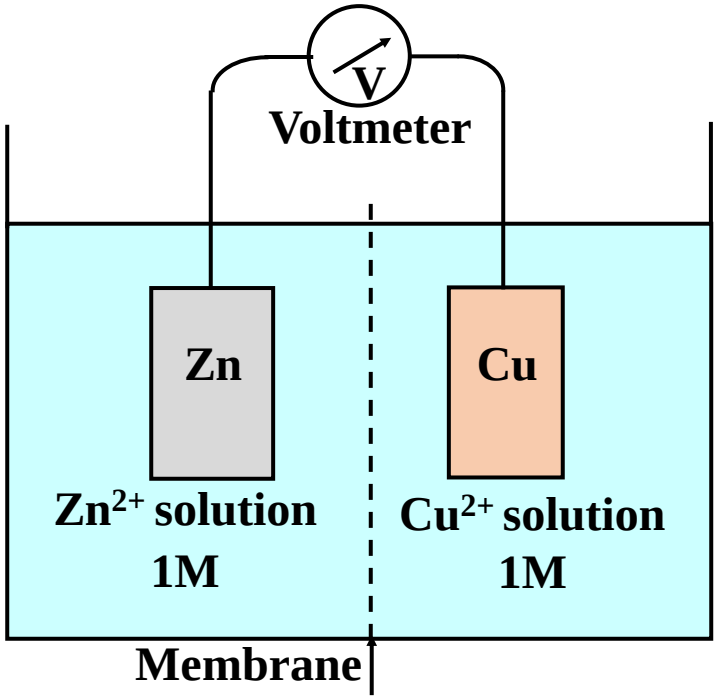






Electrode potential

Not all materials oxidize to form ions with the same degree of ease. Consider the electrochemical cell shown in the figure:



$$Zn \rightarrow Zn^{2+} + 2e^{-}$$

$$Cu^{2+} + 2e^{-} \rightarrow Cu$$

$$Cu^{2+} + Zn \rightarrow Cu + Zn^{2+}$$

On the left-hand side is a piece of pure Zinc immersed in a solution containing Zn^{+2} ions of 1 M concentration. The other side of the cell consists of a pure copper electrode in a 1M solution of Cu^{2+} ions.

Electrode potential

➢The electromotive force (emf) series is generated by coupling to the standard hydrogen electrode half-cells for various metals and ranking them according to measured voltage.

The Nernst Equation

$$E = E_0 - \frac{0.0592}{n} \log(C_{ion})$$

	Reaction	Std. Potential, V vs SHE
Noble	$Au^{3+} + 3e^- = Au$	+1.498
	$Cl_2 + 2e^- = 2Cl^-$	+1.358
	$O_2 + 4H^+ + 4e^- = 2H_2O$	+1.229
	$Pt^{3+} + 3e^- = Pt$	+1.200
	$O_2 + 2H_2O + 4e^- = 4OH^-$	+0.820
	$Ag^+ + e^- = Ag$	+0.799
	$Fe^{3+} + e^- = Fe^{2+}$	+0.771
	$Cu^{2+} + 2e^- = Cu$	+0.337
	$Sn^{4+} + 2e^- = Sn^{2+}$	+0.150
	$2H^+ + 2e^- = H_2$	0.000
	$Pb^{2+} + 2e^- = Pb$	-0.126
	$Sn^{2+} + 2e^- = Sn$	-0.136
	$Ni^{2+} + 2e^- = Ni$	-0.250
	$Co^{2+} + 2e^- = Co$	-0.277
	$Fe^{2+} + 2e^- = Fe$	-0.440
	$Cr^{3+} + 3e^- = Cr$	-0.744
	$Zn^{2+} + 2e^- = Zn$	-0.763
Active	$Zr^{4+} + 4e^- = Zr$	-1.53
	$Ti^{2+} + 2e^- = Ti$	-1.63
	$Al^{3+} + 3e^- = Al$	-1.662
	$Mg^{2+} + 2e^- = Mg$	-2.363

Strong Oxidizing agent

Strong Reducing agent

Corrosion mechanism

Four essential requirements for electrochemical corrosion

- **Anode** - Metal oxidation takes place releasing electrons
- **Cathode** – electrons are consumed to form reduced species
- **Electrical contact** between anode and cathode for electron transport
- **Electrolyte** for ionic transport between the electrodes