Al – Mustaqbal University College
Department of Chemical Engineering and Petroleum Industries

BASIC CONCEPTS IN CORROSION

Dr. KHALID OMRAN

4th Stage

2020-2021

1. Corrosion Cell

For corrosion to take place, the formation of a *corrosion cell* is essential. A corrosion cell is essentially comprised of the following four components (see Fig. 1).

- **Anode**
- **Cathode**
- **Electrolyte**
- **Metallic path**.

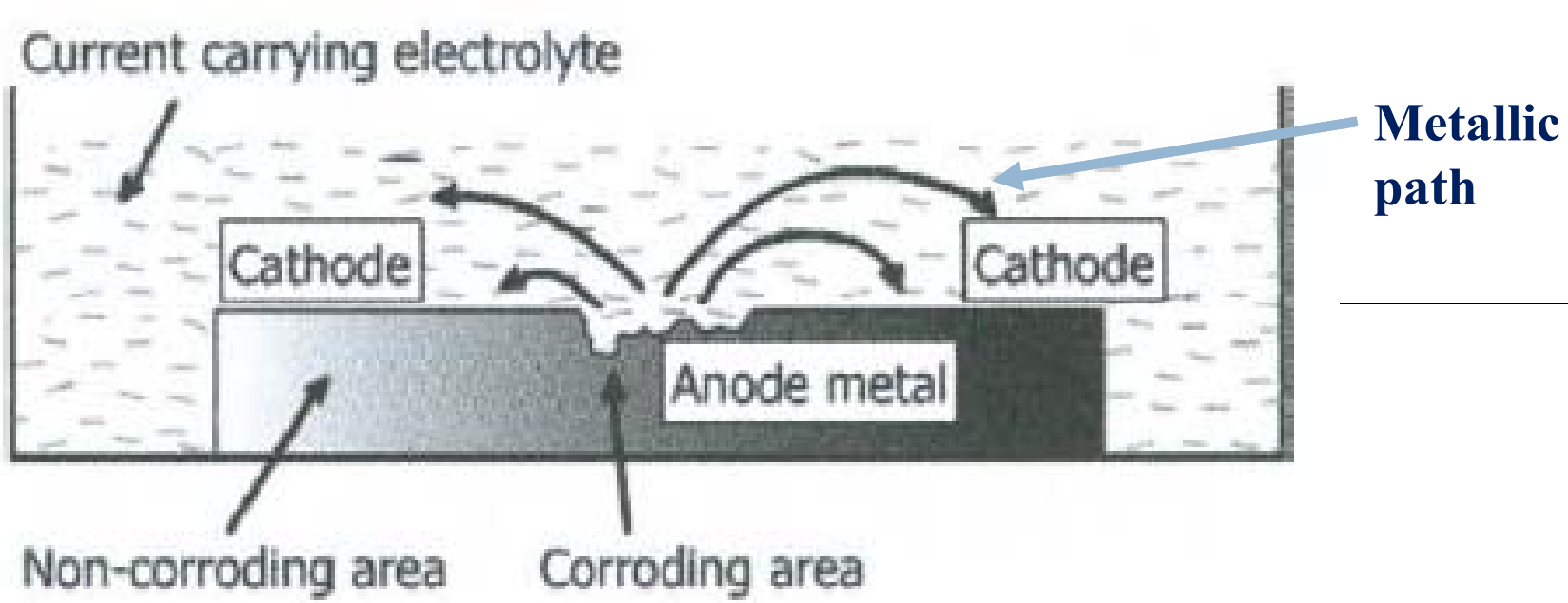


Figure 1	Corrosion cell in action

Anode

- Is one of the two metal electrodes in an electrolytic cell, represented as the **negative** terminal of the cell.
- **Electrons** are released at the anode, which is the more reactive metal.
- **Electrons** are **insoluble** in aqueous solutions and they only move, through the wire connection into the cathode.
- For example, zinc acts as the anode in Daniel cell. (see Fig. 2)

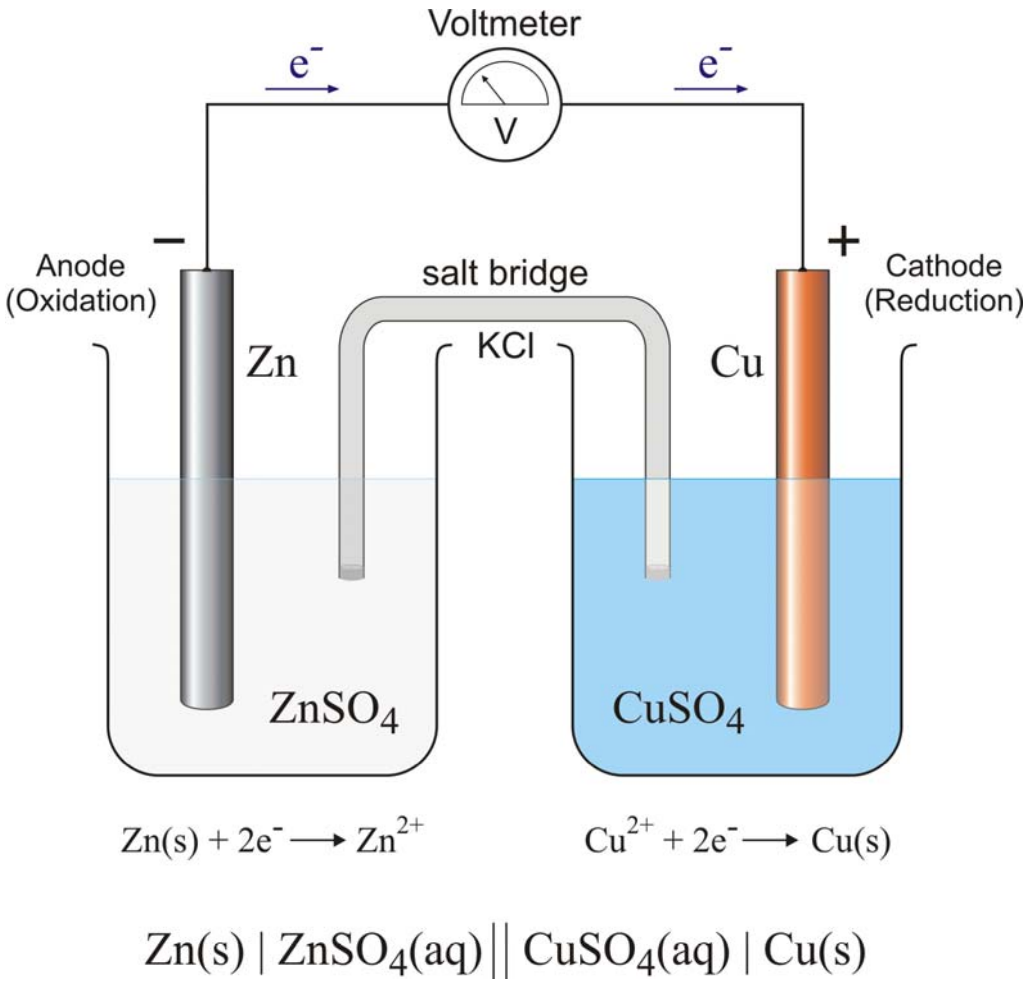


Figure 2 Daniel cell

Cathode

- Is one of the two electrodes in an electrolytic cell represented as a **positive** terminal of a cell.
- **Reduction** takes place at the cathode and electrons are consumed.
- Example, carbon electrode in a battery, copper electrode in a Daniel cell. The accompanying **Fig. 3** shows the reduction of hydrogen ion. The electron is always a reducing agent.

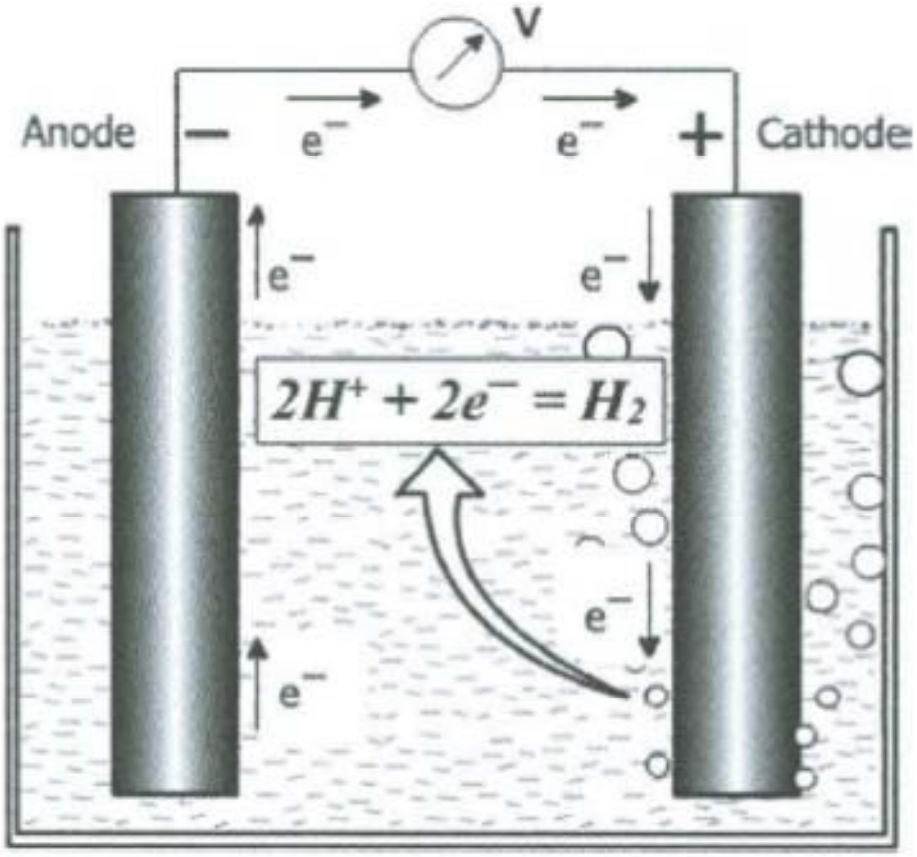


Figure 3 the reduction of hydrogen ion.

Electrolyte

- It is the ***electrically conductive solution*** (e.g. salt solution) that must be present for corrosion to occur.
- ***Note that*** pure water is a ***bad conductor of electricity***.

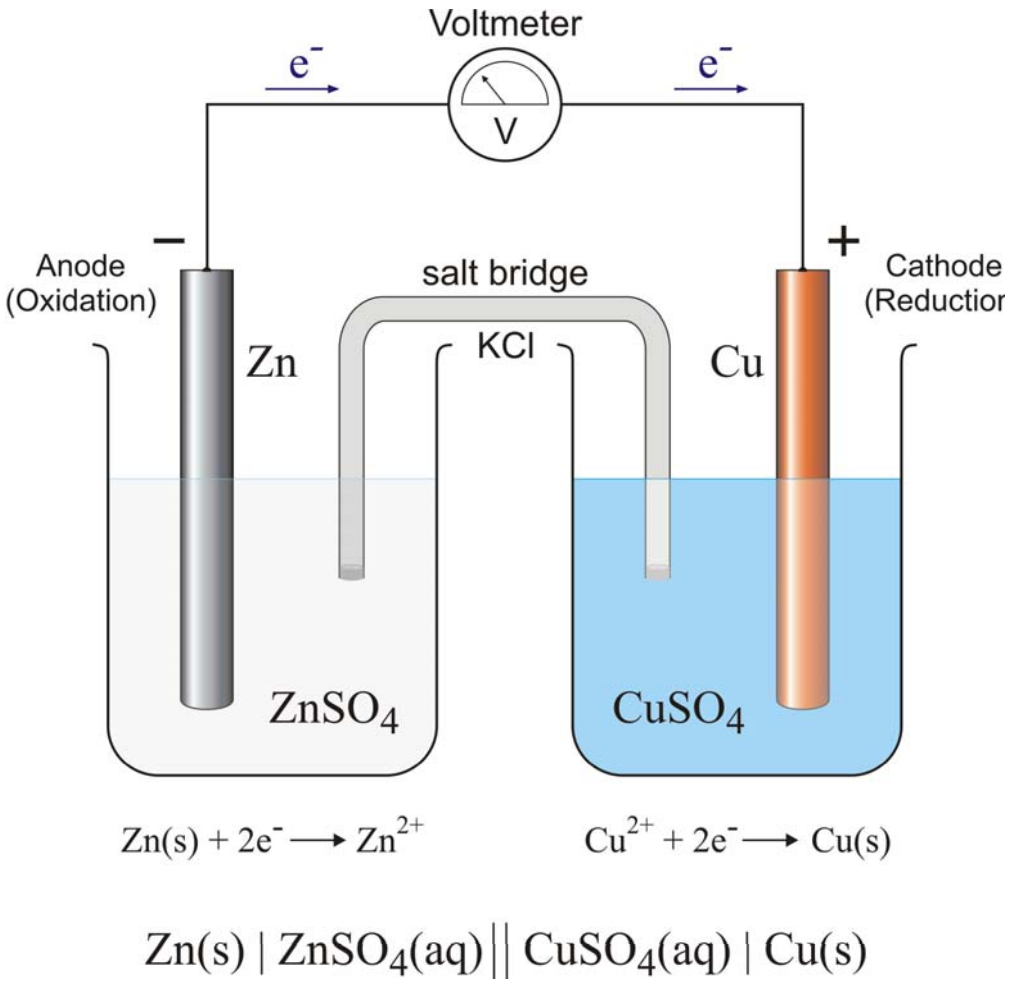


Figure 2 Daniel cell

Metallic Path

- The **two electrodes** are connected externally by a **metallic conductor**.
- In the metallic conductor, ***'conventional'*** current flows from **(+) to (—)** which is really **electrons** flowing from (—) to (+).
- Metals provide a path for the flow of conventional current which is actually passage of electrons in the opposite direction.

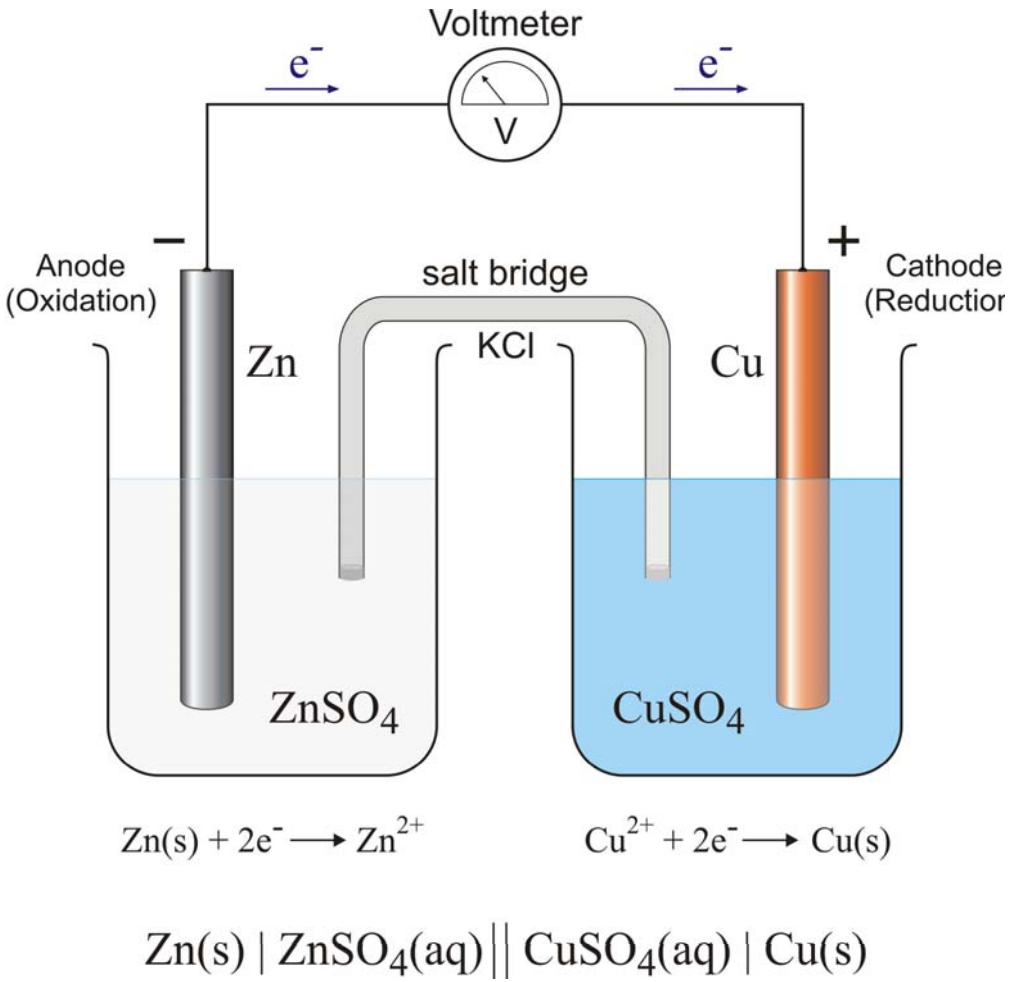


Figure 2 Daniel cell

Current Flow

- Conventional **current** flows from **anode** (—) to **cathode** (+) as **Zn^{++} ions** through the solution.
- The current is carried by these **positive charged ions.**
- The circuit is completed by passage of **electrons** from the **anode (—)** to the **cathode (+)** through the external metallic wire circuit **(outer current)**.

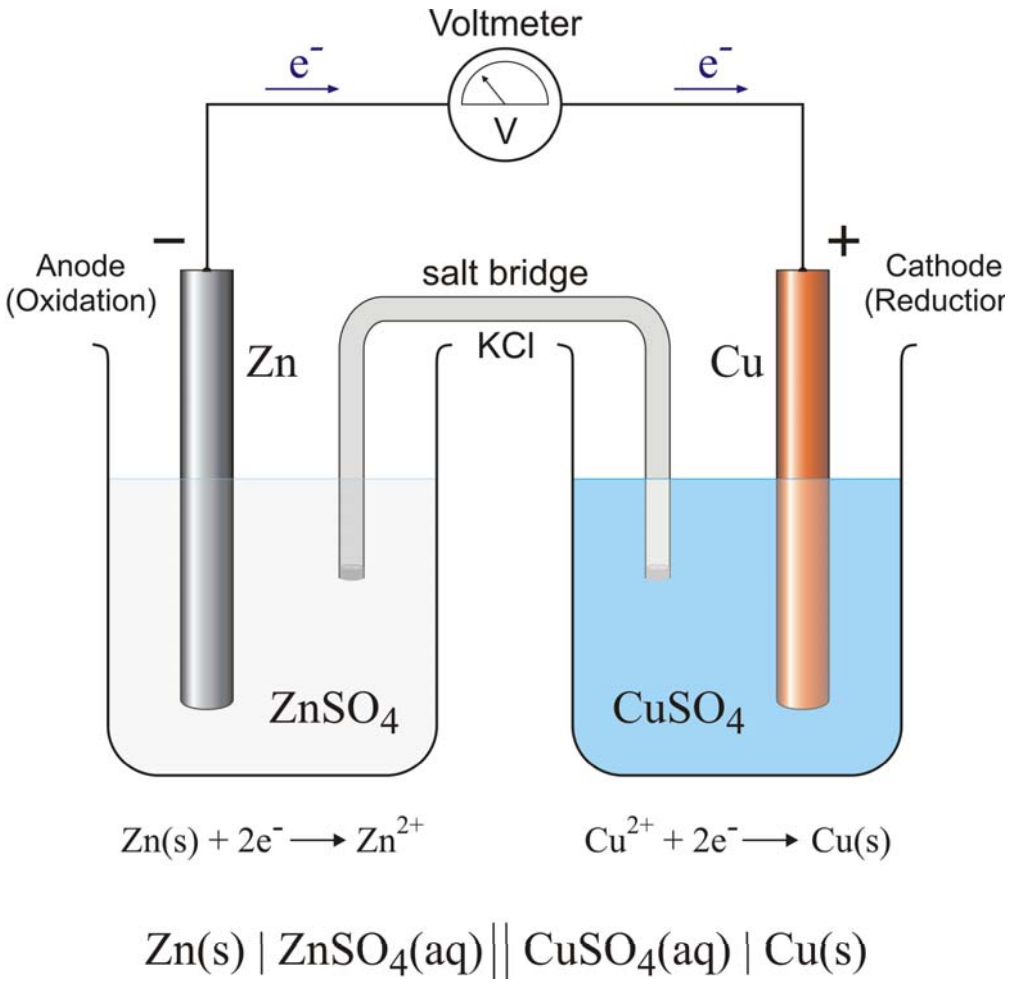


Figure 2 Daniel cell

Electron Flow

Although the **anode** (e.g. Fe or Zn) is the most **negative** of the two metals in the cell, this reaction does not occur there because its surface is emanating Fe^{++} ions which repel H^{+} ions from discharging there.

The circuit is completed by **negative ions (—)** which migrate from the **cathode (+),** through the electrolyte, towards the **anode (—)**.

Electron Flow

Current flow in an electrochemical cell is shown in Fig. 4.

Anions: Migrate towards the anode (OH^-) but precipitate as $Fe(OH)_2$ before reaching it.

Cations: Migrate towards the cathode (Fe^{2+}).

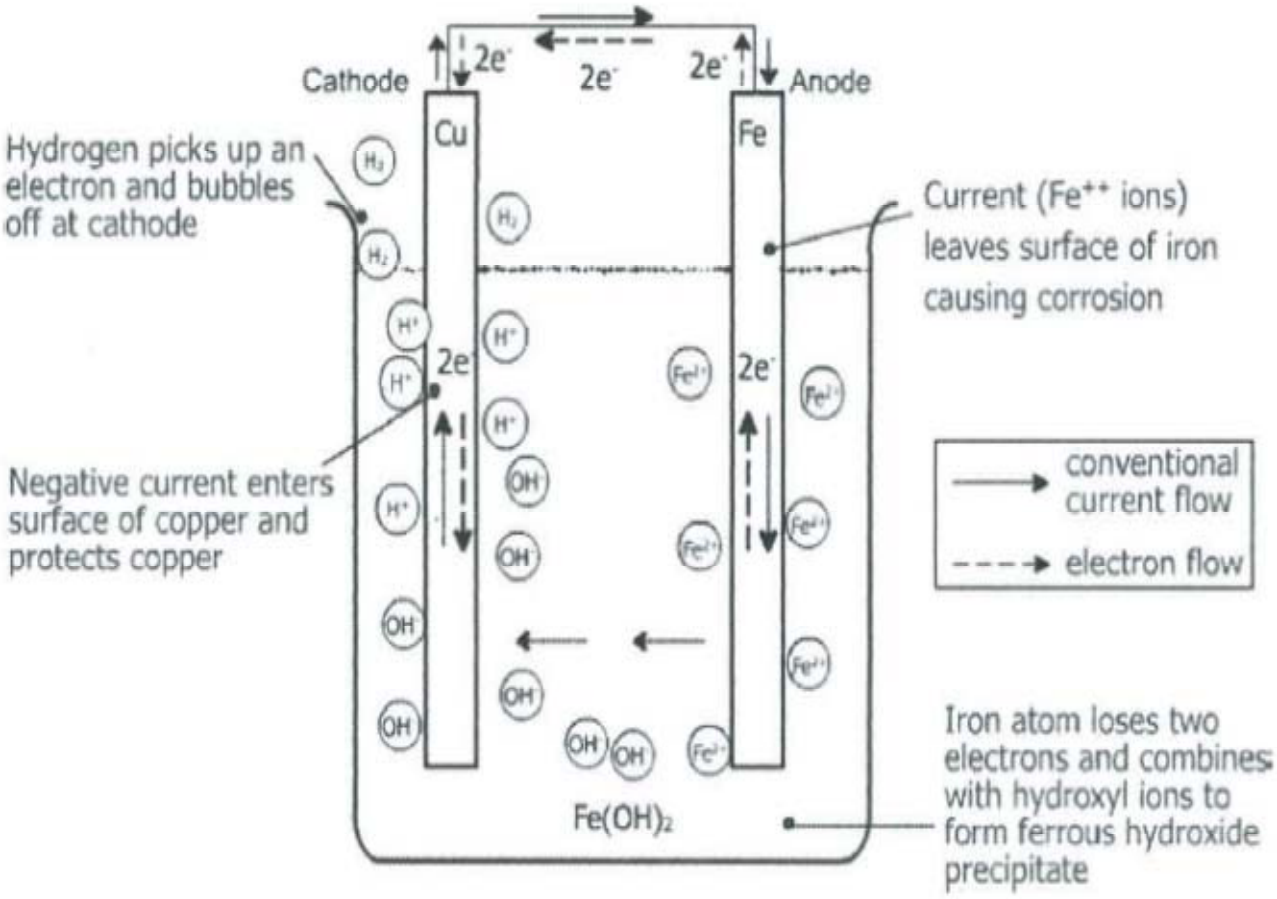


Figure 4 Current flow in an electrochemical cell.

2. Anodic And Cathodic Reactions

The anode is the area where metal is lost.

It represents the entry of metal ion into the solution, by **dissolution**, **hydration** or by **complex formation**. It also includes **precipitation** of metal ions at the metal surface.

For example,

$$Fe^{2+} + 2OH^{-} \rightarrow Fe(OH)_2$$

Ferrous hydroxide or rust formation on steel surface is a common example.

Anodic reaction in terms is written as of electron transfer.

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

3. Types of Corrosion Cells

There are several types of corrosion cells:

- Galvanic cells
- Concentration cells
- Electrolytic cell
- Differential temperature cells.

3.1 Galvanic cells

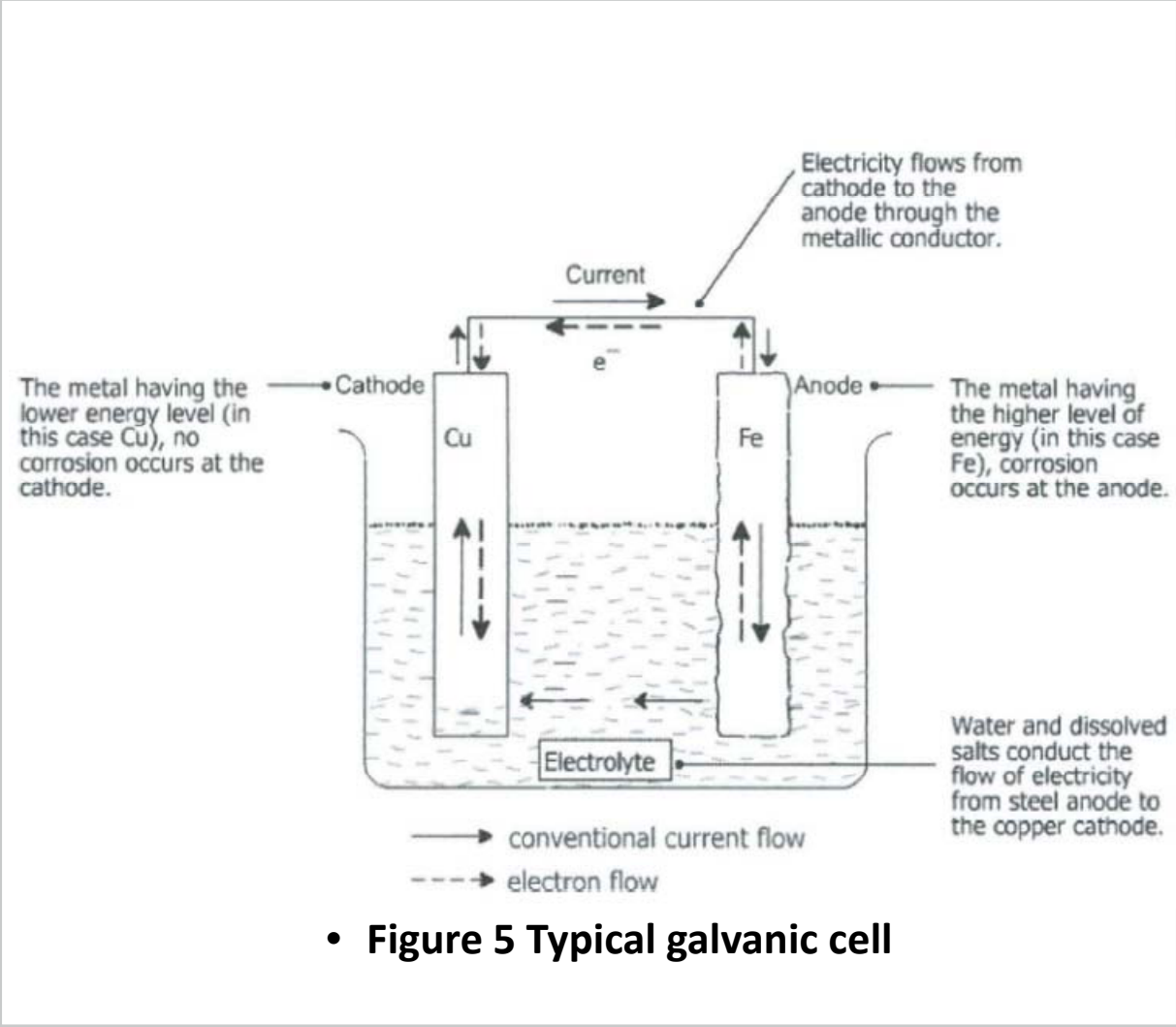
The galvanic cell may have an **anode** or **cathode** of **dissimilar** metals in an **electrolyte** or **the same metal** in **dissimilar** conditions in a **common electrolyte**.

3.1 Galvanic cells

For example, steel and copper electrodes immersed in an electrolyte (Fig. 5), represents a galvanic cell.

The more noble metal **copper** acts as the **cathode** and the more active **iron** acts as an **anode**.

Current flows from **iron anode** to **copper cathode** in the electrolyte.



- **Figure 5 Typical galvanic cell**

3.2 Concentration Cells

- This is similar to galvanic cells except with **an anode and cathode of the same metal in a heterogeneous electrolyte**.

Consider the corrosion of a pipe in the soil. Concentration cells may be set up by:

(a) Variation in the amount of oxygen in soils.

(b) Differences in moisture content of soils.

(c) Differences in compositions of the soil.

Concentration cells are commonly observed in underground corroding structures, such as buried pipes or tanks (Fig. 6).

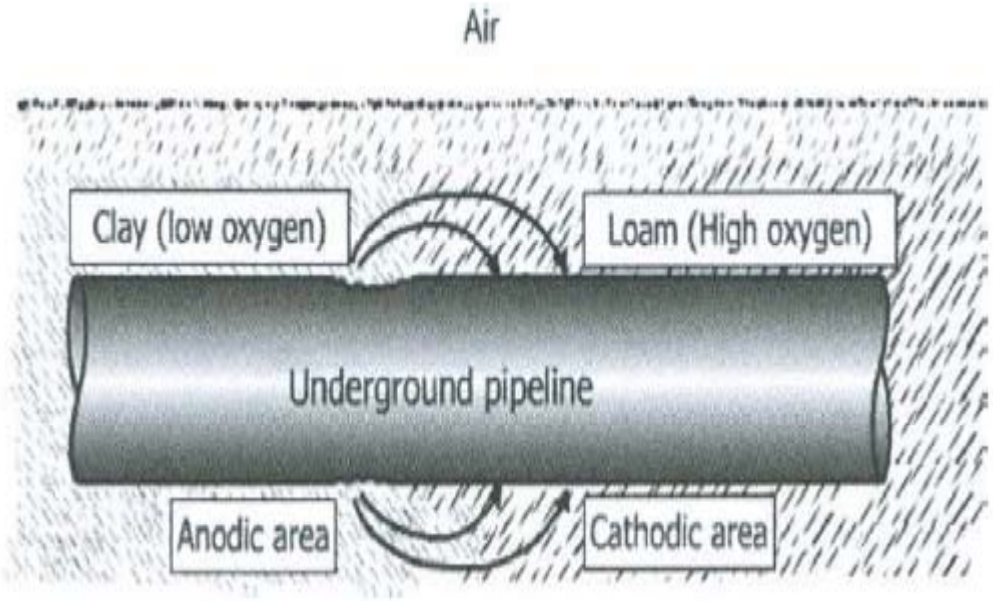


Figure 6 Concentration cell formation in an underground pipeline

3.3 Electrolytic Cell

- This type of cell is formed when an external current is introduced into the system.
- It may consist of all the basic components of **galvanic cells and concentration cells plus an external source** of electrical energy.
- Notice that anode has a (+) polarity and cathode has (—) polarity in an electrolytic cell, where external current is applied. This is the type of cell set up for electrically **protecting** the structures by **cathodic protection**.
- The polarity of an **electrolytic cell** is **opposite** to that in a **galvanic (corrosion)** cell (Fig. 7).

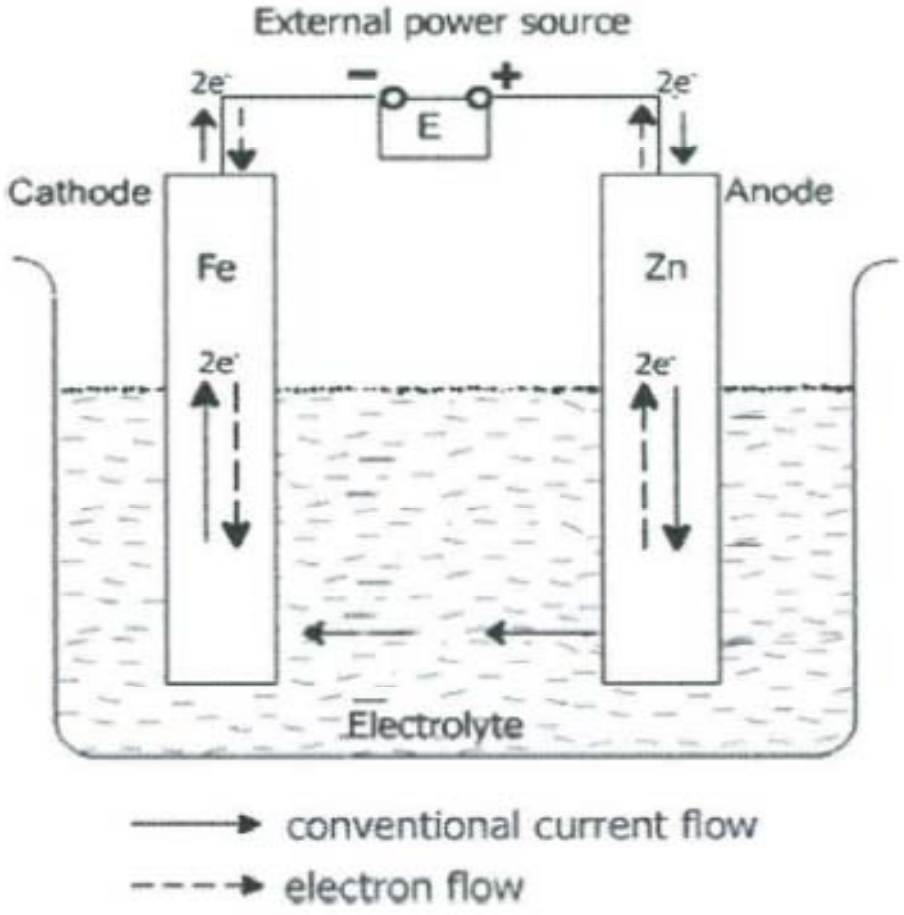


Figure 7 Electrolytic cell. The cathode and anode can be any metal.

4. Mechanism of Corrosion

- Consider a piece of iron exposed to ***humid air*** which acts as an **electrolyte**.
- **Fe^{2+} ions** are released from the **anode** by **oxidation** and **OH^- ions** from the **cathode** by **reduction** on the metal surface (see Fig. 8).

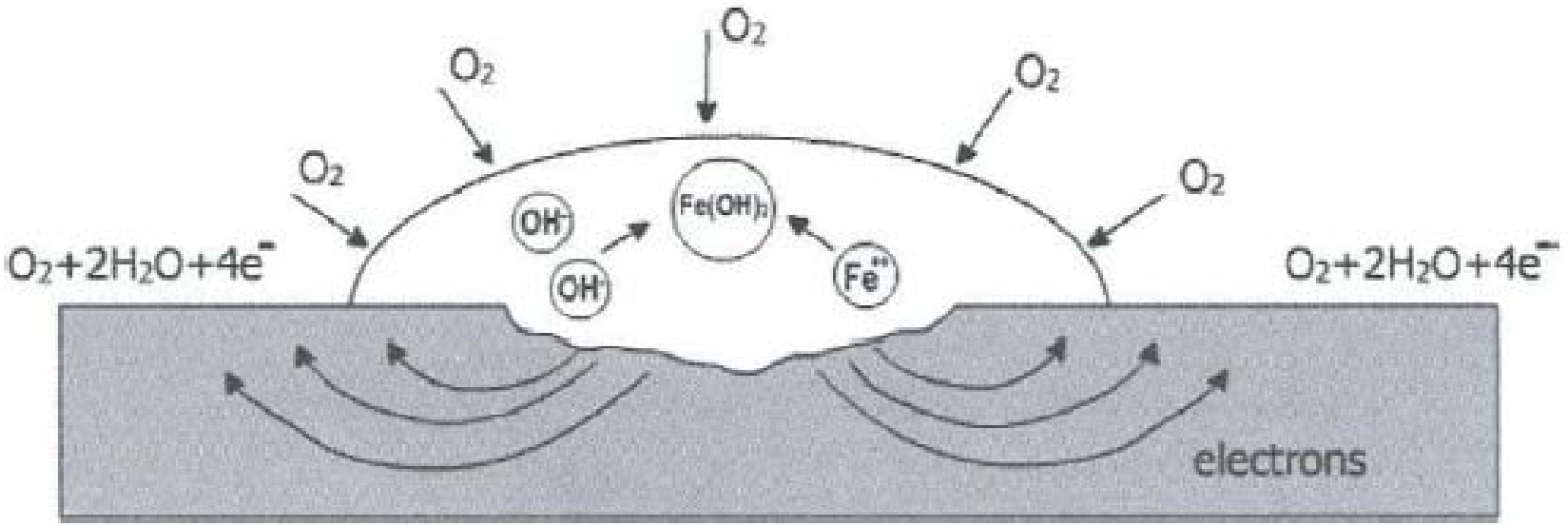


Figure 8 Formation of rust in seawater.

4. Mechanism of Corrosion

Details of reactions involved in the corrosion of iron-based materials is as follows:

(1) $Fe + H_2O \leftrightarrow FeO + 2H^+ + 2e^-$

Forms monolayer of FeO islands

(2) $Fe + 2H_2O \rightarrow Fe(OH)_2 + 2H^+ + 2e$

(3) $3FeO + H_2O \rightarrow Fe_3O_4 + 2H^+ + 2e$

(Black) **(Magnetite)**

(4) $2Fe_3O_4 + H_2O \rightarrow 3(\gamma\text{-}Fe_2O_3) + 2H^+ + 2e^-$

(Brown)

(5) $2[\gamma\text{-}Fe_2O_3] + 3H_2O \rightarrow 6(\gamma\text{-FeOOH})$

(Brown) **(Yellow hydrated oxide)**

The net reaction;

$Fe^{++} + 2OH^- \rightarrow Fe(OH)_2$

(white green precipitate)

A more familiar name of $Fe(OH)_2$ is ***rust***.