

# Electrochemical Cell

## Electrochemical Cell



# Galvanic cell

- A device is used to convert the chemical energy produced in a redox reaction into electric energy is called an electrochemical cell or simply a chemical cell.
- These are also called as **Galvanic cell** or **simply a Voltaic cell**.



# Redox Reactions.

**Reduction**

gain of electron

Loss of oxygen  
Gain of electrons

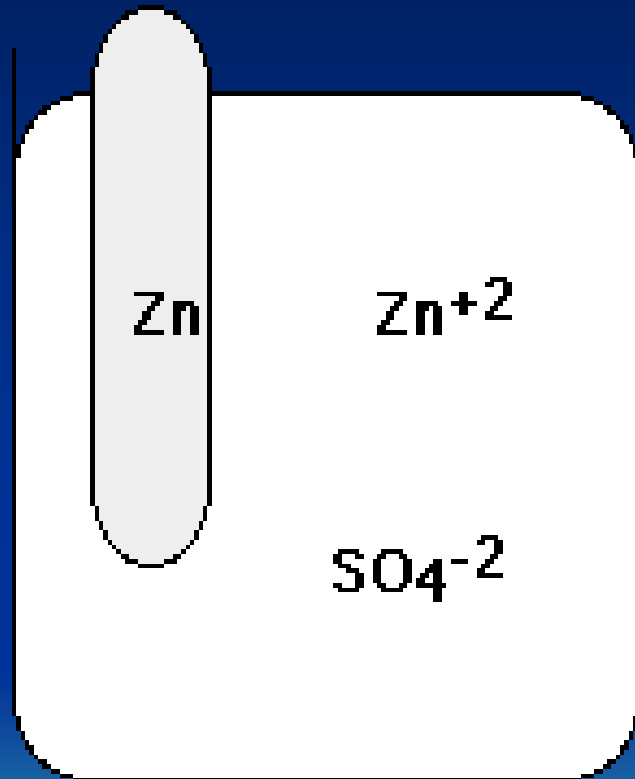
**Oxidation**

loss of electron

- An electrochemical cell is composed to two compartments or **half-cells**, each composed of an electrode dipped in a solution of electrolyte. These half-cells are designed to contain the oxidation half-reaction and reduction half-reaction separately as shown below.

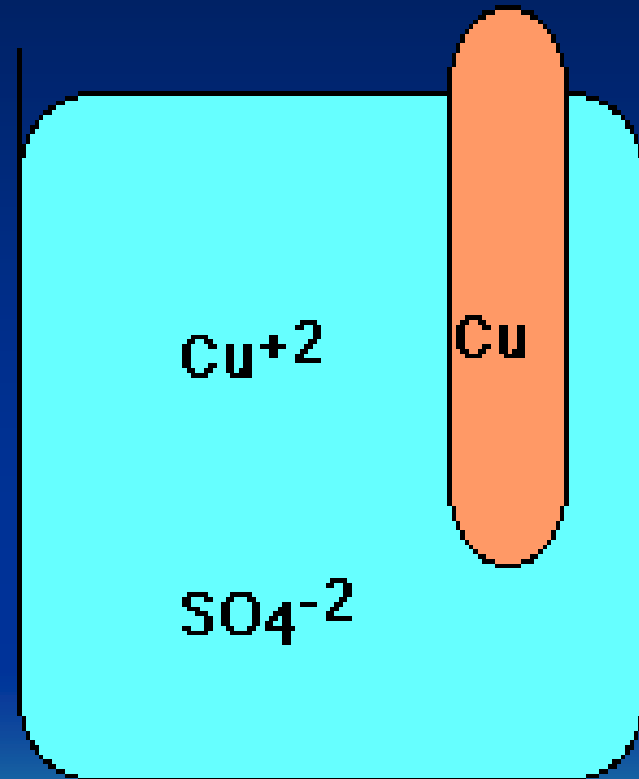


### Oxidation half-cell



Zinc electrode in zinc sulfate solution

### Reduction half-cell



Copper electrode in copper sulfate solution

- The half-cell, called the **anode**, is the site at which the **oxidation** of zinc occurs as shown below.
- $\text{Zn (s)} \longrightarrow \text{Zn}^{+2} \text{ (aq)} + 2\text{e}^{-}$
- During the oxidation of zinc, the zinc electrode will slowly dissolve to produce zinc ions ( $\text{Zn}^{+2}$ ), which enter into the solution containing  $\text{Zn}^{+2} \text{ (aq)}$  and  $\text{SO}_4^{-2} \text{ (aq)}$  ions.



- The half-cell, called the **cathode**, is the site at which **reduction** of copper occurs as shown below.
- $\text{Cu}^{+2} (\text{aq}) + 2\text{e}^{-} \longrightarrow \text{Cu} (\text{s})$
- When the reduction of copper ions ( $\text{Cu}^{+2}$ ) occurs, copper atoms accumulate on the surface of the solid copper electrode.

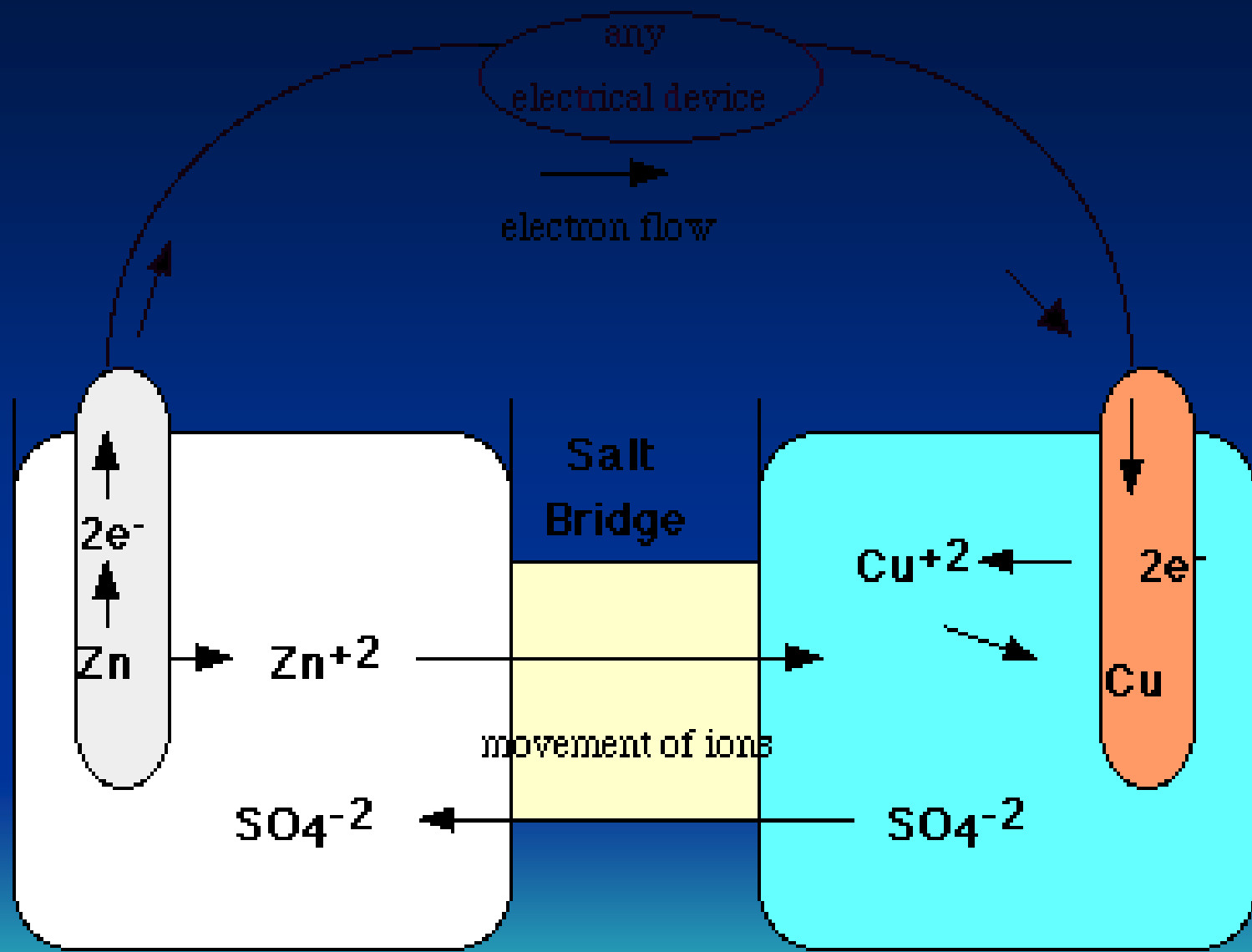


# Salt Bridge

- The reaction in each half-cell does **not** occur unless the two half cells are connected to each other.
- It is an inverted U-tube containing an electrolyte e.g  $\text{KCl}$ ,  $\text{KNO}_3$  etc it act as bridge by connecting two half cells which helps in:
  - To completing the electric circuit.
  - To prevent mixing of solutions of two half cell.
  - To help maintain electric neutrality







Zinc electrode in zinc sulfate solution

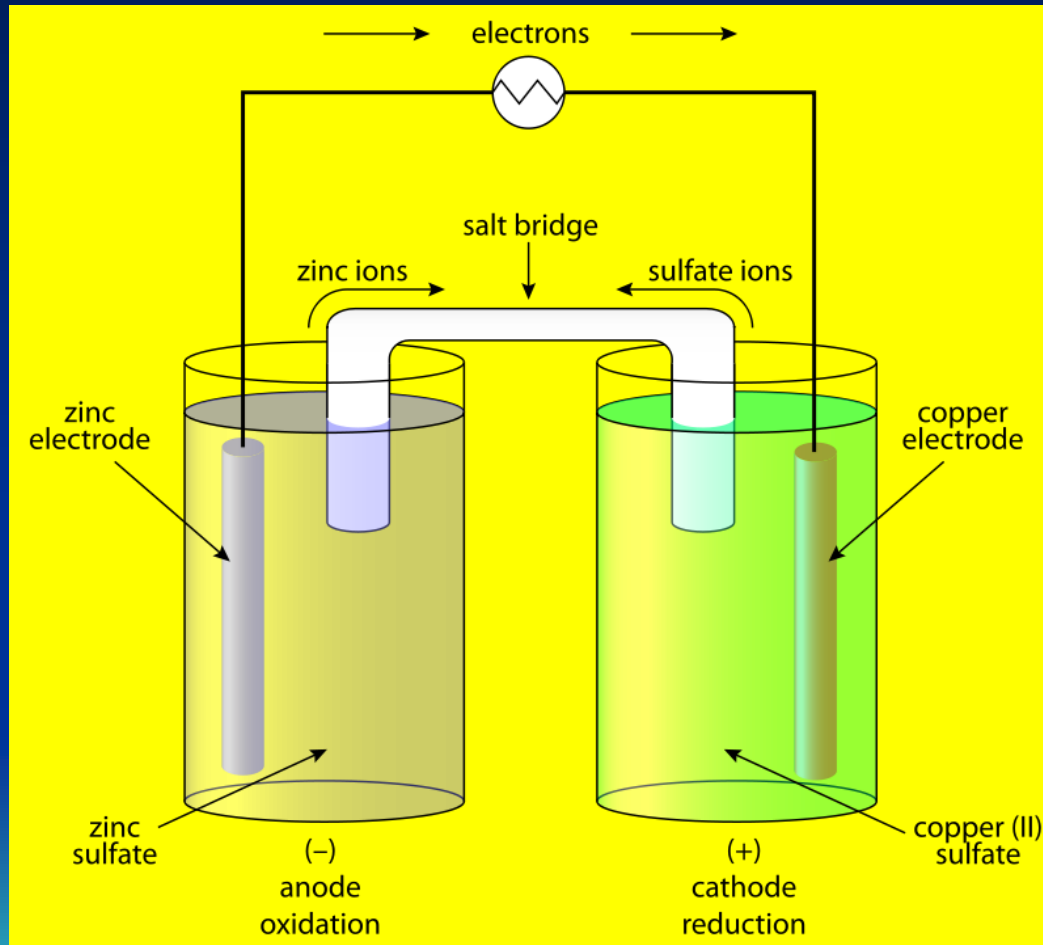
Copper electrode in copper sulfate solution

# Electrolytic Cell

- It is a device to convert Electric Energy into Chemical Energy
- An electrolytic cell is an electrochemical cell in which the energy from an applied voltage is used to drive an otherwise nonspontaneous reaction. Such a cell could be produced by applying a reverse voltage to a voltaic cell like the Daniell cell.



## Reaction in Daniell Cell



Electrolytic Cell use  
full in Electroplating



# Difference between Electrochemical cell and Eletrolytic Cell

## Electrochemical cell

Chemical energy  
converted into  
Electric energy

It is based on redox  
which is spontaneous

## Eletrolytic Cell

- Electric energy converted into Chemical energy
- It is based on redox which is non-spontaneous



# Electrode Potential

- It is the tendency of an electrode in half cell to lose or gain electrons when it is in contact with solution of its own ions
- $\text{Zn (s)} \longrightarrow \text{Zn}^{2+} \text{ (aq)} + 2\text{e}^-$       Oxidation half
- $\text{Cu}^{2+} \text{ (aq)} + 2\text{e}^- \longrightarrow \text{Cu (s)}$       Reduction half

# Electromotive force (emf)

- The two half-cells are also connected externally. In this arrangement, electrons provided by the oxidation reaction are forced to travel via an external circuit to the site of the reduction reaction. The fact that the reaction occurs spontaneously once these half cells are connected indicates that there is a difference in potential energy. This difference in potential energy is called an **electromotive force (emf)** and is measured in terms of **volts**. The zinc/copper cell has an **emf** of about 1.1 volts under standard conditions.

# Nernst Equation

- Electrochemistry deals with cell potential as well as energy of chemical reactions. The energy of a chemical system drives the charges to move, and the driving force give rise to the cell potential of a system called galvanic cell. The energy aspect is also related to the chemical equilibrium. All these relationships are tied together in the concept of **Nernst equation**.





- Walther H. Nernst (1864-1941) received the Nobel prize in 1920 "**in recognition of his work in thermochemistry**". His contribution to chemical thermodynamics led to the well known equation correlating chemical energy and the electric potential of a galvanic cell or battery.



# Nernst Equation

$$E = E^\circ + \frac{RT}{nF} \ln \frac{[Ox.]}{[Red.]}$$

or, using base-10 logarithms,

$$E = E^\circ + \frac{2.303 RT}{nF} \log \frac{[Ox.]}{[Red.]}$$

$E$  = The potential of electrode

$E^\circ$  = The standard potential of electrode

$R$  = gas constant **8.314 J/mol K**

$T$  = temperature in Kelvins **298 K**

$F$  = Faraday's constant **96,485 J/V-mol**

$n$  = number of electrons transferred

$[Ox.]$  = Molar concentration of Oxidized form

$[Red.]$  = Molar concentration of reduced form



# Nernst Equation

At room temperature (298 K), and

$$R = 8.314 \text{ J/mol K}$$

$$F = 96,485 \text{ J/V-mol}$$

$$\frac{2.303 RT}{F} = 0.0592 \text{ V}$$

The final form of the Nernst Equation becomes

$$E = E^\circ + \frac{0.0592}{n} \log \frac{[Ox.]}{[Red.]}$$

$$E_{\text{Cell}} = E_{\text{cathode}} - E_{\text{Anode}}$$

# illustrative Example

Calculate the electrode potential at a copper electrode dipped in a 0.1M solution of copper sulphate at 25°C . The standard potential of Cu<sup>2+</sup>/Cu system is 0.34 volt at 298 K.

**Solution:**  $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

We know that  $E_{\text{Cu}^{2+}/\text{Cu}} = E_{\text{Cu}^{2+}/\text{Cu}}^0 + \frac{0.0591}{n} \log_{10} [\text{Cu}^{2+}]$

Putting the values of  $E_{\text{red}}^0 = 0.34 \text{ V}$ ,  $n = 2$  and  $[\text{Cu}^{2+}] = 0.1 \text{ M}$

$$E_{\text{red}} = 0.34 + \frac{0.0591}{2} \log_{10}[0.1]$$

$$= 0.34 + 0.02955 \times (-1)$$

$$= 0.31045 \text{ volt}$$