

# Reduction-Oxidation Reactions

## (REDOX):

Oxidation- Process in which oxidation state of an element increases. Species loses electrons.

Reduction- Process in which oxidation state of an element decreases. Species gains electrons.

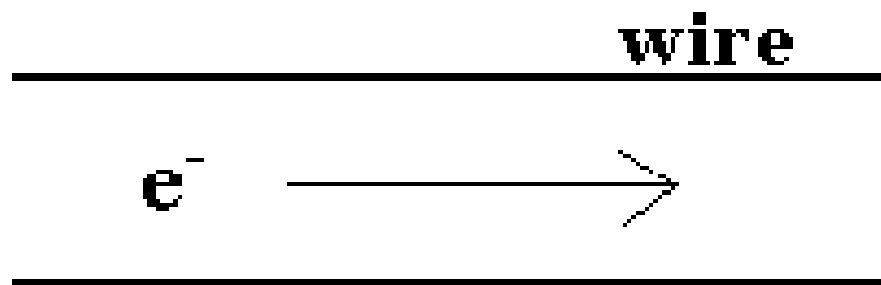
# REDOX cont...:



**Zn(s): oxidized/reducing agent.**

**Cu<sup>2+</sup>(aq): reduced/oxidizing agent.**

# Electrochemical Cells:



**I: current(flow)**

**V: voltage(pressure)**

**Consider,**



# Electrochemical Cells:

Electrode- Strip of metal.

Half cell-Strip of metal in contact with its ion.

Salt bridge- Allows passage of charge but not reactants.

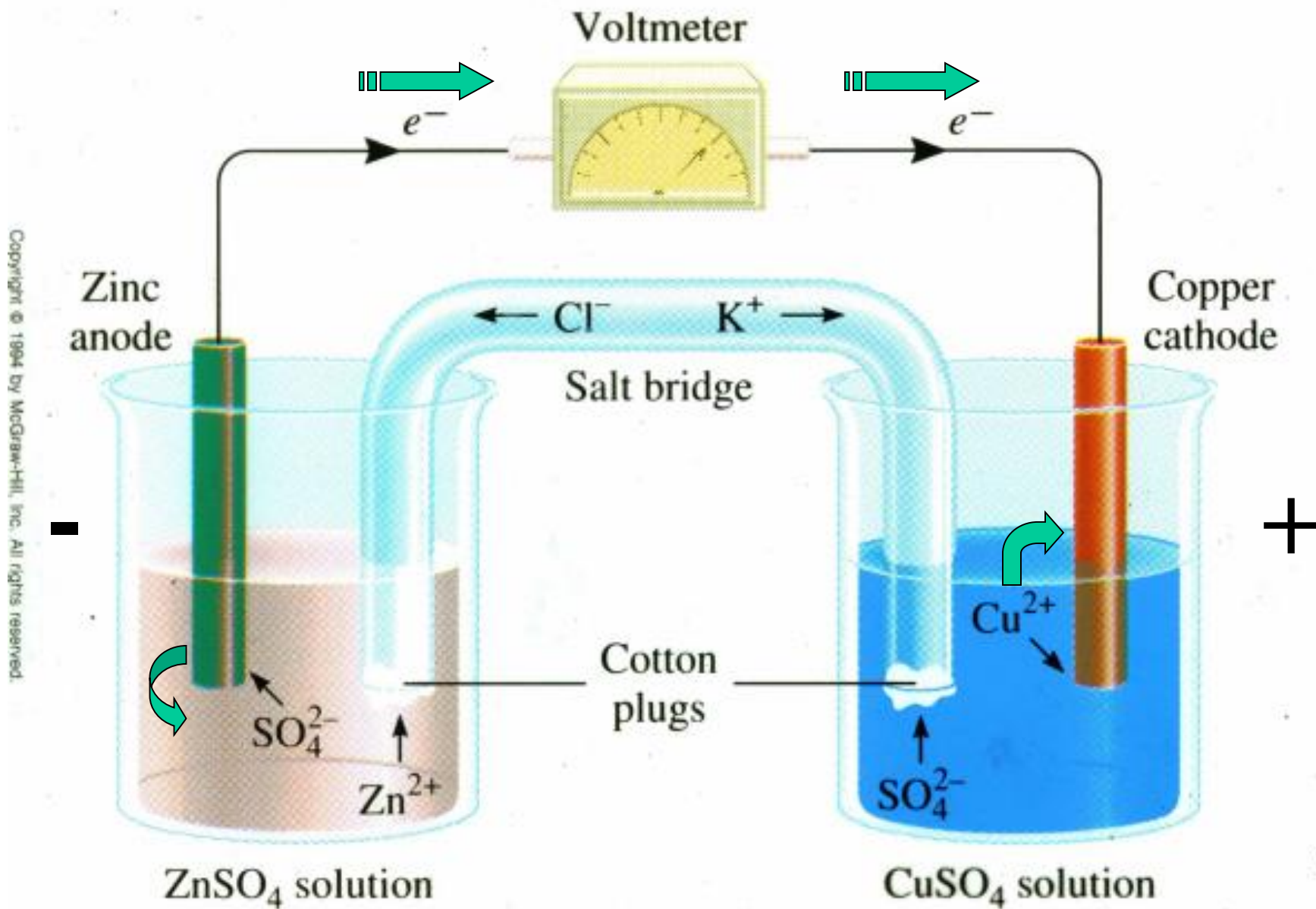
**Anode:**



**Cathode:**



# Electrochemical Cells cont...:



# Describing Electrochemical Cells

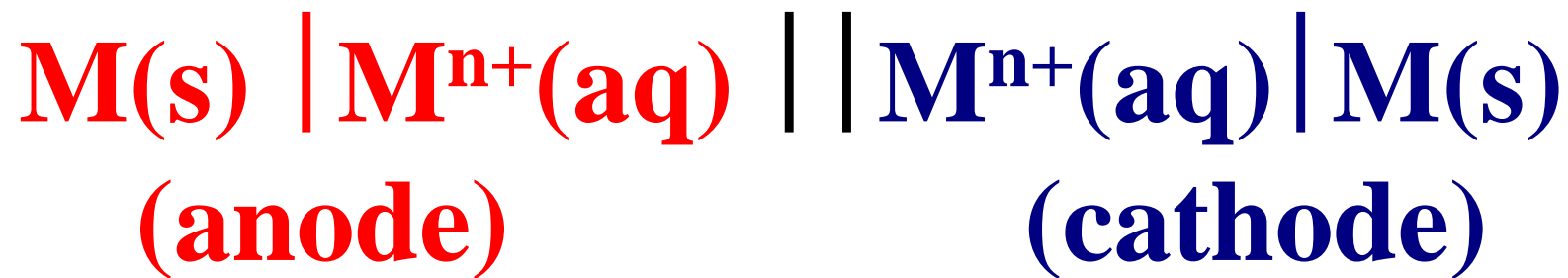
## (Cell Diagram):

**ANODE(OXIDATION) on left.**

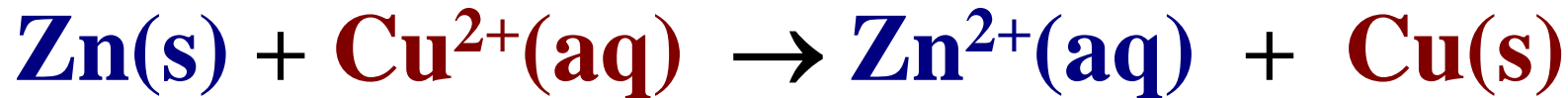
**CATHODE(REDUCTION) on right.**

| indicates phase change.

|| indicates salt bridge.



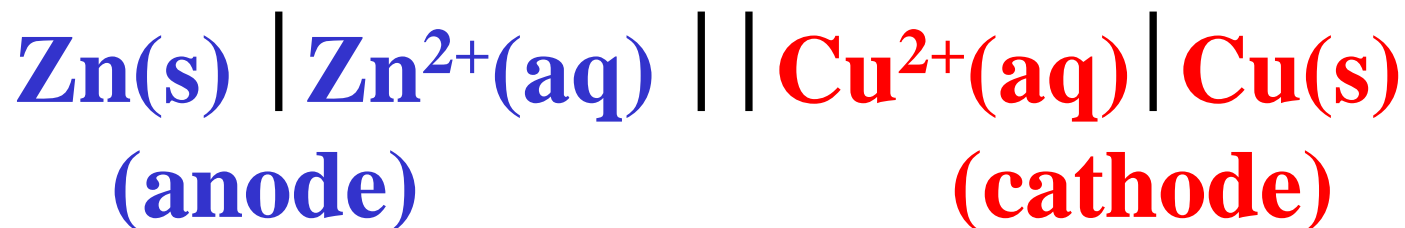
## Cell Diagram cont..:



**Anode:**



**Cathode:**



# Electrode Potentials:

The voltage recorded by an electrochemical cell is referred to as the **electromotive force(emf)** and given the following symbol.

$$E_{cell}$$



# Standard Electrode Potentials:

Reaction occurs under standard conditions:  
25°C and all substances are at unit  
concentration

(1 M for all ions and 1 atm for all gases).

Measured potential given the following  
symbol.

$$E_{cell}^{\circ}$$

The cell potential can be broken up into two  
components or half cells.

$$E_{cell}^{\circ} = E_{oxidation}^{\circ} + E_{reduction}^{\circ}$$

# Electrochemical Cells cont...:

$$E_{\text{cell}} = +1.10 \text{ V}$$

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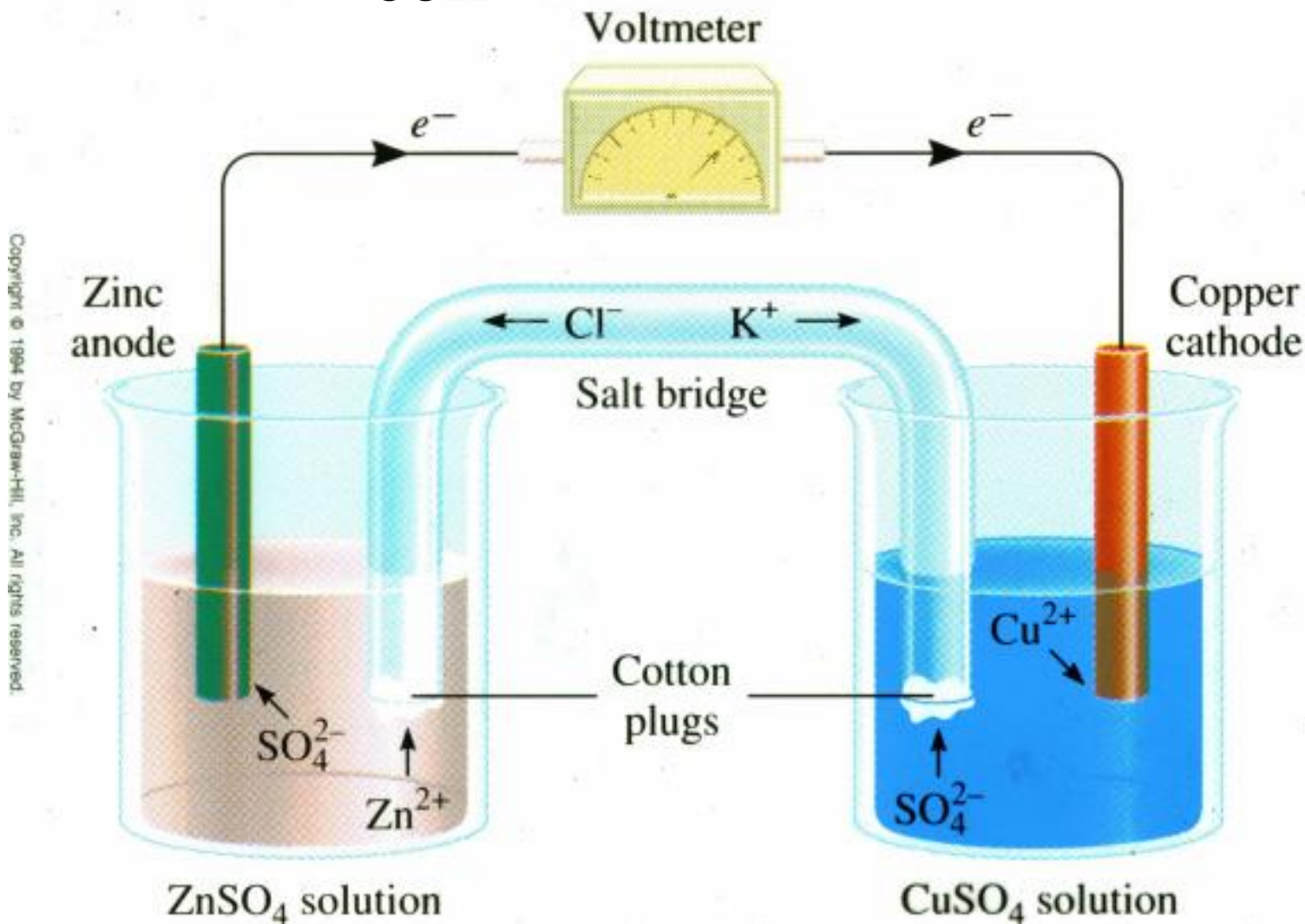
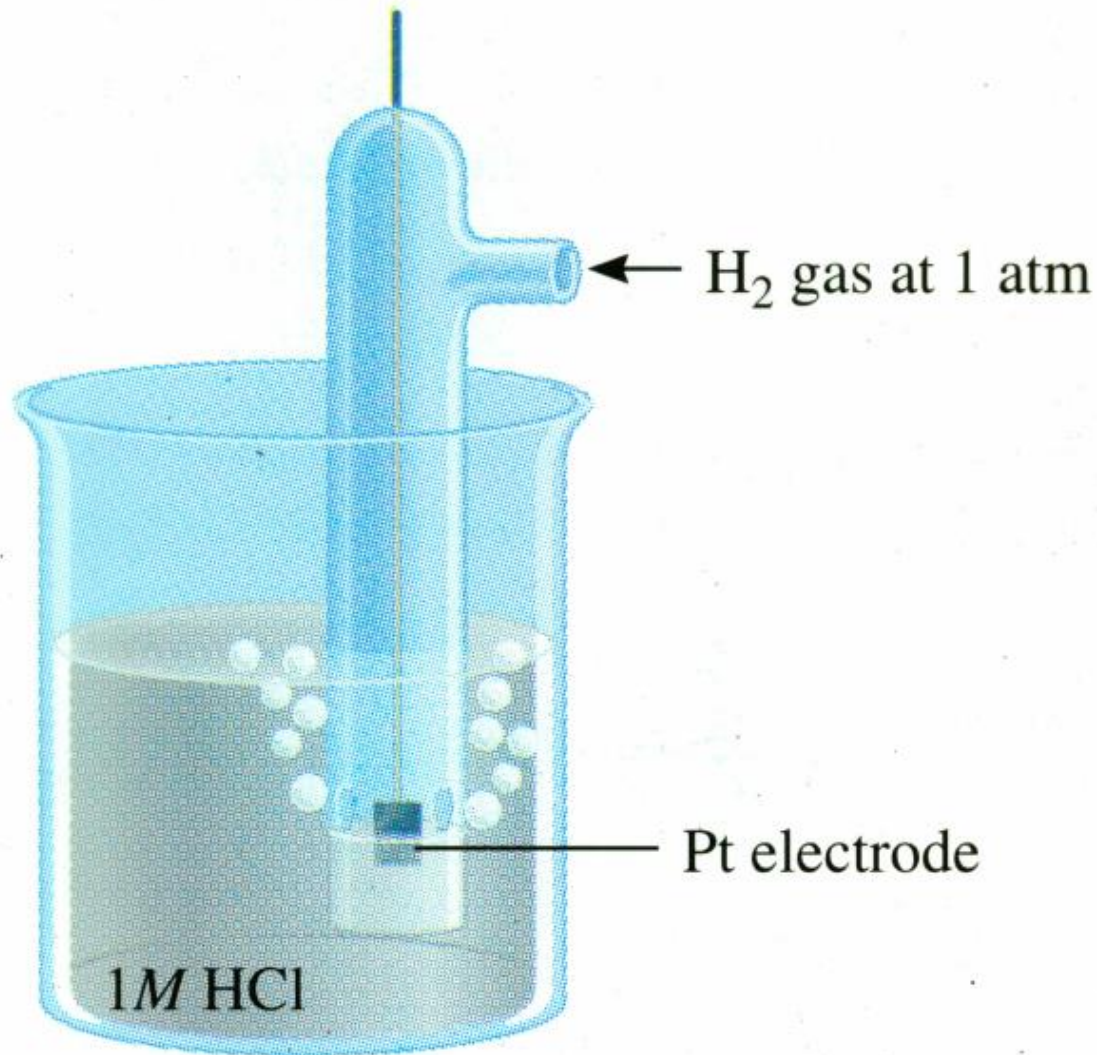
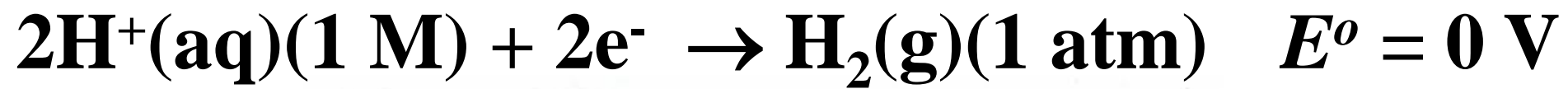
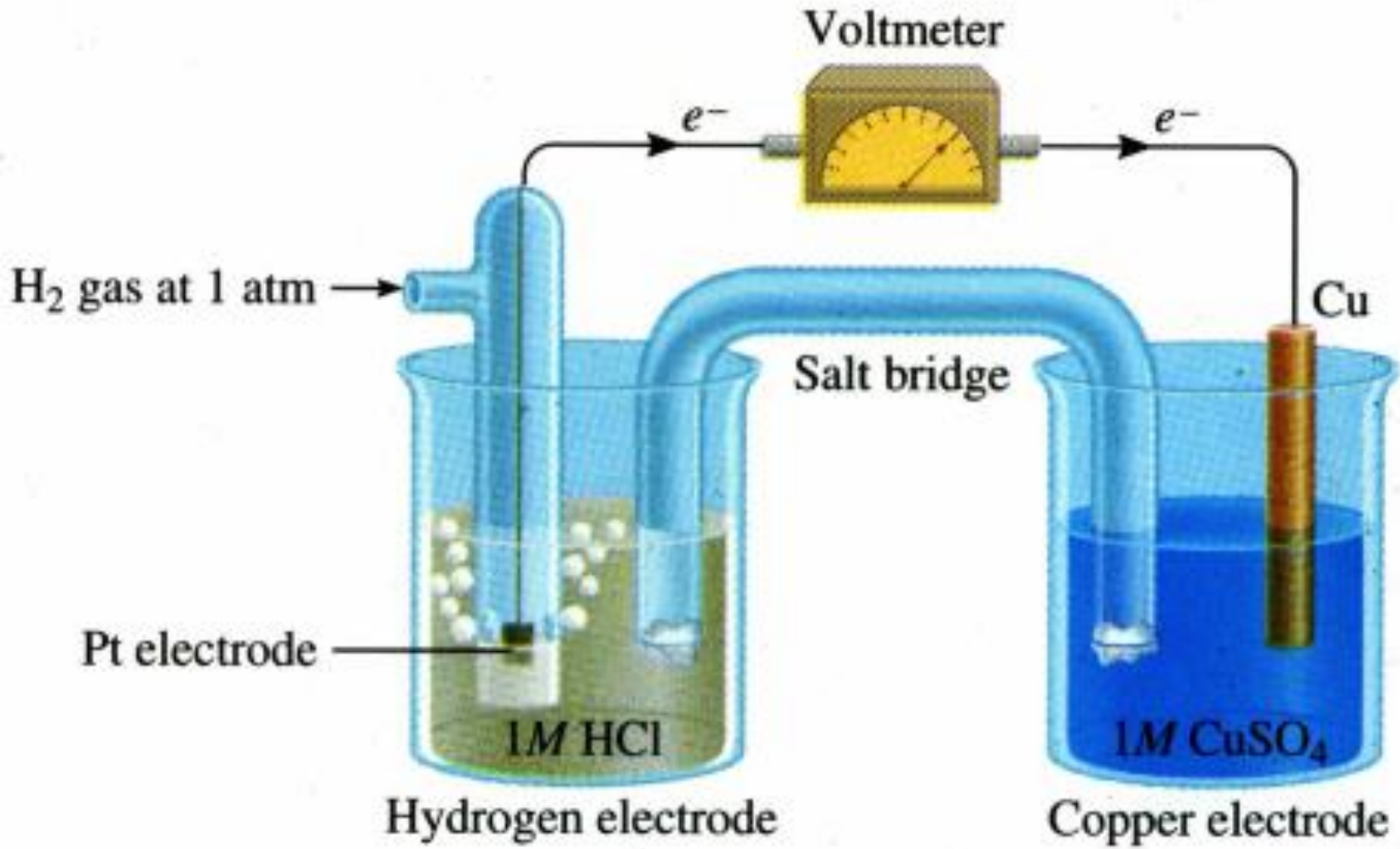


Figure 20.1

# Standard Hydrogen Electrode(SHE):

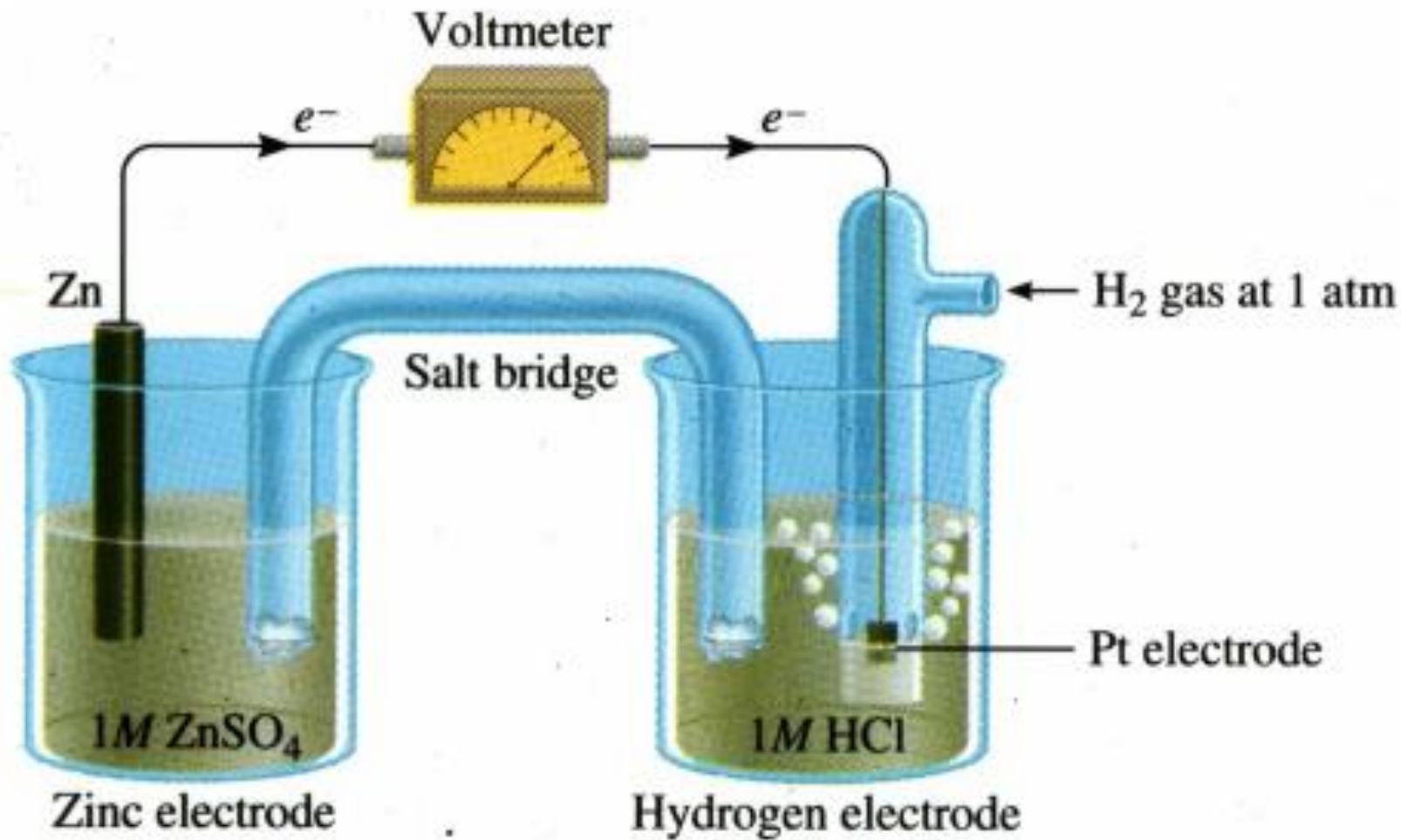


# SHE/Cu



$$E_{cell}^{\circ} = +0.337 \text{ V}$$

# Zn/SHE



$$E_{cell}^{\circ} = +0.763 \text{ V}$$

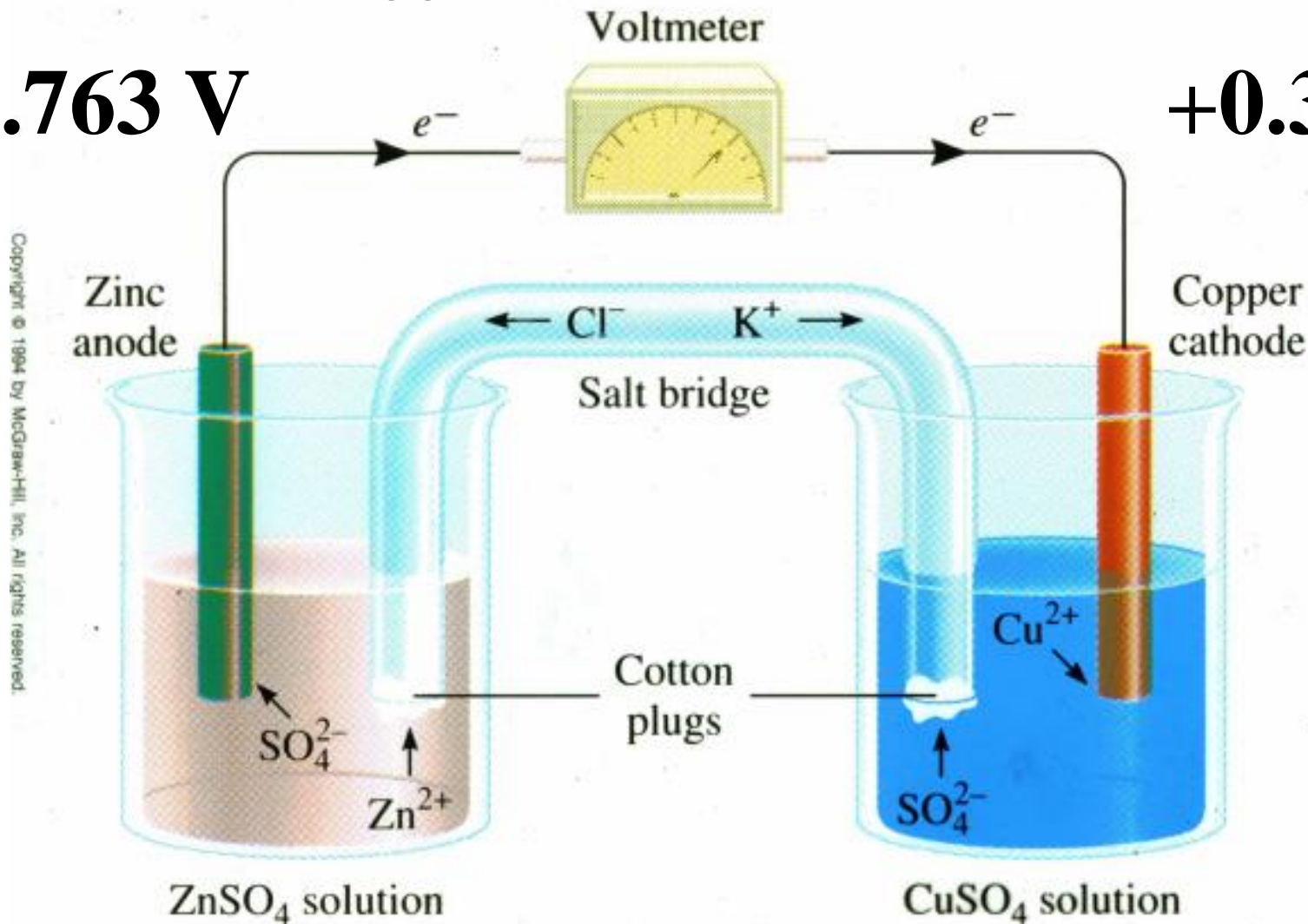


# Electrochemical Cells cont...:

$$E_{\text{cell}} = +1.10 \text{ V}$$

**+0.763 V**

**+0.337 V**



# Predicting Spontaneous Redox

## Reactions:

$$W_{\text{electrical}} = nF E_{\text{cell}} \quad \Delta G = -nFE_{\text{cell}}$$

$W_{\text{electrical}}$  = electrical work

$n$  = # moles of electrons transferred

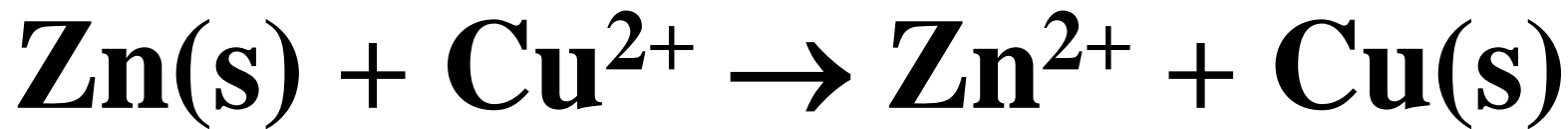
$F$  = Faraday constant (96485 C/mole)

$E_{\text{cell}}$  = Voltage of cell. NOTE: 1 J = 1 C·V

$\Delta G$  Gibbs Free Energy

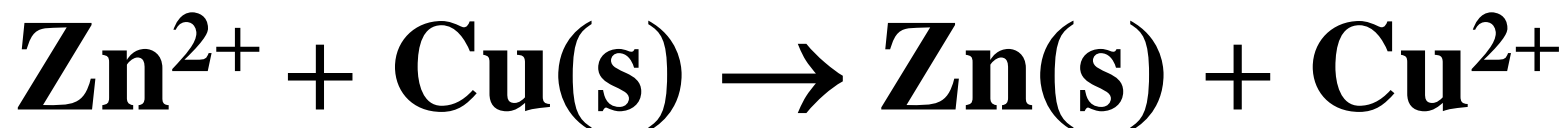
$\Delta G_{\text{cell}} < 0$  + # processes spontaneous

$\Delta G_{\text{cell}} > 0$  - # processes not spontaneous



$$E_{cell}^{\circ} = +1.10 \text{ V} \quad \text{spontaneous}$$

**Reverse reaction**



$$E_{cell}^{\circ} = -1.10 \text{ V} \quad \text{not spontaneous}$$

**Consider:**



$E_{\text{red}}^{\circ}(\text{V})$

**+0.337 V**

**-0.763 V**

**Cu<sup>2+</sup> more likely to be reduced.**



# Cell Potential as a Function of

## Concentration:

### Nernst Equation:

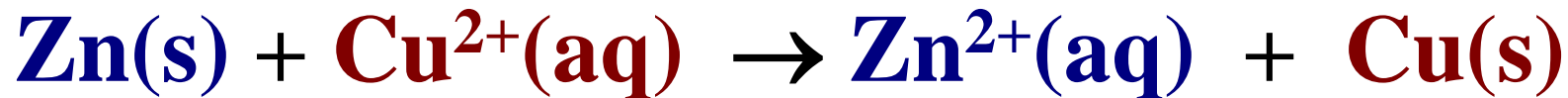
$$E_{cell} = E_{cell}^{\circ} + \frac{-0.0592V}{n} \text{Log}Q$$

Where if  $aA + bB \rightarrow cC + dD$

$$Q = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

# Nernst Equation cont:

Consider

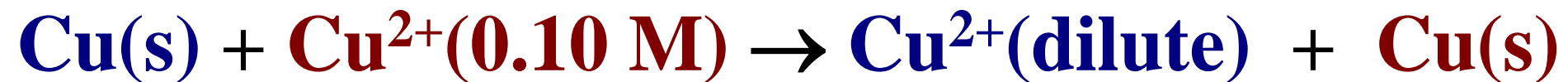


$$E_{\text{cell}} = E_{\text{cell}}^{\circ} + \frac{-0.0592 \text{ V}}{n} \text{Log} \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

$$E_{\text{cell}} = +1.100\text{V} + \frac{-0.0592 \text{ V}}{2} \text{Log} \frac{0.50 \text{ M}}{0.50 \text{ M}}$$

$$E_{\text{cell}} = +1.100\text{V}$$

# Concentration Cell:



$$E_{\text{cell}} = 0\text{V} + \frac{-0.0592\ \text{V}}{2} \text{Log} \frac{[\text{Cu}^{2+}](\mathbf{\text{dilute}})}{\mathbf{0.10\ M}}$$

Use  $E_{\text{cell}}$  to determine  $[\text{Cu}^{2+}]$  at equilibrium and thus  $K_f$  for  $\text{Cu-NH}_3$  complex.

# Electrolysis:

**Electricity is used to cause a non-spontaneous redox reaction to occur.**

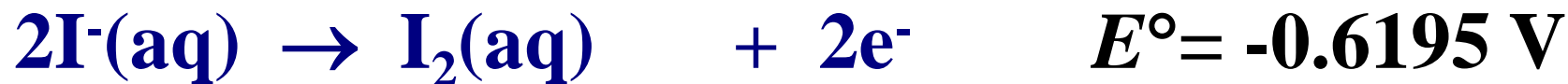
**Ex: Electrolysis of molten sodium chloride.**



**$E_{cell}^{\circ} = -4.07 \text{ V}$  NOT SPONTANEOUS!!!**

# Electrolysis of Aqueous KI:

**Anode:**



**Cathode:**



$E_{\text{cell}}^{\circ} = -1.448 \text{ V}$  not spontaneous