

Reduction potential = +0.34V  
 $E^0 = -0.34V$  ← flip sign for oxidation  
 $E^0 = -0.76V$

$E^0_{cell} = -1.10V$

Since the cell voltage is negative, the redox rxn will not occur.

Since this cell voltage is positive, the rxn will occur.

Standard state - 25°C, 1 bar pressure, 1M concentration

$\Delta G^0 = -RT \ln K$

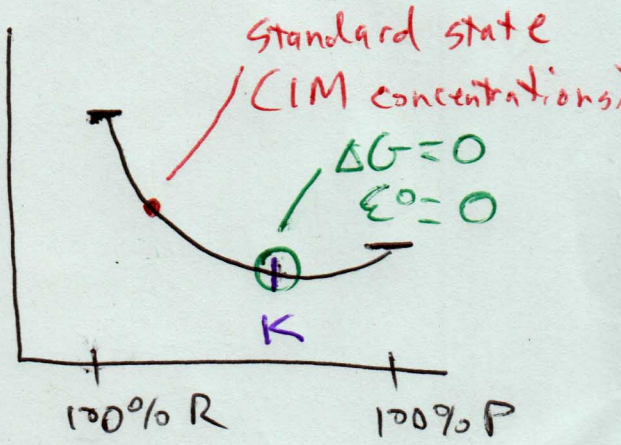
$\Delta G = \Delta G^0 + RT \ln Q$

↑ true energy available

↑ energy available at the standard state

↑ actual starting concentrations

$\Delta G$   
(E)



@ equilibrium

$0 = \Delta G^0 + RT \ln K$  (Q=K at equilibrium)

$-RT \ln K = \Delta G^0$

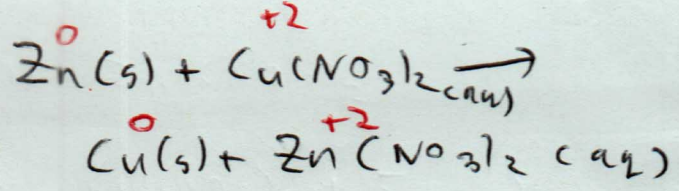
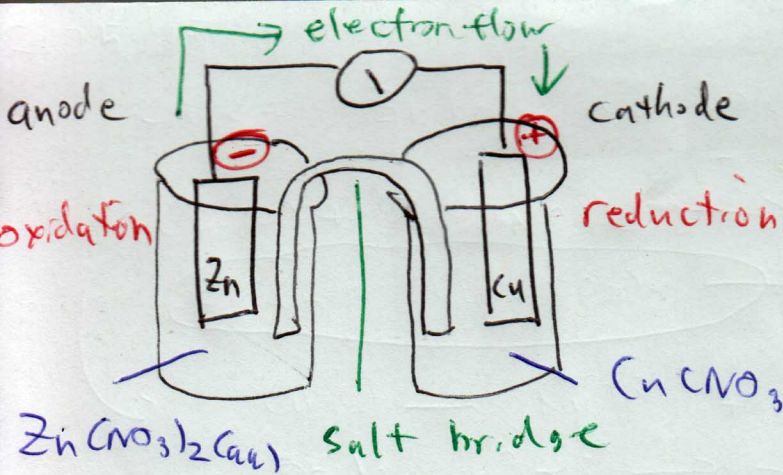
$\Delta G = -nFE^0$

$-nFE = -nFE^0 + RT \ln Q$

$E = E^0 - \frac{RT}{nF} \ln Q$

$\Delta G = \Delta G^0 + RT \ln Q$   
 $Q^0 = 1$   
 $\Delta G = \Delta G^0 + RT \ln 1$   
 $\Delta G = \Delta G^0 + 0$   
 $\Delta G = \Delta G^0$





$$Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

$$Q^{\circ} = \frac{1}{1} = 1$$



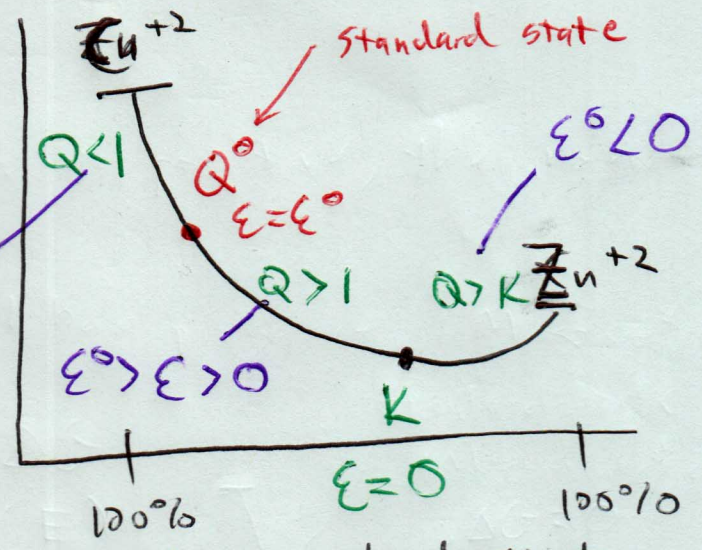
$$[\text{Zn}^{2+}] = 1.0 \times 10^{-4} \text{ M}$$

$$[\text{Cu}^{2+}] = 2.0 \text{ M}$$

$$Q = \frac{1.0 \times 10^{-4}}{2.0} = 5.0 \times 10^{-5}$$

$$Q \ll 1$$

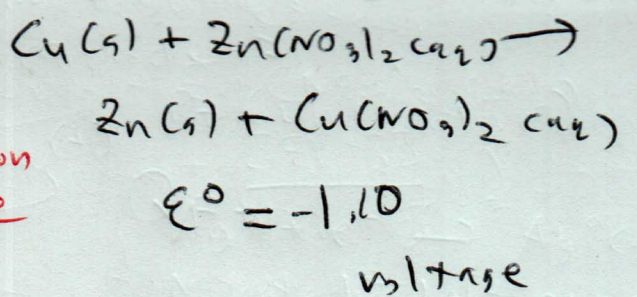
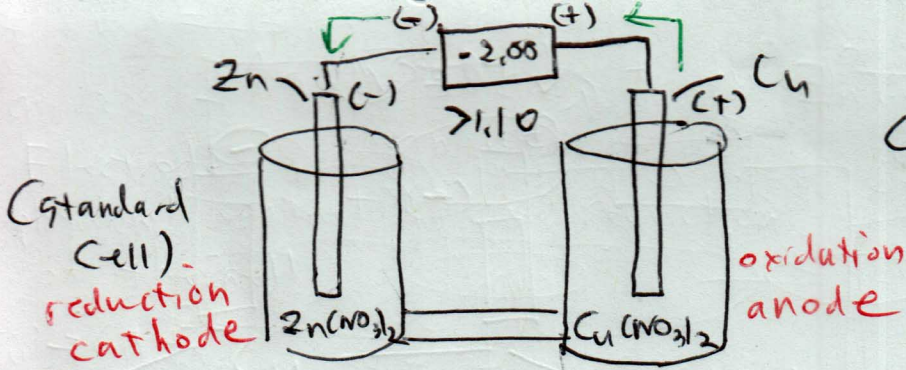
Excess of reactants  $\Delta G$



If a reaction begins with  $E > E^{\circ}$  a greater concentration of reactants than products (compared to the standard state), the rate of forward run will be greater than the standard state, and the rate of the reverse run will be small. This means the run will want to form more products, which means the run begins at higher energy than the standard state, which for an electrochemical cell means it will have higher voltage.

$$\begin{aligned} @25^{\circ}\text{C} \quad E &= E^{\circ} - \frac{0.0592}{n} \ln Q \\ &= E^{\circ} - \frac{0.0592}{2} \ln (5.0 \times 10^{-5}) = 1.23 \text{ V} \end{aligned}$$





Because when an external source with a greater than the cell potential is applied against the normal electron flow for the cell, the cell can be forced into the reverse rxn  $\rightarrow$  electrolytic cell