Chemistry 146 Lecture Problems
Zinc/Iron Electrochemical Cell
At Non-Standard Conditions

For the electrochemical system:

$Zn \mid Zn^{2+}$ (1 M) || Fe^{2+} (1 M) | Fe

Anode (oxidation occurs at the anode, in the cell notation used above the anode is on the left):

Oxidation Reaction $B_{red} \rightarrow B_{ox} + ne^-$

$Zn(s) \rightarrow Zn^{2+}$ (1 M) + $2e^-$

Reduction potential for B_{ox} $E_{anode} := -0.763 \cdot volt$

Cathode (reduction occurs at the cathode, in the cell notation used above the cathode is on the right):

Reduction Reaction $A_{ox} + ne^- \rightarrow A_{red}$

Fe^{2+} (1 M) + $2e^-$ --> Fe (s)

Reduction potential for A_{ox} $E_{cathode} := -0.409 \cdot volt$

Reaction: The overall electrochemical reaction has the form:

$a\,A_{ox} + b\,B_{red} \leftrightarrow a\,A_{red} + b\,B_{ox}$

For this problem

$Zn(s) + Fe^{2+}$ (1 M) --> Zn^{2+} (1 M) + Fe (s)

E^o_{cell} (Calculate the standard cell potential for this system):

$$E_{std\_cell} := E_{cathode} - E_{anode}$$

$$E_{std\_cell} = 0.354 \cdot volt$$

E_{cell} calculation with the Nernst Equation (This step is required any time the system is not at the "standard state". In the "standard state" concentrations are 1 M, pressure is 1 atm, and temperature is 298.15 K) In this problem you were given a series of different concentrations to use for these calculations:

Temperature $T := (273.15 + 20) \cdot K$

Faraday's Constant: $F := 96485.309 \cdot \frac{coul}{mole}$

Gas Law Constant: $R := 8.314510 \cdot \frac{joule}{mole \cdot K}$

Number of electrons in balanced redox reaction: $n := 2$

Define Molarity $M := \frac{mole}{liter}$

For each example you are given a concentration for Zn^{2+} and Fe^{2+}. The following additional calculations are required to solve:

Concentration anode: $C_{Zn2} := 1 \cdot M$

Concentration cathode: $C_{Fe2} := 1 \cdot M$

Equlibrium Quotient (Q) $Q := \frac{C_{Zn2}}{C_{Fe2}}$

Nernst Equation: $E_{cell} := E_{std\_cell} - \left(\frac{R \cdot T}{n \cdot F}\right) \cdot \ln(Q)$

The Cell potential: $E_{cell} = 0.354 \cdot volt$

Now Repeat for other concentrations

$C_{Zn2} := 1 \cdot M$ $\quad C_{Fe2} := 2 \cdot M$ $\quad Q := \frac{C_{Zn2}}{C_{Fe2}}$

$E_{cell} := E_{std\_cell} - \left(\frac{R \cdot T}{n \cdot F}\right) \cdot \ln(Q)$ $\quad E_{cell} = 0.363 \cdot volt$

Now Repeat for other concentrations

$C_{Zn2} := 2 \cdot M$ $\quad C_{Fe2} := 1 \cdot M$ $\quad Q := \frac{C_{Zn2}}{C_{Fe2}}$

$E_{cell} := E_{std\_cell} - \left(\frac{R \cdot T}{n \cdot F}\right) \cdot \ln(Q)$ $\quad E_{cell} = 0.345 \cdot volt$

Now Repeat for other concentrations

$C_{Zn2} := 0.1 \cdot M$ $\quad C_{Fe2} := 0.1 \cdot M$ $\quad Q := \frac{C_{Zn2}}{C_{Fe2}}$

$E_{cell} := E_{std\_cell} - \left(\frac{R \cdot T}{n \cdot F}\right) \cdot \ln(Q)$ $\quad E_{cell} = 0.354 \cdot volt$

Now Repeat for other concentrations

$C_{Zn2} := 1.0 \cdot M$ $\quad C_{Fe2} := 0.01 \cdot M$ $\quad Q := \frac{C_{Zn2}}{C_{Fe2}}$

$E_{cell} := E_{std\_cell} - \left(\frac{R \cdot T}{n \cdot F}\right) \cdot \ln(Q)$ $\quad E_{cell} = 0.296 \cdot volt$

Now Repeat for other concentrations

$C_{Zn2} := 0.01 \cdot M$ $\quad C_{Fe2} := 1 \cdot M$ $\quad Q := \frac{C_{Zn2}}{C_{Fe2}}$

$E_{cell} := E_{std\_cell} - \left(\frac{R \cdot T}{n \cdot F}\right) \cdot \ln(Q)$ $\quad E_{cell} = 0.412 \cdot volt$

Calculating the equlibrium concentrations for a set potential

For the electrochemical system:

Zn | Zn^{2+} (1 M) || Fe^{2+} (1 M) | Fe

Standard Potentials:

$E_{anode} := -0.7628 \cdot volt$

$E_{cathode} := -0.409 \cdot volt$

$n := 2$

Standard Cell Potential:

$E_{std\_cell} := E_{cathode} - E_{anode}$

$E_{std\_cell} = 0.354 \cdot volt$

Applied Potential $E_{applied} := 0.3538 \cdot volt$

Since the applied potential controls the cell equlibrium:

$E_{cell} := E_{applied}$

Rearanges the Nernst Equation to find Q:

NOTE: $Q = \frac{products}{reactants}$ $Q = \frac{Zn^{2+}}{Fe^{2+}}$

$$E_{cell} = E_{std\_cell} - \left(\frac{R \cdot T}{n \cdot F}\right) \cdot \ln(Q)$$

$$Q := \exp\left[\frac{-\left(E_{cell} \cdot n \cdot F - E_{std\_cell} \cdot n \cdot F\right)}{(R \cdot T)}\right]$$

$Q = 1$

Repeat calculation for other applied potentials

$E_{cell} := 0.3538 \cdot volt$ $\quad Q := \exp\left[\frac{-\left(E_{cell} \cdot n \cdot F - E_{std\_cell} \cdot n \cdot F\right)}{(R \cdot T)}\right]$ $\quad Q = 1$

$E_{cell} := 0.3530 \cdot volt$ $\quad Q := \exp\left[\frac{-\left(E_{cell} \cdot n \cdot F - E_{std\_cell} \cdot n \cdot F\right)}{(R \cdot T)}\right]$ $\quad Q = 1.065$

$E_{cell} := 0.3400 \cdot volt$ $\quad Q := \exp\left[\frac{-\left(E_{cell} \cdot n \cdot F - E_{std\_cell} \cdot n \cdot F\right)}{(R \cdot T)}\right]$ $\quad Q = 2.982$

$E_{cell} := 0.3550 \cdot volt$ $\quad Q := \exp\left[\frac{-\left(E_{cell} \cdot n \cdot F - E_{std\_cell} \cdot n \cdot F\right)}{(R \cdot T)}\right]$ $\quad Q = 0.909$

$E_{cell} := 0.3600 \cdot volt$ $\quad Q := \exp\left[\frac{-\left(E_{cell} \cdot n \cdot F - E_{std\_cell} \cdot n \cdot F\right)}{(R \cdot T)}\right]$ $\quad Q = 0.612$