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

For the electrochemical system:
\[
\text{Zn} | \text{Zn}^{2+} \ (1 \text{ M}) \ || \text{Fe}^{2+} \ (1 \text{ M}) | \text{Fe}
\]

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

\[
\text{Oxidation Reaction} \quad \text{B}_{\text{red}} \rightarrow \text{B}_{\text{ox}} + \text{ne}^{-}
\]
\[
\text{Zn}(s) \rightarrow \text{Zn}^{2+} \ (1 \text{ M}) + 2 \text{e}^{-}
\]

Reduction potential for \( \text{B}_{\text{ox}} \)
\[ E_{\text{anode}} = -0.763 \text{ volt} \]

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

\[
\text{Reduction Reaction} \quad \text{A}_{\text{ox}} + \text{ne}^{-} \rightarrow \text{A}_{\text{red}}
\]
\[
\text{Fe}^{2+} \ (1 \text{ M}) + 2\text{e}^{-} \rightarrow \text{Fe} \ (s)
\]

Reduction potential for \( \text{A}_{\text{ox}} \)
\[ E_{\text{cathode}} = -0.409 \text{ volt} \]

Reaction: The overall electrochemical reaction has the form:
\[
a \text{A}_{\text{ox}} + b \text{B}_{\text{red}} \leftrightarrow a \text{A}_{\text{red}} + b \text{B}_{\text{ox}}
\]

For this problem
\[
\text{Zn}(s) + \text{Fe}^{2+} \ (1 \text{ M}) \rightarrow \text{Zn}^{2+} \ (1 \text{ M}) + \text{Fe} \ (s)
\]

\( E_{\text{cell}} \) (Calculate the standard cell potential for this system):
\[
E_{\text{std\_cell}} := E_{\text{cathode}} - E_{\text{anode}}
\]
\[
E_{\text{std\_cell}} = 0.354 \text{ volt}
\]
The 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

Faraday's Constant:

Gas Law Constant:

Number of electrons in balanced redox reaction:

Define Molarity

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

Concentration anode: $C_{Zn^2} := 1 \cdot M$

Concentration cathode: $C_{Fe^2} := 1 \cdot M$

Equilibrium Quotient (Q)

Nernst Equation:

The Cell potential:

$E_{cell} = 0.354 \cdot \text{volt}$
Now Repeat for other concentrations

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

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

\[ E_{\text{cell}} = 0.363 \text{ volt} \]

Now Repeat for other concentrations

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

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

\[ E_{\text{cell}} = 0.345 \text{ volt} \]

Now Repeat for other concentrations

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

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

\[ E_{\text{cell}} = 0.354 \text{ volt} \]

Now Repeat for other concentrations

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

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

\[ E_{\text{cell}} = 0.296 \text{ volt} \]

Now Repeat for other concentrations

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

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

\[ E_{\text{cell}} = 0.412 \text{ volt} \]
Calculating the equilibrium concentrations for a set potential

For the electrochemical system:

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

Standard Potentials:

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

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

$n := 2$

Standard Cell Potential:

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

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

Applied Potential

$E_{\text{applied}} := 0.3538 \text{ volt}$

Since the applied potential controls the cell equilibrium:

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

Rearranges the Nernst Equation to find $Q$:

$Q = \frac{\text{products}}{\text{reactants}}$

$Q = \frac{Zn^{2+}}{Fe^{2+}}$

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

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

$Q = 1$
Repeat calculation for other applied potentials

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

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

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

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

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