Homework #19-4: The Effect of Concentration on EMF
Problems pg. 833 #29, 31, 33, 34, 95

19.29 What is the potential of a cell consisting of Zn/Zn\(^{2+}\) and Cu/Cu\(^{2+}\) half-cells at 25°C if the solutions have concentrations of [Zn\(^{2+}\)] = 0.25 \(M\) and [Cu\(^{2+}\)] = 0.15 \(M\)?

Overall Rxn:

\[E^\circ_{\text{cell}}?\]

\[E_{\text{cell}}?\]

9.31 What is the potential of a cell consisting of a Zn/Zn\(^{2+}\) half-cell and a H\(^+\)/H\(_2\) half-cell where [Zn\(^{2+}\)] = 0.45 \(M\), \(P_{H_2}\) = 2.0 atm, and [H\(^+\)] = 1.8 \(M\)?

Overall Reaction:

\[E^\circ_{\text{cell}}?\]

\[E_{\text{cell}}?\]

19.33 Complete the following steps to eventually calculate the [Cu\(^{2+}\)]/[Zn\(^{2+}\)] ratio at which the following reaction is spontaneous at 25°C: \(\text{Cu(s)} + \text{Zn}^{2+}(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + \text{Zn(s)}\)

a) Calculate \(E^\circ_{\text{cell}}\):

b) When at equilibrium, which concentration is greater—Zn\(^{2+}\) or Cu\(^{2+}\)? (Hint: Sign of \(E^\circ_{\text{cell}}\)?)

c) What is the [Cu\(^{2+}\)]/[Zn\(^{2+}\)] ratio at equilibrium. (Remember: set \(E_{\text{cell}} = 0\))

d) Thus, to be spontaneous, the [Cu\(^{2+}\)]/[Zn\(^{2+}\)] ratio must be (greater, less) than the ratio at equilibrium. Thus, is the reduction of zinc(II) by copper metal a practical use of copper?
19.34 Calculate the potential of the concentration cell depicted below:
(Mg$^{2+}$ concentrations in the solutions are 0.24M and 0.53M)

Red: \( \text{Mg}^{2+}(aq) + 2e^- \rightarrow \text{Mg(s)} \quad E_{\text{cathode}} = -2.37 \text{ V} \)

Ox: \( \text{Mg(s)} \rightarrow \text{Mg}^{2+}(aq) + 2e^- \quad E_{\text{anode}} = -2.37 \text{ V} \)

Overall Rxn:
(You must label concentrations of Mg$^{2+}$ to get overall rxn!)

\( E^{\circ}_{\text{cell}}? \)

\( E_{\text{cell}}? \)

19.95 A silver rod is placed into a saturated aqueous solution of silver oxalate, Ag$_2$C$_2$O$_4$, at 25$^\circ$C. The Ag/Ag$^+$ half-cell is connected to a standard hydrogen electrode (SHE) as shown in diagram to the right. The measured voltage is 0.589 V.

Complete the following steps to determine the solubility product constant for silver oxalate.

Overall rxn: \( 2\text{Ag}^+(aq) + \text{H}_2(g) \rightarrow 2\text{Ag(s)} + 2\text{H}^+(aq) \)

(a) \( ?\text{M} \) \( \text{ (1 atm) \quad (1M) } \)

b) Plug in the values for \( E \) and \( E^{\circ} \) into the Nernst equation to solve for the value of \( Q \):

\[ E = E^{\circ} - \frac{0.0257 \text{ V}}{n} \ln Q \]

c) Use the overall reaction to write the Q expression. Plug in all known values. Solve for [Ag$^+$].

d) Now, to find the solubility product of Ag$_2$C$_2$O$_4$:

\( \text{Ag}_2\text{C}_2\text{O}_4(s) \rightleftharpoons 2 \text{Ag}^+ + \text{C}_2\text{O}_4^{2-} \)

You know the [Ag$^+$], so what must be the [C$_2$O$_4^{2-}$]?

Write the \( K_{sp} \) expression. Plug in concentration values. Solve for the \( K_{sp} \) of Ag$_2$C$_2$O$_4$. 

Answers: 29) 1.09V 31) 0.78 V 33c) 6.7 $\times$ 10$^{-38}$ 34) 0.010 V 95b) 1.35 $\times$ 7.9 95c) 2.7 $\times$ 10$^{-4}$ M 95d) 9.8 $\times$ 10$^{-12}$