2. Answer parts (a) through (e) below, which relate to reactions involving silver ion, Ag⁺.

The reaction between silver ion and solid zinc is represented by the following equation.

\[ 2 \text{Ag}^+(aq) + \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2 \text{Ag}(s) \]

(a) A 1.50 g sample of Zn is combined with 250. mL of 0.110 M AgNO₃ at 25°C.

(i) Identify the limiting reactant. Show calculations to support your answer.

(ii) On the basis of the limiting reactant that you identified in part (i), determine the value of [Zn²⁺] after the reaction is complete. Assume that volume change is negligible.

(b) Determine the value of the standard potential, \( E^\circ \), for a galvanic cell based on the reaction between AgNO₃(aq) and solid Zn at 25°C.

Another galvanic cell is based on the reaction between Ag⁺(aq) and Cu(s), represented by the equation below. At 25°C, the standard potential, \( E^\circ \), for the cell is 0.46 V.

\[ 2 \text{Ag}^+(aq) + \text{Cu}(s) \rightarrow \text{Cu}^{2+}(aq) + 2 \text{Ag}(s) \]

(c) Determine the value of the standard free-energy change, \( \Delta G^\circ \), for the reaction between Ag⁺(aq) and Cu(s) at 25°C.

(d) The cell is constructed so that [Cu²⁺] is 0.045 M and [Ag⁺] is 0.010 M. Calculate the value of the potential, \( E \), for the cell.

(e) Under the conditions specified in part (d), is the reaction in the cell spontaneous? Justify your answer.

\[ \text{Moles of Zn} = \frac{1.50 \text{ g}}{65.4 \text{ g/mol}} = 0.0229 \text{ mol Zn} \]

\[ \text{Moles of Ag}^+ = 0.250 \text{ L} \times 0.110 \text{ M} = 0.275 \text{ mol Ag}^+ \]

The molar ratio of Ag⁺ to Zn is 2:1.

If we started with 0.0229 mol Zn, we would require

\[ 0.0229 \text{ mol Zn} \times \frac{2 \text{ mol Ag}^+}{1 \text{ mol Zn}} = 0.0458 \text{ mol Ag}^+ \]

But we require 0.0458 mol Ag⁺ and we only have 0.275 mol Ag⁺, Ag⁺ is the limiting reactant.
Since we have 250 mL of solution, and the volume change upon addition of 
Zn is negligible, \[ [Zn^{2+}] = \frac{M}{V} = \frac{0.138 \, \text{mols}}{0.250 \, \text{L}} = 0.552 \, \text{M} \]

The concentration of Zn\(^{2+}\) is 0.552 M.

\( a \) \ \ E^\circ = E^\circ_{\text{red}} (\text{cathode}) - E^\circ_{\text{red}} (\text{anode}) \)

Half reactions:

At cathode: \( \text{Ag}^+ + e^- \rightarrow \text{Ag} \) \[ E^\circ_{\text{red}} = +0.80 \, \text{V} \]

At anode: \( \text{Zn} - \rightarrow \text{Zn}^{2+}\) \( +2e^- \) \[ E^\circ_{\text{red}} = -0.76 \, \text{V} \]

\[ E^\circ = +0.80 \, \text{V} - (-0.76 \, \text{V}) = 1.56 \, \text{V} \]

\( c \) \ \ \[ \Delta G^\circ = -nFE^\circ \] \( \text{where} \ n = 2 \text{mols of electrons} \]

\[ \Delta G^\circ = -(2)(96,500)(0.46 \, \text{V}) = -92,400 \, \text{J} \]

\[ E^\circ = -0.46 \, \text{V} \]

\( d \) We use \[ E = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln Q \] but at 25°C, \[ E = E^\circ_{\text{cell}} - \frac{0.0592}{2} \log Q \]

In this case: \[ Q = \frac{[\text{Cu}^{2+}]}{[\text{Ag}^+]^2} = \frac{(0.045 \, \text{M})}{(0.010 \, \text{M})^2} = 4.5 \times 10^2 \]

\[ E = (0.46 \, \text{V}) - \frac{0.0592}{2} \log (4.5 \times 10^2) \]

\[ E = 0.38 \, \text{V} \]

\( e \) Yes, the reaction in the cell is spontaneous since \( E > 0 \). For a reaction to be 
spontaneous \( E \) must be greater than zero.

To check whether \( \Delta G \) is negative under these conditions to prove that the reaction is spontaneous:

\[ \Delta G = \Delta G^\circ + RT \ln Q \]

\[ \Delta G = -(8.9 \times 10^4 \, \text{J}) + (9.31 \, \text{mJ} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})(298 \, \text{K}) \ln (4.5 \times 10^2) \]

\[ \Rightarrow \Delta G < 0 \quad \text{the reaction is spontaneous} \]
2. Answer parts (a) through (e) below, which relate to reactions involving silver ion, \( \text{Ag}^+ \).

The reaction between silver ion and solid zinc is represented by the following equation.

\[
2 \text{Ag}^+(aq) + \text{Zn}(s) \rightarrow \text{Zn}^2+(aq) + 2 \text{Ag}(s)
\]

(a) A 1.50 g sample of \( \text{Zn} \) is combined with 250. mL of 0.110 \( M \) \( \text{AgNO}_3 \) at 25\(^\circ\)C.

(i) Identify the limiting reactant. Show calculations to support your answer.

(ii) On the basis of the limiting reactant that you identified in part (i), determine the value of \([\text{Zn}^{2+}]\) after the reaction is complete. Assume that volume change is negligible.

(b) Determine the value of the standard potential, \( E^\circ \), for a galvanic cell based on the reaction between \( \text{AgNO}_3(aq) \) and solid \( \text{Zn} \) at 25\(^\circ\)C.

Another galvanic cell is based on the reaction between \( \text{Ag}^+(aq) \) and \( \text{Cu(s)} \), represented by the equation below. At 25\(^\circ\)C, the standard potential, \( E^\circ \), for the cell is 0.46 V.

\[
2 \text{Ag}^+(aq) + \text{Cu(s)} \rightarrow \text{Cu}^2+(aq) + 2 \text{Ag}(s)
\]

(c) Determine the value of the standard free-energy change, \( \Delta G^\circ \), for the reaction between \( \text{Ag}^+(aq) \) and \( \text{Cu(s)} \) at 25\(^\circ\)C.

(d) The cell is constructed so that \([\text{Cu}^{2+}]\) is 0.045 \( M \) and \([\text{Ag}^+]\) is 0.010 \( M \). Calculate the value of the potential, \( E \), for the cell.

(e) Under the conditions specified in part (d), is the reaction in the cell spontaneous? Justify your answer.
(a) \[
\frac{0.28 \text{ mol } Ag^+}{1 \text{ mol } Zn^{2+}} 
\rightarrow 0.14 \text{ mol } Zn^{2+}
\] 
\[
0.14 \text{ mol } / 0.250 \text{ L} = 0.560 \text{ m } Zn^{2+}
\]

(b) \[
2Ag^+ + 2e^- \rightarrow 2Ag(s)
\]
\[
2Zn \rightarrow Zn^{2+} + 2e^-
\]
\[
E = \frac{0.142 \text{ V} + 0.764 \text{ V}}{2} = 0.458 \text{ V}
\]

(c) \[
\Delta G = -nFE^0 = -2(96,500 \text{ J/mol V})(0.458 \text{ V})
\]
\[
\Delta G = -88,780 \text{ J/mol}
\]

(d) \[
E_{\text{cell}} = E^0 - \frac{0.0592 \log Q}{2}
\]
\[
E_{\text{cell}} = 0.456 - \frac{0.0592 \log (0.045)}{2}
\]
\[
E_{\text{cell}} = 0.38 \text{ V}
\]

(e) Yes, it is spontaneous since the \( E \) of the cell is a positive number. This means the potential is increases, which happens when a reaction is spontaneous and needs no change or energy to start the reaction.
2. Answer parts (a) through (e) below, which relate to reactions involving silver ion, $\text{Ag}^+$. The reaction between silver ion and solid zinc is represented by the following equation.

$$2 \text{Ag}^+(aq) + \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2 \text{Ag}(s)$$

(a) A 1.50 g sample of Zn is combined with 250. mL of 0.110 $M$ AgNO$_3$ at 25°C.

(i) Identify the limiting reactant. Show calculations to support your answer.

(ii) On the basis of the limiting reactant that you identified in part (i), determine the value of $[\text{Zn}^{2+}]$ after the reaction is complete. Assume that volume change is negligible.

(b) Determine the value of the standard potential, $E^\circ$, for a galvanic cell based on the reaction between AgNO$_3$(aq) and solid Zn at 25°C.

Another galvanic cell is based on the reaction between Ag$^+(aq)$ and Cu(s), represented by the equation below. At 25°C, the standard potential, $E^\circ$, for the cell is 0.46 V.

$$2 \text{Ag}^+(aq) + \text{Cu}(s) \rightarrow \text{Cu}^{2+}(aq) + 2 \text{Ag}(s)$$

(c) Determine the value of the standard free-energy change, $\Delta G^\circ$, for the reaction between Ag$^+(aq)$ and Cu(s) at 25°C.

(d) The cell is constructed so that $[\text{Cu}^{2+}]$ is 0.045 $M$ and $[\text{Ag}^+]$ is 0.010 $M$. Calculate the value of the potential, $E$, for the cell.

(e) Under the conditions specified in part (d), is the reaction in the cell spontaneous? Justify your answer.

\[ \begin{array}{ll}
\text{(d) i)} & \text{2Ag}^+ + \text{Zn}(s) \rightarrow \text{Zn}^{2+} + 2\text{Ag}(s) \\
\text{250 mL} & \text{1.10 M} \\
\text{1.50 g Zn} & \text{1 mol} \\
65.39 g \text{Zn} & = 0.0229 mol \text{Zn} \leftarrow \text{LR} \\
\text{250 mL Ag}^+ & \text{1.10 M} \\
\text{100 mL} & = 0.0275 mol \text{Ag} \leftarrow \text{less mol used up factor} \\
\text{ii)} & \text{0.0229 mol Zn} \\
\text{1 mol Zn} & = 0.0229 mol \text{Zn}^{2+} \\
\text{1 mol Zn} & = \frac{0.0229 \text{ mol Zn}^{2+}}{250 \text{ L}} = 0.0916 M \\
\end{array} \]
ADDITIONAL PAGE FOR ANSWERING QUESTION 2.

b) \[ \text{Ag}^+ + e^- \rightarrow \text{Ag}(s), \quad E^{\circ} = 0.80 \]
\[ \text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}(s), \quad E^{\circ}_{\text{red}} = -0.76 \]
\[ \text{Zn}(s) \rightarrow \text{Zn}^{2+} + 2e-, \quad E^{\circ}_{\text{oxi}} = 0.76 \]
\[ E^{\circ}_{\text{cell}} = 0.80 - 0.76 = 0.04 \text{V} \]
\[ \Delta G^{\circ} = -nF\Delta E^{\circ} \]
\[ = (-3)(96,500 \text{ C/mol}) (0.46 \text{V}) \]
\[ = -133,170 \]

\[ E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0592}{2} \log a_a @ 25^\circ \text{C} \]
\[ 0.46 \text{V} - \frac{0.0592}{2} \log \frac{0.045 M}{0.010 M} \]
\[ E^{\circ}_{\text{cell}} = 0.44 \]

\( \therefore \) Yes, it is spontaneous because the \( E^{\circ}_{\text{cell}} \) is positive. This means the reaction will undergo without reaction.

GO ON TO THE NEXT PAGE.