Worksheet 16: Practice for Exam 3

1. For the reaction shown below, $\Delta G^\circ = -33.3 \text{ kJ/mol}$. In which direction will the reaction be spontaneous at 525 °C if $P_{N_2} = 0.75 \text{ atm}$, $P_{H_2} = 0.75 \text{ atm}$, and $P_{NH_3} = 9.0 \text{ atm}$?

   \[
   \text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})
   \]

   Calculate $Q$:
   \[
   Q = \frac{[\text{NH}_3]^2}{[\text{N}_2] \times [\text{H}_2]^3} = \frac{9.0^2}{(0.75 \times 0.75^3)} = 256
   \]

   $\Delta G = \Delta G^\circ + RT\ln Q = -33,300 \text{ J} + (8.314 \text{ J/mol} \cdot \text{K}) \times (798 \text{ K}) \times \ln(256)
   \]

   $\Delta G = -33,300 \text{ J} + 36,800 \text{ J}
   \]

   $\Delta G = 3.5 \text{ kJ}$

2. Calculate the equilibrium constant for the ammonia synthesis reaction above at 525 °C.

   $\Delta G^\circ = -33.3 \text{ kJ/mol} = -RT\ln K$ so $\ln K = -33,300 \text{ J}/[-(8.314 \text{ J/mol} \cdot \text{K}) \times 798 \text{ K}] = 5.02
   \]

   $\ln K = 5.02$ so $K = e^{5.02} = 151 = K$

3. From problems 1 and 2, compare $Q$ and $K$ with $\Delta G$ at 525 °C temperature. Does it all make sense?

   Yes, it makes sense. $K<Q$, so reaction should go backwards. At the above conditions, $\Delta G$ is positive, meaning the reverse direction is spontaneous.

4. What is the value of $E^\circ_{\text{cell}}$

   \[
   \text{Pb(s)} + \text{Al}^{3+}(\text{aq}) \rightarrow \text{Pb}^{2+}(\text{aq}) + \text{Al(s)} \quad (25 \degree \text{C and 1 M concentrations})
   \]

   \[
   \begin{align*}
   \text{Al}^{3+}(\text{aq}) + 3e^- & \rightarrow \text{Al(s)} \quad E^\circ = -1.66 \text{ V} \\
   \text{Pb}^{2+}(\text{aq}) + 2e^- & \rightarrow \text{Pb(s)} \quad E^\circ = -0.13 \text{ V}
   \end{align*}
   \]

   \[
   E_{\text{cell}} = E_{\text{red}} - E_{\text{ox}} = -1.66 \text{ V} - (-0.13 \text{ V}) = -1.53 \text{ V}
   \]

   The reaction as written is NON-spontaneous, since Al is more active than Pb

5. Calculate $\Delta G^\circ$ for #4

   \[
   \Delta G^\circ = -nFE_{\text{cell}} = 6 \text{ moles e}^- \times 9.65 \times 10^5 \text{ C/mol e}^- \times -1.53 \text{ V}
   \]

   \[
   = 8.86 \times 10^6 \text{ J} \quad \text{again, non-spontaneous}
   \]

6. What would you estimate $K$ to be for the reaction in 4, large or small?

   Small $K$, since $\Delta G$ is large and positive, and $E_{\text{cell}}$ is negative

7. Is the reaction in 4 spontaneous as written? No, $+\Delta G$ and $-E_{\text{cell}}$, Al higher on activity series

8. Draw a cell that would permit reaction 4 to occur, using any pieces of equipment that you would need, and label as much as you can.
See figure below w/ Al = cathode (reduction) and Pb = anode (oxidation)
without the power source, the cell would run in reverse.

Electron flow

Power source
> 1.53 V

Pb anode
(oxidation of Pb(s))

Al cathode
(reduction of Al^{3+})

Salt bridge

[Pb^{2+}] will increase

[Al^{3+}] will decrease