Chemistry 1B, Fall 2015   Quiz #8A KEY

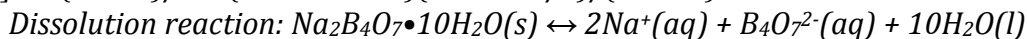**You must show your work for full credit.**

Exp 9 and 10 Questions

1. A 5.0 mL aliquot of a saturated $Na_2B_4O_7 \cdot 10H_2O$ solution is titrated and gives 0.00134 moles of $B_4O_7^{2-}$ present. What is the K value for dissolution? (**2 pts**)

$[B_4O_7^{2-}] = n(B_4O_7^{2-})/V = (0.00134 \text{ mol})(1000 \text{ mL/L})/(5.0 \text{ mL}) = 0.268 \text{ M}$

   Dissolution reaction: $Na_2B_4O_7 \cdot 10H_2O(s) \leftrightarrow 2Na^+(aq) + B_4O_7^{2-}(aq) + 10H_2O(l)$

ICE Table   I   0   0
                                                        C   +2x   +x
                                                        E   2x   x

Since $x = [B_4O_7^{2-}] = 0.268$ M, $[Na^+] = 2x = 0.536$ M and $K = [Na^+]^2[B_4O_7^{2-}] = (0.536)^2(0.268)$
K = 0.077 *(only 2 sig fig based on the volume)   Can give partial credit if parts done correctly*

**2.** When lnK for dissolution of $Na_2B_4O_7 \cdot 10H_2O$ in water is plotted vs. 1/T, a negative slope is seen. Using the equation below, explain whether dissolution of $Na_2B_4O_7 \cdot 10H_2O$ would be expected to result in the solution becoming warmer or cooler. Explain your answer. (Hint: think about how ΔH relates to temperature change from a reaction) (**1 pt**)

$$lnK = \frac{-\Delta H°}{RT} + \frac{\Delta S°}{R}$$

*Since it is reported that the lnK vs. 1/T plot gives a negative slope, this means $-\Delta H°/R$ is negative. Since R is positive, $\Delta H°$ must also be positive. When a reaction has a positive ΔH, the reaction takes heat, meaning that the **solution will get cooler**.*

3. Given the following net equation, write the anode and cathode half reactions. Be sure to include state symbols and electrons. (**2 pt**)
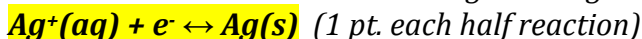
   $Cu(s) + 2Ag^+(aq) \leftrightarrow 2Ag(s) + Cu^{2+}(aq)$
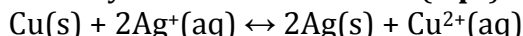
anode half reaction:
   oxidation occurs at the anode. Cu is being oxidized. Anode half reaction is:
   $Cu(s) \leftrightarrow Cu^{2+}(aq) + 2e^-$

cathode half reaction:
   reduction occurs at the cathode. Ag^+ is being reduced. Cathode half reaction is:
   $Ag^+(aq) + e^- \leftrightarrow Ag(s)$ *(1 pt. each half reaction)*

Electrochemistry problems

4. Adding H^+ and H_2O where necessary, balance the following redox reaction. To receive full credit, you need to also give two balanced half reactions. (**3 pts**)

   $HClO(aq) + Fe^{2+}(aq) \leftrightarrow Cl_2(g) + Fe^{3+}(aq)$
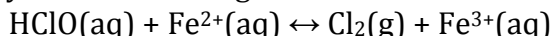
reduction: Cl (from -1 to 0) balanced half-reaction: $2HClO(aq) + 2H^+(aq) + 2e^- \leftrightarrow Cl_2(g) + 2H_2O(l)$
oxidation: $Fe^{2+}(aq) \leftrightarrow Fe^{3+}(aq) + e^-$
Combine (first multiply Fe reaction by 2 so that number of e's is the same):
   $2HClO(aq) + 2H^+(aq) + 2Fe^{2+}(aq) \leftrightarrow Cl_2(g) + 2H_2O(l) + 2Fe^{3+}(aq)$

(one more on back)   (give +1 pt for each part done correctly)

5. Given the following electrochemical reduction potential table, come up with a pair of reactions that will give the greatest potential under standard conditions and calculate the potential under standard conditions. **(2 pts)**

Reaction	E°(V)
$Fe^{3+}(aq) + e^- \leftrightarrow Fe^{2+}(aq)$	+0.771
$Cu^{2+}(aq) + 2e^- \leftrightarrow Cu(s)$	+0.337
$2H^+(aq) + 2e^- \leftrightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \leftrightarrow Pb(s)$	-0.126

The greatest potential will occur when the reaction with the highest reduction potential is coupled to the reaction with the lowest reduction potential (**so this is $Fe^{3+}(aq) + e^- \leftrightarrow Fe^{2+}(aq)$ for the reduction and $Pb(s) \leftrightarrow Pb^{2+}(aq) + 2e^-$ for oxidation**). $E_{cell}° = 0.771$ V – (-0.126 V) = **0.897 V** (1 pt for each part)