1. Calculate the pH of the buffer made with:
   
a. 2.0 g cyanic acid and 2.0 g sodium cyanate in 100.0 mL of solution.

   b. 4.5 g lactic acid and 6.5 g of potassium lactate in 200.0 mL of solution.

2. Calculate the pH of the solution that results from:
   
a. adding 39.30 mL of 0.2000 M KOH to the cyanic acid/cyanate buffer above.

   b. adding 30.37 mL of 0.2078 M HBr to the lactic acid/lactate buffer above.
3. Write the $K_{sp}$ reaction and expression for each of the following partially soluble, ionic compounds.

a. AlPO$_4$

b. BaSO$_4$

c. CdS

d. Cu$_3$(PO$_4$)$_2$

e. CuSCN

f. Hg$_2$Br$_2$

g. AgCN

h. Zn$_3$(AsO$_4$)$_2$

i. Mn(IO$_3$)$_2$

j. PbI$_2$

k. SrCO$_3$

l. Bi$_2$S$_3$

m. Fe(OH)$_3$
4. Which of the compounds in above would not be more soluble in a solution with a pH of 6.00? Explain.

5. Calculate the mass of cation in
   i. 200.0 mL of pure water
   ii. 100.0 mL of 0.50 M sodium salt of the anion for the following:
      a. \( \text{Nd}_2(\text{CO}_3)_3 \), \( K_{sp} = 1.08 \times 10^{-33} \)
      
      b. \( \text{Cd}_3(\text{PO}_4)_2 \), \( K_{sp} = 2.53 \times 10^{-33} \)

      c. \( \text{Hg}_2\text{C}_2\text{O}_4 \), \( K_{sp} = 1.75 \times 10^{-13} \)
6. A saturated solution of Co(OH)$_2$ in pure water has a pH of 9.37. Determine the $K_{sp}$ of Co(OH)$_2$.

7. For each of the following reactions of nitrogen, determine whether it is a) spontaneous at all temperatures, b) non-spontaneous at all temperatures, c) spontaneous at low temperatures, or d) spontaneous at high temperatures.

a. $\frac{1}{2} N_2(g)$ + $O_2(g)$ $\rightleftharpoons$ NO$_2$(g)

$\Delta H^\circ = +8.0$ kJ $\quad \Delta S^\circ = -14.4$ J/K

b. $\frac{1}{2} N_2(g)$ + $\frac{1}{2} O_2(g)$ $\rightleftharpoons$ NO(g)

$\Delta H^\circ = +21.6$ kJ $\quad \Delta S^\circ = +2.9$ J/K

c. $\frac{1}{2} N_2(g)$ + $\frac{3}{2} H_2(g)$ $\rightleftharpoons$ NH$_3$(g)

$\Delta H^\circ = -11.0$ kJ $\quad \Delta S^\circ = -23.7$ J/K

For the reactions above that are spontaneous at low temperatures or spontaneous at high temperatures, calculate the temperature at which the reaction reverses spontaneous direction.
8. For each of the following reactions, indicate whether ΔS should be positive, negative or indeterminate by inspection. Explain, briefly.

a. \( \text{CCl}_4(\ell) \rightleftharpoons \text{CCl}_4(\text{g}) \)

b. \( \text{CuSO}_4 \cdot 3\text{H}_2\text{O(s)} + 2 \text{H}_2\text{O(g)} \rightleftharpoons \text{CuSO}_4 \cdot 5\text{H}_2\text{O(s)} \)

c. \( \text{SO}_3(\text{g}) + \text{H}_2(\text{g}) \rightleftharpoons \text{SO}_2(\text{g}) + \text{H}_2\text{O(g)} \)

d. \( \text{H}_2\text{S(g)} + \text{O}_2(\text{g}) \rightleftharpoons \text{H}_2\text{O(g)} + \text{SO}_2(\text{g}) \) (unbalanced)

e. \( \text{CaO(s)} + \text{H}_2\text{O(\ell)} \rightleftharpoons \text{Ca(OH)}_2(\text{s}) \)

f. \( \text{2 HgO(s)} \rightleftharpoons \text{2 Hg(\ell)} + \text{O}_2(\text{g}) \)

g. \( \text{2 NaCl(\ell)} \rightleftharpoons \text{2 Na(\ell)} + \text{Cl}_2(\text{g}) \)

h. \( \text{Fe}_2\text{O}_3(\text{s}) + 3 \text{CO(g)} \rightleftharpoons 2 \text{Fe(s)} + 3 \text{CO}_2(\text{g}) \)

i. \( \text{F}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2 \text{ClF(g)} \)
9. Using $\Delta H^o$ and $S^o$ data from Appendix 2, calculate $\Delta G^o$ and $K$ for each of the following reactions at 25.0 °C.

a. $\text{SO}_3(g) + \text{H}_2(g) \rightleftharpoons \text{SO}_2(g) + \text{H}_2\text{O}(g)$

b. $\text{CaO}(s) + \text{H}_2\text{O}(l) \rightleftharpoons \text{Ca(OH)}_2(s)$

c. $\text{Fe}_2\text{O}_3(s) + 3 \text{CO}(g) \rightleftharpoons 2 \text{Fe}(s) + 3 \text{CO}_2(g)$
10. Iron ore, Fe$_2$O$_3$, is turned into steel by reduction with carbon at high temperatures:

$$2 \text{ Fe}_2\text{O}_3(s) + 3 \text{ C(g)} \rightarrow 4 \text{ Fe(s)} + 3 \text{ CO}_2(g)$$

Given $\Delta H^\circ = +467.9 \text{ kJ}$ and $\Delta S^\circ = +558.4 \text{ J/K}$, calculate the minimum temperature needed to reduce Fe$_2$O$_3$ with carbon.

11. The K$_{sp}$ reaction of AgBr has a $\Delta G^\circ$ of $+70.2 \text{ kJ}$ at 25.0 $^\circ$C.

   a. Calculate K$_{sp}$.

   b. Calculate $\Delta G$ at 50.0 $^\circ$C when the molarities of the two ions in solution are 1.5 x $10^{-6}$ M.
12. In biochemistry, $\Delta G^\circ$ is often replaced with $\Delta G^\circ'$. The only difference is standard conditions and a pH of 7.00. Calculate $\Delta G^\circ'$ of:

$$\text{glucose(aq) + ATP}^+(\text{aq}) \rightleftharpoons \text{glucose-6-phosphate}^2-(\text{aq}) + \text{ADP}^3-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$$

$\Delta G^\circ = +24.8 \text{ kJ}$

13. The lead storage battery uses the reaction

$$\text{Pb(s) + PbO}_2\text{(s) + 2 H}_2\text{SO}_4(\text{aq}) \rightleftharpoons 2 \text{PbSO}_4(\text{s}) + 2 \text{H}_2\text{O}(\ell)$$

$\Delta G^\circ$ of the reaction is $-371.4$ kJ. Calculate $\Delta G$ when the battery has “run down” to a sulfuric acid concentration of 0.100 M at a temperature of 0.0 °C.
14. A galvanic cell is constructed based on the following reactions and standard reduction potentials (V and W are both conductors):

\[ \text{V(OH)}_4^+(\text{aq}) \rightleftharpoons \text{V(s)} \quad \text{(not balanced)} \quad E^0 = -0.25 \text{ V} \]

\[ \text{WO}_2(\text{aq}) \rightleftharpoons \text{W(s)} \quad \text{(not balanced)} \quad E^0 = -0.12 \text{ V} \]

a. Calculate \( E^0_{\text{cell}} \) for the spontaneous reaction.

b. Balance the spontaneous reaction (in acid).

c. Balance the spontaneous reaction (in base).
d. **Sketch** a diagram of a galvanic cell which could be used to measure the **standard cell potential of the reaction above**. **Clearly** label the following:

1. The oxidation and reduction parts of the cell.
2. The anode and the cathode.
3. The positive and negative electrode.
4. The direction of electron flow in the external circuit and the ion flow internally.
5. The **exact** contents of the cell.
15. A galvanic cell is constructed based on the following reactions and **standard reduction potentials** (Te and Pb are both conductors):

\[
\text{TeO}_4^- \text{(aq)} \rightleftharpoons \text{Te(s)} \quad \text{(not balanced)} \quad E^0 = +0.47 \, \text{V}
\]

\[
\text{HPbO}_2^- \text{(aq)} \rightleftharpoons \text{Pb(s)} \quad \text{(not balanced)} \quad E^0 = -0.54 \, \text{V}
\]

a. Calculate $E^0_{\text{cell}}$ for the **spontaneous reaction**.

b. Balance the **spontaneous reaction** (in acid).

c. Balance the **spontaneous reaction** (in base).
d. **Sketch** a diagram of a galvanic cell which could be used to measure the *standard* cell potential of the reaction above. *Clearly* label the following:

1. The oxidation and reduction parts of the cell.
2. The anode and the cathode.
3. The positive and negative electrode.
4. The direction of electron flow in the external circuit and the ion flow internally.
5. The **exact** contents of the cell.
Polyprotic acids and fully deprotonated conjugate bases from the end of chapter 16 will not be on exam #3, but will be on the final.

1. Calculate the pH of 0.15 M solutions of the following. Show relevant chemical equations for each.
   a. \( \text{H}_3\text{PO}_4 \) (phosphoric acid)

   b. \( \text{H}_2\text{C}_3\text{H}_2\text{O}_4 \) (malonic acid)

   c. \( \text{K}_3\text{C}_6\text{H}_5\text{O}_7 \) (potassium citrate)

   c. \( \text{Na}_2\text{CO}_3 \) (sodium carbonate)