Chem 103 Second Midterm Exam Guide Spring, 2009

The test will be on Week 8 as scheduled in the syllabus, specifically: **Wednesday, May 20, 2009**. The test will mainly focus on material covered since the first midterm: namely, Chapters 17 (starting with sections 17.4), 18 and 19 (at least up to the material covered by Monday in lecture). Study your lecture notes, lecture quizzes and practice problems, written homework assignments and ehomework in that order. Expect to solve problems. Test yourselves by doing problems similar to the homework <u>under</u> time constraints. If you understand the concepts fully, you should be able to do these problems within 10-15 minutes. If not, you need to study and practice further to improve your speed and test taking skills.

Assume that chemical formulas and equations will **not** be supplied for this test.

Try the following problems:

Chapter 17:
1) Write the K_{sp} and K_f equilibria for: Ag_3PO_4 (a sparingly soluble substance) and $[Ni(NH_3)_6]^{2+}$. Which equilibrium applies to what? Solve for the concentration of the free metal in each case. (look up the values of K_{sp} and K_f in the book's appendix).
2) The K_{sp} of $AuCl_3$ is $3.2x10^{-25}$ while the K_f for $[Au(CN)_2^-]$ is $2.0x10^{38}$. Will $AuCl_3$ dissolve in NaCN? (as hint, try problem-solving example, 17.13, page 854).
3) What is amphoterism?
4) What is the solubility of $AuCl_3$ in pure water as the solvent? In .10M HCl? In 0.10M NaCl? Will the K_{sp} change? Is the solubility dependent upon pH? Upon a common ion?

Chapter 18 (thermodynamics)
Qualitative, conceptual material:
1) Contrast a spontaneous process from a nonspontaneous process by naming all the various aspects of each one (for example, the sign of ΔG, the work needed or released, etc).
2) Determine the sign of ΔS for various processes – physical and chemical – for example: based only on the changes in the molecular and atomic arrangements .
3) Be able to compare entropies for various substances under different conditions: e.g. gaseous vs dissolved in solution, etc.
4) Describe entropy and spontaneity. What is the relationship between entropy and ΔG? Is there a conservation of entropy in the universe? Describe the 2nd law.
5) What is the difference between thermodynamic stability and kinetic stability?

Quantitative:
1) Write down the various ways of calculating ΔS.
2) At the freezing point of water, what is ΔS? (assume you know the ΔH_{fusion}).
3) Given the $S°_f$ for A, B and C; what is the $\Delta S°_f$ for 3A -> 1/2 B + 5C?
4) Is $S°_f$ =0 for elements in the elemental form? How about $\Delta G°_f$ and $\Delta H°_f$? What are the units needed when doing calculations involving R?
5) What is the relationship between K_p and K_c? Which value of R will you use in this relationship (ie. R=8.314 J/mol K or R=0.0821 atm-L/molK)?

6) What are the conditions which determine whether a reaction is spontaneous only at high only at low temperatures? Describe 4 different scenarios and give an example of each one.
7) Write down the important equations involving ΔG, $\Delta G°$, Q, etc.
8) Consider the reaction: $CO(g) + 1/2\ O_2(g) \rightarrow CO_2(g)$ and the following thermodynamic values: $\Delta H_f°$: CO(g) =-111 kJ/mol, $CO_2(g)$ =-394 kJ?mol; $S°_f$: CO(g) =198 J/molK, $CO_2(g)$ =214J/molK and $O_2(g)$ = 205 J/molK. At what temperature can these 3 substances coexist at equilibrium each with a partial pressure of 1 bar? At the above temperature, what will be the value of K for this hypothetical reaction?
9) Consider the following reaction under standard conditions. $H_2(g) + 1/2\ O_2(g) \rightarrow H_2O\ (g)$. (see appendix for thermodynamic values): a) What is the maximum work one can obtain from the formation of 1 mole of water vapor? b) What is the K_c (not K_p) of this reaction under these conditions. c) by how much has the entropy of the universe increased assuming 1 mole of H_2O has been formed by the above reaction (indicate if there is an unequal sign)?

Chapter 19:

1) What is the potential for a galvanic cell with the following cell notation? $Mn(s)|Mn^{2+}(1.0\ M)||Mn^{2+}(.00001M)|Mn(s)$.

(If we did not get a chance to discuss cell notation during the Wednesday lecture, here's an explanation. The "|"'s represent the phase boundaries between the components of the galvanic cell. For example, the above notation means a Mn metal electrode immersed in 1 M Mn^{2+} solution, connected by a saltbridge ("||") to a sol'n of .00001M Mn^{2+} sol'n in which Mn(s) is immersed.)

2) Balance the most complicated half reactions in table 18-1. (start with only the redox couples)
3) Balance: $PbO_2 + SO_4^{2-} + Au \rightarrow PbSO_4 + Au^{3+}$
4) In the above, identify the reducing agent. Oxidant. Redox couples.
5) If the above were the cell reaction for a galvanic cell, which would the anode and the cathode? (what are the charges of each electrode?)
6) Write the reactions (from memory) describing the rusting process. Describe and bolster with chemical equations and potentials what happens when iron is galvanized.
7) Write down the chemical reactions present in the following batteries: dry cell, lead car battery and mercury batteries. Describe their distinguishing properties based on the chemical reactions.
8) What is equilibrium constant for a redox reaction whose standard reaction potential is .0134 V? What is the maximum work that this reaction can do?