

## Checklist for Exam 4 (not necessarily complete)

### Section 15.10. Solubility expressions.

- I know the solubility rules well enough to recognize an insoluble salt.
- I can write the solubility equilibrium expression and the  $K_{sp}$  expressions for any “insoluble” salt.

### Section 15.11. Solubility and $K_{sp}$ .

- Given  $K_{sp}$ , I can calculate the molar solubility of a salt in pure water. I know that I must follow the 3-step process of 1. write the equilibrium expression and the associated  $K_{sp}$  expression; 2. Set up the MICE table; and 3. Solve for x, the molar solubility.
- Given the molar solubility, I can calculate  $K_{sp}$  of a salt in pure water. I know that I must follow the same 3-step process as above.
- I know solubility is most conveniently given in molarity, the molar solubility, but sometimes it is given in other units such as mg/L. When this happens, I know I must be able to convert from those units to molarity.
- I can solve for solubility for two or more salts and list them in order of increasing (or decr) molar solubility.

### Section 15.12. Common ion problems.

- Given  $K_{sp}$ , I can calculate the molar solubility of an insoluble salt in the presence of one of the ions from another source. I know that I must follow the 3-step process of 1. Write the equilibrium expression and the associated  $K_{sp}$  expression; 2. Set up the MICE table (this time, one of the two ions will have a numerical value (the molarity of the common ion) in the “I” line); and 3. Solve for x, the molar solubility. Don't forget UNITS!

### Section 15.13. Will a precipitate form?

- Given  $K_{sp}$  and the initial concentrations of two ions, I can determine if a precipitate will form. This involves: 1. write the equilibrium expression and the associated  $K_{sp}$  expression; 2. Set up the MICE table (we will only use the “I” line), but this line needs to be filled in correctly (see next item); 3. Write the  $Q_{sp}$  expression for the “I” line; 4. Compare  $Q_{sp}$  and  $K_{sp}$ .
- One “issue” is that the concentrations you want to use in the “I” line may not be the ones listed if a “mutual dilution” has taken place. This can be determined by the wording of the problem. If two solutions were mixed, they mutually diluted one another and you must use the dilution formula,  $M_c V_c = M_d V_d$  in order to determine the correct concentrations to list in the “I” line.

## Chapter 16 Entropy and free energy

- I know how to write equilibrium expressions for weak acids, weak bases and water as we learned in Chapter 14.

### Section 16.1 and 2. Predicting the signs of $\Delta H$ , $\Delta S$ and $\Delta G$ .

- I know how to assess a process or chemical reaction and assign signs (+, 0, -) for each  $\Delta H$ ,  $\Delta S$ , and  $\Delta G$ .
  - I know for  $\Delta H$ , I ask myself, “Does this take heat or give off heat?”
  - For  $\Delta S$ , I ask myself, “Does this create more disorder or less disorder?”
  - For  $\Delta G$ , I ask myself, “Does this or can this process happen at the temperature indicated?”

### Sections 16.4 – 16.6 Calculating $\Delta H^\circ$ , $\Delta S^\circ$ and $\Delta G^\circ$ from the tables.

- I know how to calculate  $\Delta H^\circ$ ,  $\Delta S^\circ$  and  $\Delta G^\circ$  from the standard tables.
- Alternatively, once I have  $\Delta H^\circ$  and  $\Delta S^\circ$ , I can calculate  $\Delta G^\circ$  from the equation,  $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ .

### Sections 16.7 – 16.9 Using $\Delta G = \Delta H - T\Delta S$ and $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ .

- I know that for  $\Delta H$  and  $\Delta S$  are relatively temperature independent as long as we do not stray too far from 298 K. This allows us to estimate  $\Delta G$  at temperatures other than 298 K by using the standard tables for  $\Delta H^\circ$  and saying this is approximately the same as  $\Delta H$ . Ditto for  $\Delta S$ . We can then approximate  $\Delta G = \Delta H^\circ - T\Delta S^\circ$ .
- Given the signs for  $\Delta H$  and  $\Delta S$ , I can determine at what temperature conditions the reaction will be spontaneous, if ever, This is the “4 scenarios” view of things:
  - $\Delta H > 0$  and  $\Delta S > 0$ , the reaction is spontaneous at high temperatures but not low temperatures.
  - $\Delta H < 0$  and  $\Delta S < 0$ , the reaction is spontaneous at low temperatures but not high temperatures

- $\Delta H < 0$  and  $\Delta S > 0$ , the reaction is always spontaneous at any temperature.
- $\Delta H > 0$  and  $\Delta S < 0$ , the reaction is never spontaneous.

### Sections 16.10 – 16.11 $\Delta G$ , $K_p$ , and $Q_p$

- I know the formula  $\Delta G = \Delta G^\circ + RT \ln Q_p$  and how to use it to determine if a process is spontaneous under non-standard conditions.
- I know that at equilibrium,  $\Delta G = 0$ , so the formula  $\Delta G = \Delta G^\circ + RT \ln Q_p$  becomes  $\Delta G = 0 = \Delta G^\circ + RT \ln K_p$  or  $\Delta G^\circ = -RT \ln K_p$ . I can use this formula as well.

## Chapter 17.

### Section 4.9. Balancing oxidation-reduction reactions.

- I know how to balance redox reactions in acidic solution.
- I know how to balance redox reactions in basic solution.

### Sections 17.1 and 2. Galvanic cells.

- I know how to sketch a galvanic cell.
- I can label the cathode cell, the anode cell, and the direction of electron flow.
- I can write the cathode half-reaction, the anode half-reaction and I can combine them (overall reaction.)
- I can read and write cell notation for galvanic cells.

### Section 17.3. $E$ and $\Delta G$ and $E^\circ$ and $\Delta G^\circ$ .

- I know when and how to use these equations, provided on the exam:

$$\Delta G = -nFE$$

$$\Delta G^\circ = -nFE^\circ$$

### Section 17.4 and 5. Using a Table of Standard Reduction Potentials.

- I can use a Table of Standard Reduction Potentials in order to calculate  $E^\circ$  for any pair of half-reactions. In addition, I can write them as a spontaneous overall reaction.
- I know that I do not multiply  $E^\circ$  values even though I may have to multiply half-reactions in order to balance a pair of half-reactions.
- I can use a Table of Standard Reduction Potentials in order to determine qualitatively if one reagent will react spontaneously with another, as long as both of them are in the table.
- I know that the least common multiple is always the number of electrons transferred,  $n$ . I know this is an important number when using the Nernst and other equations.

### Section 17.6 and 7. Nernst equation.

- Given the Nernst equation on the exam, I can calculate  $E$  for any reaction as long as I know the concentrations of the aqueous species present. I can write  $Q_c$  expressions and I know that I must remember to include coefficients as exponents when writing  $Q_c$ .
- I know when the reaction contains  $H^+$  or  $OH^-$ , I could be asked to express my answer as a pH. Also, given a pH, I could convert to  $[H^+]$  or  $[OH^-]$  and solve the problem.

### Section 17.8. At equilibrium.

- I know that  $E = \Delta G = 0$  at equilibrium. This lets me convert the Nernst equation into  $E^\circ = (0.0592/n) \log K$ . As was the case with the Nernst equation, I can write  $K_c$  expressions and I know that I must remember to include coefficients as exponents when doing so. Moreover, I can solve for  $K$  given  $E^\circ$ , and visa versa.

### Section 17.12 and 14. Electrolytic cells.

- I can write half-cell reactions for the two processes occurring in any electrolytic cell.
- I can convert current, time and charge (given two, I can calculate the 3<sup>rd</sup>), using the equations, given:  
 $Charge(coul) = Current(amps) \times time(s)$   $F = 96500 C/mol = 1 \text{ faraday} = 1 \text{ mol } e^- = 96500 \text{ coul}$
- Knowing the charge, I can convert charge to moles of electrons,  $n_{e^-}$ , and then  $n_{e^-}$  into moles of an element,  $n_E$ . I will be careful to remember that for many elements, more than one equivalent of electrons are required to reduce the cation to the neutral metal.