

2017 Checklist for Exam 4 (not necessarily complete)

Section 16.10. 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 16.11. 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 16.12. 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 16.13. 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 17 Entropy and free energy

Section 17.1 and 17.2. Predicting the signs of ΔH , ΔS and ΔG .

- I know how to assess a process or chemical reaction and assign signs (+, 0, -) for each ΔH , ΔS , and ΔG .
 - I know for ΔH , I ask myself, "Does this take heat or give off heat?"
 - For ΔS , I ask myself, "Does this create more disorder or less disorder?"
 - For ΔG , I ask myself, "Does this or can this process happen at the temperature indicated?"

Sections 17.5 Calculating ΔH° , ΔS° and ΔG° from the tables.

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

Sections 17.7 – 17.9 Using $\Delta G = \Delta H - T\Delta S$ and $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$.

- I know that for ΔH and ΔS are relatively temperature independent as long as we do not stray too far from 298 K. This allows us to estimate ΔG at temperatures other than 298 K by using the standard tables for ΔH° and saying this is approximately the same as ΔH . Ditto for ΔS . We can then approximate $\Delta G = \Delta H^\circ - T\Delta S^\circ$.
- Given the signs for ΔH and ΔS , 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 17.10 – 17.11 ΔG , 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 18. Electrochemistry

Section 18.1. Balancing oxidation-reduction reactions. (or as per my lecture notes)

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

Sections 18.2 and 3. 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 18.4. E and ΔG and E° and ΔG° .

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

$$\Delta G = -nFE$$

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

Section 18.5 and 6. Using a Table of Standard Reduction Potentials.

- I can use a Table of Standard Reduction Potentials in order to calculate E° 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° 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 18.7 and 8. 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 18.9. 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° , and visa versa.