Textbook Chapters for Exam 3

Chapter 8 (you can skip polyprotic acids - section 7 and section 3 for that matter). You're responsible for any material covered in class whether it was on the homework or not. This review sheet covers most everything we have done. Also remember that no equations are given on the exams. Values for equilibrium constants are given where needed.

Chapter 8 – Buffers, pH curves, Titrations, and Indicators

• Know the basic premise for making a **buffer solution**. These are "type 2" problems according to Dr. McCord (see acid/base help sheets on our web site)

• Know HOW a good buffer will neutralize both acid and/or base. Yes, this means know the actual reactions that do the neutralizing.

• What is buffer capacity (section 8.4)? remember, the amounts (concentrations) determine buffer capacity and the ratio of conjugates determines the actually pH of the buffer.

• Be able to use and identify the **Henderson-Hasselbalch equations** for both acids and bases.

• Be able to calculate the pH of any buffer, acidic or basic.

• What is the **common ion effect**? Where an ion has more than one source. A^- comes a little bit from HA and a lot from NaA, so A^- is a common ion.

• Know the 2 ways to prepare buffers:

 1. Mix the two conjugates with proper molar ratios: e.g. HA and NaA

 2. Partial neutralization. : e.g. Neutralize a portion of HA with NaOH to MAKE the A^- needed. You are essentially doing a titration but stopping somewhere in the middle. After all, that IS where you will have a nice mixture of BOTH acid and conjugate base. This is like problems 4 and 6 on homework 5.

• Be able to calculate the new pH of a buffer AFTER the addition of strong acid or strong base. Remember that you will always be subtracting from one species and adding to the other in this calculation. For example, if my acid/conjugate base ratio is 50 mmol HA and 50 mmol A^- and I add 15 mmol of OH^-. After the addition the HA is now 50 - 15 = 35 mmol, and the A^- is now 50 + 15 = 65 mmol. That 65/35 ratio is the new ratio that governs the buffer pH. This is what is going on on problems 12 and 13 from homework 5.

• Know how to interpret a pH curve and get the stoichiometric point (aka: equivalence point) and the pK_a (or pK_b) for a weak acid (or base).

• In general, what's the pH at the stoichiometric point of a titration of a strong acid with strong base? a weak acid with a strong base? a weak base with a strong acid? Know the answer to this in general first - is the pH equal to, greater than, or less than 7.00 (no calculation necessary). Then know the answer exactly which means you'd have to actually calculate the answer.

• Be able to calculate all the points on a titration curve (pH curve). This really means you now have a full knowledge of acid/base theory from start to finish. You start with only HA (chapter 7 stuff, Type 1 equations), you titrate a bit and now have a buffer (Chapter 8, buffer stuff, Type 2 equations), you then reach the stoichiometric point (back to Type 1 equations - now the conjugate of what you started with), and finally you keep going and overshoot the end point (governed only by excess titrant).

• Know how indicators work and how to chose the correct indicator for a given titration.

• Know the approximate range of an indicator (in general, when given K_a). Remember, the center of its range is where pH = pK_a for the indicator. The range is ±1 off the pK_a value.

• What color will an indicator be at a given pH? (you would know the K_a). Is it red? yellow? orange? yellowish orange? bluish green? reddish purple? ... you get the idea. Remember to split the 2-pH range into 3 regions: ±0.3 is the center and is the "perfect" blend color (like green). Outside of that range and up to ±1 is the "-ish" colors (like bluish-green, or yellowish green). Outside the ±1 are the straight colors (blue and yellow). See problems 26-28 on homework 5.

• No exam questions with polyprotic acids or bases (section 7). Also nothing on the "Exact Treatment of Buffered Solutions" which is section 3.

• I also pointed out in class how useful a fraction of species diagram is for a titration (and buffer formation). I pointed out equivalent points on both the fraction of species diagram and a pH curve. Once again, *no polyprotics* - so the fraction of species plots are really very simple with just HA and A^- present at the various pHs.

ALSO: It should go without saying but I'll say it anyway. You still have to know Chapter 7 stuff in order for Chapter 8 to make sense. Chapter 8 concepts and principles are *built* on Chapter 7 ones.

MORE Chapter 8 - Solubility Equilibria

• Know how to calculate the molar solubility (x) from K_{sp}

• Know how to calculate K_{sp} from molar solubility. Do realize that for many problems you must first *convert* the plain ol' solubility (g/L, mg/L, g/100mL, ppm, ppb, etc...) into molar solubility first, then convert that to K_{sp}.

• Know also how to calculate the apparent solubility in the presence of a common ion. This is where one of the ions concentrations is already SET in the solubility product expression. My example in class was dissolving AgCl into water (no common ion) or into 0.01 M NaCl (common ion of Cl⁻).

• I already said this above but I'll say it again: be able to calculate molar solubility (x) from "plain" solubility which can be expressed as grams per liter or, in general, mass per unit volume like ppm (mg/L) or even g/100 mL..

• Know how to calculate all final concentrations in saturated solutions. OK, so you can calculate x from K_{sp}. Now tell me what the <u>actual</u> concentration of Mg^{2+} is, or Cl⁻, or OH⁻, or etc... sometimes it is x, sometimes it's $2x$, or $3x$, and so on. Remember, x isn't always the answer.

• REMEMBER: Not all salts have the same solution for x from K_{sp}. There are 1:1 salts, 1:2 salts, 1:3 salts, 2:3 salts and so on. Also remember that $3^3 = 27$ and not 9.

• Be able to predict whether a precipitation will occur. This is just comparing Q_{sp} to K_{sp}. If a precipitation does occur, how much does precipitate? What are the concentrations of the ions in solution after the precipitation?

• Know what will precipitate 1st, 2nd, etc.. in solutions containing many different ions that can precipitate. Remember that to do this you must calculate the minimum concentration of the added ion in order to reach the saturation point (solve the K_{sp} expression). The one that precipitates first is the one with the *lowest* concentration needed of the added ion.

• Know how to solve questions about fractional precipitation or selective precipitation. What % of the 1st precipitate is precipitated (or % NOT precipitated) when the 2nd precipitate first starts to precipitate? This is number 8 on homework 6.

• What do you add/use to get certain insoluble compounds to dissolve. Usually this involves some sort of chemical reaction with one of the ions of the salt. As the ion from the salt reacts, Le Chatlier's Principle takes over and the reaction shifts to the right thus dissolving more salt. Check numbers 14 and 16 from homework 6.

• Know what a complex ion is and how they dissociate. Know how to write a K_f or K_d expression and how to SOLVE it. Remember that K_f is a formation constant and K_d is a dissociation constant and they have an inverse relationship. Check out numbers 15 and 17 on homework 6.

Now your turn

Now you start writing about all this stuff. Read your book. Over and over. Read other books if you need more help and insight. The more you expose yourself to the concepts the more things will make sense.

Remember, I didn't purposely leave anything out here to trick you on the exam. If I talked about it in class it is fair game on the exam. Practice working problems in the book and on the homework.