1. Rank the following salts from most to least soluble: Ag_2CO_3 ($K_{sp} = 8.1 \times 10^{-12}$), BaF_2 ($K_{sp} = 1.7 \times 10^{-6}$), $Pb(OH)_2$ ($K_{sp} = 2.8 \times 10^{-16}$).

- a. $Ag_2CO_3 > Pb(OH)_2 > BaF_2$
- b. $BaF_2 > Pb(OH)_2 > Ag_2CO_3$
- c. $Ag_2CO_3 > BaF_2 > Pb(OH)_2$
- d. $BaF_2 > Ag_2CO_3 > Pb(OH)_2$

All salts have the same number of ions so you just compare the Ksp. The salt with the largest Ksp is most soluble. If they didn't have the same number of ions then you would need to calculate molar solubility.

2. What is the molar solubility of lead iodide, PbI2, in a .1 M solution of lead nitrate, Pb(NO3)2? The Ksp of lead iodide is 10[^]-9.

a. 4 x 10^-6 b. 5 x 10^-5 c. 6 x 10^-4 d. 7 x 10^-3 PbI2(s) --> Pb2+ + 2 I-.1 0 +2x ~ +χ .1+x 2x $Ksp = 10^{-9} = [Pb][1]^{2} = [.1+x][2x]^{2}$ In this case, x is really small because the Ksp is low. So, we approximate .1 + x as about equal to .1 $10^{-9} = .1^{+4}x^{2}$ $10^{-8} = 4x^{2}$ $x = ((1/4) * 10^{-8})^{1/2} = .5 * 10^{-4} = 5 * 10^{-5}$

3. After combining the following solutions, which will not produce a buffer?

a. 200 ml of .5 M ammonium chloride + 100 ml of .4 NaOH

b. 200 ml of 1 M ammonia + 200 ml of .5 M HCl

c. 200 ml of .1 M pyridinium chloride + 100 ml of .05 M NaOH

d. 100 ml of .1 M pyridine + 100 ml of .05 M NaOH

Pyridine is a base. Adding a strong base to a weak base will result in a strong base solution, not a buffer.

4. 1M solution of NaOH is used to titrate a 1M solution of acetic acid (Ka of acetic acid = 10^{-5}). What is the pH at equivalence point?

a. 11.5

b. <mark>9.5</mark>

c. 4.5

d. 2.5

this is a titration of OH- + HA --> A- (+H2O). At the equivalence point, [OH-] = [HA], so there is only [A-] left in solution. Thus we use the weak base equation to solve for pH. Ka*Kb = Kw = 10^{-14} Kb = Kw/Ka = $10^{-14}/10^{-5} = 10^{-9}$ $[OH-] = (CbKb)^{1/2} = (1*10^{-9})^{1/2} = 10^{-4.5}$ pOH = 4.5 pH = 14 - pOH = 14-4.5 = 9.5 You can also use process of elimination and get rid of C and D: the pH will be greater than 7 since A-(acetate) is a weak base.

5. In an aqueous solution of NH₄NO₂, how many equations do you need to solve all the unknowns? what is the charge balance?

a. 6, [NH₄+] + [H+] = [OH-] + [NO₂-] b. 4, [NH₄+] + [H+] = [OH-] + [NO₂-] c. 6, 4[NH₄⁺] = 2[NO₂-] d. 4, 4[NH₄⁺] + [H+] = [OH-] + 2[NO₂-] NH₄NO₂ will dissociate into NH₄+ and NO₂- in solution. Remember to also take into consideration other species in the solution: H+, OH-, NH₃, HNO₂. Since there are 6 unknowns, we need 6 equations to solve for them.

6. You mix sulfurous acid in water. What is Ka2?

- a. $K_{a2} = [HSO_3^-] / [H^+][SO_3^{-2}]$
- b. $K_{a2} = [H^+][HSO_3^-] / [H_2SO_3]$
- c. $K_{a2} = [H^+][SO_3^{-2}] / [HSO_3^{-1}]$ d. $K_{a2} = [SO_3^{-2}] / [HSO_3^{-1}]$
- $H_2SO_3 \rightleftharpoons H^+ + HSO_3^-$

 $HSO_3^- \neq H^+ + SO_3^{-2}$ Ka2 is for the second deprotonation of sulfurous acid. $K_{a2} = [H^+][SO_3^{-2}]/[HSO_3^{-1}]$

7. Calculate the pH when 100 mL of 1 M NaOH is added to 100 mL of 1 M H₂SO_{3.} The approximate acid dissociation constants for sulfurous acid are: $pK_{a1}=2 pK_{a2}=7$.

- a. 4.5
- b. 2
- c. 3
- d. 8

The remaining solution will be HSO₃⁻ only. The pH is therefore the average of the two dissociation constants.

8. ranking concentrations of polyprotic acids in solution

Rank all species of sulfurous acid at pH 4.

- a. $H_2SO_3 > HSO_3^- > SO_3^{-2}$ b. $HSO_3^- > H_2SO_3 > SO_3^{-2}$

- c. $SO_3^{-2} > H_2SO_3 > HSO_3^{-1}$ d. $H_2SO_3 > SO_3^{-2} > HSO_3^{-1}$ e. $SO_3^{-2} > HSO_3^{-2} > HSO_3^{-2}$ f. $HSO_3^{-2} > SO_3^{-2} > H_2SO_3$

At pH 4 sulfurous acid has lost one proton because the first pKa is roughly 2. HSO3⁻ is the dominant species because the second pK_a has not been reached yet. Also, since the second deprotonation has not been reached the fully deprotonated species is least present.