CHM152 – Chapter 15: Applications of Aqueous Equilibria – Homework

1. What is the pH of a solution prepared by mixing 50.00 mL of 0.10M HCN with 50.00 mL of 0.050 M NaCN? Assume that the volumes are additive and that $K_a = 4.9 \times 10^{-10}$ for HCN.

Two approaches. 1) Use H-H Equation for buffer solution OR 2) Use ICE table

1) $HCN (aq) + H_2O (l) \rightleftharpoons CN^- (aq) + H_3O^+ (aq)$

acid *c. base*

$pH = pK_a + \log([base]/[acid])$

$pH = pK_a + \log(0.025/0.050) =$ **9.01**

2) $HCN (aq) + H_2O (l) \rightleftharpoons CN^- (aq) + H_3O^+ (aq)$

0.050M		0.025	
-x		+x	+x
0.050 – x		0.025 + x	x

$4.9 \times 10^{-10} = \frac{x(0.025 + x)}{(0.050 - x)}$ Assume x is small $= 4.9 \times 10^{-10} = \frac{x(0.025)}{(0.050)}$ $x = 9.8 \times 10^{-10}$

$pH = -\log(9.8 \times 10^{-10}) =$ **9.01**

2. What is the pH of a solution prepared by mixing 50.00 mL of 0.10 M ammonia with 25.00 mL of 0.10 M ammonium chloride? Assume that the volume of the solutions are additive and that $K_b = 1.8 \times 10^{-5}$ for ammonia.

$NH_3 (aq) + H_2O (l) \rightleftharpoons NH_4^+ (aq) + OH^- (aq)$

base *c. acid*

$0.05000L(0.10\ M\ NH_3)/0.07500L = 0.067\ M\ NH_3$

$0.02500L(0.10\ M\ NH_4^+)/0.07500L = 0.033\ M\ NH_4^+$

$pH = -\log(K_w/K_b) + \log([base]/[acid])$

$pH = 9.26 + \log(0.067/0.033) =$ **9.56**

3. Which one of the following sets of compounds is not characteristic of a buffer?
   a. HNO_3 with $NaNO_3$ Buffers are weak acids with their c. base OR weak bases with their c. acids
   b. NaH_2PO_4 with Na_2HPO_4
   c. NH_3 with NH_4Cl
   d. CH_3COOH with $NaCH_3COO$

4. What is the pH of a buffer system made by dissolving 17.42 g of KH_2PO_4 and 20.41 g of K_2HPO_4 in water to give a volume of 200.0 mL? The K_{a2} for dihydrogen phosphate is 6.2×10^{-8} and the equilibrium reaction of interest is

$$H_2PO_4^- (aq) + H_2O (l) \rightleftharpoons H_3O^+ (aq) + HPO_4^{2-} (aq)$$ (Hint: What is HA? A^-?)

17.42 g KH_2PO_4 (1 mol/136.09 g) = 0.1280 mol / 0.2000L = 0.6400M $H_2PO_4^-$ (acid)

20.41 g K_2HPO_4 (1mol/174.18 g) = 0.1172 / 0.2000 L = 0.5859 M HPO_4^{2-} (c. base)

$pH = -\log(6.2 \times 10^{-8}) + \log(0.5859/0.6400) =$ **7.17**

5. What is the pH of a buffer system made by dissolving 10.70 g of CH_3NH_3Cl and 20.00 mL of 12.0 M CH_3NH_2 in enough water to make 500.0 mL of solution? $K_b = 3.7 \times 10^{-4}$ for CH_3NH_2.

$CH_3NH_2 (aq) + H_2O (l) \rightleftharpoons CH_3NH_3^+ (aq) + OH^- (aq)$

$pH = 10.57 + \log(0.48/0.31690 =$ **10.75**

6. What is the $[CH_3COO^-] / [CH_3COOH]$ ratio necessary to make a buffer solution with a pH of 4.44? $K_a = 1.8 \times 10^{-5}$ for CH_3COOH.

$4.44 = 4.74 + \log([CH_3COO^-]/[CH_3COOH]) = -0.30 = \log[CH_3COO^-]/[CH_3COOH]$ $10^{-0.30} =$ **0.50:1 ratio**

7. What is the resulting pH when 0.00500 moles of KOH is added to 0.100 L of a buffer solution that is 0.100 M in $H_2PO_4^-$ and 0.100 M HPO_4^{2-} and the $K_{a2} = 6.2 \times 10^{-8}$?
Since a strong base is added, the acid part of the buffer solution

$H_2PO_4^-$ (aq) +	OH^- (aq) →	H_2O (l) +	HPO_4^{2-} (aq)
0.0100 mol	0.0050 mol		0.0100 mol
-0.0050 mol	-0.0050 mol		+0.0050 mol
0.0050 mol	0		0.0150 mol
0.050 M			0.150 M

pH = 7.21 + log (0.15/0.050) = **7.69**

8. What is the ***change*** in pH when 0.00500 moles of HCl is added to 0.100 L of a buffer solution that is 0.200 M in CH_3COOH and 0.200 M $NaCH_3COO$? The K_a for acetic acid is 1.8×10^{-5}.
Find the initial pH of the buffer before the acid is added: Since the weak acid and it's c. base are equal molarities and volumes, the pH = pKa = 4.74.
Now find the pH after the strong acid is added. Since strong acid is being added, the c. base reacts with it.

CH_3COO^- (aq) +	H_3O^+ (aq) →	CH_3COOH (aq) +	H_2O (l)
0.0200 mol	0.00500 mol	0.0200 mol	
-0.00500 mol	-0.00500 mol	+0.00500 mol	
0.0150 mol	0	0.0250 mol	
0.150 M		0.250 M	

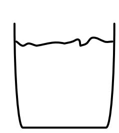
pH = 4.74 + (0.150/0.250) = **4.52 change = pH went down 0.22**

9. Describe whether the *resulting* solution when the two reagents are mixed will be acidic, basic or neutral at the equivalence point.
   a. CH_3COOH titrated with NaOH basic
   b. HF titrated with KOH basic
   c. HCl titrated with NaOH neutral
   d. CH_3NH_2 titrated with HCl acidic

10. Formic acid (HCOOH, $K_a = 1.8 \times 10^{-4}$) is the principal component in the venom of stinging ants. What is the molarity of a formic acid solution if 25.00 mL of the formic acid solution requires 29.80 mL of 0.0567 M NaOH to reach the equivalence point?
Use stoichiometry to solve:
0.02980 L (0.0567 mol NaOH/ 1 L)(1 mol formic acid/ 1mol NaOH) = 1.69×10^{-3} mol formic acid / 0.0250 L = 0.0676 M formic acid

11. What is the pH at the equivalence point of a weak base-strong acid titration if 20.00 mL of NaOCl requires 28.30 mL of 0.50 M HCl? $K_a = 3.0 \times 10^{-8}$ for HOCl.
0.02830 L(0.50 mol HCl /1 L)(1 mol OCl^-/ 1mol HCl) / 0.04830 L = 0.293 M OCl^-
Solve the ICE table for $x = 9.4 \times 10^{-5}$ pH = 4.03

12. Consider the titration of 200.0 mL of 0.150 M hydrofluoric acid in a flask with 0.175 M sodium hydroxide in a burette and answer the following: ($K_a = 3.5 \times 10^{-4}$ for HF (aq))
   a. How many mL of NaOH must be added to reach the equivalence point?
   171 mL
   b. What is the pH after 38.5 mL of NaOH is added?
   pH = 2.919 = 2.92
   c. What will the resulting solution look like at the equivalence point?
   Na^+ and F^- are the only ions that should be in solution.
   All the H^+ and OH^- are reacted to form water.



13. What is the pH of a solution made by mixing 30.00 mL of 0.10 M acetic acid with 40.00 mL of 0.10 M KOH? Assume that the volumes of the solutions are additive. $K_a = 1.8 \times 10^{-5}$ for acetic acid.

There is more OH^- than acetic acid so find the moles of OH^- left and determine pH from that.

CH_3COOH (aq) +	OH^- (aq) →	CH_3COO^- (aq) + H_2O (l)
3.0×10^{-3} mol	4.0×10^{-3} mol	
-3.0×10^{-3} mol	-3.0×10^{-3} mol	
0	1.0×10^{-3} mol	

$[OH^-]$ left = 0.0143 M so pOH = 1.85 and **pH = 12.15**

Acetate ion can further react with water to make acetic acid and hydroxide ion. So more hydroxide ion is produced but the amount is so small it is negligible.

14. What is the pH of a solution made by mixing 30.00 mL of 0.10 M acetic acid with 30.00 mL of 0.10 M KOH? Assume that the volumes of the solutions are additive. $K_a = 1.8 \times 10^{-5}$ for acetic acid.

CH_3COOH (aq) +	OH^- (aq) →	CH_3COO^- (aq) + H_2O (l)
3.0×10^{-3} mol	3.0×10^{-3} mol	
-3.0×10^{-3} mol	-3.0×10^{-3} mol	$+3.0 \times 10^{-3}$ mol
0	0	

All the acetic acid and hydroxide ion are consumed to form 3.0×10^{-3} mol of acetate ion. It can react with water:

CH_3COO^- (aq) +	H_2O (l) ⇌	CH_3COOH (aq) +	OH^- (aq)
3.0×10^{-3} mol			
- x		+x	+x
M = 0.050 M			

Use K_b: $5.6 \times 10^{-10} = x^2 / 0.050$ x = 5.3×10^{-6} M OH^- **pH = 8.72**

15. What is the chromium ion concentration for a saturated solution of $Cr(OH)_3$ if the K_{sp} for $Cr(OH)_3$ is 6.7×10^{-31}?

$Cr(OH)_3$ (s) ⇌ Cr^{3+} (aq) + 3 OH^- (aq) Ksp = $[Cr^{3+}][OH^-]^3$

x 3x

$K_{sp} = 6.7 \times 10^{-31} = x(3x)^3$ x = 1.25×10^{-8} M Cr^{3+}

16. What is the silver ion concentration for a saturated solution of Ag_2CO_3 if the K_{sp} for Ag_2CO_3 is 8.4×10^{-12}?

Ag_2CO_3 (s) ⇌ 2 Ag^+ (aq) + CO_3^{2-} (aq) Ksp = $[CO_3^{2-}][Ag^+]^2$

2x x

$K_{sp} = 8.4 \times 10^{-12} = x(2x)^2$ x = 1.28×10^{-4} M $[Ag^+] = 2x = 2.6 \times 10^{-4}$ M

17. What is the most soluble salt in the following set?
   a. AgCN with $K_{sp} = 6.0 \times 10^{-17}$
   b. $Al(OH)_3$ with $K_{sp} = 1.9 \times 10^{-33}$
   c. $Fe(OH)_3$ with $K_{sp} = 2.6 \times 10^{-39}$
   d. $Sn(OH)_2$ with $K_{sp} = 1.6 \times 10^{-19}$

18. Calculate the solubility (in g/L) of silver carbonate in water at 25°C if the K_{sp} is 8.4×10^{-12}.

Refer to question #16 for the molar solubility. x = 1.28×10^{-4} M

For solubility in g/L multiply the molar solubility by the molar mass: 1.28×10^{-4} M(275.78 g/mol) = **3.5×10^{-2} g/L**

19. What is the molar solubility of CaF_2 in 0.10 M NaF solution at 25°C if the K_{sp} is 1.5×10^{-10}.

x = 1.5 x 10-8 M

20. What is the molar solubility of $Mg(OH)_2$ in a basic solution with a pH of 12.00 at 25°C if the K_{sp} is 5.6×10^{-12}.

$Mg(OH)_2 \rightleftharpoons Mg^{2+}(aq) + 2\ OH^-(aq)$ pH =12.00 has pOH = 2.00 so $[OH^-] = 10^{-2.00} = 0.010$ M

x 0.010 – 2x

$5.6 \times 10^{-12} = x\,(0.010 - 2x)^2$ Assume x is small

$5.6 \times 10^{-12} = x\,(0.01)^2$ $x = 5.6 \times 10^{-8}$ M

If you wanted to have a little fun, you could calculate the molar solubility of magnesium hydroxide in water and see how a basic pH affects the solubility!

$Mg(OH)_2 \rightleftharpoons Mg^{2+}(aq) + 2\ OH^-(aq)$

x 2x

$5.6 \times 10^{-12} = x\,(2x)^2$ $5.6 \times 10^{-12} = 4x^3$ $x = 1.1 \times 10^{-4}$ M

So having a high pH means more OH^- around in the solution which is a common ion. It decreases the solubility of $Mg(OH)_2$!

21. What will occur when calcium nitrate is added to a saturated solution of calcium hydroxide?

Adding a common ion always decreases the solubility of the insoluble compound. In this case calcium hydroxide is slightly soluble in water and by adding calcium nitrate (Ca^{2+} is the common ion) the calcium hydroxide becomes even less soluble.