CHEM 1202 - Homework # 7 & 8 **Name** ANSWER KEY

Acids & Bases 2

Due Tuesday, Nov 28th, 2006 by Noon

Table 1. Dissociation Constants for some Acids.

Acid	pK_a Value	Acid	pK_a Value	Acid	pK_a Value
NH_4^+	10	HBF_4	–9	formic	4
HClO	8	H_2CO_3	7	benzoic	5

1. (5 pts) Which of the acids listed in Table 1, given a 0.01 M solution in water, will have a pH closest to **2**?

a) NH_4^+ b) benzoic c) H_2CO_3 d) formic e) HBF_4

2. (5 pts) Which of the acids listed in Table 1, when reacted with an equivalent amount of NaOH, will form a solution with the ***highest*** pH?

a) NH_4^+ b) benzoic c) H_2CO_3 d) formic e) HBF_4

3. (5 pts) Order the acids in Table 1 from ***strongest*** to *weakest*. Circle the correct choice.

a) benzoic > formic > H_2CO_3 > HBF_4 > NH_4^+ > HClO

b) HBF_4 > formic > benzoic > H_2CO_3 > HClO > NH_4^+

c) NH_4^+ > benzoic > HBF_4 > formic > H_2CO_3 > HClO

d) benzoic > HBF_4 > NH_4^+ > formic > H_2CO_3 > HClO

e) NH_4^+ > HClO > H_2CO_3 > benzoic > formic > HBF_4

4. (5 pts) What is the pH of a 0.01 M solution of the weak base benzylamine ($C_6H_5CH_2NH_2$)? pK_a = 8. Circle the answer below and clearly show all your work.

Initial:	**0.01M**	**pure liq**	**0M**	**0M**
	$Base(aq) + H_2O(l) \rightleftharpoons BaseH^+(aq) + OH^-(aq)$			
@ Equib:	**0.01 – x**	**pure liq**	**x**	**x**

You first need to convert the pK_a into a K_a, then into a K_b for this *basic* equilibrium:

$K_b = \frac{K_w}{K_a} = \frac{1\times10^{-14}}{1\times10^{-8}} = 1\times10^{-6}$ now you can setup your equilib: $K_b = \frac{(x)(x)}{(0.01-x)} = 1\times10^{-6}$

assume that x << 0.01, $\frac{(x)(x)}{(0.01)} = 1\times10^{-6}$ or $x^2 = 1 \times 10^{-8}$, or $x = [OH^-] = 1 \times 10^{-4}$ pOH = 4.

BUT THIS IS NOT YOUR ANSWER, since I asked for the pH!! **pH = 14 – 4 = 10**

a) 4 b) 5 c) 9 d) 10 e) 13

5. (10 pts) Will $FeCl_3$ generate an acidic, neutral, or basic solution when dissolved in water. Clearly discuss your reasoning.

$FeCl_3$ is composed of a Fe^{3+} cation and three Cl^- anions. The Cl^- anions are "*do nothing*" or "*neutral*" (acid-base property, extremely poor conjugate base) so will not make a basic solution. The Fe^{3+} cation, on the other hand, is a Lewis acid and will react with water to release H^+: $Fe^{3+}(aq) + H_2O \rightleftharpoons [Fe(OH)]^{2+} + H^+(aq)$. This is called hydrolysis and most +2 or higher cations can interact with water in a similar fashion to release H^+ and make an acidic solution. Note that the $[Fe(OH)]^{2+}$ is not a typical hydroxide base as the Fe-OH bond is reasonably strong and will not dissociate OH^-.

6. (5 pts) What is the pH of a 1 M solution of KClO? See Table 1 for pK_a values. Clearly show all your work.

The first thing to realize is that this is a **basic salt** that will generate a ***basic*** solution! K^+ is a do-nothing cation, while ClO^- is an ***active*** anion that acts as a weak base in solution.

Initial:	**1M**	**pure liq**	**0M**	**0M**
	$Base^-(aq)$ + $H_2O(l)$	⇌	**$BaseH(aq)$ + $OH^-(aq)$**	
@ Equib:	**1 – x**	**pure liq**	**x**	**x**

Next you need to convert the pK_a value I've given you into a K_a, then K_b for the conjugate base in this basic equilibrium:

$K_b = \frac{K_w}{K_a} = \frac{1\times10^{-14}}{1\times10^{-8}} = 1\times10^{-6}$ now you can setup your equilib: $K_b = \frac{(x)(x)}{(1-x)} = 1\times10^{-6}$ assume that x << 1,

$\frac{(x)(x)}{(1)} = 1\times10^{-6}$ or $x^2 = 1\times10^{-6}$, or $x = [OH^-] = 1\times10^{-3}$ pOH = 3. pH = 14 – pOH = 11 so the **pH = 11**

7. (10 pts) Consider the following list of salts:

A) NH_4Cl **B)** KI **C)** CsF **D)** potassium benzoate

E) $MoCl_4$ **F)** BaI_2 **G)** $AlBr_3$ **H)** $LiNO_3$

I) $KClO_4$ **J)** NaClO

Which salts will generate an ***acidic*** solution? **A, E, G**

Which salts will generate a ***basic*** solution? **C, D, J**

Which salts will generate a ***neutral*** solution? **B, F, H, I**

8. (5 pts) Calculate the pK_b of the weak base phenylamine if a 1 M solution has a pH = 10.

Initial:	**1M**	**pure liq**	**0M**	**0M**
	Base(*aq*) + $H_2O(l)$	⇌	**$BaseH^+(aq)$ + $OH^-(aq)$**	
@ Equib:	**1 – x**	**pure liq**	**x**	**x**

$K_b = \frac{[baseH^+][OH^-]}{[base]}$

But I have given you the pH, so you know the OH^- concentration (and $[BaseH^+]$) at equilibrium. If pH = 10, pOH = 4 and the $[OH^-] = 1\times10^{-4}$ M. This is also small enough relative to 1 – x, that we can drop the x here and really make a very simple expression to solve for K_b:

$K_b = \frac{[baseH^+][OH^-]}{[base]} = \frac{(1\times10^{-4})(1\times10^{-4})}{1} = 1\times10^{-8}$ so **pK_b = 8**

9. (10 pts) What is the pH if 800 mL of 0.125 M KOH is added to 200 mL of 0.5 M sucoloic acid (a monoprotic acid)? $pK_a = 11$ (clearly show all your work)

The first thing to realize is that sucoloic acid is a weak acid and that we are titrating it with a strong base. If there are equal amounts of each, we will be making the salt of a weak acid, which is a weak base! So we may be generating a basic solution depending on the amount of acid and base reacting. Convert the # of mL and molarity of each into moles:

moles base = (800 mL OH^-)(0.125 M OH^-) = 100 mmoles OH^-
moles weak acid = (200 mL weak acid)(0.5 M weak acid) = 100 mmoles weak acid

So we have an equal amount of each. Don't forget that we are adding 800 mL to 200 mL to make 1000 mL total solution volume. This will generate a 100 mmoles/1000 mL = 0.1 M solution that will act as a weak base:

Initial:	**0.1M**	**pure liq**	**0M**	**0M**
	$Base^-(aq)$ + $H_2O(l)$	⇌	**$BaseH(aq)$ + $OH^-(aq)$**	
@ Equib:	**0.1 – x**	**pure liq**	**x**	**x**

Next you need to convert the pK_a value I've given you into a K_a, then K_b for the conjugate base in this basic equilibrium:

$K_b = \frac{K_w}{K_a} = \frac{1\times10^{-14}}{1\times10^{-11}} = 1\times10^{-3}$ now you can setup your equilib: $K_b = \frac{(x)(x)}{(0.1-x)} = 1\times10^{-3}$ assume that x << 1,

$\frac{(x)(x)}{(0.1)} = 1\times10^{-3}$ or $x^2 = 1\times10^{-4}$, or $x = [OH^-] = 1\times10^{-2}$ pOH = 2. pH = 14 – pOH = 12 so the **pH = 12**