

CHEM 1202 - Homework # 7 & 8

Name ANSWER KEY

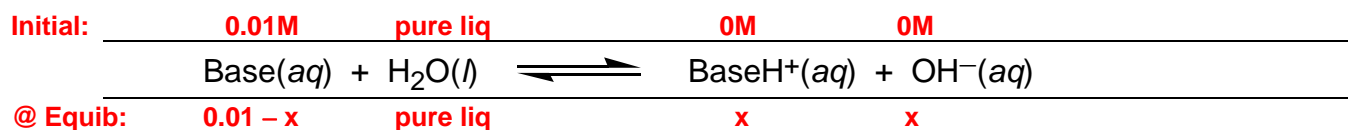
Acids & Bases 2

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	HF ₄	-9	formic	4
HClO	8	H ₂ CO 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) NH₄⁺ b) benzoic c) H₂CO₃ d) formic e) HBF₄
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) NH₄⁺ b) benzoic c) H₂CO₃ d) formic e) HBF₄
3. (5 pts) Order the acids in Table 1 from strongest to weakest. Circle the correct choice.
 a) benzoic > formic > H₂CO₃ > HBF₄ > NH₄⁺ > HClO
 b) HBF₄ > formic > benzoic > H₂CO₃ > HClO > NH₄⁺
 c) NH₄⁺ > benzoic > HBF₄ > formic > H₂CO₃ > HClO
 d) benzoic > HBF₄ > NH₄⁺ > formic > H₂CO₃ > HClO
 e) NH₄⁺ > HClO > H₂CO₃ > benzoic > formic > HBF₄
4. (5 pts) What is the pH of a 0.01 M solution of the weak base benzylamine (C₆H₅CH₂NH₂)? pK_a = 8.
 Circle the answer below and clearly show all your work.



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

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{1 \times 10^{-8}} = 1 \times 10^{-6} \quad \text{now you can setup your equilb: } K_b = \frac{(x)(x)}{(0.01 - x)} = 1 \times 10^{-6}$$

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

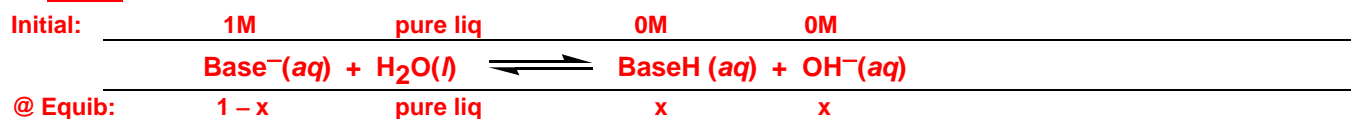
BUT THIS IS NOT YOUR ANSWER, since I asked for the pH!! **pH = 14 - 4 = 10**

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

5. (10 pts) Will FeCl₃ generate an acidic, neutral, or basic solution when dissolved in water. Clearly discuss your reasoning.

FeCl₃ is composed of a Fe³⁺ 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 Fe³⁺ cation, on the other hand, is a Lewis acid and will react with water to release H⁺: Fe³⁺(aq) + H₂O ⇌ [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.



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:

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{1 \times 10^{-8}} = 1 \times 10^{-6} \quad \text{now you can setup your equilib: } K_b = \frac{(x)(x)}{(1-x)} = 1 \times 10^{-6} \quad \text{assume that } x \ll 1,$$

$$\frac{(x)(x)}{(1)} = 1 \times 10^{-6} \quad \text{or } x^2 = 1 \times 10^{-6}, \quad \text{or } x = [OH^-] = 1 \times 10^{-3} \quad pOH = 3. \quad pH = 14 - pOH = 11 \quad \text{so the } \boxed{pH = 11}$$

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

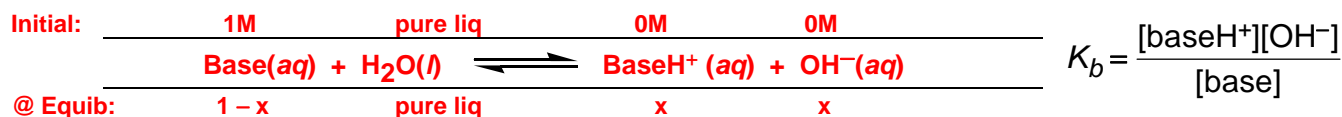
- | | | | |
|-------------|------------|-------------|-----------------------|
| A) NH_4Cl | B) KI | C) CsF | D) potassium benzoate |
| E) $MoCl_4$ | F) BaI_2 | G) $AlBr_3$ | H) $LiNO_3$ |
| I) $KClO_4$ | J) $NaClO$ | | |

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.



But I have given you the pH, so you know the OH^- concentration (and $[BaseH^+]$) at equilibrium. If pH = 10, pOH = 4 and the $[OH^-] = 1 \times 10^{-4}$ 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 :

$$K_b = \frac{[baseH^+][OH^-]}{[base]} = \frac{(1 \times 10^{-4})(1 \times 10^{-4})}{1} = 1 \times 10^{-8} \quad \text{so } \boxed{pK_b = 8}$$

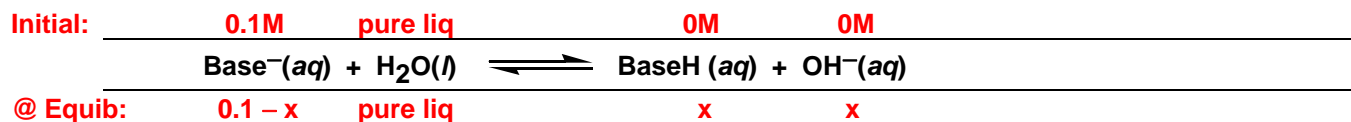
9. (10 pts) What is the pH if 800 mL of 0.125 M KOH is added to 200 mL of 0.5 M succinic acid (a monoprotic acid)? $pK_a = 11$ (clearly show all your work)

The first thing to realize is that succinic 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:

$$\# \text{ moles base} = (800 \text{ mL } OH^-)(0.125 \text{ M } OH^-) = 100 \text{ mmoles } OH^-$$

$$\# \text{ moles weak acid} = (200 \text{ mL weak acid})(0.5 \text{ M weak acid}) = 100 \text{ mmoles weak acid}$$

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:



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:

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{1 \times 10^{-11}} = 1 \times 10^{-3} \quad \text{now you can setup your equilib: } K_b = \frac{(x)(x)}{(0.1-x)} = 1 \times 10^{-3} \quad \text{assume that } x \ll 1,$$

$$\frac{(x)(x)}{(0.1)} = 1 \times 10^{-3} \quad \text{or } x^2 = 1 \times 10^{-4}, \quad \text{or } x = [OH^-] = 1 \times 10^{-2} \quad pOH = 2. \quad pH = 14 - pOH = 12 \quad \text{so the } \boxed{pH = 12}$$