CHM 1046 - JOENS WORKSHEET #8 Due date: M,W,F Monday,October 17
T,R Tuesday, October 18

* While I prefer you turn in a hard copy of the worksheet, I will accept scanned copies sent to my email address, joensj@fiu.edu

Section: (circle one) M,W,F Tu,Tr

Exam 2 is Friday, October 21 (for the M,W,F class) and Thursday, October 20 (for the T,R class) in class. It will cover material from Chapters 15 and 16 of Burdge.

For problems involving calculations you must show your work for credit. Unless otherwise stated, you may assume T = 25.0 °C.

1) The pH of a 2.4 x 10^{-3} M solution of a weak monoprotic acid HA is pH = 5.38 Based on this, find the percent dissociation for the weak acid.

The reaction of a weak monoprotic acid in water may be written as

$$HA(aq) + H_2O(\ell) \leftrightarrows H_3O^+(aq) + A^-(aq)$$

By definition, the percent dissociation is $$\% \text{ dissoc.} = \frac{\text{concentration of dissociated acid}}{\text{initial concentration of acid}} \times 100\%$$

The initial concentration of acid is 2.4 x 10^{-3} M. By looking at the balanced reaction, we can see that the concentration of dissociated acid will be equal to $[H_3O^+]$ (since when an HA molecule dissociates it forms an H_3O^+ and an A^-).

$$[H_3O^+] = 10^{-pH} = 10^{-5.38} = 4.17 \times 10^{-6} \text{ M}$$

So $$\% \text{ dissoc.} = \frac{4.17 \times 10^{-6}}{2.4 \times 10^{-3}} \times 100\% = 0.17\ \%$$

2) Consider the following list of weak acids

C_6H_5COOH	$K_a = 6.5 \times 10^{-5}$
C_6H_5OH	$K_a = 1.3 \times 10^{-10}$
HCOOH	$K_a = 1.7 \times 10^{-4}$

For each of the following questions circle the correct answer.

a) The strongest acid

C_6H_5COOH C_6H_5OH **HCOOH**

b) The strongest base

$C_6H_5COO^-$ **$C_6H_5O^-$** $HCOO^-$

The strongest acid will be the one with the largest value for K_a. Since the weaker the acid the stronger the conjugate base, the strongest conjugate base will correspond to the weakest acid.

3) For each of the following pairs of substances circle the one that is the stronger acid. You should be able to do this without looking up values for K_a for the acids.

a) **$\underline{H_2S}$** H_2O

b) HIO_2 **$\underline{HIO_3}$**

c) HIO_2 **$\underline{HBrO_2}$**

For binary acids with nonmetals in the same group, acid strength increases from top to bottom. For a ternary acid with the same nonmetal, the acid with the larger number of oxygen atoms will be the stronger acid. For ternary acids with nonmetals in the same group and the same number of oxygens, acid strength increases from bottom to top.

4) Sodium acetate ($NaCH_3COO$, MW = 82.03 g/mol) is a soluble ionic compound. It is the salt formed from the reaction of a weak acid with a strong base.

a) Give the dissolution reaction that takes place when $NaCH_3COO(s)$ is added to water.

$NaCH_3COO(s) \rightarrow Na^+(aq) + CH_3COO^-(aq)$

Since $NaCH_3COO(s)$ is a soluble ionic compound, the reaction will go completely from left to right.

b) CH_3COO^- is the conjugate base of a weak acid, and so acts as Bronsed base when added to water. Give the reaction that occurs when $CH_3COO^-(aq)$ is added to water.

This will be a Bronsted acid-base reaction

$CH_3COO^-(aq) + H_2O(\ell) \leftrightarrows CH_3COOH(aq) + OH^-(aq)$ $\qquad K_b = \dfrac{[CH_3COOH]\,[OH^-]}{[CH_3COO^-]}$

c) What is the pH of a 0.073 M aqueous solution of sodium acetate? Note that for acetic acid (CH_3COOH), $K_a = 1.8 \times 10^{-5}$

Since $K_a = 1.8 \times 10^{-5}$ for CH_3COOH, the value for K_b for the conjugate base (CH_3COO^-) can be found from the expression

$K_a K_b = 1.0 \times 10^{-14}$ $\qquad K_b = \dfrac{1.0 \times 10^{-14}}{K_a} = \dfrac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$

So we need an ICE table

	Initial	Change	Equilibrium
CH_3COOH	0	x	x
OH^-	0	x	x
CH_3COO^-	0.073	–x	0.073 – x

So $\dfrac{(x)\,(x)}{(0.073 - x)} = 5.56 \times 10^{-10}$ Assume $x << 0.073$

Then $\dfrac{x^2}{0.073} = 5.56 \times 10^{-10}$ $\qquad x^2 = (5.56 \times 10^{-10})\,(0.073) = 4.06 \times 10^{-11}$

$x = (4.06 \times 10^{-11})^{1/2} = 6.37 \times 10^{-6}$

So $[OH^-] = x = 6.37 \times 10^{-6}$ M

$pOH = -\log_{10}[OH^-] = -\log_{10}(6.37 \times 10^{-6}) = 5.20$

$pH + pOH = 14.00$, and so $pH = 14.00 - pOH = 14.00 - 5.20 = 8.80$

This is not an easy problem, mainly because it requires you to pull together a lot of different ideas. Having said that, people sometimes need to know how to find the value for pH for salts that contain conjugate acids or conjugate bases.