CHM 1046 - JOENS WORKSHEET #9 Due date: Monday, October 22^{nd}

WORKSHEETS ARE DUE AT THE BEGINNING OF CLASS ON THE DATE GIVEN ON THE WORKSHEET. LATE WORKSHEETS WILL NOT BE ACCEPTED.

NAME ______________________________ Panther ID ____________________________

For problems involving calculations you must show your work for credit.

1) It is sometimes possible to rank binary or ternary weak acids in terms of their relative strength.

a) What are the two factors that determine the relative strength of a binary or ternary acid? How do these factors influence acid strength?

The two factors that determine relative acid strength are the strength of the bond the donatable hydrogen makes with the rest of the molecule and the electronegativity of the atoms in the conjugate base.

For the donatable hydrogen atom we have an H - A single bond. The weaker the bond the easier it is to break, and so the more easily an H^+ ion will form. So the weaker the bond the stronger the acid.

For the conjugate base, the larger the value for the electronegativity of the atoms present the more easily they will be able to accommodate the extra electron charge, and so the more stable (and more easily formed) the conjugate base. This can occur either due to changes in electronegativity of the third atom in a ternary acid (when the number of oxygen atoms is the same), or in changes in the number of oxygen atoms associated with a particular third nonmetal ternary acid.

Using these two ideas we can develop general rules for the relative acid strengths of some binary and ternary acids.

b) Put the following three acids in order from strongest to weakest (strongest first): HCl, HI, H_2S

Based on our general trends we may say HI > HCl

HCl > H_2S

So the order is HI > HCl > H_2S

Since we know that HCl and HI are strong acids, and H_2S is not, this is consistent with our knowledge of the common strong acids.

2) Which of these reactions goes essentially to completion?

a) reaction of a strong acid with a strong base
b) reaction of a strong acid with a weak base
c) reaction of a weak acid with a weak base
d) both a and b
e) both a and b and c _____D______

For an acid-base (neutralization) reaction the reaction will essentially go to completion as long as one of the reactants is strong. In a both reactants are strong and in b the acid is strong, and so both of those reactions will go to completion. But in c the acid and base are both weak, and so the reaction will not go to completion.

3) Consider 500.0 mL of a 0.200 M aqueous solution of hypobromous acid at T = 25. °C.

a) What is the pH of the above solution? Note that for HOBr, $K_a = 2.5 \times 10^{-9}$.

This is just a weak acid problem.

The reaction is $HOBr(aq) + H_2O(\ell) \leftrightarrows H_3O^+(aq) + OBr^-(aq)$

$$K_a = \frac{[H_3O^+][OBr^-]}{[HOBr]} = 2.5 \times 10^{-9}$$

For our ICE table

	Initial	Change	Equilibrium
H_3O^+	0	x	x
OBr^-	0	x	x
HOBr	0.200	- x	0.200 - x

$$\frac{(x)(x)}{(0.200 - x)} = 2.5 \times 10^{-9}$$ If we assume x < 0.200, then we get

$$\frac{x^2}{0.200} = 2.5 \times 10^{-9} \qquad x^2 = (2.5 \times 10^{-9})(0.200) = 5.0 \times 10^{-10} \qquad x = (5.0 \times 10^{-10})^{1/2}$$

$$= 2.2 \times 10^{-5}$$

So our assumption that x << 0.200 was good.

So $[H_3O^+] = x = 2.2 \times 10^{-5}$ M $\qquad$ $pH = -\log_{10}(2.2 \times 10^{-5}) = 4.65$

b) 0.868 g of solid sodium hypobromite (NaOBr) is added to the above solution. What is the new pH of the solution after the addition of NaOBr?

MW(NaOBr) = 118.9 g/mol

$$\text{moles NaOBr} = 0.868 \text{ g} \; \frac{1 \text{ mol}}{118.9 \text{ g}} = 0.00730 \text{ mol NaOBr}$$

NaOBr is a soluble ionic compound and a strong electrolyte, and so

$NaOBr(s) \rightarrow Na^+(aq) + OBr^-(aq)$ $\qquad$ moles OBr^- = moles NaOBr = 0.00730 moles

$$[OBr^-] = \frac{0.00730 \text{ mol}}{0.5000 \text{ L}} = 0.0146 \text{ mol/L}$$

Since we have an appreciable amount of both a weak acid and its conjugate base present, we can use the Henderson equation to find pH.

$$pH = pK_a + \log_{10}\{[\text{base}]/[\text{acid}]\} = -\log_{10}(2.5 \times 10^{-9}) + \log_{10}\{(0.0146)/(0.200)\}$$

$$= 8.60 + (-1.14) = 7.46$$

Note we could get this same result using an ICE table, but using the Henderson equation is faster. As long as we know how to use the equation correctly this is the preferred method.