CHM 1046 - JOENS WORKSHEET #10 Due date: Monday, October 29th

The third hour exam is Friday, November 2nd. It will cover Chapter 16, sections 16.2, 16.6 to 16.12, and Chapter 17, sections 17.1 to 17.3
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) An aqueous solution at T = 25. °C contains 0.0400 M hypochlorous acid (HOCl, $K_a = 3.5 \times 10^{-8}$) and 0.0600 M sodium hypochlorite (NaOCl).

a) What is the initial pH of the above solution.

The solution contains a significant amount of a weak acid (HOCl) and conjugate base (OCl^-), and so is a buffer solution. We may find the pH from the Henderson equation.

$$pH = pK_a + \log_{10}\{[\text{base}]/[\text{acid}]\}$$

$$= -\log_{10}(3.5 \times 10^{-8}) + \log_{10}\{(0.0600)/(0.0400)\}$$

$$= 7.46 + 0.18 = 7.64$$

b) Give the correctly balanced reaction that takes place when a small amount of hydrochloric acid (HCl) is added to the above solution.

HCl is a strong acid, and so it will want to react with any base that is present. The reaction therefore is

$$HCl(aq) + OCl^-(aq) \rightarrow HOCl(aq) + Cl^-(aq)$$

The strong acid (HCl) is converted into weak acid (HOCl), minimizing the effect of the addition on pH.

2) Consider the titration of a weak aqueous base solution with hydrochloric acid (HCl), a strong acid. The pH at the equivalence point of the titration will be (circle the correct answer).

larger than 7.0 approximately 7.0 **smaller than 7.0**

3) 0.984 g of a weak monoprotic acid are dissolved in water. The weak acid is then titrated with a 0.2145 M solution of potassium hydroxide (KOH), a strong soluble base. After the addition of 33.96 mL of the KOH solution the equivalence point of the titration is reached.

What is the molecular weight of the weak monoprotic acid?

$$MW = \frac{\text{mass}}{\text{moles}}$$

Since the acid is monoprotic, the reaction is

$HA(aq) + KOH(aq) \rightarrow KA(aq) + H_2O(\ell)$

At the equivalence point moles HA = moles KOH

But moles KOH = $0.03396\ \text{L}\ \dfrac{0.2145\ \text{mol}}{\text{L soln}} = 7.28 \times 10^{-3}$ moles

So $MW = \dfrac{0.984\ \text{g}}{7.28 \times 10^{-3}\ \text{mol}} = 135.\ \text{g/mol}$

4) A solution is formed by adding 25.00 mL of a 0.1672 M solution of sodium hydroxide (NaOH), a strong soluble base, to 50.00 mL of a 0.1142 M solution of iodoacetic acid (CH_2ICOOH, $K_a = 7.6 \times 10^{-4}$), a weak monoprotic acid.

a) Is the solution that forms a buffer solution? Why or why not?

initial moles NaOH = (0.02500 L) (0.1672 mol/L) = 4.18×10^{-3} moles

initial moles CH_2ICOOH = (0.05000 L) (0.1142 mol/L) = 5.71×10^{-3} moles

The neutralization reaction is $CH_2ICOOH(aq) + NaOH(aq) \rightarrow Na^+(aq) + CH_2ICOO^-(aq) + H_2O(\ell)$

After neutralization, moles CH_2ICOOH = (5.71×10^{-3} moles) – (4.18×10^{-3} moles) = 1.53×10^{-3} moles

moles CH_2ICOO^- = 4.18×10^{-3} moles

There are significant concentrations of weak acid and conjugate base present, so the solution is a buffer.

b) What is the pH of the solution formed by the addition of the NaOH and iodoacetic acid solutions?

Since the solution is a buffer, we can use the Henderson equation

$$pH = pKa + \log10\{[\text{base}]/[\text{acid}]\} = -\log_{10}(7.6 \times 10^{-4}) + \log_{10}\{(4.18 \times 10^{-3})/(1.53 \times 10^{-3})\}$$

$$= 3.12 + 0.44 = 3.56$$

While this problem can be done using the "ICE' method, the Henderson equation is a faster and less error prone way to proceed.