Buffers and Acid-Base Titrations Worksheet (#3)

1. Write ionic equations to how each pair of compounds can serve as a buffer pair.

   a. H_2CO_3 and $NaHCO_3$ (the "carbonate" buffer in blood)

   a. NaH_2PO_4 and Na_2HPO_4 (the "phosphate" buffer inside body cells)

   b. NH_4Cl and NH_3

2. Which buffer would be able to hold a steady pH on the addition of strong acid, buffer 1 or buffer 2? Explain.

   Buffer 1: a solution containing 0.10 M NH_4Cl and 1 M NH_3

   Buffer 2: a solution containing 1 M NH_4Cl and 0.10 M NH_3

3. How many grams of sodium acetate, $NaC_2H_3O_2$, would have to be added to 1 L of 0.15 M acetic acid (pKa=4.74) to make the solution a buffer for pH 5.00? (Hint: rearrange HH: pH = pKa + log [base] – log [acid] to solve for log [base] = pH -pKa + log [acid] then take antilog

4. What ratio of molar concentration of NH_4Cl and NH_3 would buffer a solution at pH 9.25?

5. To study the effect of a weakly acidic medium on the rate of corrosion of a metal, a chemist prepared a buffer solution by making it 0.11 M $NaC_2H_3O_2$ and also 0.090 M $HC_2H_3O_2$ (pKa = 4.74). What is the pH of this solution?

Titrations

1. How many milliliters of 0.100 M HCl are required to neutralize 25.0 mL of 0.100 M $Ba(OH)_2$?

2. Exactly 50.0 mL of HOCl solution of unknown concentration was titrated with 0.100 M NaOH. An end point was reached when 38.5 mL of the base was added. Calculate the molar concentration of the HOCl solution.

3. What can make the titrated solution at the equivalence point in an acid-base titration have a pH not equal to 7.00. How does this possibility affect the choice of an indicator?

4. When 50.0 mL of 0.10 formic acid (Ka = 1.8×10^{-4}) is titrated with 50.0 mL of 0.10 M NaOH, what is the pH at the equivalence point? (Take into account the change in Volume)