Pre-lab assignment/ Titrations 1

Name ______________________________ Date _______________ Lab section _______________

1. One of the objectives in today's experiment is to standardize a NaOH solution. What does this mean? Why is this necessary?

2. A NaOH solution is found to have a concentration of 0.2107 M. If 15.13 mL of the NaOH solution is required to titrate 0.6503 g of an unknown acid dissolved in 75 mL of water, what is the molar mass of the acid? Assume the acid and NaOH react in a 1:1 mole ratio.

3. The unknown acid in question 2 is one of the following. From your answer to question 2, which is the correct acid? (Circle your answer.)

CH_3COOH HBr NaCl $C_8H_5O_4K$

Titrations: Standardization and Molar Mass Determination

Acid-Base Titrations: Determining the Concentration of a NaOH solution

Now that you are more proficient at measuring volumes with a buret, you can use that skill to gain some chemically useful information. First we need to say a word about the chemical reaction that has occurred in our previous titrations. Remember that an ***acid*** is a proton donor and a ***base*** is a proton acceptor. When a proton is transferred from acid to base, a ***neutralization*** reaction occurs:

$$H_3O^+ + OH^- \rightarrow 2\ H_2O$$

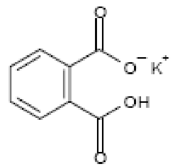
One mole of $\mathbf{H_3O^+}$ ***reacts with one mole of*** $\mathbf{OH^-}$ to form two moles of water. The phenolphthalein indicator changes color at the first slightest excess of base added. Phenolphthalein is a suitable indicator to use when we are titrating ***strong acids*** like HCl with a ***strong base*** like NaOH.
When we are dealing with solutions we need to keep in mind that the major part of those solutions is simply water. Only a relatively small part of the solutions (the solutes) is actually involved in the chemical reaction.

Since solutions are mixtures, it is important to indicate the ratio of each component (solute and solvent) present, i.e., its concentration. The concentration term which we will use here is ***molarity***, M, defined as the number of moles of solute per liter of solution. From our definition of concentration, we can see that if we multiply molarity by volume (in liters) we get moles of solute actually used in the reaction.

The first objective of this experiment is to **standardize** a solution of NaOH**,** or to determine the concentration of the solution with great accuracy. We will do that by reacting NaOH with a weighed sample of a ***monoprotic acid***. The term ***monoprotic*** means that one mole of the acid produces one mole of H_3O^+. We need to weigh samples of the acid precisely and that process of weighing will be carried out on an analytical balance capable of weighing to 0.0001 g.

Part I. Standardization of NaOH

NaOH solutions used in acid-base titrations are usually standardized against solid potassium acid phthalate (or **KHP** for short), KHP is a monoprotic acid (MW = 204.2). KHP can be obtained in a high degree of purity, is not appreciably hydroscopic (stays dry), and is reasonably soluble in water. A weighed sample is dissolved in water and is titrated to a **phenolphthalein** end point.



Potassium hydrogen phthalate: "KHP"

The stoichiometric equation that describes the chemical mixture at neutralization when the ***permanent faint pink*** color appears is:

moles acid = moles base

We can convert this statement into the equation:

$$\frac{grams\ acid}{molar\ mass\ of\ acid} = (Volume\ of\ base) \times (Molarity\ of\ base)$$

In this equation there is only one unknown; the concentration of NaOH. You will calculate that concentration.

Experimental Procedure

1. Weigh three samples of potassium acid phthalate, approximately 0.4 g in size, weighed to the nearest 0.1 mg. Dissolve each solid sample in approximately 75 mL of water. Add indicator.
2. Titrate each dissolved KHP sample with NaOH solution to a phenolphthalein end point. The NaOH solution is approximately 0.1M.
3. Calculate your experimentally determined NaOH molarity, average molarity, standard deviation, and relative average deviation.
4. If your relative standard deviation exceeds 0.7%, do one additional trial. Recalculate your precision. Repeat titrations until you have three with %rsd < 0.7%.

Part II. Determining the Molar Mass of a Monoprotic Acid

Again we will use the principle of neutralization and the fact that at this point:

moles acid = moles base

We can convert this statement for this experiment into the equation:

$$\frac{grams\ acid}{molar\ mass\ of\ acid} = (Volume\ of\ base) \times (Molarity\ of\ base)$$

Since you have already solved for the molarity of the NaOH solution in Part I, you will now use this relationship to find the **molar mass of an unknown acid.**

Experimental Procedure

The procedure is exactly the same as for the standardization, except that the unknown in the algebraic equation is the molar mass.

1. Weigh to 0.1 mg a sample of your unknown acid for your initial trial. Your instructor will tell you what size sample to aim for, depending on identity of your unknown acid. Dissolve in approximately 75 mL of deionized water. Some unknowns need to be heated slightly to dissolve the solid in water.

2. Add a few drops of phenolphthalein indicator.

3. Determine the volume of standard NaOH solution required to reach the endpoint for this first trial. (You would like a mass of unknown acid which requires approximately 20 mL of standard NaOH base to reach the endpoint.)

4. After you complete the first trial, adjust the mass of unknown acid, and carry out three additional trials. Your **rsd** should be less than or equal to 0.7 %.

In your lab notebook write-up, report your final results. Include your average **molarity ± std dev** and **rsd** for the NaOH solution you standardized. Your final result should also include the identity of the unknown NaOH solution you standardized. In addition, ***calculate the molar mass of your unknown acid***. The molar masses of the available acids lie between 80 and 250 and the acids are all monoprotic. Your conclusion should include your average molar mass ± sd, and your rsd.