Title: Analysis of Vinegar

Introduction:

Ordinary "white" vinegar is an aqueous (water) solution of approx 5% acetic acid with water. In order to ensure that the acidity is at the desired level, periodic routine analyses are run. A common method for such analyses is a *titration*, in which a strong base of known concentration is used to determine the concentration of the acid by allowing the solutes of the two solutions to react with each other. In a titration, a titrant base of known concentration is added at a controlled rate to a known amount of the acid to be analyzed. Addition continues until the neutralization reaction is complete.

An *indicator* is often used to determine when all of the acid has reacted with the base. Indicators signal that a reaction is complete by changing color. *Phenolphthalein,* the indicator that you will use in this experiment, is colorless in acidic or neutral solutions, but turns bright pink with the slightest excess of base. The first drop of base that causes the color to persist signals the end of the titration. The equation for the reaction between the sodium hydroxide and the acetic acid in the vinegar is:

$$HC_2H_3O_2\ (aq) + NaOH\ (aq) \rightarrow NaC_2H_3O_2\ (aq) + H_2O\ (l)$$

The hydroxide ion of the base (NaOH) reacts with a hydrogen ion from the acid ($HC_2H_3O_2$) to form water. These reactions are called neutralization reactions, because the acid and the base neutralize each other, producing water.

In this lab, you will prepare a sodium hydroxide solution of known concentration, then use that solution to analyze the acid content of vinegar. First you will determine the molar concentration (molarity) of the sodium hydroxide.

Because in the experiment both the mass and the concentration of the sodium hydroxide titrant are known, the number of moles of NaOH that reacts can be calculated. As the equation shows, acetic acid and sodium hydroxide react in a 1:1 mol ratio, so you can also determine the number of moles of acetic acid present in the sample which can then be converted to mass in grams. You will carry out at least 3 trials for the analysis. The amounts of the two solutions used in each of the titrations will be determined by reading the burette.

Prelaboratory Questions

1. Write the equation for the reaction between acetic acid and NaOH. Use the structural formula for acetic acid found in Chapter 16. In the formulas of the reactants, circle the atoms that form water.

2. What is the purpose of the indicator? How does it tell you when a titration is complete?

Materials

Apparatus

Balance
. Erlenmeyer flask, 10-mL (or small beakers)
burette, labeled (A) Vinegar, B (NaOH)
. Safety goggles
. Sodium hydroxide, solid pellets
. Phenolphthalein indicator
. 4 mL vinegar

Procedure

Titration of Vinegar.

1. Rince and clean a small beaker. Transfer about 1 mL of the vinegar solution to the flask and record the exact amount. The amount of vinegar used should be between 0.8 and 1.2 mL; if it is less than 0.8 mL, add a bit more. Record the exact amount of vinegar.
2. Add 1 drop of phenolphthalein indicator from the phenolphthalein pipet (labeled "PHTH") to the contents of the flask and swirl gently.
3. Record the beginning amount of the NaOH burette and add NaOH from the burette into the beaker with the vinegar swirling, add your NaOH solution a little at a time, especially after about 9 mL have been added. Once you see pink add a few drops at a time until you get a magenta color that does not fade with mixing and that lasts at least 20 seconds. The lighter the pink color, the better. Record the ending value of the NaOH burette.
4. Repeat the titration two more times, recording the beginning and ending values of the burettes before and after each trial. You will probably need to refill one or both from time to time. Try to plan so that you don't have to refill in the middle of a trial. **Note:** Use a clean flask for each titration; if necessary, wash flasks between trials, then rinse with distilled water. You need not dry the flasks.

Data Table #1 – The NaOH set-up

1. Reading of Acid before release of 1mL	
2. Reading After release of 1 mL acid	
3. Reading before the release of Base	
4. Reading of base after you see pink.	
5. Volume of Acid used	
6. Volume of Base used	

Analysis and Conclusions

Complete the **Analysis and Conclusions** section for this experiment. Show samples of all calculations.

1. Calculate the molarity (mol/L) of sodium hydroxide used in the experiment. You will need to receive this information from your instructor. How many grams of NaOH were added to mL of H_2O. Solve for molarity by converting g → mol of NaOH and converting mL → L of H_2O.

2. From your data, using the molarity of the NaOH solution used in each titration (from question #1) and the volume used you can calculate the molarity of the acetic acid present for each of your titrations. Show your work for trial 1; enter the results for all titrations in a Data Table. $M_a \times V_a = M_b \times V_b$

At the equivalence point in a titration the mols of acid equal the mols of base. In a complete titration with a monoprotic acid, where the molar ratios of acid to base are 1:1, the following formula would hold.

$$M_B \times V_B = M_A \times V_A$$

where:

m_B : molarity of base (NaOH)

V_B: volume of base used to reach end point in Liters – remember 1g is = to 1 mL for our solution

M_A: molarity of acid ($HC_2H_3O_2$)

V_A: volume of acid used in Liters

3. Determine the mass of acetic acid present in each of your titration samples. Here you will use the Molarity you solved for in #2 and solve for mol using M = mol/L . You have the M from #2 and the L in your data table. 1 g = 1 mL and 1 L = 1000 mL. Then convert for mols to grams using the molar mass of Acetic Acid (($HC_2H_3O_2$). Find the percent of acetic acid in each vinegar sample, by dividing the mass of acetic acid present by the mass of vinegar used. Convert the decimal fraction to a percent.

4. Calculate an average value for the mass percent of acetic acid in the vinegar you analyzed. Base your average on the two trials that show the closest agreement; omit the trial that deviates most greatly from the others. Remember the bottle said 5% acetic acid.

5. Calculate your percent error. You theoretical is 5% . Remember % error is

$$\frac{\text{Actual} - \text{Theoretical}}{\text{Theoretical}} \times 100$$

Conclusion:

Remember this section is where you report your results this is the MEAT of your write up and you need to be thorough!

1st draw conclusions. Give a valid conclusion based on the correct interpretation of your results and explain your results reflecting back on the target (what did you do?).

2nd evaluate procedure(s) and results including limitations, weaknesses or errors. (what worked well or what did not work)

3rd Identify weaknesses and state realistic suggestions to improve the investigation. (if you were to change this lab…. what would you change so it works better?)