The titration of Acetic Acid in Vinegar

In this laboratory exercise we will determine the percentage Acetic Acid (CH_3CO_2H) in Vinegar.

We will do this by Titrating the Acetic Acid present with a Strong Base; Sodium Hydroxide (NaOH). The Endpoint of the Titration will be detected using a Phenolphthalein indicator; an acid-base indicator that changes color from clear to pink in going from its acidic form to its basic form.

Acetic Acid (fr. Latin acetum for vinegar) is the main component of Vinegar. It is a carbon based compound with a single ionizable proton, making it an organic acid of the larger class of organic acids called Carboxylic Acids; organic compounds with a –COOH functional moiety (parts that can be split).

We will determine the **Acetic Acid content of a commercially prepared White Vinegar** solution using a type of Volumetric Analysis called Titration.

As noted, a volumetric analysis involves measuring the volume of a solution of known concentration that is needed to completely react with an analyte. The Titration is performed by slowly adding the titrant to the analyte solution via the burette until an Endpoint is reached. The Endpoint is represented by some distinct physical change in the analyte solution; typically an Indicator color change. If the indicator is chosen well, the Endpoint will represent the Equivalence Point of the Titration Reaction; the point at which the added amount of titrant is stoichiometrically equivalent to the amount of analyte. By knowing the concentration and volume of the titratnt used, the number of moles titrant can be determined. The reaction stoichiometry then allows us to determine the amount of analyte present.

Example:

Suppose we Titrate a solution of Sulfuric Acid (H_2SO_4) with a Standard 0.1054 M Solution of Sodium Hydroxide. The titration reaction is:

$$H_2SO_4(aq) + 2\ NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2\ H_2O$$

Further, suppose 12.56 mL of titrant is required to reach the Endpoint. Then,

$$\#\ \text{moles titrant used} = (0.1054\ M) \times (0.01256\ L) = 0.001324\ \text{mole NaOH}$$

and:

$$\#\ \text{mole } H_2SO_4 = \frac{1\ \text{mole } H_2SO_4}{2\ \text{moles NaOH}} \times 0.001324\ \text{mole NaOH} = 0.0006619\ \text{mole } H_2SO_4$$

Or, in grams:

$$\#\ \text{grams } H_2SO_4 = 0.0006619 \times (98.08\ \text{g/mole}) = 0.06492\ \text{g}$$

In our case, the Analyte is the Acetic Acid in the Vinegar and the Titrant is a dilute solution of the strong base Sodium Hydroxide. The Endpoint of the titration will be detected by observing the color change for vinegar.

LAB PREP:

- The following step has to be as accurate as possible since it is your standard reaction!

Use a volumetric flask to prepare the NaOH standard solution for accuracy. Add some distilled water to the volumetric flask. Mass out the NaOH in a plastic weighing dish. Carefully rinse the chemical into the flask with distilled water in a rinse bottle to get all the chemical into the flask. Swirl the flask until the chemical is completely dissolved, then add distilled water until the bottom of the meniscus is exactly on the etched mark. Stopper the flask and invert at least 25 times to mix. Be sure you make enough solution to do the complete lab.

To prepare a **0.600 M solution** of sodium hydroxide:

For 1000mL (1 liter) use **24.00 grams** NaOH
For 250mL use **6.00 grams** NaOH
For 100mL use **2.40 grams** NaOH

Procedure:

1. Make your own standardized NaOH base for 1st set of trials and repeat the following procedures using the Standardized NaOH solution for 2nd set of trials.

NOTE ON BURET TECHNIQUE: When filling a buret, check to make sure the stopcock is closed. Hold the buret in your hand with a paper towel wrapped around it to catch any spill, and fill slowly. Wipe the buret up to the top with the paper towel.

2. Pour directly from the beaker into the buret about **5 mL** of the NaOH solution. Rinse the walls of the buret thoroughly with the NaOH, allow it to drain through the stopcock, and discard it. Rinse the buret **two more times** in a similar fashion, using a new **5 mL** portion of the NaOH solution each time. Discard all rinse solutions. Fill the buret with the standard NaOH solution above the zero mark. and place the buret in one side of the double buret clamp. Withdraw enough solution from the buret to remove the air from the jet tip and bring the liquid level into the graduated region of the buret. Your starting volume does not have to be zero. Discard what you withdrew. **Label the buret**.

3. Follow steps 1 and 2 for the vinegar, and place it on the other side of the double buret clamp. **Be sure to label the buret**.

Titrations:

Make all buret readings to **0.01 mL**

4. Read the buret with the vinegar as the initial volume and record it in the **DATA TABLE** for trial 1, as the initial volume for vinegar. Obtain approximately **10 mL** of vinegar in the flask. Read the buret and record it in the **DATA TABLE** for trial 1, as the final volume for vinegar.

5. Add 10 or 15 mL of distilled water to the flask to increase the volume and make reading the equivalence point easier to read. Add **1 or 2 drops** of phenolphthalein solution to the flask to serve as an indicator.

6. Read the initial volume of the NaOH solution in the buret and record it in the **DATA TABLE** for trial 1, as the initial volume for NaOH.

7. Place the flask with the vinegar sample, on a sheet of white paper, under the buret. Now begin the titration by adding the hydroxide solution slowly, while swirling the flask. A pink color will appear in the center, but will go away as you swirl. When the pink color begins to linger, add the NaOH one drop at a time, swirling between each drop. When one drop is added, and the **faint pink color** does not disappear, you have reached the equivalence point (the point where all the acid is just neutralized by the base, but no extra base is added). Read the volume in the buret, and record it in the **DATA TABLE** for trial 1, as the final volume for NaOH.

8. Discard the liquid in the flask and rinse the flask with 4 or 5 rinses of distilled water to be sure it is clean. Repeat steps 4-7 for trials 2 and 3.

Record volumes for the burets in the **DATA TABLE** in the appropriate places.

NOTE: You do not have to refill the buret, if you are sure it will not go past the 50 mL mark in a titration, just read the initial and final volumes. If you do go below the graduations, you will have to re-do that titration, as you will not be able to read a final volume.

Data Table:

1st set DATA TABLE with student made NaOH solution (0.6M)				
	Buret Readings (mL)			
	Vinegar		**NaOH**	
Trial	**initial**	**final**	**initial**	**final**
1				
2				
3				

2nd set DATA TABLE with Standardized NaOH (0.6M)				
	Buret Readings (mL)			
	Vinegar		**NaOH**	
Trial	**initial**	**final**	**initial**	**final**
1				
2				
3				

Calculations:

Set 1 CALCULATIONS TABLE				
Trial	**Volume vinegar (mL)**	**Volume NaOH (mL)**	**Moles NaOH (mole)**	**Moles of acetic acid in vinegar (mole)**
1				
2				
3				
Set 2 CALCULATIONS TABLE				
Trial	**Volume vinegar (mL)**	**Volume NaOH (mL)**	**Moles NaOH (mole)**	**Moles of acetic acid in vinegar (mole)**
1				
2				
3				

1. Calculate the volumes of vinegar and NaOH used for each of the three trials. Record in the **CALCULATIONS TABLE**.

This is simply: **initial volume - final volume**

2. Record the molarity of the NaOH solution used in the titration.

Molarity of NaOH solution = _____ M

3. Determine the moles of NaOH used in each of the three trials. Record in the **CALCULATIONS TABLE**.

$$\textbf{By definition molarity} = \frac{\textbf{moles solute}}{\textbf{1 L solution}}$$

DON'T FORGET: You have to convert the mL of solution used into liters for the formula.

4. The balanced equation of the reaction between acetic acid (vinegar) and sodium hydroxide is:

$$\underset{\text{1 mole}}{HC_2H_3O_2} + \underset{\text{1 mole}}{NaOH} \longrightarrow \underset{\text{1 mole}}{H_2O} + \underset{\text{1 mole}}{NaC_2H_3O_2}$$

From the above relationship, you can calculate the moles of acetic acid in each titration. Record in the **CALCULATIONS TABLE**.

$$\textbf{moles NaOH X } \frac{\textbf{1 mole acetic acid}}{\textbf{1 mole NaOH}} \textbf{ = moles acetic acid}$$

Because the relationship of sodium hydroxide and acetic acid in this reaction is 1 to 1, the moles of acetic acid will be the same as the moles of sodium hydroxide.

5. Given the moles of acetic acid and the volumes of the vinegar sample for each trial, calculate the molarities for the three trials. Record in the spaces provided.

Don't forget to convert the volume to liters.

$$\frac{\textbf{moles of acid}}{\textbf{volume of vinegar}} \textbf{ = molarity of vinegar}$$

Molarity of vinegar for trial 1 = _____ M

Molarity of vinegar for trial 2 = _____ M

Molarity of vinegar for trial 3 = _____ M

14. Calculate the average molarity of the vinegar samples. Record in the space provided.

$$\frac{\textbf{molarity trial 1 + molarity trial 2 + molarity trial 3}}{\textbf{3}} \textbf{ = Average molarity}$$

Average molarity of vinegar = _____ M

6. The mass of one mole of $HC_2H_3O_2$ (acetic acid) is **60.0 g**. Remembering that molarity is equal to moles / liter, calculate the grams of acetic acid that would be in one liter of your vinegar sample by the following formula:

$$\frac{\textbf{moles } HC_2H_3O_2}{\textbf{1 L vinegar}} \textbf{ X } \frac{\textbf{60.0 g } HC_2H_3O_2}{\textbf{1 mole } HC_2H_3O_2} \textbf{ = g / liter of } HC_2H_3O_2$$

Mass of $HC_2H_3O_2$ / 1 L vinegar = _____ g

7. If we assume the density of vinegar is very close to 1.00 g/mL, then the mass of 1.00 L of vinegar is 1000 g. The percent, by mass, of acetic acid can be expressed as:

$$\frac{\textbf{mass of } HC_2H_3O_2 \textbf{ / 1 L}}{\textbf{1000 g vinegar}} \textbf{ X 100\% = \% } HC_2H_3O_2 \textbf{ (by mass)}$$

calculate the percentage of acetic acid in your vinegar sample. Record.

Set 1: % $HC_2H_3O_2$ = _____ % **Set 2: % $HC_2H_3O_2$ = _____ %**