Lab #10 – How much Acetic Acid (%) is in Vinegar?

**Purpose:** You will determine the amount of acetic acid in white vinegar (sold in grocery stores) by titration with a solution of NaOH whose concentration is known. The indicator will be **phenolphthalein** (feen-ohl-thay-leen)

**Concept:** An acid is a substance that increases the amount of hydrogen ion (H⁺) in solution, while a base increases the amount of hydroxide ion (OH⁻) in solution. When an acid and a base are mixed together, they neutralize each other, producing a salt and water.

Example:

\[
\text{KOH} + \text{HCl} \quad \rightarrow \quad \text{KCl}_{(aq)} + \text{H}_2\text{O} \\
\text{Base} \quad \text{acid} \quad \text{salt}
\]

In the reaction we will do today, acetic acid (HC₂H₃O₂) reacts with NaOH according to the equation shown below. Since the first hydrogen in acetic acid is easily lost (as H⁺), the compound is termed “acidic”. The other three hydrogens are not free to react, leaving the acetate ion (C₂H₃O₂⁻), a single polyatomic ion. When acetic acid reacts with the base, sodium hydroxide (NaOH), the products are a salt (sodium acetate) and water (H₂O).

This reaction is written as:

\[
\text{HC}_2\text{H}_3\text{O}_2 + \text{NaOH} \quad \rightarrow \quad \text{NaC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O}
\]

- You will add NaOH solution from a burette to the vinegar solution (in water) containing a few drops of the indicator.

- The indicator used in this titration is **phenolphthalein**, which is **pink in basic solutions and colorless in acidic solutions**.

- After all the acetic acid in the vinegar has reacted, the solution will turn pink with **just one drop** of NaOH solution. This point is called the **end point**.

- The end point is the **appearance of the first discernible pink color** (that persists for 30 seconds).

- At the end point, the **moles of NaOH** used to titrate the solution will be equal to the **moles of acid** you had originally.

- From this information, you will be able to calculate the **percentage of acetic acid in the vinegar**.
Prelab calculations (Problems #1-#6):

1. What is the molecular weight of $\text{H}_2\text{O}$?   ____________

2. What is the molecular weight of NaOH?   ____________

3. What is the molecular weight of $\text{HC}_2\text{H}_3\text{O}_2$?   ____________

THINGS TO REMEMBER!!!

Molarity is the number of moles of solute dissolved in one liter of solution:

$$ \text{molarity} = \frac{\text{moles}}{\text{liters}} $$

The number of MOLES in any volume of a solution can be calculated from a solution of a known molarity:

$$ \text{moles} = (\text{molarity}) \times (\text{liters}) $$

In this experiment the 1.00 $M$ NaOH solution is slowly added to the vinegar, which is of unknown acetic acid concentration. The key to the calculations in this experiment is that at the point when the reaction is just completed (the endpoint), we can use the balanced equation to determine that the number of moles of base that have been added is exactly equal to the number of moles of acid present in our original sample.

$$ \text{MOLES of ACID} = \text{MOLES of BASE (at the endpoint)} $$

In order to do a sample calculation, assume that 35.0 mL of base were added to reach the endpoint.

Since we know the concentration of the base and have measured the volume of base we have added, we can calculate the number of moles of base:

$$ \text{Moles of Base added} = \text{Volume of Base in Liters} \times \text{Concentration in Molarity} $$

4. Calculate the number of moles of base in the 35.0 mL of base (1.00 $M$ NaOH):

\# of moles of base just calculated is equal to \# of moles of acetic acid in the sample!!!!!!

Next, we can calculate the weight of acetic acid in grams in the sample:
weight of acetic acid = moles acetic acid \times \text{molecular weight acetic acid}

5. Calculate the weight of that many moles of acetic acid calculated in #4 (you determined the molecular weight of acetic acid on the previous page):

If you read the label on the vinegar bottle, the concentration of acetic acid is given in percent (by weight) and not in molarity. But since we have calculated the mass of acetic acid, we can calculate the mass percent if we know the mass of vinegar we used in the titration. Vinegar is mainly water (with a little acetic acid in it), so we'll assume that the vinegar has the same density as pure water, that is, 1.0 mL of acetic acid has the mass of 1.0 grams.

\[
\text{Mass percent} = 100\% \times \frac{\text{mass of acetic acid}}{\text{mass of vinegar}}
\]

6. Calculate the mass percent of acetic acid in the vinegar if the weight of acetic acid you calculated above (in #5) is present in 25.0 mL of vinegar:
EQUATIONS to USE

\[ \text{HC}_2\text{H}_3\text{O}_2 + \text{NaOH} \rightarrow \text{NaC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \]

1. molarity = moles/liters
2. moles = (molarity) x (liters)
3. moles of acid = moles of base added (at the endpoint – when solution turns PINK)
4. moles of base added = volume of base in Liters x concentration in Molarity
5. weight of acetic acid = moles acetic acid x molecular weight acetic acid
6. mass percent = 100% x (mass of acetic acid/mass of vinegar)

Procedure:
You will do a trial titration to determine the approximate end point. Then you will do a pair of exact titrations.

Getting Started: TRIAL TITRATION

(1) Obtain about 25.0 mL of vinegar in an Erlenmeyer flask.

(2) Add 4-5 drops of phenolphthalein indicator.

(3) Carefully fill the buret with 1.00 M NaOH solution.

(4) Record the initial burette reading (in table on next page) and the NaOH molarity

(5) Add the NaOH solution in increments of about 1 mL while swirling the flask.
   b. Observe the color of the solution after each addition.
   c. As the color change begins to persist for longer, add smaller volumes of base.

The trial titration is complete when the solution in the flask turns a faint shade of pink persists for at least 30 seconds. This is most easily visible if you place the flask on a piece of white paper. A deep pink/red color results if too much base has been added.

(6) Record the final burette reading (in the table).

(7) Calculate the volume of NaOH required for the trial titration.
EXACT TITRATIONS:
1. Repeat steps 1 – 4 of procedure for trial titration. If your buret is over half full, you will not need to refill it (step 3 above); just record the initial volume of the buret.

2. Since you found the approximate volume of base needed to reach the endpoint from the trial titrations, you know the approximate volume of NaOH you will need to reach the endpoint. If you were successful in obtaining a pale pink color for the endpoint, subtract 1 mL from the volume found in the trial titration; otherwise subtract 2mLs. Rapidly add the resulting volume of NaOH to the flask while swirling the flask.

3. Rinse the walls of the flask with distilled water from a plastic squirt bottle.

4. Continue the titration on a drop-by-drop basis. Swirl the flask rapidly after each drop. The end point is the first permanent discernible pink color. Record the final buret reading for this trial and calculate the volume of base used.

5. Repeat the procedure with another sample of vinegar.
6. Complete the calculations in the table below.

<table>
<thead>
<tr>
<th></th>
<th>Titration #1 (Trial)</th>
<th>Titration #2</th>
<th>Titration #3</th>
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<tbody>
<tr>
<td>Volume of Vinegar</td>
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<tr>
<td>Molarity of NaOH</td>
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<td>Initial Buret Reading</td>
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<tr>
<td>Final Buret Reading</td>
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<td>Volume NaOH Added</td>
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<tr>
<td>Moles NaOH added</td>
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<tr>
<td>Moles Acetic Acid</td>
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<td>Mass of Acetic Acid</td>
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<td>Mass % of Acetic Acid</td>
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7. Calculate the average percent acetic acid in vinegar using titrations #2 and #3.

When you have completed the experiment, rinse your buret two times with tap water and then once with distilled water!!! Hang your buret upside down to dry.