Vinegar is a solution of acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$). The strength or concentration of the vinegar is usually given on the label of the bottle in percent by weight or “percent acidity.” The United States Food and Drug Administration regulations require that the product called simply “vinegar” be made from apples and contain not less than 4 g of acetic acid in 100 mL of vinegar. One way to produce a cheap vinegar is to keep the concentration at this allowable minimum; however, vinegar may gradually lose strength on the shelf so the manufacturer may wish to make the product stronger than necessary in order to guarantee a good shelf life. Vinegars that are not made from apples are available, including malt vinegar, made from barley and corn, wine vinegar and rice vinegar. Whatever the source, acetic acid is the “sour”, or acid, ingredient.

In our experiment, we will determine the acetic acid concentration of your vinegar by titration with sodium hydroxide solution. We have previously standardized the sodium hydroxide and now know its concentration to 5 parts per thousand. The sodium hydroxide is about 0.1 M and is too dilute to provide a satisfactory titration with commercial vinegar. Therefore we dilute the vinegar using accurate volumetric glassware before performing the titrations. The instructions below are for 4-6% acetic acid; if your vinegar is much stronger than 6%, consult the instructor.

Different vinegars may have different subtle flavoring agents, nevertheless, the vinegar acts as a source of water-soluble acid in food preparation. (Other acid sources that are sometimes used are lemons and sour milk.) One way to evaluate the vinegar on a cost basis, then, is to determine the cost in dollars per gram of acetic acid for your vinegar.

**PROCEDURE**

Bring from home a bottle of clear or pale-colored vinegar.

Obtain a 250 mL volumetric flask, a burette, a 25 mL volumetric pipet, and a pipet bulb from the stockroom. The pipet bulb is used to fill and empty the pipet; the instructor will demonstrate its use.

Rinse the volumetric flask and the pipet with tap water and with deionized water. Pour 50 to 100 mL of commercial vinegar into a clean, dry beaker. Now rinse the pipet at least three times with vinegar, discarding the rinsing’s. Fill the pipet carefully to the mark with vinegar and allow the 25 mL sample of vinegar to flow into the volumetric flask. (Do not “blow out” the remaining solution in the tip from these “TD”, or “to deliver”, pipets.) Fill the volumetric flask carefully to the mark with deionized water and mix well by inverting the stoppered flask at least twenty times. Pour your diluted vinegar, which we will call the “acetic acid solution”, into a suitable clean, dry, stoppered container. Clean the volumetric flask and return it to the stockroom.
Rinse the burette well with tap water, deionized water and your (standardized) sodium hydroxide solution. Rinse the pipet 3 times with acetic acid solution. Fill the burette with the sodium hydroxide solution from part 12A and record the initial reading.

Carry out the following procedure for each titration:

1. Pipet 25.0 mL of solution into a clean Erlenmeyer flask. (It need not be dry.) Add 2 drops of phenolphthalein.

2. Titrate the acetic acid solution with sodium hydroxide to a phenolphthalein end point. Record the final volume on the burette.

3. Repeat the titration until 3 values for the volume of sodium hydroxide used agree within 1% (about 0.1-0.2 mL).

Use the Q test to reject any suspect values.

**CALCULATIONS**

Calculate the molarity of the acetic acid solution. The reaction is:

\[
\text{NaOH(aq)} + \text{HC}_2\text{H}_3\text{O}_2(aq) \rightarrow \text{NaC}_2\text{H}_3\text{O}_2(aq) + \text{H}_2\text{O(l)}
\]

The commercial vinegar is exactly 10 times as strong since we diluted 25.0 mL to 250. mL.

Assume that the density of the commercial vinegar is 1.00 g/mL; calculate the concentration of the vinegar in percent acetic acid by mass.

For a cost analysis, calculate the mass of acetic acid per dollar, taking the following into account:

1. Cost of the bottle of vinegar.

2. Total amount in the bottle (usually given in fluid ounces).

3. Mass of a fluid ounce of vinegar (32 fluid ounces per quart, 946 mL per quart, 1.00 g/mL of vinegar assumed).

**CONCLUSION**

Compare your analysis of the vinegar with the concentration given on the bottle. Also, collect information from other students to compare the vinegars on a cost basis. Evaluate your vinegar as to truthful labeling and economy.
Section________________________ Name_____________________

**Pre-Laboratory Assignment**

1. Write below the following information from your bottle of commercial vinegar:

   **Ingredients:**

   **Cost:**

   **Size of container (usually given in fluid ounces):**

   **% acetic acid:**

2. Calculate the molarity of a 5.1% acetic acid solution, assuming that the density of the solution is 1.00 g/mL.
Report Sheet
For each titration, you will report the following. Repeat the data table and calculations for each titration.

Initial buret reading  
Final buret reading  
Volume of NaOH used  
Average volume of NaOH  
Q test results (if used):

Calculations:
Include an example of all of your calculations with your report.

Molarity of NaOH  
Molarity of acetic acid solution  
Molarity of commercial vinegar  
% acetic acid by mass in vinegar  
Cost of vinegar: $/g  

Comparison with other students' results:

<table>
<thead>
<tr>
<th>Student</th>
<th>Brand</th>
<th>Cost: $/bottle</th>
<th>Cost: $/g acetic acid</th>
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Conclusion:
Post Laboratory Questions

1. A generic vinegar, 4.5% acetic acid, costs 95 cents for an 8 oz bottle and $2.79 for a quart bottle. Which is cheaper? Calculate the cost of each bottle in $/g acetic acid.

2a. Lemon juice contains citric acid, $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$. Write the balanced equation for the reaction of NaOH(aq) with citric acid.

b. Calculate the volume of 0.1205 M NaOH that will neutralize 10.0 mL of 3.8% (by mass) citric acid solution. Assume the density of the citric acid solution is 1.00 g/mL.