Objectives

- To determine the molarity and percent by mass of acetic acid in vinegar.

Vinegar is essentially a solution of acetic acid (\(\text{HC2H3O2}\)) in water. The concentration of acetic acid in vinegar may be expressed as a molarity (in mol/L):

\[
\text{Molarity} = \dfrac{\text{Moles of Acetic Acid}}{\text{Volume of Vinegar (in L)}}
\]

or as a mass percent

\[
\text{Mass }\% = \left(\dfrac{\text{Mass of Acetic Acid}}{\text{Mass of Vinegar}}\right) \times 100\%
\]

In this experiment, a technique known as a titration will be used to determine the concentration of acetic acid in vinegar. A titration involves performing a controlled reaction between a solution of known concentration (the titrant) and a solution of unknown concentration (the analyte). Here, the titrant is an aqueous solution of ~0.1 M sodium hydroxide (\(\text{NaOH}\)) and the analyte is vinegar. When mixed, a neutralization reaction occurs between sodium hydroxide and the acetic acid in vinegar:

\[
\text{NaOH (aq) + HC2H3O2 (aq) → NaC2H3O2 (aq) + H2O (l)}
\]

The sodium hydroxide will be gradually added to the vinegar in small amounts from a burette. A burette is a device that allows the precise delivery of a specific volume of a solution. The \(\text{NaOH}\) will be added to the vinegar sample until all the acetic acid in the vinegar has been exactly consumed (reacted away). At this point the reaction is completed, and no more \(\text{NaOH}\) is required. This is called the equivalence point of the titration.

In order to know when the equivalence point is reached, an indicator solution called phenolphthalein is added to the vinegar at the beginning of the titration. Phenolphthalein is a pH sensitive organic dye. Phenolphthalein is colorless in acidic solutions like vinegar, and deep pink in basic solutions like sodium hydroxide. At the equivalence point of the titration, just one drop of \(\text{NaOH}\) will cause the entire solution in the Erlenmeyer flask to change from colorless to a very pale pink.

As the titration is performed, the following data will be collected:

- The molarity of \(\text{NaOH}\) (aq) used
- The volume of \(\text{NaOH}\) (aq) used to neutralize the vinegar
- The volume of vinegar used.

Using this data, the molarity and mass percent of acetic acid in vinegar can be determined by performing a series of solution stoichiometry calculations (see Calculations Section).
Procedure

Materials and Equipment

50-mL burette*, 5-mL volumetric pipette*, pipette bulb*, ~ 0.1 M \(\text{NaOH} (aq)\), vinegar, phenolphthalein, burette stand, two 250-mL (or 125 mL) Erlenmeyer flasks, wash bottle with distilled water, funnel

Safety

Be especially careful when handling the sodium hydroxide base \(\text{NaOH} (aq)\), as it is corrosive and can cause chemical burns to the skin. If any NaOH spills on you, rinse immediately under running water for up to 15 minutes and report the accident to your instructor.

Titration Procedure

Your instructor will demonstrate the correct use of the volumetric pipette and burette at the beginning of the lab session. Detailed instructions on how to use a pipette are also found on the last page of this handout. Note that three titrations must be performed.

1. Obtain a 50-mL burette, 5-mL volumetric pipette and a pipette bulb from the stockroom.

Setting up the burette and preparing the \(\text{NaOH} (aq)\)

2. Rinse the inside of the burette with distilled water. Allow the distilled water to drain out through the tip in order to ensure that the tip is also rinsed.

3. Now rinse the burette with a small amount of \(\text{NaOH} (aq)\). To do this, add about 5-mL of \(\text{NaOH} (aq)\) to the burette, then twirl the burette on its side (over the sink) to rinse its entire inner surface. Then allow the \(\text{NaOH} (aq)\) to drain out through the tip.

4. Fill the burette with \(\text{NaOH} (aq)\) up to the top, between 0-mL and 5-mL. Use a funnel to do this carefully, below eye-level, and preferably over the sink. After this you will need to flush the tip of the burette — your instructor will show you how to do this. Now measure the volume at the level of the \(\text{NaOH} (aq)\) precisely, and record it as the “Initial Burette Reading” on your report. Also record the exact molarity of the NaOH (aq), which is labeled on the stock bottle.

Preparing the vinegar sample

5. The volumetric pipette used in this lab is designed to measure and transfer exactly 5.00 mL of solution. First, rinse the inside of the volumetric pipette with distilled water. Using the pipette bulb, draw the water into the pipette up above the 5-mL mark, then allow it to drain out through the tip. You may want to do this several times for practice. Then perform a final rinse, but this time use vinegar.

6. Now use the volumetric pipette to transfer 5.00-mL of vinegar into a clean 250-mL Erlenmeyer flask (see instructions on page 4). Record this volume of vinegar (precise to two decimal places) on your report. Then add about 20-mL of distilled water and 5 drops of phenolphthalein to this Erlenmeyer flask.

Performing the titration

7. Begin the titration by slowly adding \(\text{NaOH} (aq)\) from the burette to the vinegar in the Erlenmeyer flask. Swirl Erlenmeyer flask as you add the base in order to efficiently mix the chemicals. Some pinkness may appear briefly
in the flask as the base is added, but it will quickly disappear as the flask is swirled.

8. As the equivalence point is approached, the pink color will become more pervasive and will take longer to disappear. When this occurs, start to add the \( \ce{NaOH} (aq) \) drop by drop. Eventually the addition of just one drop of \( \ce{NaOH} (aq) \) will turn the solution in the Erlenmeyer flask a pale pink color that does not disappear when swirled. This indicates that the equivalence point has been reached. **Do not add any more \( \ce{NaOH} (aq) \)** at this point. Measure this volume of \( \ce{NaOH} (aq) \) precisely, and record it as the “Final Burette Reading” on your report. Then show the resulting solution in the flask to your instructor so s/he can record the final color on your report form.

9. Refill your burette with \( \ce{NaOH} (aq) \), and then repeat this procedure for a second sample of vinegar, and then a third sample of vinegar. You do not need to flush the tip of the burette again. Note that if you use less than 25-mL of \( \ce{NaOH} (aq) \) for the second titration, you do not need to refill the burette for the third titration; also that you will need to clean out and re-use one of your Erlenmeyer flasks for the third titration. You and your partner should take turns performing these titrations.

10. When finished, dispose of your chemical waste as instructed.

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**Pipetting Instructions**

1. Get the appropriate amount of the solution you wish to pipette in a clean, dry beaker. Never pipette directly out of the stock bottles of solution. This creates a contamination risk.

2. Insert the tip of the pipette into the beaker of solution so that it is about a quarter inch from the bottom. Be sure not to press the tip against the bottom of the container.

3. If you are right handed, hold the pipette in your right hand, leaving your index finger free to place over the top of the pipette. With your left hand, squeeze the pipette bulb. Press it firmly over the top of the pipette, but DO NOT
INSERT THE PIPET DEEP INTO THE BULB!

4. Release the pressure on the bulb and allow the solution to be drawn up into the pipette until it is above the volume mark. Do not allow the solution to be sucked into the bulb itself.

5. Quickly remove the bulb and place your index finger firmly over the top of the pipette. Then remove the pipette tip from the beaker of solution.

6. Slowly roll your finger to one side and allow the liquid to drain until the bottom of the meniscus is aligned with the volume mark. With practice you will be able to lower the liquid very, very slowly.

7. When the bottom of the meniscus is even with the volume mark, press your index finger firmly on the top of the pipette so no liquid leaks out. Touch the tip once to the side of the beaker to remove any hanging drops.

8. To transfer the solution, place the tip of the pipette against the wall of the receiving container at a slight angle. Then allow the liquid to drain from the pipette.

9. When the solution stops flowing, touch the pipette once to the side of the receiving container to remove any hanging drops. DO NOT blow out the remaining solution. The pipette has been calibrated to deliver the appropriate amount of solution with some remaining in the tip.

Calculations

Molarity of Acetic Acid in Vinegar

- First, using the known molarity of the \(\ce{NaOH}\) (aq) and the volume of \(\ce{NaOH}\) (aq) required to reach the equivalence point, calculate the moles of \(\ce{NaOH}\) used in the titration.

- From this mole value (of \(\ce{NaOH}\)), obtain the moles of \(\ce{HC2H3O2}\) in the vinegar sample, using the mole-to-mole ratio in the balanced equation.

- Finally, calculate the molarity of acetic acid in vinegar from the moles of \(\ce{HC2H3O2}\) and the volume of the vinegar sample used.

Mass Percent of Acetic Acid in Vinegar

- First, convert the moles of \(\ce{HC2H3O2}\) in the vinegar sample (previously calculated) to a mass of \(\ce{HC2H3O2}\), via its molar mass.

- Then determine the total mass of the vinegar sample from the vinegar volume and the vinegar density. Assume that the vinegar density is 1.000 g/mL (= to the density of water).

- Finally, calculate the mass percent of acetic acid in vinegar from the mass of \(\ce{HC2H3O2}\) and the mass of vinegar.
1. In this lab, you will perform a titration using sodium hydroxide and acetic acid (in vinegar). Write the balanced neutralization reaction that occurs between sodium hydroxide and acetic acid.

2. Specialized equipment is needed to perform a titration.
   - Consider the sodium hydroxide reactant.
     - Name the specialized device the sodium hydroxide is placed in.
     - Is the concentration of the sodium hydroxide known or unknown?
     - Is sodium hydroxide the analyte or the titrant?
   - Consider the acetic acid reactant.
     - What type of flask is the acetic acid placed in?
     - What volume of acetic acid is used?
     - What specialized device is used to obtain this precise volume?
     - Is the acetic acid the analyte or the titrant?

3. You will add sodium hydroxide to the acetic acid until all the acetic acid is consumed. This is a special point in the titration called the _________________________ point.

4. An indicator solution is used to indicate when all the acetic acid has been consumed and that the reaction is complete.
   - What is the name of the indicator solution?
   - Is this indicator mixed with sodium hydroxide or acetic acid?
   - How exactly does the indicator let you know when the reaction is complete?
**Lab Report: Titration of Vinegar**

**Experimental Data**

<table>
<thead>
<tr>
<th>Trial 1</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial Buret Reading</td>
</tr>
<tr>
<td>Final Buret Reading</td>
</tr>
<tr>
<td>Volume of (\text{NaOH}(aq)) used</td>
</tr>
<tr>
<td>Molarity of (\text{NaOH}(aq)) used</td>
</tr>
<tr>
<td>Volume of Vinegar used</td>
</tr>
</tbody>
</table>

Color at equivalence point – to be recorded by your instructor

**Data Analysis**

Write the balanced equation for the neutralization reaction between aqueous sodium hydroxide and acetic acid.

*The Molarity of Acetic Acid in Vinegar*

Use your two best sets of results (with the palest pink equivalence points) along with the balanced equation to determine the molarity of acetic acid in vinegar. Show all work for each step in the spaces provided.

<table>
<thead>
<tr>
<th>Data used ⇒</th>
</tr>
</thead>
<tbody>
<tr>
<td>Moles of (\text{NaOH}) used in titration</td>
</tr>
<tr>
<td>Moles of (\text{HC2H3O2}) neutralized in vinegar sample</td>
</tr>
<tr>
<td>Molarity of (\text{HC2H3O2}) in vinegar</td>
</tr>
</tbody>
</table>
**The Mass Percent of Acetic Acid in Vinegar**

Use your two best sets of results along with calculated values in the previous table to determine the mass percent of acetic acid in vinegar. Show all work for each step in the spaces provided.

**Questions**

1. What was the purpose of the phenolphthalein indicator in this experiment? Be specific.

2. Suppose you added 40 mL of water to your vinegar sample instead of 20 mL. Would the titration have required more, less or the same amount of \(\ce{NaOH}\) (aq) for a complete reaction? Explain.

3. Consider a 0.586 M aqueous solution of barium hydroxide, \(\ce{Ba(OH)2}\) (aq).
   - How many grams of \(\ce{Ba(OH)2}\) are dissolved in 0.191 dL of 0.586 M \(\ce{Ba(OH)2}\) (aq)?
   - How many individual hydroxide ions \(\ce{OH^-}\) are found in 13.4 mL of 0.586 M \(\ce{Ba(OH)2}\) (aq)?
   - What volume (in L) of 0.586 M \(\ce{Ba(OH)2}\) (aq) contains 0.466 ounces of \(\ce{Ba(OH)2}\) dissolved in it?

4. If 16.0 mL of water are added to 31.5 mL of 0.586 M \(\ce{Ba(OH)2}\) (aq), what is the new solution molarity?

5. Suppose you had titrated your vinegar sample with barium hydroxide instead of sodium hydroxide:

\[
\ce{Ba(OH)2 (aq) + 2 HC2H3O2 (aq) -> Ba(C2H3O2)2 (aq) + 2 H2O (l)}}
\]

   - What volume (in mL) of 0.586 M \(\ce{Ba(OH)2}\) (aq) must be added to a 5.00 mL sample of vinegar to reach the equivalence point? Use your average vinegar molarity (see page 1) in this calculation.