

Percentage of Acetic Acid in Vinegar

OBJECTIVES

- Determine the endpoint of an acid-base titration.
- Use burets for accurately measuring quantities of solution.
- Calculate the molarity of vinegar from experimental data.
- Calculate the percentage of acetic acid in vinegar.

MATERIALS

- 125 mL Erlenmeyer flask
- 100 mL beakers, 3
- 250 mL beaker (for waste solutions)
- burets, 2
- buret clamp
- NaOH solution, standardized
- phenolphthalein indicator
- ring stand
- wash bottle
- white vinegar

BACKGROUND

When sweet apple cider is fermented, the product is either an alcohol called apple jack or an acid called vinegar. If fermentation takes place without oxygen, alcohol and carbon dioxide are produced. But if oxygen is present in the fermentation process, acetic acid and carbon dioxide are produced. Most commercial vinegars have a mass percentage of acetic acid between 4.0% and 5.5%. The white vinegar you will use in this experiment is not produced by fermentation; it is obtained by the dilution of 100% acetic acid.

The percentage of acetic acid in a sample of vinegar may be found by titrating the sample against a standard basic solution. By determining the volume of sodium hydroxide solution of known molarity necessary to neutralize a measured quantity of vinegar, the molarity of the vinegar can be calculated. The molarity can be converted to the percentage CH_3COOH in vinegar.

SAFETY

Always wear safety goggles and a lab apron to protect your eyes and clothing. If you get a chemical in your eyes, immediately flush the chemical out at the eyewash station while calling to your teacher. Know the location of the emergency lab shower and eyewash station and the procedure for using them.

Do not touch any chemicals. If you get a chemical on your skin or clothing, wash the chemical off at the sink while calling to your teacher. Make sure you carefully read the labels and follow the precautions on all containers of chemicals that you use. If there are no precautions stated on the label, ask your teacher what precautions you should follow. Do not taste any chemicals or items used in the laboratory. Never return leftovers to their original containers; take only small amounts to avoid wasting supplies.

FIGURE A



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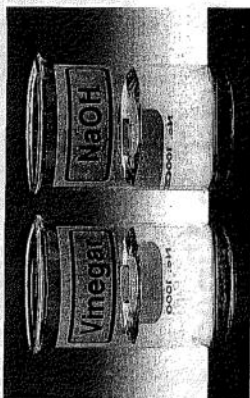


FIGURE A

Call your teacher in the event of a spill. Spills should be cleaned up promptly, according to your teacher's directions.

Never put broken glass in a regular waste container. Broken glass should be disposed of in the broken glass waste container.

PREPARATION

1. Copy the data table below in your lab notebook. Leave a space for the molarity of the standardized NaOH solution.

2. Label one clean, dry 100 mL beaker *Vinegar* and the other *NaOH*, as shown in Figure A. Label one buret *Vinegar* and the other *NaOH*.

PROCEDURE

1. Transfer approximately 80 mL of vinegar and approximately 80 mL of NaOH to the appropriately labeled beakers.

2. Pour approximately 5 mL of vinegar from the beaker into the appropriately labeled buret. Rinse the walls of the buret thoroughly with the vinegar. Allow the vinegar to drain through the stopcock into the 250 mL beaker for waste. Rinse the buret two more times in this way,

using a new 5 mL portion of vinegar each time. Collect all rinses in the waste beaker.

3. Fill the buret with vinegar above the zero mark. Withdraw enough vinegar to remove the air from the tip of the buret and bring the liquid level into the graduated region of the buret.

4. Repeat Procedure steps 2 and 3 with the sodium hydroxide solution and the appropriately labeled buret.

5. Record the initial readings of both burets, estimating the volumes to the nearest 0.01 mL.

6. Allow about 10 mL of vinegar to flow into a clean Erlenmeyer flask. Add about 10 mL of distilled water to the flask to increase the volume. This procedure will make it easier to determine the color change when the end point is reached. Add one or two drops of phenolphthalein solution to serve as an indicator.

7. Titrate the vinegar with the standard solution of sodium hydroxide. Continually swirl the flask as shown in Figure B. Stop frequently to wash down the sides of the flask with distilled water from

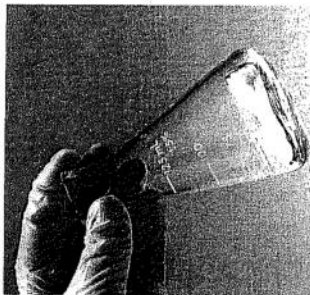


FIGURE B

Data Table 1

Trial number	Initial NaOH reading (mL)	Final NaOH reading (mL)	Initial vinegar reading (mL)	Final vinegar reading (mL)
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2				
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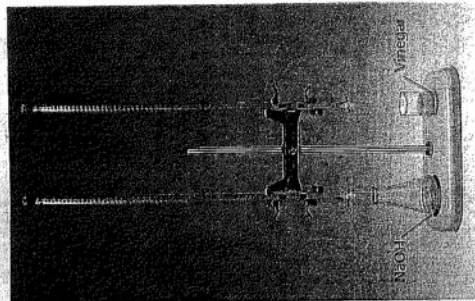


FIGURE C

your wash bottle. Add the sodium hydroxide drop by drop near the end of the titration until the last drop keeps the solution a pink color that remains after swirling.

8. Add successive quantities of both solutions, drop by drop, going back and forth from pink to colorless until the end point is clearly established.

This point is indicated by the slightest suggestion of pink coloration in the flask, as shown in Figure C. You will be able to see the pink color more clearly if the flask is resting on a sheet of white paper with a beaker of distilled water next to it for comparison. Record in your data table the final buret readings of both solutions to the nearest 0.01 mL.

9. Discard the liquid in the Erlenmeyer flask in the disposal container provided by your teacher. Rinse the flask thoroughly with distilled water, and repeat the titration two more times, following Procedure steps 5–8.

CLEANUP AND DISPOSAL

10. Clean all apparatus and your lab station. Return equipment to its proper place. Dispose of chemicals and solutions in the containers designated by your teacher. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly before you leave the lab and after all work is finished.

ANALYSIS AND INTERPRETATION

1. **Organizing Data:** Calculate the volumes of vinegar and NaOH used for each of the three trials.
2. **Organizing Data:** In your data table write the molarity of the standardized NaOH solution you used. Determine the moles of NaOH used in each of the three trials.
Hint: By definition,

$$\text{molarity} = \frac{\text{moles of solute}}{\text{L solution}}$$

$$\text{moles of solute} = \text{molarity} \times \text{liters of solution}$$

3. **Organizing Ideas:** Write the balanced equation for the reaction between vinegar and sodium hydroxide. (Hint: The formula for acetic acid is CH_3COOH .)

4. **Organizing Data:** Use the results of your calculations in Analysis and Interpretation item 2 and the mole ratio from the equation in Analysis and Interpretation item 3 to determine the moles of base used to neutralize the vinegar (acid) in each trial.

5. **Organizing Data:** Use the moles of base calculated in Analysis and Interpretation item 4 and the volumes of the acid used for each trial to calculate the molarities of the vinegar for the three trials.

6. **Organizing Data:** Calculate the average molarity of the vinegar.

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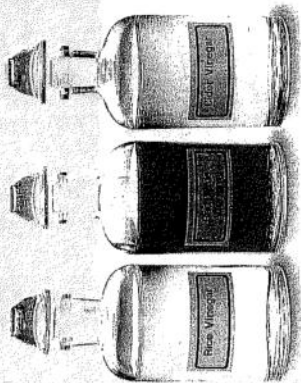
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2. Applying Conclusions: Why is it important for a company manufacturing vinegar to regularly check the molarity of its product?

3. Analyzing Methods: What was the purpose of using the phenolphthalein? Could you have titrated the vinegar sample without the phenolphthalein?

4. Analyzing Methods: At the beginning of each titration, 10 mL of vinegar was run into the Erlenmeyer flask and the vinegar was diluted with distilled water. Why was the calculated molarity of the acetic acid not affected by the water?



7. Organizing Ideas: Use the periodic table to calculate the molar mass of acetic acid, CH_3COOH .

8. Organizing Data: Use the average molarity for your vinegar sample to determine the mass of CH_3COOH in 1 L of vinegar.

CONCLUSIONS

1. Organizing Conclusions: Assume that the density of vinegar is very close to 1.00 g/mL so that the mass of 1 L of vinegar is 1000 g. Calculate the percentage of acetic acid in your vinegar sample. (Hint: The mass of acetic acid in 1 L of vinegar was calculated in Analysis and Interpretation item 8. Divide the mass of CH_3COOH in 1 L by the total mass of vinegar in a liter, then multiply by 100 to get the percentage of acetic acid in vinegar.)

EXTENSIONS

1. Evaluating Data: Share your data with other lab groups, and calculate a class average for the molarity of the vinegar.

2. Designing Experiments: What possible sources of error can you identify in this procedure? If you can think of ways to eliminate the errors, ask your teacher to approve your plan, and run the procedure again.

3. Relating Ideas: Explain the difference between the equivalence point and the end point. Can they be the same?

4. Resolving Discrepancies: An industrial chemist measured the following values when titrating 10.0 mL samples from a single vat of vinegar: 15.04 mL, 16.03 mL, and 14.98 mL. What could be the source error in these titrations?