EXPERIMENT

Titration for Acetic Acid in Vinegar

Hands-On Labs, Inc.
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Review the safety materials and wear goggles when working with chemicals. Read the entire exercise before you begin. Take time to organize the materials you will need and set aside a safe work space in which to complete the exercise.

Experiment Summary:

You will learn about acetic acid (CH_3COOH) and how to determine the concentration of acetic acid in vinegar through titration. You will explore the concepts of stoichiometric equilibrium, concentration, molarity, indicators, and mass percent.

LabPaq Science Solutions for Serious Educators

Learning Objectives

Upon completion of this laboratory, you will be able to:

- Define titration, titrant, analyte, and equivalence (or stoichiometric) point.
- Discuss why phenolphthalein is an effective pH indictor, and how it works.
- Describe the process and purpose of titration.
- Write a balanced chemical equation representing the stoichiometric equivalence point of a reaction.
- Describe how strong and weak acids differ.
- Apply titration techniques to investigate acetic acid in commercial vinegar.
- Determine the molar concentration of acetic acid in commercial vinegar.
- Calculate the average concentration and the percent (%) concentration of acetic acid in vinegar.

Time Allocation: 2.5 hours

Materials

Student Supplied Materials

Quantity	Item Description
1	Bottle of distilled water
1	Dish soap
1	Pair of scissors
1	Roll of paper towels
1	Sheet of white paper
1	Source of tap water
2-6	Textbooks

HOL Supplied Materials

Quantity	Item Description
1	Glass beaker, 100 mL
1	Graduated cylinder, 25 mL
1	Pair of safety goggles
1	Pair of gloves
1	Stopcock
1	Syringe, 10 mL
1	Test tube clamp
1	Test tube cleaning brush
1	Experiment Bag: Titration for Acetic Acid in Vinegar 1- Phenolphthalein solution, 1%, 0.5 mL 1- Vinegar, 20 mL in dropper bottle 1- Pipet, short stem 1- Sodium hydroxide, 0.5 M, 30 mL

Note: *To fully and accurately complete all lab exercises, you will need access to:*

1. A computer to upload digital camera images.
2. Basic photo editing software such as Microsoft Word® or PowerPoint®, to add labels, leader lines, or text to digital photos.
3. Subject-specific textbook or appropriate reference resources from lecture content or other suggested resources.

Note: *The packaging and/or materials in this LabPaq kit may differ slightly from that which is listed above. For an exact listing of materials, refer to the Contents List included in your LabPaq kit.*

Background

Acetic Acid

Have you ever opened a bottle of vinegar and noticed the intense scent? Now think about the sour taste you get while nibbling on a bag of salt and vinegar potato chips. You might be surprised to learn that both the intense scent of vinegar and sour taste of salt and vinegar chips are the result of acetic acid, the chemical component of vinegar. Acids are molecular substances that ionize in water to release H^+ ions. **Acetic acid** (CH_3COOH) is a **weak acid**, which means that only a small percentage of acetic acid molecules ionize when dissolved in water (only the hydrogen atom bonded to the oxygen ionizes). In contrast, the acidic hydrogen atoms in a **strong acid** are completely ionized in water. All acids, weak or strong, form a conjugate base and some number of protons when dissolved in water. See Figure 1.

$$CH_3COOH_{(aq)} \underset{}{\overset{H_2O}{\rightleftharpoons}} H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

$$HNO_3 \xrightarrow{H_2O} H^+_{(aq)} + NO_3^-{}_{(aq)}$$

Figure 1. Strong and weak acids. **A.** The weak acid, acetic acid (CH_3COOH), only partially ionizes in water. Notice that of the four total hydrogen molecules in acetic acid, only one H^+ ion is produced in solution. **B.** The strong acid, nitric acid (HNO_3), completely ionizes in water.

The amount of acetic acid in common table vinegars, apple cider vinegar, balsamic vinegar, white vinegar, and sherry vinegar, varies from approximately 4-6%. However, some varieties of vinegar, such as pickling vinegar or distilled vinegar, contain approximately 5-8% acetic acid.

Titration

The exact concentration of acetic acid in a bottle of vinegar can be determined through titration with a strong base. **Titration** is a quantitative, volumetric technique where a solution of a known concentration (**titrant**) is added to a solution of an unknown concentration (**analyte**) until the equivalence point is reached. See Figure 2. The **equivalence point** of a titration, also known as a **stoichiometric point** or end point, is the moment in a titration where exactly enough titrant has been added to completely react with the analyte. In an acid-base titration, the equivalence point can be identified through the use of a pH indicator.

Figure 2. Laboratory titration apparatus. © Alexander Raths

A **pH indicator** is a substance that changes color when the pH of a solution changes, allowing scientists to qualitatively measure the moment when the analyte has completely reacted with the titrant. A common indicator for a titration between a weak acid and a strong base is phenolphthalein. See Figure 3. **Phenolphthalein** is a pH indicator, which turns bright-pink at a basic pH of 8.2 and higher, allowing for equivalence points in titrations to be marked by the analyte changing in color from colorless to bright pink.

Figure 3. Phenolphthalein added to solution. © Kesu

When designing a titration between an acid and a base, it is important to know the reaction between titrant and analyte (chemical equation), the exact stoichiometric equivalence point (balanced chemical equation), and the concentration of the titrant. To determine the exact concentration of acetic acid in vinegar, a titration between acetic acid (CH_3COOH) and sodium hydroxide (NaOH) will be performed, using phenolphthalein as the indicator. The balanced chemical equation for the reaction between CH_3COOH and NaOH is shown below:

$$CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaC_2H_3O_{2(aq)}$$

From the balanced equation, we know that the reaction between acetic acid and sodium hydroxide is 1:1, thus the number of moles of NaOH required to cause a change in the indicator color is equal to the number of moles of CH_3COOH in the vinegar. The stoichiometric equivalence point (end point) of the titration will be visualized when the phenolphthalein causes the solution to change from a colorless solution to a bright-pink colored solution.

The titration will provide the number of moles of CH_3COOH present in the vinegar sample. In order to determine the %CH_3COOH in the sample, the **mass percent** (grams of acetic acid/grams of vinegar) will be determined using the molar mass of CH_3COOH (60.05g CH_3COOH/1 mol CH_3COOH) and the density of vinegar (1.00 g/mL).

Titration is used in many different ways to measure chemical concentrations in solutions: Nutritional analysis often uses titration methods to measure acidity in a food (for example, orange juice); Blood testing and pregnancy tests may use methods of titration to measure chemical levels; Aquarium water testing uses titration methodologies to ensure a healthy ecosystem for aquatic pets; Winemakers may use fairly simple titration kits to measure the acidity of wine; The development of medications (pharmacology) by pharmaceutical companies frequently involves the use of specialized titration equipment to accurately measure chemical quantities; Acid rain evaluation employs titration processes to quantify the degree of contamination in natural rain water or snow; Wastewater analysis often uses specialized titration equipment to analyze contamination, and thus determine the requirements for filtering and cleaning.

Exercise 1: Determining the Concentration of Acetic Acid

In this exercise, you will determine the concentration of acetic acid (CH_3COOH) in the vinegar provided in your lab kit.

Note: *Please read all steps and safety information before starting the procedure.*

Procedure

1. Gather the test tube holder, small stopcock, 10-mL syringe (titrator), and either 2 thick textbooks and the lab kit box **or** 5-6 thick textbooks. See Figure 4.

Figure 4. Titrator and small stopcock.

2. Remove the plunger from the titrator and place it back in your lab kit box.
3. Attach the stopcock to the tip of the titrator by placing the larger, clear, plastic end of the stopcock into the tip of the titrator and then twist the stopcock into place. The stopcock should fit tightly into the titrator so that the liquid will not leak. See Figure 5.

Figure 5. Fitting the stopcock into the titrator.

4. Stack the 5 textbooks or stack 2 textbooks on top of the lab kit box.
5. Clamp the test tube holder around the middle of the titrator and slide the long end under the top 2 books in the stack. Place a sheet of white paper next to the bottom of the stack and set the 100 mL beaker on the sheet of white paper. The end of the stopcock should be located near the top of the beaker, approximately 1 cm above to 1 cm below the top of the beaker. See Figure 6.

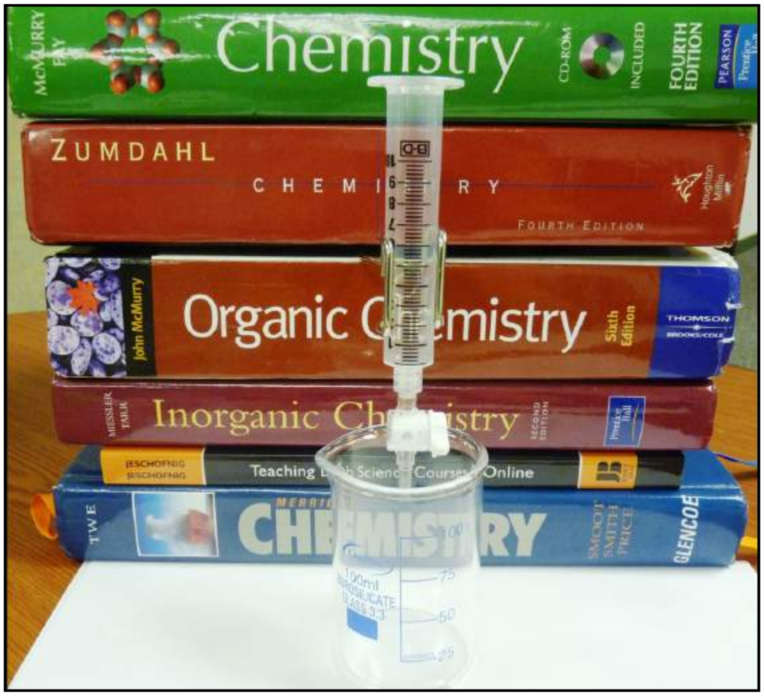

Figure 6. Titration setup.

Note: It is important that the placement of the titrator allows for the white knob to be easily adjusted. If this is not the case, then either adjust the location of the books in the stack or slightly adjust where in the test tube clamp the titrator is located.

6. Use the pipet to fill the titrator with 7 - 9 mL of distilled water.

Note: It is important that you use distilled water for this step and not tap water.

7. Using both hands, one hand on the titrator and your other hand on the stopcock, practice releasing water from the titrator into the beaker. The goal is to be comfortable releasing only 1 drop at a time from the titrator. See Figure 7.

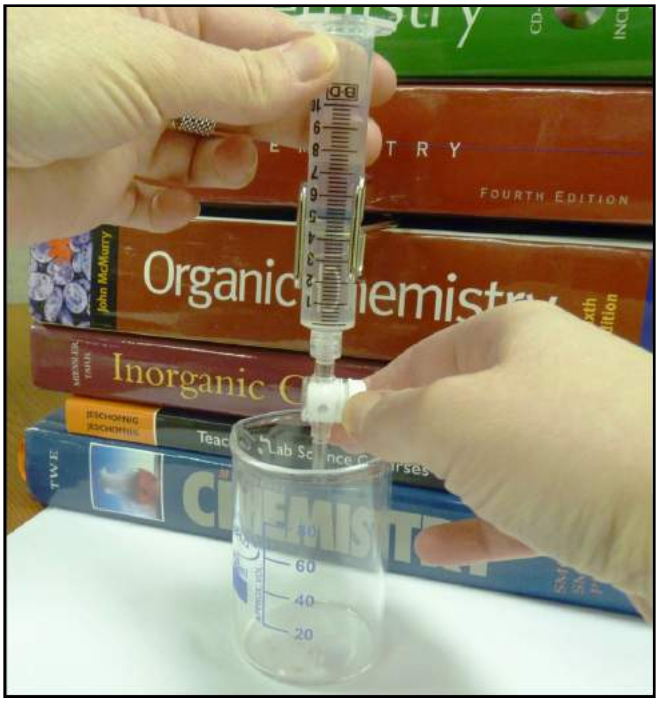

Figure 7. Proper hand positioning for titration. When the open circle is facing you, the titrator is closed, when the open circle is directly under the titrator spout, the titrator is open and liquid will flow.

8. When you are comfortable using the titrator, pour the water in the beaker down the drain, remove the titrator from the test tube clamp, and remove the stopcock from the titrator. Thoroughly dry each of these 3 items with paper towels.

9. When all items are **completely** dry, reassemble the titration setup, as shown in Figure 6.

10. Put on your safety gloves and goggles.

11. With the stopcock in the closed position, fill the titrator with 9 - 10 mL of the 0.5M NaOH.

12. Move the beaker away from the titrator and place a crumpled paper towel directly below the titrator.

13. Using the stopcock, allow a few drops of the NaOH to flow through the titrator into the paper towel. This will fill the tip of the titrator with NaOH solution and remove any air bubbles from the titrator.

14. Place the paper towel with the NaOH drops into the trash and reposition the clean, dry 100 mL beaker back in the titration setup, under the titrator.

15. Use the graduated cylinder to measure exactly 5 mL of vinegar.

16. Pour the 5 mL of vinegar into the completely dry 100 mL beaker.

17. Cut off the tip of the pipet with scissors and add 2 drops of phenolphthalein to the 5 mL of vinegar in the beaker.

18. Carefully swirl the mixture in the beaker to ensure that the indicator is incorporated into the vinegar; the solution will be colorless and clear.

19. Read the volume of NaOH in the titrator and record in **Data Table 1** of your **Lab Report Assistant** under “Initial NaOH Volume (mL),” next to “Trial 1”.

20. Open the stopcock and add 1 drop of NaOH to the colorless and clear vinegar sample in the beaker. After the drop is added, gently swirl the beaker and observe the color for 5 seconds.

Note: *It is important to add the NaOH 1 drop at a time to avoid overshooting the titration.*

21. Continue adding NaOH to the beaker, 1 drop at a time, swirling and observing after each drop until the color changes to a bright-pink color for at least 5 seconds. See Figure 8.

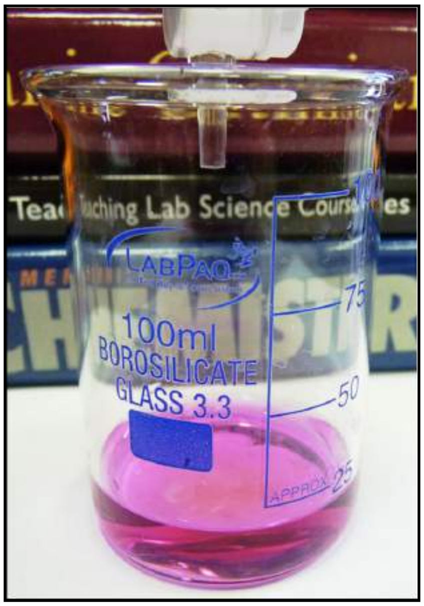

Figure 8. Endpoint of titration. The titration is complete when the color changes to a bright pink for at least 5 seconds.

22. Read the volume of the NaOH solution remaining in the titrator and record this volume in **Data Table 1** under “Final NaOH Volume (mL)”, next to “Trial 1”.

23. Determine the total volume of NaOH used by subtracting the final NaOH volume from the initial NaOH volume and record the total volume in **Data Table 1**.

24. Leave the titrator assembly intact. You will need it for future titrations in this experiment.

25. Pour the contents of the beaker down the drain and flush the drain with water. Thoroughly wash the beaker with soap and water to remove all of the vinegar/NaOH/indicator solution from the beaker. When the beaker is clean, rinse the beaker with distilled water and then thoroughly dry.

26. Place the clean and thoroughly dry beaker back in the titration setup, under the titrator.

27. If necessary, add more NaOH to the titrator.

Note: *It is only necessary to add more NaOH to the titrator if there is less than 1 mL more than the total volume of NaOH used in the previous trial. For example, if the total volume of NaOH used in Trial 1 was 2.1 mL, then there needs to be at least 3.1 mL of NaOH in the titrator.*

28. Repeat Steps 15 through 27 two additional times (Trial 2 and Trial 3), using the vinegar provided in the lab kit.

29. Average the results from the three trials and record in **Data Table 1** and **Data Table 2** of your **Lab Report Assistant**.

30. Using the following equation, determine the average concentration (moles per liter) of acetic acid (CH_3COOH) present in your vinegar. Record the concentration in **Data Table 2**.

$$\text{M } CH_3COOH = \text{L NaOH} \times \frac{\text{mol NaOH}}{\text{L}} \times \frac{1 \text{ mol } CH_3COOH}{1 \text{ mol NaOH}} \times \frac{1}{0.005 \text{ L}}$$

31. Using the following equation, determine % concentration (mass percent) of acetic acid (CH_3COOH) in the vinegar and record it in **Data Table 2**.

Note: *The density of vinegar is 1.00 g/ml and the molecular mass of vinegar is 60.05 g/mol.*

$$\% \ CH_3COOH = \frac{\text{mol } CH_3COOH}{1 \text{ L vinegar}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1 \text{ mL vinegar}}{1.00 \text{ g vinegar}} \times \frac{60.0 \text{ g } CH_3COOH}{1 \text{ mol } CH_3COOH} \times 100\%$$

Cleanup:

32. Properly dispose of remaining chemicals.

33. Wash and dry all equipment and return to the lab kit.

Questions

A. The manufacturer of the vinegar used in the experiment stated that the vinegar contained 5.0% acetic acid. What is the percent error between your result and the manufacturer's statement?

B. What challenges would you encounter with the titration if you had used apple cider vinegar or balsamic vinegar as the analyte instead of white vinegar?

C. How would your results have differed if the tip of the titrator was not filled with NaOH before the initial volume reading was recorded? Explain your answer.

D. How would your results have differed if you had over-titrated (added drops of NaOH to the analyte beyond the stoichiometric equivalence point)?

E. If a 7.0 mL sample of vinegar was titrated to the stoichiometric equivalence point with 7.5 mL of 1.5M NaOH, what is the mass percent of CH_3COOH in the vinegar sample?

F. Why is it important to do multiple trials of a titration, instead of only one trial?

