Determination of Acetic Acid in Vinegar

The principal component of vinegar besides water is the weak acid acetic acid, HC₂H₃O₂ (Kₐ = 1.78 x 10⁻⁵). In this experiment you will determine the concentration of acetic acid in a vinegar sample by titrating the acetic acid with the strong base sodium hydroxide. The stoichiometry of the neutralization reaction is as follows:

Molecular:  \[ \text{HC}_2\text{H}_3\text{O}_2(aq) + \text{NaOH}(aq) \rightarrow \text{NaC}_2\text{H}_3\text{O}_2(aq) + \text{H}_2\text{O} \] (1)

Net Ionic: \[ \text{HC}_2\text{H}_3\text{O}_2(aq) + \text{OH}^- (aq) \rightarrow \text{C}_2\text{H}_3\text{O}_2^- (aq) + \text{H}_2\text{O} \] (2)

The equivalence point of the titration is the condition when the initial number of moles of H⁺ from the acid has been exactly neutralized by an equal number of moles of OH⁻ ion from the sodium hydroxide:

\[ \# \text{ of mols H}^+ = \# \text{ of mols OH}^- \] (3)

Thus, the only species present at the equivalence point are the products sodium acetate and water. However, salts of weak acids contain the conjugate base of the acid and undergo hydrolysis, or reaction with water (Zumdahl, sections 14.8 and 15.4):

\[ \text{C}_2\text{H}_3\text{O}_2^- (aq) + \text{H}_2\text{O} \rightarrow \text{HC}_2\text{H}_3\text{O}_2(aq) + \text{OH}^- (aq) \] (4)

Although this equilibrium lies largely to the left side, a small amount of hydroxide ion is formed. Therefore, at the equivalence point of a weak acid-strong base titration the solution is basic. This is in contrast to strong acid-strong base titrations, which produce a neutral solution at the equivalence point. Consequently, an indicator must be chosen which changes color in the basic pH region (Zumdahl, section 15.5). Either phenolphthalein (colorless in acid, red in base) or thymol blue (yellow in acid, blue in base) may be used. Both indicators change color in the range ca. pH 8.0-9.6.

Procedure

1. Obtain about 35 mL of a vinegar sample in a clean, dry beaker and record the identification or code number. This sample must be diluted before titration.
2. Rinse a clean 25.00-mL pipet with a small portion of the vinegar and discard.
3. Pipet 25.00 mL of vinegar into a clean 250.0-mL volumetric flask, dilute to the mark with deionized water, and mix well.
4. Rinse your pipet with a small portion of this diluted solution and discard.
5. Pipet 25.00-mL aliquots (samples of precisely known volume) of the diluted vinegar into each of three 250-mL Erlenmeyer flasks.
6. To each flask add 50 mL of deionized water and 3 drops of indicator solution and swirl well to mix.
7. Obtain about 100 mL of standard sodium hydroxide solution in a clean, dry beaker. Record the molarity of the solution.
8. Rinse a clean burette with a small portion of standard NaOH solution, then fill the burette to a point near the zero mark.
9. Remove the burette funnel before titrating. Slowly drain enough of the NaOH solution into a waste beaker to ensure that the tip of the burette is filled and that no air bubbles are present in the tip. Record the initial volume reading to the nearest 0.01 mL.

10. Titrate the first sample with good mixing to the indicator endpoint. The color change should persist for 30 seconds (it will eventually fade). Use the split-drop technique to add very small increments as you get close to the endpoint (the color lingers for a while before fading) and rinse down the sides of the flask with your wash bottle just before you think the endpoint will be reached.

11. Record the final burette reading to the nearest 0.01 mL. As you refill the burette for subsequent titrations, use no more of the NaOH solution than is necessary! When finished, clean up all glassware and your bench space.

12. Calculate the molarity of each of your samples of diluted vinegar and then calculate the average molarity and relative average deviation. By taking the dilution into account, calculate the molarity of the original, undiluted vinegar.

13. Finally, given that the molar mass of acetic acid is 60.05 g/mol, calculate the number of grams of acetic acid contained in 100.0 mL of the original undiluted vinegar; this number is commonly referred to as the “percent” concentration of the vinegar.

14. Obtain from the instructor the correct value for the molarity of the original vinegar (undiluted) and calculate the percent error in your result.

15. Include a detailed discussion of systematic errors in your discussion of results. In particular, you should focus upon those errors which would make the result too high if yours was too high, or errors which would lead to a low result if yours was too low. In addition, include answers to the following questions in your report.
Beginning questions & ideas: What do you want to know or already know? (6 pts)

Tests & Procedures: What did you do to answer your questions or prove your idea? (4 pt)

Data & Observations: What did you observe from each test and procedure? (10 pts)
Claims: What are the answers to your questions or the ideas that you claimed? (4 pts)

Evidence: How do your data and observations support your claims? Be specific and show calculations as necessary. (18 pts)
Reflection: Explain any discrepancies between your results and what was expected. What theories or principles helped you to complete this experiment? (3 pts)

Additional Questions (5 pts)

For the following conditions, explain how the resultant molarity of acetic acid would be affected; be specific in explaining your answer.

- If the tip of the burette was not filled with sodium hydroxide before the initial volume reading was recorded.

- If a few drops splashed out of the Erlenmeyer flask during the titration.

- If the last drop of the vinegar solution were blown out of the pipet into the Erlenmeyer flask.

- If the volume of water added to the Erlenmeyer flask was slightly larger than 50 mL.

- Why does the exact volume of water added to the titration flask not need to be known?