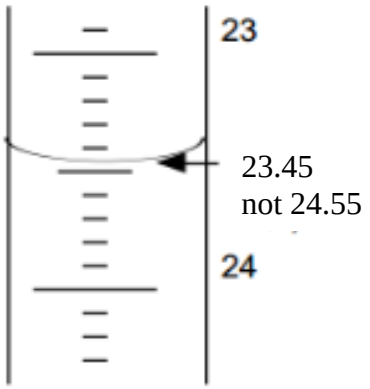
Chemistry Lab 28: Concentration of Acetic Acid in Commercial Vinegar (Titration)

Introduction: The neutralization of an acid by a base can be done using titration. In titration, a solution of known acidity, or pH, is added gradually to a basic solution of unknown pH. When the unknown solution is exactly neutralized, as shown by the color of an acid-base indicator or by the reading of a pH meter, it is said that the number of hydronium ions equals the number of hydroxide ions. *moles* H_3O^+ = *moles* OH^- In titration, the solutions are dispensed from burettes. The volume used of each solution is calculated by subtracting the volume read before the titration from the volume read after the titration. The volume in a buret can be read accurately to +/- 0.01 mL. Before the titration of the unknown can be done the titrant (solution being used for the titration) must standardized (calculating the accurate concentration). A solution whose concentration is known to a high degree of accuracy is known as a standard solution. The purpose of this lab is to use solid $KHC_8H_4O_4$, potassium hydrogen phthalate (KHP), to prepare a very accurate solution using a titration. KHP is a monoprotic acid. This solution (titrant) will be used to standardize a solution of sodium hydroxide. This standardized solution of sodium hydroxide will then be used to titrate vinegar (analyte) to find the molarity of acetic acid in commercial vinegar.



P1. Write the equation for the dissociation of acetic acid in aqueous solution. Identify the conjugate base and the relative strength of that base.

P2. Write the equation for the dissociation of NaOH in aqueous solution. Identify the conjugate acid and the relative strength of that acid.

P3. Sketch a titration curve to plot the pH of the analyte solution (acetic acid) versus the volume of the titrant (sodium hydroxide) added. Label the equivalence point on the graph, which should be

Procedure: Part A. Standardization of Solution of Sodium Hydroxide

1. Make 150 mL of a 0.1 M solution of NaOH. Use distilled water, prepare, seal, and store this in an Erlenmeyer flask.
2. Clean a 50 mL buret thoroughly with distilled water. Mount the buret vertically in a buret clamp attached to a ring stand. Place a white sheet of paper beneath the buret.
3. Rinse the buret with 5 mL of the 0.1 M NaOH solution. (This step is done so that any water that might be in the buret prior to the experiment will not dilute your NaOH during the titration.) Fill the buret with 0.1 M NaOH.
4. Get approximately 0.5 grams of $KHC_8H_4O_4$ into a ~125 mL Erlenmeyer flask. Make sure that you know the exact mass of your solid acid.
5. Dissolve the $KHC_8H_4O_4$ into about 50 mL of water and add 2-3 drops of phenolphthalein. Make sure that it is completely dissolved.
6. Now, slowly titrate the NaOH into the KHP, swirling at all times, until the solution turns the lightest shade of pink possible. The color needs to stay a pale pink for 15-30 seconds. This is the endpoint.
7. Do one "quick & dirty" titration to get an approximate amount.

8. Record all information so that you know exactly how much NaOH was added to the flask.
9. Do at least three trials, with usable data, to insure accurate results. **Rinse out your buret three times with tap water after completion of this experiment and open the stopcock.**
10. Calculate the molarity of the exact NaOH solution using the mass of the $KHC_8H_4O_4$. Use stoichiometry to do this with the end result being a "standardized" NaOH solution. **SAVE** this solution for part B.

Part B: Determination of the Molarity of Acetic Acid in Vinegar

1. Pipette 20 mL of commercial vinegar into a 200 mL graduated cylinder (or 25 mL into a 250 mL) and dilute with distilled water. Transfer this solution to an Erlenmeyer and label it 10% vinegar solution.
2. Pipette 25 mL of the 10% vinegar to a flask. Add 2 drops phenolphthalein indicator. Titrate the vinegar with your standardized NaOH from part A.
3. Do one "quick & dirty" titration to get an approximate amount and then 3 trials. Use a pH digital probe to find the pH of your sample at the endpoint for each trial? **Rinse out your buret with tap water after completion of this experiment and then open the stopcock.**
4. Calculate the molarity of the 10% solution of vinegar.
5. Multiply by 10 to get the concentration of commercial vinegar (show all work).
6. Calculate the "accepted value" of the commercial vinegar.
7. Determine the percentage error.
8. Discuss at least three sources of systematic and/or random error.
9. Write your conclusion and 'self-assess" your lab using the lab rubric checklist (click here)

Acid Base Lab #2 Data Sheet (used to guide you through page 1 of this lab, if needed)

Data:

Part A: (Numbering corresponds to procedural steps)

1. Calculate **volume** of ___M NaOH needed to make 150ml of 0.1M NaOH. (The molarity is the stock solution concentration)

   $$\mathbf{M_1V_1=M_2V_2}$$

   Then dilute to $\mathbf{V_1}$ 150ml in Erlenmeyer flask.

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	Trial 1	Trial 2	Trial 3
Volume NaOH Endpoint (ml)			
Mass $KHC_8H_4O_4$ (g)			

11. Calculate molarity of NaOH from mass of $KHC_8H_4O_4$. (The equation is given below)

$$NaOH + KHC_8H_4O_4 \rightarrow NaKC_8H_4O_4 + HOH$$

Trial 1 =

Trial 2 =

Trial 3 =

Average of 3 trials = _____
