Acid-Base Titration

EXPERIMENT 20A

Titration is a laboratory technique that can be used to determine the concentration of certain substances. A standard solution of known molarity is titrated against (reacted with) a solution of unknown concentration. An indicator can signal the completion of the reaction and the concentration can be quantitatively determined.

Acid-base titrations involve the neutralization reaction between the hydronium and hydroxide ions. These ions combine to form the neutral water molecule:

$$H_3O^+(aq) + OH^-(aq) \rightleftarrows 2H_2O(l)$$

When equal amounts of hydronium and hydroxide ions are present within the titration flask, the equivalence point is achieved. The endpoint in acid-base titrations is accompanied by a rapid change in solution pH. This shift may be recognized by its effect upon an appropriate chemical indicator.

Indicators are weak acids and bases whose colors are dependent upon solution pH. Some indicators are naturally occurring pigments found within plant leaves.

In this experiment, you will extract a chemical indicator from red cabbage leaves. You will test acidic and basic solutions with the indicator and record the color changes.

You will then titrate a sample of vinegar against a sodium hydroxide solution of known concentration. Vinegar is a solution of acetic acid in water. Using the titration technique, you can determine the concentration of acetic acid in a commercial vinegar solution. The equivalence point of the titration will be recognized by its action upon the cabbage juice indicator. The equation for this neutralization reaction may be written as follows:

$$CH_3COOH(aq) + NaOH(aq) \rightleftarrows CH_3COONa(aq) + H_2O(l)$$

In Part III of the experiment, household ammonia will be titrated against a HCl standard. The reaction may be illustrated as follows:

$$NH_3(aq) + H_3O^+(aq) \rightleftarrows NH_4^+(aq) + H_2O(l)$$

The concentration of NH_3 in household ammonia can be determined from the titration data.

OBJECTIVES

1. to extract an indicator from red cabbage leaves
2. to observe the macroscopic properties of the indicator in acidic and basic environments
3. to titrate an acetic acid solution (vinegar) with standardized 0.5*M* NaOH
4. to titrate a household ammonia solution with standardized 0.5*M* HCl
5. to utilize the titration data and calculate the molarities and percent composition of the vinegar and ammonia solutions

MATERIALS

Apparatus

beaker (600-mL)
beaker (250-mL)
wire gauze
laboratory burner
ring stand and ring
5 small test tubes
test tube rack
suction bulb
2 delivery pipets
2 Erlenmeyer flasks (250-mL)
2 burets
1 buret clamp
graduated cylinders (10-mL and 50-mL)
lab apron
safety goggles
beaker tongs

Reagents

red cabbage leaves
distilled water
0.5*M* NaOH
0.5*M* HCl
white vinegar
household ammonia solution

PRELAB

Answer questions 1-6 on the Report Sheet.

PROCEDURE

Part I

1. Put on your lab apron and safety goggles.
2. Place several small pieces of red cabbage leaf into the bottom of the 600-mL beaker.
3. Add approximately 150 mL of distilled water to the beaker. If necessary, press the leaf pieces against the bottom of the beaker to be sure they are submerged.

CAUTION: Handle the hot beaker with care.

CAUTION: The hydrochloric acid solution is corrosive to skin, eyes, and clothing. Wear safety goggles and lab apron when handling the acid in this experiment. Wash any spills and splashes off your skin and clothing immediately with plenty of water. Call your teacher.

CAUTION: The vinegar solution is mildly irritating. Keep it off your skin and out of your eyes. Wash any spills and splashes immediately with plenty of water.

4. Place the beaker on a ring stand with a ring and wire gauze. Use the laboratory burner to heat the system. Boil the cabbage mixture until the water acquires a deep purple color. Remove the beaker from the heat source and allow it to cool for several minutes.
5. Pour approximately 100 mL of the cabbage extract into a 250-mL beaker. Set aside for Parts II and III of this experiment.
6. Measure 5 mL quantities of the remaining cabbage extract to be poured into 5 separate test tubes labeled *A* to *E*.
7. Test tube *A* serves as a control. To test tube *B*, add one drop of standardized 0.5*M* NaOH.
8. To test tube *C*, add 1 drop of standardized 0.5*M* HCl.
9. To test tube *D*, add 1 drop of vinegar.
10. To test tube *E*, add 1 drop of household ammonia.
11. Record the colors in tubes *A* through *E* in Table 1 of the Report Sheet.

Part II

CAUTION: The household ammonia solution is mildly corrosive. Keep it off your skin and out of your eyes. Wash any spills or splashes immediately with plenty of water. Call your teacher.

CAUTION: The NaOH solution is mildly corrosive. Keep it off your skin and out of your eyes. Wash any spills or splashes immediately with plenty of water. Call your teacher.

1. Using a suction bulb, pipet 5.0 mL of white vinegar into a 250-mL Erlenmeyer flask.
2. Add 20 mL of the cabbage extract obtained in Part I in the flask. Record the color of the indicator.
3. Rinse a clean buret with approximately 15 mL of the standardized NaOH solution. Drain the buret and refill with standardized NaOH solution. Record the initial volume of the buret.
4. Gradually dispense some of the standardized NaOH solution into the titration flask. Swirl the flask constantly. See Figure 20A-1. Continue adding NaOH, slowly noting any changes in the flask.
5. As the equivalence point is approached, the green color will dissipate more slowly. Now add the NaOH dropwise. Stop the titration once the addition of a single drop causes the solution to remain green for 30 seconds. Record the volume of NaOH needed to reach the equivalence point in Table 2 of the Report Sheet.
6. Repeat Steps 1 through 5 using a second 5.0-mL sample of vinegar.

Part III

1. Using a suction bulb, pipet 5.0 mL of a household ammonia solution into a 250-mL Erlenmeyer flask.
2. Add 20 mL of the cabbage extract obtained in Part I.
3. Rinse a clean buret with approximately 15 mL of the standardized HCl solution. Drain the buret and refill with standardized HCl solution.
4. Gradually dispense the HCl solution into the titration flask. Continuously swirl the flask.
5. As the equivalence point is approached, the solution will turn purple. Now add the HCl dropwise. Stop the titration once the addition of a single drop causes the solution to remain pink for 30 seconds. Record the volume of HCl needed to reach this equivalence point in Table 3 of the Report Sheet.
6. Repeat Steps 1 through 5 using a second 5.0-mL sample of the ammonia solution.
7. Review the Post Lab Discussion. Use this information and your textbook to complete Tables 2 and 3 and answer the questions on the Report Sheet.
8. Before leaving the laboratory, wash your hands thoroughly with soap and water; use a fingernail brush to clean under your fingernails.

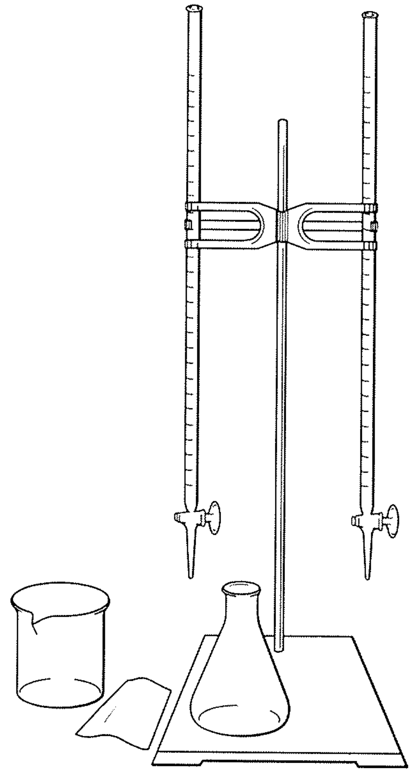

Figure 20A-1

POST LAB DISCUSSION

During the vinegar titration, the hydroxide ions liberated from the standardized NaOH solution reacted with the ionizable protons of the acetic acid to form neutral water molecules. When the concentration of both ion species approached equivalence, the titration end point was reached. Since the molarity and volume of standardized NaOH was known, the total number of reactant moles can be calculated as shown in the following equation:

$$(\text{Volume}_{NaOH})\ (\text{Molarity}_{NaOH}) = \text{Reactant Moles}_{NaOH}$$
$$= \text{Reactant Moles}_{CH_3COOH}$$

Ammonia is a weak base that is neutralized by its reaction with titrating HCl solution. When the concentrations of hydroxide and hydronium ions are equal, the molarity of the unknown ammonia solution can be quantitatively determined.

$$(\text{Volume}_{HCl})\ (\text{Molarity}_{HCl}) = \text{Reactant Moles}_{HCl}$$
$$= \text{Reactant Moles}_{NH_3}$$

The percent composition of CH_3COOH in vinegar or NH_3 in household ammonia can be calculated as follows:

$$\text{percent composition} = \frac{\text{mass solute}}{\text{mass solution}}$$

The mass of each solute is equal to the product of its reactant moles and the molar mass. The mass of solution may be determined by multiplying the solution density by the sample volume. Assume the density of both the vinegar and ammonia solutions to be 1.00 g/mL.

Acid-Base Titration

Name ______________________
Class __________ Date __________

PRELAB QUESTIONS

1. Write the equation for the neutralization reaction between ammonia and HCl. ______________________

2. Write the equation for the neutralization reaction between acetic acid and sodium hydroxide. ______________________

3. Why should a minimal amount of water be used to cover the red cabbage leaves in Part I? ______________________

4. What is the function of test tube *A* in Part I? ______________________

5. When performing the titrations, why should the Erlenmeyer flask be constantly swirled? ______________________

6. Why were the titrations in Part II and Part III repeated? ______________________

OBSERVATIONS AND DATA

Part I

	REAGENTS ADDED	OBSERVATION
Tube *A*	none	
Tube *B*	1 drop 0.5*M* NaOH	
Tube *C*	1 drop 0.5*M* HCl	
Tube *D*	1 drop vinegar	
Tube *E*	1 drop household ammonia	

Part II

	TRIAL 1	TRIAL 2
Volume of vinegar		
Molarity of NaOH		
Volume of NaOH (initial buret reading)		
Volume of NaOH (final buret reading)		
Volume of NaOH dispensed		
Moles of NaOH needed to neutralize sample		
Moles of acetic acid in vinegar		
Molar mass of acetic acid		

Mass of acetic acid in vinegar		
Mass of vinegar sample (density 1 g/mL)		
Percent composition of acetic acid in vinegar		

Part III

	TRIAL 1	TRIAL 2
Volume of ammonia		
Molarity of HCl		
Volume of HCl (initial buret reading)		
Volume of HCl (final buret reading)		
Volume of HCl dispensed		
Moles of HCl needed to neutralize sample		
Moles of NH_3 in household ammonia		
Molar mass of NH_3		
Mass of NH_3 in household ammonia		
Mass of household ammonia sample (density 1 g/mL)		
Percent composition of NH_3 in household ammonia		

CONCLUSIONS

1. Would you use an acid or base as a standard when titrating against a solution of soda pop? Why?

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2. How would the results in Part III differ if a polyprotic acid such as H_2SO_4 were used instead of the standardized HCl solution? ____

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3. When selecting an appropriate chemical indicator, one should choose an indicator that will undergo a shift to a darker color. Why? ____

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SYNTHESIS

1. When solutions approach the titration equivalence point, they demonstrate an increased sensitivity to hydronium ion equilibrium. Why? ____

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2. Describe how you could use the red cabbage extract as a quantitative method for determining solution pH. ____

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