EXPERIMENT 20C

Acid-Base Titration

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 aqueous hydrogen and hydroxide ions. These ions combine to form the neutral water molecule:

$$H^+(aq) + OH^-(aq) \rightleftharpoons H_2O(l)$$

The indicator phenolphthalein will be used to show when the number of moles of acid equals the number of moles of base. This point is called the equivalence point.

Titration is one of the most common operations performed by the chemist. We all depend upon chemical analysis, and it is with this branch of chemistry that the average citizen is most likely to come into contact. Decisions involving huge sums of money, or even life and death, depend upon the accuracy and speed of chemical analysis, whether in hospital lab testing, environmental pollution monitoring, or crime detection.

OBJECTIVES

1. to titrate a hydrochloric acid solution of unknown concentration with standardized 0.50*M* sodium hydroxide
2. to titrate an acetic acid solution (vinegar) with standardized 0.50*M* sodium hydroxide
3. to utilize the titration data to calculate the molarity of the hydrochloric acid, and the molarity and percentage composition of the vinegar

MATERIALS

Apparatus

suction bulb
2 delivery pipets (10 mL)
2 Erlenmeyer flasks (250 mL)
1 buret
1 buret clamp and stand
lab apron
safety goggles

Reagents

standardized NaOH solution (approx. 0.50*M*)
unknown HCl solution (approx. 0.50*M*)
white vinegar
phenolphthalein solution

PROCEDURE

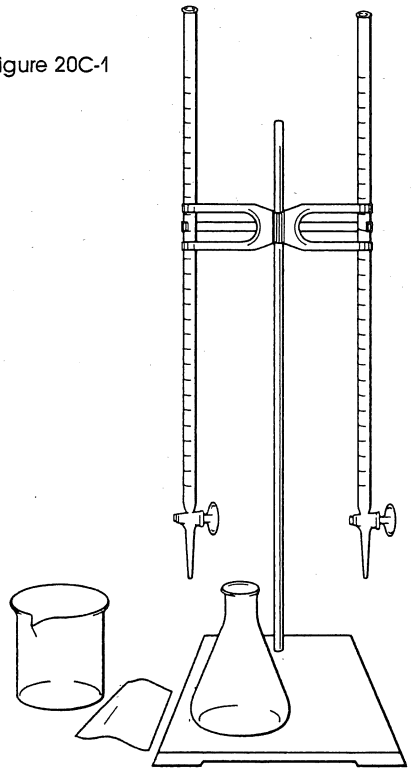
Part I Determination of Molarity of Hydrochloric Acid Solution

1. Put on your lab apron and safety goggles.
2. Obtain about 50 mL of the hydrochloric acid solution of unknown concentration and about 100 mL of the standardized NaOH solution. Your teacher will provide you with the exact molarity of the NaOH. Record this value in your copy of Table 1 in your notebook.
3. Using a suction bulb, pipet 10.0 mL of the HCl solution into a 250 mL Erlenmeyer flask, after rinsing your pipet with a small amount of HCl first.
4. Add 3 drops of phenolphthalein solution.
5. 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 NaOH in the buret in your copy of Table 1 in your notebook.
6. Gradually dispense some of the standardized NaOH solution into the titration flask. Swirl the flask constantly. (See Figure 20C-1.) Continue adding NaOH, noting any changes in the flask..

CAUTION: The NaOH solution is corrosive to skin, eyes, and clothing. Wash any spills or splashes immediately with plenty of water.

CAUTION: Phenolphthalein is poisonous and flammable. Do not get any in your mouth; keep well away from flame.

Figure 20C-1



7. As the equivalence point is approached, a pinkish color will appear, and dissipate more slowly as the titration proceeds. Now add the NaOH drop by drop. Stop the titration when the addition of a single drop causes the solution to remain pinkish for 30 s. Record the volume of NaOH needed to reach the equivalence point in your copy of Table 1. The most accurate reading is one in which the solution is the faintest possible pink, but still remaining that color.

8. Repeat steps 3 through 7 using a second 10.0 mL sample of the HCl. Knowing the volume obtained in your first titration, you can be extra careful when you are within 1 mL of the previous value, and add the NaOH a drop at a time, shaking after each drop. This lessens the likelihood of your overshooting the mark.
9. If the two values differ widely, it would be a good idea to do one more titration if you have time.

Part II Determination of Percentage Composition of Vinegar

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

1. Obtain approximately 30 mL of white vinegar (acetic acid solution).
2. Using the same buret of NaOH as was used in Part I, do more titrations, but this time use 10.0 mL portions of vinegar instead of HCl. Follow exactly the same procedure as in Part I (Procedures 3 to 9). Record your observations in your copy of Table 2.

REAGENT DISPOSAL

Mix any leftover acids and bases together to neutralize, and pour down the sink with plenty of water. Do not return any solutions to their original containers.

POST LAB DISCUSSION

During the HCl titration, the hydroxide ions liberated from the standardized NaOH solution reacted in a 1-to-1 ratio with the H^+ ions from HCl to form neutral water molecules. When the concentration of both ion species were the same, the equivalence point was reached. Since the molarity and volume of the standardized NaOH is known, the total number of moles reacting can be calculated as shown in the following equation:

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

Knowing the volume of HCl used originally, the molarity of the HCl can be calculated from the formula:

$$(\text{Molarity}_{HCl}) = \text{Moles}_{HCl} \div \text{Volume of HCl (in L)}$$

In Part II, calculate the molarity of the acetic acid in vinegar in the same manner as shown for HCl. In addition, calculate the percentage composition of the vinegar. This is given by:

$$\text{Percentage composition} = \frac{\text{mass solute}}{\text{mass solution}} \times 100\%$$

The mass of the acetic acid (CH_3COOH) is obtained from the number of moles times the molar mass. The mass of the solution is obtained from the measured volume of the solution times its density. For this experiment, you may assume that the density of the vinegar is 1.00 g/mL.