

Chemistry 12

Lab 20C Acid-Base Titration

NAME _____

Partner _____

DATE _____ BLOCK _____

Pre-Lab Questions:

1. Write the formula equation for the neutralization reaction between hydrochloric acid and sodium hydroxide solution. Include phase subscripts.

2. Write the corresponding net ionic equation for the reaction in question #1.

3. Write the net ionic equation for the reaction between acetic acid and sodium hydroxide solution.

4. When performing titrations, why is it not necessary to wash **and dry** the flask in which the titration will take place? (In other words, excess water in the flask doesn't matter.)

5. What is the function of the indicator phenolphthalein in the titration?

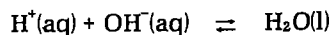
6. The point during a titration at which the number of moles of H^+ equals the number of moles of OH^- is called what?

7. It takes 37.54 mL of 0.200 M NaOH to neutralize 25.00 mL of unknown M HCl. Showing your work, calculate the molarity of the acid.

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:



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 0.50M sodium hydroxide
2. to titrate an acetic acid solution (vinegar) with standardized 0.50M 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

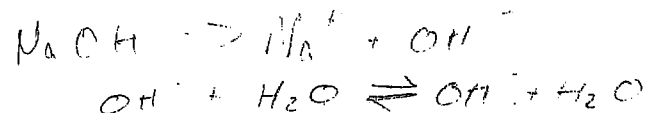
MATERIALS

Apparatus

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

Reagents

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



PROCEDURE

Part I Determination of Molarity of Hydrochloric Acid Solution

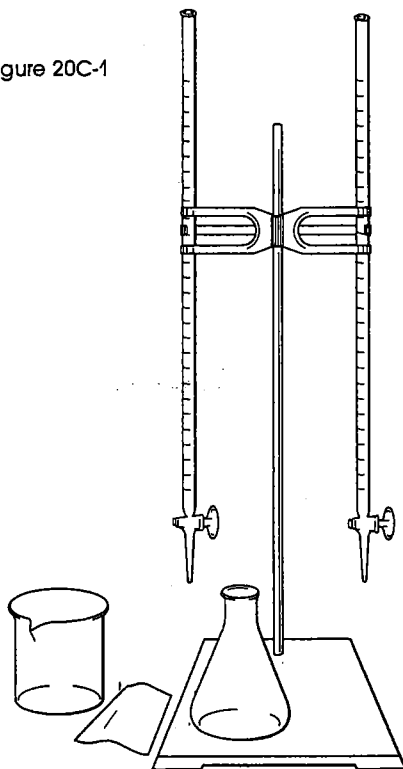
1. Put on your lab apron and safety goggles.
2. Obtain about 50 mL of the hydrochloric acid solution of unknown concentration and about 100 mL of the standardized NaOH solution. Your teacher will provide you with the exact molarity of the NaOH. Record this value in your copy of Table 1 in your notebook.
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.

- 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.
- 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.

- Obtain approximately 30 mL of white vinegar (acetic acid solution).
- 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:

$$\begin{aligned} (\text{Volume}_{\text{NaOH}}) (\text{Molarity}_{\text{NaOH}}) &= \text{Reactant Moles}_{\text{NaOH}} \\ &= \text{Reactant Moles}_{\text{HCl}} \end{aligned}$$

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

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

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 (CH₃COOH) 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.

Experimental Procedure:

Follow the procedure as given on pages 237 and 238.

Your results of the titrations should agree to within 0.20 mL. If they don't, repeat the titration until you get results that agree to within this limit.

Table 1: Molarity of NaOH = _____ M

	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
buret reading before titration					
buret reading after titration					
volume of NaOH used					
Average volume NaOH		← (average your two closest volumes)			

1. Calculate the moles of NaOH from the average volume used in Part 1 and the given molarity.

2. What is the ratio of OH^- in NaOH to H^+ in HCl in the balanced neutralization reaction?

3. Calculate the molarity of the unknown HCl solution.

Table 2: Molarity of NaOH = _____ M

	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
buret reading before titration					
buret reading after titration					
volume of NaOH used					
Average volume NaOH		← (average your two closest volumes)			

1. Calculate the moles of NaOH from the average volume used in Part 1 and the given molarity.
2. What is the ratio of OH^- in NaOH to H^+ in CH_3COOH in the balanced neutralization reaction?

3. Calculate the molarity of the vinegar sample.
4. Calculate the mass of acetic acid in 1 L of solution.
(Refer to the Post Lab discussion on page 238.)
5. Calculate the percentage of acetic acid in the vinegar.
6. By law, vinegar must be not less than 4% by mass acetic acid. Did your sample meet this specification?