EXPERIMENT 22

Determination of a Chemical Equation

Performance Goals

22–1 By dilution, prepare a known molarity solution of sulfuric acid and sodium hydroxide.

22–2 Using the dilute solutions, titrate a given volume of acid with the base.

22–3 Based on the titration data, calculate the mole ratio of base to acid and write the chemical equation corresponding to this mole ratio.

CHEMICAL OVERVIEW

The concept of molarity and the method of titration are discussed in detail in the introduction to Experiments 20 and 21 on pages 267–68. Please read this material carefully before beginning this experiment.

In this experiment you will be supplied two ''concentrated'' solutions of known molarity, from which you will prepare, by dilution, solutions that you will use in the titration. You will then calculate the molarity of the dilute solutions (see Example 1 in the Sample Calculations).

To perform the titration, first you will carefully pipet a given volume of sulfuric acid into three Erlenmeyer flasks. Then, you will titrate each sample with the dilute sodium hydroxide. A minimum of three titrations will be carried out. Since all acid samples have the same volume, the amount of sodium hydroxide should be nearly the same. If the three runs are not within the range of 0.2 mL, additional titrations will have to be performed.

Using the volumes and molarity of the sulfuric acid and sodium hydroxide, you will calculate the number of moles contained in each. Using these values, the mole ratio can be calculated (see Example 2 in the Sample Calculations). *Based on this ratio*, you will then write the chemical equation representing the reaction that has taken place.

SAMPLE CALCULATIONS

Example 1

15.0 mL of a 3.12 M solution is diluted to 250.0 mL. Calculate the molarity of the dilute solution.

The number of moles of solute is the same in both solutions, since only water was added. This can be expressed mathematically by the equation

$$M_1V_1 = M_2V_2$$

where M_1 and V_1 are the molarity and volume of the concentrated solution, respectively, and M_2 and V_2 those of the dilute solution. Using the equation above and solving for M_2, we get

$$M_2 = \frac{M_1V_1}{V_2}; \qquad M_2 = \frac{3.12 \text{ moles}}{\text{L}} \times \frac{15.0 \text{ mL}}{250.0 \text{ mL}} = 0.187 \text{ M}$$

Note that we do not necessarily have to convert the volumes to liters, since the volume ratio will remain the same regardless of what units are used—as long as they are the same!

Example 2

In a titration, 15.0 mL of 0.152 M sulfuric acid required 31.5 mL of 0.145 M sodium hydroxide to reach the end point. Calculate the mole ratio of base to acid in this titration.

First, calculate the number of moles of acid and base, using the relationship

$$\text{moles} = M \times V; \qquad \text{moles} = (\text{moles/L}) \times \text{L}$$

Note, in this case you *must* convert the volume to liters!

$$\text{moles } H_2SO_4 = \frac{0.152 \text{ mole}}{\text{L}} \times 0.0150 \text{ L} = 0.00228 \text{ mole}$$

$$\text{moles NaOH} = \frac{0.145 \text{ mole}}{\text{L}} \times 0.0315 \text{ L} = 0.00457 \text{ mole}$$

$$\frac{\text{moles NaOH}}{\text{moles } H_2SO_4} = \frac{0.00457}{0.00228} = 2$$

This means that 2 moles of NaOH reacts with 1 mole of H_2SO_4. Your experimental values may not yield an exact whole number; round your ratios to the nearest whole number if you are within 0.15 of a whole number (e.g., 2.89 rounds to 3; 0.95 rounds to 1).

SAFETY PRECAUTIONS AND DISPOSAL METHODS

Both sodium hydroxide and sulfuric acid are very corrosive and harmful substances. If some of the sodium hydroxide solution, either concentrated or dilute, comes in contact with your skin, it will feel slippery. Wash your skin immediately with lots of cold water until the slippery feeling is gone. Sulfuric acid will burn when in contact with skin; rinse the exposed area with plenty of cold water.

Obviously, you must avoid any contact with your eyes. Be sure to wear goggles throughout the experiment and during cleanup. Also, always use a pipeting device, never use mouth suction!

Dispose of the solutions as directed by your instructor.

PROCEDURE

NOTE: Record all volume measurements in milliliters to the nearest 0.1 mL.

1. Preparation of NaOH Solution

A. Obtain a 15.0-mL pipet, a pipet bulb (or similar device), and a 250.0-mL volumetric flask. Rinse the pipet and the flask with deionized water. Next, rinse the pipet (only!) with about 5 mL of the ''concentrated'' NaOH, and discard the rinse. Pipet 15.0 mL of the NaOH into the volumetric flask. Add deionized water from a beaker until the water level is about 1 cm below the etch line on the neck of the flask. Add the remaining water drop by drop using an eye dropper until the *bottom* of the meniscus just touches the line (see Figure LP.4). Be careful, because if you overfill the flask, you will have to start over again! Place the stopper in the flask and mix the solution very thoroughly. Turning the flask upside down, at least five times, while gently shaking it will accomplish this. This step is *extremely* important, because a non-homogeneous solution will never yield proper duplication of titration values.

B. Transfer the diluted NaOH to a *dry*, capped storage bottle. Keep this bottle closed, except when pouring from it, because evaporation loss or absorption of CO_2 will change the concentration.

C. Record the molarity of the ''concentrated'' NaOH and calculate the concentration of the diluted solution. Record your value on the work page.

2. Preparation of H_2SO_4 Solution

A. Rinse the pipet and the volumetric flask with deionized water. Rinse the pipet (only!) with the ''concentrated'' sulfuric acid solution. Follow the procedure in Part 1, except use ''concentrated'' sulfuric acid.

B. Store the diluted acid in a *dry*, 250-mL Erlenmeyer flask.

3. Titration of H_2SO_4 with NaOH

A. Set up three 250-mL Erlenmeyer flasks. Rinse your pipet first with deionized water, then with about 5 mL of your *diluted* H_2SO_4 solution. Discard the rinse. Now, pipet 15.0-mL portions of the acid into each of the Erlenmeyer flasks. Add about 75 mL of deionized water and 3–5 drops of phenolphthalein indicator to each flask.

B. Rinse a 50-mL buret first with deionized water, then with two 10-mL portions of the *diluted* NaOH, discarding the rinse. Always use a funnel to fill the buret. Make sure the tip of the buret is also rinsed. Fill up the buret with the *diluted* NaOH, set the bottom of the meniscus on a line (not necessarily a whole milliliter), and read the volume. Record this value on your work page (this is the initial NaOH). Be sure the tip of the buret is filled and that you don't see a bubble under the stopcock. If there is one, let down some solution from the buret and set the level again.

C. If a magnetic stirring apparatus is available, place a stirring bar into the flask and position the tip of the buret inside the neck of the Erlenmeyer flask (see Figure 20.1). If you do not have magnetic stirring available, you will have to swirl the flask manually throughout the titration.

At the beginning of the titration you may add the base in larger portions, slowing down as the time for the pink color to disappear gets longer. The end of the titration is reached when the pink color persists for 30 seconds. Record the buret reading on your work page (this is the final NaOH).

D. Repeat Step C with the remaining two samples. Be sure to fill up the buret if you judge that not enough solution is available for the next titration. Adjust the level on a whole line (never start between lines!). Remember, the bottom part of the buret is not calibrated!

E. If the three titrations are not in the required range of 0.2 mL, pipet two more acid samples and repeat the titration. You should have three values within 0.2 mL of each other.

F. When you are finished, dispose of the solutions as instructed. Return the magnetic stirring bar to the container from which you obtained it.

CALCULATIONS

From the titration data and the diluted molarities, calculate the number of moles of H_2SO_4 and NaOH present in the volumes used (see Example 2 in the Sample Calculations). Then, calculate the mole ratio for each valid run, moles of NaOH/moles of H_2SO_4. Round your values to the nearest whole number. Use the average of the mole ratios to write the chemical equation for the reaction that has occurred.

Name *Date* *Section*

Experiment 22

Advance Study Assignment

1. How many moles of H_3PO_4 are contained in 75.0 mL of 0.215 M solution?

2. A 20.0-mL sample of 0.875 M HCl solution was diluted to 150.0 mL. Calculate the molarity of the resultant solution.

3. Why do you have to rinse the pipet with the ''concentrated'' H_2SO_4 before using it to prepare the dilute solution?

4. Why is it advisable to keep the storage bottle containing the NaOH solution closed?

Name *Date* *Section*

Experiment 22

Work Page

Concentration Data

Solution	*Concentrated*	*Dilute*
H_2SO_4		
NaOH		

Titration Data

Sample	*1*	*2*	*3*	*4*	*5*	*6*
Final NaOH (mL)						
Initial NaOH (mL)						
Volume NaOH (mL)						
Moles NaOH						
Volume H_2SO_4 (mL)						
Moles H_2SO_4						
Mole Ratio $\frac{\text{Moles NaOH}}{\text{Moles } H_2SO_4}$						

Show calculations for one valid titration run:

Write the chemical equation that corresponds to your mole ratio:

Name ____________________ Date __________ Section __________

Experiment 22

Report Sheet

Concentration Data

Solution	*Concentrated*	*Dilute*
H_2SO_4		
NaOH		

Titration Data

Sample	*1*	*2*	*3*	*4*	*5*	*6*
Final NaOH (mL)						
Initial NaOH (mL)						
Volume NaOH (mL)						
Moles NaOH						
Volume H_2SO_4 (mL)						
Moles H_2SO_4						
Mole Ratio $\frac{\text{Moles NaOH}}{\text{Moles } H_2SO_4}$						

Show calculations for one valid titration run:

Write the chemical equation that corresponds to your mole ratio: