Experiment

26

Heat of Neutralization

OBJECTIVE

To measure, using a calorimeter, the energy changes accompanying neutralization reactions.

APPARATUS AND CHEMICALS

Apparatus

Bunsen burner
Styrofoam cups (2)
cardboard square with hole in center
400-mL beaker
wire gauze
thermometers (2)
50-mL graduated cylinder
split one-hole rubber stopper
250-mL beaker
ring stand and ring

Chemicals

1 *M* HCl
1 *M* acetic acid ($HC_2H_3O_2$)
1 *M* NaOH

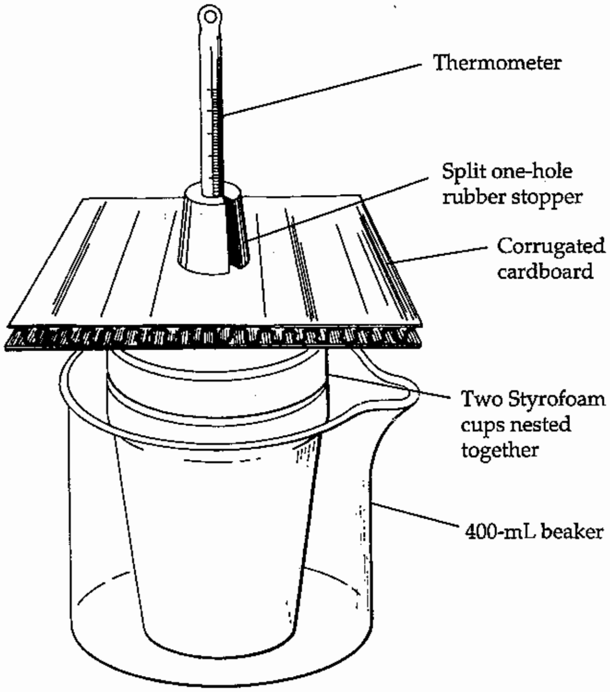
WORK IN PAIRS, BUT EVALUATE YOUR DATA INDIVIDUALLY.

DISCUSSION

Every chemical change is accompanied by a change in energy, usually in the form of heat. The energy change of a reaction that occurs at constant pressure is termed the *heat of reaction* or the *enthalpy change*. The symbol ΔH (the symbol Δ means "change in") is used to denote the enthalpy change. If heat is evolved, the reaction is *exothermic* ($\Delta H < 0$); and if heat is absorbed, the reaction is *endothermic* ($\Delta H > 0$). In this experiment, you will measure the heat of neutralization (or the enthalpy of neutralization) when an acid and a base react to form water.

This quantity of heat is measured experimentally by allowing the reaction to take place in a thermally insulated vessel called a *calorimeter*. The heat liberated in the neutralization will cause an increase in the temperature of the solution and of the calorimeter. If the calorimeter were perfect, no heat would be radiated to the laboratory. The calorimeter you will use in this experiment is shown in Figure 26.1.

Because we are concerned with the heat of the reaction and because some heat is absorbed by the calorimeter itself, we must know the amount of heat absorbed by the calorimeter. This requires that we determine the heat capacity of the calorimeter. By "heat capacity of the calorimeter" we mean the amount of heat (that is, the number of joules) required to raise its temperature 1 kelvin, which is the same as 1°C. In this experiment, the temperature of the calorimeter and its contents is measured before and after the reaction. The change in the enthalpy, ΔH, is equal to the negative product of the temperature change, ΔT, times the heat capacity of the calorimeter and its contents:



▲ **FIGURE 26.1** A simple calorimeter.

$$\Delta H = -\,\Delta T\,(\text{heat capacity of calorimeter} + \text{heat capacity of contents}) \quad [1]$$

Note that the *numerical difference* on the Celsius scale is the same as the *numerical difference* on the kelvin scale where ΔT is the difference between the final and initial temperatures: $\Delta T = T_f - T_i$. Because ΔH is negative for an exothermic reaction whereas ΔT is positive, a negative sign is required in Equation [1].

The heat capacity of the calorimeter is determined by measuring the temperature change that occurs when a known amount of hot water is added to a known amount of cold water in the calorimeter. The heat lost by the warm water is equal to the heat gained by the cold water and the calorimeter. (We assume no heat is lost to the laboratory.) For example, if T_1 equals the temperature of a calorimeter and 50 mL of cooler water, if T_2 equals the temperature of 50 mL of warmer water added to it, and if T_f equals the temperature after mixing, then the heat lost by the warmer water is

$$\text{heat lost by warmer water} = (T_2 - T_f) \times 50\text{ g} \times 4.18\text{ J/K-g} \quad [2]$$

The specific heat of water is 4.184 J/K-g, and the density of water is 1.00 g/mL. The heat gained by the cooler water is

$$\text{heat gained by cooler water} = (T_f - T_1) \times 50\text{ g} \times 4.18\text{ J/K-g} \quad [3]$$

The heat lost to the calorimeter is the difference between heat lost by the warmer water and that gained by the cooler water:

$$(\text{heat lost by warmer water}) - (\text{heat gained by cooler water}) \\ = \text{heat gained by the calorimeter}$$

Substituting Equations [2] and [3] we have

$$[(T_2 - T_f) \times 50\text{ g} \times 4.18\text{ J/K-g}] - [(T_f - T_1) \times 50\text{ g} \times 4.18\text{ J/K-g}] = (T_f - T_1) \times \text{heat capacity of calorimeter} \quad [4]$$

Note that the heat lost to the colorimeter equals its temperature change times its heat capacity. Thus by measuring T_1, T_2, and T_f, the heat capacity of the calorimeter can be calculated from Equation [4]. This is illustrated in Example 26.1.

EXAMPLE 26.1

Given the following data, calculate the heat lost by the warmer water, the heat lost to the cooler water, the heat lost to the calorimeter, and the heat capacity of the calorimeter:

Temperature of 50.0 mL warmer water: 37.92°C = T_2

Temperature of 50.0 mL cooler water: 20.91°C = T_1

Temperature after mixing: 29.11°C = T_f

SOLUTION: The heat lost by the warmer water, where ΔT = 37.92°C − 29.11°C, is

$$8.81\text{ K} \times 50\text{ g} \times 4.18\text{ J/K-g} = 1840\text{ J}$$

The heat gained by the cooler water, where ΔT = 29.11°C − 20.91°C, is

$$8.20\text{ K} \times 50\text{ g} \times 4.18\text{ J/K-g} = 1710\text{ J}$$

The heat lost to the calorimeter is

$$1840\text{ J} - 1710\text{ J} = 130\text{ J}$$

The heat capacity of the calorimeter is, therefore,

$$130\text{ J}/8.20\text{ K} = 16.0\text{ J/K}$$

Once the heat capacity of the calorimeter is determined, Equation [1] can be used to determine the ΔH for the neutralization reaction. Example 26.2 illustrates such a calculation.

EXAMPLE 26.2

Given the following data, calculate the heat gained by the solution, the heat gained by the calorimeter, and the heat of reaction:

Temperature of 50.0 mL of acid before mixing: 21.02°C

Temperature of 50.0 mL of base before mixing: 21.02°C

Temperature of 100.0 mL of solution after mixing: 27.53°C

Assume that the density of these solutions is 1.00 g/mL.

SOLUTION: The heat gained by the solution, where ΔT = 27.53°C − 21.02°C, is

$$6.51\text{ K} \times 100\text{ g} \times 4.18\text{ J/K-g} = 2720\text{ J}$$

The heat gained by the calorimeter, where ΔT = 27.53°C − 21.02°C, is

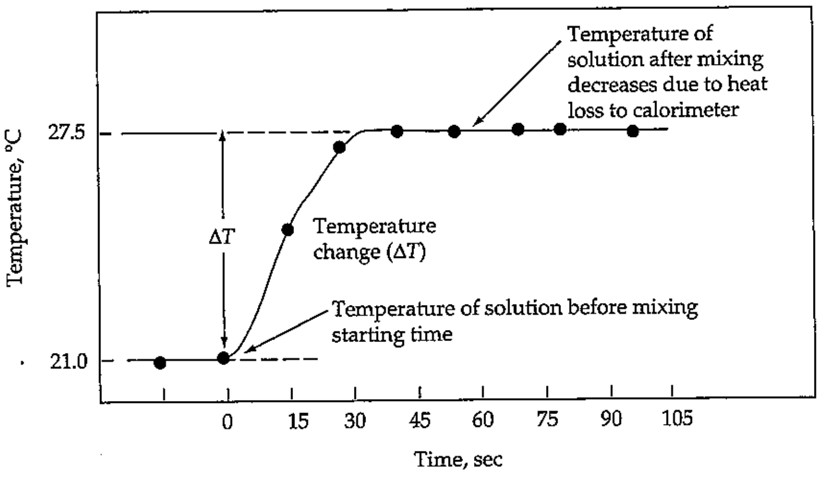
$$6.51\text{ K} \times 16.0\text{ J/K} = 104\text{ J}$$

The heat of reaction is therefore

$$2720\text{ J} + 104\text{ J} = 2824\text{ J}$$

or

$$2.82\text{ kJ}$$



▲ FIGURE 26.2 Temperature as a function of time.

PROCEDURE

A. Heat Capacity of Calorimeter

Construct a calorimeter similar to the one shown in Figure 26.1 by nesting two Styrofoam cups together. Use a cork borer to make a hole in the lid just big enough to admit the thermometer and slip the thermometer into a split one-hole rubber stopper to prevent the thermometer from entering too deeply into the calorimeter. The thermometer should not touch the bottom of the cup. Rest the entire apparatus in a 400-mL beaker to provide stability.

Place exactly 50.0 mL of tap water in the calorimeter cup and replace the cover and thermometer. Allow 5 to 10 min for the system to reach thermal equilibrium; then record the temperature to the nearest 0.1°C.

Place exactly 50.0 mL of water in a 250-mL beaker and heat the water with a low flame until the temperature is approximately 15° to 20°C above room temperature. Do not heat to boiling, or appreciable water will be lost, leading to an erroneous result. Allow the hot water to stand for a minute or two; quickly record its temperature to the nearest 0.1°C and pour it as completely as possible into the calorimeter. Replace the lid with the thermometer and carefully stir the water with the thermometer. Observe the temperature for the next 3 min and record the temperature every 15 s. Plot the temperature as a function of time, as shown in Figure 26.2. Determine ΔT from your curve and then do the calculations indicated on the report sheet.

B. Heat of Neutralization of HCl—NaOH

Dry the calorimeter and the thermometer with a towel. Carefully measure 50.0 mL of 1.0 *M* NaOH and add it to the calorimeter. Place the lid on the calorimeter but leave the thermometer out. Measure out exactly 50.0 mL of 1.0 *M* HCl into a dry beaker. Allow it to stand near the calorimeter for 3 to 4 min. Measure the temperature of the acid, rinse the thermometer with tap water, and wipe dry. Insert the thermometer into the calorimeter and measure the temperature of the NaOH solution.

The temperatures of the NaOH and the HCl should not differ by more than 0.5°C. If the difference is greater than 0.5°C, adjust the temperature of the HCl by *either* warming it by holding the beaker in your hands or cooling the outside of the beaker with tap water until the temperature of the HCl is within 0.5°C of that of the NaOH.

Record the temperature of the NaOH solution. Lift the lid and carefully add the 1.0 *M* HCl all at once. Be careful not to splash any on the upper sides of the cup. Stir the solution gently with the thermometer and record the temperature as a function of time every 15 s for the next 3 min. Construct a temperature-versus-time curve and determine ΔT. Calculate the heat of neutralization per mole of water formed. You may assume that the NaCl solution has the same density and specific heat as water.

C. Heat of Neutralization of $HC_2H_3O_2$—NaOH

Follow the same procedure as in Part B, but substitute 1.0 *M* $HC_2H_3O_2$ for 1.0 *M* HCl. Calculate the heat of neutralization per mole of water formed.

Waste Disposal Instructions Handle the stock solutions carefully. You may use a wet sponge or paper towel to clean up any spills. The reaction mixtures produced in the Styrofoam cups contain harmless salts. They should be disposed of in the designated receptacles.

REVIEW QUESTIONS

Before beginning this experiment in the laboratory, you should be able to answer the following questions:

1. Define endothermic and exothermic reactions in terms of the sign of ΔH.
2. A 625-mL sample of water was cooled from 50.0°C to 10.0°C. How much heat was lost?
3. Define the term *heat capacity*.
4. How many joules are required to change the temperature of 60.0 g of water from 23.3°C to 38.8°C?
5. Define the term *specific heat*.
6. Calculate the final temperature when 50 mL of water at 60°C are added to 25 mL of water at 20°C.
7. Describe how you could determine the specific heat of a metal by using the apparatus and techniques in this experiment.
8. A piece of metal weighing 5.10 g at a temperature of 48.6°C was placed in a calorimeter into 20.00 mL of water at 22.1°C, and the final equilibrium temperature was found to be 28.2°C. What is the specific heat of the metal?
9. If the specific heat of methanol is 2.51 J/K-g, how many joules are necessary to raise the temperature of 250 g of methanol from 18°C to 33°C?
10. When a 3.25-g sample of solid sodium hydroxide was dissolved in a calorimeter in 100.0 g of water, the temperature rose from 23.9°C to 32.0°C. Calculate ΔH (in kJ/mol NaOH) for the solution process:

$$NaOH(s) \longrightarrow Na^{+}(aq) + OH(aq)$$

Assume it's a perfect calorimeter and that the specific heat of the solution is the same as that of pure water.

Name ______________________ Desk ____________

Date ____________ Laboratory Instructor ______________________

REPORT SHEET | EXPERIMENT

Heat of Neutralization | 26

A. Heat Capacity of Calorimeter

1. Temp. of calorimeter and water before mixing ____________ °C
2. Temp. of warm water ____________ °C
3. Maximum temp. determined from your curve ____________ °C
4. Heat lost by warm water (temp decrease × 50.0 g × 4.18 J/K-g)= ____________ J

5. Heat gained by cooler water (temp. increase × 50.0 g × 4.18 J/K-g)= ____________ J

6. Heat gained by the calorimeter [(4)−(5)]= ____________ J

7. Heat capacity of calorimeter:

$$\frac{\text{Heat gained by the calorimeter}}{\text{Temperature increase}} =$$ ____________ J/K

B. Heat of Neutralization of HCl—NaOH

1. Temp. of calorimeter and NaOH ____________ °C
2. ΔT determined from your curve after adding HCl to the NaOH ____________ °C
3. Heat gained by solution (temperature increase × 100 g × 4.18 J/K-g) ____________ J

4. Heat gained by calorimeter (temperature increase × heat capacity of calorimeter) = ____________ J

5. Total joules released by reaction [(3) + (4)] = ____________ J

6. Complete: HCl + NaOH ⟶ ____________

7. The number of moles of HCl in 50 mL of 1.0 *M* HCl (show calculations): ____________ mol

8. The number of moles of H_2O produced in reaction of 50 mL 1.0 *M* HCl and 50 mL 1.0 *M* NaOH (show calculations): ____________ mol

9. Joules released per mole of water formed:

$$\frac{\text{Total joules released (5)}}{\text{Number of moles water produced (8)}} = \text{\_\_\_\_\_\_\_\_\_\_\_\_ kJ/mol}$$

C. Heat of Neutralization of $HC_2H_3O_2$—NaOH

1. Temperature of calorimeter and NaOH ____________ °C
2. ΔT determined from cooling curve after adding $HC_2H_3O_2$ to NaOH ____________ °C
3. Heat gained by solution (temp. increase × 100 g × 4.18 J/K-g) = ____________ J

4. Heat gained by calorimeter (temp. increase × heat capacity of calorimeter) = ____________ J

5. Total joules released by reaction [(3)+(4)] = ____________ J

6. Complete: $HC_2H_3O_2 + NaOH \longrightarrow$ ____________
7. The number of moles of H_2O produced in reaction of 50 mL 1.0 *M* $HC_2H_3O_2$ and 50 mL 1.0 *M* NaOH (show calculations): ____________ mol

8. Joules released per mole of water formed:

$$\frac{\text{Total joules released (5)}}{\text{Number of moles water produced (7)}} = \text{\_\_\_\_\_\_\_\_\_\_\_\_ kJ/mol}$$

QUESTIONS

1. What is the largest source of error in the experiment?

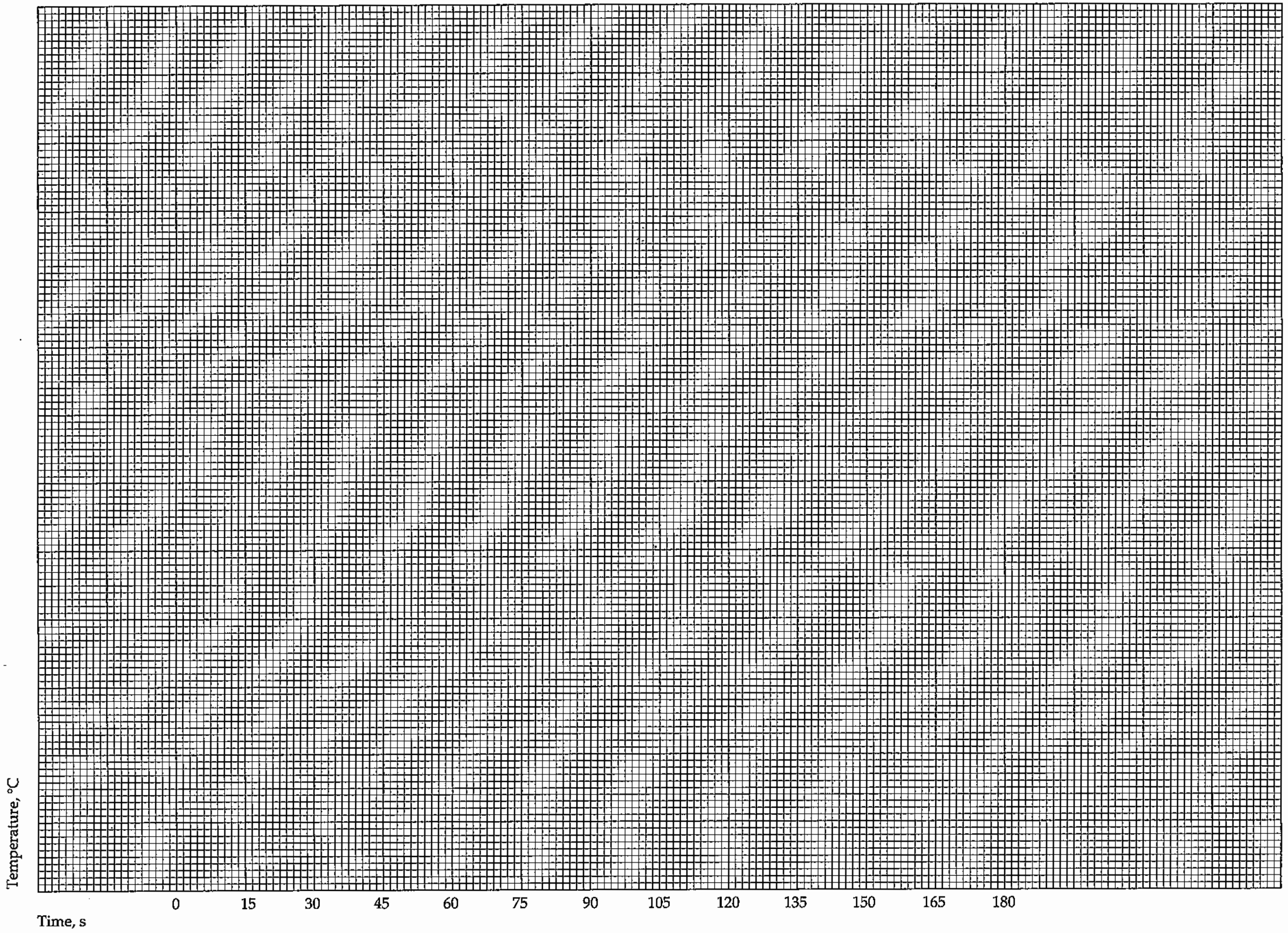
2. How should the two heats of reaction for the neutralization of NaOH and the two acids compare? Why?

3. The experimental procedure has you wash your thermometer and dry it after you measure the temperature of NaOH solution and before you measure the temperature of the HCl solution. Why?

4. A 50.0-mL sample of a 1.00 *M* solution of $CuSO_4$ is mixed with 50.0 mL of 2.00 *M* KOH in a calorimeter. The temperature of both solutions was 20.2°C before mixing and 26.3°C after mixing. The heat capacity of the calorimeter is 12.1 J/K. From these data calculate ΔH for the process

$$CuSO_4(1\ M) + 2KOH(2\ M) \longrightarrow Cu(OH)_2(s) + K_2SO_4(0.5\ M)$$

Assume the specific heat and density of the solution after mixing are the same as those of pure water.


Temperature, °C
0
15
30
45
60
75
90
105
120
135
150
165
180
Time, s