

## Heat of Neutralization

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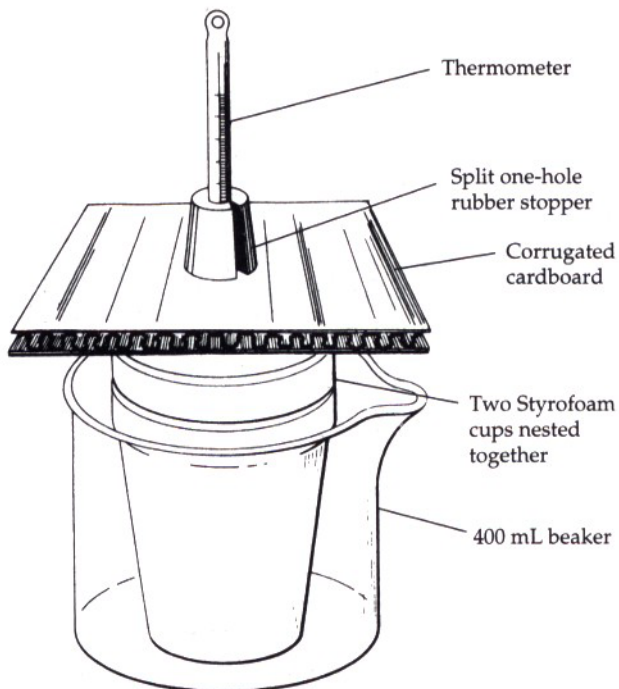
Purpose: To measure, using a calorimeter, the energy changes accompanying neutralization reactions.

Equipment: 2 Styrofoam (polystyrene) cups, cardboard square with hole in center, split one-hole rubber stopper, thermometer, 400-mL beaker, 250-mL beaker, 50-mL graduated cylinder

Materials: 1 M HCl, 1 M HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> (acetic acid), 1 M NaOH

Introduction: 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.” If heat is evolved, the reaction is exothermic. If heat is absorbed, the reaction is endothermic. In this experiment, two exothermic reactions will be observed, and the heat of reaction for each will be found. The reaction studied will be the “heat of neutralization,” which is the enthalpy change produced when an acid and a base react to form water.

In order to measure the amount of heat produced by a reaction, an instrument called a calorimeter must be used. The calorimeter used in this experiment will be somewhat rudimentary. It will be constructed as shown in the following picture.



▲ FIGURE 26.1 A simple calorimeter.

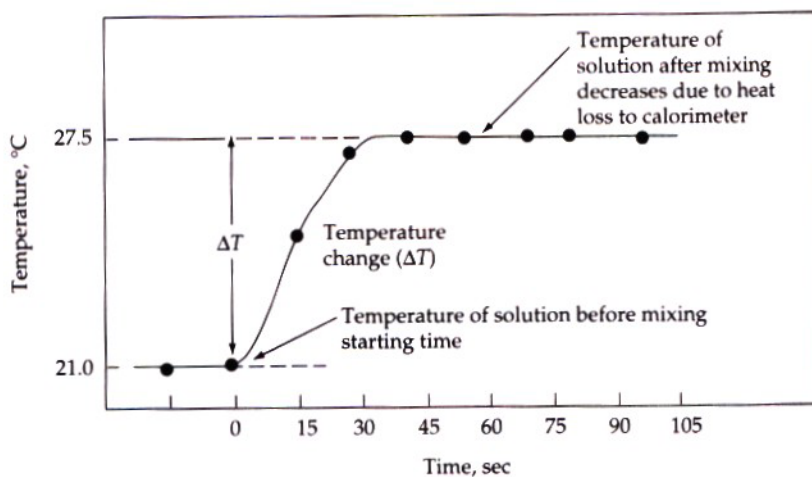
This calorimeter consists of two polystyrene cups nested together, which are then placed in a 400-mL beaker. A cardboard square is added on top to keep heat inside the calorimeter and to stabilize the thermometer which juts out of the cardboard into the solution below.

The heat liberated in the neutralization reaction occurring within the calorimeter 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. Two forms of enthalpy must be measured in this experiment: the heat that causes a temperature increase in the solution and the heat that causes a temperature increase in the calorimeter.

In order to determine the amount of heat absorbed by the calorimeter, one must first determine the heat capacity of the calorimeter. The heat capacity is the amount of heat required to raise its temperature by 1 Kelvin, or 1 degree Celsius, which are equivalent increments. However, the heat capacity must be found using standard solutions. In

this case, tap water will be used to find the heat capacity of the calorimeter as outlined in the Procedure.

In order to interpret the data obtained, two final steps must be taken. All of the temperatures observed by the thermometer must be adjusted according to the calibration curve. This calibration curve was constructed during a previous experiment, so the same graph will be used and attached to this report. Also, the change in temperature of a solution must be found in order to determine the heat absorbed by solution. In order to find the change in temperature, a graph of temperature versus time must be constructed (as outlined in the Procedure). This graph will look like the graph presented in Figure 26.2:



▲ FIGURE 26.2 Temperature as a function of time.

Figure 26.2 serves to show the graph and explain its components.

In this experiment, after finding the heat capacity of the calorimeter to be used, two acids will be neutralized and the heat of neutralization will be calculated. Those two acids are hydrochloric acid and acetic acid. 1.0 M solutions of each will be used in this experiment.

Procedure:

### **A. Heat Capacity of Calorimeter**

1. A calorimeter similar to the one shown in Figure 26.1 in the Introduction was constructed by nesting two polystyrene cups together. A thermometer was placed into the hole in the cardboard square and was stabilized by a split one-hole rubber stopper.
2. Exactly 50.0 mL of tap water was placed in the calorimeter cup. The water was allowed to sit with a thermometer in it for 10 minutes to reach thermal equilibrium. The temperature was measured and recorded.
3. Exactly 50.0 mL of hot tap water was placed in a 250-mL beaker. The temperature of the hot water was measured and recorded.
4. The hot water was poured into the calorimeter. The lid of the calorimeter was replaced and the water was stirred gently with the thermometer. The temperature of the water was recorded every 15 seconds.
5. A graph of temperature as a function of time was constructed.

### **B. Heat of Neutralization of HCl—NaOH**

1. The calorimeter and the thermometer were dried with a towel. 50.0 mL of 1.0 *M* NaOH was measured and added to the calorimeter. The lid of the calorimeter was replaced without the thermometer.
2. Exactly 50.0 mL of 1.0 *M* HCl was measured into a dry beaker. It was allowed to stand near the calorimeter for 4 minutes. The temperature of the acid was measured. The thermometer was rinsed with tap water and dried. The temperature of the base solution was measured and recorded.
3. The HCl was added to the NaOH. The solution was stirred gently and the temperature was recorded every 15 seconds.

4. A graph of temperature as a function of time was constructed.

### **C. Heat of Neutralization of HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub>—NaOH**

1. The calorimeter and the thermometer were dried with a towel. 50.0 mL of 1.0 M NaOH was measured and added to the calorimeter. The lid of the calorimeter was replaced without the thermometer.

2. Exactly 50.0 mL of 1.0 M HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> was measured into a dry beaker. It was allowed to stand near the calorimeter for 4 minutes. The temperature of the acid was measured. The thermometer was rinsed with tap water and dried. The temperature of the base solution was measured and recorded.

3. The HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> was added to the NaOH. The solution was stirred gently and the temperature was recorded every 15 seconds.

4. A graph of temperature as a function of time was constructed.

#### Observations:

The temperature was calibrated for each of these observations using a previously-made calibration curve. It is attached with interpolation lines drawn in.

#### **A. Heat Capacity of Calorimeter**

*Temperature of calorimeter and water before mixing (observed):*  $21.5 \pm .2 \text{ }^\circ\text{C}$

*Temperature of calorimeter and water before mixing (adjusted):*  $18.5 \pm .2 \text{ }^\circ\text{C}$

*Temperature of warm water (observed):*  $46.0 \pm .2 \text{ }^\circ\text{C}$

*Temperature of warm water (adjusted):*  $43.0 \pm .2 \text{ }^\circ\text{C}$

#### **B. Heat of Neutralization of HCl—NaOH**

*Temperature of calorimeter and NaOH (observed):*  $23.0 \pm .2 \text{ }^\circ\text{C}$

*Temperature of calorimeter and NaOH (adjusted):*  $20.0 \pm .2 \text{ }^\circ\text{C}$

### C. Heat of Neutralization of $\text{HC}_2\text{H}_3\text{O}_2$ — $\text{NaOH}$

*Temperature of calorimeter and NaOH (observed):*  $22.0 \pm .2 \text{ }^\circ\text{C}$

*Temperature of calorimeter and NaOH (adjusted):*  $19.0 \pm .2 \text{ }^\circ\text{C}$

→Graphs of temperature versus time for each of the three parts of the experiment can be found at the end of this report, clearly labeled. These graphs were used for the results in the next section. All of these graphs represent adjusted temperatures as interpolated on the calibration curve.

#### Results:

##### A. Heat Capacity of Calorimeter

*Maximum temperature* (determined from curve) =  $28.0 \pm .2 \text{ }^\circ\text{C}$

*Heat lost by warm water* (temperature decrease times mass of water times specific heat of water) =  $(43.0 \pm .2 \text{ }^\circ\text{C} - 28.0 \pm .2 \text{ }^\circ\text{C})(50.0 \text{ g})(4.18 \text{ J/}^\circ\text{C-g}) = (43.0 \pm .47\% \text{ }^\circ\text{C} - 28.0 \pm .71\% \text{ }^\circ\text{C})(50.0 \text{ g})(4.18 \text{ J/}^\circ\text{C-g}) = 3135 \pm 1.18\% \text{ J} = 3140 \pm 40 \text{ J}$

*Heat gained by cooler water* (temperature increase times mass of water times specific heat of water) =  $(28.0 \pm .2 \text{ }^\circ\text{C} - 18.5 \pm .2 \text{ }^\circ\text{C})(50.0 \text{ g})(4.18 \text{ J/}^\circ\text{C-g}) = (28.0 \pm .71\% \text{ }^\circ\text{C} - 18.5 \pm 1.08\% \text{ }^\circ\text{C})(50.0 \text{ g})(4.18 \text{ J/}^\circ\text{C-g}) = 1986 \pm 1.79\% \text{ J} = 1990 \pm 40 \text{ J}$

*Heat gained by the calorimeter* (heat lost by warm water minus heat gained by cooler water) =  $3140 \pm 40 \text{ J} - 1990 \pm 40 \text{ J} = 1150 \pm 80 \text{ J}$

*Heat capacity of calorimeter* (heat gained by the calorimeter divided by temperature increase) =  $1150 \pm 6.96\% \text{ J} / (28.0 \pm .71\% \text{ }^\circ\text{C} - 18.5 \pm 1.08\% \text{ }^\circ\text{C}) = 121 \pm 8.75\% \text{ J/K} = 120 \pm 10 \text{ J/K}$

## B. Heat of Neutralization of HCl—NaOH

*Change in temperature* (determined from curve) =  $6.0 \pm .2 \text{ }^\circ\text{C}$

*Heat gained by solution* (temperature increase times mass of solution times specific heat of water) =  $(6.0 \pm .2 \text{ }^\circ\text{C})(100 \text{ g})(4.18 \text{ J}^\circ\text{C-g}) = (6.0 \pm 3.33\% \text{ }^\circ\text{C})(100 \text{ g})(4.18 \text{ J}^\circ\text{C-g}) = 2508 \pm 3.33\% \text{ J} = 2510 \pm 80 \text{ J}$

*Heat gained by calorimeter* (temperature increase times heat capacity of calorimeter) =  $(6.0 \pm 3.33\% \text{ }^\circ\text{C})(120 \pm 8.33\% \text{ J/K}) = 720 \pm 11.66\% \text{ J} = 720 \pm 80 \text{ J}$

*Total joules released by reaction* (heat gained by solution plus heat gained by calorimeter) =  $2510 \pm 80 \text{ J} + 720 \pm 80 \text{ J} = 3230 \pm 160 \text{ J}$

*Balanced equation* (simple derivation) =  $\text{HCl} + \text{NaOH} \rightarrow \text{NaOH} + \text{H}_2\text{O}$

*Number of moles of HCl in 50 mL of 1.0 M HCl* (molarity divided by liters of solution) =  $1.0 \text{ M} / .05 \text{ L} = .05 \text{ mol HCl}$

*Number of moles of H<sub>2</sub>O produced in reaction of 50 mL 1.0 M HCl and 50 mL 1.0 M NaOH* (number of moles of HCl times ratio of moles of H<sub>2</sub>O to HCl) =  $.05 \text{ mol HCl} (1 \text{ mol H}_2\text{O} / 1 \text{ mol HCl}) = .05 \text{ mol H}_2\text{O}$

*Joules released per mole of water formed* (total joules released divided by number of moles of water produced) =  $3230 \pm 160 \text{ J} / .05 \text{ mol H}_2\text{O} = (3.230 \pm 4.95\% \text{ kJ}) / .05 \text{ mol H}_2\text{O} = 64.6 \pm 4.95\% \text{ kJ/mol} = 65 \pm 3 \text{ kJ/mol}$

## C. Heat of Neutralization of HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub>—NaOH

*Change in temperature* (determined from curve) =  $7.0 \pm .2 \text{ }^\circ\text{C}$

*Heat gained by solution* (temperature increase times mass of solution times specific heat of water) =  $(7.0 \pm .2 \text{ }^\circ\text{C})(100 \text{ g})(4.18 \text{ J}^\circ\text{C-g}) = (7.0 \pm 2.86\% \text{ }^\circ\text{C})(100 \text{ g})(4.18 \text{ J}^\circ\text{C-g}) = 2926 \pm 2.86\% \text{ J} = 2930 \pm 80 \text{ J}$

*Heat gained by calorimeter* (temperature increase times heat capacity of calorimeter) =  
 $(7.0 \pm 2.86\% \text{ } ^\circ\text{C})(120 \pm 8.33\% \text{ J/K}) = 840 \pm 11.19\% \text{ J} = 840 \pm 90 \text{ J}$

*Total joules released by reaction* (heat gained by solution plus heat gained by calorimeter) =  $2930 \pm 80 \text{ J} + 840 \pm 90 \text{ J} = 3770 \pm 170 \text{ J}$

*Balanced equation* (simple derivation) =  $\text{HC}_2\text{H}_3\text{O}_2 + \text{NaOH} \rightarrow \text{NaC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O}$

*Number of moles of H<sub>2</sub>O produced in reaction of 50 mL 1.0 M HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> and 50 mL 1.0 M NaOH* (number of moles of HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> times ratio of moles water to moles HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub>) =  
 $.05 \text{ mol HC}_2\text{H}_3\text{O}_2 (1 \text{ mol H}_2\text{O} / 1 \text{ mol HC}_2\text{H}_3\text{O}_2) = .05 \text{ mol H}_2\text{O}$

*Joules released per mole of water formed* (total joules released divided by number of moles of water produced) =  $(3770 \pm 170 \text{ J}) / .05 \text{ mol} = 3.770 \pm 4.51\% \text{ kJ} / .05 \text{ mol} = 75.4 \pm 4.51\% \text{ kJ/mol} = 75 \pm 3 \text{ kJ}$

Discussion: The values obtained are experimental data that must be compared to the true value in order to gauge the accuracy of the experiment. The true value of the joules released per mole of water for the HCl reaction is 58.2 kJ/mol. The true value of the joules released per mole for the acetic acid reaction is 55.8 kJ/mol.

*Percent error, HCl reaction.*

$$(65 \text{ kJ} - 58.2 \text{ kJ}) / 58.2 \text{ kJ} * 100 =$$

**11.68% error.**

*Percent error, acetic acid reaction.*

$$(75 \text{ kJ} - 55.8 \text{ kJ}) / 55.8 \text{ kJ} * 100 =$$

**34.41% error.**

These are astronomical percent errors. There were major mistakes with equipment and procedure. A few sources of error that could account for the discrepancies are the measuring devices themselves. A graduated cylinder was used to measure volume. A buret or pipet should have been used for maximum accuracy. Using a graduated cylinder leaves room for a crucial error in volume determination, which would then lead to errors in determination of mass, molar content of the solution, and every other derivative formula. Not only could an inaccurate amount of solution be poured into the cylinder, but not all of this solution may be poured out. There are always extra drops of solution clinging to the walls of a cylinder. This would have negatively influenced volume readings. In addition, the thermometer had to be calibrated, which improves accuracy but is itself an imprecise technique. A better thermometer could have prevented errors in temperature readings. Also, it is assumed that the specific heat of the solution is the same as the specific heat of water, and that the tap water used in the experiment contains negligible impurities. Neither of these assumptions should be made, as slight discrepancies can cause larger discrepancies in calculations later in the experiment.

The theory associated with this experiment deals with potential energy in molecules. Heat was released by molecules in this reaction. This heat came from potential energy stored in chemical bonds. When the bonds were broken and new bonds were formed, energy was released. This experiment serves to verify the atomic theory and the observations made about chemical bonds. In addition, various parts of this experiment relied on the atomic theory, such as the determination of the number of moles of solute in a solution.

The ramifications for this experiment are broad. First, experience in the laboratory was gained. Secondly, the energy of a system was studied. Anything can be classified as a system and its energy can be studied. This experiment related the way that energy behaves in a system. Energy is released by commonplace reactions; if this energy could be harnessed, it could be used. This is the driving force behind hydrogen fuel cells. Hydrogen and oxygen are mixed, producing water, but also producing massive amounts of energy. The formation of water was the driving force of this metathesis reaction, and it produced energy.

Questions:

1. The largest source of error in this experiment was the gross imprecision of the measuring instruments. Exactly 50.0 mL of solutions should have been used, as many derivative calculations depend on that amount being precise. The plus/minus values provide cause for concern. The calorimeter itself had a high heat capacity. This experiment could have gone better if more precise measuring instruments and calorimeters had been used.
2. The two heats of neutralization should be the same. The same number of moles of NaOH are acting on the acid in each case. Thus, while initial and final temperatures may vary, the change in temperature should be the same in each case.
3. The procedure calls for the washing of the thermometer between measuring the temperature of NaOH and HCl because the drops of NaOH on the thermometer would react with an equivalent of HCl, causing a discrepancy, however small, in moles of reactant solutions, and in initial temperature. The change in temperature would be changed accordingly, leading to larger errors.

4. In order to find  $\Delta H$ , several steps must be undertaken. First, the heat absorbed by solution must be found.

*Heat absorbed by solution* (temperature increase times mass of solution times specific heat) =  $(6.1\text{ }^{\circ}\text{C})(100\text{ g})(4.18\text{ J/}^{\circ}\text{C-g}) = 2550\text{ J}$

Then the heat absorbed by the calorimeter must be calculated.

*Heat absorbed by calorimeter* (temperature increase times heat capacity of calorimeter) =  $(6.1\text{ }^{\circ}\text{C})(12.1\text{ J/K}) = 74\text{ J}$

Now the two values must be added.

*Heat produced by reaction* (heat absorbed by solution plus heat absorbed by calorimeter) =  $2550\text{ J} + 74\text{ J} = 2624\text{ J} = 2.624\text{ kJ}$

Conclusion: The experiment was successful to the degree that the procedure was carried out correctly. Equipment problems and other unavoidable sources of error served to cause a large percent error for each part of the experiment. However, the techniques associated with calorimetry were practiced, providing valuable experience.