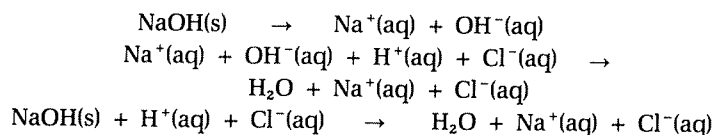# Thermodynamics

The formation or destruction of chemical bonds is always accompanied by an energy exchange between the reactant molecules and the immediate environment. The term ΔH is used to describe the resulting *enthalpy changes*. When atoms or ions form new bonds, energy is released into the environment and the reaction products possess lower enthalpy than the reactants. This energy release may be observed as a measurable increase in the temperature of the surroundings. *Exothermic* reactions are associated with negative ΔH values. Conversely, when chemical bonds are broken, energy must be absorbed from the environment and the products possess a higher enthalpy than the reactants. The ΔH value is positive. This is called an *endothermic* reaction, which is characterized by a measurable decrease in the temperature of the surroundings.

Germain Henry Hess, a Swiss-Russian chemist examined heat transformations in reactions involving several steps. He found that the enthalpy changes for a series of reactions can be added together to describe the energy change for the overall reaction. Known as Hess's law, this relationship will be studied in this laboratory exercise. By measuring the heat transformations generated from a sequence of reactions, you will obtain experimental evidence for the law of enthalpy additivity. Determinations and comparisons of the heat that evolved in the following three exothermic reactions will be performed:

$$NaOH(s) \quad \rightarrow \quad Na^+(aq) + OH^-(aq)$$
$$Na^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \quad \rightarrow$$
$$H_2O + Na^+(aq) + Cl^-(aq)$$
$$NaOH(s) + H^+(aq) + Cl^-(aq) \quad \rightarrow \quad H_2O + Na^+(aq) + Cl^-(aq)$$

A simple calorimeter will be constructed using a plastic foam cup. Noticeable rises in water temperature will be used to calculate the ΔH value for each of the reactions.

OBJECTIVES

1. to construct a simple calorimeter
2. to determine the amount of heat energy released or absorbed in three separate reactions
3. obtain experimental evidence for the additivity of reaction heats

MATERIALS

Apparatus

graduated cylinder	stirring rod
plastic foam cup	lab apron
tongs or forceps	safety goggles
centigram balance	plastic gloves
thermometer	full face shield
1 beaker (250-mL)	glazed paper
	(optional)

Reagents

NaOH pellets
distilled water
1.0M HCl
0.5M HCl
1.0M NaOH

CAUTION: *Your thermometer is made of glass and can break, leaving sharp edges that cut. Handle your thermometer gently. For example, do not use it to crush a piece of ice. The glass around the mercury bulb is very thin. Mercury vapor and liquid are very poisonous. If your thermometer breaks, call your teacher.*

PRELAB

Answer questions 1-9 on the Report Sheet.

Laboratory Experiments

PROCEDURE

Part I

CAUTION: Sodium hydroxide pellets are very corrosive to the skin, eyes, and clothing; if even one gets in your mouth or is swallowed, it will cause serious damage. Wear safety goggles, full face shield, and gloves. Do not touch a pellet, even while wearing gloves; use tongs. If any pellets are spilled, call your teacher. Do not attempt to pick up or move spilled pellets. Immediately wash pellets and residues of pellets off your skin or clothing, using plenty of water. Call your teacher.

1. Put on your lab apron, safety goggles, plastic gloves, and full face shield.
2. Measure 100 mL of distilled water into a graduated cylinder and pour it into a clean, dry, plastic foam cup.
3. Obtain approximately 2 g of NaOH pellets. Record their mass to the nearest 0.01 g on glazed paper. Since NaOH absorbs water from the air, it is necessary to proceed to Steps 4 and 5 as soon as the mass of the pellets has been found.
4. Measure and record the temperature of the water in the calorimeter to the nearest 0.1 °C.
5. Add the NaOH pellets to the water and constantly check the solution temperature. Use a stirring rod to gently stir the solution until all the solid has dissolved.
6. Record the highest solution temperature reached on the Report Sheet. Discard the solution into a waste container and rinse all apparatus with distilled water.

Part II

CAUTION: Hydrochloric acid is corrosive to skin, eyes, and clothing. When handling 1M hydrochloric acid, wear safety goggles and lab apron. Immediately wash spills and splashes off your skin and clothing, using plenty of water. Call your teacher.

1. Measure 50 mL of a 1.0M hydrochloric acid solution into a graduated cylinder. Carefully pour this solution into a clean, dry, plastic foam cup.
2. Measure 50 mL of 1.0M sodium hydroxide solution into a graduated cylinder. Carefully pour it into a 250-mL beaker.
3. Record the temperature of each solution to the nearest 0.1 °C. Use the same thermometer for both readings, but be sure to rinse it with distilled water between measurements.
4. Pour the sodium hydroxide into the plastic foam cup containing the hydrochloric acid solution. Using a stirring rod, stir the solution gently. Record the highest temperature reached on the Report Sheet. Discard the solution into a waste container and rinse all apparatus with distilled water.

Part III

CAUTION: Sodium hydroxide solution is corrosive to the skin, eyes, and clothing. When handling 1M solutions, wear safety goggles and gloves. Immediately wash spills and splashes off your skin and clothing, using plenty of water. Call your teacher.

1. Measure 100 mL of 0.5M hydrochloric acid into a graduated cylinder. Pour the solution into a clean, dry, plastic foam cup.
2. Obtain 2 g of NaOH pellets. Record their mass to the nearest 0.01 g.
3. Measure and record the temperature of the hydrochloric acid solution to the nearest 0.1 °C.
4. While monitoring the solution temperature, add the NaOH pellets to the HCl. Using a stirring rod, gently stir the solution until all the solid has dissolved.
5. Record the highest solution temperature reached. Discard the solution.
6. Before you leave the laboratory, wash your hands thoroughly with soap and water; use a fingernail brush to clean under your fingernails.

Laboratory Experiments

POST LAB DISCUSSION

The heat liberated by each of the three reactions was released into the calorimeter environment. Since the quantity of heat absorbed by the plastic foam cup or transferred to the atmosphere is negligible, you can assume that all evolved energy was absorbed by the solution. To convert the observed rise in solution temperatures into ΔH values, the collected data for each reaction must be substituted into the following equation:

$$\text{heat of reaction}(\Delta H) = \text{temperature change}(\Delta t)$$
$$\times \text{ mass of solution} \times \text{specific heat of solution}(c)$$

The mass of the solution can be calculated from its final volume if its density is known. Since the solutions in this experiment are dilute aqueous solutions, their densities can be taken to be the same as that of water, 1.0 g/mL. The specific heat (c) of a dilute aqueous solution can be taken to be the same as that of water, 4180 J/kg·°C.

Since the heats of reaction are commonly expressed as heat per *mole*, an additional calculation is needed. The heat of reaction per mole of NaOH can be obtained by dividing the heat given off in a reaction by the number of moles of NaOH that reacted.

The energy value obtained in Part I represents the energy of solution for one mole of NaOH(s). The value represented for Part II represents the heat of reaction for one mole of H^+(aq) and one mole of OH^-(aq). Part III can be considered to be the sum of the reactions occurring in Parts I and II. Its heat of reaction should therefore equal the heat of solution of NaOH(s) plus the heat of reaction of H^+(aq) and OH^-(aq).

Thermodynamics

PRELAB QUESTIONS

1. What is meant by the symbol, ΔH?_____

_____

_____

2. What is the significance of a negative ΔH?_____

_____

3. Write the equation for the reaction between solid sodium hydroxide and water. _____

_____

_____

4. State Hess's law. _____

_____

_____

5. Write the equation for the reaction between solid sodium hydroxide and dilute hydrochloric acid.

_____

_____

6. Differentiate between exothermic and endothermic reactions. _____

_____

7. Why should all procedures involving solid NaOH be completed quickly? _____

_____

8. Why is it important to have all the solutions that are used in this experiment at room temperature?

_____

_____

9. In Step 3 of Part II, why should the thermometer be rinsed between readings? _____

_____

DATA
Part I

Original temperature of water (t_1)	
Final temperature of solution (t_2)	
Temperature change ($t_2 - t_1 = \Delta t$)	
Mass of 100 mL of water	
Heat evolved by reaction (ΔH)	
Mass of NaOH(s)	
Moles of NaOH	
Energy per mole of NaOH(x_1)	

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Part II

Original temperature of HCl(aq)	
Original temperature of NaOH(aq)	
Average original temperature (t_1)	
Final temperature of solution (t_2)	
Temperature change ($t_2 - t_1 = \Delta t$)	
Total mass of solution (assume 1 mL = 1 g)	
Heat evolved by reaction (ΔH)	
Molarity of NaOH solution	
Volume of NaOH solution	
Moles of NaOH	
Energy per mole of NaOH (x_2)	

Part III

Original temperature of HCl(aq) (t_1)	
Final temperature of solution (t_2)	
Temperature change ($t_2 - t_1 = \Delta t$)	
Mass of HCl (assume 1 mL = 1 g)	
Heat evolved by reaction (ΔH)	
Mass of NaOH(s)	
Moles of NaOH	
Energy per mole of NaOH (x_3)	

CONCLUSIONS

1. Compare the calculated enthalpy change of each of the three reactions. _____

_____

_____

2. Write the net ionic equations for the reactions in Part II and Part III.

_____

_____

3. How does the value of x_3 differ from the sum of x_1 and x_2? _____

_____

4. Calculate the percent deviation between x_3 and the sum of x_1 and x_2. Assume x_3 to be correct.

5. Suggest several reasons for the observed percent difference. _____

_____

6. Suppose you doubled the mass of NaOH used in Part I. How would this affect the number of joules released in this reaction? What would be the effect on x_1? _____

_____

_____

SYNTHESIS

1. The dissolving of ammonium nitrate in water is a highly endothermic process. How could this fact be utilized in the development of a cold pack used to treat athletic injuries? _____

_____

_____

_____