Endothermic & Exothermic Reactions Wet Lab

Note: This lab is advanced and requires extra supplies not included

Standards:

GLE 0807.Inq.1 Design and conduct open- ended scientific investigations.
GLE 0807.Inq.2 Use appropriate tools and techniques to gather, organize, analyze, and interpret data.

This lab requires some lab equipment that is not included in the kit. If there is not access to these supplies, the class cannot do this lab. This lab assumes 6 groups to a class. There can be more groups if there are more supplies. This lab is along the lines of a high school chemistry class procedure. It is not recommended for classes younger than 7^{th} grade or classes with poor classroom management.

Supplies Needed but **Not** Included:

- 6 electric balances
- 6 400 ml beakers
- 6 50mL Graduated Cylinders
- 1-6 Spatula's/Scoop
- 6 Thermometers

Supplies Included:

- Carboy with DI water
- 12 Styrofoam Cups
- Ammonium Nitrate NH4NO3, 24g per group
- Hydrochloric Acid Solution 1M HCl, 50mL per group
- Sodium Hydroxide solution 1M NaOH, 50mL per group
- Sodium Hydroxide Flakes NaOH, 16g per group

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Exothermic & Endothermic Reactions

An Energy Changes Laboratory Kit

Introduction

Energy changes occur in all chemical reactions—energy is either absorbed or released. If energy is released in the form of heat, the reaction is called *exothermic*. If energy is absorbed, the reaction is called *endothermic*. Can these energy changes be measured? Let's investigate the amount of heat absorbed or released from three different reactions.

Chemical Concepts

- Thermodynamics
- Exothermic and endothermic reactions
- Heat of solution

Background

Thermodynamics is the study of energy changes in a system. One important application of thermodynamics in chemistry is the study of heat transfer that accompanies a chemical reaction or a change of state. *Enthalpy* is a measure of the heat content of a system. The transfer of heat into or out of a system results in a change in enthalpy, symbolized by ΔH (delta H). The enthalpy change during a chemical reaction is the difference in the enthalpy of the reactants compared to the enthalpy of the products, according to Equation 1.

$$\Delta H = H_{products} - H_{reactants}$$ *Equation 1*

When discussing the enthalpy change of a system, it is important to understand the difference between the system and the surroundings during a chemical reaction. Typically, the system consists of the reactants and products of the reaction. The solvent, the container, the atmosphere above the reaction (in other words, the rest of the universe) are considered the surroundings.

Exothermic Reactions

When a system releases heat to the surroundings during a reaction, the temperature of the surroundings increases and the reaction vessel feels warm. This type of reaction is called an *exothermic reaction,* where *exo* means "out of" and *thermo* means "heat"—heat flows out of the system. For this type of reaction, the enthalpy of the products is lower than the enthalpy of the reactants, and heat is released. The general form of an exothermic equation is represented in Equation 2.

$$\text{Reactants} \rightarrow \text{Products} + \text{Heat Energy} \qquad \Delta H = \text{negative } (-) \text{ value}$$ *Equation 2*

An example of an exothermic reaction is the combustion of fuel. In the reaction of gasoline with oxygen, for example, energy in the form of heat and flames is produced. Heat flows from the reaction to the surroundings and the surroundings feel hotter.

The reaction pathway (Figure 1) graphically represents the relative enthalpies of the reactants and products in an exothermic reaction. Notice that the products are at a lower overall enthalpy than the reactants due to the loss of heat. The ΔH value is a negative value—characteristic of an exothermic reaction.

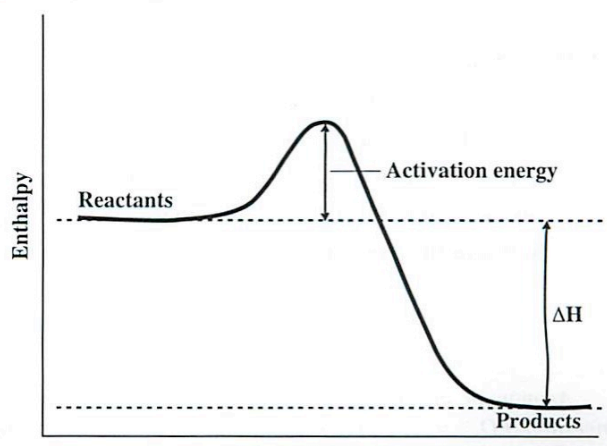

Figure 1. Exothermic Reaction Pathway

Endothermic Reactions

Sometimes a reaction or process requires heat in order to proceed. In this case the system will absorb heat from the surroundings. Since heat is being removed from the surroundings, the reaction vessel will feel cool. This type of reaction is called an *endothermic* reaction, where *endo* means "into" and *thermo* means "heat"—heat flows into the system. For this type of reaction, the enthalpy of the products is higher than the enthalpy of the reactants, and heat is consumed. The general form of an endothermic reaction is represented in Equation 3.

$$\text{Reactants} + \text{Heat Energy} \longrightarrow \text{Products} \quad \Delta H = \text{positive (+) value} \qquad \textit{Equation 3}$$

A common example of an endothermic process is the melting of ice. Solid water (ice) needs heat energy to help break apart the forces holding it together as a solid. Heat from the surroundings flows into the ice, leaving the surroundings with less heat and feeling cooler.

The reaction pathway for an endothermic reaction (Figure 2) shows the products at a higher enthalpy than the reactants, leading to a positive ΔH value—characteristic of an endothermic reaction. The positive ΔH value corresponds to the amount of heat absorbed by the system.

For reactions carried out at a constant pressure, the change in enthalpy (ΔH) of a system is equal to the amount of heat transfer, symbolized by q, and is represented by Equation 4.

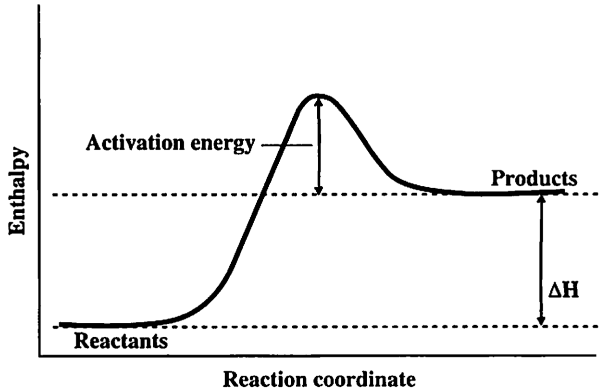

Figure 2. Endothermic Reaction Pathway

$$\Delta H \text{ (at constant pressure)} = q \qquad \textit{Equation 4}$$

Thus, for exothermic reactions, heat is transferred out of a material and the sign of q is negative. For endothermic reactions, heat is absorbed by a material and the sign of q is positive. According to the *Law of Conservation of Energy*, the magnitude of heat released by the system must be equal to the heat absorbed by the surroundings. Incorporating the sign convention, this gives Equation 5.

$$q_{system} = -q_{surroundings} \qquad \textit{Equation 5}$$

The heat, or energy flow, of a system (q) can be determined by multiplying the temperature change of the solution (*ΔT*), the mass of solution (m), and the specific heat capacity (s) of the solution, according to Equation 6.

$$q = m \times \Delta T \times s \qquad \textit{Equation 6}$$

The *specific heat capacity* (or specific heat) of a material is the ability of that material to absorb heat energy; it is defined as the amount of heat needed to raise the temperature of one gram of a substance by one degree Celsius. The SI units for specific heat are given in J/g °C, or cal/g °C. The specific heat capacity of water is 4.184 J/g °C (or 1 cal/g °C). Using Equation 6, the units for heat energy (q) are Joules or calories. To make accurate measurements of heat transfer and to prevent heat loss to the surroundings, an insulating device known as a calorimeter is used. In this laboratory activity, two insulating Styrofoam® cups, one nested inside the other, will be used as the calorimeter.

For each of the three reactions that will be performed in the following lab activity, two materials of known mass will be combined. The temperature change of the mixture will be measured and the heat (q) will be determined. The equation for each reaction will be written and each will be classified as exothermic or endothermic.

Materials

- Ammonium nitrate, NH_4NO_3, 24 g
- Hydrochloric acid solution, 1.0 M, HCl, 50 mL
- Sodium hydroxide solution, 1.0 M, NaOH, 50 mL
- Sodium hydroxide flakes, NaOH, 16 g
- Water, distilled or deionized
- Balance, 0.1 g precision
- Beaker, 400-mL
- Graduated cylinder, 50-mL
- Spatula or scoop
- Styrofoam® cups, 2
- Thermometer

Safety Precautions

Ammonium nitrate is a strong oxidizer and may explode if heated under confinement. Ammonium nitrate is slightly toxic by ingestion, LD_{50} = 2217 mg/kg, and may be a body tissue irritant. Hydrochloric acid solution is toxic by ingestion or inhalation, and is severely corrosive to skin and eyes. Sodium hydroxide, both as solid and in solution, is corrosive and may cause skin burns.

Much heat is evolved when sodium hydroxide is added to water. It is very dangerous to eyes. Avoid all body tissue contact with all chemicals. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory.

Procedure

Part 1. Reaction between HCl and NaOH

1. Nest one Styrofoam cup inside another. Set the nested cups inside a 400-mL beaker for support. This assembly, together with the thermometer, will serve as your calorimeter.
2. Use a graduated cylinder to measure 25 mL of 1.0 M hydrochloric acid solution. Record the exact volume, to the nearest tenth of a milliliter, in Data Table 1. Pour this solution into the calorimeter.
3. Rinse the graduated cylinder and fill it with 25 mL of 1.0 M sodium hydroxide solution. Do not pour this solution into the calorimeter. Record the exact volume, to the nearest tenth of a milliliter, in Data Table 1.
4. Measure the initial temperature (T_i) of the HCl solution (in the calorimeter) and of the NaOH solution (in the cylinder) to the nearest 0.1 °C. Record the temperatures in Data Table 1. (*Note:* Be sure to rinse the thermometer between measurements.)
5. Carefully pour the NaOH solution into the calorimeter. Stir the mixture very gently with the thermometer. (*Caution:* Hold the thermometer at all times; do not allow the thermometer to stand up in the calorimeter as it may tip over.) Record the highest or lowest temperature (final temperature, T_f) that the solution reaches.
6. Rinse the neutral solution down the drain with water. Thoroughly rinse and dry the cups. Repeat Steps 1–5 for Trial II.
7. Rinse the solution down the drain with water. Thoroughly rinse and dry the cups for use in Part 2.

Part 2. Reaction between NH_4NO_3 and H_2O

8. Obtain the two nested Styrofoam cups (the calorimeter) from Part 1, being sure they are completely dry. Weigh out approximately 12 grams of ammonium nitrate into the calorimeter. Record the exact mass to the nearest 0.1 gram in Data Table 2. Set the cups inside a 400-mL beaker for support.
9. Use a graduated cylinder to measure 25 mL of distilled or deionized water. Record the exact volume of the water, to the nearest tenth of a milliliter, in Data Table 2.
10. Measure and record the temperature of the water in the cylinder to the nearest 0.1 °C.
11. Carefully pour the water into the calorimeter containing the ammonium nitrate. Stir the mixture gently with the thermometer until all the solid is dissolved. (*Caution:* Hold the thermometer at all times; do not allow the thermometer to stand up in the calorimeter as it may tip over.) Record the highest or lowest temperature that the solution reaches.
12. Rinse the solution down the drain with water. Thoroughly rinse and dry the cups. Repeat Steps 8–11 for Trial II.
13. Rinse the solution down the drain with water. Thoroughly rinse and dry the cups for use in Part 3.

Part 3. Reaction between NaOH and H_2O

14. Obtain the two nested Styrofoam cups (the calorimeter) from Part 2, being sure they are completely dry. Weigh out approximately 8 grams of sodium hydroxide flakes into the calorimeter. Record the exact mass to the nearest 0.1 gram in Data Table 3. Set the cups inside a 400-mL beaker for support.
15. Use a graduated cylinder to measure 50 mL of distilled or deionized water. Record the exact volume of the water, to the nearest tenth of a milliliter, in Data Table 3.
16. Measure and record the temperature of the water in the cylinder to the nearest 0.1 °C.
17. Carefully pour the water into the calorimeter containing the sodium hydroxide. Stir the mixture gently with the thermometer until all solid is dissolved. (*Caution:* Hold the thermometer at all times; do not allow the thermometer to stand up in the calorimeter as it may tip over.) Record the highest or lowest temperature that the solution reaches.
18. Return the solution to the instructor to be disposed of properly. Thoroughly rinse and dry the cups. Repeat Steps 14–17 for Trial II.
19. Thoroughly rinse and dry the cups and return them to your instructor for use in future labs.

Name: ______________________

Exothermic & Endothermic Reactions

Data Table 1. HCl and NaOH

	Trial I	Trial II	Average
Volume 1.0 M HCl (mL)			
Volume 1.0 M NaOH (mL)			
T_i of HCl before mixing (°C)			
T_i of NaOH before mixing (°C)			
Average T_i before mixing (°C)			
T_f of mixture (°C)			
ΔT (°C)			
Specific Heat (J/g °C)	4.184	4.184	
Heat, q (J)			

Data Table 2. NH_4NO_3 and H_2O

	Trial I	Trial II	Average
Mass of NH_4NO_3 (g)			
Volume of H_2O (mL)			
T_i of H_2O (°C)			
T_f of mixture (°C)			
ΔT (°C)			
Specific Heat (J/g °C)	4.184	4.184	
Heat, q (J)			
Heat per gram (J/g)			

Data Table 3. NaOH and H_2O

	Trial I	Trial II	Average
Mass of NaOH (g)			
Volume of H_2O (mL)			
T_i of H_2O (°C)			
T_f of mixture (°C)			
ΔT (°C)			
Specific Heat (J/g °C)	4.184	4.184	
Heat, q (J)			
Heat per gram (J/g)			

Calculations and Post-Lab Questions

Perform the following calculations and answer the questions on a separate sheet of paper. Show all work and include units.

Reaction 1. HCl and NaOH

1. Determine the average initial temperature for each trial by averaging the initial temperatures of the HCl and the NaOH solutions before mixing. Record as Average T_i before mixing in Data Table 1.
2. Calculate the change in temperature (ΔT) by subtracting the average initial temperature from the final temperature of the mixture ($\Delta T = T_f - T_i$). Record in Data Table 1.
3. Did the temperature rise or fall as the two solutions were mixed? Explain this in terms of heat transfer.
4. Calculate the heat of neutralization (q in joules) for the reaction using Equation 6 ($q = m \times \Delta T \times s$). Assume that the density and specific heat of the solution are the same as that of pure water (density of H_2O = 1.0 g/mL). Record in Data Table 1. Calculate the average heat (q) by averaging Trials I and II. Record in Data Table 1.
5. Classify the reaction as either exothermic or endothermic. Give evidence for your answer.
6. Write the chemical equation for the reaction that took place. Include heat on the appropriate side of the equation.

Reaction 2. NH_4NO_3 and H_2O

7. Calculate the change in temperature (ΔT) for each trial. Record in Data Table 2.
8. Did the temperature of the water rise or fall when the NH_4NO_3 was added? Explain this in terms of heat transfer.
9. Calculate the heat (q in Joules) for the reaction. Record q for each trial in Data Table 2.
10. Calculate the heat absorbed or released per gram of solute added to the water (in joules/g). Record in Data Table 2. Calculate the average heat per gram by averaging Trials I and II. Record in Data Table 2.
11. Classify the reaction as either exothermic or endothermic. Give evidence for your answer.
12. Write the chemical equation for the reaction that took place. Include heat on the appropriate side of the equation.
13. What could be some possible uses for a chemical such as ammonium nitrate?

Reaction 3. NaOH and H_2O

14. Calculate the change in temperature (ΔT) for each trial. Record in Data Table 3.
15. Did the temperature of the water rise or fall when the NaOH was added? Explain this in terms of heat transfer.
16. Calculate the heat (q in Joules) for the reaction. Record q for each trial in Data Table 3.
17. Calculate the heat absorbed or released per gram of solute added to the water (in joules/g). Record in Data Table 3. Calculate the average heat per gram by averaging Trials I and II. Record in Data Table 3.
18. Classify the reaction as either exothermic or endothermic. Give evidence for your answer.
19. Write the chemical equation for the reaction that took place. Include heat on the appropriate side of the equation.

Teacher's Notes

Exothermic & Endothermic Reactions

Materials Included in Kit

Ammonium nitrate, NH_4NO_3, 500 g
Hydrochloric acid solution, 1.0 M, HCl, 800 mL
Sodium hydroxide solution, 1.0 M, NaOH, 800 mL
Sodium hydroxide, flakes, NaOH, 300 g
Styrofoam® cups, 30

Additional Materials Needed (for each lab group)

Balance, 0.1 g precision
Beaker, 400-mL
Graduated cylinder, 50-mL
Spatula or scoop
Thermometer
Water, distilled or deionized

Safety Precautions

Ammonium nitrate is a strong oxidizer and may explode if heated under confinement. Ammonium nitrate is slightly toxic by ingestion, LD_{50} = 2217 mg/kg, and may be a body tissue irritant. Hydrochloric acid solution is toxic by ingestion or inhalation, and is severely corrosive to skin and eyes. Sodium hydroxide, both as solid and in solution, is corrosive and may cause skin burns. Much heat is evolved when sodium hydroxide is added to water. It is very dangerous to eyes. Avoid all body tissue contact with all chemicals. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures governing the disposal of laboratory waste. Ammonium nitrate solution and the neutralized HCl/NaOH solution may be disposed of down the drain according to Flinn Suggested Disposal Method #26b. The sodium hydroxide solution may be disposed of by dilution with water, neutralization, and then by flushing it down the drain according to Flinn Suggested Disposal Method #10.

Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

Unifying Concepts and Processes: Grades K–12
Systems, order, and organization
Evidence, models, and explanation

Content Standards: Grades 5–8
Content Standard B: Physical Science, properties and changes of properties in matter, transfer of energy

Content Standards: Grades 9–12
Content Standard B: Physical Science, structure and properties of matter, chemical reactions, conservation of energy and increase in disorder, interactions of energy and matter

Tips

- Enough materials are provided in this kit for 30 students working in pairs, or for 15 groups of students. All reactions in this laboratory activity can reasonably be completed in one 50-minute class period.
- For calculation purposes, an assumption must be made that the container does not absorb or leak any heat. Two nested Styrofoam® cups are used as the calorimeter; while Styrofoam has insulating properties, some heat may be lost to the environment.
- The acid and base solutions can be treated as if they were pure water with a density of 1.0 g/mL and a specific heat capacity of 4.184 J/g °C. Using this assumption about density, there is no need to measure the mass. However, if it is feasible in your lab to have students weigh the materials, the mass will be more accurate.
- A common confusion among students is the use of the terms system and surroundings. In these experiments, the reactants are the system while the water is the surroundings.

Sample Data

Data Table 1. HCl and NaOH

Volume 1.0 M HCl (mL)	25.0 mL
Volume 1.0 M NaOH (mL)	25.0 mL
T_i of HCl before mixing (°C)	21.1 °C
T_i of NaOH before mixing (°C)	21.5 °C
Average T_i before mixing (°C)	21.3 °C
T_f of mixture (°C)	28.4 °C
ΔT (°C)	7.1 °C
Specific Heat (J/g °C)	4.184
Heat, q (Joules)	–1500 J

Data Table 2. NH_4NO_3 and H_2O

Mass of NH_4NO_3 (g)	13.45 g
Volume of H_2O (mL)	25.0 mL
T_i of H_2O (°C)	20.9 °C
T_f of mixture (°C)	–1.1 °C
ΔT (°C)	–22.0 °C
Specific Heat (J/g °C)	4.184
Heat, q (Joules)	+2300 J
Heat per gram (J/g)	+171 J/g

Data Table 3. NaOH and H_2O

Mass of NaOH (g)	8.58 g
Volume of H_2O (mL)	50.0 mL
T_i of H_2O (°C)	20.9 °C
T_f of mixture (°C)	51.1 °C
ΔT (°C)	30.2 °C
Specific Heat (J/g °C)	4.184
Heat, q (Joules)	–6320 J
Heat per gram (J/g)	–737 J/g

Possible Answers to Questions

Reaction 1. HCl and NaOH

1. Answer: Average initial temperature

   T_i (average) = (21.1 °C + 21.5 °C)/2

   T_i (average) = 21.3 °C

2. Answer: Change in temperature

   $\Delta T = T_f - T_i$ = 28.4 °C − 21.3 °C

   ΔT = 7.1 °C

3. Answer: The temperature increased as the acid and base solutions were mixed.

   Heat was transferred from the system (the neutralization reaction) to the surroundings.

4. Answer: Heat of neutralization

   $q = m \times \Delta T \times s$ = (50.0 g) × (7.1 °C) × 4.184 J/g °C = 1490 J = 1500 J

   Since 1500 J of heat were gained by the surroundings, then 1500 J of heat were lost by the system. Since $q_{system} = -q_{surroundings}$, then

   q = −1500 J

5. Answer: The reaction is exothermic. Heat was released to the surroundings as evidenced by the increase in temperature.

6. Answer: $Na^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \rightarrow H_2O + Na^+(aq) + Cl^-(aq)$ + heat

   Note that the heat energy is shown as a product since it was released during the chemical reaction.

Reaction 2. NH_4NO_3 and H2O

7. Answer: Change in temperature

   $\Delta T = T_f - T_i$ = (−1.1 °C) − (20.9 °C)

   ΔT = −22.0 °C

8. Answer: The temperature decreased when the NH_4NO_3 was added to the water.

   Heat was transferred from the surroundings into the system.

9. Answer: Heat

   $q = m \times \Delta T \times s$ = (25.0 g) × (−22.0 °C) × 4.184 J/g °C = −2300 J

   Since 2300 J of heat were lost from the surroundings, then 2300 J of heat were gained by the system. Since $q_{system} = -q_{surroundings}$, then

   q = +2300 J

10. Answer: Heat per gram of solute

    q/gram = +2300 J/13.45 g = +171 J/g

11. Answer: The reaction is endothermic. Heat was absorbed from the surroundings as evidenced by the decrease in temperature.

12. Answer: $NH_4NO_3(s)$ + heat $\rightarrow NH_4^+(aq) + NO_3^-(aq)$

    Note that the heat energy is shown as a reactant since heat was needed by the system to break up the crystal structure of the salt.

13. Answer: Uses for ammonium nitrate include an instant cold pack or an ice pack. Answers will vary.

Reaction 3. NaOH and H_2O

14. Answer: Change in temperature

$\Delta T = T_f - T_i = (51.1\ °C) - (20.9\ °C)$

$\Delta T = 30.2\ °C$

15. Answer: The temperature increased sharply when the NaOH was added to the water. Heat was transferred from the system to the surroundings, causing an increase in temperature.

16. Answer: Heat

$q = m \times \Delta T \times s = (50.0\ g) \times (30.2\ °C) \times 4.184\ J/g\ °C = 6320\ J$

Since 6320 J of heat were gained by the surroundings, then 6320 J of heat were lost by the system. Since $q_{system} = -q_{surroundings}$, then

$q = -6320\ J$

17. Answer: Heat released per gram of solute

$q/gram = -6320\ J/8.58\ g = -737\ J/g$

18. Answer: The reaction is exothermic. Heat was released to the surroundings as evidenced by the increase in temperature.

19. Answer: $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq) + heat$

Note that the heat energy is shown as a product since it was released during the chemical reaction.

Discussion

Reaction 1

The reaction between an acid solution (HCl) and a base solution (NaOH) is called a neutralization reaction. The heat that is produced during a neutralization reaction is called the heat of neutralization. Since heat is released from the system and transferred to the surroundings, the temperature of the reaction vessel rises, and heat is written as a product of the reaction (See Equation 7 below). The reaction is exothermic and the sign of q is negative.

$$Na^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \rightarrow H_2O + Na^+(aq) + Cl^-(aq) + heat \quad \textit{Equation 7}$$

Since both hydrochloric acid and sodium hydroxide are strong electrolytes, they are written in ionic rather than molecular form, as shown above.

Reaction 2

The second reaction involves the addition of ammonium nitrate to water. This is a very common and useful endothermic process. The process creates an instant cold (ice) pack similar to those used commercially by sports trainers or first aid personnel. It must be pointed out that when ammonium nitrate is added to water, no actual chemical reaction occurs. The process is a simple dissolution process in which the ammonium nitrate (solute) dissolves in the water (solvent) according to Equation 8.

$$NH_4NO_3(s) + heat \rightarrow NH_4^+(aq) + NO_3^-(aq) \quad \textit{Equation 8}$$

Since heat is required to break apart the forces holding ammonium nitrate together as a crystal, heat must be put into the system in order for dissolving to occur. The process is endothermic, heat is written as a reactant, and the sign of q is positive.

Reaction 3

The third reaction involves the addition of sodium hydroxide to water. The reaction is an extremely exothermic reaction, as noted by the immediate and large temperature rise, and the sign of q is negative. Heat is released from the system to the surroundings.

$$NaOH(s \rightarrow Na^+(aq) + OH^-(aq) + heat \quad \textit{Equation 9}$$

The *Exothermic & Endothermic Reactions—An Energy Changes Laboratory Kit* is available from Flinn Scientific, Inc.

Catalog No.	Description
AP6128	Exothermic & Endothermic Reactions—An Energy Changes Laboratory Kit

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.