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Heats of Reaction and Hess's Law

Small-Scale Calorimetry

Introduction

The reaction of magnesium metal with air in a Bunsen burner flame provides a dazzling demonstration of a combustion reaction. Magnesium burns with an intense flame that produces a blinding white light. This reaction was utilized in the early days of photography as the source of "flash powder" and later in flashbulbs. It is still used today in flares and fireworks. How much heat is produced when magnesium burns?

Concepts

- Heat of reaction
- Heat of formation
- Hess's law
- Calorimetry

Background

Magnesium reacts with oxygen in air to form magnesium oxide, according to Equation 1.

$$Mg(s) + ½O_2(g) \rightarrow MgO(s) + \text{heat} \qquad \textit{Equation 1}$$

A great deal of heat and light are produced—the temperature of the flame can reach as high as 2400 °C. The amount of heat energy produced in this reaction cannot be measured directly in the high school lab. It is possible, however, to determine the amount of heat produced by an indirect method, using Hess's Law.

The heat or enthalpy change for a chemical reaction is called the heat of reaction (ΔH_{rxn}). The enthalpy change—defined as the difference in enthalpy between the products and reactants—is equal to the amount of heat transferred at constant pressure and does not depend on how the transformation occurs. This definition of enthalpy makes it possible to determine the heats of reaction for reactions that cannot be measured directly. According to Hess's Law, if the same overall reaction is achieved in a series of steps rather than in one step, the enthalpy change for the overall reaction is equal to the sum of the enthalpy changes for each step in the reaction series. There are two basic rules for calculating the enthalpy change for a reaction using Hess's Law:

- Equations can be "multiplied" by multiplying each stoichiometric coefficient in the balanced chemical equation by the same factor. The heat of reaction (ΔH) is proportional to the amount of reactant. Thus, if an equation is multiplied by a factor of two to increase the number of moles of product produced, the heat of reaction must also be multiplied by a factor of two.
- Equations can be "subtracted" by reversing the reactants and products in the balanced chemical equation. The heat of reaction (ΔH) for the reverse reaction is equal in magnitude but opposite in sign to that of the forward reaction.

Consider the following three reactions:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g) \qquad \textit{Equation A}$$

$$MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l) \qquad \textit{Equation B}$$

$$H_2(g) + ½O_2(g) \rightarrow H_2O(l) \qquad \textit{Equation C}$$

It is possible to express the combustion of magnesium (Equation 1) as an algebraic sum of Equations A, B and C. Applying Hess's Law, therefore, it should also be possible to determine the heat of reaction for Equation 1 by combining the heats of reaction for Equations A–C in the same algebraic manner. *Note:* Chemical equations may be combined by addition, subtraction, multiplication and division.

Experiment Overview

The purpose of this experiment is to use Hess's Law to determine the heat of reaction for the combustion of magnesium (Equation 1). The heats of reaction for Equations A and B will be measured by calorimetry. These heats of reaction will then be combined algebraically with the heat of formation of water (Equation C) to calculate the heat of reaction for the combustion of magnesium.

Pre-Lab Questions

1. Review the *Background* section. Arrange Equations A–C in such a way that they add up to Equation 1.
2. Use Hess's Law to express the heat of reaction for Equation 1 as the appropriate algebraic sum of the heats of reaction for Equations A–C.
3. The heat of reaction for Equation C is equal to the standard heat of formation of water. The heat of formation of a compound is defined as the enthalpy change for the preparation of one mole of a compound from its respective elements in their standard states at 25 °C. Chemical reference sources contain tables of heats of formation for many compounds. Look up the heat of formation of water in your textbook or in a chemical reference source, such as the *CRC Handbook of Chemistry and Physics*.

Materials

Hydrochloric Acid, HCl, 1 M, 60 mL

Magnesium oxide, MgO, 0.40 g

Magnesium ribbon, Mg, 7-cm strip

Balance, centigram (0.01 g precision)

Calorimeter, small-scale

Digital thermometer or temperature sensor

Forceps

Graduated cylinder, 25- or 50-mL

Metric ruler, marked in mm

Scissors

Spatula

Stirring rod

Wash bottle and water

Weighing dish

Safety Precautions

Hydrochloric acid is toxic by ingestion and inhalation and is corrosive to skin and eyes. Magnesium metal is a flammable solid. Keep away from flames. Do not handle magnesium metal with bare hands. Wear chemical splash goggles and chemical-resistant gloves and apron. Wash hands thoroughly with soap and water before leaving the lab.

Procedure

Record all data for Parts A and B in the data table.

Part A. Reaction of Magnesium with Hydrochloric Acid

1. Obtain a 7-cm strip of magnesium ribbon, and cut it into two pieces of unequal length, roughly 3- and 4-cm each. *Note:* Handle the magnesium ribbon using forceps.
2. Measure the exact length of each piece of magnesium ribbon to the nearest 0.1 cm.
3. Multiply the length of each piece of Mg ribbon by the conversion factor (g/cm) provided by your teacher to obtain the mass of each piece of Mg.
4. Mass a clean, dry calorimeter to the nearest 0.01 g.
5. Using a graduated cylinder, add 15 mL of 1 M hydrochloric acid to the calorimeter, and measure the combined mass of the calorimeter and acid.
6. Using a digital thermometer or a temperature sensor, measure the initial temperature of the hydrochloric acid solution to the nearest 0.1 °C.
7. Add the first (shorter) piece of magnesium ribbon to the acid, and stir the solution until the magnesium has dissolved and the temperature of the solution remains constant.
8. Record the final temperature of the solution to the nearest 0.1 °C.
9. Rinse the contents of the calorimeter down the drain with excess water.
10. Dry the calorimeter, and mass it again to the nearest 0.01 g.
11. Repeat steps 5–9 using the second (larger) piece of magnesium ribbon.

Part B. Reaction of Magnesium Oxide with Hydrochloric Acid

12. Mass a clean, dry calorimeter to the nearest 0.01 g.
13. Using a graduated cylinder, add 15 mL of 1 M HCl to the calorimeter, and measure the combined mass of the calorimeter and hydrochloric acid.
14. Tare a small weighing dish, and add about 0.20 g of magnesium oxide. Measure the exact mass of magnesium oxide to the nearest 0.01 g.
15. Using a digital thermometer or a temperature sensor, measure the initial temperature of the hydrochloric acid solution to the nearest 0.1 °C.
16. Using a spatula, add the magnesium oxide to the acid. Stir the reaction mixture until the temperature remains constant for several five-second intervals. Record the final temperature of the solution to the nearest 0.1 °C.
17. Pour the reaction mixture down the drain with excess water. Rinse and dry the calorimeter.
18. Repeat steps 12–17 using a second sample of magnesium oxide.

Name ____________________

Class/Lab Period ____________________

Heats of Reaction and Hess's Law

Data Table

	Reaction A ($Mg + HCl$)		Reaction B ($MgO + HCl$)	
	Trial 1	Trial 2	Trial 1	Trial 2
Mass of Calorimeter (g)				
Mass of Calorimeter + HCl Solution (g)				
Mass of Mg (Reaction A) or MgO (Reaction B) (g)				
Initial Temperature (°C)				
Final Temperature (°C)				

Post-Lab Calculations and Analysis

Show all work on a separate sheet of paper.

Construct a results table to summarize the results of all calculations. For each reaction and trial, calculate the following:

1. Mass of hydrochloric acid solution.
2. Total mass of the reactants.
3. Change in temperature, $\Delta T = T_{final} - T_{initial}$.
4. Heat (q) absorbed by the solution in the calorimeter. *Note:* $q = m \times s \times \Delta T$, where s is the specific heat of the solution in J/g•°C. Use the total mass of reactants for the mass (m) and assume the specific heat is the same as that of water, namely, 4.18 J/g•°C.
5. Number of moles of magnesium and magnesium oxide in Reactions A and B, respectively.
6. Enthalpy change for each reaction in units of kilojoules per mole (kJ/mole).
7. Average enthalpy change (heat of reaction, ΔH_{rxn}) for Reactions A and B. *Note:* The enthalpy change is positive for an endothermic reaction, negative for an exothermic reaction.
8. Use Hess's Law to calculate the heat of reaction for Equation 1. *Hint:* See your answer to *Pre-Lab Question* 2.
9. The heat of reaction for Equation 1 is equal to the heat of formation of solid magnesium oxide.
   a. Look up the heat of formation of magnesium oxide in your textbook or a chemical reference source.
   b. Calculate the percent error in your experimental determination of the heat of reaction for Equation 1.

Teacher's Notes

Heats of Reaction and Hess's Law

Master Materials List

(for 30 students working in pairs)

Hydrochloric acid, HCl, 1 M, 900 mL

Magnesium oxide, MgO, 6 g

Magnesium ribbon, Mg, 7-cm strips, 15

Balances, centigram (0.01 g precision), 3

Calorimeters, small-scale, 15

Digital thermometers or temperature sensors, 15

Forceps, 15

Graduated cylinders, 25- or 50-mL, 15

Metric rulers, 15

Scissors, 3

Spatulas, 15

Stirring rods, 15

Wash bottles and water, 15

Weighing dishes, 30

Preparation of Solution

- Hydrochloric Acid, 1 M: Place about 500 mL of distilled or deionized water in a flask, and add 83 mL of 12 M hydrochloric acid. Stir to mix, then dilute to 1 L. *Note:* Always add acid to water.

Safety Precautions

Hydrochloric acid is toxic by ingestion and inhalation and is corrosive to skin and eyes. Magnesium metal is a flammable solid. Keep away from flames. Do not handle magnesium metal with bare hands. Wear chemical splash goggles and chemical-resistant gloves and apron. Wash hands thoroughly with soap and water before leaving the lab. Consult current 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, and review all federal, state and local regulations that may apply, before proceeding. The final solutions may be disposed of down the drain with excess water according to Flinn Suggested Disposal Method #26b.

Lab Hints

- The experimental work for this lab can reasonably be completed in one 50-minute period. Each trial should take no more than 10 minutes. If time is a concern, consider performing the experiment as a cooperative class activity in which each group performs two trials of either Reaction A or Reaction B. Students calculate the heats of reaction for their two trials and record their results, along with those of the rest of the class, on the board or overhead projector. The class results for both Reactions A and B are averaged, and the average heats of reaction are used to calculate the heat of reaction for Equation 1.

- The remarkable accuracy of the tested results is due to the use of the special, small-scale calorimeters. Expect much greater percent errors in this experiment if conventional coffee-cup calorimeters are used.

- The small-scale calorimeters recommended in the *Materials* section are available from Flinn Scientific (Catalog No. AP5928). These calorimeters are manufactured from dense, 2″-thick polystyrene and lined with a specially formulated coating that permeates the pores of the foam. The lining maximizes heat efficiency and limits water absorption, which means that the calorimeter constant is small and does not change. The maximum temperature recommended for these small-scale calorimeters is 50 °C. Small-scale calorimeters work well with the microscale quantities used in this experiment.

- The thermometer may be used as the stirrer with small-scale calorimeters because the soft foam is not likely to damage any type of thermometer. Advise students, however, not to punch holes or indentations into the bottom or sides of the calorimeter. These indentations may trap liquid and thus interfere with both mixing the solution and drying the calorimeter for multiple trials.

- Because of the small-scale nature of this experiment, the maximum time required for any given trial, including measuring and weighing the reagents, is less than 10 minutes. This factor allows students to carry out several trials in a single class period. The maximum temperatures in this experiment are usually reached within one minute after mixing and are stable for at least 20–30 seconds. The temperatures then begin to decrease at a rate of about 0.1 °C every 20 seconds.

- For best results, use fresh magnesium ribbon. If the metal ribbon appears to be oxidized, rub it with some steel wool to remove the oxide coating. Measure the mass of magnesium ribbon after it has been cleaned, if necessary.

- The best thermometers for small-scale calorimetry are digital electronic thermometers (Flinn Scientific Catalog No. AP8716) or temperature sensors connected to a computer- or calculator-based interface system, such as LabPro™ or CBL-2. Digital thermometers are reasonably inexpensive, update every second and are precise to the nearest ±0.1 °C. Temperature measurements may be a significant source of error in calorimetry experiments. The *Supplementary Information* section contains instructions for adapting this experiment to the use of technology for computer-interface data collection and analysis.

- For greater accuracy in the results, add a correction term to each enthalpy calculation to account for the heat loss to the calorimeter. See the *Supplementary Information* section.

- The procedure may be scaled up if conventional Styrofoam® cup calorimeters are used. If so, use 100 mL of hydrochloric acid per trial, and increase the amount of reagents to 0.5 g of magnesium ribbon and 1.0 g of magnesium oxide in Parts A and B, respectively.
- In order to obtain two significant figures in the calculations involving the mass of Mg, it is necessary to give the students a conversion factor in g/cm for the mass of Mg. Measure the exact length and exact mass of a 20-cm piece of Mg ribbon. Divide the mass of the ribbon by its length to obtain the conversion factor. *Example:* A 20.0-cm piece of Mg weighed 0.19 g. The conversion factor is 0.0095 g/cm. Notice two significant figures are allowed in the recorded mass.

Teaching Tips

- For a review of the basics of calorimetry, see the *Supplementary Information* section.
- Consider leading into this experiment with a visual demonstration of the combustion of magnesium ribbon in a Bunsen burner flame. This simple demonstration will arouse interest and provides a counterbalance against the abstract nature of Hess's Law. When performing the demonstration for the students, instruct them not to look directly or stare at the bright white flame. The bright light emitted by the burning magnesium is UV light and can damage the eyes. Students should look at the flame "out of the corners of their eyes" with their peripheral vision.
- Using the total mass of the reaction mixture (hydrochloric acid solution plus magnesium or magnesium oxide) in the heat equation calculations ($q = m \times s \times \Delta T$) may be confusing to some students. Students may argue that they are measuring the temperature increase in the surroundings, not the system, and thus they should not include the mass of the reactants and products. Using the combined mass of the reaction mixture is traditional in these types of calorimetry exercises and may compensate for the fact that the specific heat of the solution is assumed to be equal to 4.18 J/g•°C, the same as that of water.
- It is very important that students complete the *Pre-Lab Questions* before doing the experiment. We recommend that the teacher check the students' work before allowing them to do the lab. Students who do not understand the Hess's Law calculations will not really "get" the experiment.
- What makes Hess's Law a law? This may be a good time to review with students the definition of a natural law. A law is not engraved in stone in nature—it is the expression of the results of many experiments repeated for many different systems. The "law" is a generalization that has been widely tested and has been found to be true for every reaction that has been tested. Hess's Law is also known as the Law of Additivity of Reaction Heats.
- Bank statements provide an analogy that many teachers find helpful in explaining Hess's Law. Consider two students who have ending bank balances of $150 each. One student may have started with $50 in the account, made three deposits of $50 each and made two withdrawals of $25 each. The other student may have started with $50 and made only one deposit of $150 and one withdrawal of $50. Their beginning and ending bank balances were both the same and did not depend on how they got there.

Answers to Pre-Lab Questions

Student answers will vary.

1. Review the *Background* section. Arrange Equations A–C in such a way that they add up to Equation 1.

$$\mathrm{Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)} \qquad \textit{Equation A}$$

$$\mathrm{MgCl_2(aq) + H_2O(l) \rightarrow MgO(s) + 2HCl(aq)} \qquad - \textit{Equation B}$$

$$\mathrm{H_2(g) + ½O_2(g) \rightarrow H_2O(l)} \qquad + \textit{Equation C}$$

$$\mathrm{Mg(s) + ½O_2(g) \rightarrow MgO(s)} \qquad \textit{Equation 1}$$

2. Use Hess's Law to express the heat of reaction for Equation 1 as the appropriate algebraic sum of the heats of reaction for Equations A–C.

$$\Delta H_A - \Delta H_B + \Delta H_C = \Delta H_1$$

3. The heat of reaction for Equation C is equal to the standard heat of formation of water. The heat of formation of a compound is defined as the enthalpy change for the preparation of one mole of a compound from its respective elements in their standard states at 25 °C. Chemical reference sources contain tables of heats of formation for many compounds. Look up the heat of formation of water in your textbook or in a chemical reference source, such as the *CRC Handbook of Chemistry and Physics*.

   The standard heat of formation of liquid water at 25 °C is equal to –285.8 kJ/mole at 25 °C. (*CRC Handbook of Chemistry and Physics*)

Sample Data

Student data will vary.

Data Table

	Reaction A ($Mg + HCl$)		**Reaction B ($MgO + HCl$)**	
	Trial 1	Trial 2	Trial 1	Trial 2
Mass of Calorimeter (g)	2.38	2.43	2.21	2.21
Mass of Calorimeter + HCl Solution (g)	16.97	16.84	17.23	17.15
Mass of Mg (Reaction A) or MgO (Reaction B) (g)	0.028	0.038	0.20	0.18
Initial Temperature (°C)	21.2	21.2	21.1	21.1
Final Temperature (°C)	30.0	32.4	31.0	30.6

Answers to Post-Lab Calculations and Analysis

Student answers will vary.

Construct a results table to summarize the results of all calculations. For each reaction and trial, calculate the following:

1. Mass of hydrochloric acid solution.

   Sample calculation for Reaction A, Trial 1:
   16.97 g − 2.38 g = 14.59 g

   See sample results table for results of all other calculations.

2. Total mass of the reactants.

   Sample calculation for Reaction A, Trial 1:
   14.59 g + 0.03 g = 14.62 g

   See sample results table for results of all other calculations.

3. Change in temperature, $\Delta T = T_{\text{final}} - T_{\text{initial}}$.

   Sample calculation for Reaction A, Trial 1:
   30.0 °C − 21.2 °C = 8.8 °C

   See sample results table for results of all other calculations.

4. Heat (q) absorbed by the solution in the calorimeter. *Note:* $q = m \times s \times \Delta T$, where s is the specific heat of the solution in J/g•°C. Use the total mass of reactants for the mass (m) and assume the specific heat is the same as that of water, namely, 4.18 J/g•°C.

   Sample calculation for Reaction A, Trial 1:
   q = 14.62 g × 4.18 J/g•°C × 8.8 °C = 540 J

   See sample results table for results of all other calculations.

5. Number of moles of magnesium and magnesium oxide in Reactions A and B, respectively.

   Sample calculation for Reaction A, Trial 1:
   0.028 g Mg × (1 mole/24.3 g) = 1.2×10^{-3} moles Mg

   See sample results table for results of all other calculations.

6. Enthalpy change for each reaction in units of kilojoules per mole (kJ/mole).

   Sample calculation for Reaction A, Trial 1:
   –540 J/(1.2×10^{-3} moles) × 1 kJ/1000 J = –450 kJ/mole

   Notice the negative sign for the enthalpy change for the reaction. Reactions A and B are both exothermic reactions. The heat absorbed by the solution in the calorimeter is equal in magnitude but opposite in sign to the heat released by the reaction. See sample results table for results of all other calculations.

7. Average enthalpy change (heat of reaction, ΔH_{rxn}) for Reactions A and B. *Note:* The enthalpy change is positive for an endothermic reaction, negative for an exothermic reaction.

   Sample calculation for Reaction A:
   $[-450 + (-420)]/2 = -440$ kJ/mole
   See sample results table for results of all other calculations.

8. Use Hess's Law to calculate the heat of reaction for Equation 1. *Hint:* See your answer to *Pre-Lab Question* 2.

   $\Delta H_1 = \Delta H_A - \Delta H_B + \Delta H_C$

   $\Delta H_1 = [-440 - (-130) + (-286)]$ kJ/mole $= -600$ kJ/mole (rounded to two significant figures)

9. The heat of reaction for Equation 1 is equal to the heat of formation of solid magnesium oxide.

   a. Look up the heat of formation of magnesium oxide in your textbook or a chemical reference source.

   The standard heat of formation of solid magnesium oxide at 25 °C is equal to –601.6 kJ/mole. (*CRC Handbook of Chemistry and Physics*)

   b. Calculate the percent error in your experimental determination of the heat of reaction for Equation 1.

$$\text{percent error} = \frac{|\text{experimental} - \text{theoretical}|}{\text{theoretical}} \times 100\% = \frac{|-600 - (-602)|}{602} = 0.3\%$$

Sample Results Table

	Reaction A (Mg + HCl)		Reaction B (MgO + HCl)	
	Trial 1	Trial 2	Trial 1	Trial 2
Mass of Hydrochloric Acid	14.59 g	14.41 g	15.02 g	14.94 g
Total Mass of Reactants	14.62 g	14.45 g	15.22 g	15.12
Temperature Change	8.8 °C	11.2 °C	9.9 °C	9.5 °C
Heat Absorbed by Solution	540 J	676 J	630 J	600 J
Moles of Mg or MgO	0.0012 moles	0.0016 moles	0.0050 moles	0.0045 moles
Enthalpy Change per Mole of Mg or MgO*	–450 kJ/mole	–420 kJ/mole	–126 kJ/mole	–133 kJ/mole
Average Enthalpy Change	–440 kJ/mole		–130 kJ/mole	

* The calculated enthalpy changes are all negative values. These are both exothermic reactions—the heat released by the system resulted in a temperature increase in the surroundings. The sign of the enthalpy change is a common source of student error.

Supplementary Information

Calorimetry Basics

Calorimetry experiments are carried out by measuring the temperature change in water that is in contact with or surrounds the reactants and products. In a typical calorimetry experiment, the reaction of a known mass of reactant is carried out either directly in or surrounded by a known quantity of water and the temperature increase or decrease in the surrounding water is measured. The temperature change (ΔT) produced in the water is related to the amount of heat energy (q) absorbed or released by the reaction system according to the following equation:

$$q = m \times s \times \Delta T$$

where m is the mass of the aqueous solution, s is the specific heat of water and ΔT is the observed temperature change. The specific heat of water is defined as the amount of heat required to increase the temperature of one gram of water by 1 °C and is equal to 4.18 cal/g•°C.

Measuring the Calorimeter Constant of the Calorimeter

When equal volumes of hot and cold water are combined, the new temperature is the average of the two initial temperatures if there is no heat loss to the calorimeter. In actual practice, the new temperature will be slightly less than the average due to heat loss to the calorimeter. Use the following procedure to determine the calorimeter constant in J/°C.

1. Label two calorimeters "cold water" and "warm water." Mass each calorimeter to the nearest 0.01 g. Add 8 mL of cold tap water to the cold water calorimeter, and find its mass to the nearest 0.01 g. Add 8 mL of warm tap water to the warm water calorimeter, and find its mass to the nearest 0.01 g.

2. Stir the water in each calorimeter, and record the initial temperature in each calorimeter to the nearest 0.1 °C (the initial temperature should be stable for at least 20 seconds).

3. Pour the cold water into the warm water calorimeter, and record the resulting final temperature of the mixture to the nearest 0.1 °C.

4. Rinse and dry the calorimeter and repeat steps 1–3 at least once.

5. Calculate the mass of water in each calorimeter.

6. Subtract the initial temperature of the cold and warm water from the final temperature to determine the temperature change for both the cold and warm water (ΔT_{cold} and ΔT_{warm}).

7. Calculate the energy gained by the cold water:
   $q_{\text{cold}} = m_{\text{cold}} \times s \times \Delta T_{\text{cold}}$

8. Calculate the energy lost by the warm water:
   $q_{\text{warm}} = m_{\text{warm}} \times s \times \Delta T_{\text{warm}}$

9. Determine the absolute difference between the energy lost by the warm water and the energy gained by the cold water. Calculate the calorimeter constant in J/°C using the following equation:

$$\text{calorimeter constant} = \frac{|q_{\text{warm}} - q_{\text{cold}}|}{\Delta T_{\text{warm}}}$$

10. For each trial in Parts A and B, multiply the observed temperature change ΔT by the calorimeter constant to determine the heat absorbed by the calorimeter.

11. Calculating the heat released by the exothermic reactions in Parts A and B, add a term to correct for the heat absorbed by the calorimeter:

    heat released by reaction =
    –(heat absorbed by solution + heat absorbed by calorimeter)

Alternative Procedure with Computer-Interface Data Collection and Analysis

The following instructions are provided for adapting the experiment to the use of technology (computer-interface data collection and analysis).

1. Connect the interface (LabPro™, CBL-2 system, etc.) to a computer or calculator.

2. Plug a temperature probe into the interface.

3. Open and set up a graph in your data collection software so the *y*-axis reads temperature in degrees Celsius. Set the minimum and maximum temperature values at 20 degrees and 50 degrees, respectively.

4. The *x*-axis should be set for time in minutes. Set the minimum and maximum time values at 0 seconds and 240 seconds, respectively.

5. The time interval should be set so a temperature reading is taken every 10 seconds.

6. Obtain 15 mL of hydrochloric acid in a graduated cylinder, and carefully transfer the acid to the calorimeter.

7. Place the temperature probe in the acid solution. Allow the probe to equilibrate at the initial temperature for 1 minute, then press Start to begin collecting temperature data.

8. Stir the solution with the temperature sensor, and add the pre-weighed amount of magnesium ribbon or magnesium oxide in Parts A and B, respectively.

9. Continue stirring solution and collecting temperature data for 4 minutes (240 seconds).

10. The system will automatically record data for the allotted time (240 seconds), then stop.

11. Print computer-generated data table and graph, if possible, and use the data to complete the data table and the *Post-Lab Calculations and Analysis* section.