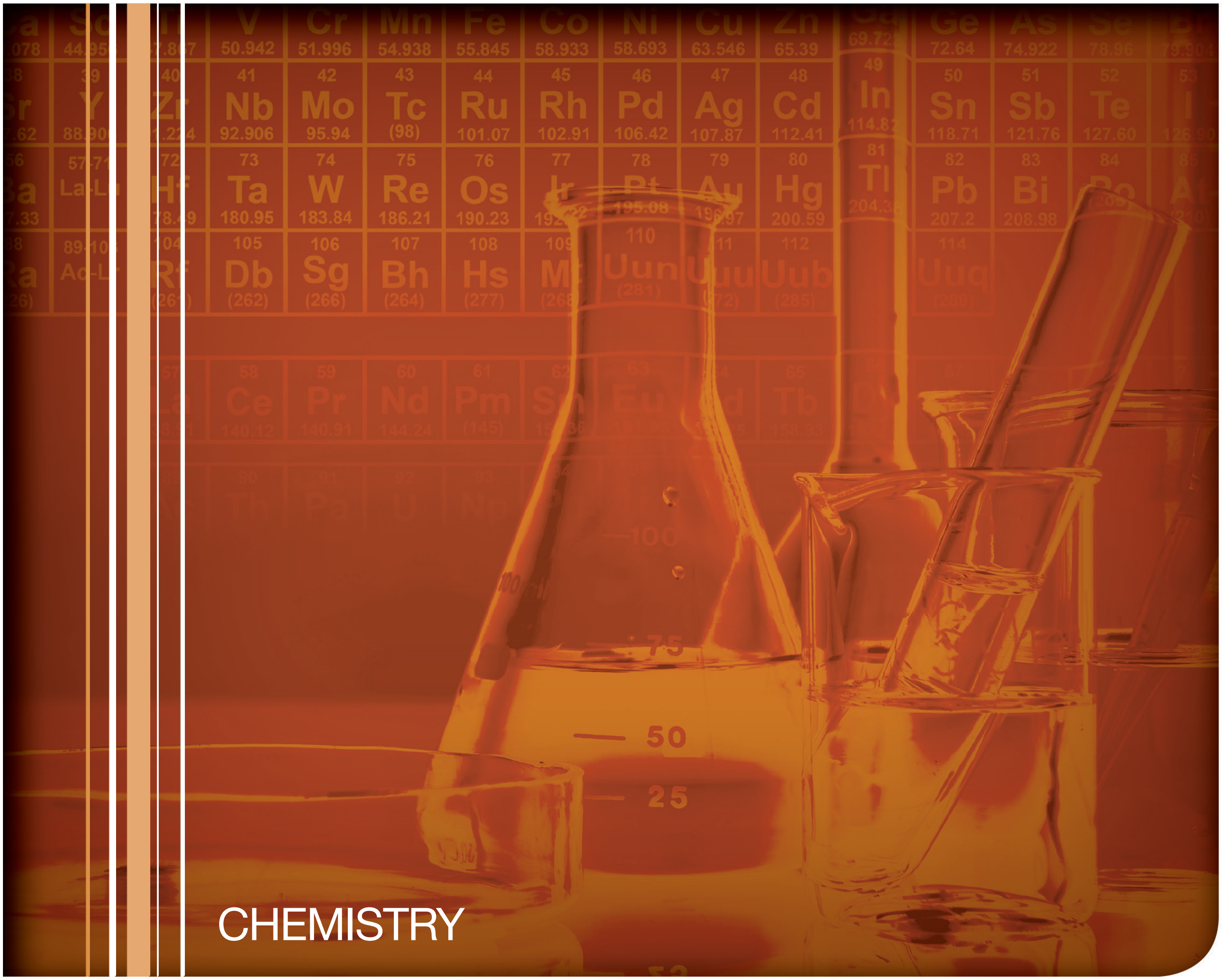

CHEMISTRY

Enthalpy of Formation Using Hess's Law

Investigation Manual

CAROLINA
DISTANCE LEARNING

ENTHALPY OF FORMATION USING HESS'S LAW

Table of Contents

Overview

In this investigation, the student will determine the enthalpy of formation of magnesium oxide using Hess's law. This requires the student to first calculate the calorimeter constant. The student will then perform two separate reactions, determining the heat generated in each reaction to calculate the enthalpy of formation of magnesium oxide.

Outcomes

- Calculate the heat capacity of a calorimeter.
- Deduce the changes in temperature in a chemical reaction.
- Determine the enthalpy of formation of a molecule using Hess's law.

Time Requirements

Preparation .... 5 minutes
Activity 1: Heat Capacity of the Calorimeter .... 45 minutes
Activity 2: Determination of ΔH_{rxn} Mg and HCl .... 45 minutes
Activity 3: Determination of ΔH_{rxn} MgO and HCl .... 45 minutes

Key

Personal protective equipment (PPE)


goggles


gloves


apron


follow link to video


photograph results and submit

stopwatch required


warning


corrosion


flammable


toxic


environment


health hazard

Background

Chemical thermodynamics is the study of energy changes that accompany change of physical state and chemical transformations. These changes occur through the generation or absorption of a quantifiable amount of heat **(q)**. If heat is generated during the chemical reaction or physical change, it is an **exothermic** reaction. If the heat is absorbed or needs to be added to the chemical reaction or physical change, then it is called an **endothermic** reaction.

Scientists use the term **enthalpy (H)** to measure the total energy of a system (in **joules**). Enthalpy also accounts for any change in volume during a reaction, which is important when dealing with gas reactions. The equation for the enthalpy change of a reaction is:

$$\Delta H = \Delta q + P\Delta V$$

Where ΔH is the change in enthalpy, Δq is the heat change (generation or absorption), P is pressure, and ΔV is the change in volume. In most calorimeters, pressure and volume are constant. Therefore, the equation becomes:

$$\Delta H = \Delta q$$

Enthalpy is dependent on the amount of substance present. The enthalpy change for a reaction is generally written as a molar quantity. To calculate the molar enthalpy of a reaction (**ΔH_{rxn}**), divide the reaction heat, (**Δq_{rxn}**) by the number of moles of reactant or product.

$$\Delta H_{rxn} = \Delta q_{rxn}/mol$$

Every chemical reaction either absorbs or emits heat when chemical bonds are formed or broken. Each chemical compound has a heat change associated with its formation. This heat change is **the standard enthalpy of formation (ΔH_f^o),** and is the heat associated with the formation of one mole of a compound at one atmosphere of pressure at 25 °C.

In this investigation, the standard enthalpy of formation (ΔH_f^o) will be determined for magnesium oxide (MgO). The overall equation for the formation of MgO is as follows:

$$Mg(s) + \frac{1}{2} O_2(g) \rightarrow MgO(s)$$

By definition the enthalpy of reaction is the sum of the enthalpies of formation of the products minus the sum of the enthalpies of formation of the reactants.

$$\Delta H_{rxn} = \Sigma\, \Delta H_f\, (Products) - \Sigma\, \Delta H_f\, (Reactants)$$

Pure elements, such as magnesium and oxygen, are assigned a ΔH_f of 0. The reactants in the formation of magnesium oxide are pure elements, meaning that the enthalpy of reaction will be equal to the enthalpy of formation for magnesium oxide. However, the reaction of solid magnesium with oxygen gas is not possible in normal lab conditions and would not be safe in a home environment. An alternative method for determining the ΔH_f^o for magnesium oxide is thus required. Fortunately, Hess's law states that a reaction can be performed in steps, and the sum of the ΔH_{rxn}^o for each of those steps is equal to the ΔH_f^o for the compound in question (in this case, magnesium oxide). The synthesis of magnesium oxide is performed in the following three steps:

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Background continued

Equation 1	$Mg(s) + 2\,H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$	ΔH_1
Equation 2	$Mg^{2+}(aq) + H_2O(l) \rightarrow MgO(s) + 2\,H^+(aq)$	ΔH_2
Equation 3	$H_2(g) + \frac{1}{2}\,O_2(g) \rightarrow H_2O(l)$	ΔH_3
Overall Equation	$Mg(s) + \frac{1}{2}\,O_2(g) \rightarrow MgO(s)$	$\Delta H_f = \Delta H_1 + \Delta H_2 + \Delta H_3$

By adding equations 1, 2, and 3, and crossing out molecules found on both sides of the equation, the overall equation is obtained. Summation of their enthalpies gives the enthalpy of formation for MgO.

This investigation will be conducted using calorimetry. **Calorimetry** is the method by which heat energy is measured. The magnitude and the direction of heat transfer can be determined using a **calorimeter**, an insulated apparatus designed to create a closed system by preventing heat from flowing in or out. In this lab, the calorimeter is composed of two test tubes that are nested together with paper towels sandwiched between them. The paper towel between the two test tubes and the air surrounding the test tubes will insulate the system.

No calorimeter is a perfectly insulated system, and a fraction of the energy is absorbed by the calorimeter. The joules of energy absorbed per 1 °C temperature change represent the **heat capacity** (C) of the calorimeter. In Activity 1, the heat capacity of the calorimeter will be determined by analyzing the temperature change of cold water in the calorimeter when hot water is added. The difference between the heat lost by the hot water and the heat gained by the cold water will be due to the heat absorbed by the calorimeter (i.e., the heat capacity of the calorimeter).

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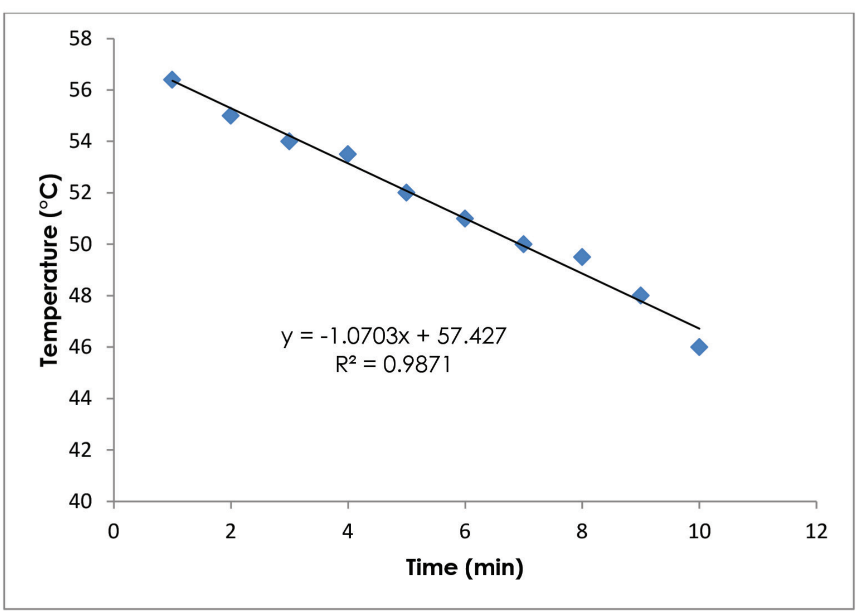
Figure 1.



A calibration curve is used to determine the heat capacity of the calorimeter. The calibration curve is a plot of the temperature (in °C) of the water in the calorimeter as a function of time. Plotting the temperature as a function of time will determine the cooling rate of the mixture. The initial mixing temperature is accurately calculated by fitting the data with a linear trend line using Excel™. Figure 1 shows a calibration curve generated from adding 50 g of water measuring 96 °C to 50 g of water measuring 24 °C. The initial temperature of the mixture should have been 60 °C [(96 + 24)/2 = 60]. The y-intercept from the linear trend curve is 57.4 °C and represents T_{final}.

$$\Delta T_{(hot\ water)} = 57.4\ °C - 96\ °C = -38.6\ °C$$

$$\Delta T_{(cold\ water)} = 57.4\ °C - 24\ °C = 33.4\ °C$$

Δq is calculated by multiplying the heat capacity (C) of the substance by the change in temperature and the mass of the substance. The energy lost by this reaction ($\Delta q_{(hot\ water)} + \Delta q_{(cold\ water)}$) will be gained by the calorimeter ($\Delta q_{(cal)}$). Each calorimeter has a unique heat capacity based on the material of the calorimeter and the mass of solution within the calorimeter. In this investigation, the same volume of liquid will be used for each activity. Since the mass is constant for each activity, the mass of the calorimeter is not relevant to calculating $\Delta q_{(cal)}$.

$$\mathbf{\Delta q = C \times \Delta T \times m}$$

$$\mathbf{\Delta q_{(cal)} = C_{(cal)} \times \Delta T}$$

$$\mathbf{C_{water} = 4.18\ J/g\ °C}$$

$$\Delta q_{(hot\ water)} = 4.18\ J/g\ °C \times -38.6\ °C \times 50\ g = -8067\ J$$

$$\Delta q_{(cold\ water)} = 4.18\ J/g\ °C \times 33.4\ °C \times 50\ g = 6981\ J$$

Energy lost to calorimeter = -1086 J

$$\Delta q(cal) = 1086\ J = C(cal) \times 33.4\ °C$$

$$C_{cal} = 32.5\ J/°C$$

For each reaction, $\mathbf{\Delta q_{rxn}}$ is calculated as the sum of the heat absorbed by the solution and the heat absorbed by the calorimeter.

$$\mathbf{\Delta q_{rxn} = \Delta q_{soln} + \Delta q_{(cal)}}$$

In this investigation, it will be assumed that the heat capacity of the solution is the same as that of water (4.18 J/g-K).

In Activity 2, students will react magnesium metal with hydrochloric acid, as described in Equation 1. Students will measure the temperature change and calculate the heat generated during this reaction. In Activity 3, students will react magnesium oxide with hydrochloric acid. This reaction is the reverse of Equation 2, so the $\mathbf{\Delta H_2}$ will have the opposite sign when calculations are performed.

The reaction described in Equation 3 will not be performed because the heat of reaction of hydrogen and oxygen gases cannot be measured using a simple calorimeter such as the one that will be used in this activity. Instead, students will use the accepted value for the enthalpy change of this reaction (–285.8 kJ/mol) when calculating $\mathbf{\Delta H_f}$.

ENTHALPY OF FORMATION USING HESS'S LAW

Materials

Included in the materials kit:

Magnesium metal

Styrofoam® cup

Magnesium oxide powder

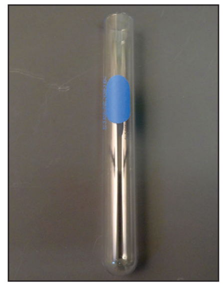
Hydrochloric acid, 1 M

Needed but not supplied:
- Paper towel
- Stopwatch or other timing device

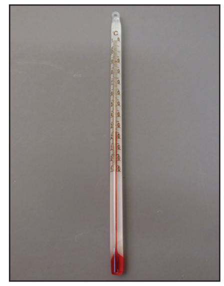
Reorder Information: Replacement supplies for the Enthalpy of Formation Using Hess's Law investigation can be ordered from Carolina Biological Supply Company, kit 580348.

Call 800-334-5551 to order.

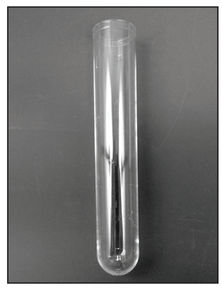
Needed from the equipment kit:



Test tube, 13 x 100 mm

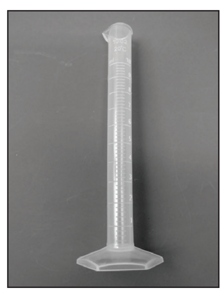


Thermometer



Test tube, 17 x 100 mm

Balance



Graduated cylinder, 10 mL

Weigh boat

Safety

Wear your safety goggles, chemical apron, and gloves at all times while conducting this investigation.

Read all the instructions for this laboratory activity before beginning. Follow the instructions closely and observe established laboratory safety practices, including the use of appropriate personal protective equipment (PPE) described in the Safety and Procedure section.

Magnesium and magnesium oxide are flammable. Keep both chemicals away from any heat or flame sources.

Magnesium, magnesium oxide, and hydrochloric acid are ingestion hazards. Keep chemicals away from food storage locations to avoid accidental ingestion.

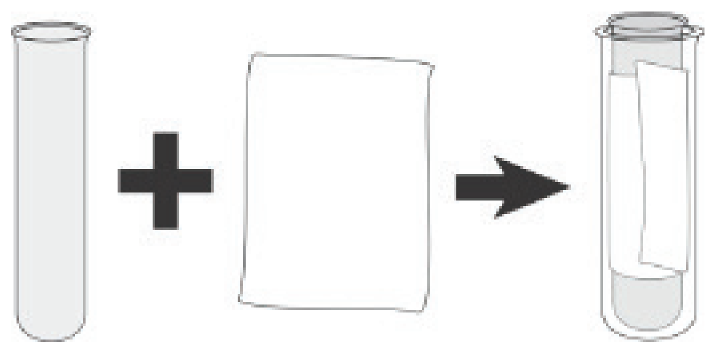
Hydrochloric acid is corrosive. In the event of contact with skin or eyes, the affected area should be immediately rinsed with water for 15 minutes.

Do not eat, drink, or chew gum while performing this investigation. Wash your hands with soap and water before and after performing the investigation. Clean up the work area with soap and water after completing the investigation. Keep pets and children away from lab materials and equipment.

Preparation

1. Read each activity description thoroughly.
2. Clean the workspace.
3. Gather all needed materials.
4. Take a 13 x 100 mm test tube and wrap it tightly in a paper towel, leaving the opening exposed.
5. Place the wrapped test tube inside a 17 x 100 mm test tube. This will complete construction of the calorimeter that will be used for the activities.
6. Be prepared to record data: print out data tables.

Figure 2.



ACTIVITY 1

A Heat Capacity of the Calorimeter

1. Measure 2.0 mL of water with a graduated cylinder and pour into the inner test tube of the calorimeter. This is the cold water.

> Remember that 1 mL of water = 1 g of water.

2. Place the calorimeter in the test tube rack for stability.
3. Stir the water with the thermometer until temperature is constant, and record it as the initial temperature of the cold water (Tc) in Data Table 1.
4. Fill the Styrofoam® cup 1/4 full with the hottest tap water possible.
5. Measure the temperature of the hot water. If the temperature is less than 75 °C, heat in a microwave with short bursts until the temperature reaches 75 °C.
6. Measure 2.0 mL of the hot water with a graduated cylinder.
7. Place a thermometer in the cylinder and gently stir until you obtain a constant temperature reading. The temperature should be 75–80 °C. Record this as the initial temperature of the hot water (Th) in Data Table 1.
8. Prepare the stopwatch or timer.
9. Quickly pour the 2.0 mL of hot water into the calorimeter containing the cold water and start a timer.
10. Gently stir the mixed water with the thermometer and record the temperature at 1-minute intervals for 10 minutes in Data Table 1.
11. Discard the water, dry the inner test tube with a paper towel, and place it back in the calorimeter.
12. Repeat this activity two additional times for a total of three trials.
13. Calculate the heat capacity, as described on page 5 of the Background, of the calorimeter for each of the three trials and record the value in Data Table 1.
14. Calculate the average heat capacity of the calorimeter and record the value in Data Table 1.

ACTIVITY 2

A Determination of ΔH_{rxn} Mg and HCl

1. Place the calorimeter in the test tube rack for stability.
2. Measure 4 mL of 1 M HCl with a graduated cylinder and pour it into the calorimeter.
3. Stir the HCl with a thermometer until the temperature is constant.
4. Record the initial temperature of the HCl in Data Table 2.
5. Place a weigh boat on the scale and tare (zero) the balance. Collect approximately 0.06 g of the magnesium metal strip and record the mass in Data Table 2.
6. Calculate the moles of magnesium (24.305 grams per mole) in the sample and record in Data Table 2.
7. Prepare the stopwatch or timer.
8. Place the magnesium strip into the calorimeter with the HCl and immediately start the timer.
9. Record the temperature at 1-minute intervals for a total of 10 minutes in Data Table 2.
10. Periodically stir the test tube with the thermometer as the reaction occurs, taking care not to spill any of the solution.
11. After 10 minutes, empty the contents into the sink with running water and rinse and dry the calorimeter.
12. Repeat this activity two additional times for a total of three trials.
13. For each trial, create a graph with temperature on the Y-axis and time on the X-axis.
14. Using a graphing software, calculate the linear regression for each graph utilizing only the temperatures from the portion of the graph where temperature is decreasing.
15. Record the equation of each linear regression line obtained with the graphing software in step 14 in Data Table 2.
16. The y-intercept of each line represents the final temperature for each corresponding trial. Record the final temperature in Data Table 2.
17. Calculate the temperature change (ΔT) for each trial and record in Data Table 2.
18. Calculate change in reaction heat (Δq_{rxn}) for each trial using the calorimeter constant calculated in Activity 1.
19. Calculate the average change in enthalpy (ΔH_1).

ACTIVITY 3

A Determination of ΔH_{rxn} MgO and HCl

1. Place the calorimeter in the test tube rack for stability.
2. Measure 4 mL of 1 M HCl with a graduated cylinder and pour it into the calorimeter.
3. Stir the HCl with a thermometer until the temperature is constant
4. Record the initial temperature in Data Table 3.
5. Place a weigh boat on the scale and tare (zero) the balance. Collect approximately 0.09 g of magnesium oxide and record the mass in Data Table 3.
6. Calculate the moles of magnesium oxide in the sample and record in Data Table 3.
7. Prepare the stopwatch or timer.
8. Place the magnesium oxide into the calorimeter with the HCl and immediately start the timer.
9. Record the temperature at 1-minute intervals for a total of 10 minutes in Data Table 3.
10. Periodically stir the test tube with the thermometer as the reaction occurs, taking care not to spill any of the solution.
11. After 10 minutes, empty the contents into the sink with running water and rinse and dry the calorimeter.
12. Repeat the procedure two additional times for a total of three trials.
13. For each trial, graph the temperature vs. time.
14. Calculate the linear regression for each graph utilizing only the temperatures from the portion of the graph where temperature is decreasing.
15. Record the equation of each linear regression line in Data Table 3.
16. The y-intercept of each line represents the final temperature for each corresponding trial. Record the final temperature in Data Table 3.
17. Calculate ΔT for each trial and record in Data Table 3.
18. Calculate Δq for each trial using the calorimeter constant calculated in Activity 1.
19. Calculate the average ΔH_2.
20. Calculate ΔH_f.

Disposal and Cleanup

1. Dispose of solutions by pouring them into the sink with the water running. Allow the water to run for a few minutes to dilute the solutions.
2. Rinse and dry the lab equipment and return the materials to your equipment kit.
3. Sanitize the workspace.

ACTIVITY

Data Table 1: Heat Capacity of the Calorimeter

Time (min)	Trial 1 Temp. °C	Trial 2 Temp. °C	Trial 3 Temp. °C
1			
2			
3			
4			
5			
6			
7			
8			
9			
10			

	Trial 1	Trial 2	Trial 3
Initial temperature of cold water, T_c			
Initial temperature of warm water, T_h			
Average temperature, T_a $(T_h + T_c)/2$			
Temperature at time 0 from graph, T_0			
Heat lost by hot water			
Heat gained by cold water			
Heat lost by water to the calorimeter in J			
Temperature change of the calorimeter			
Heat capacity (C) of the calorimeter in J/°C			
Average heat capacity (C) of calorimeter in J/°C			

Data Table 2: Determination of H_{rxn} Mg and HCl

	Trial 1	Trial 2	Trial 3
Mass of magnesium (g)			
Moles of magnesium			
Initial temperature			
T at 1 minute			
T at 2 minutes			
T at 3 minutes			
T at 4 minutes			
T at 5 minutes			
T at 6 minutes			
T at 7 minutes			
T at 8 minutes			
T at 9 minutes			
T at 10 minutes			
Equation of line			
Final temperature			
ΔT			
Δq_{soln}			
Δq_c			
ΔH_1 (kJ/mol)			
Average ΔH_1			

ACTIVITY

Data Table 3: Determination of H_{rxn} MgO and HCl

	Trial 1	Trial 2	Trial 3
Mass of magnesium oxide (g)			
Moles of magnesium oxide			
Initial temperature			
T at 1 minute			
T at 2 minutes			
T at 3 minutes			
T at 4 minutes			
T at 5 minutes			
T at 6 minutes			
T at 7 minutes			
T at 8 minutes			
T at 9 minutes			
T at 10 minutes			
Equation of line			
Final temperature			
ΔT			
Δq_{soln}			
Δq_c			
ΔH_2 (kJ/mol)			
Average ΔH_2			
Average ΔH_1			
ΔH_3	–286 kJ/mol		
H_f of magnesium oxide			

NOTES

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