

# HEATS OF REACTION AND HESS'S LAW

## Introduction:

When reactions take place, energy in the form of heat is either absorbed (endothermic) or released (exothermic). The amount of heat involved in a reaction depends upon the particular reaction studied or the amount of reactant that is used. Sometimes it is difficult to measure the heat of a reaction directly. Hess's Law tells us that the amount of heat involved in a reaction is the same, whether the reaction takes place in one or several steps. A modification of this law permits us to find the heat of one step of a reaction if the overall heat and the heats for the other steps are known. In this experiment, you will determine the heat of reaction for magnesium oxide by studying a series of reactions of magnesium and magnesium oxide.

## Purpose:

The purpose of this experiment is to calculate the heat of formation of magnesium oxide using Hess's Law.

## Equipment/Materials:

plastic cup calorimeter	1.0 M HCl
50 mL beaker	magnesium ribbon
thermometer	magnesium oxide
balance	

## Safety:

- An apron and goggles should be worn at all times in the lab.

**Procedure:**

1. Obtain a clean piece of magnesium. Measure the length of the ribbon accurately. Using the conversion factor on the label of the magnesium container, calculate the mass of the magnesium used.
2. Add 10.0 mL of the HCl to the plastic cup calorimeter. The calorimeter is made by nesting one plastic cup inside another with a rubber band spacer between them. The calorimeter should be placed in a 50 mL beaker to increase its stability. Measure the temperature of the acid. Record the temperature of the acid to the nearest 0.1 degree.
3. Drop the magnesium ribbon into the acid and gently swirl the cup and its contents. When all of the magnesium has reacted, record the temperature of the solution. Repeat the procedure.
4. Assuming the heat capacity of the HCl is the same as that of water, calculate the amount of heat liberated when the magnesium reacted. You should recall that the formula for calculating heat is:  
$$\text{Heat} = \text{mass} \times \text{temperature change} \times \text{specific heat.}$$
5. Calculate the number of moles of magnesium that reacted. Calculate the heat of reaction that would have been observed if 1.0 mole of magnesium had reacted. Average the results for the trials.
6. Weigh out to the nearest milligram about 0.05 grams of MgO. Again, add 10.0 mL of the HCl to the plastic cup calorimeter. Measure and record the temperature of the acid solution. Carefully transfer the MgO to the cup. Gently stir the reaction mixture, and record the temperature when it is constant. Note: This reaction does not evolve as much heat as did the Mg and the HCl.

7. Again, assuming the heat capacity of the acid is the same as that of water, calculate the amount of heat liberated when the MgO reacted.
  
8. Calculate the number of moles of magnesium oxide that reacted. Determine the amount of heat that would have been released if 1.0 mole of MgO had reacted. Average the values for the trials.
  
9. A third heat of reaction is required. Refer to the table of heats of formation in your textbook and find the amount of heat that would be released if one mole of water (liquid) were formed.

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### Data:

Conversion factor for magnesium ribbon \_\_\_\_\_ g/meter

### Part 1

	Trial 1	Trial 2
Initial temp. of HCl		
Final temp. of HCl		
Change in temperature		
Volume of HCl		
Mass of HCl		
Heat released		
Mass of Mg		
Moles of Mg		
Heat/Mole Mg		

### Part 2

	Trial 1	Trial 2
Initial temp. of HCl		
Final temp. of HCl		
Change in temperature		
Volume of HCl		
Mass of HCl		
Heat released		
Mass of MgO		
Moles of MgO		
Heat/Mole MgO		

**Part 3:**

1. Heat of formation for H<sub>2</sub>O (liquid) \_\_\_\_\_

**Calculations:**

1. Write the equations for the three reactions used in this experiment. Include the heat of reaction in each equation. Add the reactions, reversing any if necessary, to determine the heat of reaction for:  $\text{Mg}_{(s)} + \frac{1}{2} \text{O}_{2(g)} \rightarrow \text{MgO}_{(s)}$

2. What is the accepted value for this reaction? What is the percent error for your determination?