

Heats of Reaction and Hess's Law

Small-Scale Calorimetry

Introduction

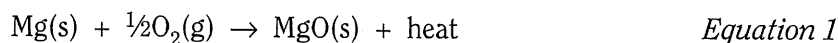
The reaction of magnesium metal with air in a Bunsen burner flame provides a dazzling demonstration of a combustion reaction. Magnesium burns with an intense flame that produces a blinding white light. This reaction was utilized in the early days of photography as the source of “flash powder” and later in flashbulbs. It is still used today in flares and fire-works. 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.

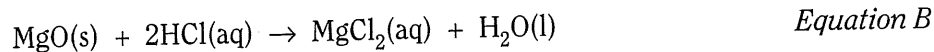
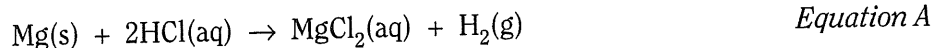


As mentioned above, a great deal of heat and light are produced—the temperature of the flame can reach as high as 2400 °C. The amount of heat energy produced in this reaction cannot be measured directly in the high school lab. It is possible, however, to determine the amount of heat produced by an indirect 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, then 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:



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. The heats of reaction for these two reactions 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	Graduated cylinder, 25- or 50-mL
Magnesium ribbon, Mg, 7-cm strip	Metric ruler, marked in mm
Magnesium oxide, MgO, 0.40 g	Scissors
Balance, centigram (0.01 g precision)	Spatula
Calorimeter, small-scale	Stirring rod
Digital thermometer or temperature sensor	Wash bottle and water
Forceps	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 12–16 using a second sample of magnesium oxide.
19. Wash the contents of the calorimeter down the drain with excess water.

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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:

- Mass of hydrochloric acid solution.
- Total mass of the reactants.
- Change in temperature, $\Delta T = T_{\text{final}} - T_{\text{initial}}$.
- 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 $J/g \cdot ^\circ C$. Use the total mass of reactants for the mass (m) and assume the specific heat is the same as that of water, namely, $4.18 J/g \cdot ^\circ C$.
- Number of moles of magnesium and magnesium oxide in Reactions A and B, respectively.
- Enthalpy change for each reaction in units of kilojoules per mole (kJ/mole).
- 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.
- Use Hess's Law to calculate the heat of reaction for Equation 1. *Hint:* See your answer to PreLab Question #2.
- The heat of reaction for Equation 1 is equal to the heat of formation of solid magnesium oxide.
 - Look up the heat of formation of magnesium oxide in your textbook or a chemical reference source.
 - Calculate the percent error in your experimental determination of the heat of reaction for Equation 1.