EXPERIMENT 6

KEEP YOUR EYES ON THE JOULES

A Hess's Law Experiment

Contents:

ACKNOWLEDGEMENTS

The creation of this experiment and its support materials would not have been possible without the scientific and pedagogical expertise of dedicated educators active in the field. Adam Equipment extends both acknowledgement and appreciation to the following teachers for their assistance in making this classroom activity available to the education community:

Penney Sconzo (Westminster High School, Atlanta, GA) – project leader and experiment author

Adrian Dingle (Westminster High School, Atlanta, GA) - procedure verification and reviewer

Maureen Miller (Westminster High School, Atlanta, GA) – peer reviewer and advisor

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Teacher Guide

KEEP YOUR EYES ON THE JOULES

A Hess's Law Experiment

AGE LEVEL

This experiment is designed for ages 15-18. It can be adapted for physical science or introductory chemistry classes, but it is most appropriate for advanced classes.

SUBJECTS

1. Applying the Law of Conservation of Energy to Hess's Law
2. Enthalpy of a reaction (ΔH) and thermochemical equations
3. Calculating the energy change for a reaction when the energy change cannot be measured directly
4. Energy diagrams for reactions
5. State functions

PURPOSES

1. To show how the enthalpy of a reaction changes when the reaction's is modified
2. To apply the Law of Conservation of energy to a series of related reactions
3. To use an energy diagram to explain Hess's Law
4. To interpret Hess's Law

TIME NEEDED

One laboratory period.

ACTIVITY OVERVIEW

The heat of formation of magnesium oxide is not easily measured in the laboratory. In order to obtain the enthalpy for forming magnesium oxide, two related reactions will be conducted in the lab. Summing the equations leads to the net overall equation, $2Mg + O2 \rightarrow 2MgO$. Hess's Law states that if a series of equations can be summed to give the net equation for a reaction, then the energy change associated with each equation can also be summed to give the net energy change for the overall reaction.

This lesson is designed to be easily and quickly integrated into existing curriculum. It comes complete with a two-page student hand-out, standards correlations, vocabulary lists, plus extensions and assessment tools with answer keys. This experiment is appropriate for remedial, review, reinforcement or extension purposes.

HELPFUL ADVICE TO MAXIMIZE SAFETY AND STUDENT SUCCESS

- As always, students must wear safety goggles during any laboratory experiment.
- Instruct students on how to measure a small quantity of a loose powder like MgO. MgO is a basic anhydride.
- Exercise caution with magnesium oxide near any part of the face (eyes, nose, mouth).
- Wash hands after completing the lab. If any powder is spilled, vinegar can be used to neutralize the base.
- Discuss the accuracy of the balances used for this experiment.

SCIENCE SKILLS AND ABILITIES

SCIENCE AS INQUIRY

Abilities necessary to do scientific inquiry:

1. Formulate scientific explanations and models using logic and evidence (ages 14-18).
2. Use technology and mathematics to improve investigations and communications (ages 10-18).

PHYSICAL SCIENCE

Understanding chemical reactions (ages 14-18) and understanding conservation of energy and increase in disorder (ages 14-18).

DATA ANALYSIS, PROBABILITY AND DISCRETE MATHEMATICS

Understanding and applying data collection, organization and representation to analyze and sort data (ages 10-18).

GEOMETRY AND MEASUREMENT

Understand and apply appropriate units of measure, measurement techniques, and formulas to determine measurements (ages 10-18).

PURPOSE

Hess's Law states that when two or more reactions can be added together to give a net chemical reaction, the energy associated with each of the reactions can also be added together to give the net enthalpy (ΔH) change. The law can be used to obtain thermodynamic data that cannot be measured directly. In this lab, students will become familiar with how chemical reactions and the energy of each of those chemical reactions can be summed up to give the overall change in energy for the net chemical equation.

BACKGROUND INFORMATION

The first law of thermodynamics states there is a fixed amount of energy in a system. New energy cannot be created. No energy can be destroyed. However, energy can be converted from one form of energy to another. The thermodynamic energy, or heat, in a system is known as enthalpy, H. What can be measured in the laboratory is the change in energy, ΔH.

During a chemical reaction, changes in energy can be monitored by measuring changes in volume, temperature and pressure. These variables are not always easy to measure in reactions. Suppose you wanted to find the heat of a reaction, but you could not perform the reaction. For example,

- The reactants are unstable and there could be an accident in the lab.
- A protective oxide layer forms on a reactant and the formation of product ceases.
- The reaction is too slow to be measured. It is difficult to form the product you desire. The reaction you are investigating goes to completion and you are interested in the heat of forming an intermediate in the pathway. For example, it is very difficult to control the burning (oxidation) of graphite to give pure carbon monoxide, CO. The reaction for the burning of graphite to form carbon dioxide, CO_2 is easily performed and the enthalpy change easily measured. Stopping the reaction in midstream when CO has formed is difficult. The reaction wants to proceed to form carbon dioxide when enough oxygen is present.
- The enthalpy change for the reaction cannot be measured directly. It is too hot to measure the formation of magnetic iron oxide (Fe_3O_4).

What can you do? It is often possible to calculate the enthalpy for a reaction by summing ΔH values of other reactions. This way, you can avoid doing the experiment. You perform reactions in the lab that are step reactions which can be combined to yield the reaction you are avoiding. By performing arithmetic operations on these thermochemical equations that can be carried out in the lab, there is a practical way to find the heat of that difficult reaction. Use Hess' Law which states the heat involved in a chemical process is the same whether it takes place in one step or multiple steps. First proposed by Henri Hess in 1840, this is sometimes referred to as the law of constant heat summation and is a consequence of the law of conservation of energy.

Never expect more or less energy from a chemical reaction by changing the method of carrying out the reaction. A chemist might take a sample of food and burn it to in order to measure the amount of heat given off. Upon consumption, the same amount of heat will be given off in the human body when that food sample is metabolized.

The thermochemical equations used are just like other balanced equations except they also specify the heat for the reaction. The heat flow is listed just to the right of the equation using the symbol ΔH. Kilojoules, kJ, is a scientific unit for heat.

The heat flow of a reaction is directly proportional to the amount of material undergoing the reaction.

$$C(graphite) + 0.5\ O_2 \rightarrow CO, \qquad H° = -110\ kJ/mol$$
$$2\ C(graphite) + O_2 \rightarrow 2\ CO, \qquad \Delta H° = -220\ kJ/mol\ (multiplied\ by\ 2)$$
$$6\ C(graphite) + 3\ O_2 \rightarrow 6\ CO, \qquad \Delta H° = -660\ kJ/mol\ (multiplied\ by\ 6)$$

The heat of a reaction is equal in magnitude but opposite in sign for the reverse reaction.

$$CO \rightarrow C(graphite) + 0.5\ O_2, \qquad \Delta H° = 110\ kJ/mol$$
$$2\ CO \rightarrow 2\ C(graphite) + O_2, \qquad \Delta H° = 220\ kJ/mol$$

An exothermic reaction releases heat that will flow to the surroundings. An increase in temperatureis usually observed in the solvent, the container, or other close surroundings. In a perfect system, the heat flow out of the reaction is equal to the heat flow into the solution and the calorimeter. In this experiment, the calorimeter absorbs so little heat that the heat lost by the reaction will be considered equal to the heat gained by the solution in the calorimeter. The temperature change caused by the heat transfer will depend on the specific heat and the mass of the substance in the calorimeter. The change in temperature is calculated as final temperature – initial temperature.

$$Q = C_{sp}\ m\ \Delta T$$

The heat for a reaction is independent of path. Enthalpy, ΔH, is a state function. Such functions only depend on the initial and final state of the reactants and products. Making the products in a single step or a multi-step mechanism is not important as far as the enthalpy of the reaction is concerned. The same transfer of heat occurs.

In chemical equations, coefficients refer to the number of moles of each substance. The enthalpy of each substance depends on whether it is a solid (s), a liquid (l) or a gas (g). The symbol (aq) is used for a solution in water. The phase of the reactants and products must be specified so the correct enthalpy value can be obtained from a chart or table. The enthalpy of a substance also depends upon temperature. On a table of Heats of Formation, the temperature is usually specified to be 25°C. In the lab, temperature may not be 25°C and thermochemical calculations can be harder to accurately calculate.

KEY VOCABULARY

FIRST LAW OF THERMODYNAMICS: Based on the Law of Conservation of Energy, the 1st Law of Thermodynamics states that energy can only be transferred from one form of energy to another.

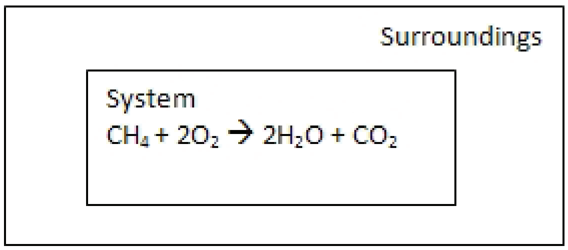
SECOND LAW OF THERMODYNAMICS: The universe is moving towards chaos, randomness and disorder.

ENDOTHERMIC REACTION: A reaction where more energy goes into the reaction (along with the reactants) than comes out of the reaction (along with the products), a $+\Delta H$. Heat is supplied to the system by the surroundings.

EXOTHERMIC REACTION: More energy comes out of the reaction (along with the products) compared to the amount of energy that goes into a reaction, a $-\Delta H$. Such processes give off heat which is then transferred to the surroundings.

THE UNIVERSE: In thermodynamics, the universe comprises two parts: the system and the surroundings. The system is the part of the universe that interests us. For a chemist, the system might be a chemical reaction. The rest of the universe outside the system is the surroundings.

The Universe



Heat moving from the system to the surroundings is an exothermic process, a $-\Delta H$.

CALORIMETER: A device used to measure heat changes such as the heat of a reaction. Heat in measured in units of joules or kilojoules.

THERMOMETER: A device used to measure the temperature of a substance. In chemistry, temperature is usually measured in units of Celcius.

ENTHALPY: This state function describes heat changes taking place at constant pressure. Enthalpy measures the heat of that reaction, ΔH_{rxn}.

STATE FUNCTION: A state function is independent of the path of a reaction. Energy, pressure, volume, and temperature are state functions – properties that are determined by the state of the system. State functions depend only on the initial and final states of the system and not on how the change occurred.

STANDARD CONDITIONS: In thermodynamics, standard conditions are consider to be one atmosphere of pressure, 25°C, and a 1 molar concentration for any solution.

TEACHER NOTES

How to make and use a simple calorimeter for this experiment

Rather than using Styrofoam cups to make a calorimeter, try using two small plastic salsa cups and a lid for the cups. You can get these at many fast food restaurants. Or try your school cafeteria!

1. Set one salsa cup inside the other, but separate them with a rubber band. This is done by slipping a rubber band around one of the cups. Then, insert the cup with the rubber band into the second cup. This creates an air pocket between the two cups. The air pocket is sealed off from the atmosphere by the rubber band. Air is a good insulator! The air pocket will slow down the transfer of heat between the salsa cup calorimeter and the environment.

2. Use a hole-punch to form a small opening in the lid. Place the lid on the inner salsa cup. A thermometer can be inserted through the opening to measure the temperature of the inner contents. Initially, this is the solution of HCl.

3. After adding the solid substance to the hydrochloric acid, the reaction begins and there will be a heat change. To get an accurate measurement of the heat change, be sure to mix or stir the reaction mixture in the calorimeter. Be careful not to spill or slosh any of the mixture out of the calorimeter.

4. Monitor the change in temperature on a thermometer. Temperature will change but the temperature should settle at a certain temperature before starting to lose more heat and returning to room temperature.

ADDITIONAL RESOURCES

Visit adamequipment.com/education regularly for new classroom resources.

ABOUT ADAM EQUIPMENT

Adam Equipment's world headquarters is located in Milton Keynes, United Kingdom, with facilities in the United States, Australia, South Africa and China. The company's balances have been trusted by professionals worldwide for 40 years. Contact Adam Equipment at education@adamequipment.com or online at www.adamequipment.com/education.

ADAM EQUIPMENT BALANCES RECOMMENDED FOR THIS EXPERIMENT

Highland Portable Precision Balance

Models recommended for this experiment:
HCB302 (300g capacity x 0.01g readability)
or HCB123 (120g capacity x 1mg readability)

Complete with more features and accessories than any other in its class, Adam Equipment's Highland Portable Precision Balances have what it takes for school and college applications. The reliable Highland provides the latest in weighing technology, 15 weighing units with four weighing modes and it is easy enough for novice students. It features Adam's unique patented ShockProtect™ overload protection to withstand up to 200kg, and HandiCal™ internal calibration with built-in mass. Calibrate whenever you want without external masses or use your own masses. USB and RS-232 interfaces are both included with cables. The rechargeable battery (adapter/charger included), removable draft shield and brilliant backlit display with capacity tracking make Highland the most complete portable precision balance available. Available in seven models from 150g x 0.001g to 3000g x 0.1g. For complete product details, visit www.adamequipment.com/education.

Core Portable Compact Balance

Model recommended for this experiment:
CQT202 (200g capacity x 0.01g readability)

Compact and durable, no other balance can beat the Core for basic weighing value. The tough, durable ABS housing is designed to withstand classroom environments, while being easy to clean and protected from accidental spills. With built-in ShockProtect™ overload protection, Core balances can handle excessive overloads without a problem. The simple keypad with dual tare keys and a brilliant backlit display make Core balances easy to use. Complete with a removable draft shield, AC adapter, and integral security slot, Core balances are ready to work right out of the box.

GETTING INVOLVED IN ADAM EQUIPMENT'S EXPERIMENTS

Feedback On This Adam Experiment

If you have feedback on this Adam Experiment that would be valuable to other teachers, we encourage you to share your thoughts. Please email your comments to Adam's education division at education@adamequipment.com.

Submitting Your Own Experiment

If you have an idea for a useful educational resource that you would like to share with other teachers, Adam Equipment is interested in hearing from you. Initial submissions need to include only a simple description of the activity with the activity's purpose, subject, and grade level. Please contact Adam's education division by e-mail at education@adamequipment.com to determine if your particular activity will fit into our experiment library. Adam Equipment will respond promptly to all inquiries.

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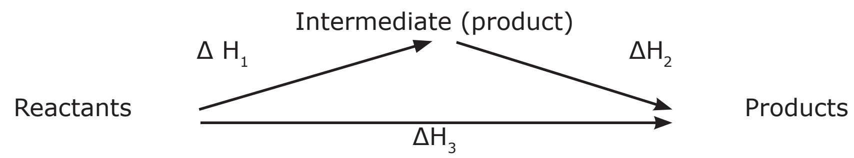
Student Section

KEEP YOUR EYES ON THE JOULES

A Hess's Law Experiment

HESS'S LAW

Hess's Law states that the change in enthalpy for any chemical reaction is constant. The pathway from reactants to products could involve one step or multiple steps, but the total enthalpy change from reactants to products would be the same, regardless of the pathway.



Thermochemical data for these reactions can be treated algebraically. This means that a reaction requiring two steps will involve the same amount of heat flow as the sum of the heat involved when the two steps occur independently. The reaction requiring two steps would also involve the same amount of heat if the reaction occurred in a single step.

$$\Delta H3 = \Delta H1 + \Delta H2$$

Let's look at a problem where Hess's Law might be used to calculate the enthalpy change. Magnesium ribbon burns in air (oxygen) to produce magnesium oxide according to the equation:

$$2\ Mg_{(s)} + O2_{(g)} \rightarrow 2\ MgO_{(s)}$$

The combustion of magnesium can be readily observed in the high school laboratory but the heat of combustion cannot be directly measured. By examining the enthalpy for a series of reactions, the heat of combustion for magnesium oxide can be determined indirectly by using Hess's Law.

MATERIALS

- Plastic cups with lids (holds 55-60 mL of liquid)
- Rubber band
- Hole-punch
- Thermometer
- Magnesium ribbon, Mg
- Magnesium oxide, MgO
- Balance

PROCEDURE

Part A

1. Put together a calorimeter as instructed by your teacher.

2. Mass 20.0 g of 1.0 M HCl into the calorimeter. Record the temperature of the acid.

   Initial temperature of the acid = ____________ °C

3. Using magnesium ribbon, record the mass of 1.00 meter of magnesium. Measure 4 cm of magnesium and cut into smaller pieces. Quickly add the pieces of magnesium to the acid and replace the lid on the calorimeter. Insert thermometer.

   Mass of 1.00 meters of magnesium ribbon = _______________ g

   Actual length of cut piece of Mg ribbon = _______________________ cm

   Calculations for mass of cut piece of Mg ribbon =

   If using magnesium turnings, mass 0.20 grams of magnesium turnings and add to the acid.

4. Gently stir the calorimeter and monitor the temperature. Record the highest temperature reached. (Note: this reaction gives off heat for quite awhile and should be allowed to react until the temperature begins to fall.)

   Final temperature of the acid = _____________ °C

Part B

1. Mass 20.0 g of 1.0 M HCl into the empty calorimeter. Record the temperature of the acid.

   Initial temperature of the acid = ____________ °C

2. Mass 0.35 g of magnesium oxide, MgO. Quickly, add the magnesium oxide powder to the acid in the calorimeter and replace the lid on the calorimeter. Insert thermometer.

   Actual mass of magnesium oxide = ______________ g

3. Gently swirl and record the highest temperature reached.

   Final temperature of acid = _______________ °C

DATA ANALYSIS

Part A

Mass of 1.0 M HCl solution used	g
Mass of magnesium ribbon used	g
Total mass of HCl and Mg	g
Initial temperature of 1.0 M HCl solution before mixing	°C
Final temperature of solution after Mg added	°C
Change in solution temperature	°C

Part B

Mass of 1.0 M HCl solution used	g
Mass of magnesium oxide used	g
Total mass of HCl and MgO	g g
Initial temperature of 1.0 M HCl solution before mixing	°C
Final temperature of solution after MgO added	°C
Change in solution temperature	°C

1. Write balanced equations for the reactions:

   Part A ____

   Part B ____

2. For the reactions in part A and B, calculate the heat of the reaction using:

$$\Delta H = q_{rxn} = (mc\Delta T + C'\Delta T)$$

   For c, use the specific heat for the mixture = 3.97J/g°C. If you do not know the calorimeter constant, C′, assume it to be zero. Show calculations!

   q (part A) = _____________J   Calculations:

   q (part B) = _____________J   Calculations:

3. Using values of J/g, convert to kJ/mol for the Q values. Show calculations! Pay careful attention to units when calculating from the previous answers.

   q (part A) = _____________kJ/mol   Calculations:

   q (part B) = _____________kJ/mol   Calculations:

DATA ANALYSIS

4. It is necessary to use a third reaction. The heat for formation of water from hydrogen (gas) and oxygen (gas) has a value of -285.5kJ/mol. Write the equation for the heat of formation of water.

5. Combine this reaction with the reactions from part A and B and cancel the same expressions that appear on the opposite sides of the equation. You should arrange the equations to end up with the equation:

$Mg + ½ O_2 \rightarrow MgO$ $\Delta H =$ ______

$Mg + ½ O_2 \rightarrow MgO$ $\Delta H =$ ______

6. Use Hess's Law to determine the heat of reaction for the net equation. This value is the heat of formation of magnesium oxide.

ΔH_f of MgO = ______________ kJ/mol

7. Compare the experimental results with the heat of formation in the thermodynamic tables (-603 kJ/mol) and calculate the percent error.

Percent error = ____________ %

8. Was the net outcome of the reaction in this experiment endothermic or exothermic? Use your data to support your answer.

9. What sources of error could account for your percent error?

10. Could the net reaction of a three-reaction system be endothermic if two of the reactions in the system are exothermic? Explain.

Sample Data and Answer Key

Part A. #3 Mass of 1.00 meters of magnesium ribbon = .973g
Calculations for mass of cut piece of Mg ribbon = 0.0404g

4.15cm	0.973g
	100cm

DATA ANALYSIS ANSWER KEY

Part A

Mass of 1.0 M HCl solution used	20.11g
Mass of magnesium ribbon used	0.0404 g
Total mass of HCl and Mg	20.15g
Initial temperature of 1.0 M HCl solution before mixing	23.7°C
Final temperature of solution after Mg added	32.6°C
Change in solution temperature	8.9°C

Part B

Mass of 1.0 M HCl solution used	19.95g
Mass of magnesium oxide used	0.348g
Total mass of HCl and MgO	20.298g (20.3)g
Initial temperature of 1.0 M HCl solution before mixing	23.7°C
Final temperature of solution after MgO added	31.7°C
Change in solution temperature	8.0°C

1. Write balanced equations for the reactions:

   Part A $Mg + 2\ HCl \rightarrow MgCl_2 + H_2$

   Part B $MgO + 2HCl \rightarrow MgCl_2 + H_2O$

2. For the reactions in part A and B, calculate the heat of the reaction using:

$$\Delta H = q_{rxn} = (mc\Delta T + C'\Delta T)$$

For c, use the specific heat for the mixture = 3.97J/g°C. If you do not know the calorimeter constant, C′, assume it to be zero. Show calculations!

Answers here are to hundredths place.

-q (part A) = - 711.96 J Calculations: (factor label)

20.15g	3.97J	8.9°C
	g °C	

-q (part B) = -644.64 J Calculations: (factor label)

20.298g	3.97J	8.0°C
	g °C	

DATA ANALYSIS ANSWER KEY

3. Using values of J/g, convert to kJ/mol for the Q values. Show calculations! Pay careful attention to units when calculating from the previous answers.

   -q (part A) = -427.99 kJ/mol Calculations: (factor label)

711.96 J	1 kJ	24.3 g Mg
.0404 g Mg	1000 J	1 mol Mg

   -q (part B) = -74.66 kJ/mol Calculations: (factor label)

-644.64 J	1 kJ	40.3g MgO
.348 g MgO	1000 J	1 mol Mg

4. It is necessary to use a third reaction. The heat for formation of water from hydrogen and oxygen has a value of -285.5 kJ/mol. Write the equation for the heat of formation of water.

   ANSWER: $H_2 + ½ O_2 \rightarrow H_2O$

5. Combine this reaction with the reactions from part A and B and cancel the same expressions that appear on the opposite sides of the equation. You should arrange the equations to end up with the equation:

   $Mg + ½ O_2 \rightarrow MgO$ ΔH = ? kJ
   $Mg + 2 HCl \rightarrow MgCl_2 + H_2$ ΔH = - 427.99 kJ
   MgCl2 + H2O → MgO + 2HCl ΔH = + 74.66 kJ
   H2 + ½ $O_2 \rightarrow H_2O$ ΔH = - 285.5 kJ

   Mg + ½ O2 → MgO **Δ H = -638.8**

6. Use Hess's Law to determine the heat of reaction for the net equation. This value is the heat of formation of magnesium oxide.

   ANSWER: ΔH_f of MgO = -638.8 kJ/mol

7. Compare the experimental results with the heat of formation in the thermodynamic tables (-603 kJ/mol) and calculate the percent error.

   ANSWER: Percent error = 6.1%

8. Was the net outcome of the reaction in this experiment endothermic or exothermic? Use your data to support your answer.

   ANSWER: Exothermic

9. What sources of error could account for your percent error?

10. Could the net reaction of a three-reaction system be endothermic if two of the reactions in the system are exothermic? Explain.