NAME PERIOD DATE

11. ENERGY IN CHEMICAL REACTIONS

Initial Question

The First Law of Thermodynamics states that energy is neither lost nor gained in a chemical process. This is paraphrased as "energy is conserved." Reactions that release energy are known as exothermic reactions. Reactions that absorb energy are endothermic reactions.

The amount of heat energy involved in a process is referred to as *enthalpy*. Although the amount of enthalpy cannot be measured directly, scientists can determine how much it changes. In this lab, you will use the First Law of Thermodynamics to determine the change of energy in various reactions and combine the results to determine the enthalpy change of a related reaction.

How do you find the change of enthalpy in chemical reactions?

Materials and Equipment

Calorimeter for Model 1, Model 2, Model 3, and Applying Your Knowledge

- Data collection system
- Stainless steel temperature sensor
- Polystyrene cup, 8 oz
- Ring stand
- Beaker, 250-mL
- Clamp, utility
- Graduated cylinder, 50-mL or 100-mL
- 10 cm × 10 cm cardboard lid

Model 1

- Ammonium nitrate (NH_4NO_3), solid[1]
- Distilled water[1]

[1]The mass and volume needed depend on the reaction assigned to your group.

Model 2

- 1.0 M Sodium hydroxide (NaOH), 100 mL
- 1.0 M Hydrochloric acid (HCl), 100 mL
- Distilled water, 100.0 mL

Model 3

- 1.0 M Sodium hydroxide (NaOH), 100 mL[1]
- 1.0 M Hydrochloric acid (HCl), 100 mL[1]
- Sodium hydroxide (NaOH), 4.0 g

[1]The solution and volume needed depend on the reaction assigned to your group.

Applying Your Knowledge

- 2.0 M Hydrochloric acid (HCl), 100 mL
- Magnesium ribbon (Mg), about 0.5 g

Safety

Add these important safety precautions to your normal laboratory procedures:

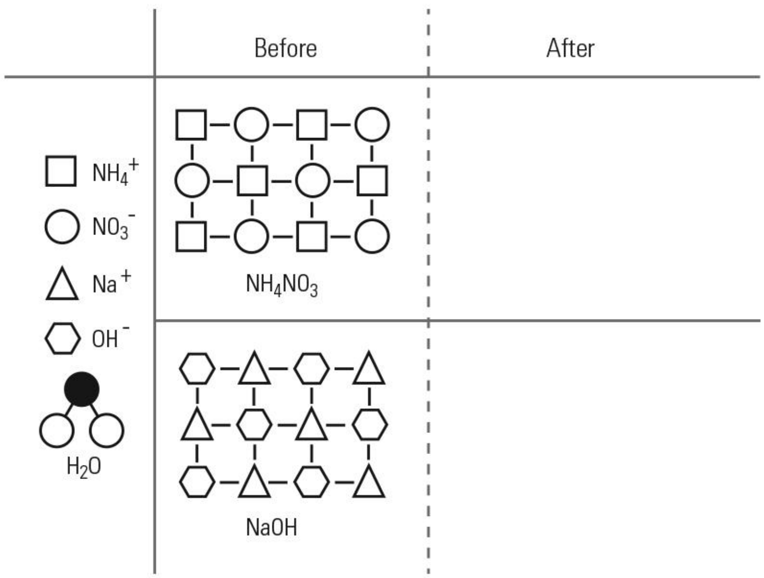
- Do not touch the solid NaOH and handle its resulting solution with care. NaOH is caustic and will cause skin burns and burn holes in clothing.
- When NaOH or HCl solutions come in contact with your skin or eyes, rinse immediately with a large amount of running water.

Getting Your Brain in Gear

1. A coffee cup calorimeter will be used to measure the heat of reaction (enthalpy of reaction) of several different reactions. The coffee cup minimizes, but does not eliminate, heat transfer with the surroundings. If the temperature of the room was 22.00 °C on the day of the experiment, what transfer of heat will take place between the solution and room, based on the temperature of the solution?

2. Samples of solid ammonium nitrate and sodium hydroxide are depicted below.

   a. Draw the changes that take place when the following compounds dissolve in water.



   b) Describe the changes to the intramolecular forces holding the components together when each salt is dissolved in water.

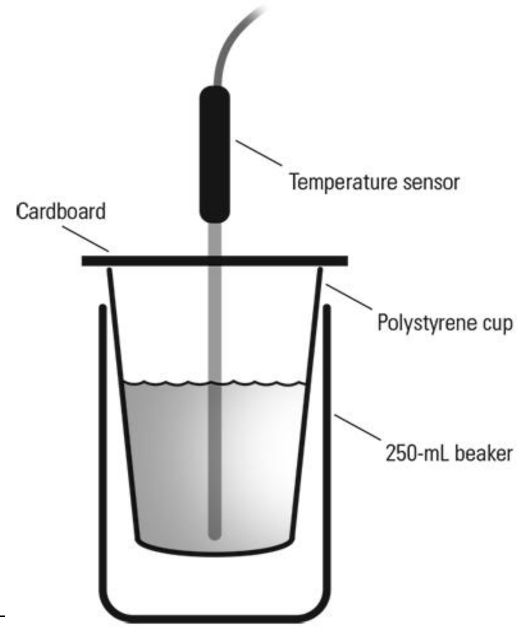
3. The reaction below is exothermic. What must be true about the total energy in the bonds of the reactants as compared to the total energy of the bonds in water? The formula for the *heat of reaction* is $\Delta H_{rxn} = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}$.

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l) \qquad \Delta H = -285.8\ \text{kJ}$$

4. In science, there is no such thing as negative energy but when we calculate the heat energy absorbed or released by a chemical reaction using the equation $q = mc\Delta T$, the temperature change is calculated as $T_{\text{final}} - T_{\text{initial}}$. When a chemical reaction causes a solution to cool, the change in temperature is negative ($T_{\text{final}} < T_{\text{initial}}$) and the resulting value of q is negative. q is assigned an artificial negative sign to force a relationship with ΔH. What is the relationship between the algebraic sign of q and the algebraic sign of ΔH?

MODEL 1

Building Model 1 – Dissolution of NH_4NO_3



1. Start a new experiment on the data collection system.

2. Connect a temperature sensor to the data collection system.

3. Create a graph display of temperature (°C) versus time (s).

4. Place the polystyrene cup in the 250-mL beaker.

5. Mount the temperature sensor on the ring stand and set it into the cup so it is about half an inch from the bottom.

6. How will placing the cup in a beaker help improve the accuracy of temperature measurements?

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7. Check with your teacher to determine which reaction you will carry out.

Table 1: Reactant volumes for Model 1 reactions

Reaction	Mass of NH_4NO_3 (g)	Volume of Water (mL)
A	2.00	50.0
B	2.00	100.0
C	4.00	50.0
D	4.00	100.0
E	6.00	50.0
F	6.00	100.0

8. Rinse the graduated cylinder with deionized water.

9. Measure the amount of water specified in Table 1 for your reaction. Record the volume in the Model 1 Data Table and then pour the water into the calorimeter.

10. Measure the solid reactant and record the mass to at least the nearest 0.01 g in the Model 1 Data Table.

11. Start recording data.

12. When the temperature readings stabilize, carefully transfer the solid reactant into the calorimeter. Swirl the calorimeter gently to stir.

13. When the temperature readings stabilize again, stop recording data.

14. Dispose of the solution properly, wash the cup and graduated cylinder, and rinse them with deionized water.

15. Repeat the procedure to verify your result.

16. Determine the temperature change for each data run and record it in the Model 1 Data Table.

17. Save your experiment and clean up according to your teacher's instructions. Then exchange data with your classmates and enter it into the Model 1 Data Table.

Model 1 – Dissolution of NH_4NO_3

Table 2: Model 1 Data Table—Determining the heat of reaction for the dissolution of NH_4NO_3

Reaction		Mass of NH_4NO_3 (g)	Volume of Water (mL)	ΔT (°C)	Heat q Gained or Lost, (kJ)	Number of Moles of NH_4NO_3 (mol)	ΔH_{rxn} (kJ/mol)	Average ΔH_{rxn} (kJ/mol)
A	Run 1							
	Run 2							
B	Run 1							
	Run 2							
C	Run 1							
	Run 2							
D	Run 1							
	Run 2							
E	Run 1							
	Run 2							
F	Run 1							
	Run 2							

Analyzing Model 1 – Dissolution of NH_4NO_3

18. Describe what happened to the temperature of the solution as you increased the mass of ammonium nitrate and used the same quantity of water (Reactions A, C, and E, or B, D, and F).

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19. For each reaction, use the specific heat capacity of water to determine the heat of reaction, as follows:

    a. Calculate the heat gained by or lost to the solution using the equation $q = mc\Delta T$ where m is the mass of water plus the mass of the solute and for simplicity assume the mass of 1.00 mL of water is 1.00 g, c is the specific heat capacity of water, 4.184 J/g °C, and $\Delta T_{water} = T_{final} - T_{initial}$. Record your answer, in kilojoules, in the Model 1 Data Table.

    b. Using the mass of ammonium nitrate, calculate the number of moles of ammonium nitrate present in each reaction.

    c. Using the number of moles of ammonium nitrate and the heat released or absorbed, calculate the heat of reaction [$\Delta H = -(q/\text{mole of product})$] for the dissolution of one mole of ammonium nitrate.

20. Will energy appear as a reactant or a product in the following equation? Write the word "energy" on the appropriate side. Explain your reasoning.

$$NH_4NO_3(s) + \text{water} \rightarrow NH_4^+(aq) + NO_3^-(aq)$$

21. Considering the ratios of solute to solvent in Model 1, do large ratios of solute-to-solvent or small ratios of solute-to-solvent cause the greatest temperature changes? Use experimental results to support your answer.

22. Compare the heat q and heat of reaction ΔH. What are the main differences between the two quantities?

23. Identify and explain any trends in the heat of reaction for Model 1.

24. Consider an additional reaction for Model 1 using 10 g of ammonium nitrate in 50.0 mL of water.

    a. How would this mass and volume affect the change in temperature compared to the other trials?

    b. Speculate on the quantity of heat q that would have been gained or lost by the solution in this reaction.

    c. What effect would this have had on your calculation of the heat of reaction for Model 1?

Model 2

Building Model 2 – Limiting Reactants

1. Set up the calorimeter as you did for Model 1.

2. Check with your teacher to determine which reaction you will carry out.

3. What is the relationship between the amount of HCl and the amount of NaOH for the five reactions in Table 3?

Table 3: Model 2 reactions using different ratios of reactant volumes

Ratio	Reactant 1 1.0 M HCl (aq) (mL)	Reactant 2 1.0 M NaOH (aq) (mL)
1	10.0	50.0
2	20.0	40.0
3	30.0	30.0
4	40.0	20.0
5	50.0	10.0

4. For your assigned reaction, measure the specified volume of Reactant 1. Record the value in the Model 2 Data Table and pour Reactant 1 into the calorimeter.

5. Rinse the graduated cylinder. Measure Reactant 2 and record the volume in the Model 2 Data Table.

6. Start recording data.

7. When the temperature readings stabilize, pour Reactant 2 into the calorimeter. Swirl the calorimeter gently to stir.

8. When the temperature readings stabilize again, stop recording data.

9. Dispose of the solution properly, wash the cup and graduated cylinder, and rinse them with deionized water.

10. Conduct the experiment again to obtain a second run.

11. Determine the initial and maximum temperature for each data run. Record the values in Model 2 Data Table.

12. Save your experiment and clean up according to your teacher's instructions. Then exchange data with your classmates and enter it into the Model 2 Data Table so you have the results for each reaction ratio.

Model 2– Limiting Reactants

Table 4a: Model 2 Data Table – Limiting Reactants: Measurements

Reaction		Volume (mL)		Temperature (°C)	
Ratio	Run	1.0 M HCl (aq)	1.0 M NaOH (aq)	Initial	Final
1	Run 1				
	Run 2				
2	Run 1				
	Run 2				
3	Run 1				
	Run 2				
4	Run 1				
	Run 2				
5	Run 1				
	Run 2				

Table 4b: Model 2 Data Table – Limiting Reactants: Calculation of the heat of reaction

Reaction		Temperature Change Δ*T* (°C)	Heat *q* (kJ)	Number of Moles		Heat of Reaction Δ*H* (kJ/mol)	Average Heat of Reaction Δ*H* (kJ/mol)
Ratio	Run			HCl	NaOH		
1	Run 1						
	Run 2						
2	Run 1						
	Run 2						
3	Run 1						
	Run 2						
4	Run 1						
	Run 2						
5	Run 1						
	Run 2						

Analyzing Model 2– Limiting Reactants

13. a. What is the equation for this reaction?

b. Calculate the heat q absorbed or released by the solution. Record your results in the Model 2 Data Table.

c. Circle the limiting reagent (the one with the least number of moles) in Table 3, copied below, for each reaction.

Ratio	Reactant 1 1.0 M HCl (aq) (mL)	Reactant 2 1.0 M NaOH (aq) (mL)
1	10.0	50.0
2	20.0	40.0
3	30.0	30.0
4	40.0	20.0
5	50.0	10.0

d. What will happen to the reaction when you run out of limiting reactant?

e. Which should you use, the number of moles of the limiting reactant or the number of moles of the excess reagent, when calculating the molar heat of reaction (kJ per mole of product)? Explain your answer.

14. Calculate the heat of reaction ΔH as you did in Model 1, using the following steps. Record all results in the Model 2 Data Table.

a. Calculate the number of moles of HCl and NaOH used in each reaction. How many moles of product are formed?

b. Calculate the heat of reaction ΔH.

15. Describe the changes in the heat q with respect to the ratios of reactants present.

16. Compare your results with the results of other students with respect to the trends and values of the heat q and the molar heat of reaction, ΔH.

17. Which ratio had the least amount of leftover reactants? Explain, citing data from the experiment.

18. How did the amount used of each reactant affect the heat of reaction ΔH? Why do you think this is?

MODEL 3

Building Model 3 – Additive Nature of Reactions

1. Set up the calorimeter as you did for the previous models.

2. Check with your teacher to determine which reaction you will carry out.

Table 5: Reactions and reactants for Model 3

Rxn #	Reactions	Reactant 1	Reactant 2
1	$NaOH(s) \rightarrow NaOH(aq)$	100.0 mL water	4.00 g NaOH
2	$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$	100.0 mL 1.0 M HCl(aq)	100.0 mL 1.0 M NaOH(aq)
3	$NaOH(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$	100.0 mL 1.0 M HCl(aq)	4.00 g NaOH

3. Rinse a graduated cylinder with deionized water and obtain Reactant 1 based on your assigned reaction.

4. Why is it important to rinse the graduated cylinder with deionized water between each use?

5. Carefully transfer the contents of the graduated cylinder into the calorimeter.

6. Obtain Reactant 2.

   NOTE: Since solid sodium hydroxide readily picks up moisture from the air, obtain its mass and proceed to the next step without delay.

7. Start recording data.

8. When the temperature readings stabilize, add Reactant 2 to the calorimeter.

9. Swirl gently to stir. When the temperature readings stabilize again, stop recording data.

10. Why is it important to wait for the readings to stabilize before adding Reactant 2?

11. Dispose of the solutions according to your teacher's instructions, wash the cup and graduated cylinder, and rinse them with deionized water.

12. Conduct the experiment again to obtain a second run.

13. Determine the initial and maximum temperature for each data run. Record the values in the appropriate Model 3 Data Table.

14. Clean up according to your teacher's instructions. Then exchange data with your classmates and enter it into the Model 3 Data Table so you have the results for each reaction.

Model 3 – Additive Nature of Reactions

Table 6a: Model 3 Data Table—Reaction 1

Parameters	Reaction 1 $NaOH(s) \rightleftharpoons Na^{+}(aq) + OH^{-}(aq)$	
	Run 1	Run 2
Initial temperature (°C)		
Final temperature (°C)		
Mass of solid NaOH (g)		
Volume of water (mL)		
Mass of solution (g)		
Change of temperature (°C)		
Heat *q* (kJ)		
Number of moles of NaOH (mol)		
Molar heat of reaction Δ*H* (kJ/mol)		
Average molar heat of reaction Δ*H* (kJ/mol)		

Table 6b: Model 3 Data Table—Reaction 2

Parameters	Reaction 2 $NaOH(aq) + HCl(aq) \rightleftharpoons NaCl(aq) + H_2O(l)$	
	Run 1	Run 2
Initial temperature (°C)		
Final temperature (°C)		
Volume of 1.0 M NaOH (mL)		
Volume of 1.0 M HCl (mL)		
Mass of solution (g)		
Change of temperature (°C)		
Heat *q* (kJ)		
Number of moles of NaOH (mol)		
Number of moles of HCl (mol)		
Molar heat of reaction Δ*H* (kJ/mol)		
Average molar heat of reaction Δ*H* (kJ/mol)		

Table 6c: Model 3 Data Table—Reaction 3

Parameters	**Reaction 3** $NaOH(s) + HCl(aq) \rightleftharpoons NaCl(aq) + H_2O(l)$	
	Run 1	**Run 2**
Initial temperature (°C)		
Final temperature (°C)		
Mass of solid NaOH (g)		
Volume of 1.0 M HCl (mL)		
Mass of solution (g)		
Change of temperature (°C)		
Heat q (kJ)		
Number of moles of NaOH (mol)		
Number of moles of HCl (mol)		
Molar heat of reaction ΔH (kJ/mol)		
Average molar heat of reaction ΔH (kJ/mol)		

Analyzing Model 3 – Additive Nature of Reactions

15. Calculate the change of temperature due to the reaction you carried out. Record your results in the corresponding Model 3 Data Table.

16. Calculate the heat q as you did in Model 1 and Model 2. Record your results in the corresponding Model 3 data table.

17. Calculate the number of moles of the reactants in each reaction. Record the results in the corresponding Model 3 Data Table. What is the limiting reagent for each reaction?

    NOTE: For Reaction 2, remember that the units of molarity are moles/liter and the answer should be in moles.

18. Compare your results to the results of other students. Note similarities and differences with respect to the heat q and the molar heat of reaction ΔH.

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19. Determine the number of moles of product that will be created. Then calculate the molar change of heat ΔH for each reaction (kJ per mole of product) and indicate if the reaction was endothermic or exothermic. Record your results in Model 2 data table.

20. Write the net ionic equation for each of the three reactions.

    Net ionic equation for Reaction 1: ______________________________

    Net ionic equation for Reaction 2: ______________________________

    Net ionic equation for Reaction 3: ______________________________

21. Demonstrate the relationship between the three ionic equations by combining them so the addition of the first two reactions equals the third. In the space below, algebraically add the ions from two of the reactions to equal the third reaction.

22. By citing the heats of reaction you obtained, mathematically verify that you added the correct two reactions together.

23. The previous two questions outline fundamental points of *Hess's Law*—the additive nature of heats of reactions. Describe Hess's Law.

24. How would the numerical value for ΔH_{rxn} change if the calorimeter was made of a conducting material and absorbs a significant amount of energy?

25. How would the numerical value of ΔH_{rxn} change if your lab partner put too much of the limiting reagent in the calorimeter?

Connecting to Theory

Thermochemistry studies are based on measuring the heat released or absorbed in a chemical process. The First Law of Thermodynamics states that energy is conserved in a process; therefore, any energy released or absorbed as heat can be measured by its direct effect on the environment. When a reaction is carried out in an aqueous solution, the energy given off or taken in by the process is transferred to or from the water although a small percentage may be lost to the calorimeter.

For example, if a reaction releases heat, an exothermic reaction, then the temperature of the water will increase. On the other hand, an endothermic reaction will absorb heat from the water, thus causing a decrease in the temperature of the water. This allows for a simple calculation of the heat of the reaction by first measuring the temperature change for the water, and then using the following equation to calculate the heat q absorbed or released by the dilute solution:

$$q = mc\Delta T$$

where m is the mass of the solution (assume the mass of 1.00 mL of solution is 1.00 g), c is the specific heat capacity of water: 4.184 J/g °C, and $\Delta T = T_{final} - T_{initial}$ of the solution. When the solutions used are dilute, they are assumed to have the same thermal properties as water.

In the Model 3 experiment, a polystyrene-cup calorimeter was used to measure the heat released by three different reactions. One of the reactions can be expressed as the combination of the other two reactions. Therefore, the heat of reaction of the one reaction should be equal to the sum of the heats of reaction of the other two. This concept is sometimes referred to as Hess's Law, or the additivity of heats of reaction.

Applying Your Knowledge – Applying Hess's Law

The following reaction cannot be performed directly:

$$Mg + \frac{1}{2}O_2 \rightleftharpoons MgO$$

To calculate the energy of formation for magnesium oxide, Hess's Law must be applied. A pathway that can be used is:

$Mg + 2HCl$	$\rightleftharpoons$	$MgCl_2 + H_2$	$\Delta H_1 = ?$
$MgCl_2 + H_2O$	$\rightleftharpoons$	$MgO + 2HCl$	$\Delta H_2 = 133$ kJ/mol
$H_2 + \frac{1}{2}O_2$	$\rightleftharpoons$	H_2O	$\Delta H_3 = -286$ kJ/mol
$Mg + \frac{1}{2}O_2$	$\rightleftharpoons$	MgO	$\Delta H_4 = ?$

Design an experiment using approximately 0.5 g of magnesium ribbon and 2.0 M HCl to determine ΔH_1. Be sure to record the actual amounts you used in the data table. Show your process of determining the limiting reactant. Then use Hess's law to determine the energy required to form one mole of MgO. Make sure the magnesium dissolves fully.

Table 7: Determination of ΔH_1

Parameters	Reaction 1 $Mg + 2HCl \rightleftharpoons MgCl_2 + H_2$	
	Run 1	Run2

The molar heat of reaction for the formation of MgO is ____________________ (ΔH_4).

Compare your results to known values and calculate the percent error. The following website may be helpful: http://bilbo.chm.uri.edu/CHM112/tables/thermtable.htm.