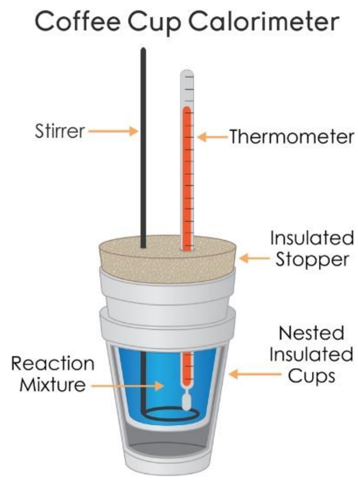
Lab 05.6
Enthalpy of Reaction

AP Chemistry

BACKGROUND

During any chemical reaction, heat can be either taken in from the environment or released out into it. The heat exchange between a chemical reaction and its environment is known as the enthalpy of reaction, or *H*. However, *H* can't be measured directly — instead, scientists use the *change* in the temperature of a reaction over time to find the *change* in enthalpy over time (denoted as Δ***H***). With Δ*H*, a scientist can determine whether a reaction gives off heat (or "is *exothermic*") or takes in heat (or "is *endothermic*"). In general, **Δ*H* = *m* x *c* x ΔT**, where m is the mass of the reactants, c is the specific heat of the product, and ΔT is the change in temperature from the reaction.



Enthalpies of formation are set Δ*H* values that represent the enthalpy changes from reactions used to create given chemicals. If you know the enthalpies of formation required to create products and reactants in an equation, you can add them up to estimate the enthalpy of an overall reaction.

For example, let's consider the reaction $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$. In this case, we know the enthalpies of formation for the following reactions:

$C_2H_5OH \rightarrow 2C + 3H_2 + 0.5O_2 = 228$ kJ/mol
$2C + 2O_2 \rightarrow 2CO_2 = -394 \times 2 = -788$ kJ/mol
$3H_2 + 1.5\ O_2 \rightarrow 3H_2O = -286 \times 3 = -858$ kJ/mol

Since we can add these equations up to get $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$, the reaction we're trying to find the enthalpy for, we can simply add up the enthalpies of the formation reactions above to find the enthalpy of this reaction as follows:

228 + -788 + -858 = **-1418 kJ/mol**.

The purpose of this experiment is to verify Hess's Law. Three acid-base reactions, chosen so that the third reaction equation equals the first reaction equation minus the second, are measured for temperature change by calorimetry. The values of heat change and enthalpy of reaction are calculated for each reaction. The measured value for the third reaction is then compared to the value calculated by subtracting the enthalpy of reaction for reaction two from the enthalpy of reaction of reaction one.

The solutions used in this lab ionize as follows:

$NH_4OH \rightarrow NH_3(aq) + H_2O(l)$
$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$
$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
$NH_4Cl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

Species	ΔH^o_f (kJ/mol)
$H^+(aq)$	0
$OH^-(aq)$	-230.0
$H_2O(l)$	-285.9
$NH_4^+(aq)$	-132.5
$NH_3(aq)$	-80.3

Safety

- The ammonia compounds used in Parts II and III will release vapors that are eye and respiratory irritants. Students should carry out reactions involving ammonia in the fume hood and should be careful not to inhale fumes.
- Both the hydrochloric acid and sodium hydroxide solutions are fairly concentrated and will be corrosive to skins and eyes.
- The bases in this lab will make glassware extremely slippery if spilled on the outside of the glassware.
- Spills should be neutralized and cleaned up immediately. Any solutions that contact skin should be rinsed off with plenty of water.
- Goggles and aprons must be worn.

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I. PURPOSE

Calculate the enthalpy of reactions experimentally and compare the results to another reaction to observe Hess' Law.

II. MATERIALS

1. styrofoam cups
2. 100 mL graduated cylinders
3. glass stirring rod
4. 2.0*M* HCl(aq)
5. 2.0*M* NaOH(aq)
6. 2.0*M* NH_4Cl (aq)
7. 2.0*M* NH_4OH (aq)
8. thermometer

III. PROCEDURES

<u>Part I – HCl and NaOH</u>

1. Obtain 50.0 mL of 2.0*M* HCl and pour it carefully into the styrofoam cup (calorimeter). Record the temperature of the HCl solution in the data table. *CAUTION: Handle the hydrochloric acid with care.*
2. Rinse the 100mL graduated cylinder repeatedly with tap water and rinse with distilled water.
3. Obtain 50.0 mL of 2.0*M* NaOH solution. Record the temperature of the NaOH solution in the data table. *CAUTION: Handle the sodium hydroxide solution with care.*
4. Add the 50.0 mL of NaOH solution to the styrofoam cup all at once. Stir the mixture throughout the reaction with the thermometer.
5. Record the temperature every 20 seconds. Stop recording the temperature change when 3 successive measurements are the same.
6. Rinse and dry the thermometer, styrofoam cup, and the stirring rod. Dispose of the solution as directed.

<u>Part II – NaOH and NH_4Cl</u> ***<u>Conduct this reaction in a fume hood or in a well-ventilated area.</u>***

1. Obtain 50.0 mL of 2.0*M* NaOH and pour it carefully into the styrofoam cup (calorimeter). Record the temperature of the NaOH solution in the data table.
2. Rinse the 100mL graduated cylinder repeatedly with tap water and rinse with distilled water.
3. Obtain 50.0 mL of 2.0*M* NH_4Cl solution. Record the temperature of the NH_4Cl solution in the data table.
4. Add the 50.0 mL of NH_4Cl solution to the styrofoam cup all at once. Stir the mixture throughout the reaction with the thermometer.
5. Record the temperature every 20 seconds. Stop recording the temperature change when 3 successive measurements are the same.
6. Rinse and dry the thermometer, styrofoam cup, and the stirring rod. Dispose of the solution as directed.

<u>Part III – HCl and NH_3</u> ***<u>Conduct this reaction in a fume hood or in a well-ventilated area.</u>***

1. Obtain 50.0 mL of 2.0*M* HCl and pour it carefully into the styrofoam cup (calorimeter). Record the temperature of the HCl solution in the data table.
2. Rinse the 100mL graduated cylinder repeatedly with tap water and rinse with distilled water.
3. Obtain 50.0 mL of 2.0*M* NH_3 solution. Record the temperature of the NH_3 solution in the data table.
4. Add the 50.0 mL of NH_3 solution to the styrofoam cup all at once. Stir the mixture throughout the reaction with the thermometer.
5. Record the temperature every 20 seconds. Stop recording the temperature change when 3 successive measurements are the same.
6. Rinse the thermometer, styrofoam cup, and the stirring rod. Dispose of the solution as directed.

IV. PRE-LAB QUESTIONS

1. Identify the acid and base in each reaction.
2. Write a molecular equation for each of the reactions being carried out. (Look up the product of Part III online.)
3. For each reaction, write a net ionic equation.
4. Use the table of Enthalpies of Formation (H^o_f) in the Background section to determine the ΔH^o for each reaction.

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V. DATA & CALCULATIONS

A. DATA

Time	Reaction 1 Temperature (^oC)	Reaction 2 Temperature (^oC)	Reaction 3 Temperature (^oC)
Reactant 1 – Time 0	HCl =	NaOH =	HCl =
Reactant 2 – Time 0	NaOH =	NH_4Cl =	NH_3 =
Mixture – Time (20s)			
Mixture – Time (40s)			
Mixture – Time (60s)			
Mixture – Time (80s)			
Mixture – Time (100s)			
Mixture – Time (120s)			
Mixture – Time (180s)			
Total Temperature change (ΔT)			

B. CALCULATIONS

1. Calculate the amount of heat energy, *q*, produced in each reaction. Use 1.03 g/mL for the density of all solutions. Use the specific heat of water, 4.18 J/(g•°C), for all solutions.

	$q_{solution\ calculation}$	$q_{reaction}$
ΔH_{rxn1}		
ΔH_{rxn2}		
ΔH_{rxn3}		

2. Calculate the enthalpy change, ΔH, for each reaction in kJ/mol.
3. Use the enthalpy changes from Calculation #2 to calculate a molar enthalpy for Reaction 3 using Hess' Law.
4. Compare the values of ΔH_{rxn3} by calculating a "percent error" for the calculated value (C#3) compared to the measured value (C#2).
5. Calculate two percent errors comparing the calculated ΔH_{rxn3} and measured ΔH_{rxn3} to the value determined in Pre-lab Question # 4.

VI. QUESTIONS & DISCUSSION OF ERROR

A. QUESTIONS

1. Is the system open, closed, or isolated? Explain.
2. Is the system constant pressure or constant volume? How does this affect the ΔH of the system?
3. Why is it necessary to know the density of the solutions to complete the enthalpy calculations?
4. How might the procedure be improved to achieve better results?

B. DISCUSSION OF ERROR

VII. CONCLUSION