# 7V-DETERMINING THE ENTHALPY OF A CHEMICAL REACTION

# (Using the Vernier Lab Quest)

All chemical reactions involve an exchange of heat energy. If the energy of the system increases at the expense of the surroundings, the reaction is **endothermic.** If energy flows from the system to the surroundings, the reaction is **exothermic.**  It is tempting to determine the thermodynamics of a reaction by measuring the enthalpy change (*∆H*). However, it is often not possible to directly measure the heat energy change of the reactants and products (the *system*). We can measure the heat change that occurs in the *surroundings* by monitoring temperature changes. We know that the Law of Conservation of Energy states that the sum of the heats gained by one component and lost by the other equals zero. If we conduct a reaction between two substances in aqueous solution where the solvent water serves as the surroundings for the reacting species, then the enthalpy of the reaction can be indirectly calculated with the following equation:

qsurr  = cp × m × ∆T = -qsys

The term **q** represents the heat energy that is gained or lost for the trial, **cp** is the specific heat of water, **m** is the mass of water, and **∆T** is the temperature change of the reaction solution. The specific heat and mass of water are used because it is the water in the solution that will either gain or lose heat energy during a reaction that occurs in aqueous solution. Note that q is an extensive value and depends on the amount of the participants. Combining q and the quantity of reacting species for a reaction at constant pressure yields **ΔH**, the **intensive enthalpy change** for the process, usually given in kJ/mol.

Furthermore, according to a principle known as **Hess’s Law**, the enthalpy changes of a series of reactions can be combined to calculate the enthalpy change of a reaction that is the sum of the components of the series. When one cannot measure the enthalpy change for a single step reaction, often one can create a series of reactions that sum to the target one. Summing the series gives the same overall enthalpy change as a direct single step reaction.

The purpose of this experiment is to use Hess’s Law to determine the enthalpy change of the reaction between aqueous ammonia and aqueous hydrochloric acid. You will then compare your theoretical enthalpy change values with the experimental results. To achieve this:

(1) First, we must account for the fact that every calorimeter absorbs some of the heat released in the reaction. The heat required to raise the temperature of the calorimeter by 1°C is the calorimeter constant, Ccal. We must determine Ccal for a particular calorimeter (a Thermos), then use the same calorimeter for the rest of the experiment.

(2) You will measure the temperature change of three chemical reactions, and then calculate each individual experimental enthalpy value, applying the correction for the calorimeter constant.

(3) You will use Hess’s Law to determine the **experimental** enthalpy change, **ΔH** of a third reaction, by using the experimental enthalpy values of the first two reactions. You will compare your experimental sum of the first two (partial) reactions with the measured value for the third (complete) reaction.

(4) You will use Hess’s Law to determine the **theoretical** enthalpy change, **ΔH** of a third reaction, by using the theoretical enthalpy values of the first two reactions as calculated in the pre-laboratory assignment. You will then compare the experimental and theoretical values for the third reaction and evaluate any sources of inaccuracy.

The following three reactions will be studied in this experiment. In the pre-lab report sheet, write balanced net ionic reaction equations for each from the descriptions. Use the table of thermodynamic data in your text (or another approved resource) to calculate the molar enthalpy of the reactions.

Reaction 1: An aqueous solution of sodium hydroxide reacts with an aqueous solution of hydrochloric acid, yielding water.

Reaction 2: An aqueous solution of sodium hydroxide reacts with an aqueous solution of ammonium chloride, yielding aqueous ammonia, NH3, and water.

Reaction 3: An aqueous solution of hydrochloric acid reacts with aqueous ammonia, NH3, yielding aqueous ammonium chloride.

| Required materials: |  |
| --- | --- |
| LabQuest Temperature Probe | 2.0 M hydrochloric acid, HCl, solution  2.0 M sodium hydroxide, NaOH, solution |
| Lab Quest App | 2.0 M ammonium chloride, NH4Cl, solution |
| glass stirring rod | 2.0 M ammonium hydroxide, NH4OH, solution |
| Soup thermos (calorimeter) | ring stand |
| 250 mL or 400 mL beaker | utility clamp |
| 50 mL or 100 mL graduated cylinders | fume hood |

## PROCEDURE:

1. Obtain and wear goggles. It is best to conduct this experiment in a fume hood, or in a   
well-ventilated room.

2. Connect the Temperature Probe to LabQuest and choose New from the File menu. If you have an older sensor that does not auto-ID, manually set up the sensor.

3. Use a utility clamp to suspend the Temperature Probe from a ring stand. Adjust the ring stand so that the probe can be easily lowered into a reaction mixture then raised after the measurement has been taken.

### **Part 1: Determination of the Calorimeter Constant.**

Check out a calorimeter and record its code number.

Measure out 50.0 mL of cool water, pour it into the calorimeter and allow it to remain for three minutes to reach thermal equilibrium with the calorimeter. Use the LabQuest temperature probe to read and record the temperature of the cool water (Tc) to the nearest 0.1°C.

Heat another sample of water to approximately 60°C then remove the flame or hot plate. Pour 50.0 mL of hot water into a graduated cylinder, read and record its temperature (Th) then IMMEDIATELY pour the hot water into the calorimeter. Stopper the calorimeter to prevent heat loss, swirl to mix for approximately one minute then read the temperature of the mixture of hot and cold waters (Thc). Record the value for this trial. Repeat the procedure for at least 3 trials.

**Part I: Calculation of the calorimeter constant:**

The general formula for heat absorbed or released by a substance is:

q = cs m ΔT

where q is the heat, cs is the specific heat of the substance in J ⋅ g–1 ⋅ °C–1, m is the mass of the substance and ΔT is the temperature change, ΔT = Tfinal – Tinitial. Since this experiment is performed under constant pressure conditions, q can also be related to ΔH.

The heat lost by the hot water equals the heat gained by the cold water plus the heat gained by the calorimeter (Law of Conservation of Energy):

0 = q hot water + q cold water + q calorimeter

or, 0 = (c m ΔT)hot water + (c m ΔT)cold water + (Ccal)(ΔTcal)

(ΔTcal) is the same as (ΔTcold water) since they were originally in contact. Since the only unknown is Ccal, we can calculate the heat capacity of the calorimeter. No mass is needed to find the heat absorbed by the calorimeter since its mass is constant throughout the experiment. (We are finding its heat capacity, C, not its specific heat, c.)

### **Part II: Conduct Reaction 1 between solutions of NaOH and HCl**

Measure out 50.0 mL of 2.0 M HCl solution into the thermos. Lower the Temperature Probe into the solution. **CAUTION:** *Handle the hydrochloric acid with care. It can cause painful burns if it comes in contact with the skin*. Start data collection on LabQuest and obtain the initial temperature of the HCl solution.

Measure out 50.0 mL of NaOH solution, record its temperature using a thermometer but do not add it to the HCl solution yet. It should be at approximately the same room temperature value as the HCl solution in the Thermos. **CAUTION:** *Handle the sodium hydroxide solution with care. Be especially careful not to get any in your eyes. Notify the instructor immediately if you experience any discomfort or feel a slippery, soapy sensation on your skin.*

After three or four readings have been recorded at the same temperature for the HCl solution, add the 50.0 mL of NaOH solution to the thermos all at once. Lightly cap and swirl the thermos continuously until the temperature reaches a peak and begins to fall. Collect temperature reading for about 3 minutes. (If the temperature readings are no longer changing, you may stop the trial early.) If you collected the readings using the LabQuest, you can examine the data pairs on the displayed graph. To extract data, click on any data point then scan the graph. As you move the “examine” line, the temperature values of each data point are displayed to the right of the graph. Record the initial temperature of the HCl solution and the maximum temperature of the mixed solutions in your data table.

Rinse and dry the Temperature Probe, and thermos. Dispose of the contents (the reaction products) as directed.

**Part III: Conduct Reaction 2 between Solutions of NaOH and NH4Cl**

Measure out 50.0 mL of 2.0 M NaOH solution into the thermos. Lower the tip of the Temperature Probe into the cup of NaOH solution. Start data collection on LabQuest and obtain the initial temperature of the NaOH solution.

Measure out 50.0 mL of 2.0 M NH4Cl solution, but do not add it to the NaOH solution yet.

Determine the initial temperature of the NaOH. Use a thermometer to check the temperature of the NH4Cl solution. After three or four readings have been recorded at the same temperature for the NaOH solution, add the 50.0 mL of NH4Cl solution to the thermos all at once. Swirl the reaction mixture continuously until the temperature reaches a peak and begins to fall. Data collection will end after three minutes or when the new temperature becomes constant. Examine the graph as before to determine and record the initial and maximum temperatures of the reaction.

Rinse and dry the Temperature Probe and the thermos. Dispose of the contents as directed.

Part IV: Conduct Reaction 3 Between Solutions of HCl and NH3

NOTE: It is best to conduct this reaction in a fume hood or in a well-ventilated area.

Measure out 50.0 mL of 2.0 M HCl solution into thermos and lower the tip of the Temperature Probe into the cup of HCl solution. Start data collection with the LabQuest as before, this time collecting the initial temperature of the HCl solution.

Measure out 50.0 mL of 2.0 M NH3 solution, measure its temperature with a thermometer, but do not add it to the HCl solution yet.

After three or four readings have been recorded at the same temperature for the HCl solution, add the 50.0 mL of NH3 solution to the thermos all at once. Swirl continuously until the temperature reaches a peak and begins to fall. Data collection will end after three minutes or when the new temperature becomes constant. Examine the graph as before to determine and record the initial and maximum temperatures of the reaction.

Rinse and dry the Temperature Probe and the thermos. Dispose of the contents as directed.

Detach the Temperature Probe, clear the program from the Lab Quest and turn it off. Return the Lab Quest and Temperature Probe.

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**Report Sheet**

**PRE-LAB ASSIGNMENT - THEORETICAL DETERMINATION**: Before attending lab, write the balanced net ionic reaction equations and then calculate the enthalpy change for each reaction by summing enthalpy of formation values (ΔHf), as found in your textbook.

| Reaction | Balanced reaction equation | Δ*H* (kJ/mol) |
| --- | --- | --- |
| 1 |  |  |
| 2 |  |  |
| 3 |  |  |

**EXPERIMENTAL DETERMINATION**: Record your measurements in the table below.

|  | CCal trial 1 | CCal trial 2 | CCal trial 3 | Reaction 1 | Reaction 2 | Reaction 3 |
| --- | --- | --- | --- | --- | --- | --- |
| Total volume (mL) |  |  |  |  |  |  |
| Total mass (g) |  |  |  |  |  |  |
| Maximum temperature (°C) |  |  |  |  |  |  |
| Initial temperature (°C) |  |  |  |  |  |  |
| Temperature change (∆*T*) |  |  |  |  |  |  |

**DATA ANALYSIS:** Show only a sample calculation for one of the reactions here. Then, for each in either Word or Excel, create a table with all of your final result values. Attach a copy of that table to this report sheet.

**SAMPLE CALCULATIONS**

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

qproducts = -[(c m ΔT)soln + (Ccal)(ΔTcal)]

2. Now calculate the enthalpy change, ∆*H*, for each reaction in terms of kJ/mol of product. Watch your SIGN so the result correctly shows whether the reaction is endothermic or exothermic! Show your stoichiometry calculation for the amount of each reactant consumed and each product formed.

3. Use your answers from above for reactions 1 and 2 and Hess’s Law to determine the experimental molar enthalpy for Reaction 3.

4. Next use Hess’s Law, to sum the theoretical values of Δ*Hrxn* for reactions 1 and 2 and so to calculate the Δ*Hrxn* for Reaction 3. How does the sum of these theoretical values compare to your experimental value? Discuss.

5. Does this experimental process support Hess’s law? Suggest ways of improving your results.

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**Post-Laboratory Questions**

1. When pure sulfuric acid dissolves in water, a great deal of heat is liberated. To determine the enthalpy of solution, 175 g of water was placed in an open calorimeter and cooled to 20.0**°**C. Then 5.00 g of pure, concentrated sulfuric acid, H2SO4(l), also at 20.0**°**C, was added and mixed in. The temperature rose rapidly to 25.4**°**C. Assume that the specific heat of the solution is

4.184 J ⋅ g–1 ⋅ C–1. Ccal = 19 J/**°**C. Calculate q. Then, calculate the ΔH of solution for one mole of H2SO4.

2. During a determination of ΔHrxn for the reaction: **Mg2+(aq) + H2O(l) 🡪 MgO(s) + 2 H+(aq**), using a method similar to this experiment, solid MgO was added to aqueous HCl solution. This reaction was exothermic. Suppose not all the MgO(s) dissolved in the HCl solution before the final temperature reading was obtained. What effect would this have on your calculated ΔHrxn value for the target reaction?

3. The standard enthalpy of formation for MgO(s) is -601.7 kJ/mol. Calculate the theoretical values for Equations 1, 2 and 3 and use to determine the value of the Overall Equation.

Equation 1 Mg(s) + 2 H+(aq) 🡪 Mg2+(aq) + H2(g) ΔH1 = \_\_\_\_\_\_\_\_\_\_

Equation 2 Mg2+(aq) + H2O(l) 🡪 MgO(s) + 2 H+(aq) ΔH2 = \_\_\_\_\_\_\_\_\_\_

Equation 3 H2(g) + ½O2(g) 🡪 H2O(l) ΔH3 = \_\_\_\_\_\_\_\_\_\_

Overall Equation Mg(s) + ½ O2(g) 🡪 MgO(s)

Calculate (and compare) using Hess’s Law ΔHf for MgO(s) = ΔH1+ΔH2+ΔH3 = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

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## Pre-laboratory Assignment

1. Fill out the theoretical table in the report sheet before class.

2. How much heat is absorbed by a 550. g sample of water as it warms from 22.5**°**C to 47.1**°**C?

3. Calculate ΔH**°** for the reaction: GeO(s) + ½ O2(g) 🡪 GeO2(s) by combining the enthalpies of these two known equations, using Hess’s Law.

Ge(s) + ½O2(g) 🡪 GeO(s) ΔH**°** = -255 kJ

GeO2(s) 🡪 Ge(s) + O2(g) ΔH**°** = +535 kJ

4. During Reaction 1, the reaction of NaOH(aq) and HCl(aq), the thermometer registers an increase in temperature for the mixed solutions in the calorimeter. What change occurs to the chemical potential energy of the reactants, the NaOH(aq) and the HCl(aq), in the experiment?