Thermochemistry

Introduction

All chemical reactions and phase changes involve energy. One form of energy is heat: when a change in the energy of a system results in a temperature difference, we say that energy has been transferred as heat. An equation that expresses heat in the internationally accepted energy unit of Joules (J) is

Heat Released or Absorbed = mass of sample x specific heat x change in temperature

$$q = ms\Delta T$$

$$\mathbf{J = g \times [J/(g \times {}^\circ C)] \times {}^\circ C}$$

In this lab, you'll perform chemical reactions in a calorimeter, which is a device scientists use to determine the amount of heat produced (or consumed) by a process. Lest this seem like an impossible task that must be left entirely to KU's thermochemistry wizards, we'll begin by taking a "sneak peek" at the calculations. First, you'll "plug in" numerical values for the variables on the right hand side of the above equation to calculate the quantity of heat ***q*** released or absorbed, in J, for each process you study. You'll then express the results in terms of a quantity called heat of reaction, ΔH (also known as enthalpy of reaction). However, considering that the atmospheric pressure in lab will not change significantly during the course of your experiments, you can assume that $q=\Delta H$, as explained in section 6.7 of your Tro Chemistry: A Molecular Approach textbook. (For the reactions in Part 2, you'll convert your ΔH values to units of **kilojoules per mole**, kJ/mol).

Processes that absorb energy are referred to as endothermic, whereas processes that release energy are exothermic. This is a particularly important convention to consider when you report your experimental data: endothermic processes correspond to *positive* values of ΔH, whereas exothermic processes correspond to *negative* values of ΔH.

In the above preview of this week's calculations, you may be left wondering what ***s*** means -- the specific heat. In general, the **specific heat *s*** of a substance refers to the amount of heat gained or lost when one gram of that substance changes temperature by one Celsius degree. For example, the specific heat of water is 4.184 J/g C°**.** Thus, 4.184 Joules of energy is required to raise the temperature of 1 gram of water by 1 Celsius degree. Water is the primary component of each of the reaction mixtures to be studied in this lab, so we can substitute 4.184 J/g C° for ***s*** in all of our calculations.

Pre-lab

Safety

Caution! Precautions should be taken with when working with acids and bases. These substances are corrosive when they come in contact with skin and may cause damage if splashed into your eyes. **Goggles must be worn at all times as required by state law.** If any contact with skin or eyes occurs, the area should be thoroughly flushed with water for 10 minutes or more. There is an eye wash in every lab. Any clothing that becomes contaminated should be removed from contact with the body immediately and not worn again until it is thoroughly washed.

Caution! Sodium hydroxide (NaOH) *denatures* proteins. (It causes them to lose their natural conformation by altering the interactions among their constituent amino acids.) **YOU** are made up of proteins. If you have been working with strong bases and the skin on your hands feels soapy, you should wash your hands thoroughly with water to remove this caustic material. Take particular care when dispensing or working with solid sodium hydroxide. Be sure to pick up and properly dispose of any pellets that may be dropped. Check with your TA for proper handling. (Keep NaOH containers capped!)

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1. List the chemicals you will use for the hands-on portion of this week's lab experiment. For each chemical, list specific safety precaution(s) that must be followed.

For problem #2 and #3, fill in the blanks to complete the chemical equations.

2. $HC_2H_3O_2$ + NaOH → ______ + ______

3. HCl + NaOH → ______ + ______

4. Given the following data,

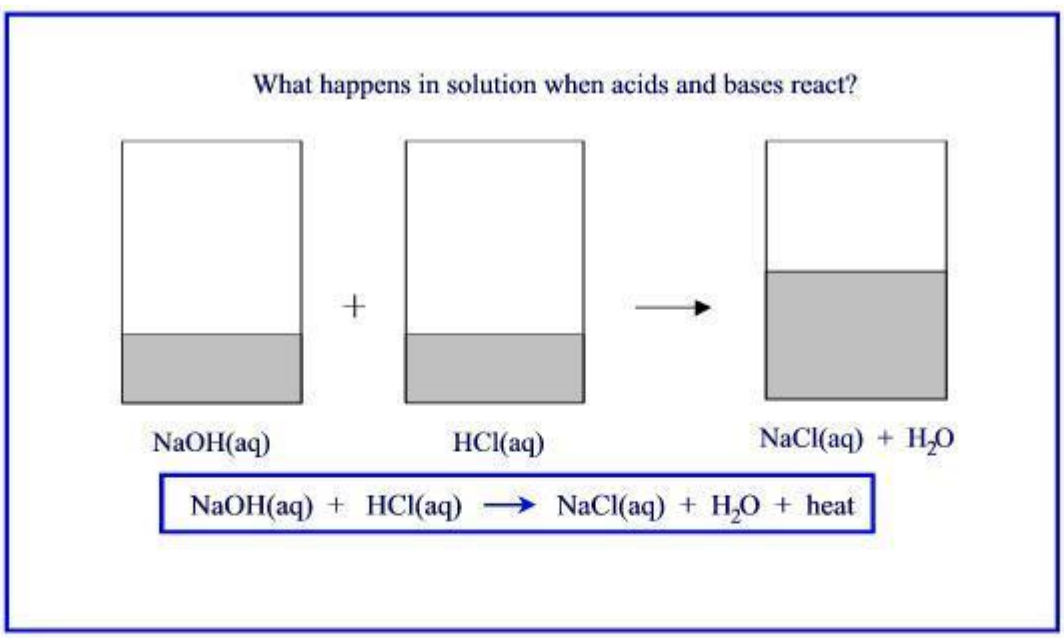
$NO_2(g) \rightarrow \frac{1}{2}N_2(g) + O_2(g) \quad \Delta H_1 = -33.9$ kJ

$NO_2(g) \rightarrow NO(g) + \frac{1}{2}O_2(g) \quad \Delta H_2 = 56.5$ kJ

What is ΔH for the reaction $\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightarrow NO(g)$?

(Hint for question #4: concepts explained in section 6.8 and 6.9 of Chang and Tro's Chemistry: A Molecular Approach textbook are helpful in solving this problem.)

5. Consider the reaction below. What species are present in each of the beakers? Draw a representation of what is going on at the molecular level in the each of the beakers.



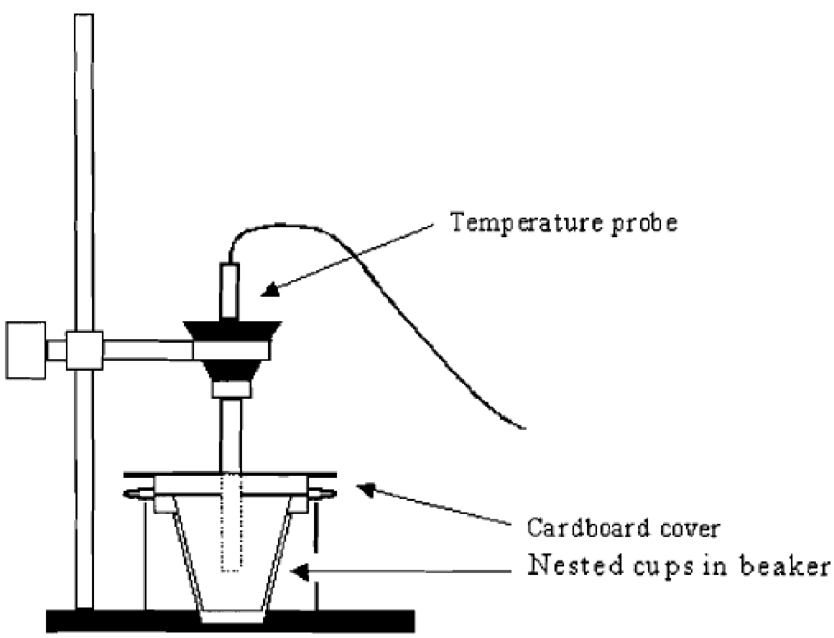
Procedure

Part 1 – Introduction to Calorimetry

Reagents: 1.0 M NaOH, 1.0 M HCl, and distilled water

In Part 1 you will measure the temperature change that occurs when NaOH and HCl are combined in a calorimeter. Rather than asking students to attempt to combine pure solid NaOH and pure HCl gas, Gary the "stockroom guy" has prepared aqueous (i.e., water-based) solutions of each of these chemicals for your use. Each solution contains 1.0 mole of the chemical (either NaOH or HCl) per liter of solution. In scientific vocabulary, a solution containing 1.0 mole of solute per liter is called 1.0 M, or "one molar." (If one substance is dissolved in a second substance to make a solution, the two components of the solution are referred to as the solvent and the solute. For all of our solutions, water is considered the solvent. The substance dissolved in the water--whether HCl, NaOH, or something else--is therefore the solute.)

The laboratory device we will use to study the heat produced during the mixing of these solutions is the **calorimeter** shown below. This device measures the transfer of heat into a known mass of water by monitoring the change in temperature of the solution. Today, you will make a simple device from two nested Styrofoam cups in a beaker and a cardboard cover. Styrofoam is a very good insulator, so only a small amount of heat is "lost" in warming the calorimeter itself. For your investigation, you can consider this small amount of heat negligible.



Each team will be assigned one of the following combinations of NaOH and HCl to examine.
Note: Auto-pipetters will be set up for your use. The auto-pipetter for the HCl will be set for 5 mL; take multiples of 5 mL. **Do NOT reset the auto-pipetters.** If you have a problem, see your TA.

Reagent	Team 1	Team 2	Team 3	Team 4	Team 5
NaOH 1.0 M	25 mL	25 mL	25 mL	25 mL	25 mL
HCl 1.0 M	15 mL	20 mL	25 mL	30 mL	35 mL
Distilled Water	20 mL	15 mL	10 mL	5 mL	0 mL

Collect data for the neutralization reaction that occurs between the volumes of acid and base assigned to your group as follows:

1. Construct your calorimeter as shown in the above diagram. Prepare your computer for data collection with the temperature probe by opening the experiment entitled “18 Hess's Law” from the *Chemistry with Vernier* experiment folder of *LoggerPro*.

2. Place all required reagents **except for the base, NaOH,** into the calorimeter. *Do you know the mass of the reagent(s) you are adding?*

3. When you are ready to begin collecting temperature data, click the Collect button, which looks like the "play" icon on a DVD player. You will see the temperature data being collected; this *initial* temperature--established by a series of consecutive identical readings--is designated algebraically as T_i.

4. Now, add the NaOH to the calorimeter. *Do you know the mass of the NaOH solution you are adding?* To homogenize the solution, swirl gently and continuously until the maximum temperature has been reached. The highest temperature reached will be designated as the *final* temperature T_f.

5. When the temperature begins to drop, click the Stop button, which looks like the "stop" icon on a DVD player. You may click on the Statistics (STAT) button on the tool bar, and the minimum and maximum temperatures will be listed in the statistics box on the graph. Be sure to record the initial and final temperatures in your lab notebook.

6. Compute your results for the heat of reaction ΔH in units of J; write these results on the chalkboard. Also, be sure to record the total mass of your solution and determine and record the number of moles of NaOH and HCl involved.

7. Finally, pour your solution into a beaker, add a drop of indicator to the beaker, agitate gently, and record your observation. This particular indicator is a pH indicator that remains colorless in acidic solutions, but becomes brightly colored in basic solutions. Keep this solution until directed to dispose of it by your TA. Discuss the following questions with your group, and with your classmates, as directed by your TA.

What do your measurements and observations after adding the indicator tell you about the reaction between your samples of NaOH(aq) and HCl(aq)?

Write a chemical equation for the reaction.

How do the results of the various groups compare? How do you explain any differences?

Can one particular group's combination of volumes be considered the "correct" stoichiometric ratio of HCl to NaOH? Why?

During your class discussion, your TA will select the solution that the class decides has the "correct" stoichiometric ratio. To a **portion** of this solution, several drops of acid or base, as appropriate, may be added drop wise, with gentle agitation. *What do you observe, and why?*

Part 2 – Studying ΔH Values with Hess's Law

Reagents: 1.0 M NaOH, 1.0 M HCl, 1.0 M $HC_2H_3O_2$, solid NaOH, and distilled water
The experiment goals for Part 2 A-D are described in brief below. Details, such as quantities of material to combine, are not provided. Use your understanding of the methods employed in Part 1, above, to develop strategies for Part 2.

Which team in Laboratory Part 1 used amounts of acid and base that resulted in the complete neutralization of the acid? How many moles of acid and base were used in this reaction?

Part 2A

Each group should repeat the experiment with the volumes of 1.0 M NaOH and 1.0 M HCl that gave the "correct" stoichiometric ratio in Part 1. Determine the heat of reaction, ΔH, using the approach employed in Part 1.

Part 2B

*Would there be any observable difference if the acid were mixed with **solid** NaOH rather than the aqueous NaOH you have been using?* Devise a method to investigate this question and perform the investigation. Calculate the heat of reaction.

Part 2C

What would happen if you performed a variation of **Part 2B** *by dissolving the same amount of solid NaOH in distilled water, with no HCl present? Is there a detectable heat of reaction associated with this process?* Try it and find out!

Now write chemical equations for the processes (Parts **2A, 2B,** and **2C**) you have investigated thus far. Look carefully at your ΔH values to compare the heat produced upon mixing NaOH(aq) with HCl(aq), NaOH(s) with H_2O, and NaOH(s) with HCl(aq). **Is neutralization of the acid the only process producing heat in these reactions? Is there any evidence that more than one process is at work? Can you apply the general method used in prelab question #4 to support the idea that the heat produced by chemical processes is additive?**

Part 2D

Would there be an observable difference if you repeated the reaction in **Part 2A** *above using a different acid, such as acetic acid (CH_3COOH)?* Perform the investigation and calculate the heat produced by this reaction. **How does the heat involved in the reaction of acetic acid with sodium hydroxide compare with the analogous reaction of hydrochloric acid and sodium hydroxide? Could you write a general equation that describes the neutralization reaction of any acid with any strong base like NaOH?**