1. Calorimetry

Objective

The calorimeter constant for a simple coffee-cup calorimeter will be determined, and then the calorimeter will be used to measure the quantity of heat that flows in several physical and chemical processes.

Choice I. Determination of a Calorimeter Constant

Introduction

Chemical and physical changes are always accompanied by a change in *energy.* Most commonly, this energy change is observed as a flow of heat energy either into or out of the system under study. Heat flows are measured in an instrument called a **calorimeter**. There are specific types of calorimeters for specific reactions, but all calorimeters contain the same basic components. They are *insulated* to prevent loss or gain of heat energy between the calorimeter and its surroundings. For example, the simple calorimeter you will use in this experiment is made of a heat-insulating plastic foam material. Calorimeters contain a *heat sink* that can absorb or provide the energy for the process under study. The most common material used as a heat sink for calorimeters is water, because of its simple availability and large heat capacity. Calorimeters also must contain some device for the measurement of *temperature,* because it is from the temperature change of the calorimeter and its contents that the magnitude of the heat flow is calculated. Your simple calorimeter will use an ordinary thermometer for this purpose.

To determine the heat flow for a process, the calorimeter typically is filled with a weighed amount of water. The process that releases or absorbs heat is then performed within the calorimeter, and the temperature of the water in the calorimeter is monitored. From the *mass of water* in the calorimeter, and from the *temperature change* of the water, the quantity of heat transferred by the process can be determined.

The calorimeter for this experiment is pictured in Figure 1-1. The calorimeter consists of two nested plastic foam coffee cups and a cover, with thermometer and stirring wire inserted through holes punched in the cover. As you know, plastic foam does not conduct heat well and will not allow heat generated by a chemical reaction in the cup to be lost to the room. (Coffee will not cool off as quickly in such a cup compared to a china or paper cup.) Other sorts of calorimeters might be available in your laboratory; your instructor will demonstrate such calorimeters if necessary. The simple coffee-cup calorimeter generally gives quite acceptable results, however.

Although the plastic foam material from which your calorimeter is constructed does not conduct heat well, it does still absorb some heat. In addition, a small quantity of heat may be transferred to or from the metal wire used for stirring the calorimeter's contents or to the glass of the thermometer used to measure

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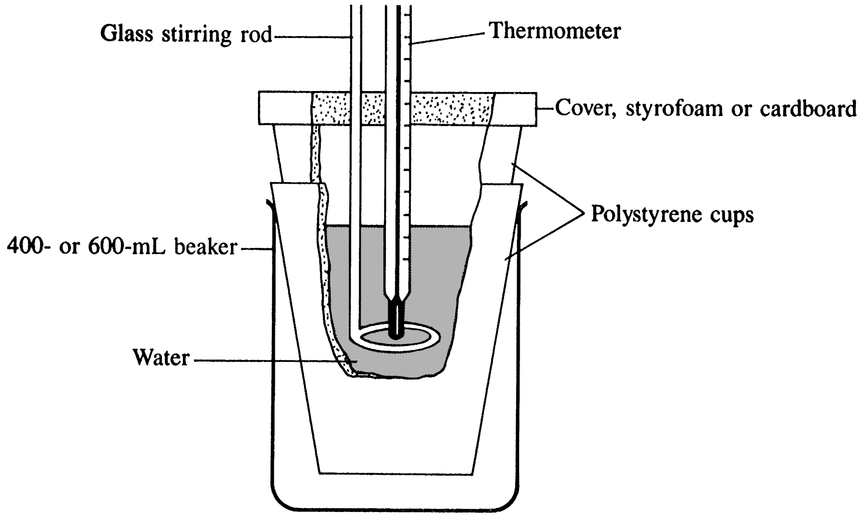
FIGURE 1-1
A simple calorimeter made from two nested plastic foam coffee cups. Make certain the stirring wire can be agitated easily.



temperature changes. Some heat energy may also be lost through the openings for these devices. Therefore, the calorimeter will be *calibrated* using a known system before it is used in the determination of the heat flows in unknown systems.

As mentioned earlier, there are several mechanisms by which a calorimeter can absorb or transmit heat energy. Rather than determining the influence of each of these separately, a function called the **calorimeter constant** can be determined for a given calorimeter. The calorimeter constant represents what portion of the heat flow from a chemical or physical process conducted in the calorimeter goes to the apparatus itself, rather than to affecting the temperature of the heat sink (water). Once the calorimeter constant has been determined for a given apparatus, the value determined can be applied whenever that calorimeter is employed in subsequent experiments.

As discussed, the temperature changes undergone by the heat sink are used to calculate the quantity of heat energy that flows during a chemical or physical process conducted in the calorimeter. When a sample of *any* substance changes in temperature, the quantity of heat, Q, involved in the temperature change is given by

$$Q = mC\Delta T \qquad (1\text{-}1)$$

where m is the mass of the substance, ΔT is the temperature change, and C is a quantity called the **specific heat** of the substance. The specific heat represents the quantity of heat required to raise the temperature of one gram of the substance by one degree Celsius. (Specific heats for many substances are tabulated in handbooks of chemical data.) Although the specific heat is not constant over all temperatures, it remains constant for many substances over fairly broad ranges of temperatures (such as in this experiment). Specific heats are quoted in units of kilojoules per gram per degree, kJ/g°C (or in molar terms, in units of kJ/mol°C).

To determine the calorimeter constant for the simple coffee-cup apparatus to be used in the later choices of this experiment, we will make use of the conservation of energy principle: Energy cannot be created or destroyed during a process, but can only be transformed from one form to another or transferred

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from one part of the universe to another. A measured quantity of cold water is placed in the calorimeter to be calibrated and is allowed to come to thermal equilibrium with the calorimeter. Then a measured quantity of warm water is added to the cold water in the calorimeter. Since the energy contained in the hot water is conserved, we can make the following accounting of energy:

$$Q_{\text{warm water}} = -[Q_{\text{cold water}} + Q_{\text{calorimeter}}] \tag{1-2}$$

The minus sign in this statement is necessary because the warm water is *losing* energy, whereas the cold water and calorimeter are *gaining* energy (these processes have the opposite sense from one another). Since the calorimeter is considered a complete single unit, the amount of heat absorbed by the calorimeter, $Q_{\text{calorimeter}}$, can be written as

$$Q_{\text{calorimeter}} = C_{\text{calorimeter}}\,\Delta T \tag{1-3}$$

in which ΔT is the temperature change undergone by the calorimeter, and $C_{\text{calorimeter}}$ is the calorimeter constant, which represents the number of kilojoules of heat required to warm the calorimeter by 1°C.

Applying Equations 1-1 and 1-3 to the accounting of the energy transferred in the system as given in Equation 1-2, we can say the following:

$$(mC\Delta T)_{\text{warm water}} = -[(mC\Delta T)_{\text{cold water}} + (C_{\text{calorimeter}}\Delta T)] \tag{1-4}$$

Since the specific heat of water is effectively constant over the range of temperatures in this experiment (C_{water} = 4.18 J/g°C), determination of the calorimeter constant amounts simply to making two measurements of mass and two measurements of changes in temperature.

SAFETY PRECAUTIONS

- **Wear safety glasses at all times while in the laboratory.**
- **Use tongs or a towel to protect your hands when handling hot glassware.**

Apparatus/Reagents Required

Plastic foam coffee cups and covers, thermometer, wire for use as a stirrer, one-hole paper punch

Procedure

Record all data and observations directly in your notebook in ink.

Nest two similar-sized plastic foam coffee cups for use as the calorimeter chamber. If the cups have been rinsed with water, dry them out.

Obtain a plastic lid that tightly fits the coffee cups. Using the paper punch, make two small holes in the lid. Make one hole near the center of the lid (for the thermometer), and one hole to the side (for the stirring wire). Assemble the stirring wire and thermometer as indicated in Figure 1-1.

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Since the density of water over the range of temperatures in this experiment is very nearly 1.00 g/mL, the amount of water to be placed in the calorimeter can be more conveniently measured by volume.

With a graduated cylinder, place 75.0 ± 0.1 mL of *cold* water into the calorimeter. Cover the calorimeter with the thermometer/stirrer apparatus.

Measure 75 ± 0.1 mL of water into a clean, dry beaker, and heat the water to 70–80°C. Stir the water with a glass rod occasionally during the heating to ensure that the temperature is as uniform as possible.

While the water is heating, monitor the temperature of the cold water in the calorimeter for 2–3 minutes to make certain that it has become constant. Record the temperature of the cold water in the calorimeter to the nearest 0.2°C.

When the water being heated has reached 70–80°C, use tongs or a towel to remove the beaker from the heat. Allow the beaker to stand on the lab bench for 2–3 minutes, stirring the water occasionally during this time period. After the standing period, record the temperature of the hot water to the nearest 0.2°C.

Quickly remove the lid from the calorimeter, and pour the hot water into the cold water in the calorimeter. Immediately replace the lid of the calorimeter, stir the water with the stirring wire for 30 seconds to mix, and begin monitoring the temperature of the water in the calorimeter. Record the *highest* temperature reached by the water in the calorimeter, to the nearest 0.2°C.

From the masses (volumes) of cold and hot water used, and from the two temperature changes, calculate the calorimeter constant for your calorimeter.

Repeat the experiment twice to obtain additional values for the calorimeter constant. Use the mean value of the three determinations of the calorimeter constant for the other choices of this experiment.

Choice II. Specific Heats of Metals and Glass

Introduction

The specific heat, *C*, of a substance represents the quantity of heat energy (in joules) required to warm one gram of the substance by one Celsius degree. Although the specific heats of many substances are relatively constant over broad ranges of temperatures (such as those likely to be encountered in the general chemistry laboratory), the specific heat is dependent on the temperature. Generally the temperature range over which a particular value of the specific heat applies is quoted in the literature.

Metallic substances generally have numerically small specific heats. Metals are good conductors of heat energy and require very little input of heat energy to cause an increase in their temperature. Insulating substances, on the other hand, are very poor conductors of heat energy and have much larger specific heats. For example, the plastic foam used in the construction of the coffee-cup calorimeter in this experiment is an insulator.

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When any sample of substance undergoes a temperature change, the amount of heat energy (Q) involved in causing the temperature change is given by

$$Q = mC\Delta T$$

where m is the mass of the sample of substance, C is the specific heat of the substance, and ΔT is the temperature change undergone by the sample.

In this choice, you will determine the specific heats of an unknown metallic substance and of ordinary glass. The method used is essentially the same as in Choice I, with the sample of metal or glass replacing the hot water. A measured sample of cold water is placed in the calorimeter while a weighed metal/glass sample is being heated to 100°C (boiling water bath). The hot metal or glass is then poured into the cold water in the calorimeter, and the maximum temperature reached by the cold water as it absorbs heat from the metal/glass is determined. From the masses of cold water and metal/glass used, and from the temperature changes undergone, the specific heat of the metal/glass may be calculated:

$$Q_{\text{substance}} = -[Q_{\text{water}} + Q_{\text{calorimeter}}] \qquad (1\text{-}5)$$

$$(mC\Delta T)_{\text{substance}} = -[(mC\Delta T)_{\text{water}} + (C\Delta T)_{\text{calorimeter}}] \qquad (1\text{-}6)$$

SAFETY PRECAUTIONS

- **Wear safety glasses at all times while in the laboratory.**
- **Use tongs or a towel to protect your hands when handling hot glassware.**
- **Metal pellets are very expensive and will be collected by the instructor. Do not spill the metal pellets on the floor of the laboratory. (Clean up any accidents immediately to prevent any possible injury.)**
- **Use caution when handling glass fragments. The edges of the glass may be very sharp. Be careful not to spill the glass, and clean up any spills immediately.**

Apparatus/Reagents Required

Coffee-cup calorimeter (as designed in Choice I), metal pellets, glass beads or rings

Procedure

Record all data and observations directly in your notebook in ink.

Choice I, in which the calorimeter constant is determined for the apparatus, must be performed before this portion of the experiment. Use the same calorimeter assembly here.

Since the density of water over the range of temperatures in this experiment is nearly 1.00 g/mL, the amount of water to be placed in the calorimeter can be more conveniently measured by volume.

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With a graduated cylinder, place 75.0 ± 0.1 mL of cold water into the calorimeter. Cover the calorimeter with the thermometer/stirrer apparatus.

Obtain an unknown metal sample and record its identification code number in your notebook and on the report page. Weigh out approximately 50 g of the metal sample, and record the precise weight taken to the nearest 0.1 g.

Set up a 600-mL beaker half-filled with water on a ringstand and heat the water to boiling. When the water is boiling, record its temperature to the nearest 0.1°C.

Transfer the unknown metal sample to a clean, *dry* test tube, and heat the test tube in the boiling water bath for at least 10 minutes to allow the metal to reach the temperature of the boiling water.

While the metal sample is heating in the boiling water, monitor the temperature of the cold water in the calorimeter for 2–3 minutes to make certain that the temperature is constant. Record the temperature of the cold water to the nearest 0.2°C.

Remove the cover from the calorimeter, and quickly transfer the metal pellets to the calorimeter. Cover the calorimeter, stir the water for 60 seconds, and record the highest temperature reached to the nearest 0.2°C.

From the mass of metal sample used and the mass (volume) of cold water in the calorimeter, and from the temperature changes and calorimeter constant (see Choice I), calculate the specific heat of the unknown metal.

Repeat the determination twice more, and calculate a mean value from your three determinations for the specific heat of the unknown metal. The same metal sample may be used for these repeat determinations, but the pellets must be dried completely on a paper towel before reheating in the boiling water bath. If no pellets are lost during any point of the procedure, it is not necessary to reweigh the metal sample.

Make three determinations of the specific heat of glass beads or rings by the same method used for the metal pellets. Be careful of any sharp edges on the glass rings, and clean up any spills of glass immediately.

Choice III. Heat of Acid/Base Reactions

Introduction

Many chemical reactions are performed routinely with the reactant species dissolved in water. The use of such solutions has many advantages over the use of "dry chemical" methods. The presence of a solvent matrix permits easier and more intimate mixing of the reactant species, solutions may be measured by volume rather than mass, and the presence of water may act as a moderating agent in the reaction.

Heats of reaction between species dissolved in water are especially easy to measure because the measurement may be performed in a simple calorimeter

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as described in the earlier choices of this experiment. A measured volume of solution of one of the reagents is placed in the calorimeter cup and its temperature is determined. The second reagent solution is prepared in a separate container and is allowed to come to the same temperature as the solution in the calorimeter. When the two solutions have come to the same temperature, they are combined in the calorimeter, and the temperature of the calorimeter contents is monitored. If the chemical reaction that takes place is **exergonic**, the temperature of the water in the calorimeter will *increase* as energy is transferred to it from the reagents. If the chemical reaction is **endergonic**, the temperature of the water in the calorimeter will *decrease* as thermal energy is drawn from the water into the reactant substances.

For the purpose of tabulating heats of reaction in the chemical literature, such heats are usually converted to the basis of the number of *kilojoules of heat energy* that flows in the reaction *per mole* of reactant (or product). A typical experimental determination may only use a small fraction of a mole of reactant, and so only a few joules of heat energy will be involved in the experiment, but the results are converted to the basis of one mole. When such a determination of heat flow is conducted in a calorimeter that is equilibrated with the constant pressure of the atmosphere, the heat flow in kJ/mole is given the symbol ΔH, and is referred to as the **enthalpy change** for the reaction. Handbooks of chemical data list the enthalpy changes for many reactions.

You will measure the heat of reaction for four "different" reactions:

$$HCl + NaOH \rightarrow$$
$$HCl + KOH \rightarrow$$
$$HNO_3 + NaOH \rightarrow$$
$$HNO_3 + KOH \rightarrow$$

Each of these reactions represents the neutralization of an acid by a base, and although the reactions formally appear to involve different substances, the net reaction occurring in each case is the same:

$$H^+ \text{ (\textit{from the acid})} + OH^- \text{ (\textit{from the base})} \rightarrow H_2O$$

The actual reaction that occurs in each situation is the combination of a proton with a hydroxide ion, producing a water molecule. For this reason, the heat flows for each of these reactions should be the same.

SAFETY PRECAUTIONS

- **Wear safety glasses at all times while in the laboratory.**
- **Although the acids and bases used in this experiment are in relatively dilute solution, the acids/bases will concentrate if spilled on the skin and the water is allowed to evaporate. Wash immediately if these substances are spilled.**

Apparatus/Reagents Required

2 *M* hydrochloric acid, 2 *M* nitric acid, 2 *M* sodium hydroxide, 2 *M* potassium hydroxide, calorimeter/thermometer/stirrer apparatus as used in the previous choices

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Procedure

Record all data and observations directly in your notebook in ink.

Choice I, in which the calorimeter constant is determined for the apparatus, must be performed before this portion of the experiment. Use the same calorimeter assembly here.

The procedure that follows is written in terms of the reaction between hydrochloric acid and sodium hydroxide. Perform this determination first; then repeat the procedure for each of the other three acid/base combinations indicated in the equations in the Introduction.

Obtain 75 mL of 2 *M* NaOH and place it in the calorimeter. Obtain 75 mL of 2 *M* HCl in a clean dry beaker. Allow the two solutions to stand until their temperatures are the same (within ±0.5 degree). Be sure to rinse off and dry the thermometer when transferring between the solutions to prevent mixing of the reagents prematurely. Record the temperature(s) of the solutions to the nearest 0.2°C.

Add the HCl from the beaker *all at once* to the calorimeter, cover the calorimeter quickly, stir the mixture for 30 seconds, and record the highest temperature reached by the mixture (to the nearest 0.2°C).

From the change in temperature undergone by the mixture upon reaction, the total mass (volume) of the combined solutions, and the calorimeter constant for the apparatus (see Choice I), calculate the quantity of heat that flowed from the reactant species into the water of the solution.

Calculate the number of moles of water produced when 75 mL of 2 *M* HCl reacts with 75 mL of 2 *M* NaOH.

Calculate ΔH in terms of the number of kilojoules of heat energy transferred when 1 mol of water is formed in the neutralization of aqueous HCl with aqueous NaOH.

Repeat the determination of ΔH for the HCl/NaOH reaction twice, and calculate a mean value for ΔH for the reaction.

Repeat the procedure for the other combinations of acids and bases:

$$HCl + KOH \rightarrow$$
$$HNO_3 + NaOH \rightarrow$$
$$HNO_3 + KOH \rightarrow$$

Calculate a mean value and standard deviation for ΔH for the net reaction occurring in each of the mixtures, i.e.,

$$H^+ + OH^- \rightarrow H_2O$$

from your twelve determinations.

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Choice IV. Heat of Metal/Acid Reactions

Introduction

Some metallic elemental substances are easily able to displace hydrogen from its compounds. For example, the alkali metals and some of the alkaline earth metals react with cold water, producing the metal hydroxide and releasing gaseous elemental hydrogen:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$

Other less reactive metals, although not able to displace hydrogen from water, will evolve elemental hydrogen gas from solutions of acids. For example,

$$Mg + 2HCl \rightarrow MgCl_2 + H_2 \qquad (Mg + 2H^+ \rightarrow Mg^{2+} + H_2)$$

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2 \qquad (Zn + 2H^+ \rightarrow Zn^{2+} + H_2)$$

In Choice III, the enthalpy changes for various strong acids reacting with strong bases were determined, and it was noted that, because the net reaction occurring was the same in each case, the enthalpy changes should be the same for each combination. For the reactions of metallic substances with acids to be studied in this choice, based on the net reactions shown above, the enthalpy change determined should depend on which metal is used, but should not be a function of which strong acid is used.

SAFETY PRECAUTIONS

- **Wear safety glasses at all times while in the laboratory.**
- **Although the acids used in this experiment are in relatively dilute solution, the acids may concentrate if allowed to evaporate. Wash immediately if the acids come in contact with skin, and clean up any spills in the laboratory with large amounts of water.**
- **The acids used in this experiment are damaging to clothing.**
- **Hydrogen gas is generated in this experiment. Hydrogen gas is extremely flammable. Make certain that the room is well ventilated. No flames will be permitted in the room while hydrogen is being generated.**

Apparatus/Reagents Required

1 *M* hydrochloric acid, 1 *M* nitric acid, 1 *M* sulfuric acid, calorimeter/thermometer/stirrer apparatus as used in the previous choices, magnesium turnings, mossy or granulated zinc

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Procedure

Record all observations and data directly in your lab notebook in ink.

Choice I, in which the calorimeter constant is determined for the apparatus, must be performed before this portion of the experiment. Use the same calorimeter assembly here.

For the dilute acid solutions to be used in this experiment, it may be assumed that the densities of the solutions are very nearly 1.0 g/mL. For this reason, a particular mass of solution may be conveniently measured out by volume.

Obtain 150 ± 1 mL of 1 *M* HCl and place it in the calorimeter. After covering the calorimeter, determine the temperature of the acid to the nearest 0.2°C. Monitor the temperature of the acid over a 5-minute period to make certain that the temperature is constant.

Weigh out a 0.5-g sample of magnesium turnings and record the exact mass to the nearest milligram (0.001 g).

Remove the cover of the calorimeter, quickly add the magnesium turnings, and replace the cover. Stir the mixture and monitor the temperature as the reaction occurs. Record the highest temperature reached.

Repeat the reaction between magnesium and hydrochloric acid twice.

Taking the specific heat of the dilute hydrochloric acid solution to be the same as that of water, calculate the number of joules of heat energy released for each of the individual Mg/HCl reactions.

Based on the mass of magnesium taken in each individual run, calculate the enthalpy change for each run, in terms of the number of kJ of heat energy transferred per mole of magnesium reacting.

Calculate the mean value for the enthalpy change for the reaction of magnesium with hydrochloric acid.

Perform similar sets of determinations for the reaction between magnesium and 1 *M* nitric acid, and for the reaction between magnesium and 1 *M* sulfuric acid.

Based on these nine determinations, calculate a mean value and standard deviation for the reaction between magnesium metal and aqueous hydrogen ion.

Repeat the processes, determining the enthalpies of reaction of zinc metal with each of the three acids available.

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Choice V. Heat of Solution of a Salt

Introduction

When salts dissolve in water, the positive and negative ions of the salt *interact* with water molecules. Water molecules are highly polar, and arrange themselves in a layer around the ions of the salt so as to maximize electrostatic attractive forces. Such a layer of water molecules surrounding an ion is called a **hydration sphere**.

For example, consider dissolving the salt potassium bromide, KBr, in water. As the ions enter solution, water molecules would orient their dipoles in a particular manner. The potassium ions would become surrounded by a layer of water molecules in which the *negative* ends of the water dipoles would be oriented toward the positive potassium ions. Similarly, the bromide ions would become surrounded by a layer of water molecules in which the *positive* ends of the dipoles would be oriented toward the *negative* bromide ions. The development of such hydration spheres can have a large effect on the properties of ions in solutions, which you will study in depth if you go on to take a course in physical chemistry.

Formation of a hydration sphere around an ion necessarily involves an energy change. In this Choice, you will determine the energy change for such a process.

SAFETY PRECAUTIONS

- **Wear safety glasses at all times while in the laboratory.**
- **The salts used in this Choice may be toxic. Wash after handling them and dispose of them as directed by the instructor.**

Apparatus/Reagents Required

Salt sample, calorimeter/thermometer/stirrer apparatus as used in the previous choices

Procedure

Record all data and observations directly in your notebook in ink.

Choice I, in which the calorimeter constant is determined for the apparatus, must be performed before this portion of the experiment is begun. Use the same calorimeter assembly here.

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Place 75.0 mL of distilled water in the calorimeter. Determine and record the temperature of the water to the nearest 0.2°C. Monitor the temperature of the water for 3 minutes to make certain that its temperature does not change.

Obtain a salt sample from your instructor. If the salt's identity is known, record the name and formula of the salt. If the salt is presented as an unknown, record the code number of the unknown sample. If an unknown is presented, the instructor will also tell you the molar mass of the unknown salt (record).

Using a clean, dry beaker, weigh out a 5–6-gram sample of the salt, recording the exact mass taken to the nearest 0.01 g.

Remove the lid from the calorimeter and quickly add the weighed salt sample. Replace the lid of the calorimeter and immediately agitate the solution with the stirring wire. Monitor the temperature of the solution while continuing to stir the solution. Record the highest—or lowest—temperature reached as the salt dissolves in the water.

From the mass of salt taken and the mass (volume) of water used, and from the temperature change and calorimeter constant, calculate the quantity of heat that flowed during the dissolving of the salt. From this quantity of heat, and from the molar mass of the salt, calculate the enthalpy change for the dissolving of the salt (heat of solution).

Dispose of the salt solution as directed by the instructor.

Clean out and dry the calorimeter, and repeat the determination twice more. Calculate a mean value for the heat of solution of your salt from your three determinations.

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Calorimetry

Date:	....	Student name:	....
Course:	....	Team members:	....
Section:	....		....
Instructor:	....		....

Prelaboratory Questions

Choice I. Determination of a Calorimeter Constant

1. What is the definition of the *joule* in terms of the basic SI units?

2. A calorimeter is to be calibrated: 51.203 g of water at 55.2°C is added to a calorimeter containing 49.783 g of water at 23.5°C. After stirring and waiting for the system to equilibrate, the final temperature reached is 37.6°C. Calculate the calorimeter constant.

Choice II. Specific Heats of Metals and Glass

1. Using a handbook of chemical data, look up the specific heat capacities of the following materials. Give your references.

Substance	Specific heat	Reference
Al(*s*)	....	....
C(*s*)	....	....
Fe(*s*)	....	....
$H_2O(l)$	....	....
Pb(*s*)	....	....

2. What are the *units* of specific heat capacity?

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Choice III. Heat of Acid/Base Reactions

1. The acids and bases to be determined in this experiment are all classified as *strong* acids or bases. Use your textbook to find what is meant by the word *strong* in this context.

2. What is meant by *neutralization*?

Choice IV. Heat of Metal/Acid Reactions

1. If adding a 0.552 g sample of Metal X to 150 mL of 1.0 *M* HCl caused a temperature increase of 7.33°C, calculate the quantity of heat that has flowed.

2. If the atomic weight of Metal X is 113, calculate ΔH for the above process.

Choice V. Heat of Solution of a Salt

1. Use your textbook or a chemical encyclopedia to write a specific definition of a *salt*.

2. What does it mean to say that an ion becomes *hydrated* when a salt is dissolved in water?

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Calorimetry

Date:	....	Student name:	....
Course:	....	Team members:	....
Section:	....		....
Instructor:	....		....

Results/Observations

Choice I. Determination of a Calorimeter Constant

	Trial 1	**Trial 2**	**Trial 3**
Mass (volume) of cold water	....	....	....
Temperature of cold water	....	....	....
Mass (volume) of hot water	....	....	....
Temperature of hot water	....	....	....
Final temperature reached	....	....	....
Temperature change, ΔT	....	....	....
Calorimeter constant	....	....	....

Mean value of calorimeter constant ....

Questions

1. What effect on the calorimeter constant calculated would be observed if the calorimeter cup were made of a conducting material (such as metal) rather than plastic foam?

2. Why is water typically used as the heat-absorbing liquid in calorimeters?

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3. What is the significance of the minus sign in Equation 1-2 in the Introduction of Choice I of this experiment?

4. A unit of heat energy that was formerly used frequently was the *calorie*. Look up the definition of the calorie in your textbook or a handbook and record it here.

5. Calculate the calorimeter constant for your calorimeter in cal/°C.

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Calorimetry

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Course:	....	Team members:	....
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Results/Observations

Choice II. Specific Heats of Metals and Glass

Specific heat of a metal

Metallic substance used ....

	Trial 1	**Trial 2**	**Trial 3**
Mass of substance taken	....	....	....
Mass (volume) cold water used	....	....	....
Initial temperature of water	....	....	....
Final temperature of water	....	....	....
Temperature change, ΔT	....	....	....
Specific heat of substance	....	....	....

Mean value of specific heat of substance ....

Literature value .... Reference ....

Percent error in specific heat determination ....

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Specific heat of glass

	Trial 1	Trial 2	Trial 3
Mass of glass taken	....	....	....
Mass (volume) cold water used	....	....	....
Initial temperature of water	....	....	....
Final temperature of water	....	....	....
Temperature change, ΔT	....	....	....
Specific heat of glass	....	....	....

Mean value of specific heat of glass ....

Literature value .... Reference ....

Percent error in specific heat determination ....

Questions

1. In the pre-lab question, you were asked to look up the specific heat of ice [i.e., $H_2O(s)$]. Why is the value for ice not the same as for liquid water?

2. You determined the specific heat of glass. There are many types of glass. Use a chemical handbook to find the chemical composition of three types of glass, and tell how their specific heats differ.

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Calorimetry

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Course:	....	Team members:	....
Section:	....		....
Instructor:	....		....

Results/Observations

Choice III. Heat of Acid/Base Reactions

Reaction of HCl and NaOH

	Trial 1	**Trial 2**	**Trial 3**
Volume of 2 *M* NaOH used			
Initial temperature of NaOH			
Volume of 2 *M* HCl used			
Initial temperature of HCl			
Final temperature reached			
Total mass (volume) of mixture			
Temperature change, ΔT			
Heat flow, joules			
Moles of water produced			
ΔH (kJ/mol water)			

Mean value of ΔH .... Literature value ....

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Reaction of HCl and KOH

	Trial 1	Trial 2	Trial 3
Volume of 2 *M* KOH used			
Initial temperature of KOH			
Volume of 2 *M* HCl used			
Initial temperature of HCl			
Final temperature reached			
Total mass (volume) of mixture			
Temperature change, ΔT			
Heat flow, joules			
Moles of water produced			
ΔH (kJ/mol water)			

Mean value of ΔH .... Literature value ....

Reaction of HNO_3 and NaOH

	Trial 1	Trial 2	Trial 3
Volume of 2 *M* NaOH used			
Initial temperature of NaOH			
Volume of 2 *M* HNO_3 used			
Initial temperature of HNO_3			
Final temperature reached			
Total mass (volume) of mixture			
Temperature change, ΔT			
Heat flow, joules			
Moles of water produced			
ΔH (kJ/mol water)			

Mean value of ΔH .... Literature value ....

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Reaction of HNO_3 and KOH

	Trial 1	Trial 2	Trial 3
Volume of 2 *M* KOH used			
Initial temperature of KOH			
Volume of 2 *M* HNO_3 used			
Initial temperature of HNO_3			
Final temperature reached			
Total mass (volume) of mixture			
Temperature change, ΔT			
Heat flow, joules			
Moles of water produced			
ΔH (kJ/mol water)			

Mean value of ΔH .... Literature value ....

Questions

1. The heat flows measured in this experiment were actually not for the simple neutralization of a proton and hydroxide ion (as indicated in the Introduction for Choice III); rather they include contributions based on the fact that these species are hydrated in aqueous solution. What does it mean to say that a proton is hydrated, and how will the heat of hydration affect the measured heat flow for a neutralization reaction?

2. Based on your 12 measurements, what is the mean value and the standard deviation for ΔH for the reaction between H^+ and OH^- in aqueous solution?

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Calorimetry

Date:	....	Student name:	....
Course:	....	Team members:	....
Section:	....		....
Instructor:	....		....

Results/Observations

Choice IV. Heat of Metal/Acid Reactions

Reaction of HCl and Mg metal

	Trial 1	Trial 2	Trial 3
Mass of Mg used, g	....	....	....
Volume of 1 M HCl used	....	....	....
Initial temperature of HCl	....	....	....
Final temperature reached	....	....	....
Temperature change, ΔT	....	....	....
Heat flow, joules	....	....	....
Moles of Mg used	....	....	....
ΔH (kJ/mol Mg)	....	....	....

Mean value of ΔH .... Literature value ....

Reaction of HNO_3 and Mg metal

	Trial 1	Trial 2	Trial 3
Mass of Mg used, g	....	....	....
Volume of 1 M HNO_3 used	....	....	....
Initial temperature of HNO_3	....	....	....
Final temperature reached	....	....	....
Temperature change, ΔT	....	....	....

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	Trial 1	Trial 2	Trial 3
Heat flow, joules			
Moles of Mg used			
ΔH (kJ/mol Mg)			

Mean value of ΔH ________ Literature value ________

Reaction of H_2SO_4 and Mg metal

	Trial 1	Trial 2	Trial 3
Mass of Mg used, g			
Volume of 1 M H_2SO_4 used			
Initial temperature of H_2SO_4			
Final temperature reached			
Temperature change, ΔT			
Heat flow, joules			
Moles of Mg used			
ΔH (kJ/mol Mg)			

Mean value of ΔH ________ Literature value ________

Reaction of HCl and Zn metal

	Trial 1	Trial 2	Trial 3
Mass of Zn used, g			
Volume of 1 M HCl used			
Initial temperature of HCl			
Final temperature reached			
Temperature change, ΔT			
Heat flow, joules			
Moles of Zn used			
ΔH (kJ/mol Zn)			

Mean value of ΔH ________ Literature value ________

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Reaction of HNO_3 and Zn metal

	Trial 1	**Trial 2**	**Trial 3**
Mass of Zn used, g			
Volume of 1 M HNO_3 used			
Initial temperature of HNO_3			
Final temperature reached			
Temperature change, ΔT			
Heat flow, joules			
Moles of Zn used			
ΔH (kJ/mol Zn)			

Mean value of ΔH .... Literature value ....

Reaction of H_2SO_4 and Zn metal

	Trial 1	**Trial 2**	**Trial 3**
Mass of Zn used, g			
Volume of 1 M H_2SO_4 used			
Initial temperature of H_2SO_4			
Final temperature reached			
Temperature change, ΔT			
Heat flow, joules			
Moles of Zn used			
ΔH (kJ/mol Zn)			

Mean value of ΔH .... Literature value ....

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Questions

1. Why would one not expect the heat of reaction to depend on which acid is used for the reaction?

2. Which metal, zinc or magnesium, releases more heat energy per gram when reacted with acid? Which metal releases more heat energy per mole?

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Calorimetry

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Section: ....

Instructor: ....

Results/Observations

Choice V. Heat of Solution of a Salt

	Trial 1	Trial 2	Trial 3
Volume of water used, mL			
Initial temperature of water, °C			
Mass of salt taken, g			
Moles of salt taken, mol			
Highest/Lowest temperature, °C			
Calculated heat flow, J			
Enthalpy change (ΔH), kJ/mol			

Mean heat of solution, kJ/mol ....

Questions

1. If the identity of the salt was provided by the instructor, use a chemical handbook to find the literature value of the heat of solution for your salt. Calculate the percent error between the literature value and your mean experimental value for ΔH of solution for your salt.

   Literature value for ΔH .... Reference .... % error ....

2. Why is the dissolving of a salt sometimes an exothermic and sometimes an endothermic process?

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