# Calorimetry

#### **Performance Goals**

**8–1** Calculate the specific heat of an unknown solid element by measuring the heat exchanged in a calorimeter.

#### CHEMICAL OVERVIEW

Heat, a form of energy, can be gained or lost by an object. When the object cools, it loses heat energy; when it is heated, it gains energy. The unit in which heat is measured is the **joule**, **J.** A joule is a derived unit, having base units of  $kg \cdot m^2/sec^2$ . The joule is a very small amount of heat, so the **kilojoule**, **kJ**, is commonly used.\*

**Heat flow** is the change in "heat content" of an object as heat energy passes between the object and its surroundings. It is proportional to the mass of the object and its change in temperature. The proportional relationship becomes an equation if the proportionality constant called **specific heat** is introduced:

$$Q = (\text{mass})(\text{specific heat})(\Delta T)$$
  
=  $m \times c \times \Delta T$  (8.1)

where Q is the heat flow in joules, m is mass in grams, c is specific heat in joules per gram degree, or  $J/g \cdot {}^{\circ}C$ , and  $\Delta T$  is the temperature change, or final temperature minus initial temperature. The Greek letter delta,  $\Delta$ , indicates a change in a measured value and always means final value minus initial value:  $\Delta X = (X_f - X_i)$ .

Specific heat is a property of a pure substance. It is the number of joules of energy that are required to raise the temperature of 1 g of the substance by 1°C. The specific heat of water is  $4.18 \, \text{J/g} \cdot ^{\circ}\text{C}$ . This value is used often in calorimetry experiments; it is one you should remember.

When a "hot" object comes into contact with a "cold" object—that is, when an object at higher temperature comes into contact with an object at lower temperature—heat flows from the hot object to the cold object. The hot object raises the temperature of the cold object, and the cold object cools the hot object. Eventually they reach the same intermediate temperature.

<sup>\*</sup>The joule is still in the process of replacing the calorie as the standard heat unit in chemistry, and the calorie is still in common use. The calorie is now defined as exactly 4.184 joules.

Heat flow between objects can be measured in a **calorimeter**. A perfect calorimeter is an isolated segment of the universe that allows no heat to flow to or from its contents during an experiment. It follows from the law of conservation of energy that, in a perfect calorimeter,

$$\sum Q = 0 \tag{8.2}$$

where  $\sum Q$  is the sum of all individual changes in heat content within the calorimeter. We will assume that the calorimeters used in this experiment are "perfect." This means that any heat transferred between its contents and the surroundings or between its contents and the calorimeter itself are negligible and may be disregarded.

In this experiment, you will measure the specific heat of two metals, one known and the other unknown. The known metal will be copper, which has a specific heat of 0.38 J/g· $^{\circ}$ C. You will calculate the percent error from this known value. The same experimental method will be used to find the specific heat of the unknown metal.

The laboratory procedure is to heat a weighed metal sample in boiling water until it reaches the temperature of the water. This is the initial temperature of the metal. The mass and initial temperature of water in a "coffee cup" calorimeter are measured. The metal is placed into the calorimeter, and the final temperature reached by both is recorded. According to Equation 8.2, the heat flow of the metal plus the heat flow of the water add up to zero:

$$Q_w + Q_m = 0 (8.3)$$

$$Q_w = -Q_m \tag{8.4}$$

What Equation 8.4 says is that the heat lost by the metal is gained by the water. Substituting the expressions of Equation 8.1 into Equation 8.4 gives

$$m_w \times c_w \times \Delta T_w = -(m_m \times c_m \times \Delta T_m) \tag{8.5}$$

All values are known except the specific heat of the metal. Note the negative sign on the right side of the equation.

#### SAFETY PRECAUTIONS

No hazardous chemicals are used in this experiment—only hazardous temperatures. You will be working with fairly large quantities of boiling water, pouring from one container to another, and dropping pieces of metal into boiling water. Do it carefully, anticipating and avoiding the conditions that might cause you to lose control of a beaker and burn yourself or others as a result. Also in this experiment you will be working with Bunsen burners. If you have long hair, be sure to tie it back.

#### PROCEDURE

**NOTE:** Record all mass measurements in grams to the nearest 0.01 g. Record all temperature measurements in degrees Celsius to the nearest 0.1°C. The smallest graduation on your thermometer must be no larger than 0.2°C for satisfactory results on this experiment.

#### 1. Preparation for Calorimeter Runs

- **A.** Place about 500 mL of tap water into a beaker. Set it aside so it will come to room temperature. *Do not place it near an operating Bunsen burner*. This will be your source of calorimeter water throughout the experiment.
- **B.** Fill a beaker, 600 mL or larger, to about 80% of capacity with deionized water. (The deionized water is not required from a purity standpoint, but its use avoids the buildup of hard-to-remove scale that forms when tap water is boiled in a glass vessel.) Mount the beaker over wire gauze on a ring stand or, if available, use a hot plate and heat the water to boiling. Proceed to Parts IC and 1D while waiting.

From time to time in the experiment you will have to replace the water that has boiled out of the beaker with other water *already at the boiling temperature*. If some "community" source of boiling water is not available, establish your own by setting up a second beaker in the same manner as the first.

- **C.** Select a piece of copper and an unknown metal sample from among those provided. Record the identification number of the unknown on your data sheet. Measure the mass of each piece and record it in grams to the nearest centigram.
- **D.** Using crucible tongs, place each piece of metal in the water being heated. Position them in such a way that they will be easy to pick up and transfer to the calorimeter later. Be careful not to drop the metal in such a manner that it hits the bottom with sufficient force to break the beaker. The metal piece should remain in the boiling water for at least 30 minutes. As you proceed with the next steps, keep watch over the water level, which must be high enough to cover the metal completely. If too much water boils away, replenish it from the community source of boiling water, or from your second beaker, as explained in Part 1B.
- E. When the water in the heating bath is boiling, determine and record its temperature to the nearest 0.1°C. Set the thermometer aside and allow it to cool to room temperature. Two assumptions are made about the temperature of the boiling water: first, that it will remain constant throughout the experiment; and second, that the metal immersed in it will reach the same temperature, which is the initial temperature of the metal for later calculations.
- F. Select and prepare a place on the laboratory desk for your calorimeter runs. It should be as far as possible from an operating Bunsen burner from which it might absorb heat, but not so far that the metal loses temperature while being transferred from the boiling water to the calorimeter. It is convenient to place a split rubber stopper on your thermometer (see Figure LP-6 in the Laboratory Procedures section) and clamp it on a ring stand, or to place your thermometer in a buret clamp, as depicted in Figure 8.1.

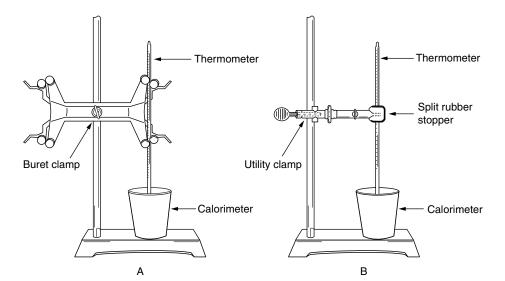


Figure 8.1

#### 2. First Calorimeter Run

- **A.** Measure exactly 100 mL of your calorimeter water in a graduated cylinder and carefully transfer it into your calorimeter. Assuming the density of water to be 1.00 g/mL, the mass of the water is then 100 grams.
- **B.** Place the thermometer into the calorimeter water, as shown in Figure 8.1. When the temperature has remained constant for about 1 minute, record that temperature to the nearest 0.1°C. In this and all other temperature measurements, be sure the thermometer bulb is totally immersed in the water.
- C. (The next step is critical. The period in which the metal is out of water must be held to an absolute minimum to keep heat loss to the air as low as possible.) Using crucible tongs, lift the piece of copper from the boiling water and hold it above the water level but below the top of the beaker for about 2 seconds to allow the boiling water to drain off. Then, as quickly as possible, place the metal into the calorimeter water. Be sure it is completely submerged. Be careful not to splash water out of the calorimeter. If this does occur, return the metal to the boiling water and start again at Step 2A.
- **D.** Continuously and gently stir the water with a glass stirring rod, being careful not to break the bulb of the thermometer. Watch the thermometer. Read and record to the nearest 0.1 °C the highest temperature that is reached before it begins to drop again. This may take several minutes.
- **E.** Pour the water from the calorimeter into the sink. Place the copper back into the boiling water to be heated for a second run.

## 3. Subsequent Calorimeter Runs

Again pour exactly 100 mL of water into your calorimeter. Repeat Steps 2B through 2E, using the unknown metal. The metal should be used alternately in subsequent runs, giving each sample enough time in the boiling water to reach its temperature. Make two runs with your copper sample and two with the unknown or as directed by your instructor. At the end of the experiment, dry the metal pieces and return them to the place indicated by your instructor.

### CALCULATIONS

Calculate the specific heat of the metal for each run of the experiment, using Equation 8.5, in which the specific heat is the only unknown. Find the percent error for each run with the copper, using the equations

$$\begin{aligned} \text{Percentage error} &= \frac{\text{error}}{\text{accepted value}} \times 100 \\ \text{Percentage error} &= \frac{|\text{experimental value} - \text{accepted value}|}{\text{accepted value}} \times 100 \end{aligned}$$

Find the average specific heat for the two runs with the unknown metal.

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Name	Date	Section
Experiment 8		
Advance Study Assignment		
<b>1. A.</b> Define specific heat.		
<b>B.</b> Why do you have to transfer the c	copper rapidly from the boiling	water into the calorimeter?
<b>C.</b> What error would be committed if	f the boiling water did not cove	r the metal completely?
Ci vinat ciror would be committed in	the coming water and not cover	t the mean completely.
<b>2.</b> A student places 1.38 g of unknown system reaches a uniform temperature		
<b>3.</b> If the actual specific heat of the meta	al in Problem 2 is 0.25 J/g·°C, c	alculate the percentage error.

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Name	Date	Section

## **Experiment 8**

### Work Page

### Data

Run number			
Metal or unknown number			
Mass of metal (g)			
Temperature of boiling water (°C)			
Initial temperature of calorimeter water (°C)			
Final temperature (°C)			

### Calculations and Results

Record all results in the table on the following page. Show below and on the next page all calculations for one column of data and results.

**104** Introduction to Chemical Principles: A Laboratory Approach ■ Weiner and Harrison

### Calculations and Results

Mass of water (g)			
Δ <i>T</i> of water (°C)			
Δ <i>T</i> of metal (°C)			
Heat flow (J) (absolute value)			
Specific heat (J/g·°C)			
Percentage error (copper only)			

Name	Date	Section

## **Experiment 8**

### **Report Sheet**

### Data

Run number			
Metal or unknown number			
Mass of metal (g)			
Temperature of boiling water (°C)			
Initial temperature of calorimeter water (°C)			
Final temperature (°C)			

### Calculations and Results

Record all results in the table on the following page. Show below and on the next page all calculations for one column of data and results.

**106** Introduction to Chemical Principles: A Laboratory Approach ■ Weiner and Harrison

### Calculations and Results

Mass of water (g)			
$\Delta T$ of water (°C)			
$\Delta T$ of metal (°C)			
Heat flow (J) (absolute value)			
Specific heat (J/g⋅°C)			
Percentage error (copper only)			