Specific Heat of Metals

Heat flows from a warmer object to a cooler object. As heat flows, the temperature of the warmer object decreases and the temperature of the cooler object increases. The magnitude of the temperature change depends in part upon what each object is made of. Objects that experience a large temperature change when they absorb or release a given amount of heat have a low specific heat. Objects that experience a small temperature change for a given amount of heat transfer have a high specific heat.

**Problem**
How can you use water to measure the specific heat of metals?

**Objectives**
- **Construct** a calorimeter.
- **Measure** changes in the temperature of water in the calorimeter when warmer metals are added.
- **Calculate** the specific heat of each metal.

**Materials**
- aluminum, iron, and lead samples
- plastic-foam cups (2)
- 400-ml beakers (2)
- 100-ml graduated cylinder
- large test tubes (3)
- boiling chips
- scissors
- balance
- hot plate
- thermometer
- test-tube rack

**Safety Precautions**
- Always wear safety goggles, gloves, and a lab apron.
- Broken glassware may have sharp edges. Be careful when using scissors.
- Use caution and proper protection when handling hot objects.
- Avoid spilling water on the hot plate power cord.
- Wipe up any water spills immediately to avoid slipping.

**Pre-Lab**
1. Define specific heat.
2. A sample of substance X has a mass of 123 g. When the sample releases 795 J of heat, its temperature falls from 45.1°C to 17.6°C. What is the specific heat of substance X?
3. What is a calorimeter?
4. Why is it important that a calorimeter be made of an insulating material?
5. Read the entire laboratory activity. Form a hypothesis about which of the three metals will cause the largest change in water temperature for its mass. Explain your reasoning. Record your hypothesis on page 46.

**Procedure**
1. Use scissors to remove the lip from one of the plastic-foam cups. As shown in Figure A, this cup will be the top of the calorimeter. Invert it and set it on top of the other cup, which will be the bottom of the calorimeter.
2. Use a pencil to punch a hole in the center of the top of the calorimeter. The hole should be large enough to hold the thermometer.
3. Place the calorimeter in one of the beakers to keep it from tipping over.
4. Measure the mass of each metal sample to the nearest 0.1 g. Record the masses in Data Table 1. Place each metal in a separate test tube, and label each tube.
5. Add about 300 mL of water and a few boiling chips to the second beaker. Place all three test tubes in the beaker, as shown in Figure B.

6. Be sure the water level in the beaker is above the tops of the metal samples. Add more water to the beaker if necessary, but do not allow any water to get into the test tubes.

7. Set the beaker of water and test tubes on the hot plate. Turn the hot plate on and heat the water to a boil.

8. While the water is heating, pour about 75 mL of cold water into the graduated cylinder. Measure the volume to the nearest 0.1 mL. Record the volume in the data table column for aluminum.

9. Pour the cold water into the calorimeter. With the top of the calorimeter off, measure the temperature of the water every minute until it stays the same for 3 min. Record this temperature as the initial water temperature in the data table column for aluminum.

10. After the water has been boiling for 10 min, measure the temperature of the boiling water. You can assume that the metal samples are at the same temperature as the water. Record the temperature as the initial metal temperature in the data table columns for each metal. Keep the water boiling.

11. Use the test-tube holder to remove the test tube that contains the aluminum sample. CAUTION: Carefully slide the aluminum into the water in the calorimeter without splashing. Quickly put the top on the calorimeter.

12. Insert the thermometer through the hole in the calorimeter top until the tip of the thermometer touches the bottom of the calorimeter.

13. Gently swirl the beaker containing the calorimeter for 30 s while you monitor the temperature. Do not allow the metal sample to hit the thermometer. Record the highest temperature attained by the water as the final temperature in the data table column for aluminum.

14. Remove the aluminum sample from the calorimeter. Pour the water down the drain.

15. Repeat steps 8, 9, and 11–14 for the iron and lead samples.

**Hypothesis**

**Cleanup and Disposal**

1. Turn off the hot plate. After the boiling water has cooled, pour it down the drain.

2. Dry the metal samples with a paper towel.

3. Make sure your balance is left in the same condition as you found it.

4. Return the metal samples and lab equipment to their proper places.

5. Wash your hands thoroughly with soap or detergent before you leave the lab.
Data and Observations

<table>
<thead>
<tr>
<th>Data Table 1</th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Aluminum</td>
<td>Iron</td>
</tr>
<tr>
<td>Mass of metal (g)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Volume of water (mL)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Initial water temperature (°C)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Initial metal temperature (°C)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Final temperature (°C)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Change in water temperature (°C)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Change in metal temperature (°C)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Heat gained by water (J)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Specific heat, J/(g·°C)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

1. Calculate the change in water temperature caused by each metal by subtracting the initial water temperature from the final temperature. Record the results in **Data Table 1**.

2. Calculate the change in each metal's temperature by subtracting the final temperature from the initial metal temperature. Record the results in **Data Table 1**.

3. Use the equation in Section 16.1 of your textbook to calculate the amount of heat gained by the water from each metal. (To determine the mass of water, assume the density of water is 1.00 g/mL.) Record the results in **Data Table 1**.

4. Use the same equation to calculate the specific heat of each metal. (Rearrange the equation to solve for specific heat, and assume that the amount of heat lost by the metal equals the amount of heat gained by the water.) Record the results in **Data Table 1**.

**Analyze and Conclude**

1. **Applying Concepts** To calculate each metal's specific heat, you assumed that the amount of heat lost by the metal equals the amount of heat gained by the water. What factors determine whether this assumption is valid or not? (Hint: Identify the system and the surroundings in this experiment.)
2. **Drawing a Conclusion**  For each metal, divide the change in water temperature by the mass of the metal. Use the results of this calculation to evaluate your hypothesis.

3. **Observing and Inferring**  Which of these metals must release the most heat to experience a given decrease in temperature per gram of metal? Explain.

4. **Error Analysis**  Compare the specific heats you calculated for aluminum, iron, and lead with the values given in Table 16-2 of your textbook. Calculate the percent error if any. Explain possible sources of error in the lab.

---

**Real-World Chemistry**

1. Would a fishing sinker dropped into an ice-covered lake reach the temperature of the lake water more quickly if the sinker was made of iron or lead? Use your data on the specific heats of these metals to explain your answer. (Assume the sinkers have a starting temperature of 37°C and have the same shape and mass.)

2. It is possible to remove a sheet of aluminum foil from a hot oven with your bare hands without burning yourself. However, you will surely burn yourself if you touch a thick aluminum pan in the same oven with your bare hands. Why?

3. One way to identify the composition of metal fragments found at the site of an explosion is to measure the specific heat of the fragments. Suppose a fragment is found to have a specific heat of 0.129 J/(g·°C). Would this information alone be enough to identify the metal in the fragment? Explain why or why not. If not, suggest a method for identifying the metal that would not require any additional equipment.