Basic Calorimetry Set

TD-8557C

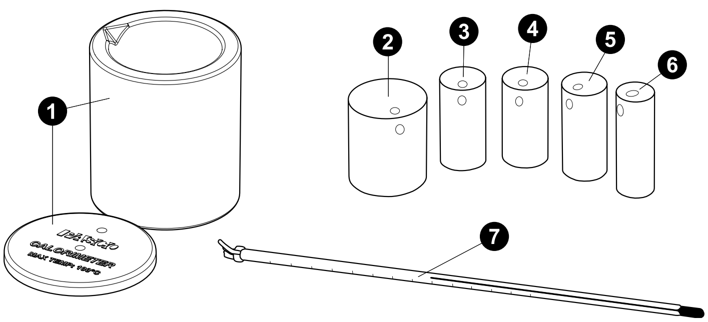
Introduction

The Basic Calorimetry Set is an affordable introduction to thermodynamics, the study of the role of heat in physical processes. With the addition of a mass balance, ice, and a heat source, the Basic Calorimetry Set provides the equipment needed to perform a variety of calorimetry activities.

The calorimeter cup has two holes in its lid for inserting a thermometer, tubing from a steam generator, or other appropriate components. The rim of the cup has a pouring notch, which makes it easier to pour liquids out of the calorimeter. The cup has 1.3 cm thick walls, is 10 cm deep, and has an inside diameter of 7.5 cm. The five metal samples each have approximately the same mass of 80 g (0.080 kg), and each sample has a pair of holes to which to tie a string, allowing you to hang the sample in water.

Equipment

Included equipment:



1. Calorimeter cup and lid
2. Stainless steel sample
3. Brass sample
4. Aluminum sample
5. Zinc sample
6. Copper sample
7. Alcohol thermometer (20 °C to 110 °C)

Recommended equipment:

- Triple beam balance, such as the Ohaus Triple-Beam Balance (SE-8707 or SE-8723)

About calorimetry

A calorimeter is a vessel or device that thermally isolates an experiment from its surroundings. Ideally, this means that the results of an experiment performed in a calorimeter are independent of the temperature of the surroundings, because no heat can flow into or out of the calorimeter.

However, no calorimeter is perfect, and there is always some unwanted and unaccountable heat flow affecting the results of any calorimetric experiment. To minimize unwanted heat flow, always follow these rules during an experiment:

- **Minimize the time between the measurements of initial and final temperatures.** In other words, perform the critical portion of the experiment quickly, so that there is as little time as possible for unwanted heat flow between measurements. (Do not rush, but carefully plan the experiment to minimize the delay.)
- **Whenever possible, ensure room temperature is approximately midway between the beginning and ending temperatures of the experiment.** When the experimental temperature is colder than room temperature, heat flows from the surroundings into the calorimeter. When the experimental temperature is hotter than room temperature, heat flows from the calorimeter into the surroundings. If the experimental temperature varies above and below room temperature by equal amounts, the heat gained and lost to the environment will be approximately equal, minimizing the net influence on the experiment.
- **Make mass measurements of liquids as near to the critical temperature measurements as possible.** This reduces the effects of mass loss by evaporation. One useful technique for this is measuring liquid masses by taking appropriate differences. (See the individual experiment instructions.)

Experiments

This document includes three important introductory experiments:

- **Experiment 1: What is a Calorie?**

  This lab provides an introduction to the ideas of temperature and heat, and a demonstration of the conservation of energy.

- **Experiment 2: Thermal Capacity and Specific Heat**

  In this lab, you will measure the specific heats of the aluminum, brass, copper, stainless steel, and zinc samples, as well as antifreeze.

- **Experiment 3: Latent Heat of Fusion**

  In this lab, you will investigate the role of heat transfer in the conversion of ice into water.

Experiment 1: What is a Calorie?

Required equipment

- Calorimeter cup
- Alcohol thermometer
- Triple beam balance
- Hot and cold water
- Insulated cup (about 450 mL)

Introduction

When two systems or objects of different temperature come into contact, energy is transferred from the warmer system to the cooler system in the form of heat. This transfer of heat raises the temperature of the cooler system and lowers the temperature of the warmer system. Eventually, the two systems reach a common intermediate temperature, at which point the heat transfer stops.

The standard unit for measuring heat transfer is the *calorie*. A calorie is defined as the amount of energy required to raise the temperature of one gram of water from 14.5 °C to 15.5 °C. However, for our purposes, we can generalize this definition to say that a calorie is the amount of energy required to raise the temperature of one gram of water by one degree Celsius, as the variation with temperature is slight.

In this experiment, you will combine samples of hot and cold water of known temperature and mass. Using the definition of the calorie, you can determine the amount of heat energy that is transferred in the process of bringing the hot and cold water to a common temperature, and thereby determine whether heat energy is conserved in this process.

Procedure

1. Determine M_{cal}, the mass of the empty calorimeter, and M_{cup}, the mass of the empty insulated cup. Record both values in Table 1.1 under **Trial 1**.
2. Fill the calorimeter cup approximately 1/3 full of cold water. Measure the mass of the calorimeter and water together to determine $M_{\text{cal + cold water}}$. Record this value in Table 1.1 under **Trial 1**.
3. Fill the insulated cup approximately 1/3 full of hot water. The water should be at least 20 °C above room temperature. Weigh the cup and water together to determine $M_{\text{cup + hot water}}$. Record this value in Table 1.1 under **Trial 1**.
4. Measure T_{hot} and T_{cold}, the temperatures of the hot and cold water respectively. Record these values in Table 1.1 under **Trial 1**.
5. Immediately after measuring the temperatures, add the hot water to the cold water. Stir with the thermometer until the temperature stabilizes. Record T_{final}, the final temperature of the mixture, in Table 1.1 under **Trial 1**.
6. Measure M_{final}, the final mass of the calorimeter and mixed water. Record this mass in Table 1.1 under **Trial 1**.
7. Repeat Steps 1 through 6 twice more with different masses of water at different temperatures. Record the results of these trials in Table 1.1 under **Trial 2** and **Trial 3** respectively.

> **TIP:** Try adding cold water to hot instead of hot water to cold to see if any variation occurs.

Data

Measurement	Trial 1	Trial 2	Trial 3
M_{cal}			
M_{cup}			
$M_{\text{cal + cold water}}$			
$M_{\text{cup + hot water}}$			
T_{cold}			
T_{hot}			
T_{final}			
M_{final}			

Calculations

1. Use your data in Table 1.1 to calculate $M_{\text{cold water}}$, the original mass of the cold water, and $M_{\text{hot water}}$, the original mass of the hot water, by finding the difference between each empty container's mass and the mass of that container with water. Record your data in Table 1.2 for each trial.
2. For each trial, calculate the change in temperature experienced by the hot water and the cold water when mixed by taking the difference between T_{final} and the initial temperatures. Record these values in Table 1.2 as ΔT_{hot} and ΔT_{cold} respectively.
3. Use the equations below to calculate ΔQ_{cold} and ΔQ_{hot}, the heat gained by the cold and hot water respectively. Enter your results in Table 1.2.

$$\Delta Q_{\text{cold}} = (M_{\text{cold water}})(\Delta T_{\text{cold}})(1\ \text{cal/(g K)})$$

$$\Delta Q_{\text{hot}} = (M_{\text{hot water}})(\Delta T_{\text{hot}})(1\ \text{cal/(g K)})$$

Measurement	Trial 1	Trial 2	Trial 3
$M_{\text{cold water}}$			
$M_{\text{hot water}}$			
ΔT_{cold}			
ΔT_{hot}			
ΔQ_{cold}			
ΔQ_{hot}			

Questions

1. Which had more thermal energy: the two cups of water before they were mixed together, or the calorimeter cup after the water samples were mixed? Was energy conserved during the process?
2. Discuss any unwanted sources of heat loss or gain that might have affected the experiment.
3. If 200 g of water at 85 °C were added to 150 g of water at 15 °C, what would the final equilibrium temperature of the mixture be?

Experiment 2: Specific Heat

Required equipment

- Calorimeter cup
- Alcohol thermometer
- Samples (aluminum, brass, copper, stainless steel, zinc)
- Triple beam balance
- Antifreeze (~100 g)
- Boiling water and cool water
- Thread

Introduction

The *specific heat* of a substance, usually represented by the symbol c, is the amount of heat required to raise the temperature of one gram of the substance by 1 °C (or 1 K). From the definition of the calorie given in Experiment 1, we can see that the specific heat of water is 1.0 cal/(g K). If an object is made of a substance with specific heat equal to c_{sub}, then the heat ΔQ required to raise that object's temperature by an amount ΔT is given by:

$$\Delta Q = (\text{mass of object})(c_{\text{sub}})(\Delta T)$$

In Part 1 of this experiment, you will measure the specific heats of aluminum, brass, copper, stainless steel, and zinc. In Part 2, you will measure the specific heat of antifreeze.

⚠ **CAUTION:** This experiment involves the use of boiling water and the handling of hot metal objects. Work carefully to avoid burns!

Part 1: The specific heats of five metals

1. Measure M_{cal}, the mass of the calorimeter cup. Ensure the calorimeter is empty and dry. Record your result in Table 2.1.
2. Measure the masses of the aluminum, brass, copper, stainless steel, and zinc samples. Record these masses in Table 2.1 in the row labeled M_{sample}.
3. Attach a thread to each of the metal samples and suspend each of the samples in boiling water. Allow a few minutes for the samples to heat thoroughly.
4. Fill the calorimeter cup approximately half full of cool water. Use enough water to fully cover any of the samples.
5. Select **one** of the metal samples to be used for the next five steps.
6. Measure T_{cool}, the temperature of the cool water. Record this value in Table 2.1 in the column for your chosen sample.
7. Immediately following your temperature measurement, remove your chosen sample from the boiling water, quickly wipe it dry, and then suspend it in the cool water in the calorimeter. The sample should be completely covered but should not touch the bottom of the calorimeter.
8. Swirl the water, then measure T_{final}, the highest temperature attained by the water as it comes into thermal equilibrium with the metal sample. Record this value in Table 2.1 in the column for your chosen sample.
9. Immediately after measuring T_{final}, measure and record M_{total}, the total mass of the calorimeter, water, and metal sample together.
10. Remove the sample from the calorimeter cup, then empty and dry the calorimeter.
11. Repeat Steps 4 through 10 for the other metal samples.

Part 2: The specific heat of antifreeze

Repeat Part 1 of this experiment, but instead of using the metal samples, use 100 g of antifreeze. Begin by heating the antifreeze to about 60 °C. Measure and record the temperature, then quickly pour the antifreeze into the calorimeter cup containing cool water. Stir until the highest stable temperature is reached (which should take about 1 minute) and measure the temperature. Record the following data in the provided blanks:

- M_{cal}, the mass of the calorimeter cup when empty
- $M_{\text{cal + water}}$, the mass of the calorimeter cup plus cool water
- T_{cool}, the temperature of the cool water
- $T_{\text{antifreeze}}$, the temperature of the heated antifreeze immediately before mixing
- M_{total}, the mass of the calorimeter, water, and antifreeze
- T_{final}, the temperature of the water plus antifreeze

Data and calculations

Part 1:

1. For each metal tested, use the equations below to determine M_{water}, the mass of the water used; ΔT_{water}, the temperature change of the water when it came into contact with the metal sample; and ΔT_{sample}, the temperature change of the metal sample when it came into contact with the water. Record your results in Table 2.1.

$$M_{water} = M_{total} - (M_{cal} + M_{sample})$$
$$\Delta T_{water} = T_{final} - T_{cool}$$
$$\Delta T_{sample} = 100° - T_{final}$$

2. From the law of energy conservation, the heat lost by the metal sample must be equal to the heat gained by the water. In other words:

$$\text{Heat lost by sample} = (M_{sample})(c_{sample})(\Delta T_{sample}) = (M_{water})(c_{water})(\Delta T_{water}) = \text{Heat gained by water}$$

where c_{water} is the specific heat of water, which is 1.0 J/(g K). Use the above equation and your collected data to calculate the specific heats of aluminum, brass, copper, stainless steel, and zinc. Record your results in Table 2.1.

Measurement	Aluminum	Brass	Copper	Stainless steel	Zinc
M_{cal}					
M_{sample}					
T_{cool}					
T_{final}					
M_{total}					
M_{water}					
ΔT_{water}					
ΔT_{sample}					
c					

Part 2:

M_{cal} = _________________

$M_{cal + water}$ = _________________

T_{cool} = _________________

$T_{antifreeze}$ = _________________

M_{total} = _________________

T_{final} = _________________

Perform similar calculations to those you performed in Part 1 to determine $c_{antifreeze}$, the specific heat of antifreeze. Note that, in this case, you will use $T_{antifreeze}$ for the initial temperature rather than assuming it to be 100 °C.

$c_{antifreeze}$ = _________________

Questions

1. How do the specific heats of the samples compare with the specific heat of water?
2. Discuss any unwanted heat loss or gain that might have affected your results.
3. From your measured specific heat for antifreeze, which should be the better coolant for an automobile engine, antifreeze or water? Why is antifreeze used as an engine coolant?

Experiment 3: Latent Heat of Fusion

Required equipment

- Calorimeter cup
- Alcohol thermometer
- Warm water
- Small chunks of ice

Introduction

When a substance changes phases, the arrangement of its molecules changes. If the new arrangement has a higher internal energy, the substance must absorb heat in order to make the phase transition. Conversely, if the new arrangement has a *lower* internal energy, heat will be released as the transition occurs. For example, water has a higher internal energy content than ice. It takes a certain amount of energy for the water molecules to break free of the forces that hold them together in the crystalline formation of ice. This same energy is released when the water molecules come together and bond to form an ice crystal.

In this experiment, you will measure the difference in internal energy between one gram of ice at 0 °C and one gram of water at 0 °C. This difference in energy is referred to as the *latent heat of fusion* of water.

Procedure

1. Measure T_{rm}, the room temperature.
2. Determine M_{cal}, the mass of the empty, dry calorimeter cup.
3. Fill the calorimeter cup approximately half full of warm water, about 15 °C above room temperature.
4. Measure $M_{cal + water}$, the combined mass of the calorimeter and the warm water.
5. Measure $T_{initial}$, the initial temperature of the warm water.
6. Add small chunks of ice to the warm water, wiping the excess water from each piece of ice immediately before adding it. Add the ice slowly, stirring continuously with the thermometer until each chunk melts.
7. When the temperature of the mixture is as far below room temperature as the warm water was initially above room temperature and all the ice has melted, measure T_{final}, the final temperature of the water.
8. Immediately after measuring T_{final}, weight the calorimeter and water to determine M_{final}, the combined mass of the calorimeter and mixed water.
9. *Optional:* Repeat Steps 1 through 8, but instead of ordinary ice, use the material which is packaged in metal or plastic containers to be frozen and used in picnic coolers.

Data

$$T_{rm} = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_$$

$$M_{cal} = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_$$

$$M_{cal + water} = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_$$

$$T_{initial} = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_$$

$$T_{final} = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_$$

$$M_{final} = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_$$

Calculations

1. Calculate the mass of the ice added to the cup using the following equation:

$$M_{ice} = M_{final} - M_{cal + water} = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_$$

2. According to conservation of energy, the quantity of heat absorbed by the ice as it melts and then heats up to T_{final} must be equal to the quantity of heat released by the warm water as it cools down to T_{final}. Mathematically, this can be expressed as:

$$(M_{ice})(L_f) + (M_{ice})(1\ \text{cal/(g K)})(T_{final} - 0\ °\text{C}) = (M_{water})(1\ \text{cal/(g K)})(T_{initial} - T_{final})$$

Use your data and the above equation to calculate L_f, the latent heat of fusion per gram of water.

Questions

1. What advantage might the commercially packaged coolant material have over ice, other than that it produces less mess? (If you did not perform the optional step, what properties would a material need in order to be a better coolant than ice?)
2. Design an experiment to determine which of two substances (for example, ice and packaged coolant) will keep an insulated food cooler:
    a. cool for the longest time
    b. at a lower temperature

Technical support

Need more help? Our knowledgeable and friendly Technical Support staff is ready to answer your questions or walk you through any issues.

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For a description of the product warranty, see the Warranty and Returns page at www.pasco.com/legal.