Instructions
Activities
Experiments

Instruction Manual

No. 012-08808A

Energy Transfer Calorimeter

ET-8499

PASCO®

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Table of Contents

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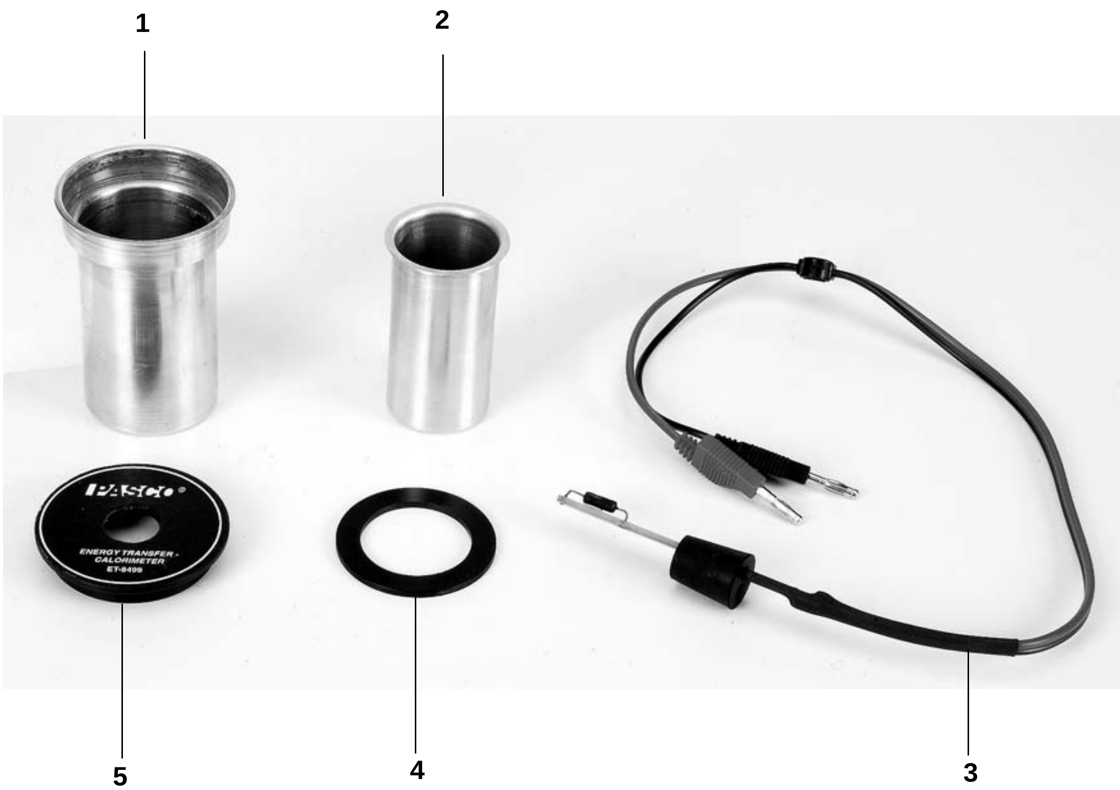
Model No. ET-8499



Included Equipment	Replacement Part Number
1. Outer cup	648-08748
2. Inner cup	648-08747
3. Heating resistor assembly	003-08803
4. Insulating spacer	648-08750
5. Lid	648-08749
6. One-hole stopper (not shown)	699-129
7. Experiment setup CD (not shown)	013-08983

Introduction

The Energy Transfer Calorimeter incorporates the classic design of two nested aluminum cups with an air space in between for insulation. Instead of the traditional coil, a 10 ohm, circuit-board-mounted heating resistor is provided for heating water. However, the unit can also be used without the heating resistor. The size of the calorimeter cup has been scaled to work well with any 1 amp, 10 volt power supply. Typical applications include the calorimetry of mixtures, the heat of fusion of ice, and the electrical equivalent of heat.

Component specifications

- The outer aluminum cup is a holder and acts as an insulator. Its mass does not need to be included in any calculations.
- The inner calorimeter cup is made of aluminum, and its mass should be included in heat transfer equations.
- The heating resistor is permanently mounted in a two-hole stopper, where the second hole provides access for a temperature probe. The cable end plugs directly into a power supply with banana terminals.
- The insulating spacer aligns the inner and outer cup to leave a small air space in between that acts as an insulator. The spacer rests on the step inside the outer cup.
- The insulating lid features a hole for a rubber stopper. A #2, one-hole stopper is included that can be used to hold the temperature probe when the heating resistor is not used.

Activities

The two copy-ready labs included with this manual cover the topics of calorimetry of mixtures, latent heats of transformation, and entropy. The experimental setups do not specify the type of thermometer used. This allows the instructor the flexibility to utilize any available equipment.

Equipment Setup

Important!

- Never apply power to the heating resistor unless the resistor is immersed in water.
- Never touch the resistor. It gets hot!
- Do not apply more than 10 volts.

1. Place the spacer inside the larger, outer cup so it is seated on the lip.

2. Place the inner cup into the outer cup through the hole in the spacer.

3. Place the lid on top.

4. Feed the heating resistor assembly through the hole in the lid. Fit the stopper snugly into the hole to insulate the system and to hold a temperature probe. **Note:** If the heating resistor is not being used, insert the one-hole stopper to insulate the system and to hold a temperature probe.

Activity 1:
Electrical Equivalent of Heat

Equipment Required
Calorimeter with heating resistor
Hand-cranked generator
Water

This is the classic experiment where electrical energy (provided by an external power supply) is converted to thermal energy in a container of water. This experiment is described in detail in the electronic workbook "Electrical Equivalent of Heat" included on the experiment setup CD. Students use a hand cranked generator and the calorimeter with heating resistor, to raise the temperature of the water in the cup by several degrees C. The electrical energy added to the system is compared to the resulting increase in thermal energy of the water and the inner calorimeter cup.

The workbook is written utilizing computer probeware, but a traditional approach is also possible. If a power supply is used, the current and voltage change so little during the experiment that a standard multimeter will suffice to determine the input energy, and a glass thermometer can easily be used to measure the change in temperature.

Activity 2: Calorimetry

This experiment teaches the concept of conservation of energy: In an isolated system, the sum of the heat that flows out of the objects that cool down, must equal the sum of the heat that flows into the objects that heat up. Following is a list of required equpment for this activity.

Equipment Required
Calorimeter
Steam generator or Hot plate
Small object of known material (approx 10 cc)
Thermometer or Temperature sensor (0.1° resolution)
String

Procedure

1. Prepare hot water for heating the object (70 - 100°C).
2. Measure the mass of the cup and the object.
3. Pour cool water, a few degrees below room temperature, into the inner cup.
4. Measure the combined mass of the cup and cool water to determine the net mass of the water.
5. Tie the object to a string and heat the object in a hot water bath. Allow enough time for the object to come to equilibrium with the water.
6. Measure the temperatures of the hot and the cold water.
7. Transfer the object to the cup.
8. Swirl the cup to mix, and record final temperature of the water.

Compare the heat lost by the object to the heat gained by the cup and water to investigate conservation of energy. Or work the problem backwards to find the specific heat of an "unknown" sample.

Experiment 1: Heat of Fusion of Ice

Equipment Required
Calorimeter
Thermometer or Temperature sensor (0.1 Degree resolution)
Ice cubes (solid, not crushed)
Scale

Purpose

This lab investigates the concept of conservation of energy as applied to calorimetry: In a closed system, the sum of all materials that gain (absorb) energy must equal the sum of all materials that lose energy. The relationship between temperature and heat is discussed, including the topics of specific heat and latent heat of transformation.

Theory

Heat added to (or removed from) a solid or a liquid can change the material's internal thermal energy and thus change its temperature. The relationship between the heat flow and the resulting change in temperature is given by

$$\mathbf{Q = mc\Delta T}$$

Q = Heat

m = mass

c = specific heat

ΔT = change in Temperature

In addition, when a material is at a boundary state between phases, heat flow can cause a change in phase. For example, in this experiment heat added to ice at 0°C causes a change in phase from solid to liquid. The relationship between the heat flow and the resulting change in phase is given by

$$\mathbf{Q = mL_f}$$

Q = Heat

m = mass

L_f = (Latent) Heat of Fusion

In this experiment, the water and the inner aluminum Calorimeter cup lose energy and cool down, while the ice gains energy first in melting, and then in warming that melted ice-water up to the final equilibrium temperature of the system. The water and cup undergo the same temperature change, but the melted ice-water undergoes a different temperature change. It is assumed that the initial temperature of the ice is 0°C.

Setup

This experiment requires a source of ice that is not crushed, preferrably small chunks of about 5 grams. Some care must be taken with the ice. Ice that is taken from the freezer will be much colder than 0°C and cannot be used immediately straight out of the freezer. You must wait until the ice has warmed up to 0°C and is melting. On the other hand, if the ice sits around the lab for a hour in an icy water bath, there will be pockets of water within the ice that have already melted. The ideal is to take a chunk of ice out of the freezer, place it on napkin, and wait until it is melting.

Procedure

1. Measure the ambient temperature of the room.
2. Measure the mass of the **inner** aluminum cup from the Calorimeter. Only the inner cup changes temperature and is part of the experiment. The outer (bigger) cup acts only as a holder, and due to the air gap in between, helps to insulate the inner cup.
3. Prepare some water that is about 8°C above room temperature. Add about 40 g of this water to the inner cup. Measure the combined mass of the cup and water to determine the exact mass of the water. After the cup and water have come to equilibrium and you are ready to start the experiment, you will want the temperature to be at least 5°C above room temperature.
4. Measure the mass of the cup plus water, and calculate the mass of the water.
5. Assemble the Calorimeter, using the spacer to suspend the inner cup inside the bigger cup. Put on the lid, and use the one-hole stopper in the lid to hold the thermometer.
6. Prepare about 5 grams of ice. It must be a solid chunk, not crushed. If you take the ice cube directly from the freezer, you must wait a few minutes for it to warm up and start to melt. You do not need to know the exact mass now, but it should be about 5 grams.
7. Check the temperature of the water again. If the water is not about 5°C above room temperature, you can place the cup and water in a water bath to change the temperature slightly. Wipe off any moisture on the cup before placing it back into the outer cup. Gently swirl the cup to stir the water and wait for equilibrium before adding the ice.
8. Record the initial temperature of the water and cup to a resolution of at least 0.1°C

9. Dry off the ice with a paper towel, and then place it in the Calorimeter. Replace the lid and watch the temperature of the water. Gently swirl the Calorimeter every 10 seconds to ensure an even temperature.

10. Continue to gently swirl the Calorimeter every 10 seconds as you watch the temperature decrease, until it reaches its lowest value. Lift the lid and look inside to see that all the ice has melted. If not, replace the lid and continue. Record the lowest temperature.

11. Remove the inner cup and determine the mass of cup and water. Use this to calculate the mass of ice you added.

Analysis

1. Calculate the change in temperature, ΔT, of the cup and water.

2. Calculate the change in temperature of the melted ice water.

3. Calculate the total amount of heat, Q_{lost}, lost from the cup and water. Use the proper specific heat for each.

4. Calculate the total amount of heat, Q_{gained}, added to the ice. You should have two terms.

5. What percentage of the energy supplied by the water and cup is delivered to the ice? Calculate the percent difference.

$$\%\ \text{difference} = \frac{Q_{lost} - Q_{gained}}{Q_{lost}} \text{ X } 100(\%)$$

6. Did the <u>system</u> "lose" energy or "gain" energy. Explain your results using the concept of conservation of energy.

Questions

1. Why did we go to the trouble to start with water in the cup that was above room temperature? What does that accomplish?

2. Why does Q for the aluminum cup include only the inner cup? Why do you not include the spacer? Why is the spacer not made of aluminum?

3. Why does the gap between the inner and outer cup provide insulation? In theory, what would ideally be in this gap?

4. Why did you wipe the water off the ice to dry it before adding it to the cup.

Further Study

Use "cold" ice directly out of the freezer. Assume conservation of energy, and work the equations backwards to find the initial temperature of the ice. **Hint:** The specific heat of ice is not the same as that of water. If you have a non-glass thermometer, you can freeze the temperature probe in the ice, and watch it warm up when it is added to the cup.

Experiment 2: Entropy

Equipment Required
Calorimeter
(2) Thermometer or Temperature sensor (0.1° resolution)
Water (hot and cold)
Scale

Purpose

This lab investigates the concept of entropy as applied to calorimetry. The relationship between entropy and spontaneous processes is investigated.

Theory

In a closed system, heat energy flows spontaneously from the hot objects to the cooler objects and the total change in energy for the system is zero. The heat added to one part of the system equals the heat lost by the other parts. However, the total change in **entropy** of the system is not zero. For any real, spontaneous process, the total change in entropy of the entire system must be greater than zero.

The change in entropy (ΔS) of an object is defined in terms of the energy transferred (dQ) and the temperature of the object (T).

$$\Delta S = S_f - S_i = \int_i^f \frac{dQ}{T} \qquad (1)$$

For a calorimetry experiment, $dQ = mc\,dT$, where

m = mass

c = specific heat

Thus, equation (1) can be reduced to

$$\Delta S = mc\int_i^f \frac{dT}{T} \qquad (2)$$

Completing the integral yields the final relationship

$$\Delta S = mc\ln\left(\frac{T_f}{T_i}\right) \quad (3)$$

Note that entropy is a state function. The change in entropy depends only on the initial and final conditions, not on the process involved.

In this experiment, hot water is mixed with cold water in an aluminum calorimeter cup. The cold water and aluminum cup have the same increase in temperature and can be considered as one part of the system. The hot water which is added to the cup is the other part of the system. The change in entropy of each part of the system is calculated using equation (3).

The sign of ΔS will not be positive for both parts of the system. In general, a positive change in entropy indicates a process which tends towards a higher state of disorder. As you can see from equation (1), if dQ is positive, the change in entropy is positive, and that means (for this experiment) that the temperature is increasing. An object at a higher temperature has faster-moving molecules, and in general is in a more "disordered" state.

Procedure

1. Measure the mass of the inner aluminum cup from the Calorimeter. Only the inner cup changes temperature and is part of the experiment. The outer (bigger) cup acts only as a holder, and due to the air gap in between, helps to insulate the inner cup.

2. Add about 30 g of "ice" water to the inner cup. You want very cold water, but make sure there is no ice in the cup.

3. Measure the mass of the cup plus water and calculate the mass of the cold water.

4. Assemble the cup, using the spacer to suspend the inner cup inside the bigger cup. Put on the lid, and use the one hole stopper in the hole to hold the thermometer.

5. Pour about 100 g of hot water into an insulated cup. Place a second thermometer in the hot water.

6. Gently stir the water, and record the initial temperatures of the hot and cold water to a resolution of at least 0.1°C. Pour about 30 g of hot water into the calorimeter cup. You want at least 1 cm above the water line to the top of the cup. Replace the lid and watch the temperature of the water. Gently swirl the water and record the final equilibrium temperature.

7. Place the other thermometer (from the hot water) in the same water (in the Calorimeter cup), and check to see if it reads the same value. If not, measure the offset and adjust your values accordingly.

8. Remove the inner cup and determine the mass of cup and water. Use this to calculate the mass of hot water you added

Analysis Part A: Energy

1. Calculate the change in temperature, ΔT, of the cup and cold water.
2. Calculate the change in temperature of the hot water.
3. Calculate the total amount of heat, Q_{gained}, added to the cup and cold water. Use the proper specific heat for each.
4. Calculate the total amount of heat, Q_{lost}, given up by the hot water.
5. What percentage of the energy supplied by the hot water is delivered to the cold water and the cup? Calculate the percent difference.

$$\% \text{ difference } = \frac{Q_{lost} - Q_{gained}}{Q_{lost}} \text{ X } 100(\%)$$

6. If you were very careful, this percentage should be down close to 1 or 2 percent. If not, check for mistakes. If your percentage is over 5 percent, repeat the experiment before going on to part B.

Analysys Part B: Entropy

1. Use equation (3) to calculate the change in entropy of the cup and cold water. Remember to use Kelvin for the temperatures. Pay careful attention to which is the initial and which is the final temperature: It affects the sign of your answer.
2. Use equation (3) to calculate the change in entropy of the hot water. Remember to use Kelvin for the temperatures. Pay careful attention to which is the initial and which is the final temperature: It affects the sign of your answer.
3. Calculate the total change in entropy of the system.

Questions

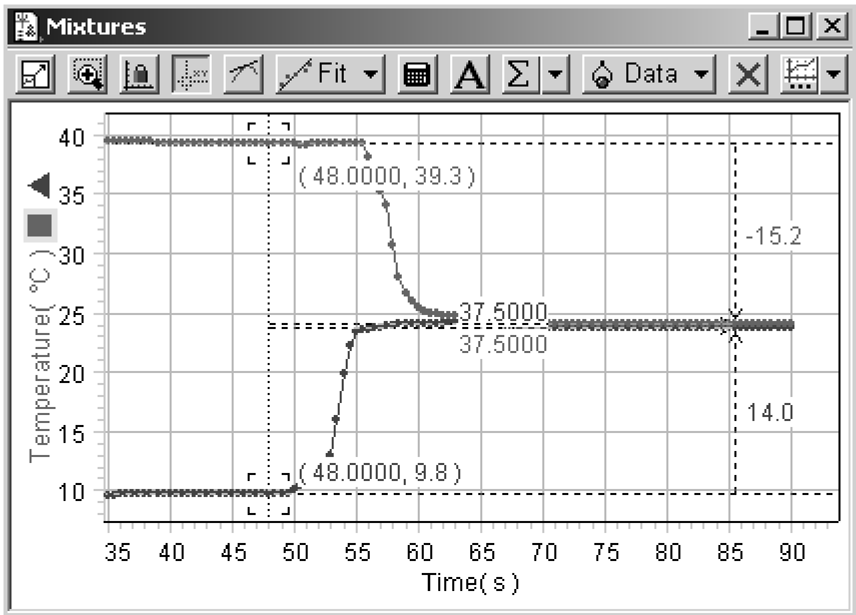
1. Evaluate the sign of the change in entropy for both parts of the system. Explain in terms of "disorder" and the definition of entropy.
2. Evaluate the sign of the change in entropy for the entire system. Explain in terms of the definition of entropy, and the concept of spontaneous processes.
3. When calculating heat flow (as in part A) it is not necessary to change the temperatures to Kelvin, but when calculating entropy (as in part B) it is necessary. Explain why this is true.

Further Study

Repeat this experiment by adding ice to hot water in the calorimeter cup. **Hint:** Ice melts at a constant temperature of 0°C, and thus equation (3) does not apply for the melting part. When ice melts, is its change in entropy positive or negative?

Teacher's Notes

Results for Exp 2



Since these two temp probes were not calibrated together first, I need to account for the offset.

39.3 - 15.2 =24.1

9.8 + 14.0 = 23.8

This results in an offset of 0.3 which I will add to the lower number.

Cup and cold water: T_o = 9.8 +.3 = 10.1°C

T_f = 24.1°C

ΔT=14.0°C

Hot water: T_o = 39.3°C

T_f = 24.1°C

ΔT=15.2°C

M_{cup} = 22.0 g

M_{CW} = 30.2 g

M_{HW} = 1.9 g

Part A:

Q_{gained} = ((30.2)(4.186)+(22)(.90))(14) = 2.05 kJ

Q_{Lost} = (31.9)(4.186)(15.2) = 2.03 kJ

Difference of 1 %, is acceptable, therefore it is OK to proceed to part B.

Part B:

ΔS= (31.9)(4.186) ln (297.1/283.1) = + 7.06 J/K

This is the value for the cup and cold water. The change in entropy is positive for an increase in temperature. The molecules are moving faster, moving around more, and occupying more space. This is an increase in "disorder".

ΔS= ((30.2)(4.186) + (22)(.9)) ln (297.1/312.3) = - 6.66 J/K

This is the value for the hot water. The change in entropy is negative for a decrease in temperature and is a decrease in "disorder".

Total ΔS= +7.06 + (-6.66) = +0.40 J/K

Note that the sign is positive. For any real, spontaneous process, the total change in entropy of the entire system must be greater than zero.

Safety

Read the instructions before using this product. Students should be supervised by their instructors. When using this product, follow the instructions in this manual and all local safety guidelines that apply to you.

Technical Support

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