ENTROPY

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

Materials and Equipment

- Calorimeter (ET-8499)
- Temperature Sensor or Thermometer (2)
- Balance
- Cold Water (~0 °C)
- Hot Water

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.

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

For a calorimetry experiment, $dQ = mc\,dT$, where m = mass and c = specific heat. Thus, equation 1 can be reduced to

$$\Delta S = mc\int_i^f \frac{dT}{T}$$

Completing the integral yields the final relationship

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

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.

Safety

Follow these important safety precautions in addition to your regular classroom procedures:

- Wear safety goggles at all times.
- Use caution with hot water.

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.
2. Add about 30 g of cold water to the inner cup. Make sure there is no ice in the cup.
3. Measure the mass of the cup plus water. 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 30 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

Data Collection

Table 1. Mass Measurements

Cup	
Cup and Cold Water	
Cold Water	
Cup and Mixed Water	
Hot Water	

Table 2. Temperature Measurements

Cold Water	
Hot Water	
Mixed Water	

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.

$$\% \text{ difference} = \frac{Q_{\text{lost}} - Q_{\text{gained}}}{Q_{\text{lost}}} \times 100$$

6. 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.

7. 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

8. Calculate the total change in entropy of the system.

Questions and Analysis

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 it is not necessary to change the temperatures to Kelvin, but when calculating entropy it is necessary. Explain why this is true.