Heat of Fusion of Ice

In an earlier experiment (Experiment 2B) you learned that temperature stays constant during a phase change of a substance while the heat is either absorbed or released during the phase change. In this experiment you will determine the actual amount of heat required for a phase change of a substance, namely the melting of ice. The experiment will be done in a simple calorimeter, an apparatus that contains the material undergoing the temperature change and enables the temperature to be recorded easily.

The *temperature* of a substance is a measure of the average amount of kinetic energy per particle of the substance. The higher the temperature, the higher the average kinetic energy, and the faster the molecules are moving. *Heat,* on the other hand, is a measure of the *total* amount of energy contained by a substance and depends not only on the temperature, but what the substance is and how much of it you have. The quantity of heat (q) required to raise the temperature of a substance in one phase is given by the equation:

$$q = m \times c \times \Delta T$$

where q = quantity of heat (kJ)
m = mass of substance (kg)
c = specific heat capacity (kJ/kg•°C)
ΔT = temperature change (°C)

Water has one of the highest specific heat capacities of any substance, at 4.18 kJ/kg•°C. Iron, for example, has a value approximately one tenth of that of water, at 0.450 kJ/kg•°C. In other words, water requires a relatively high quantity of heat to change its temperature. Another way of looking at this concept is to consider that if the same amount of heat is supplied to equal amounts of two different materials, the one with the lower specific heat capacity undergoes a much larger increase in temperature. This explains why a very rapid temperature rise occurs if an iron saucepan is inadvertently left on a heating element without any water in it. Furthermore, the large specific heat capacity of water is the reason why climate is much less extreme in locations near an ocean or another large body of water such as one of the Great Lakes. Summers are not as hot and winters are not as cold, compared to locations well away from a large body of water. Similarly, the large amount of water in our bodies also helps us to withstand temperature extremes.

When a solid is heated, the molecules eventually possess enough energy to overcome the intermolecular forces of attraction in the solid and break free from one another. This is the process of *melting* or *fusion*, and the temperature at which it occurs is the *melting point* of the substance. The substance remains at this temperature until all of it has melted, and the energy required for this change of state is called the heat of fusion, and is usually quoted with the units of kJ/kg.

In this experiment you will determine the heat of fusion of ice. Ice cubes at 0°C will be placed in warm water at a measured temperature and the final temperature will be recorded when the system has come to thermal (temperature) equilibrium. The heat lost by the warm water in cooling to the final temperature will be equal to the heat required to melt the ice and then raise the temperature of the resulting water to the final temperature. The calorimeter you will use is simply comprised of a stirring rod and thermometer in a covered styrofoam cup. Styrofoam has a very low specific heat capacity and, consequently, is an excellent heat insulator which minimizes heat loss to the surroundings and heat absorption by the cup itself.

OBJECTIVES

1. to calculate the energy absorbed by a mass of ice as it melts
2. to calculate the heat of fusion of ice in kJ/kg

SUPPLIES

Equipment

centigram or digital balance
styrofoam cup (large, approx. 300 mL)
lid for styrofoam cup (with two holes, one in middle, one close to edge)
400 mL beaker
source of warm water (kettle, or hot water tap)
thermometer
stirring rod
lab apron
safety goggles

Chemical Reagents

ice cubes

PROCEDURE

CAUTION

Your thermometer is made of glass and breaks easily, leaving sharp edges that cut. Handle your thermometer gently. Do not use it to crush or stir ice. If your thermometer breaks, call your instructor. If it contains mercury, be aware that mercury liquid and vapor are very poisonous.

1. Put on your lab apron and safety goggles.
2. Determine the mass of a clean, dry styrofoam cup and lid and enter it and all future readings in your copy of Table 1 in your notebook.
3. Using the method given by your instructor, obtain about 200 mL of warm water in your beaker. Adjust the temperature to about 40°C with cold water if necessary.
4. Pour the water into your styrofoam cup, add the lid, and record the mass of the cup, lid, and water.
5. Obtain two medium-sized ice cubes and dry them with a paper towel.
6. Record the temperature of the water in the cup to the nearest 0.1°C.
7. Quickly remove the lid from the cup, add the two ice cubes, and replace the lid. Insert a thermometer in the side hole, letting it go to the bottom of the cup, and insert a stirring rod in the center hole. Stir the cup's contents slowly and constantly with the stirring rod and measure the temperature as soon as the ice has melted. (You will have to peek under the lid occasionally.) Continue to monitor the temperature for another minute, to make sure the contents reach a constant temperature.

8. Remove the thermometer and stirring rod. Record the new mass of the cup, lid, and contents, in order to be able to determine the mass of ice added.

9. Empty your containers of water and wash your hands thoroughly with soap and water before leaving the laboratory.

REAGENT DISPOSAL

No chemicals other than water are used in this experiment.

POST LAB CONSIDERATIONS

In this experiment, the heat lost by the warm water equals the heat gained by the ice. The heat gained by the ice is used in two ways: first the ice melts and then the water produced by the melting of the ice has its temperature raised from 0°C to the final temperature of the mixture. In other words,

$$q_{(temp\ loss)} = q_{(melting)} + q_{(temp\ gain)}$$

The measured masses in grams must be converted to kilograms in order to calculate the heat values. Remember that the specific heat capacity of water is 4.18 kJ/kg•°C

EXPERIMENTAL RESULTS

Table 1

Mass of empty cup and lid	g	kg
Mass of cup, lid, and warm water	g	kg
Mass of warm water	g	kg
Temperature of warm water, T_{warm}(°C)		
Final temperature after melting T_{final}(°C)	COMPLETE IN YOUR NOTEBOOK	COMPLETE IN YOUR NOTEBOOK
ΔT(loss) = $T_{warm} - T_{final}$ (°C)		
ΔT(gain) = $T_{final} - 0°C = T_{final}$ (°C)		
Final mass of cup, lid, and contents	g	kg
Mass of ice added	g	kg

ANALYSIS OF RESULTS

If necessary, consult the Introduction and Post Lab Considerations for formulas (relationships) required.

1. Calculate the quantity of heat lost, $q_{(temp\ loss)}$, by the warm water solution, in kilojoules.
2. Calculate the quantity of heat absorbed, $q_{(temp\ gain)}$, by the melted ice, in going from 0°C to the final equilibrium temperature, in kilojoules.
3. Calculate the amount of heat, $q_{(melting)}$, that must have been absorbed in the melting process.
4. Calculate the heat of fusion for ice, H_{fus}, in kJ/kg.

FOLLOW-UP QUESTIONS

1. The accepted value for the heat of fusion of ice, H_{fus}, is 334 kJ/kg. By wha percentage does your value vary from the accepted value?
2. Suggest any possible assumptions or sources of error that could accour for any difference observed for your value of H_{fus} and whether the would make your answer larger or smaller than the accepted value.
3. Heat is given off in the process in which a liquid solidifies to form a soli The heat of solidification has the same numerical value as the heat fusion but is opposite in sign. When water is frozen into ice cubes in refrigerator, where does the released heat go?

CONCLUSION

State the result of Objective 2.