Name **Section #**

Latent Heat of Fusion, *Hf*

Introduction

Ice and water are at equilibrium at 0°C, yet 334 kilojoules of energy are required to melt 1 kilogram of ice at 0°C and change it into water at 0°C. (It is for this reason that ice is so much more effective than an equivalent amount of ice–cold water in cooling a drink.) The energy per kilogram required to completely melt a substance, already at its melting temperature, is called the latent heat of fusion. In the case for ice, $H_f = 334$ kJ/kg.

You will add a mass m of ice at 0°C to a mass M of warm water at about 20°C. By measuring ΔT_m and ΔT_M for the ice and the water, respectively, you will be able to determine the heat gained by the ice and by the warm water. (The heat gain by the warm water will be negative since it actually cools down.) Using the relationships discussed in the theory section below, you will be able to obtain an experimental value for H_f and compare to the standard value.

Theory

ΔQ_f is the heat gained by the ice in melting into water at 0°C,

$\Delta Q_f = m(H_f)$;

ΔQ_m represent the heat gained by the ice-cold water in warming up to the equilibrium temperature T,

$\Delta Q_m = c_w m \Delta T_m$, $c_w = 4.186$ kJ/kg, and $\Delta T_m = (T - 0^\circ \text{C})$;

ΔQ_M is the heat change of the warm water cooling down from the original temperature T_i to the equilibrium temperature T,

$\Delta Q_M = c_w M \Delta T_M$, *where* $\Delta T_M = (T - T_i)$ *and is negative since* $T < T_i$.

By conservation of energy all these changes should total to zero,

$\Delta Q_f + \Delta Q_m + \Delta Q_M = 0$ so that, $\Delta Q_f (= mH_f) = -\Delta Q_M - \Delta Q_m$.

Rearranging the terms we obtain, $H_f = \dfrac{-\Delta Q_M - \Delta Q_m}{m}$

Procedure

- Determine the mass of an empty cup then pour in warm water until the cup is ¾ full, remeasure the mass and subtract the mass of the cup. Enter your result in kilograms in the table below under the column marked M (for the water's mass.) For example, a number such as 290 g should be entered as 0.290 kg. Carefully measure the water's original temperature T_i and enter it in the appropriate column.
- Take an ice cube, dry it, and then put it into the warm water. Stir the water in your cup with the thermometer until the temperature reaches a minimum and rises. Record the minimum temperature in the column labeled T on the next page. It is a good idea to cover the cup with a piece of paper to minimize heat exchange with the room. (You can insert the thermometer into a hole punched into the paper.)

- Measure the mass of the mixture after the ice has been added. The difference between the original and final measurements is the ice's mass. Enter this result in the column labeled m. (Note: A mass such as 24 g should be entered as 0.024 kg.) The temperature change of the warm water (which will be negative) should be entered in the column labeled ΔT_M, while that of the melted ice's temperature change, in warming up from 0°C to the final equilibrium temperature T, should be placed under ΔT_m.
- Calculate the (negative) heat 'gained' by the water, the heat gained by the by the melted ice, and enter your data in the columns labeled ΔQ_M and ΔQ_m, respectively.
- Calculate the latent heat of fusion for ice and enter your result under H_f. Repeat the above procedure for two or three additional trials, using new warm water each time.
- Determine the average value for ice's latent heat of fusion and summarize your results in the second table.

Trial	M	T_i	T	ΔT_M	ΔQ_M	m	ΔT_m	ΔQ_m	H_f
1									
2									
3									
4									

Average H_f	
Accepted Value	334 kJ/kg
Percent Difference	

Questions

1. What kind of internal energy change would account for the large value of H_f ?

2. The *Physics Reference Table* lists different values of H_f for different substances. Offer a plausible explanation why this should be so.

3. Why is it important that the ice cube be dried before it is placed in the warm water?

4. Why would using cold water or hot water for the experiment make the results less accurate?