

# Experiment 15 - Heat of Fusion and Heat of Solution

Phase changes and dissolving are physical processes that involve heat. In this experiment, you will determine the heat of fusion of ice (the energy required to melt ice) and the heat of solution of two different ionic compounds.

## Part 1: The Heat of Fusion of Ice

The melting (fusion) of any solid substance is an endothermic process. A substance must absorb heat in order to be converted from the solid phase to the liquid phase. In this experiment, the amount of heat required to melt one gram of solid H<sub>2</sub>O (ice) will be determined. This quantity is called the **Heat of Fusion** of ice.

When a substance is heated, the temperature rises because the heat energy is used to increase the kinetic energy of the molecules of the substance. Increasing kinetic energy means that the molecules of the substance are moving faster (since kinetic energy =  $\frac{1}{2}mv^2$ , with  $m$  = mass and  $v$  = velocity), and increased molecular velocity is observed and measured as increased temperature.

When a substance is at its melting point or boiling point, however, its temperature will not change as it is heated. Instead, it will undergo a **phase change** -it will either melt or boil. Therefore, phase changes are processes that occur at constant temperature. The heat energy is used in these cases to increase the potential energy of the molecules. In melting, energy is required to break apart the crystalline lattice of the solid state. In boiling, the molecules must be separated from each other (and the bulk liquid) to vaporize. Heat energy is used to overcome intermolecular forces holding the molecules together. This heat is called the **latent heat, or the heat of fusion**.

The amount of heat energy that is required to melt a sample of ice can be measured by allowing the ice to melt in a known amount of water. When the ice is placed in water, an exchange of thermal energy takes place. The ice absorbs energy from the water and melts (potential energy increases), while the water gets colder (the kinetic energy of the molecules decreases) as it loses thermal energy to the ice. Once the ice has melted, it then warms up (the newly melted H<sub>2</sub>O molecules now gain kinetic energy) until its temperature becomes the same as that of the water initially in the calorimeter. This means we have two sources of water: one initially in the calorimeter, which will be an energy source and one that arises from the process of the ice melting. We will use the water initially in the calorimeter as an energy source to melt the ice, and warm the water that arises from the ice.

The heat flow for the ice (any anything that came from the ice),  $q_{\text{(ice)}}$ , must be equal, but opposite in sign, to the heat flow from the initial water in the calorimeter,  $q_{\text{(H2O in cal)}}$ . This can be expressed by the equation:

$$q_{\text{(ice)}} + q_{\text{H2O in cal}} = 0 \text{ and, therefore, } q_{\text{ice}} = -q_{\text{(H2O in cal)}}$$

The amount of heat energy lost by the water in the calorimeter can be calculated by measuring its mass ( $m_{\text{(cal H2O)}}$ ) and its temperature change ( $\Delta T_{\text{cal}} = T_{\text{f cal}} - T_{\text{i cal}}$ ). The total amount of heat energy gained or lost by a sample of water in changing temperature is as follows:

$$q = C_p m \Delta T.$$

When this equation is applied using the mass and the temperature change of the water initially in the calorimeter, then  $q_{\text{(H}_2\text{O in cal)}}$  can be found. Changing the sign of this value gives  $q_{\text{i (ce)}}$ , the energy needed to melt the ice and warm the resulting water to the final temperature.

Our goal will be to determine the amount of energy per gram required to melt ice. To do this, we must realize that  $q_{\text{ice}}$ , heat that the ice absorbs, has two terms: the energy absorbed by the ice as it is melting at  $0^\circ\text{C}$  (we'll call this  $q_{\text{ice melting}}$ ), and the energy absorbed by the resulting water as it warms up from  $0^\circ\text{C}$  to the final temperature (we'll call this  $q_{\text{cold H}_2\text{O warmed}}$ ). To reach this goal, we must consider the system as isolated even though it is actually an open system. We also must assume that the water has a constant density and constant specific heat. The ice is suspect in this experiment. We assume the ice to be uniformly at  $0^\circ\text{C}$ ; however, the center of the ice is much colder than this temperature. This means that the surface of the ice has cold water which is added to the system. In order to simplify calculations we must assume that the surface of the ice is absolutely dry.

$$q_T = q_{\text{ice}} = q_{\text{ice melting}} + q_{\text{cold H}_2\text{O warmed}}$$

While  $q_{\text{ice melting}}$  is related to the heat of fusion,  $q_{\text{cold H}_2\text{O warmed}}$  can be calculated from collected data. Since  $q_{\text{cold H}_2\text{O warmed}}$  represents the amount of heat involved in changing the temperature of a sample of liquid water, it can be calculated using  $q_{\text{cold H}_2\text{O warmed}} = C_p m_{\text{ice}} \Delta T$ , where  $C_p$  is the specific heat of water,  $m$  is the mass of the water that was once ice, undergoing the temperature change, and  $\Delta T$  is the temperature change of that water. The initial temperature of this water sample is  $0^\circ\text{C}$  (exactly), and the final temperature is the measured final temperature for the entire water sample at the end of the experiment.

Therefore, you will be able to calculate the value for  $q_{\text{(H}_2\text{O in cal)}}$  and  $q_{\text{cold H}_2\text{O warmed}}$ . To find the heat of fusion of ice, you need to know  $q_{\text{ice melting}}$ .

### **RECAP OF TERMS AND HOW THEY COME TOGETHER IN A FORMULA:**

- $q_{\text{ice}}$  = energy needed to melt the ice and warm the water that came from the ice.
- $q_{\text{cold H}_2\text{O warmed}}$  = energy needed to warm the water that came from the ice
- $q_{\text{ice melting}}$  = energy needed to melt the ice
- $q_{\text{H}_2\text{O in cal}}$  = energy stored in the water in the calorimeter

$$q_{\text{ice}} + q_{\text{H}_2\text{O in cal}} = 0$$

$$q_{\text{ice}} = -q_{\text{H}_2\text{O in cal}}$$

$$q_{\text{ice}} = q_{\text{ice melting}} + q_{\text{cold water warmed}}$$

Combining these equations gives us

$$q_{\text{ice melting}} + q_{\text{cold water warmed}} = -q_{\text{H}_2\text{O in cal}}$$

Rearranging the equation for  $q_{\text{ice melting}}$  gives us:

$$q_{\text{ice melting}} = -q_{\text{(H}_2\text{O in cal)}} - q_{\text{cold water warmed}}$$

$$q_{\text{ice melting}} + C_s m_{\text{cold H}_2\text{O}} \Delta T_{\text{cold H}_2\text{O warmed}} = -C_s m_{\text{H}_2\text{O (l)}} \Delta T_{\text{H}_2\text{O (l)}}$$

$$q_{\text{ice melting}} = -C_s m_{\text{cold H}_2\text{O}} \Delta T_{\text{cold H}_2\text{O warmed}} - C_s m_{\text{H}_2\text{O (l)}} \Delta T_{\text{H}_2\text{O (l)}}$$

$$\frac{q_{\text{ice melting}}}{m_{\text{ice}}} = \Delta H_{\text{fus}} = 333.6 \frac{\text{J}}{\text{g}} \quad (\text{accepted value})$$

This is the amount of energy used to melt the amount of ice you used. To find the heat of fusion, divide the  $q_{\text{ice melting}}$  by the number of grams of ice used, and you will have the heat of fusion in units of J/g.

## Part 2: Heat of Solution of Two Solutes

When some solute dissolves in a solvent, energy can be absorbed or released. In this part of the experiment, you will be dissolving two different ionic solids in a known amount of water and determine the heat of solution of each substance. (The heat of solution is the energy involved in dissolving a specific amount of the substance in a solvent).

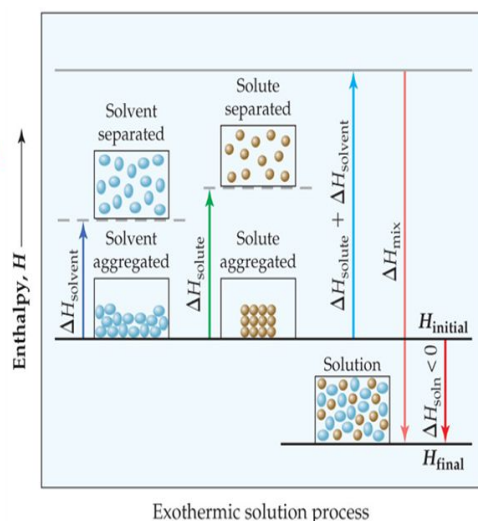
When water and an ionic compound are mixed to form a solution, the heat of solution ( $\Delta H_{\text{solution}}$ ) depends not only on the attractions between water and the ions in the solution, but also on the water-water attractions in pure water, and on the ion-ion attractions when the pure crystal.

For some substances, the heat of solution is endothermic, and for others the solution process is exothermic (see below). The overall energy change depends on two main factors. Energy must be added to break the solid apart into separate ions, and energy is released when the water molecules hydrate the solute ions. When these two amounts of energy are added together, the result can be either positive or negative (endothermic or exothermic) overall. (See Section 13.1 in text)

## Energy Changes in Solution

- Three processes affect the energetics of the process:

- Separation of solute particles
- Separation of solvent particles
- New interactions between solute and solvent



If the attractions in the solution are stronger than the attractions in the pure substances, the dissolving process will be exothermic. By dissolving, the water and the ions will reach a more stable state (lower potential energy). The difference in potential energy will appear as heat (kinetic energy). The solution will be hotter than the pure substances were before mixing.

### **Prelab-questions**

1. As the water in the calorimeter cools down, explain what happens in terms of energy on a molecular level.
2. In this lab, when the water loses energy, where does the energy go? [this applies to both parts of the lab.
3. Why do we use a Styrofoam cup and not a metal cup as a calorimeter?
4. What are 4 assumptions that we are making that influence the experiment.

### **Safety Precautions:**

- Wear your safety goggles.

### **Waste Disposal:**

- The waste from this experiment may be safely disposed of in the sink with plenty of running water.

### **Common errors to watch for:**

- **STIRRING:** It is important to stir the water and ice mixture to ensure that the temperature throughout the water is uniform. Not stirring the ice and water mixture causes the final temperature to be too warm and gives an experimental value of the Latent Heat of Fusion that is too low.
- **THERMOMETER:** The thermometer should not touch the Styrofoam calorimeter. This contact causes the final temperature to be too warm and gives an experimental value of the Latent Heat of Fusion that is too low. Also, we have seen in past labs that thermometers have a error.
- **DRYING THE ICE:** If the ice is not dried there will be water at 0°C on the ice. The added water will contribute to the final mass of liquid but it will not gain the amount of heat that an equivalent amount of ice would gain. The initial temperature of the water in the calorimeter will not have to drop as far. Hence the final temperature will be too high. The result will be an experimental value of the Latent Heat of Fusion that is too low
- **TARING THE BALANCE OR NOT USING IT CORRECLTY.** If the balance is not properly zeroed, the mass will not reflect the mass of the material.

### **Procedure – Part 1: Determining the Heat of Fusion of Ice (you will repeat this three times)**

1. Weigh the empty Styrofoam cups (glass and metal can change the energy of the isolated system) (use two nested together) that will be used to hold the ice and water.

2. Add about 100 - 150 mL of water to the Styrofoam cup. The water should be at or slightly above room temperature before it is added.
3. Weigh the cup again to determine the weight of the water.
4. Set up a ring stand with an iron ring. Place the cups in a beaker (for extra stability) and place the beaker below the ring.
5. Between temperature measurements, the thermometer can rest against the inside of the iron ring so that it won't fall over and break.
6. Measure the temperature of the water in the cup ( $\pm 0.1\text{ }^{\circ}\text{C}$ ). Be sure that the temperature has reached a constant value and is not changing. Record this temperature.

The next step is to add ice to the Styrofoam cup. The exact weight of the ice will need to be known, but it is inconvenient to weigh the ice **before** adding it to the water, since it will begin to melt as it is being weighed.

Probably the largest source of error in this experiment would be to add liquid water (on the surface of the ice) and count this liquid water as ice. Liquid water on the ice's surface has already absorbed its heat of fusion.

To minimize the melting of the ice before it reaches the water in the cup, the ice will simply be added without weighing it. The weight of the ice added will be found later, at the end of the experiment, by weighing the cup to see how much weight has been added.

7. Take about 10 - 15 grams of ice (estimated by eye as about 10 - 15 mL of ice), dry it quickly with a paper towel, and add it to the water in the cup. Be sure that the ice is as dry as possible before adding it. Allow the ice to melt, stirring constantly with the thermometer.
8. Record the lowest temperature reached as the final temperature ( $\pm 0.1\text{ }^{\circ}\text{C}$ ). This should be reached just as the ice has completely melted.
9. Weigh the cup and its contents to obtain the mass of melted ice.
10. Repeat the entire procedure two more times

## **Procedure – Part 2-(You will repeat this three times for each solute)**

### **NH<sub>4</sub>Cl**

1. Obtain 2 Styrofoam cups, and nest one in the other. Weigh the cups.
2. Transfer about 100 mL of deionized water into the cup, and weigh again.
3. Weigh out about 5 grams of solid anhydrous ammonium chloride (NH<sub>4</sub>Cl) on a piece of weighing paper or in a weighing boat. (Put the weighing paper or boat on the top-loading centigram balance, press the tare button to zero the balance, and then transfer some of the solid to the paper or boat.) Record the precise mass of solid used.
4. Carefully measure and record the temperature of the water in the cup to  $\pm 0.1^{\circ}\text{C}$ .
5. Add the ammonium chloride to the cup, and vigorously stir the mixture with the thermometer until the entire sample of solid has dissolved. Watch the temperature reading, and record the reading that differs the most from the initial temperature.
6. Dump the solution down the sink, and rinse and dry the calorimeter.

### Na<sub>2</sub>CO<sub>3</sub>

1. Obtain 2 Styrofoam cups, and nest one in the other. Weigh the cups.
2. Transfer about 100 mL of deionized water into the cup, and weigh again.
3. Weigh out about 10 grams of anhydrous sodium carbonate on a piece of weighing paper or in a weighing boat. (Put the weighing paper or boat on the top-loading centigram balance, press the tare button to zero the balance, and then transfer some of the solid to the paper or boat.) Record the precise mass of solid used.
  - a. NOTE: The sodium carbonate will not dissolve as easily as the ammonium chloride, so be sure to stir the solution vigorously. The sodium carbonate must dissolve in less than 40 seconds, or your results will be inaccurate.
4. Carefully measure and record the temperature of the water in the cup to  $\pm 0.1^\circ\text{C}$ .
5. Add the ammonium chloride to the cup, and vigorously stir the mixture with the thermometer until the entire sample of solid has dissolved. Watch the temperature reading, and record the reading that differs the most from the initial temperature.
6. Dump the solution down the sink, and rinse and dry the calorimeter.

## **Calculations – Part 1 (You need to do the calculations for each trial)**

1. Calculate the mass of water initially placed in the calorimeter.
2. Calculate the mass of ice that was used.
3. Calculate the temperature difference that the water initially in the calorimeter underwent as it cooled from the initial temperature to the final temperature of the system.
4. Calculate the temperature difference that the water that was initially ice underwent as it warmed from  $0^\circ\text{C}$  to the final temperature of the system.
5. Calculate  $-Q_{\text{H}_2\text{O in cal}}$  using the formula  $C_p m \Delta T$ . Use the correct mass and temperature change here.
6. Calculate  $q_{\text{cold H}_2\text{O warmed}}$  using the formula  $C_p m \Delta T$ . Use the correct mass and temperature change here.
7. Calculate  $q_{\text{ice melting}}$  from the difference between the two heats.
8. Calculate the percent difference, the percent error using the theoretical value of  $\Delta H_{\text{fusion}}$  for ice (333 J/g) and your average value of  $\Delta H_{\text{fusion}}$  as the experimental value.

## **Calculations – Part 2**

Note: Do these calculations separately for each trial.

1. Determine the mass of water and mass of solution in the cup.
2. From the mass of solution, the temperature change the solution underwent, and the heat capacity of the solution (assume the heat capacity of the solution is very close to the heat capacity of water), determine the amount of energy absorbed by/given off by the water/solution in the cup.
3. Determine the amount of energy absorbed by or given off by the dissolving solid dissolving in water.

- Determine the number of moles of solid used.
- Calculate the “heat of solution” of this compound ( $\Delta H_{\text{soln}}$ ) in units of kilojoules per mole of solute. Make sure to include the appropriate sign.
- Calculate the average and percent difference for each compound.

## Questions:

### Part 1

- Suppose that one were to mix 30.0 g of aluminum pellets, originally at 97.0 °C with 100.00g of water, originally at 23.0°C, in a perfect calorimeter. What will be the equilibrium, or final temperature of the water (and the aluminum metal)?
- How could you determine the heat of vaporization of water at 100°C in a similar experiment? (Think steamed milk). Explain the procedure you would use along with calculations you would make. This is a perfect opportunity to study at the local coffee shop. Watch how your latte is made!
- A student and her partner performed the same experiment as you. Sadly, her partner was in a hurry in trial 2 and did not zero the balance when weighing the water in the calorimeter. This introduced a 5.00% error in the weight of the water in the calorimeter which was transferred throughout the trial date. If the mass of the water was 105.0 g instead of 100.0g, the  $\Delta T$  of the water was -9.51°C, the mass of the water from ice is 10.00g, and the  $\Delta T$  of water from ice is 15.48°C, calculate the  $\Delta H_{\text{fus}}$  induced from not taring the balance. Use the theoretical value of 333.6J/g to calculate the % error.
- We assumed that the surface of the ice did not have melted water. Suppose some of the ice melted before it was transferred to the water at room temperature. If we introduce a 5.00% error for the mass of ice, then 0.500 grams of it would be already in the form of water at 0.00°C. **Calculate the final temperature of the system** if the mass of the water in the calorimeter is 100.0 grams, the initial temperature of the water in the calorimeter is 25.0°C, the mass of the water from the ice cube is 10.0 grams, at 0.0°C, but the ice is 0.500 grams less than expected. Use the theoretical value of 333.6J/g as the heat of fusion. If we assumed the ice did not melt until it was transferred in to the calorimeter, the final temperature will be 15.48°C; use this value to calculate the % error of the temperature.
  - Is the final temperature higher or lower than the temperature if the ice had not melted?
  - Does the calculated final temperature with the 5% error in the mass of ice make sense?
  - How is the heat distributed to support the final temperature reflected by the 5% error in the ice?

- How do we justify the mass of the water from the ice and the ice not having the same mass?

## **Part 2**

1. Which solute gave off heat as it dissolved? Explain where this heat came from.