# Experiment 15 - Heat of Fusion and Heat of Solution

Phase changes and dissolution are physical processes that absorb or release 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.

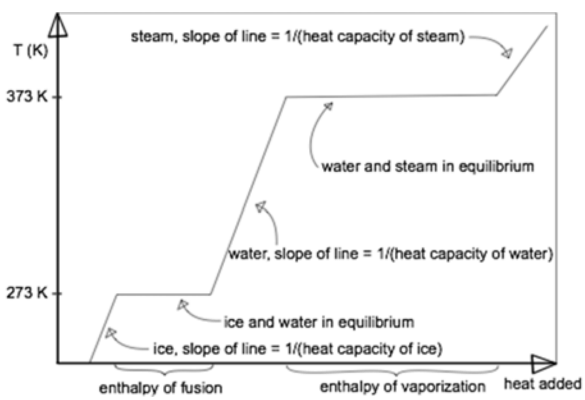
**Part 1: The Heat of Fusion of Ice**

The melting (fusion) of any solid substance is an endothermic process. Heat must be absorbed by a substance 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 $H_2O$ (ice) will be determined. This quantity is called the **heat of fusion** of ice.

Normally, when a substance is heated, the temperature rises. This is 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, and increased average 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 - and 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 in order to vaporize. Heat energy is used to overcome intermolecular forces holding the molecules together, increasing the potential energy (a process analogous to increasing the gravitational potential energy of a weight by lifting it).

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 liquid 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 $H_2O$ molecules now gain kinetic energy) until its temperature becomes the same as that of the originally liquid water.




If the process of ice melting in water occurs in an insulated container, no heat energy will enter the system from the outside. All of the energy gained by the ice in melting and warming will come from the originally liquid water. In other words, the heat absorbed by the ice, $q_{abs}$, must be equal, but opposite in sign, to the heat released by the liquid water, $q_{rel}$. This can be expressed by the equation:

$$q_{abs} + q_{rel} = 0 \qquad \text{Eq 1}$$

The amount of heat energy lost by the originally liquid water can be calculated by using the **specific heat**, temperature change and mass of the water. The specific heat, or heat capacity per gram, has units of J/g·°C. The calorimetry equation, which relates specific heat to heat, is

$$q = m \cdot C_s \cdot \Delta T \qquad \text{Eq 2}$$

where $q$ is the heat required to cause temperature change $\Delta T$ (where $\Delta T = T_{final} - T_{initial}$) in $m$ grams of a substance with specific heat $C_s$. The definition of the calorie states that 1.00 calorie is required to heat one gram of water by 1 °C. Since we will be working in joules, the specific heat of water is 4.18 J/g°C, which implies that 1 cal = 4.18 J.

Our goal will be to determine the amount of energy per gram required to melt ice. In order to do this, we must realize that $q_{abs}$ has two terms: the energy absorbed by the ice as it is melting at 0 °C (we'll call this $q_{melt}$), and the energy absorbed by the resulting water as it warms up from 0 °C to the final temperature (we'll call this $q_{warm}$).

$$q_{abs} = q_{melt} + q_{warm} \qquad \text{Eq 3}$$

While $q_{melt}$ is related to the heat of fusion, $q_{warm}$ can be calculated from the Eq 2, where $m$ will be equal to the mass of ice, since this is the water that was originally ice. The initial temperature of this water sample is 0 °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 $q_{rel}$ and $q_{warm}$. To find the heat of fusion of ice, you need to know $q_{melt}$.

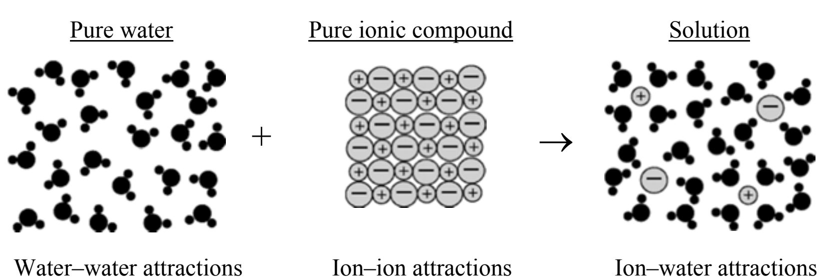
$$-q_{rel} = q_{abs} = q_{melt} + q_{warm} \qquad \text{Eq 4}$$

Use the above relationships to solve for $q_{melt}$. This is the amount of energy used to melt the amount of ice you used. To find the heat of fusion, divide the $q_{melt}$ 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**

When a solute dissolves in a solvent, energy can be absorbed or released. In this experiment, you will be dissolving two different ionic solids in water and determining 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 particular solvent).

When water and an ionic compound are mixed to form a solution, the heat of solution ($\Delta H_{soln}$) 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 in the pure crystal.



For some substances, the heat of solution is endothermic, and for others the solution process is exothermic. The overall energy change depends on two main factors. Energy must be added to break the solid apart into separate ions and to separate the solvent molecules to make room for the ions, and energy is released when ions are hydrated by water molecules. When these two amounts of energy are added together, the result can be either positive or negative (endothermic or exothermic) overall.

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. If the attractions in the solution are *weaker* than the attractions in the separate solute and solvent, then the dissolving process will be **endothermic**, and will absorb heat from the solution, making the solution colder than it started. However, even if the dissolution is endothermic, the substance may be soluble because there is a natural tendency toward mixing.

The names and symbols here can get confusing. Since we don't yet know what's releasing and absorbing energy, let's call the heat released or absorbed by the process of dissolution $q_{dis}$, and the heat released to or absorbed from the solution as a result $q_{sol}$. Then, just like Eq 1,

$$q_{dis} + q_{sol} = 0 \qquad \text{Eq 5}$$

You can find $q_{sol}$ using Eq 2, and $q_{dis}$ is the heat of solution multiplied by the number of moles.

**<u>Safety Precautions:</u>**
- Wear your safety goggles.

**<u>Waste Disposal:</u>**
- The waste from this experiment may be safely disposed of in the sink with plenty of running water.
- Rinse and save the cups unless they are broken.

## Procedure – Part 1

The procedure here is similar to that used in Experiment 8. We will add some ice to a known mass of water at a known temperature, and record the final minimum temperature after all the ice has melted. 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. A large 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, since melted ice 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 can be found later, at the end of the experiment, by weighing the cup to see how much weight has been added.

1. Set up an appropriate table in your notebook to hold the data.
2. Weigh the empty Styrofoam cups (use two stacked cups) that will be used to hold the ice and water.
3. Add 100–150 mL of water to the Styrofoam cup. The water should be at or slightly above room temperature before it is added. 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. Between temperature measurements, the thermometer can rest against the inside of the iron ring so that it won't fall over and break.
5. Measure the temperature of the water in the cup (± 0.1 °C). Be sure that the temperature has reached a constant value and is not changing. Record this temperature.
6. Take 10–15 grams of ice (estimated by eye, this is about 3 connected pieces or 5 separate pieces from our machine), 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.
7. Allow the ice to melt, **stirring constantly** with the thermometer. Record the lowest temperature reached as the final temperature (±0.1 °C). This should be reached just as the ice has completely melted.
8. Weigh the cup and its contents to obtain the mass of melted ice by subtraction.
9. Repeat the entire procedure.

## Calculations – Part 1

1. Calculate the heat of fusion of ice in J/g for each trial using the equations in the introduction. Average the two $\Delta H_{fus}$ values.
2. Calculate the percent error using the theoretical value of $\Delta H_{fus}$ for ice (333 J/g) and your average value of $\Delta H_{fus}$ as the experimental value.
3. Write down your results for $\Delta H_{fus}$ on the board, along with your name. The instructor will distribute the results for the class so you can use them in your post-lab.

## Procedure – Part 2

1. Dry and reweigh your cups if you have doubt of their mass.
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 ($NH_4Cl$) on a piece of weighing paper or in a weighing boat. (Put the weighing paper or boat on the top-loading 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 ±0.1°C.
5. Add the ammonium chloride to the cup, and vigorously stir the mixture with the thermometer until all of the 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.
7. Repeat steps 1–6, using about 10 grams of anhydrous sodium carbonate ($Na_2CO_3$) instead of 5 g of $NH_4Cl$. 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.

## 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, 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 or released by the water/solution in the cup using Eq 2. This is $q_{sol}$ in Eq 5.

3. Determine the amount of energy absorbed or released by the dissolving solid, $q_{dis}$, using Eq 5.
4. Determine the number of moles of solid used.
5. Calculate the “heat of solution” of this compound ($\Delta H_{soln}$) in units of kJ/mol of solute. Make sure to include the appropriate sign.

Name: Section:

**Experiment 15 Pre-Lab Sheet**

1. (2 pts) Fill in the appropriate variables for these statements.
   Choose from $q_{abs}$, $q_{rel}$, $q_{melt}$, $q_{dis}$, $q_{sol}$, $m_{ice}$, $m_{water}$, $m_{solute}$, $m_{solution}$, $n_{ice}$, $n_{water}$, $n_{solute}$, and $n_{solution}$.

   To find the heat of fusion of ice in J/g, you divide __________ by the ______________.

   To find the heat of solution in kJ/mol, you divide ___________ by the ______________.

2. (1 pts) Why should you dry the ice before adding it to the calorimeter?

3. (1 pts) Why is it important to stir constantly during your trials?

4. (3 pts) On the simple heating diagram, label the following:

   A: where the "ice" starts if it is wet,
   B: where the ice starts if it is below 0°C,
   C: where the ice starts if it is dry at 0°C,
   D: where the water starts (approximately),
   E: where everything finishes, approximately.


5. (3 pts) Sample calculation: You add 5 g of ice to 100. g of water. The initial temperature of the water is 23 °C. The minimum temperature of the water is 19 °C. Based on this data, you should find that the enthalpy of fusion of water is 255 J/g. Do the calculation, showing your work clearly, and make sure that your answer is 255 J/g.