Enthalpy of Fusion of Ice

Revised, 10-12-17

Enthalpy of Fusion is defined as the amount of energy required to melt one gram of a solid which is already at its melting point. This property is a constant for a particular substance which is dependent on the attractions which hold that substance together. The stronger the attraction, the greater the amount of energy needed to melt it. Water has a significantly higher enthalpy of fusion than other substances made of molecules with similar size and structure, due to hydrogen bonding. This especially strong dipole-dipole attraction is also responsible for water’s unexpectedly high melting point, boiling point, enthalpy of vaporization, and its unexpectedly low vapor pressure. The structure of ice involves four hydrogen “bonds” around each water molecule. Each oxygen atom attracts a hydrogen atom in TWO other molecules and each hydrogen attracts ONE oxygen. In the process of changing ice to liquid water, ***some*** of these hydrogen bonds must be overcome to break down the structure of ice to the point where molecules can flow over and around each other.

**Objectives:**

In this experiment, you will;

* use calorimetry to measure the thermal energy transfer which occurs as ice melts in a cup of warm water,
* calculate the enthalpy of fusion of ice (kJ/g) as well as the molar enthalpy of fusion (kJ/mol), and
* estimate the percentage of hydrogen bonds which break during the melting process.

**Procedure:**

1. Connect a Lab Pro interface to your computer, connect the power supply for the Lab Pro and connect a temperature probe via the CH1 port. You will use the probe as your thermometer whenever the lab calls for a temperature measurement. (OPTIONAL: You may also use a regular thermometer to confirm the probe readings.)
2. Open the most recent version of Logger Pro on your laptop. Throughout the lab you are encouraged to experiment with/familiarize yourself with the adjustments, parameters and labeling you can do with Logger Pro.
3. Heat about 150 mL of tap water in a beaker over a Bunsen burner to about 60°C.
4. Record the mass of an empty calorimeter (insulated cup with a cardboard lid).
5. Fill the calorimeter about half full with the tap water you heated. Replace the lid, then measure and record the mass.
6. Precisely measure and record the temperature of the water in the calorimeter.
7. Add 2-3 ice cubes to the calorimeter. Measure and record the mass.
8. Monitor the temperature as the ice melts. Measure and record the temperature as soon as the ice is completely melted.

**Calculations:**

1. Calculate the heat lost by the hot tap water. (cH2O=4.184 J/gC°)
2. Calculate the enthalpy of fusion of ice in kJ/g.
3. Convert the enthalpy of fusion to the molar enthalpy of fusion (kJ/mol).
4. Calculate the energy needed to break ALL of the hydrogen bonds in one mole of ice. (HINT: Think about the number of H-bonds per molecule (your first thought is likely to be wrong). The bond energy for the H-bonds in water is 3.5x10-20J/bond.)
5. Estimate the percentage of H-bonds that are broken during the melting process.

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**Data:**

This experiment was performed on \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

|  |  |
| --- | --- |
| Mass of empty calorimeter and lid |  g |
| Mass of calorimeter, lid and water |  g |
| Mass of water |  g |
| Mass of calorimeter, lid, water and ice |  g |
| Mass of ice |  g |
| Initial temperature of hot water |  °C |
| Final temperature, after ice melts |  °C |
| Change in temperature |  C ° |
| Heat lost by hot water |  J |
| Heat of fusion of ice |  kJ/g |
| Molar heat of fusion |  kJ/mol  |
| Energy needed to break ALL H-bonds in one mole of ice |  kJ  |
| % of H-bonds broken during melting |  % |

Questions:

1. Define enthalpy of fusion and molar enthalpy of fusion.
2. What is an isothermal change? Why is melting not an automatic result of heating a solid to its melting point? Why is freezing not an automatic result of cooling a liquid to its freezing point?
3. Explain why energy must be added to a solid at its melting point in order to melt it.
4. Explain why water’s enthalpies of fusion and vaporization and melting and boiling points are so much higher than other substances with molecules of similar size and structure.
5. Propose at least three ways in which our environment, weather patterns and life as we know it would be (potentially devastatingly) different without hydrogen bonding.