

Measuring Energy Changes

Heat of Fusion

Introduction

A physical change of state or a chemical reaction may be either exothermic or endothermic. An exothermic reaction releases heat to its surroundings, while an endothermic reaction absorbs heat from its surroundings. What is the source of the heat energy that is released in an exothermic reaction? What happens to the heat energy that is absorbed in an endothermic reaction? Can the amount of heat energy be measured?

Concepts

- Heat vs. temperature
- Exothermic vs. endothermic
- Heat of fusion
- Heat of vaporization
- Enthalpy change

Background

Our everyday experience tells us that energy in the form of heat is needed to melt ice or boil water. Imagine a beaker of water at room temperature on a hot plate. As the beaker is heated, the temperature of the water will increase steadily from 25 °C to 100 °C, the boiling point of water. If the water is heated further, it will begin to boil, but the temperature of the water in the beaker will remain constant at 100 °C until all of the water has been converted to steam. All of this time, of course, while the temperature has remained constant, heat has been added to the system. The heat absorbed by the water has been used to break apart the forces (e.g., hydrogen bonding) between water molecules in the liquid phase. The amount of heat that must be absorbed to vaporize a specific quantity of liquid (usually one gram or one mole) is called the *heat of vaporization*. The heat of vaporization for water is 2260 Joules per gram. In a similar manner, heat is also required to melt ice. The amount of heat that must be absorbed to melt a specific quantity of solid is called the *heat of fusion*.

Vaporization and fusion are examples of endothermic physical changes. The reverse physical processes must therefore be exothermic. When steam condenses to water, for example, an amount of heat energy equal to the heat of vaporization must be released to the surroundings. This explains why steam burns are more dangerous than hot water burns. Heat is also released when water freezes to ice.

The amount of heat transferred in these processes depends on the difference in the energy stored in each substance. This stored energy is called the heat content or enthalpy, and is represented by the symbol H . The *enthalpy change* (ΔH) for a physical process or a chemical reaction is defined as the heat change that occurs at constant pressure. This is convenient, because most of the reactions that are carried out in the lab are in flasks or containers that are open to the atmosphere—that is, they take place at a constant pressure equal to the barometric pressure.

Equation 1 shows the equality between the change in enthalpy (ΔH) of a system and the amount of heat transferred, symbolized by q_p , for a reaction carried out at constant pressure.

$$\Delta H = q_p \quad \text{Equation 1}$$

The amount of heat (q_p) transferred to a substance or object depends on three factors: the mass (m) of the object, its specific heat (s), and the resulting temperature change (ΔT). See Equation 2.

$$q_p = m \times s \times \Delta T \quad \text{Equation 2}$$

The specific heat (s) of a substance reflects its ability to absorb heat energy and is defined as the amount of heat needed to raise the temperature of one gram of a substance by one degree Celsius. The specific heat of water is equal to $4.18 \text{ J/g}\cdot^\circ\text{C}$. The temperature change (ΔT) is equal to the difference between the final temperature and the initial temperature ($\Delta T = T_{\text{final}} - T_{\text{initial}}$).

In most laboratory situations, the temperature change is measured not for the system itself (the reactants and products), but for the surroundings (the solution and the reaction vessel). The amount of heat released by the system must be equal to the amount of heat absorbed by the surroundings. The sign convention in Equation 3 reveals that the heat change occurs in the opposite direction.

$$q(\text{system}) = -q(\text{surroundings}) \quad \text{Equation 3}$$

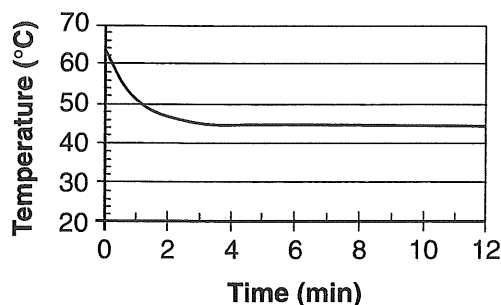
For an exothermic reaction, the heat released by the system results in a temperature increase for the surroundings (ΔT is positive) and the heat absorbed by the surroundings will be a positive quantity. The heat released by the system must have the reverse sign—it must be a negative quantity. According to this convention, the enthalpy change for an exothermic reaction is always a negative value. For an endothermic reaction, in contrast, the heat absorbed by the system results in a temperature decrease for the surroundings (ΔT is negative) and the heat released by the surroundings will be a negative quantity. The heat absorbed by the system must have the opposite sign—it must be a positive quantity. According to this convention, the enthalpy change for an endothermic reaction is always a positive value.

Experiment Overview

The purpose of this experiment is to determine the temperature and heat changes that occur when ice melts. In Part A a heating curve will be obtained by melting ice in a beaker on a hot plate and measuring temperature as a function of time. In Part B ice will be placed in a beaker of warm water and the temperature change that is produced as the ice melts will be measured and used to calculate the heat of fusion of water (the heat required to melt ice).

Pre-Lab Questions

A sample of lauric acid—an organic compound used to make soap—was heated above its melting point in a test tube and then cooled in a bath of water until it solidified. The temperature of the lauric acid was measured as a function of time and the following graph was recorded.



1. What is the physical significance of the flat region (plateau) on the curve?
2. Use the graph to estimate the melting point of lauric acid.
3. Is heat being absorbed or released by the lauric acid sample as it solidifies?

Materials

Beakers, 400-mL, 2	Digital thermometer
Distilled or deionized water	Hot plate
Graduated cylinders, 100- and 250-mL	Insulated foam (Styrofoam™) cup, 6 oz
Ice, crushed, about 200 mL volume	Stirring rod
Ice cubes, 2	Beaker tongs or Hothands®

Safety Precautions

Exercise caution when using the hot plate and handling hot glassware. Remember that a “hot” hot plate looks exactly the same as a “cold” hot plate. Use beaker tongs or Hothands® to handle hot glassware. Wear chemical splash goggles whenever chemicals, heat, or glassware are used in the laboratory.

Procedure

Part A. Temperature and Phase Changes

1. Obtain a 400-mL beaker and fill the beaker to the 150-mL mark with crushed ice.
2. Place the digital thermometer in the ice and measure its temperature (it should be between 0 and 2 °C). Record the lowest temperature as the initial temperature (time 0) in Data Table A.
3. Place the beaker on the hot plate and adjust the heat setting to the halfway point (for example, if the heat dial goes from 0–10, adjust the setting to 5). Stir the ice constantly using a stirring rod.

4. Measure and record the temperature of the ice every minute. Note in Data Table A the temperature at which all of the ice has melted.
5. When all of the ice has melted, adjust the hot plate setting to its maximum value (10 in the above example) and continue heating and stirring the water. Do not allow the thermometer to touch the bottom of the beaker.
6. Record the temperature of the water every minute. Note in Data Table A the temperature at which the water begins to boil. Continue heating until the water has boiled for three minutes.
7. Turn off the hot plate and carefully remove the beaker using beaker tongs or Hothand® heat protectors. ...

Part B. Energy Needed to Melt Ice

8. Obtain 100 mL of *warm* water (about 50 °C) in a graduated cylinder. Measure the initial volume of water to the nearest 0.2 mL and record this value in Data Table B.
9. Place a Styrofoam cup in a 400-mL beaker for stability and pour the water into the cup. Measure and record the temperature of the warm water to the nearest 0.1 °C in Data Table B.
10. Obtain several ice cubes. Shake any excess water off the ice cubes and carefully add the ice cubes to the warm water bath.
11. Stir the ice/water mixture until the temperature is around 0 °C (within 2–4 °C). Add more ice, if necessary, to cool the water to this temperature.
12. When the temperature has reached its lowest value (again, this should be between 0 and 4 °C), record the temperature and immediately remove any unmelted ice from the water bath using tongs.
13. Carefully pour the water from the Styrofoam cup into a 250-mL graduated cylinder. Measure and record the final volume of water to the nearest 1 mL in Data Table B.
Note: If a 250-mL graduated cylinder is not available, measure the volume of water in two batches using a 100-mL graduated cylinder.

Name: _____

Class/Lab Period: _____

Measuring Energy Changes

Data Table A. Temperature and Phase Changes

Time (min)	Temperature (°C)	Notes	Time (min) continued	Temperature (°C) continued	Notes
0 (initial)					

Data Table B. Energy Needed to Melt Ice

Initial Volume (Warm Water)	Initial Temperature (Warm Water)	Final Volume (Ice Water)	Final Temperature (Ice Water)

Post-Lab Calculations and Analysis (Use a separate sheet of paper to answer the following questions.)

Part A. Temperature and Phase Changes

- Using your data, draw a graph of temperature (y-axis) versus time (x-axis).
- Note on the graph the temperature at which all of the ice has melted and the temperature at which the water starts to boil.
- Are there any temperature plateaus (flat regions of the curve) on the temperature versus time graph? Name the physical property of water corresponding to each temperature plateau.
- Was heat being added to the system during the times that the temperature remained relatively constant? Use the law of conservation of energy to describe what happened to the heat energy that was absorbed during this time.
- Was heat being added to the system during the times that the temperature was rising? Use the law of conservation of energy to describe what happened to the heat energy that was absorbed during this time.

Part B. Energy Needed to Melt Ice

- Use the appropriate density value from the following table of density versus temperature to calculate the mass of warm water that was cooled by the ice in Part B.

Temperature	40 °C	45 °C	50 °C	55 °C
Density of water	0.992 g/mL	0.990 g/mL	0.988 g/mL	0.985 g/mL

- Use Equation 2 in the *Background* section to calculate the amount of heat in joules released by the warm water as it cooled.
- Use Equation 3 to calculate the amount of heat absorbed by the ice as it melted.
- Determine the volume of ice melted in Part B and calculate its mass. *Note:* The density of water at 0 °C is 1.00 g/mL. Divide the result in Question #8 by the mass of ice melted to determine the amount of energy absorbed per gram of ice as it melted.
- (a) Use the gram formula weight of water to calculate the heat of fusion of water in *kilojoules per mole* (kJ/mole).
 (b) Write a chemical equation for melting ice and include the energy term in kJ/mole.
 (c) The literature value for the heat of fusion of water is 6.02 kJ/mole. Calculate the percent error in the experimentally determined heat of fusion.

$$\text{Percent error} = \frac{|\text{experimental} - \text{literature}|}{\text{literature}} \times 100\%$$