

Liquids and Solids

The Vapor Pressure and Enthalpy of Vaporization of Water

The Enthalpy of Fusion of Water

Purpose

- Determine the vapor pressure of water as a function of temperature.
- Use the vapor pressure data to make a Clausius-Clapeyron plot [$\ln(P)$ versus $1/T$].
- Determine the enthalpy of vaporization of water from the slope of the $\ln(P)$ versus $1/T$ plot.
- Measure the enthalpy of fusion of ice using a Styrofoam-cup calorimeter.

Introduction

Heat and molecular motion

Our accumulated knowledge about the effect of heat provides a rather detailed picture of what happens when you heat (or transfer energy to) a substance. On an atomic or molecular scale, heating increases the motion of the atoms. If the atoms are linked by chemical bonds to form a molecule, we can focus our attention not on individual atoms but on the molecule as a whole.

If we heat a molecule, it moves about faster. This motion of the molecule's center of mass is called *translational motion* (meaning "motion from place to place"). Gaseous molecules have a substantial fraction of their total energy stored as kinetic energy of translation. They move about rapidly, bouncing off one another and off the walls of their container.

Molecules can also rotate about their centers of mass; as a gas is heated, the molecules rotate faster. Energy is stored in the rotating molecules, just as energy is stored in a rotating flywheel.

Individual atoms in molecules are linked by chemical bonds that are "springy," constraining the atoms that are bonded together to move like masses connected by springs. As the molecules are heated, the energies and amplitudes of their vibrations increase.

To summarize, the energy we put into a molecule is stored as *translational, rotational, and vibrational motions* of the molecules.

When we cool a substance, we remove energy from it, and molecular motions slow down. We reach the lowest temperature possible, called *absolute zero*, when we have removed as much energy as possible from the substance. At this temperature, molecular motion nearly ceases.

As we remove energy from gaseous molecules, at some temperature they will condense into a liquid and eventually into a solid, because there are attractive forces between molecules caused by dipole-dipole (or induced-dipole) attraction. To separate the molecules in a solid to the distances characteristic of those in a gas, we must do work against the opposing attractive forces. (Recall that $\text{work} = \text{force} \times \text{distance}$.) This work is the energy required to melt the solid and vaporize the liquid.

Now imagine the following experiment. Suppose we begin adding energy to ice at a uniform rate while measuring the temperature. For water, a plot of temperature versus energy (or time) would look like Figure 21-1. As we warm the solid, the molecules begin to move. On average, the molecules in the solid are not going anywhere. They simply begin to bounce around in the cage formed by their surrounding neighbors. In effect, they are like gas molecules confined to a very small container not much larger than the molecule itself. At the same time, the amplitude

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Scott C. Dulebohn



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Chemistry in the Laboratory James M. Postma, Julian L. Roberts, J. Leland Hollenberg, 2009-08-07 For nearly 40 years Chemistry in the Laboratory has been meeting the needs of teachers and students This new edition builds on that legacy while addressing cutting edge trends in the chemistry laboratory including forensic chemistry and environmental and green chemistry As always the new edition of Chemistry in the Laboratory offers precise easy to follow instructions helpful illustrations and an emphasis throughout on laboratory safety Again throughout a Consider This feature encourages users to expand the principles of the experiment into interesting applications open ended experiments or unexplored corners Most experiments in the manual can be completed in one lab session but some can be linked or extended for a multi lab project

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