Chapter 7

STATES OF MATTER

Chapter Check-In

- ☐ Learning the characteristics of the three states of matter: solid, liquid, and gas
- ☐ Understanding how pressure and temperature can determine and change the physical state of a substance

Nearly every substance can exist as a solid, a liquid, or a gas. These are the three common states of matter. Whether a substance is a solid or a liquid or a gas depends on its temperature and the pressure placed on it. At room temperature (about 22° C) and at the normal pressure exerted by the atmosphere, water exists as a liquid, which can flow from one container to another. But if its temperature is lowered to about –10° C, liquid water freezes to solid ice. Going the opposite direction in temperature and at this same pressure, water changes to a gas when the temperature exceeds 100° C. Changes in state can also occur by changing the pressure while holding temperature constant. The relationship between temperature and pressure and the three states of matter is easier to see when displayed in a phase diagram. Because phase diagrams provide so much information, they are known for thousands of substances.

Any change in phase is accompanied by the taking in or release of heat energy because, as change takes place, the attractive forces between molecules are being broken down or being formed. As solid water converts to liquid water, heat is absorbed as the forces between water molecules weaken, allowing the liquid to flow. The energy involved in phase changes is accurately known for many substances. The heat energy needed to warm or cool solids, liquids, and gases without changing phase is also accurately known.

Solids, Liquids, and Gases

The familiar compound H_2O provides the evidence that substances occur in three different physical classes called **states.** At room temperature,

H_2O is a dense fluid called a **liquid.** When this liquid is chilled to 0° C, it changes to a rigid **solid.** If the liquid is heated to 100° C, however, it abruptly expands to a tenuous fluid called vapor or **gas.**

Such different states of matter are not unique to H_2O. Almost all substances can exist in two or three of the fundamental states. Table 7-1 defines the states in terms of the shape and volume of substances. Because both liquids and gases flow readily, they are collectively referred to as **fluids.**

Table 7-1 Definitions of the States of Matter

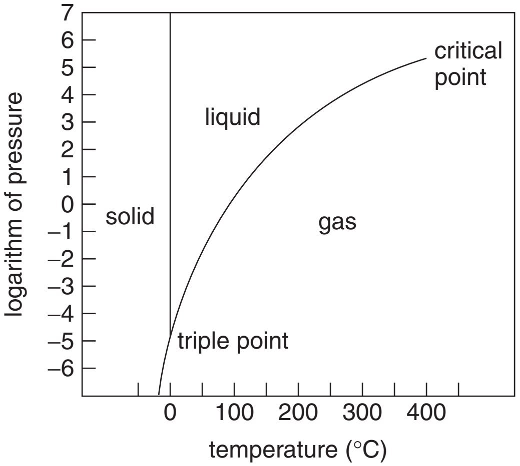
State of Matter	*Shape of Substance*	*Volume of Substance*
Solid	Definite	Definite
Liquid	Indefinite	Definite
Gas	Indefinite	Indefinite

These states have different properties because they have distinct structures on the atomic or molecular scale. In a solid, the atoms are strongly bonded to the surrounding atoms so each is in a fixed position; if the solid structure has a regular pattern that is repeated throughout the solid, it is described as a *crystalline structure.* The atoms or molecules in a liquid are less strongly bonded to one another than in a solid of the same chemical composition, and consequently, they may shift their positions. The bonds between molecules in a liquid are, nevertheless, strong enough so the molecules stay in contact with surrounding molecules. In a gas, the bonding between individual molecules is essentially zero, and individual molecules may move in all directions, allowing the vapor to expand throughout any container.

Phase Diagrams

Although the introductory example of H_2O mentioned changes of state caused by varying the temperature, it is known that variation of pressure can also produce such changes. In laboratory experiments, these two environmental factors—temperature and pressure—can each be varied or held constant, they are referred to as *independent variables.* Figure 7-1 assigns these variables to axes to form a plot that describes the physical condition at each point in the graph. The vertical axis is the natural logarithm (ln) of the pressure measured in atmospheres (atm).

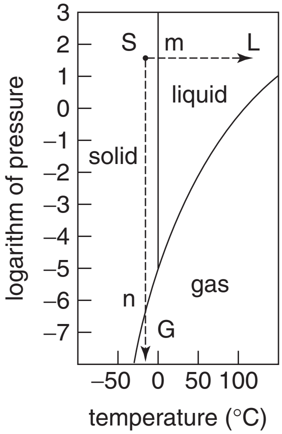
Figure 7-1 The phase diagram for water.



A temperature-pressure graph showing the various states of matter is a phase diagram. **Phase** refers to a single homogeneous physical state. Different phases have either different compositions or different physical states. In the preceding figure, there are 3 phases with the same composition solid, liquid, and gas.

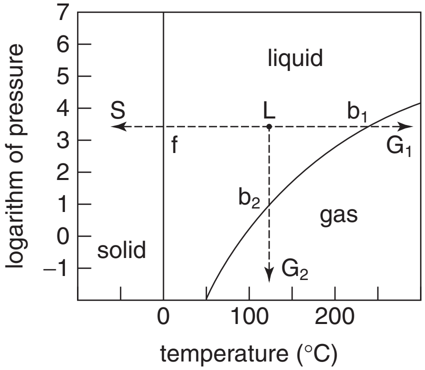
Begin studying how both temperature and pressure determine the state of H_2O by taking some ice at a temperature of –10° C and pressure of 5 atmospheres, labeled *S* in Figure 7-2. If the pressure is held constant but the temperature is increased, the substance heats up along the dashed line marked *L,* melting to a liquid at point *m,* about –0.01° C. Alternatively, if you decrease the pressure on the initial solid *S*, while holding the temperature constant at –10° C, the conditions change downward along path *G*, and the ice vaporizes abruptly when the pressure has fallen to the point marked *n*, about 3×10^{-3} atm. Such a direct change from a solid to a gas is called **sublimation;** notice that there was no intervening liquid state.

Figure 7-2 Changing the phase of solid water.



In the graph in Figure 7-3, study the possible state changes of an initial liquid marked *L*. The liquid is assumed to begin at 120° C and 30 atmospheres. The high pressure allows this liquid to exist at a temperature exceeding the 100° C boiling point at 1 atmosphere. If the pressure is maintained at a constant 30 atm, cooling the liquid *L* will produce a change to the left along path *S*, and the liquid will freeze at point *f* (about –0.01° C) to solid ice. A second course with constant pressure is heating *L* toward G_1, and the liquid will abruptly vaporize at boiling point b_1 (about 235° C). Returning to the initial liquid *L*, you can imagine holding the temperature constant at 120° C while decreasing the pressure toward G_2. When the pressure falls to approximately 2 atm, the liquid will boil at point b_2. Boiling has been induced without heating the liquid.

Figure 7-3 Changing the phase of liquid water.

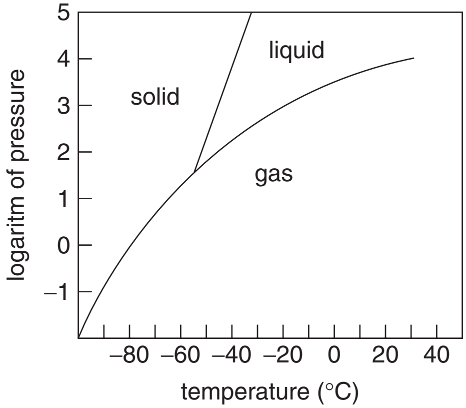


In summary, a change of state can be caused by varying only the temperature, or varying only the pressure, or varying both temperature and pressure. Most random combinations of temperature and pressure fall within the three areas of a phase diagram in which only a single state is stable. The special temperature-pressure combinations plotted as lines in the phase diagram of H_2O (see Figure 7-2) are where two states can coexist. For example, both solid ice and liquid water are stable at precisely O° C and 1 atm.

Look back at the large phase diagram (Figure 7-1) and notice the intersection of the three lines at 0.01° and 6×10^{-3} atm. Only at this **triple point** can the solid, liquid, and vapor states of H_2O all coexist. Now find the point at 374° C and 218 atm where the liquid/gas boundary terminates. This **critical point** is the highest temperature and highest pressure at which there is a difference between liquid and gas states. At either a temperature or a pressure over the critical point, only a single fluid state exists, and there is a smooth transition from a dense, liquid-like fluid to a tenuous, gas-like fluid.

Each substance has its own phase diagram to display how temperature and pressure determine its properties. Figure 7-4 is the phase diagram for carbon dioxide.

Figure 7-4 The phase diagram for carbon dioxide.



Use Figure 7-4 to answer the two practice problems.

> **Problem 15:** What is the minimum pressure in atmospheres at which CO_2 can occur as a liquid?
>
> **Problem 16:** If pressure is held at a uniform 3 atmospheres, at what temperature does solid CO_2 become unstable? What phase begins to appear at this temperature?

Heat Capacities and Transformations

For chemical reactions and phase transformations, the energy absorbed or liberated is measured as **heat.** The principal unit for reporting heat is the **calorie,** which is defined as the energy needed to raise the temperature of 1 gram of water at 14.5° C by a single degree. The term *kilocalorie* refers to 1,000 calories. Another unit of energy is the **joule** (rhymes with school), which is equal to 0.239 calories. Conversely, a calorie is 4.184 joules. The translation of calories to joules, or kilocalories to kilojoules, is so common in chemical calculations that you should memorize the conversion factors.

If a substance is heated without a change of state, the amount of heat required to change the temperature of 1 gram by 1° C is called the **specific heat capacity** of the substance. Similarly, the **molar heat capacity** is the amount of heat needed to raise the temperature of 1 mole of a substance by 1° C. Table 7-2 shows the heat capacities of several elements and compounds.

Table 7-2 Heat Capacities

Substance	***Calories per Degree per Gram***	***Calories per Degree per Mole***
$CaCO_3$	0.205	20.52
H_2O (liquid)	1.000	18.02
H_2O (solid)	0.485	8.74
MgO	0.208	8.38
Pb	0.031	6.32
Fe	0.108	6.01
Al	0.213	5.74

As an example of the use of the heat capacity values, calculate the calories required to heat 1 kilogram of aluminum from 10° C to 70° C. Multiply the grams of metal by the 60° C increase by the specific heat capacity:

$$1{,}000 \text{ grams} \times 60° \text{ C} \times 0.213 \text{ cal/deg-g} = 12{,}780 \text{ calories}$$

It therefore requires 12.78 kilocalories of energy to heat this particular piece of aluminum. Conversely, if a kilogram of the same metal cooled from 70° to 10°, 12.78 kcal of heat will be released into the environment.

You will realize that there is an abrupt change of energy when one state of matter is transformed into another. A considerable amount of energy is required to transform a low energy state to a higher energy state, like melting a solid to a liquid or vaporizing a liquid to a gas. The same quantity of energy is released upon the reverse transformation from a high energy state to a lower energy state, like condensing a gas to a liquid or freezing a liquid to a solid. Table 7-3 shows these energy values for H_2O.

Table 7-3 Heats of Transformation for H_2O

Change of State	***Associated Energy Term***	***Calories per Gram***	***Calories per Mole***	***Heat Flow***
Solid → liquid	Heat of fusion	79.8	1,436	Absorbed
Liquid → solid	Heat of crystallization	79.8	1,436	Released
Liquid → gas	Heat of vaporization	539.4	9,715	Absorbed
Gas→ liquid	Heat of condensation	539.4	9,715	Released

Bear in mind that such transformations of state are *isothermal;* that is, they take place without any change in temperature of the substance. It takes 79.8 calories to change 1 gram of ice at 0° C to 1 gram of water at 0° C; the 79.8 calories are used to rearrange the molecules from the crystalline order in the solid to the more irregular order in the liquid.

The data in the two previous tables do permit some complex calculations of energy for changes of both state and temperature. Take a mole of water vapor at 100° C and cool it to ice at 0°. The energy released, which must be removed by the refrigeration process, comes from three distinct changes listed in Table 7-4.

Table 7-4 Example of Heat Calculation

Initial	*Final*	*Calories*	*Energy Source*
Vapor @ 100° C →	Liquid @ 100° C	9,715	Heat of condensation
Liquid @ 100° C →	Liquid @ 0° C	1,802	Heat capacity of water
Liquid @ 0° C →	Solid @ 0° C	1,436	Heat of crystallization
		12,953	Total heat released

You should make sure that you understand how each of the values in the third column is obtained. For example, the 1,802 calories is the molar heat capacity of water (18.02 cal/deg) multiplied by the 100-degree change in temperature.

Notice especially that of the total heat released in this example, only 13.9% comes from lowering the temperature. Most of the heat comes from the two transformations of state—condensation and crystallization. For H_2O, the fact that the heat of condensation is almost 7 times greater than the heat of crystallization may be interpreted as meaning that the molecular description of the liquid state is much more like the solid than the gas.

> **Problem 17:** Use the data for H_2O in Table 7-2 and Table 7-3 to calculate the calories required to change 100 grams of ice at –40° C to water at 20° C.

Chapter Checkout

Q&A

1. The name of the process that converts a solid directly into a gas is called ______.

a. boiling
b. melting
c. sublimation

2. Which phase of a substance has a definite volume but an indefinite shape?

a. solid
b. liquid
c. gas

3. The point on a phase diagram at which the solid, liquid, and gaseous phases exist simultaneously is the ______.

a. critical point
b. triple point
c. heat capacity

4. Which process absorbs heat?

a. freezing a liquid to a solid
b. converting gaseous water to liquid water
c. vaporizing liquid water to gaseous water

5. If it requires 0.108 calories of heat to raise the temperature of 1.0 gram of iron 1.0° C, how many calories will be required to raise the temperature of 1.0 g of iron 15.0° C?

a. 1.62 calorie
b. 15 calories
c. 0.0072 calorie

Answers: **1.** c **2.** b **3.** b **4.** c **5.** a