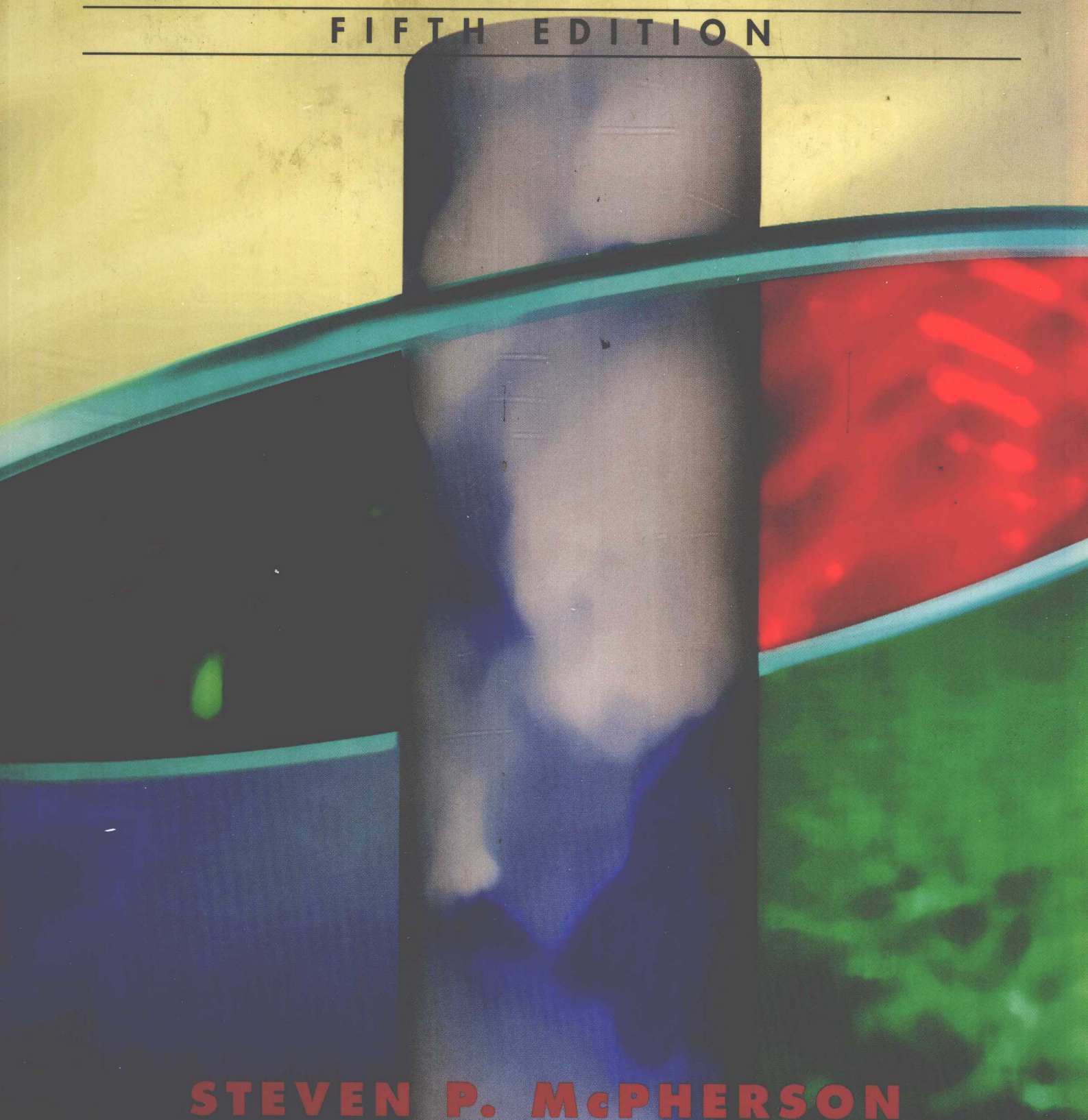


RESPIRATORY CARE EQUIPMENT

FIFTH EDITION



STEVEN P. McPHERSON

RESPIRATORY CARE EQUIPMENT

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FIFTH EDITION



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FIFTH EDITION

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Preface

McPherson's *Respiratory Care Equipment* has been and remains the most comprehensive and widely used text and reference dealing with the design, function, and application of respiratory care equipment. The fifth edition carries on this tradition of excellence and service to the respiratory care student, practitioner, and their patients.

While the overall concept of the previous editions of McPherson's *Respiratory Care Equipment* has been left intact, many changes in presentation and content have been introduced into the fifth edition. For example, each chapter now has specific learning objectives and a listing of key words and phrases in addition to the traditional chapter outline. A glossary has also been added for easy reference to these key terms. Additionally, each chapter has been reviewed and rewritten by a member of the editorial board into more current and "user-friendly" language.

The text "boxes" found in past editions have been reviewed, revised, and expanded into Key Point Highlight Boxes that are now found throughout the text. These are identified by a shaded background for ease of recognition. In addition, new summaries have been written for each chapter, and summary tables of ventilator characteristics have been developed or redesigned for each chapter discussing specific types of ventilators. Furthermore, the editorial board members have examined each figure, table, and legend for accuracy and currency. Some figures have been redrawn or updated. In many instances, tables have been modified and redesigned, and legends have been reworded for easier comprehension.

As each chapter was reviewed much old and obsolete equipment was deleted from the discussion. However, in many cases the editorial board elected to retain discussion of older devices under two circumstances. First, if the device provided a rational approach to discussion of an important principle or application; and second, if the device was judged to be "in wide use." Thus even devices not currently being manufactured and sold are included if they fit one of these categories.

The ventilator chapters (8-15), the chapter on resuscitators (6), and the chapter on bedside monitoring (7) have all seen substantial revision. Other chapters, such as those on gas regulation (3), humidifiers and nebulizers (4), and artificial airways (5) have had moderate changes introduced. Although rewritten in many places, the contents in Chapter 1 on gas physics are relatively unchanged; but Chapter 2 on primary gas systems has undergone a major restructuring. The bulk of the detailed regulatory requirements has been moved to a chapter appendix.

The editorial board, who were responsible for reviewing and rewriting specific chapters, included Christina Blazer, RRT (Community Memorial Hospital, Sidney Montana), Bethene Gregg, MS, RRT (University of Kansas), Kenneth Watson, Med, RRT (Boston Children's Hospital/Northeastern University), Douglas McIntyre, BS, RRT (St. Charles Hospital), and Paul Mathews EdS, RRT, FCCM (University of Kansas).

Paul J. Mathews, EdS, RRT, FCCM

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CHAPTER

1

Gas Physics

Chapter Outline

States of Matter
Changes of State
Sublimation
Evaporation
Gas vs Vapor
Kinetic Theory
Gas Pressure Measurement
Pressure Equivalents
Characteristics of Gases

Gas Laws
Boyle's Law
Charles' Law
Gay-Lussac's Law
Universal Gas Law
Dalton's Law
Diffusion
Fick's Law
Graham's Law

Henry's Law
Gas Flow
Bernoulli's Principle
Torricelli's Law
Venturi's Principle
Pitot Principle
Avogadro's Law
Poiseuille's Law
Reynolds' Number

Chapter Objectives

Upon completion of this chapter, the reader should be able to accomplish the following:

1. Identify the basic differences in the three states of matter of a substance.
2. Define the following terms: melting point, freezing point, and boiling point.
3. Describe the difference between sublimation and evaporation.
4. Describe the effects that temperature and pressure have on the states of matter of a substance.
5. Describe how temperature affects vapor pressure.
6. Describe the difference between a gas and a vapor.
7. Identify how each of the following measure pressure: a mercury barometer, an aneroid barometer, and a Bourdon gauge.
8. Given values for any of the following units of pressure, centimeters of water (cm H₂O), millimeters of mercury (mm Hg), kilopascals (kPa), or pounds per square inch (psi), convert to any of the other units.
9. Identify the descriptions of the following gas laws: Boyle's law, Charles' law, Gay-Lussac's law, the Universal gas law, and Dalton's law.
10. Given problems involving the use of gas laws, apply the appropriate law from those listed above and identify the law used to solve the problem.
11. Identify definitions for each of the following: Fick's law, Graham's law, and Henry's law.
12. Describe two general factors that influence gas flow.
13. Describe the relationship of velocity and pressure as it applies to Bernoulli's principle and Venturi's principle.
14. Use Avogadro's law to calculate the density of various gases.
15. Identify what happens to resistance to gas flow when each of the following is either increased or decreased: flow, radius or length of a tube, and the viscosity of a gas.
16. Use Poiseuille's law to calculate the change in resistance to gas flow when the radius of a tube is decreased by one half.
17. Describe how Reynolds' number relates to laminar and turbulent gas flow.

Key Terms

absolute zero	FiO ₂ (fraction of inspired oxygen)	PAO ₂ (partial pressure of alveolar oxygen)
aneroid barometer	freezing point	PH ₂ O (partial pressure of water vapor in the respiratory tract)
atmospheric barometric pressure	gas	pressure
atmospheric pressure	gram molecular weight (gmw)	real gas
baffle	humidity	solubility coefficient
boiling point	ideal gas	specific gravity
Bourdon gauge	inertia	sublimation
concentration gradient	jet	temperature
critical pressure	kinetic activity	van der Waal forces
critical temperature	latent heat	vapor
density	melting point	vapor pressure
diffusion	mercury barometer	viscosity
evaporation	PaCO ₂ (partial pressure of arterial carbon dioxide)	

Because respiratory therapy involves the use of gases, it is reasonable to consider some of the physics of gases and their effect on the apparatus that respiratory therapists use. In addition,

a clear grasp of the physical characteristics of gases provides a strong foundation for the understanding of the equipment used daily in the clinical setting.

States of Matter

In our natural environment there are three states of matter: solids, liquids, and gases. The state of matter of a substance largely depends on the **kinetic activity** (*motion*) of the molecules of that substance. The degree of motion is most dependent on the temperature of those molecules (Fig. 1-1).

As the temperature of the substance increases, its molecules move faster. This acceleration (average velocity \times time) causes molecules to exert more force when they hit something because of their inertia, such as when they collide with each other (**inertia** [force] = *mass* \times *velocity*¹). This increased force tends to move the molecules farther apart, causing expansion. In a closed container the gas is unable to expand freely, and the increase in motion results in increased pressure. In these cases, the additional pressure results in increased temperature caused by more frequent molecular collisions.

Changes of State

When the molecules have moved far enough apart, a change in their state will occur (solid to liquid or liquid to gas).

As molecules of a solid gain heat, they tend to expand and become liquid. As molecules in the liquid state absorb more heat, they tend to become gas.

Lord Kelvin, a professor of physics at Glasgow University in the late 1800s, calculated absolute zero.² **Absolute zero** is the calculated point where molecular kinetic activity ceases; that is, there is a cessation of random thermal motion. Based on that calculation, he devised a method of conversion between the Celsius and Kelvin

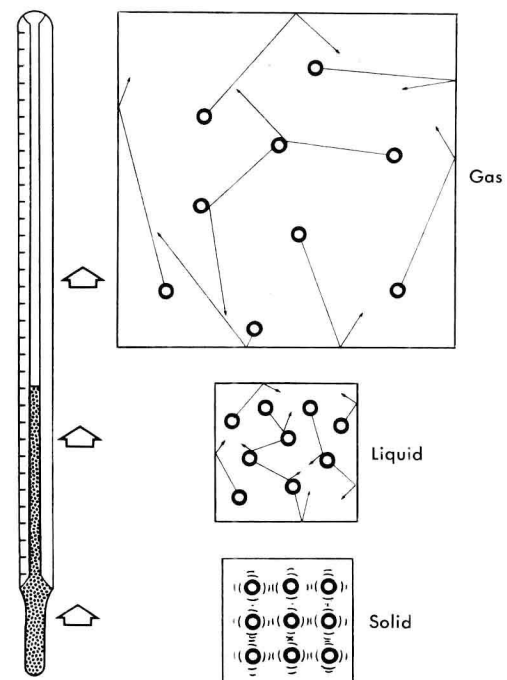


FIGURE 1-1 Increased heat causes increased molecular motion and force. This results in a tendency to expand and change to a more fluidlike state of matter. Solids will change to liquids and liquids to gases as heat and molecular motion increases.

scales by adding 273 degrees to the Celsius value ($K = 273 + ^\circ C$). See Fig. 1-2. To be exact, the Celsius scale is slightly adjusted to the former centigrade scale by 0.15° . Thus absolute zero is $-273.15^\circ C$.

A predictable succession of events occur when a substance is heated. As any substance is brought above absolute zero, its molecules begin to move at random. When a substance is still in its solid form, the molecules simply vibrate. As they gain more heat, their kinetic activity

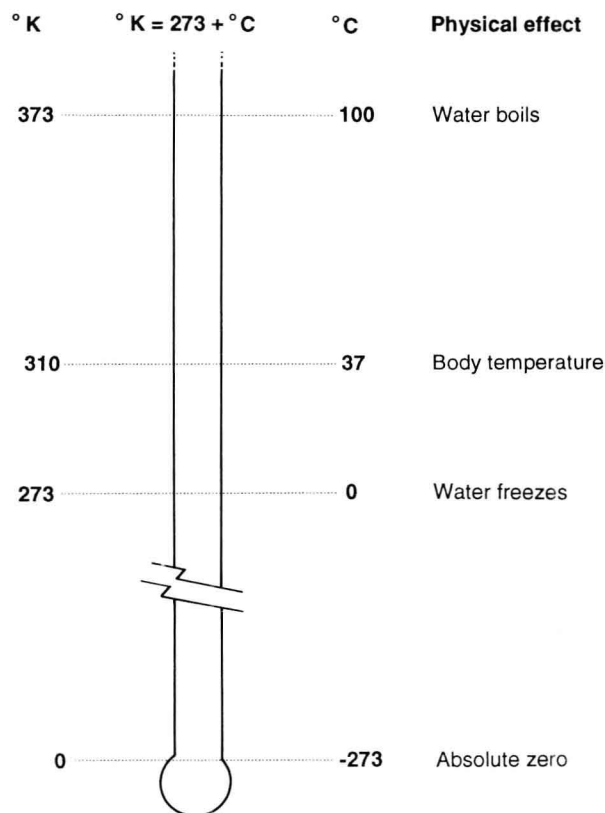


FIGURE 1-2 Comparison of Kelvin and Celsius Scales.

increases, and their vibrational energy increases proportionally. At the melting point of a solid the kinetic activity is high enough for the molecules to break free from their own mass attraction sufficiently to escape and to travel at random within the space of their containing vessel. The substance is now considered to be in the liquid state. The **melting point** is defined as *the temperature at which transition occurs from a solid to a liquid state*. The **freezing point** of that same substance occurs at the same temperature, but the transition is in reverse (Fig. 1-3).

As the temperature of the substance increases further, the molecules of the liquid gain more and more kinetic activity and then travel around more freely and exert more force. Ultimately they reach a point at which they

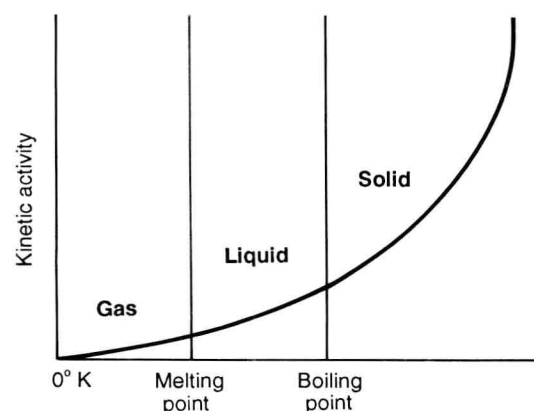


FIGURE 1-3 As heat content of molecules increases above absolute zero, they start to move. When heat is sufficient to cause enough vibrational motion that molecules break free from each other, they change to liquid state. This temperature is the *melting point*. As heat content increases further, molecules can gain enough force that they break free and travel around as individual gas molecules. At atmospheric pressure this temperature is considered to be the boiling point.

TABLE 1-1 Physical constants of medical gases

Substance	Melting point	Boiling point	Critical temperature	Critical pressure
Unit of Measure	$^{\circ}C$	$^{\circ}C$	$^{\circ}C$	psig*
Gas at room temperature				
Oxygen	-281.40	-182.96	-118.40	715.87
Nitrogen	-209.86	-195.80	-147.00	477.50
Helium	-272.20	-268.90	-267.90	18.50
Hydrogen	-259.14	-252.78	-239.90	176.10
Ethylene	-169.89	-103.03	9.44	730.30
Liquid/vapor at room temperature				
Cyclopropane	-127.20	-32.69	124.70	780.30
Carbon dioxide	-78.40*	-78.40*	31.00	1057.40
Nitrous oxide	-90.80	-88.50	36.11	1039.30
Water	0.00	100.00	374.10	5484.50

can break free from the attraction of the liquid. At this point in temperature the molecules all begin to convert to a gas (Fig. 1-3).

This temperature, at atmospheric pressure, is the substance's boiling point. The **boiling point** is defined as *that point at which a transition occurs between a liquid and gaseous state at atmospheric pressure*. Every substance has its own melting and boiling point. Table 1-1 lists some common examples of medical gases.

As substances change states of matter, there is progressively more noticeable expansion involved than is evident by merely heating the substance in the previous state. Thus the expansion that takes place when a substance changes from a liquid to a gas increases by a greater expansion factor than when that same substance converts from solid to liquid (Fig. 1-1).

Sublimation

Under certain conditions *molecules can completely bypass the liquid state turning from solid to gas*. As the heat content increases and the molecules vibrate more vigorously, they may break loose below the melting point and become free gas molecules. This process is called **sublimation** (Fig. 1-4). Solid carbon dioxide (dry ice) is the most common example of this process.

Evaporation

Latent heat is the amount of heat required to change a substance's state of matter. Only after all matter has changed form does the addition of more heat elevate the temperature of the new state of matter.

Evaporation occurs when molecules break through the surface of the liquid and convert to free gaseous molecules below the boiling point (Fig. 1-4).

This is commonly seen in simple humidifying devices. The frequency with which this occurs is directly dependent on the liquid's temperature: the higher its temperature, the more force is exerted when the molecules hit the liquid's surface and the more likely the molecules are to escape. A measure of this force is called **vapor pressure**, which is *the pressure the molecules of the liquid exert as they hit the surface and escape*. Vapor pressure increases with temperature as molecular velocity (force) increases (Fig. 1-5).

The humidity content possible for a gas increases with a rise in its temperature. The water molecules in the liquid state gain more velocity as their temperature increases, then they exert more force at the liquid's surface, and more of them escape. Therefore more *molecules of gaseous water (humidity)* are present in the gas above the water's surface.

The forces opposing the molecule's escape are (1)

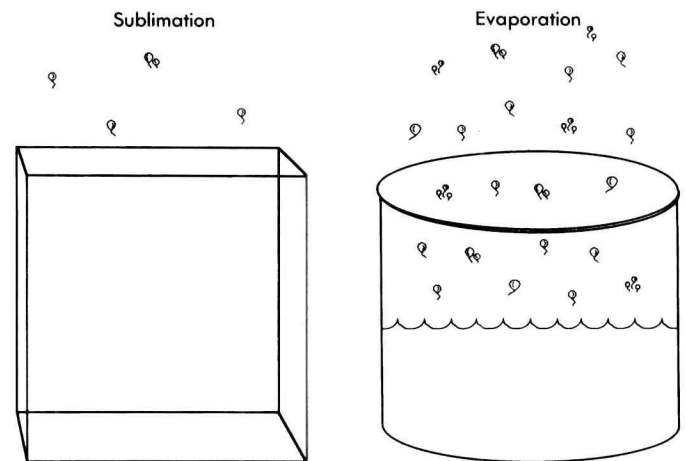


FIGURE 1-4 Occasionally molecules from a solid may vibrate loose and become free vapor molecules; this process is called *sublimation*. Molecules of liquids may become free vapor molecules below the boiling point as they gain enough energy to escape the liquid surface. This process is called *evaporation*.

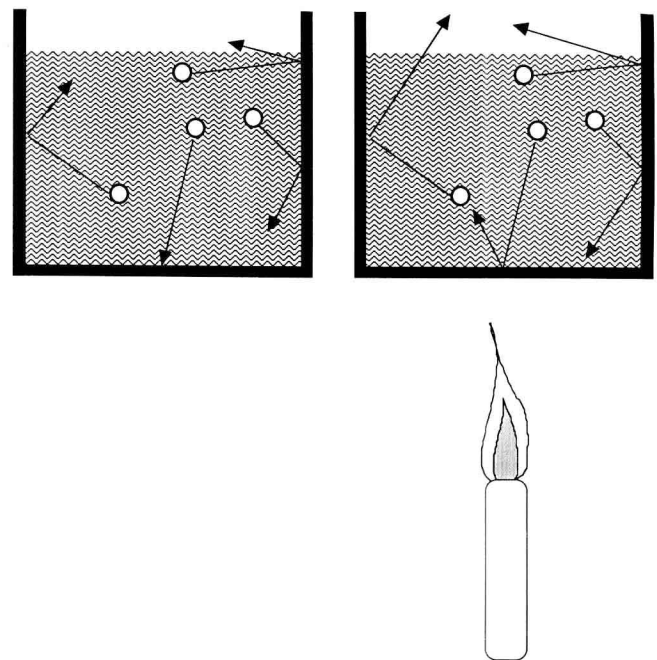


FIGURE 1-5 As liquid molecules are heated, their velocity increases. They then hit the surface more often and with more inertia, which increases their force and the frequency of their escape as vapor.

mass attraction of the molecules for each other (Van der Waal forces) and (2) the pressure of the gas above the liquid's surface. We have seen how a rise in temperature increases the velocity and force of the molecules hitting each other and moves them farther apart. The mass attraction is then decreased because of the increased distance between the molecules (see Fig. 1-1). Therefore the molecules escape more easily and more frequently with a temperature rise. In evaporation, if the gas pres-

sure above the liquid decreases, there is less force opposing the escape of molecules, and they will escape more easily (Fig. 1-5). This concept explains why the boiling point of water, for example, is lower on Mt. Everest than at sea level; this occurs because the vapor pressure is lower, and the force (pressure) opposing the escape of molecules is decreased.

If a liquid can be kept in a closed container, the force of the molecules trying to escape from the liquid will eventually reach an equilibrium with the water vapor pressure and no more liquid molecules will escape (Fig. 1-6). In other words, the force of the molecules in the liquid is not sufficient to break loose from an increased opposing force resulting from the increased vapor pressure. Now one water vapor molecule must accidentally hit the surface of the water and return to a liquid state before another liquid molecule can escape to become a gas or vapor molecule. The forces of molecular mass attraction plus gas pressure and vapor pressure are in equilibrium. However, as the temperature of the liquid increases, the increased force of its molecules reduces mass attraction. The vapor pressure, in turn, increases as a higher opposing force is necessary to equilibrate the molecules' escape from the liquid state. At its boiling point the force of the molecules in the liquid equals the surrounding pressure (atmospheric), and they may all escape. So, in essence,

boiling point is that temperature at which the force exerted by the molecules of the liquid trying to escape equals the forces opposing escape (atmospheric pressure and mass attraction).

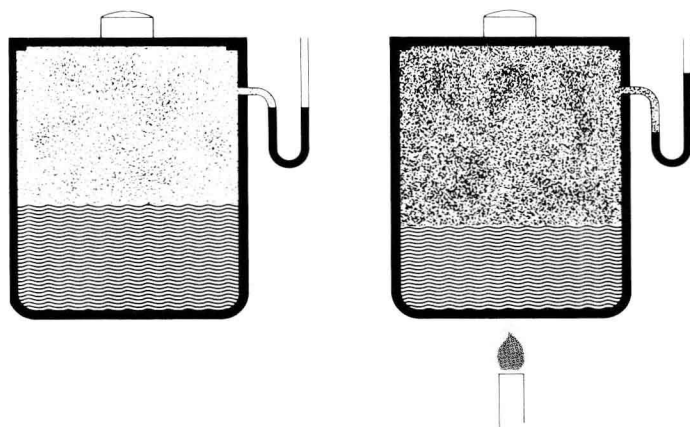


FIGURE 1-6 Increased temperature causes liquid molecules to escape (evaporate) with more force. Once this force is counterbalanced by the gas and vapor pressure forces above the liquid, surface equilibrium exists and no more liquid molecules can escape until some of the gas molecules return to the liquid state. This pressure increase is related to the vapor pressure of the force with which liquid molecules escape to gaseous state.

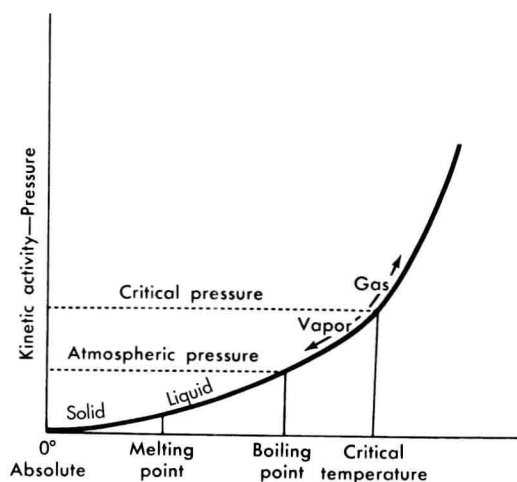


FIGURE 1-7 Example of the effects of changes of temperature and pressure on states of matter.

As gas molecules are heated above the boiling point and their force increases, the force (pressure) required to convert them back to a liquid also increases (Fig. 1-7).

Ultimately, a point can be reached at which the gaseous molecules cannot be converted back to a liquid, no matter what pressure is exerted on them. That point would be considered the **critical temperature**. At a temperature between the critical temperature and the *boiling point* a pressure can be exerted to overcome the kinetic activity of the gas molecules sufficiently to convert them back into the liquid state. As the temperature of a substance rises above its boiling point toward its critical temperature, the pressure necessary to change it to a liquid increases proportionally. The *pressure required to convert a gas back into a liquid at its critical temperature* is called its **critical pressure**. After the gas passes its critical temperature, no amount of pressure is sufficient to convert it back into a liquid.

A review of Table 1-1 will show that some of the substances that are common to the respiratory care practitioner are **vapors**—they *can be liquified under pressure*. Water is probably the most common example. Other examples of vapors are carbon dioxide and nitrous oxide. Both of these substances can be converted to liquid at room temperature if enough pressure is applied. In fact, both gases are supplied to hospitals in pressurized cylinders where most of the vapor is converted to liquid.

Cylinders containing substances below their critical temperature (liquids) are weighed to determine the amount remaining in the cylinder as opposed to reading the pressure level within the cylinder.

Oxygen, nitrogen, and helium are examples of substances above critical temperature (gas) whose remaining cylinder contents can be determined by gas pressure.