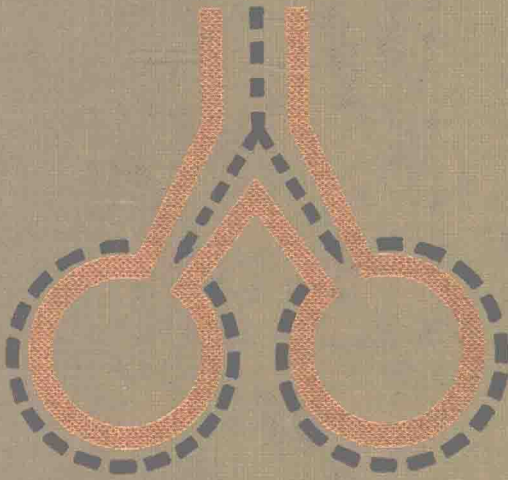


FUNDAMENTALS OF RESPIRATORY THERAPY



DONALD F. EGAN

Third edition

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Third edition

with 227 illustrations

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PREFACE

Since publication of the second edition of *Fundamentals of Respiratory Therapy*, three events have occurred that are of great importance to respiratory therapists and therapy. First, in April of 1973 the standards of hospital respiratory therapy services, as proposed by the Joint Commission on Accreditation of Hospitals, were accepted and are now in operation. Brief though they are and lacking substance in many sections, the standards nevertheless represent a large step toward protecting the patient with pulmonary problems and giving much needed support to respiratory therapy and its technical and medical personnel. If hospitals are to maintain high accreditation, respiratory therapy must be recognized and accepted as a legitimate health source, not just a counterbalance to other hospital services that are fiscal liabilities. We all know that no dramatic overnight change will take place, and it will be a few years yet before the full objectives of the standards are realized. Still, overcoming inertia and making a start is a major achievement, and it is reassuring to know that some people of influence in our field have foresight and the willingness to act.

Second, perhaps the most stabilizing event in the recent history of respiratory therapy was the birth of the National Board for Respiratory Therapy. Certainly, housing the therapist and technician credentialing procedures under one roof makes the greatest of sense. It is most impressive that, despite the misunderstandings, suspicions, charges, and countercharges attending the emergence of the original technician certification program, the concept of the NBRT was able to flourish on reason and not founder on emotionalism. A mature move that should be easily recognized as such, creation of the NBRT will give added professionalism and credibility to both personnel levels involved. Credit is due those who worked to bring the Board to fruition, as many of us are aware of the obstacles with which they contended. Still, this is the kind of action that counts and that identifies a structure built on a strong foundation.

Third, and more representative of a trend than of a specific event is the growing criticism of the scientific validity of much of respiratory therapy. This is a topic deserving of an in-depth scholarly essay, and I recognize that a superficial treatment may tend to belittle its importance. It is important—very important. Many of the complaints against respiratory therapy are more than justified. I can recall no other facet of medicine that has so permeated hospital patient care, involved such critically ill

patients, subjected these patients to physiological alterations as significant, and done this with such a positive economic impact, all with as meager a scientific background, as has respiratory therapy. This does not mean that all respiratory therapy is invalid or that much of it is not well founded, but we do know that many of the things we do are empirical or based on the somewhat nebulous support of "clinical observation." Many of the users are convinced of the value of respiratory therapy, but it remains for us who are responsible for it to demonstrate why it is effective or to delete those procedures, if any, that are no more useful than no therapy at all.

These developments were noted as a means of stressing that with growth and maturity comes increased responsibility. The stature of the respiratory therapist in the hospital should be enhanced, as his services now count toward his hospital's accreditation standing. His registration or certification will have more prestige as hospital co-workers and administrators become better acquainted with the credentialing mechanism. Accordingly, the registered therapist and the certified technician will be expected to demonstrate a more professional attitude. In the face of current criticism, the therapist's deportment must be above reproach, and his technical expertise confidence-inspiring.

It was with these considerations in mind, and as an attempt to meet their implied needs, that a host of notes, sketches, observations, and opinions were brought together, hopefully in harmony, to produce this third edition of *Fundamentals of Respiratory Therapy*. The objective of the work was to compact as much currently relevant essential material about respiratory therapy as possible into one book. On this third go-around, I have affirmed that while writing a text is very hard work, attempting to keep it updated is no easier. Respiratory therapy continues to grow, but I do not believe at quite the frenetic pace of the mid-sixties to early seventies. At the same time, literature pertinent to respiratory therapy has been sounding greater depths, slowly diminishing the wrench-and-bolt dominance of the recent past. This burdens an author with an increasing load of reading, reviewing, and judgmental weighing of material to decide what to remove from the old text and what to include in the new.

In this edition the section on basic sciences has been augmented, especially with a review of some aspects of chemistry with which it is felt the respiratory therapist should be familiar. In no way are the few pages devoted to chemistry intended to supplant a standard text or a well-taught course on the subject. Many therapists and students, however, are afraid of chemistry, perhaps from a past experience in which they found no relevance between it and their needs. I would hope that some of these might be stimulated to pursue its study further and reap the rich harvest of a good background of scientific knowledge. Concepts of some of the currently popular techniques such as positive end-expiratory pressure and intermittent mandatory ventilation have been explained and illustrated and a miscellany of comments made and opinions expressed on preclinical and clinical topics. For example, the abbreviation for gram has been changed from *gm* to *g*, conforming to standards described in *Units of Weights and Measure*, by L. J. Chisholm of the U.S. Department of Commerce/National Bureau of Standards, Miscellaneous Publication 286, 1967.

Reference is also made to terminology changes recommended by a special joint Committee of the American College of Chest Physicians and the American Thoracic Society.³² These should be of interest to all in the field.

Some changes from the previous edition may involve only a sentence or two, but each was made for a specific purpose, and it is felt that this edition is just about as current as is possible at this time in basics and essentials. Like its predecessors, the third edition is directed primarily toward the respiratory therapy student and the therapist and technician reviewing for credentialing examinations. It should also be of help to nurses involved in respiratory care, to house staff and pulmonary fellows, and last, but not least, to the nonpulmonary specialist attending physician who wishes to be sure that his patients with respiratory diseases are properly managed. To all, happy reading.

Donald F. Egan

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Chapter 1

GASES, THE ATMOSPHERE, AND THE GAS LAWS

A gas, which may not be seen, cannot be felt, and has no inherent confining boundaries, almost invokes an impression of nothingness until we become aware of the tremendous activity of its components and its great flexibility. It can be compressed, can expand, can produce heat, can cool, and can be liquefied. From its behavior, we can infer that a gas consists of much empty space, with its substance in minute particle form, and it is with the general relationship between such space and particles that we will be concerned.

MOBILITY OF GASES

As a start, let us recall that all matter is composed of *atoms*, the characteristics of which differentiate the many elements that comprise our world. In chemical reactions it is atoms that combine with like or different atoms, separate from yet others, or re-group into new combinations. Compounds are made from such joining of atoms, the resulting products possessing characteristics distinctly different from those of their constituents.

The smallest particle of a substance that retains all of its properties is referred to as a *molecule*. This may consist of one or several atoms, but at least as many as are represented in the chemical formula of the substance. The term *gram molecular weight (gmw)* is frequently encountered in chemistry, and it is the weight in grams equal to the molecular weight of a substance, or the sum of all the atomic weights in its molecular formula. It must be noted, however, that not all substances have true molecular structures, depending upon the type of bonding which holds together the atoms. Those classified as *ionic compounds* are essentially mixtures of fixed proportions of minute electrically charged particles called ions, without any basic molecular structure. Atomic bonding and further details of ions will be discussed in Chapter 2. Because the expression, gram molecular weight, is inappropriate for ionic compounds, we speak of *gram formula weight (gfw)*, which has the same meaning and is calculated in the same manner as gram molecular weight. It is proper to use either molecular weight or formula weight for substances that form molecules, but only formula weight is acceptable for ionic compounds.¹

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It is possible to differentiate between molecular and ionic substances by inspection of their formulas, but it is not our purpose to delve into such details of chemistry. Our interest is only to make use of those principles of chemistry that relate to our needs of this book. Actually, the problem is partially solved for us, since gases, which are our chief concern, do consist of molecules. Throughout the text, then, when referring to the structure or chemistry of gases, we will use the terms *molecules* and *gram molecular weights* as needed, but for all other substances we will use the unit of *gram formula weight*.

Thus the substance of gases consists of aggregates of small molecular particles that are in constant motion, called *kinetic activity*. The density with which atoms and molecules are packed determines the solidity or fluidity of all matter. Substances with a high density of particles are solids, while those with easily mobile particles are fluids. Gases and liquids fall into this second category. It should not be supposed that the atoms or molecules of solids cannot move, for they respond to such stimuli as vibrations, bending, stretching, and temperature. Fluidity, however, implies the easy flowing of particles over one another and the ability to conform readily to the confines of a container. Mercury, although a metal, can thus be described as fluid.

Particles that make up the substance of gases are minute in size, within the range of 10^{-8} to 10^{-7} cm in diameter, with weights varying from 10^{-23} to 10^{-20} g. The kinetic theory tells us that these particles are in *constant rapid* motion, following completely random paths, and that their speed is phenomenal. Hydrogen particles move 1.84×10^5 cm/sec (greater than 1 mile per second), and oxygen 4.6×10^4 cm/sec (one third of a mile per second).² During this intense activity the particles "collide" with one another and with the surface of enclosing containers. Actually, the particles probably repulse one another before physical contact, but the forces involved can best be visualized in the mind as collisions. The average number of collisions per second for each molecule of hydrogen is 1×10^{10} , for oxygen 4.6×10^9 , and for carbon dioxide 6.2×10^9 . The *mean free path* of gas molecules describes the average distance traveled by the molecules between collisions. Again, for hydrogen this distance is 1.66×10^{-5} cm, for oxygen 8.8×10^{-6} cm, and for carbon dioxide 5.8×10^{-6} cm.³ Very fine particles of an insoluble substance such as carbon or metal dust suspended in water, if viewed under a microscope, can be seen to move about in an erratic random manner. This is called *Brownian movement* and is produced by the kinetic activity of water molecules striking the suspended material.

To view the phenomenon of kinetic activity in familiar quantitative terms, let us imagine oxygen molecules in a pure sample of that gas to be the size of Ping-Pong balls. We can see them, in the mind's eye, in large numbers bouncing off walls, ceiling, floor, and each other, never stopping and never settling to the floor. Considering the relative sizes of oxygen molecules and Ping-Pong balls, the mean free path, and the average number of collisions of the molecules, the Ping-Pong balls would travel an average distance of 40 feet between collisions. This gives us some concept of the great distance between molecules in relation to their size as well as the mass of "nothingness" that makes up a gas.

PRESSURE AND TEMPERATURE OF GASES

All gases exert pressure, whether free in the atmosphere, enclosed in a container, or dissolved in a liquid such as blood. In physiology this pressure is frequently referred to as the *tension* of a gas. Gas pressure is dependent upon molecular kinetic activity and is the result of molecular bombardment upon any confining surface, be it a steel cylinder or the earth's surface, and we may consider such pressure as the striking force of molecules attempting to escape. In addition, the force of the earth's gravity, by its effect on the molecular masses of the gas, augments the gas pressure against the dependent confining surface of a gas volume. Thus, in a container of gas, although the travel of molecules is random in all directions, pressure in the bottom of the vessel is somewhat higher than elsewhere, as the force of molecular impingement is aided by gravity. The amount of pressure exerted by a gas depends on the *number* of particles present and the *frequency* of their collisions. The frequency, in turn, is related to the *velocity* of the gas particles, since the greater the speed of travel, the greater will be the number of collisions per unit of time, the greater the force of collisions, and the greater the gas tension.

Gas particle velocity is not a constant value but is directly related to gas temperature, and as temperature rises the kinetic activity accelerates, molecular collisions increase in number, and the pressure of the gas rises. Conversely, with dropping temperature molecular activity declines, particle velocity and collision frequency drop, and pressure is lowered. The familiar increase in automobile tire pressure while driving on a hot pavement is an example of the relationship between gas tension and temperature. This relationship between molecular activity and temperature can be graphically illustrated by a special temperature scale, which the student will put to practical use when he studies the gas laws. We will describe the concept of *absolute temperature* and the two subscales by which it is calibrated. There is a temperature at which all molecular activity ceases, a theoretical value arrived at by projection and calculation, which, although closely approximated, has not actually been attained. If we are interested in the relative kinetic behavior of gases at various temperatures, a point of no activity produces a logical zero on which to build a scale. This is called *absolute zero* (0°_{abs}) and is the origin of the absolute temperature scale. If it is calibrated in Celsius temperature units, it is called the *Kelvin scale* (K), and if in Fahrenheit units, the *Rankine scale* (R).

Kelvin scale

In Celsius units, molecular activity stops at about -273°C . Therefore $0^{\circ}\text{K} = -273^{\circ}\text{C}$, and $0^{\circ}\text{C} = 273^{\circ}\text{K}$, since 0°C is 273 temperature units above 0°K . When used as symbols in formulas, Celsius temperatures are often designated by a small t and absolute temperatures by a capital T or a capital K . A simple equation to keep in mind is $^{\circ}\text{K} = ^{\circ}\text{C} + 273$. In other words, to convert Celsius degrees to Kelvin, add 273. Thus:

$$\begin{aligned} 25^{\circ}\text{C} &= 25 + 273 = 298^{\circ}\text{K} \\ 37^{\circ}\text{C} &= 37 + 273 = 310^{\circ}\text{K} \\ -15^{\circ}\text{C} &= -15 + 273 = 258^{\circ}\text{K} \end{aligned}$$

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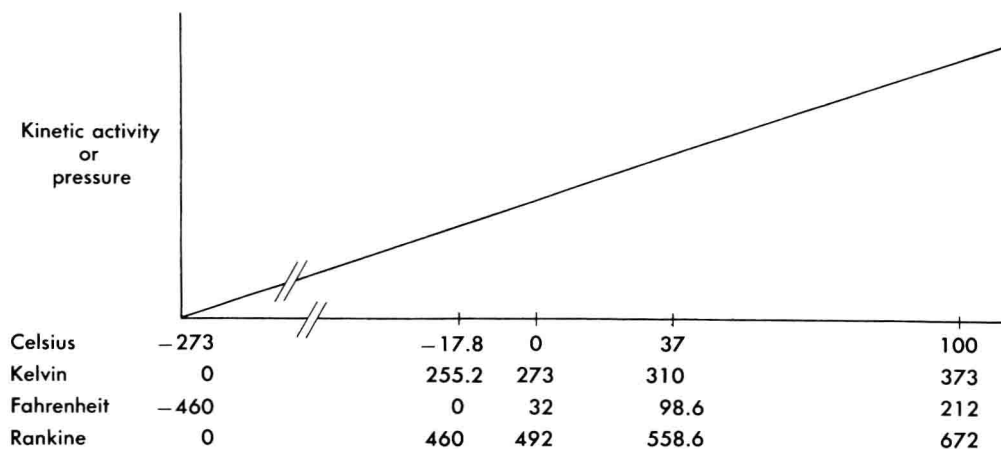


Fig. 1-1. Linear relationship between gas molecular activity, or pressure, and temperature. Comparable readings of the four scales are indicated for five temperature points.

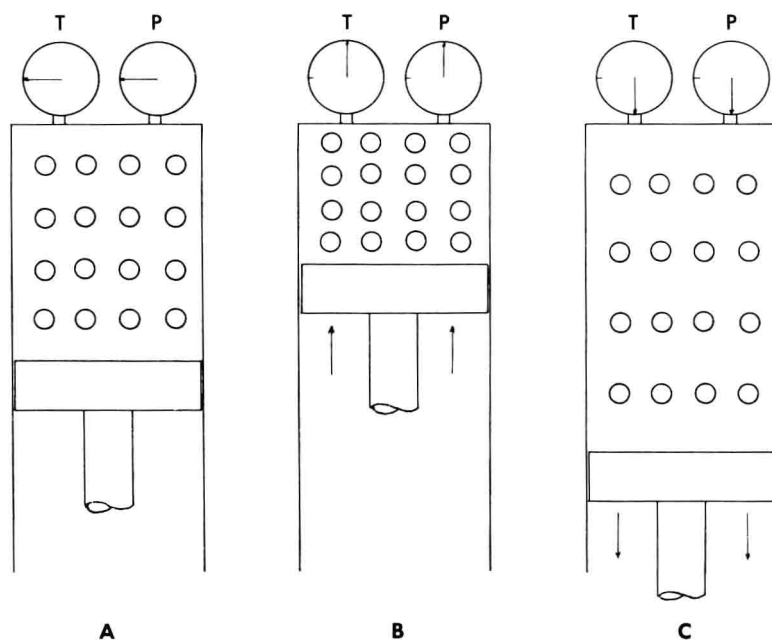


Fig. 1-2. A mass of gas in the resting state exerts a given pressure at a given temperature, in cylinder A. In B, as the piston compresses the gas, the molecules are crowded closer together, and the increased energy of molecular collisions is reflected in a rise of both temperature and pressure. Conversely, retraction of the piston in C allows the gas to expand, and the temperature and pressure drop as molecular interaction decreases.

Rankine scale

Used frequently in engineering but rarely in medical science, the Rankine scale is based on Fahrenheit units. Since $-273^{\circ}\text{C} = -460^{\circ}\text{F}$ (refer to formulas to convert between Celsius and Fahrenheit scales), then $0^{\circ}\text{R} = -460^{\circ}\text{F}$, and $^{\circ}\text{R} = ^{\circ}\text{F} + 460$. Fig. 1-1 is a scalar representation of the relation between gaseous kinetic activity, or pressure, and five commonly used temperatures of the four related scales.

Because of the relatively great distances between molecules, gases possess the quality of *compressibility*. When pressure is exerted on a gas, the molecules can be brought closer together as their intervening spaces are narrowed. Conversely, if the container of a volume of gas enlarges, the gas therein *expands* to accommodate the new volume, and its molecules range further apart. Fig. 1-2 illustrates the relationship between compression and expansion of a given mass of gas molecules and corresponding temperature and pressure changes. Because the tremendous energy of molecular collision is expended as heat, compression of a gas produces *heat* as well as a buildup of pressure among the molecules. As compression brings the molecules closer together, the frequency of collisions increases, and both heat and pressure increase. Thus the heat of compression may be considered as a means of dissipating the great increase in kinetic energy that accompanies compression. It should be apparent that expansion of a gas produces a drop in temperature as molecular collision frequency decreases. Cooling of expansion is utilized in refrigerating systems and is part of the natural phenomenon of cooling through expansion of large air masses.

MOLAR VOLUME OF GASES

One of the major principles of physics and chemistry is *Avogadro's law* (Count Amadeo Avogadro, 1776-1856), which tells us that equal volumes of all gases at the same temperature and pressure contain the same number of molecules or, conversely, that at constant temperature and pressure equal numbers of molecules of all gases occupy the same volume. Further, it has been established that the weights of all atoms in grams corresponding to their atomic weights, the weights of all molecules in grams corresponding to their molecular weights, and the weights of ions of all non-molecular compounds in grams corresponding to the formula weights of the compounds always contain the same number of their respective particles, 6.02×10^{23} . This is known as *Avogadro's number*. Although these quantities are often referred to as "gram atomic weights," "gram molecular weights," and "gram formula weights," they are each technically known as a *mole*. To put it another way, any quantity of matter that contains 6.02×10^{23} atoms, molecules, or ions is called a mole. Later we will discuss physiologically active substances in concentrations so small that it will be more convenient to refer to them in terms of *thousandths of a mole*, or *millimole* (*mM*). Just as moles are the weights of substances in grams equal to atomic weights (*gaw*), molecular weights (*gmw*), or formula weights (*gfw*), so are millimoles the weights of substances in milligrams equal to atomic weights (*mgaw*), molecular weights (*mgmw*), or formula weights (*mgfw*).

The volume occupied by one mole of a gas (1 gram molecular weight; 6.02×10^{23} molecules) is the *molar volume*. It is important because it will allow us to calculate

Table 1-1. *Molar volume of selected gases under standard conditions**

<i>Gas</i>	<i>Symbol</i>	<i>Molar volume in liters</i>
"Ideal gas"		22.414
Ammonia	NH ₃	22.094
Carbon dioxide	CO ₂	22.262
Carbon monoxide	CO	22.402
Chlorine	Cl ₂	22.063
Helium	He	22.426
Hydrogen	H ₂	22.430
Hydrogen chloride	HCl	22.248
Nitrogen	N ₂	22.402
Oxygen	O ₂	22.393
Sulfur dioxide	SO ₂	21.888

*Modified from Pimental, G. C., editor: *Chemistry, an experimental science*, San Francisco, 1963, W. H. Freeman Co., Publishers.

densities of gases and gas mixtures and to convert values for dissolved gases from volumes percent to moles per liter, a topic to be considered in some detail later. It is customary, for the sake of uniformity in comparing values, to measure molar volumes under what are termed *standard conditions*, a temperature of 0° C and an ambient pressure of 1 atmosphere. This is an artificial situation, however, since we do not ordinarily handle gases in the difficult environment of zero degree Celsius. When molar volumes are measured under common ambient conditions, then calculated to what they would be at standard (a technique known as "correcting a gas volume," which we will discuss shortly), we find that the natural intermolecular behavior of each gas causes its molar volume to deviate from the universal low pressure value. Table 1-1 compares molar volumes of several gases corrected from ambient to standard or measured directly at standard conditions.

If the student refers to the subject of molar gas volumes in the average chemistry textbook, he will probably find only the value 22.4 *ℓ*/mole given for all gases. Actually, for most purposes this is usually adequate, and we see in Table 1-1 that it is representative of the molar volumes of oxygen, nitrogen, and carbon monoxide, gases with which we will be concerned as we continue our study. Even carbon dioxide, a most important respiratory gas, has a value rounded off to 22.3 *ℓ*/mole, close to the universal normal. As a summary and general statement let us say that under standard conditions of a temperature of 0° C and a pressure of 1 atm, moles of all gases occupy 22.4 *liters*. In this text the only exception will be found in a calculation involving carbon dioxide (p. 136), where the molar volume used for that gas is the more accurate one of 22.3 *liters*.

DENSITY OF GASES

Before considering the density of gases specifically, we will need some definitions. We must learn the relationship between the widely used physical terms *mass*, *weight*, and *density*. The word *mass* refers to the substance of an object, the quantity of matter it contains, and the number and nature of its molecules, and is characterized by

Table 1-2. *Examples of gas densities (D) under standard conditions*

$D_{O_2} = \frac{\text{gmw}}{22.4} = \frac{32}{22.4} = 1.43 \text{ g/}\ell$	$D_{He} = \frac{\text{gmw}}{22.4} = \frac{4}{22.4} = 0.1785 \text{ g/}\ell$
$D_{N_2} = \frac{\text{gmw}}{22.4} = \frac{28}{22.4} = 1.25 \text{ g/}\ell$	$D_{CO_2} = \frac{\text{gmw}}{22.4} = \frac{44}{22.4} = 1.965 \text{ g/}\ell$

having inertia and by being subject to the pull of gravity. *Weight* is the gravitational pull of the earth upon a body; thus the greater the mass, the greater the weight. Mass is thereby proportional to weight and is measured in arbitrary units as weight, against such standards as the kilogram and the pound. It should be noted that weight varies with the position of mass relative to the surface of the earth, decreasing both toward the earth's center and away from its surface. The *inertia* of a body, that quality of mass which requires force to start it in motion from a resting state or to change its velocity once in motion, remains unchanged no matter where it is located.

Density may be defined as the amount of mass per unit volume of a body, the concentration of its molecules, and is usually employed as *weight density*. Density is thus the *weight of a body per unit volume* and in our field of interest is most often described in grams per cubic centimeter for solids and liquids, and grams per liter for gases. Other units, such as pounds per cubic foot, can also be used. A mass weighing 15 g and measuring 3 cc has a density of 5 g/cc. A ton of feathers and a ton of bricks weigh the same, but the obvious difference in the volumes of similar weights of these substances makes for widely differing densities. *Specific gravity* is a variation of density measurement whereby the density of solids and liquids is calibrated against the density of water used as a standard of unity, and gases against oxygen or hydrogen. A liquid with specific gravity of 1.5, for example, has a density half again as great as that of water. Specific gravity values of gases play a negligible role in pulmonary physiology, whereas gas densities are of great importance.

Since density equals weight divided by volume, the density of any gas can readily be calculated by dividing its *gram molecular weight* by the universal molar volume of 22.4 liters. The quotient is expressed as *grams per liter*. Examples of gas densities are shown in Table 1-2.

Densities of gas mixtures are easily calculated if the percentage composition of the mixture is known. Given the following mixed gases:

Gas A = 10%

Gas B = 60%

Gas C = 30%

$$D = \frac{(0.10 \times \text{gmw A}) + (0.60 \times \text{gmw B}) + (0.30 \times \text{gmw C})}{22.4}$$

Calculate the density of a mixture of 30% hydrogen bromide (HBr) and 70% ethylene (C_2H_4):

$$D = \frac{(0.3 \times 81) + (0.7 \times 28)}{22.4} = \frac{43.9}{22.4} = 1.955 \text{ g/}\ell$$

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Exercise 1-1. Calculate densities of the following, rounding off at three decimal places:

- | | |
|--|--|
| (a) C_2H_2 (acetylene) | (f) 5% CO_2 + 95% O_2 |
| (b) NH_3 (ammonia) | (g) 80% He + 20% O_2 |
| (c) SiF_4 (silicon fluoride) | (h) 70% He + 30% O_2 |
| (d) CO (carbon monoxide) | (i) 25% CH_4 + 75% C_4H_{10} |
| (e) SO_2 (sulfur dioxide) | (j) 3% SO_2 + 15% N_2 + 82% O_2 |

COMPOSITION OF THE ATMOSPHERE

The atmosphere upon which we, as oxygen-breathing creatures, are completely dependent is a mixture of many gases plus water vapor. The elements composing the atmosphere, with the exception of water vapor, which will be discussed separately later, have the approximate concentrations shown in Table 1-3.

The atmosphere is divided into two major segments, three subsegments, and several layers, each with certain physical and/or chemical properties,⁴ outlined in Table 1-4:

1. The first major segment is the *inner atmosphere*, extending from the earth's surface to an altitude of about 600 miles; it is composed of the following sub-segments called spheres:
 - a. The *troposphere* extends from the earth's surface to an outer border called the tropopause, an average distance of some 8 miles up but varying with the latitude of the earth. It is higher over the equator than over the poles.

Table 1-3. Approximate composition of the atmosphere

Element	Percent	
Nitrogen (N_2)	78.08	99.99%
Oxygen (O_2)	20.95	
Argon (Ar)	0.93	
Carbon dioxide (CO_2)	0.03	
Neon (Ne)	1.8×10^{-3}	0.01%
Helium (He)	5.0×10^{-4}	
Krypton (Kr)	1.0×10^{-4}	
Hydrogen (H_2)	1.0×10^{-4}	
Xenon (Xe)	1.0×10^{-5}	
Ozone (O_3)	1.0×10^{-5}	
Radon (Rn)	6.0×10^{-18}	

Table 1-4. Summary of atmospheric divisions

Atmosphere	Strata	Approximate height in miles
Free space		Above 1200
Outer	Exosphere	600-1200
	Ionosphere	50-600
	Stratosphere	8-50
Inner	Troposphere	0-8

- The troposphere is characterized by decreasing temperatures with altitude, reaching a low of approximately -55°C (-67°F), and has much turbulence.
- b. The *stratosphere* continues from 8 to about 50 miles above the earth. The first layer of the stratosphere, from 8 to 15 miles up, is one of constant temperature around -55°C (-67°F) and has little turbulence. The next layer, from 15 to 30 miles, shows an increase in temperature, reaching a high of 10°C (50°F). The third and last layer of the stratosphere, from 30 to 50 miles, has a sharp temperature drop to -72°C (-100°F), and there is a return of turbulence. At approximately 50 miles of altitude the stratopause separates the stratosphere from the next sphere.
 - c. The *ionosphere* reaches from a distance of 50 miles outward to a distance of 600 miles. Here there are several layers of ions, resulting from photochemical reactions between solar ultraviolet radiation and atmospheric molecules. The ionosphere is important as a reflector for the electromagnetic waves of radio communication. Temperatures in this sphere soar up to 2000°C (3600°F), but because the density of the air molecules is so low in this region, such temperatures have little meaning in our usual concept of temperature. As with the other spheres, a boundary called the ionopause delineates the end of the ionosphere.
2. The second major segment is the outer atmosphere, which is also called the *exosphere*. This region extends from the 600-mile limit to about 1200 miles from earth, where it blends with the vacuum of *free space*. It is a marginal area where molecular collisions become progressively more rare.

The gravitational pull of the earth on atmospheric gas molecules produces the greatest density of molecules close to its surface, a density that decreases steadily outward to the vacuum of free space. It is speculated, however, that despite decreasing density, the percentage composition of the atmosphere, as described earlier, remains fairly constant to a height of at least some *60 miles*. Beyond this limit, with a decrease in mass air movement to keep the gases well mixed, there is a separation of the elements on the basis of their molecular weights. This phenomenon, called diffusion separation, disrupts the composition of the air as we know it on earth.

MEASUREMENT OF AIR PRESSURE

In cardiopulmonary physiology and therapy of cardiopulmonary diseases, we are constantly dealing with the principles of gas pressure, and it is vitally important that the student clearly understand this aspect of gas physics. Pressure, in any context, is defined as a *force* applied to a specific *surface area*. For our purposes such force is usually expressed as grams per square centimeter (g/cm^2) or pounds per square inch (lb/in^2) (psi). The force exerted by gases is a result of their kinetic molecular bombardment already discussed, and in a mixture of gases such as air this force is the sum of molecular activity of all the constituent gases. If we visualize the atmospheric mantle enveloping the earth, described above, we can understand that the molecular activity of atmospheric gases will exert a force against the surface of the earth. To view it another way, we can see that the many miles of atmosphere rest upon the earth as an