Physics of Diving

2

SECTION PAGE

Physics of Diving

2

2.0 GENERAL

In all diving operations, safety is the primary consideration. One key to safety is a clear understanding of the physics of diving. Physics is the field of science dealing with matter and energy and their interactions. This chapter explores physical laws and principles that pertain to the diving environment and its influence on the diver. Gravity is passive, vision and hearing may be misleading, color perception changes at varying depth, and breathing dynamics are ever changing. The principles of physics provide the keystone for understanding the reasons for employing various diving procedures and the operation of associated equipment. Many of these principles receive further elaboration in other sections of the NOAA Diving Manual.

2.1 PRESSURE

Pressure is force acting on a unit area. Stated mathematically,

Pressure = force/area P = F/A

In the United States, pressure is typically measured in pounds per square inch (psi). Under water, two kinds of pressure affect a person, the weight of the surrounding water and the weight of the atmosphere over that water. One concept that must be remembered at all times is: a diver, at any depth, must be in pressure balance with the forces at that depth.

At all depths, the diver must compensate for the pressure exerted by the atmosphere, by the water, and by the gases being used for breathing under water. This compensation must always be thought of in terms of attaining and maintaining a balance between the pressure inside the body and the external pressure.

2.1.1 Atmospheric Pressure

Atmospheric pressure is the pressure exerted by the earth's atmosphere; it decreases with altitude above sea level. At sea level, atmospheric pressure is equal to 14.7 pounds per square inch (psi) or one atmosphere (atm). The higher the altitude above sea level, the lower the atmospheric pressure. For example, at 18,000 ft. (5,486 m), atmospheric pressure is 7.35 psi, or half that at sea level (see Figure 2.1). At sea level, atmospheric pressure is considered constant and universal; that is, anywhere on the earth at sea level, the pressure is 14.7 psi. The pressure inside a person's lungs is the same as the pressure outside.

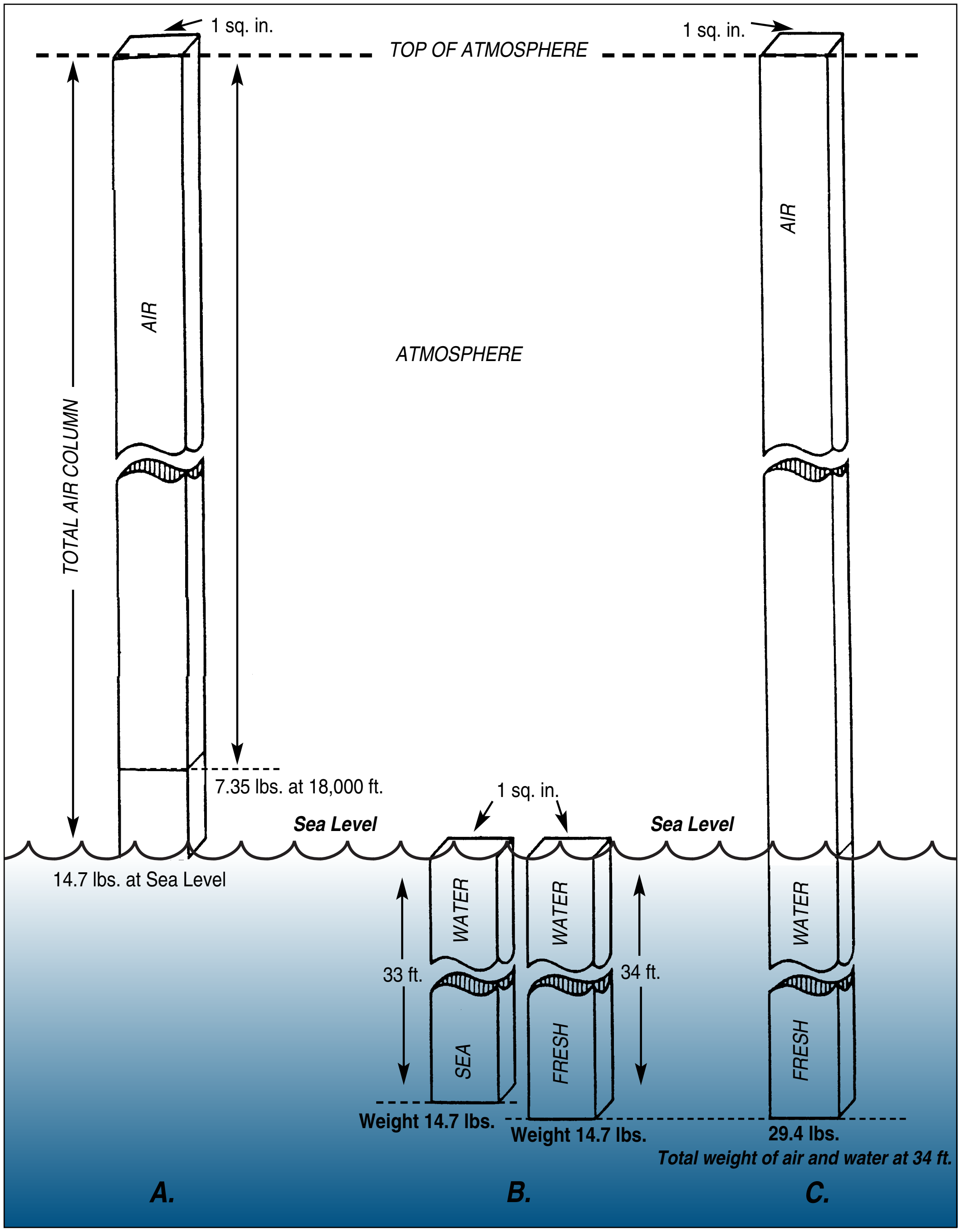
2.1.2 Hydrostatic Pressure

Pressure due to the weight of water is called "hydrostatic pressure." The weight of water is cumulative; the deeper the dive, the more water there is above the diver and the greater the weight of that water. This weight affects a diver from all sides equally and increases at a rate of 0.445 psi per foot of seawater. Thus, at a depth of 33 ft. (10.1 m) of seawater (fsw), the hydrostatic pressure is 14.7 psi, or one atmosphere, the same pressure as atmospheric pressure at sea level. In freshwater, 34 ft. (10.4 m) equals 14.7 psi or 0.432 psi per foot of freshwater (ffw). Thereafter, for every 34 ft. of additional depth in freshwater, the hydrostatic pressure increases by one atmosphere (see Figure 2.1).

2.1.3 Absolute Pressure

The sum of atmospheric pressure plus hydrostatic pressure is called the "absolute pressure." Absolute pressure can be expressed in many ways, including "pounds per square inch absolute" **(psia)**, "atmospheres absolute" **(ata)**, feet of seawater absolute **(fswa)**, feet of freshwater absolute **(ffwa)**, or millimeters of mercury absolute **(mmHga)**.

To understand the effects of absolute pressure on a diver, consider this: the feet of a 6-foot tall man standing under water will be exposed to pressure that is almost three pounds per square inch greater than that exerted at his head.



A. **A one square inch column of air extending from sea level to the top of the atmosphere weighs 14.7 lbs. One half of the weight is contained in the first 18,000 ft. (5,486 m or 3 1/2 miles) of the column.**

B. **A one-inch square column of seawater 33 ft. (10.1 m) deep and a column of freshwater 34 ft.(10.4 m) deep each weigh 14.7 lbs.**

C. **At a depth of 34 ft. (10.4 m) of freshwater, the sum of atmospheric and hydrostatic pressures equal 29.4 lbs.**

FIGURE 2.1
Weight of Air and Water

2.1.4 Gauge Pressure

The difference between atmospheric pressure and the pressure being measured is “gauge pressure.” Consider the pressure gauge on a scuba tank, for instance. The zero reading on the gauge before it is attached actually represents the ambient atmospheric pressure. To put it another way, at sea level, the zero on the tank gauge actually represents 14.7 psia. Thus, the pressure in the tank is referred to in terms of "pounds per square inch gauge" (psig). To convert gauge pressure to absolute pressure, add 14.7.

2.1.5 Partial Pressure

In a mixture of gases, the proportion of the total pressure contributed by each gas in the mixture is called the “partial pressure.” Although traces of other gases are also present in atmospheric air, for our discussion here, we can approximate that atmospheric air is composed of 21% oxygen and 79% nitrogen, for a total of 100%, or one atmosphere absolute. The impact of partial pressures upon the diver is explained in detail later in this chapter under Dalton's Law.

The body can function normally only when the pressure difference between the inside of the body and the outside is very small.

2.2 DENSITY

Density can be defined as weight per unit volume. Expressed mathematically,

Density = Weight/Volume or D = W/V

Density is expressed in pounds per cubic foot (lbs/ft^3) or in grams per cubic centimeter (g/cm^3).

Gas density is related to absolute pressure. As depth increases, the density of the breathing gas increases and becomes heavier per unit volume. High gas density increases the effort required to breathe and limits a diver's ability to ventilate the lungs adequately, especially during strenuous exercise and at deeper depths (see Table 2.1).

Freshwater has a density of 62.4 pounds per cubic foot. Seawater has a density of 64.0 pounds per cubic foot (see Figure 2.2). As a result, freshwater floats on top of seawater and a diver floats easier in seawater than in freshwater.

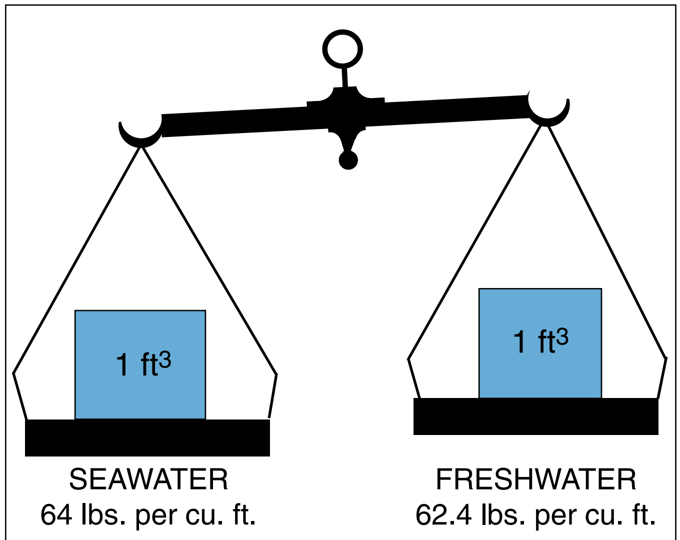


FIGURE 2.2
Seawater and Freshwater Density

TABLE 2.1
Pressure Chart

Depth	Pressure		Gas Volume	Gas Density
Sea Level	14.7 psia	1 ata	1ft^3	1x
33 feet	29.4 psia	2 ata	1/2 ft^3	2×
66 feet	44.1 psia	3 ata	1/3 ft^3	3×
99 feet	58.8 psia	4 ata	1/4 ft^3	4×
132 feet	73.5 psia	5 ata	1/5 ft^3	5×
165 feet	88.2 psia	6 ata	1/6 ft^3	6×
297 feet	147.0 psia	10 ata	1/10 ft^3	10×

1. **Pressure of each atmosphere is equal to approximately 15 psi, i.e., at three atmospheres of pressure it is approximately 45 psi, at six atmospheres it is approximately 90 psi, etc.**

2. **Gas volume is inversely proportional to the depth in atmospheres absolute (ata), i.e., any gas volume at four ata is one-fourth of the sea level volume; at six ata it is one-sixth, etc.**

3. **Gas density is directly proportional to the pressure in atmospheres absolute (ata), i.e., when a gas mixture at sea level is taken to two atmospheres absolute, each gas in the mixture is twice as dense; at three atmospheres absolute it is three times as dense, etc.**

psia = pounds per square inch absolute
ata = atmospheres absolute

2.2.1 Specific Gravity

Specific gravity is the ratio of the weight of a given volume of a substance (density) to that of an equal volume of another substance (water [for liquids and solids] and air [for gases] are used as standards). Water has a specific gravity of 1.0 at 39.2°F (4C). Substances that are more dense than freshwater have a specific gravity greater than 1.0. Thus, the specific gravity of seawater is 64.0/62.4 = 1.026.

2.3 WATER

Physical laws that act upon a person above the surface of water also apply below the surface. As a diver descends into the water, those forces increase; the diver should be aware of these effects.

2.3.1 Freshwater

Water, H_2O, is a major constituent of all living matter. It is an odorless, tasteless, very slightly compressible liquid oxide of hydrogen, which freezes at 32°F (0C), and boils at 212°F (100C). In its purest form, water is a poor conductor of electricity.

2.3.2 Seawater

Seawater contains just about every substance known. Sodium chloride (common table salt) is the most abundant chemical. Because of its components, seawater is a good conductor of electricity.

2.3.3 pH

The pH of an aqueous solution expresses the level of acids or alkalis present. The pH of a liquid can range from 0 (strongly acidic) to 14 (strongly alkaline), with a value of seven representing neutrality. The pH balance in blood signals to the brain the need to breathe. Too much carbon dioxide in the blood causes the pH level in the blood to change, making it more acidic. One of the ways the body can reduce the acidity of the blood is to increase ventilation, which reduces the CO_2 level and thus reduces the acidity. The importance of pH in diving is covered in Chapter 3, Diving Physiology.

2.4 UNITS OF MEASUREMENT

How much air do we have? How deep are we? How much longer can we stay on the bottom? Divers must have a common system of communicating the answers to these questions.

There are two systems for specifying force, length, and time: the English System and the International System of Units (SI), also known as the Metric System. The English System is based on the pound, the foot, and the second, and is widely used in the United States. The International System of Units is used virtually everywhere else, and is based on the kilogram, the meter, and the second. Every diver will eventually encounter the International System of Units and should be able to convert units of measurement from one system to the other (see Tables 2.2, 2.3, and 2.4).

2.4.1 Length

The principle SI unit of length is the meter (39.37 inches). Smaller lengths are measured in centimeters (cm) or millimeters (mm). Greater lengths are measured in kilometers (km).

$$39.37 \text{ in} \times 1 \text{ ft}/12 \text{ in} = 3.28 \text{ ft} = 1 \text{ m}$$

Example: Convert 10 feet to meters.
Solution: 10 ft × 1 m/3.28 ft = 3.05 m

Example: Convert 10 meters to feet.
Solution: 10 m × 3.28 ft/1 m = 32.8 ft

2.4.2 Area

In both the English and International System of Units (SI), area is expressed as a length squared. For example, a room that is 12 feet by 10 feet would have an area that is 120 square feet (12 ft x 10 ft).

2.4.3 Volume

Volume is expressed in units of length cubed. Using the room example from paragraph 2.4.2 but adding a third dimension—an eight-foot ceiling would result in a volume of 960 cubic feet (120 ft^2 x 8 ft). The English System, in addition to using cubic feet, uses other units of volume such as gallons. The International System of Units (SI) uses the liter (1). A liter equals 1000 cubic centimeters (cm^3) or 0.001 cubic meters (m^3), which is one milliliter (ml).

2.4.4 Weight

The pound is the standard measure of weight in the English System. The kilogram is the standard measure of weight in the International System of Units. One liter of water at 4C weighs one kilogram or almost 2.2 lbs.

$$1 \text{ liter(l)} = 1 \text{ kg} = 2.2 \text{ lbs.}$$

Example: Convert 180 pounds to kilograms.
Solution: 180 lbs × 1 kg/2.2 lbs = 81.8 kg

Example: Convert 82 kilograms to pounds.
Solution: 82 kg × 2.2 lbs/1 kg = 180.4 lbs

2.5 TEMPERATURE

Body temperature is a measure of the heat retained in the human body. Heat is associated with the motion of molecules. The more rapidly the molecules move, the higher the temperature.

Temperature is usually measured either with the Fahrenheit **(°F)** scale or with the Celsius, or Centigrade, **(C)** scale.

Temperatures must be converted to absolute when the gas laws are used. The absolute temperature scales, which use Rankine **(R)** or Kelvin **(K)**, are based upon the absolute zero (the lowest temperature that could possibly be reached) (see Figure 2.3). Note that the degree symbol (°) is used only with Fahrenheit temperatures.

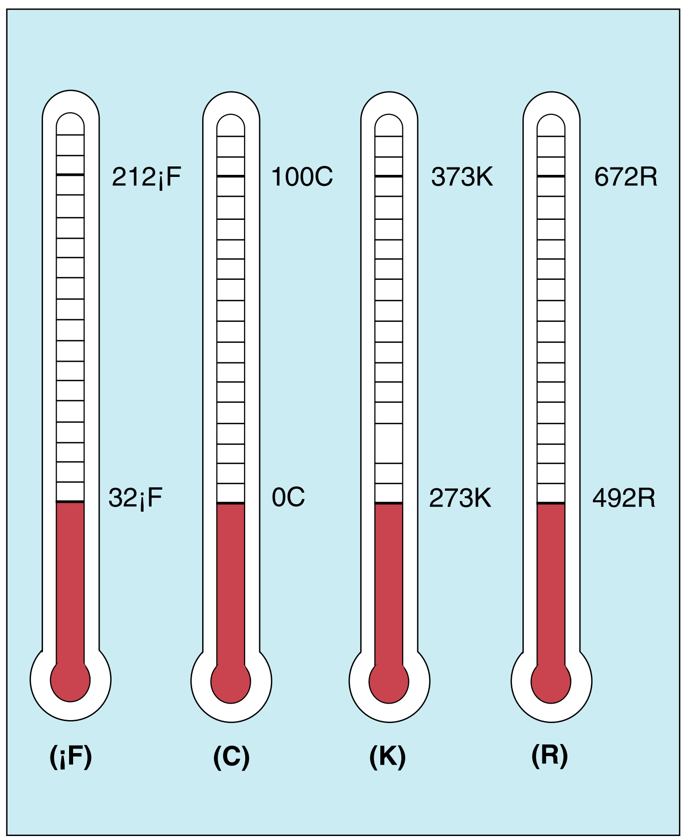


FIGURE 2.3
Freezing and Boiling Points of Water

TABLE 2.2
Conversion Factors, Metric to English Units

To Convert From SI Units	To English Units	Multiply By
PRESSURE		
1 gm/cm^2	inches of freshwater	0.394
1kg/cm^2	pounds/square inch (psi)	14.22
1kg/cm^2	feet of freshwater (ffw)	32.8
1kg/cm^2	inches of mercury (in. Hg)	28.96
1 cm Hg	pounds/square inch	0.193
1 cm Hg	foot of freshwater (ffw)	0.447
1 cm Hg	foot of seawater (fsw)	0.434
1 cm Hg	inches of mercury	0.394
1 cm of freshwater	inch of freshwater	0.394
VOLUME AND CAPACITY		
1 cc or ml	cubic inch (in^3)	0.061
1 m^3	cubic feet (ft^3)	35.31
1 liter	cubic inches	61.02
1 liter	cubic foot	0.035
1 liter	fluid ounces (fl oz)	33.81
1 liter	quarts (qt)	1.057

To Convert From SI Units	To English Units	Multiply By
WEIGHT		
1 gram	ounce (oz)	0.035
1 kg	ounces	35.27
1 kg	pounds (lb)	2.205
LENGTH		
1 cm	inch	0.394
1 meter	inches	39.37
1 meter	feet	3.28
1 km	mile	0.621
AREA		
1 cm^2	square inch	0.155
1 m^2	square feet	10.76
1 km^2	square mile	0.386

TABLE 2.3
Conversion Table for Barometric Pressure Units

		atm	N/m^2 or Pa	bars	mb	kg/cm^2	gm/cm$_2$ (cm H_2O)	mm Hg	in. Hg	lb/in^2 (psi)
1 atmosphere	=	1	1.013X10^5	1.013	1013	1.033	1033	760	29.92	14.70
1 Newton (N)/m$_2$ or Pascal (Pa)	=	.9869X10^{-5}	1	10^{-5}	.01	1.02X10^{-5}	.0102	.0075	.2953X10^{-3}	.1451X10^{-3}
1 bar	=	.9869	10^5	1	1000	1.02	1020	750.1	29.53	14.51
1 millibar (mb)	=	.9869X10^{-3}	100	.001	1	.00102	1.02	.7501	.02953	.01451
1 kg/cm^2	=	.9681	.9807X10^5	.9807	980.7	1	1000	735	28.94	14.22
1 gm/cm^2 (1 cm H_2O)	=	968.1	98.07	.9807X10^{-3}	.9807	.001	1	.735	.02894	.01422
1 mm Hg	=	.001316	133.3	.001333	1.333	.00136	1.36	1	.03937	.01934
1 in. Hg	=	.0334	3386	.03386	33.86	.03453	34.53	25.4	1	.4910
1 lb/in^2 (psi)	=	.06804	6895	.06895	68.95	.0703	70.3	51.70	2.035	1

TABLE 2.4
Barometric Pressure Conversions

Units	psig	psia	atm	ata	fsw	fswa	ffw	ffwa
psig →		Add 14.7	Divide 14.7	Add 14.7, Divide 14.7	Divide .445	Divide .445 Add 33	Divide .432	Divide .432 Add 34
psia →	Minus 14.7		Minus 14.7 Divide 14.7	Divide 14.7	Minus 14.7 Divide .445	Divide .445	Minus 14.7 Divide .432	Divide .432
atm →	Times 14.7	Times 14.7 Add 14.7		Add 1	Times 33	Times 33 Add 33	Times 34	Times 34 Add 34
ata →	Minus 1 Times 14.7	Times 14.7	Minus 1		Times 33 Minus 33	Times 33	Times 34 Minus 34	Times 34
fsw →	Times .445	Times .445 Add 14.7	Divide 33	Add 33 Divide 33		Add 33	Times 1.03	Times 1.03 Add 34
fswa →	Minus 33 Times .445	Times .445	Minus 33 Divide 33	Divide 33	Minus 33		Minus 33 Times 1.03	Times 1.03
ffw →	Times .432	Times .432 Add 14.7	Divide 34	Add 34 Divide 34	Times .97	Add 34 Times .97		Add 34
ffwa →	Minus 34 Times .432	Times .432	Minus 34 Divide 34	Divide 34	Minus 34 Times .97	Times .97	Minus 34	

Either of the absolute temperature scales, Rankine or Kelvin, may be used in gas law calculations.

To convert from Fahrenheit to absolute temperature Rankine, use the following equation:

$$°F + 460 = R$$

To convert from Celsius, or Centigrade, to absolute temperature Kelvin, use the following equation:

$$C + 273 = K$$

The Fahrenheit (°F) and Rankine (R) temperature scales are used in the English System. The Celsius (C) and Kelvin (K) temperature scales are used in the International System of Units. The Celsius and Fahrenheit scales are based on the temperature of melting ice as 0C (32°F) and the temperature of boiling water as 100C (212°F).

To convert from Fahrenheit to Celsius, use the following equation:

$$C = 5/9 \times (°F - 32) \quad \text{or} \quad C = .56 \times (°F - 32)$$

To convert from Celsius to Fahrenheit, use the following equation:

$$°F = (9/5 \times C) + 32 \quad \text{or} \quad °F = (1.8 \times C) + 32$$

2.5.1 Heat

Water temperature is an important consideration in all diving operations. Human beings function effectively within a narrow range of internal temperatures, becoming chilled when the water temperature drops below 75°F (23.9C) and overheated when body temperature rises above 98.6°F (37C). Below that temperature, body heat loss occurs faster than it can be replaced. A person who has become chilled cannot work efficiently or think clearly and may be more susceptible to decompression sickness.

A cellular neoprene wet suit loses a portion of its insulating property as depth increases and the material is compressed (see Figure 2.4). As a consequence, it is often necessary to employ a thicker suit, a dry suit, or a hot water suit to compensate for extended exposures to cold water.

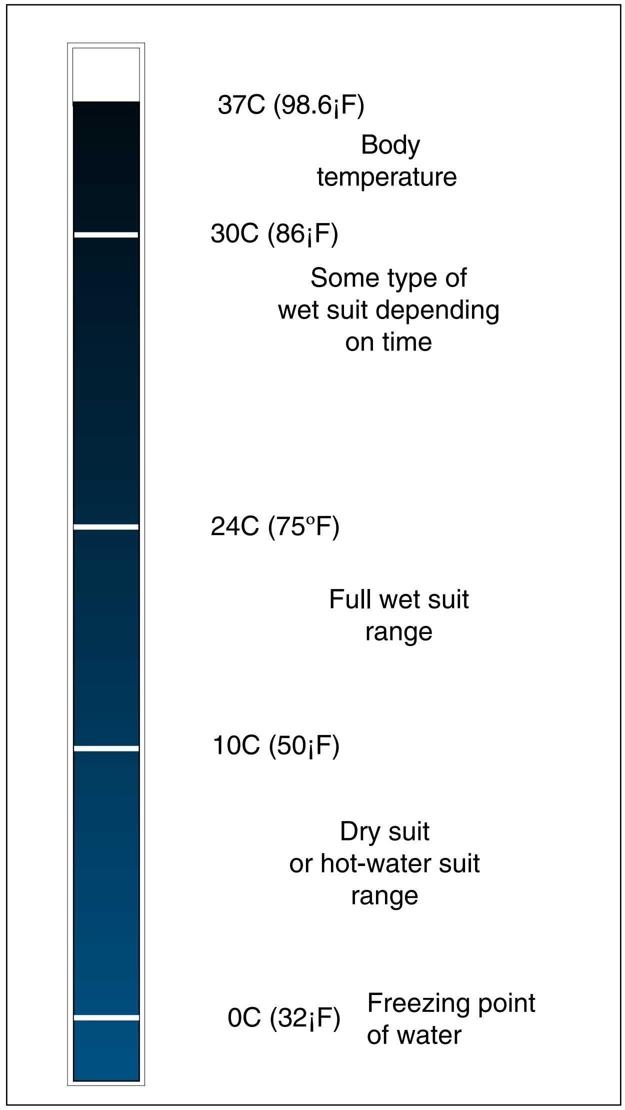


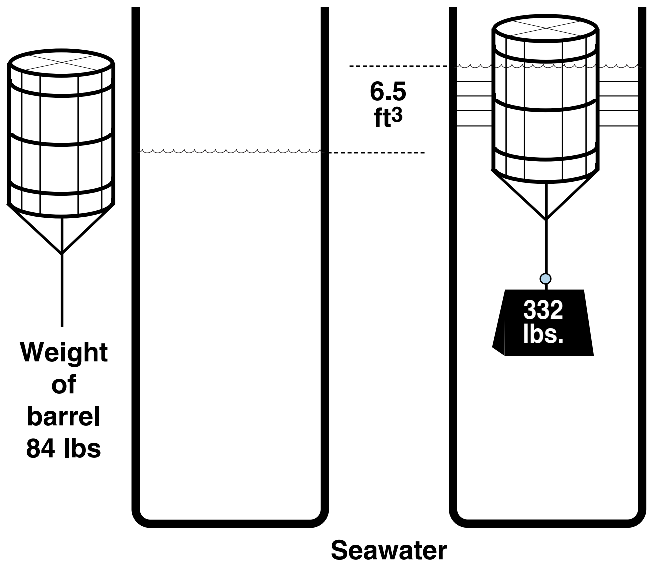
FIGURE 2.4
Recommended Thermal Protection

2.6 BUOYANCY (Archimedes' Principle)

A Greek mathematician named Archimedes determined why things float 2000 years ago. He established that ***"Any object wholly or partly immersed in a fluid is buoyed up by a force equal to the weight of the fluid displaced by the object."*** This explains why a steel ship floats, but its anchor does not. The more water displaced, the greater the buoyancy (see Figure 2.5).

If the weight of the displaced water (total displacement) is greater than the weight of the submerged body, the buoyancy is **positive** and the object floats. If the weight of the displaced water is less than the weight of the object, then the buoyancy is **negative** and the object sinks. If the weight of the object is equal to the weight of the displaced water, then buoyancy is **neutral** and the object is suspended. Neutral buoyancy is the state frequently used when diving.

Buoyancy is dependent upon the density of the surrounding liquid. Seawater has a density of 64.0 pounds per cubic foot, compared to 62.4 pounds per cubic foot for freshwater. Therefore, each cubic foot of seawater that is displaced by a volume of air in a container has a lifting force of 64 pounds. The greater the density, the greater the buoyancy force. Thus, it is easier to float in seawater than in a freshwater lake.



The partially immersed 84 lb. barrel with 332 lbs. attached has displaced 6.5 ft^3 of seawater.

6.5 ft^3 $\times$ 64 lbs/ft^3 = 416 lbs

416 lbs Ð 84 lbs = 332 lbs positive buoyancy

FIGURE 2.5
ArchimedesÕ Principle

An understanding of buoyancy serves the diver in a number of ways. By using weights, by expanding the air in a buoyancy compensator, or by increasing the size of a variable-volume diving suit, a diver can manipulate his buoyancy to meet operational needs. When working on the bottom, for example, a slightly negative buoyancy provides better traction and more stability on the sea floor. Buoyancy is also an invaluable aid to lifting heavy items in salvage operations.

2.7 GASES USED IN DIVING

Air breathed on the surface (atmospheric air) is also the most common gas breathed under water. Gases react in specific ways to the effects of pressure, volume, and temperature.

2.7.1 Atmospheric Air

The components of dry atmospheric air are given in Table 2.5. Depending upon the location and weather conditions, atmospheric air may also contain industrial pollutants. The most common pollutant is carbon monoxide, often present around the exhaust outlets of internal-combustion engines. Diving safety is jeopardized if pollutants are not filtered from compressed air prior to diving.

TABLE 2.5
Components of Dry Atmospheric Air

	Concentration	
Component	Percent by Volume	Parts per Million (ppm)
Nitrogen	78.084	
Oxygen	20.946	
Argon	0.934	
Carbon Dioxide	0.033	
Rare Gases	0.003	30.00*
Neon		18.18*
Helium		5.24*
Carbon Monoxide		2.36*
Methane		2.0 *
Krypton		1.14*
Hydrogen		0.5 *
Nitrous Oxide		0.5 *
Xenon		0.08*

*Approximate

Besides atmospheric air, divers use various mixtures of oxygen, nitrogen, and helium. In some diving applications, special mixtures of one or more of the gases may be blended with oxygen. The physiological effects of each gas alone, or in combination with other gases, must be taken into account to insure that no harm is done to body organs and functions. The so-called "inert" gases breathed from the atmosphere, or those in gas mixtures we breathe when diving, serve only to dilute and mix with oxygen.

2.7.2 Oxygen (O_2)

Oxygen is the most important of all gases and is one of the most abundant elements on earth. Fire cannot burn without oxygen and people cannot survive without oxygen. Atmospheric air contains approximately 21% oxygen, which exists freely in a diatomic state (two atoms paired off to make one molecule). This colorless, odorless, tasteless, and active gas readily combines with other elements. From the air we breathe, only oxygen is actually used by the body. The other 79% of the air serves to dilute the oxygen. Pure 100% oxygen is often used for breathing in hospitals, aircraft, and hyperbaric medical treatment facilities. Sometimes 100% oxygen is used in shallow diving and in certain phases of diving. For storage in saturation and for deeper diving the percentage may be less; in fact, it may be too low to be safely breathed at sea level. Mixtures low in oxygen require special labeling and handling to ensure that they are not breathed unintentionally. Breathing a mixture with no oxygen will result in unconsciousness, brain damage, and death within a few minutes. Besides its essential metabolic role, oxygen is fundamental to decompression. Still, the gas can also be toxic. Breathing pure oxygen under pressure may affect the central nervous system in short exposures, and various other parts of the body and nervous system, particularly the lungs, from longer exposures.

Oxygen is the agent responsible for most oxidation that takes place on this planet. The gas itself does not bum, nor does it explode. In order for combustion to take place there has to be oxygen, fuel, and a source of ignition. Material burns more vigorously in an oxygen-enriched environment, and a great deal faster and more intensely in pure oxygen. With several methods for mixing gases, it is necessary to handle pure oxygen appropriately.

Oxygen comes in three basic grades: aviator's oxygen (Grade A), medical/industrial oxygen (Grade B or USP grade [Medical Grade]), and technical oxygen (Grade C). Aviator's oxygen is ultra-dry in order to prevent freezing of regulators, but otherwise it is the same as medical oxygen. Technical (welding) oxygen comes from the same source as medical, but the containers may not be evacuated prior to filling and may contain pre-existing contaminants and objectionable odors; accordingly, it is not recommended for diving. However, if this is the only oxygen available in a case where decompression sickness needs to be treated, it is far better to use technical oxygen than not to breathe oxygen when it is needed. Significant contamination in oxygen is quite rare. Oxygen cylinders should never be completely emptied, but should be maintained with a minimum of 25 psi cylinder pressure to prevent contamination from entering the cylinder.

In the United States, oxygen is shipped in gas cylinders that are color-coded green. This is the only gas for which there is a uniform color-coding, and this green color applies only in the U.S. Color-coding should never be relied upon to make positive identification of the gas in any cylinder. If the oxygen is Grade A or B, the label on the cylinder should clearly state this; Grade C may have no identification other than that it is oxygen.

2.7.3 Nitrogen (N_2)

Nitrogen gas, which forms the largest proportion of the air we breathe, is also taken through the lungs into the blood stream, but it plays no part in metabolism. In breathing air under pressure it is the nitrogen portion that plays the major role in decompression. For diving, nitrogen may be used to dilute oxygen, but it has several disadvantages compared with other gases. When breathed at increased partial pressures, typically at least 100 ft. (31m) or deeper, it has a distinct anesthetic effect, producing an intoxicated state characterized by a loss of judgment and disorientation called *nitrogen narcosis*.

2.7.4 Helium (He)

Helium is used extensively in deep diving as a diluting agent for oxygen. Helium has a lower density than nitrogen and it does not cause the same problems associated with nitrogen narcosis. A lower-density gas, in deep diving, reduces breathing resistance (see Section 2.2).

However, helium does have several disadvantages. Breathing helium-oxygen mixtures impairs voice communication when diving by creating the "Donald Duck" voice. Helium also has a high thermal conductivity, which can cause rapid heat loss from the body, especially via the respiratory system.

2.7.5 Carbon Dioxide (CO_2)

Carbon dioxide is a natural by-product of metabolism that is eliminated from the body by breathing. Although generally not considered poisonous, excessive carbon dioxide can cause unconsciousness that can be fatal for divers. The two major concerns with carbon dioxide are control of the quantity in the breathing supply, and its removal after exhalation. Elevated carbon dioxide levels may further predispose a diver to nitrogen narcosis, oxygen toxicity, and decompression sickness.

Carbon dioxide is considered biologically "active" since it directly influences the pH level of blood. In addition, recent advances in medicine indicate that carbon dioxide may be involved chemically in changes with dilation of blood vessels.

With divers using closed or semi-closed breathing systems, it is absolutely essential to remove carbon dioxide from the breathing gas.

2.7.6 Carbon Monoxide (CO)

Carbon monoxide is a poisonous gas which interferes with the blood's ability to carry oxygen. Because it is colorless, odorless, and tasteless, it is difficult to detect and it acts as a cellular poison. It is produced by the incomplete combustion of fuels and is most commonly found in the exhaust of internal-combustion engines, and by overheated oil-lubricated compressors.

The usual carbon monoxide problem for divers is contamination of the air supply because the compressor intake is too close to the compressor-motor exhaust. The exhaust gases, including carbon monoxide, are sucked in with the air at the intake of the compressor. The effects can be lethal.

2.7.7 Argon (Ar), Neon (Ne), Hydrogen (H_2)

Argon, neon, and hydrogen have been used experimentally to dilute oxygen in breathing-gas mixtures, but normally these gases are not used in diving operations. Argon has narcotic properties and a density that make it inappropriate for use as a breathing gas. However, it is frequently used for inflation of variable-volume dry suits for warmth, because its higher density reduces the conduction of heat.

Neon causes less voice distortion than helium and has lower thermal conductivity. As a breathing gas, however, neon is expensive and causes increased breathing resistance at moderate or heavy workloads.

Hydrogen has two important advantages as a breathing gas: it is readily available and it produces less breathing resistance at depth than other gases. However, the explosive properties of hydrogen are a significant disadvantage.

2.8 GAS LAWS

Definitions have been provided of terms, units of measurement, and the properties of the gases divers use under water. What follows are various physical laws that directly and indirectly affect underwater activity.

Gases are subject to three interrelated factors: pressure, volume, and temperature. A change in one results in a measurable change in the others. This is true whether we are dealing with a pure gas or with a gas mixture. The relationships among these three factors have been defined as the gas laws.

A diver needs a basic understanding of the gas laws. As a diver moves up and down the water column, pressure changes affect the air in dive equipment and in the diver's lungs. It is also essential to determine whether the air compressor on deck has the capacity to deliver an adequate supply of air to a proposed operating depth, and to be able to interpret the reading on the depth gauge of the pneumofathometer hose as conditions of temperature and depth vary.

2.8.1 Boyle's Law

"For any gas at a constant temperature, the volume of the gas will vary inversely with the pressure." If an inverted bucket is filled with air at the surface where the pressure is one atmosphere (14.7 psi), and then taken under water to a depth of 33 fsw (10.1 msw), or two atmospheres (29.4 psi), it will be only half full of air. Any compressible air space, whether it is in a diver's body or in a flexible container, will change its volume during descent and ascent. Ear and sinus clearing, diving mask volume, changes in buoyancy, functioning of a scuba regulator, descent or ascent, air consumption, decompression—all are governed by Boyle's Law (see Figure 2.6).

Examples of Boyle's Law

An open-bottom diving bell with a volume of 24 cubic feet is lowered into the water from a surface support ship. No air is supplied to or lost from the bell. Calculate the volume of the air space in the bell at depths of 33, 66, and 99 fsw (10.1, 20.3, and 30.4 msw, respectively).

Boyle's Equation:

$$P_1 V_1 = P_2 V_2$$

P_1 = initial pressure surface absolute
V_1 = initial volume in cubic feet (ft^3)
P_2 = final pressure absolute
V_2 = final volume in cubic feet (ft^3)

Example 1 - Boyle's Law

Transposing to determine the volume (V_2) at 33 ft.:

$$V_2 = \frac{P_1 V_1}{P_2}$$

P_1 = 1 ata
P_2 = 2 ata
V_1 = 24 ft^3

$$V_2 = \frac{1 \text{ ata} \times 24 \text{ ft}^3}{2 \text{ ata}}$$

$$V_2 = 12 \text{ ft}^3$$

NOTE: The volume of air in the open bell has been compressed from 24 to 12 ft^3 in the first 33 ft. of water.

Example 2 - Boyle's Law

Using the method illustrated above to determine the air volume at 66 ft.:

$$V_3 = \frac{P_1 V_1}{P_3}$$

P_3 = 3 ata

$$V_3 = \frac{1 \text{ ata} \times 24 \text{ ft}^3}{3 \text{ ata}}$$

$$V_3 = 8 \text{ ft}^3$$

NOTE: The volume of air in the open bell has been compressed from 24 to 8 ft^3 at 66 ft.

Example 3 - Boyle's Law

For the 99 ft. depth, using the method illustrated previously, the air volume would be:

$$V_4 = \frac{P_1 V_1}{P_4}$$

P_4 = 4 ata

$$V_4 = \frac{1 \text{ ata} \times 24 \text{ ft}^3}{4 \text{ ata}}$$

$$V_4 = 6 \text{ ft}^3$$

How is it that a breath-hold diver can return to the surface from a depth of several hundred feet with no problem, but a scuba diver at a very shallow depth who comes to the surface holding his breath may develop an air embolism and die?

Assume that a breath-hold diver is going from the surface down to 99 ft. During descent, the gas in his lungs will be compressed, until at 99 ft. it will be reduced to one-fourth the original volume. As he ascends, the volume of gas in his lungs expands back to the original amount, thus there is no change in the original volume.

A scuba diver at 99 ft. is in a pressure/volume balance with his environment. He takes a breath of air, discards his

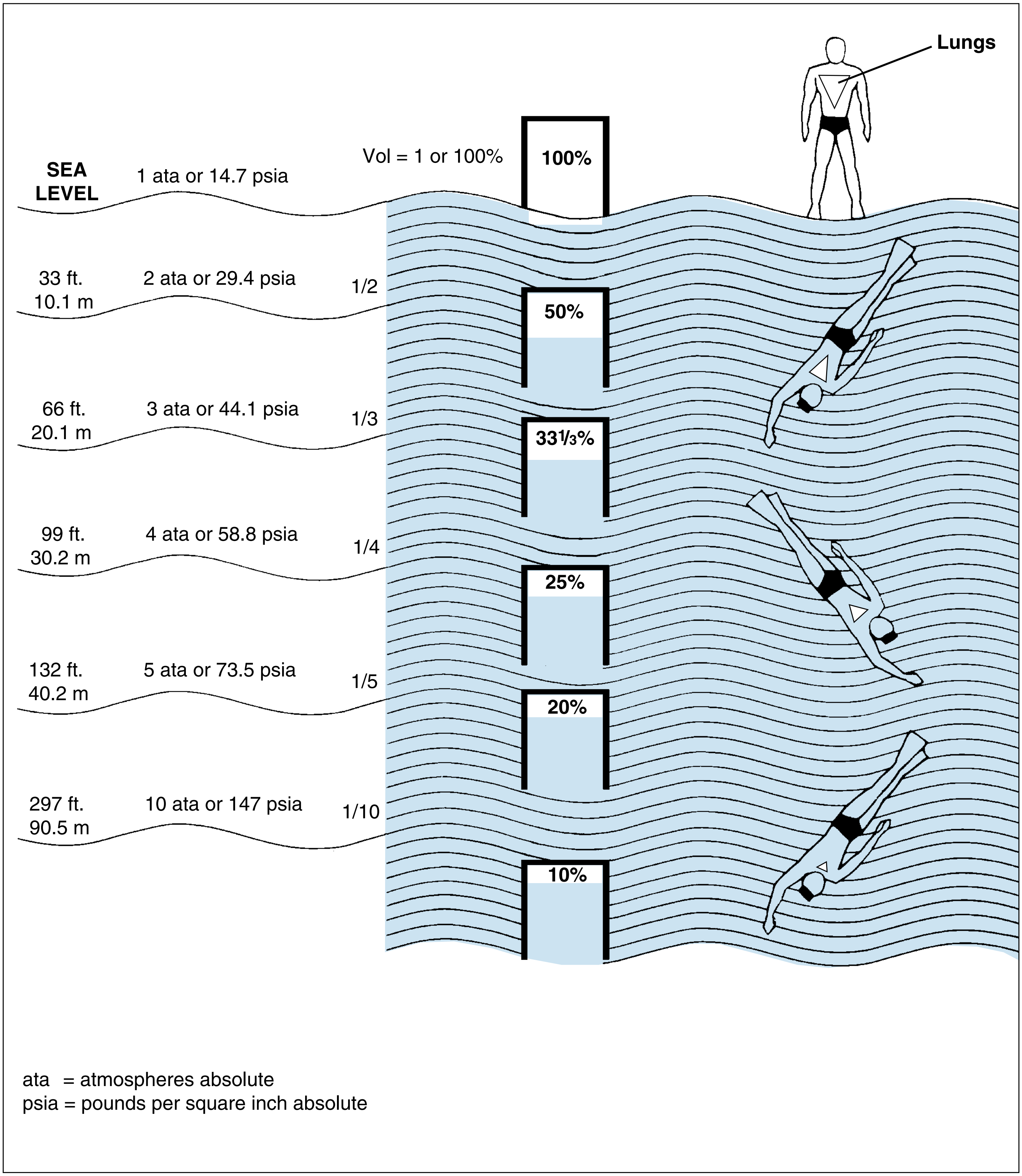


FIGURE 2.6
BoyleÕs LawApplied to Depth Versus Volume and Pressure

scuba gear and ascends holding his breath. As he ascends, his body is affected by Boyle's Law. By the time he reaches 33 ft., the air in his lungs will have increased in volume to match the decrease in water pressure. If he continues to ascend without releasing air from his lungs, the effect of increasing volume in his lungs may actually rupture the lungs with fatal consequences.

2.8.2 Charles'/Gay-Lussac's Law

Temperature has an effect on the pressure and volume of a gas. It is essential to know the effect of temperature since temperature at depth is often different from that at the surface.

"For any gas at a constant pressure, the volume of the gas will vary directly with the absolute temperature or for any gas at a constant volume, the pressure of the gas will vary directly with the absolute temperature."

Example 1: Charles' Law – Volume Change

To illustrate Charles' Law, consider a balloon with the capacity of 24 ft^3 of air which is lowered into the water to a depth of 99 ft. At the surface the temperature is 80°F, at depth the temperature is 45°F. What is the volume of the gas at 99 ft?

From Example 3 in the illustration of Boyle's Law above, we know that the volume of the gas (24 ft^3) was compressed to six cubic feet when lowered to the 99 ft. level. The application of Charles' equation illustrates the further reduction of volume due to temperature effects.

Charles' Equation:

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} \quad \text{(pressure remains constant)}$$

where

V_1 = volume at 99 ft. = 6 ft^3
T_1 = 80°F + 460 = 540 Rankine
T_2 = 45°F + 460 = 505 Rankine
V_2 = unknown

Transposing:

$$V_2 = \frac{V_1 T_2}{T_1}$$

$$V_2 = \frac{6\ ft^3 \times 505\ R}{540\ R}$$

$$V_2 = 5.61\ ft^3$$

NOTE: The volume within the balloon at 99 ft. was reduced further due to the drop in temperature.

Example 2: Gay-Lussac's Law – Pressure Change

A scuba cylinder contains 3,000 psig (3,014.7 psia) at 64°F. It is left on the boat deck on a hot summer day. What will the cylinder pressure be if the temperature of the air inside reaches 102°F?

Stated mathematically:

$$\frac{P_1}{P_2} = \frac{T_1}{T_2} \quad \text{(volume constant)}$$

P_1 = 3,014.7 psia
T_1 = 64°F + 460 = 524 Rankine
T_2 = 102°F + 460 = 562 Rankine
P_2 = Unknown

Transposing:

$$P_2 = \frac{P_1 T_2}{T_1}$$

$$P_2 = \frac{3{,}014.7\ \text{psia} \times 562\ R}{524\ R}$$

$$P_2 = 3{,}233.3\ \text{psia}$$

Converting to gauge pressure yields:

3,233.3 psia - 14.7 psi = 3,218.6 psig

Note that a scuba cylinder is a non-flexible, constant-volume container. As kinetic energy increases with increased temperature, the molecules travel faster. They hit the vessel walls harder and more often. This means pressure within the cylinder increases as the temperature is raised, and there is an increase in pressure associated with heating a scuba cylinder. To prevent this increase in pressure (and possible rupture of the scuba valve safety disc), it is especially important to store full scuba cylinders in a cool place.

2.8.3 Dalton's Law

The human body has a wide range of reactions to various gases or mixtures under different conditions of pressures. Dalton's Law is used to compute the partial pressure differences between breathing at the surface and breathing at various depths in the water.

"The total pressure exerted by a mixture of gases is equal to the sum of the pressures of each of the different gases making up the mixture, with each gas acting as if it alone was present and occupied the total volume." In other words, the whole is equal to the sum of the parts, and each part is not affected by any of the other parts.

According to Dalton's Law, the total pressure of a mixture of gases is the sum of the partial pressures of the components of the mixture.

Stated mathematically:

$$\mathbf{P_t = PP_1 + PP_2 + PP_3, etc}$$

P_t = Total Pressure
PP_1, etc. = Partial Pressure of the first gas, etc.

Partial pressure of a given quantity of a particular gas is the pressure it would exert if it alone occupied the total volume. The figure P_x is used to indicate partial pressure. The subscript x represents the specific gas (i.e. PO_2 for the partial pressure of oxygen). To determine the partial pressure of a gas in a mixture, use the following equation:

Partial Pressure = (Percent of Component) × (Total Pressure [absolute])

Stated mathematically:

$$\mathbf{P_x = Gas\ \% \times P_t}$$

(The pressure can be stated in psi, atm, etc.)
Gas % = Percent of Component (decimal)
P_t = Total Pressure
P_x = Partial Pressure of Gas

Imagine a container at atmospheric pressure (1 atm or 14.7 psi). If the container is filled with oxygen alone, the partial pressure of the oxygen will be 14.7 psi (1 atm). If the container is filled with dry atmospheric air, the total pressure will also be 14.7 psi (1 atm), as the partial pressures of all the constituent gases contribute to the total pressure (see Table 2.6).

Example 1: Dalton's Law

Calculate the partial pressure of nitrogen in a total mixture as follows:

$$P_x = \text{Gas \% (decimal)} \times P_t$$

$$PN_2 = .7808 \times 14.7 \text{ psi} = 11.478 \text{ psi}$$

or in atmospheres:

$$PN_2 = .7808 \times 1.0 \text{ atm} = .7808 \text{ atm}$$

If a scuba cylinder is filled to 2,000 psi with atmospheric air, the partial pressure of the various components will reflect the increased pressure in the same proportion as their percentage of the gas (see Table 2.7).

Example 2: Dalton's Law

What is the partial pressure of nitrogen within the scuba cylinder filled to 2,000 psi?

$$P_x = \text{Gas \% (decimal)} \times P_t$$

$$PN_2 = .7808 \times 2{,}000 \text{ psi} = 1{,}561.6 \text{ psi}$$

or in atmospheres:

$$PN_2 = 2{,}000 \text{ psi} \times 1 \text{ atm}/14.7 \text{ psi} = 136.05 \text{ atm}$$

$$PN_2 = .7808 \times 136.05 \text{ atm} = 106.23 \text{ atm}$$

Observe in Tables 2.6 and 2.7 that, while the partial pressures of some constituents of the air (particularly CO_2) were negligible at 14.7 psi, these partial pressures have increased to significant levels at 2,000 psi.

The implications for divers are important. If surface air is contaminated with 2% (PCO_2 .02 ata) of carbon dioxide, a level a person can easily accommodate at one atm, the partial pressure at depth will be dangerously high.

The Dalton's Law correlation in gas density with oxygen and nitrogen richness is illustrated in Figure 2.7.

2.8.4 Henry's Law

"The amount of any given gas that will dissolve in a liquid at a given temperature is proportional to the partial pressure of that gas in equilibrium with the liquid and the solubility coefficient of the gas in the particular liquid." If one unit of gas is dissolved at one atm, then two units will be dissolved at two atm, three units at three atm, etc.

Henry's Law stated mathematically:

$$\mathbf{\frac{VG}{VL} = \alpha\, P_1}$$

where
VG = Volume of gas dissolved at STPD (standard temperature pressure dry)

VL = Volume of the liquid

α = Solubility coefficient at specified temperatures

P_1 = Partial pressure of that gas above the liquid

When a gas-free liquid is first exposed to a gas mixture, gas molecules will diffuse into the solution, pushed by the partial pressure of each individual gas. As the gas molecules enter the liquid, they add to a state of "gas tension," a way of identifying the partial pressure of the gas in the liquid. The difference between the gas tension and the partial pressure of the gas outside the liquid is called the pressure gradient, which gives an indication of the net rate at which the gas tends to enter or leave the solution. When the gradient for diffusion into tissue is high, with low tension and high partial pressure, the rate of absorption into the liquid is high. As the number of gas molecules in the liquid increases, the gas tension increases until it reaches an equilibrium value equal to the outside partial pressure. At that point, the liquid is "saturated" with the gas molecules, and the pressure gradient is zero. Unless there is some change in temperature or pressure, the net rate at which gas molecules enter or leave the liquid will be zero, and the two states will remain in balance.

How does this phenomenon apply to divers? To begin with, a large percentage of the human body is water. Whenever a gas is in contact with a liquid, a portion of the gas will dissolve in the liquid until equilibrium is reached.

TABLE 2.6
Air at 14.7 psi (1 atm)

Air	Percent of component	Partial Pressure atm	Partial Pressure psi
N_2	78.08%	.7808 atm	11.478 psi
O_2	20.95%	.2095 atm	3.080 psi
CO_2	.03%	.0003 atm	.004 psi
Other	.94%	.0094 atm	.138 psi
Total	100.00%	1.000 atm	14.700 psi

TABLE 2.7
Air at 2,000 psi (136.05 atm)

Air	Percent of component	Partial Pressure atm	Partial Pressure psi
N_2	78.08%	106.23 atm	1561.6 psi
O_2	20.95%	28.50 atm	419.0 psi
CO_2	.03%	.04 atm	.6 psi
Other	.94%	1.28 atm	18.8 psi
Total	100.00%	136.05 atm	2000.0 psi

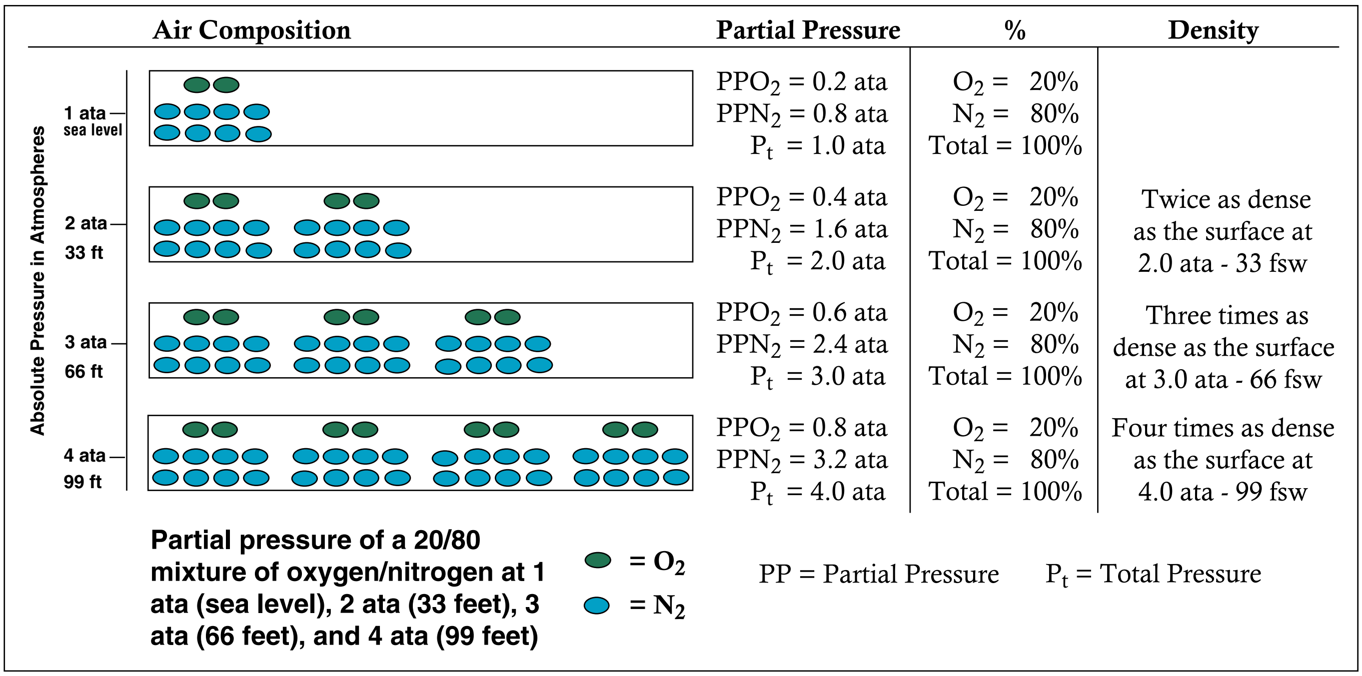


FIGURE 2.7
Partial Pressure

Gas can dissolve in water and fat in the human body as they make up a large percentage of the body's total mass. The deeper one dives, the greater the pressure exerted upon the body, and the higher the total pressure of the breathing gas. It follows that more gas will dissolve in the body tissues. During ascent, the dissolved gases will begin to be released.

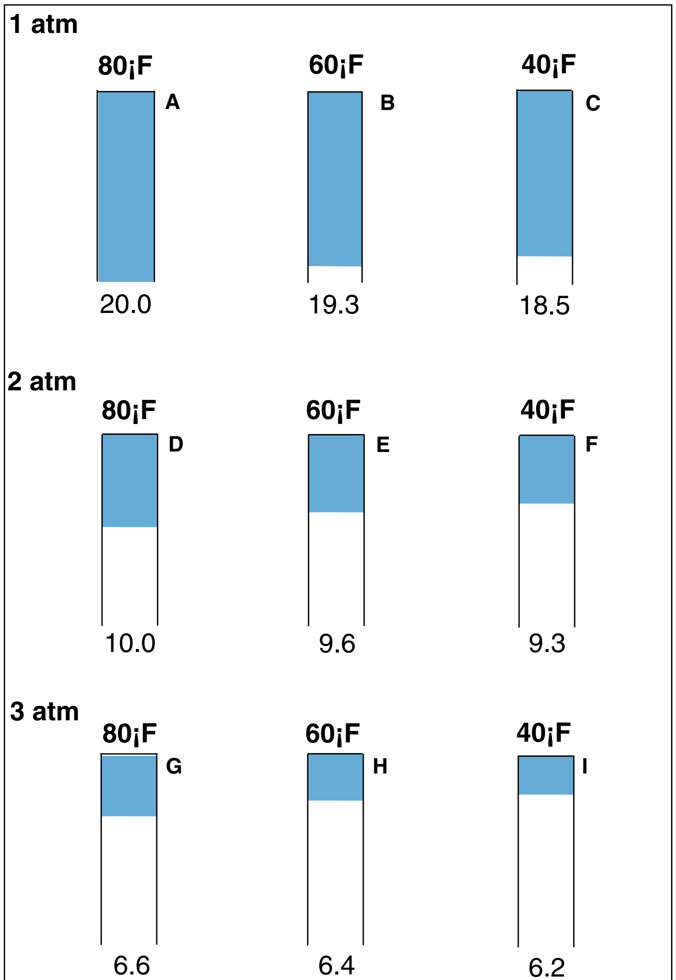
If a diver's rate of ascent (including decompression stops) is controlled properly, the dissolved gas will be carried to the lungs by the tissue's blood supply and will be exhaled before it accumulates and forms bubbles in the tissues. If, on the other hand, ascent is too rapid and/or decompression stops are missed or reduced so that the pressure is reduced at a rate higher than the body can accommodate, gas bubbles may form, disrupting body tissues and systems, and producing a condition known as decompression sickness (the bends).

The various gases are dissolved in the body in proportion to the partial pressure of each gas in the breathing medium. The amount of gas dissolved is also governed by the length of time and the pressure at which you breathe it. However, as gases vary in their solubility, the exact amount dissolved depends on the specific gas in question. If a diver breathes a gas long enough, his body will become saturated; but this occurs slowly. Depending on the gas, it will take anywhere from 8 to 24 hours.

Some gases are more soluble than others and some liquids are better solvents than other liquids. For example, nitrogen is five times more soluble (on a weight-for-weight basis) in fat than in water. These facts and the differences in blood supply have led to the postulate of tissues with different saturation halftimes (5-minute tissues, 10-minute tissues, 20-, 40-, 75-, etc.). This serves as the basis for calculating decompression tables.

2.8.5 General Gas Law

Pressure, volume, and temperature are interrelated. A change in one factor must be balanced by a change in one



Imagine a uniform-bore tube, sealed on one end, is inverted in a container of water at 80¡F and one atmosphere. The volume of air in the tube will be affected by changes in temperature and pressure in accordance to the following gas laws:

CharlesÕ Law:
(A, B, C) (D, E, F) (G, H, I) illustrate the reduction in volume caused by a reduction temperature at a constant pressure.

BoyleÕs Law:
(A, D, G) (B, E, H) (C, F, I) illustrate the reduction in volume caused by an increased pressure at a constant temperature.

The General Gas Law (CharlesÕ and BoyleÕs Laws Combined):
(A, E, I) (C, E, G) illustrate that a change in either volume, temperature, or pressure causes changes to the others.

FIGURE 2.8
Gas Laws

or both of the remaining factors. The General Gas Law (also known as the Ideal Gas Law) is a convenient combination of Charles' and Boyle's laws. It is used to predict the behavior of a given quantity of gas when changes may be expected in any or all of the variables (see Figure 2.8).

The general gas law is stated mathematically as follows:

$$\frac{\mathbf{P_1\,V_1}}{\mathbf{T_1}} = \frac{\mathbf{P_2\,V_2}}{\mathbf{T_2}}$$

where

P_1 = initial pressure (absolute)
V_1 = initial volume
T_1 = initial temperature (absolute)

and

P_2 = final pressure (absolute)
V_2 = final volume
T_2 = final temperature (absolute)

Example of General Gas Law:

Consider the open diving bell of 24 ft^3 capacity lowered to 99 ft. in seawater. Surface temperature is 80°F and depth temperature is 45°F. Determine the volume of the gas in the bell at depth.

The General Gas Law states:

$$\frac{P_1\,V_1}{T_1} = \frac{P_2\,V_2}{T_2}$$

P_1 = 14.7 psia
V_1 = 24 ft^3
T_1 = 80°F + 460 = 540 Rankine
P_2 = 58.8 psia
T_2 = 45°F + 460 = 505 Rankine
V_2 = Unknown

Transposing: $$V_2 = \frac{P_1\,V_1\,T_2}{T_1\,P_2}$$

$$V_2 = \frac{(14.7\text{ psia})\,(24\text{ ft}^3)\,(505\text{ R})}{(540\text{ R})\,(58.8\text{ psia})}$$

$$V_2 = 5.61\text{ ft}^3$$

NOTE: The volume was reduced, due to the drop in temperature and the increase in outside pressure.

2.9 MOISTURE IN BREATHING GAS

Breathing gas should have sufficient moisture for comfort. Too much moisture in a system can increase breathing resistance and produce lung congestion; too little can cause an uncomfortable sensation of dryness in the mouth, throat, nasal passages, and sinus cavities. Air or other breathing gases supplied from surface compressors or tanks can be assumed to be dry. This dryness can be reduced by removing the mouthpiece and rinsing the mouth with water or by introducing a small amount of water inside a full-face mask. It can be dangerous to use gum or candy to reduce dryness while diving. Do not remove your mouthpiece in seawater or freshwater that may be polluted.

2.9.1 Humidity

Water vapor (a gas) behaves in accordance with the gas laws. However, because water vapor condenses at temperatures we are likely to encounter while diving, the effects of humidity are important considerations.

2.9.2 Condensation in Breathing Hoses or Mask

Exhaled gas contains moisture that may condense in breathing hoses of a rebreather or in your mask. This water is easily blown out through the exhaust valve and, in general, presents no problem. However, in very cold water, this condensation may freeze, disrupting normal functioning of a scuba regulator. The dive should be aborted if such a condition occurs.

2.9.3 Fogging of the Mask

Masks become fogged because of the moisture in exhaled breath, or because of evaporation through facial skin. This fogging can be prevented by moistening the face plate with saliva, liquid soap, or commercial anti-fogging products. Exhalation through the mouth, instead of the nose, will reduce face mask fogging.

2.10 LIGHT

The sense of sight allows perception of electromagnetic energy (light). Human beings can perceive only the very narrow range of wave lengths from 380 to 800 nanometers (see Chapter 3). Eyes function by collecting light that is emitted or reflected by an object. Some light is absorbed by the object, making the object appear colored. The energy waves that are received by the eye are turned to electrical impulses in nerves and sent to the brain via the optic nerve. The brain interprets the signals and we "see."

Under water, the eyes continue to function by collecting light reflected off objects, but the light itself changes. Water slows the speed at which light travels. As light enters or leaves water, this change in speed causes light rays to bend, or refract (see Figure 2.9 and 2.10). That is why a pencil in a glass of water looks bent. Seen through a diving mask, refraction affects close vision, creating distortions that affect eye-hand coordination and the ability to grasp objects under water.

By placing a pocket of air (i.e., a facemask) between the water and the eyes, the light rays are refracted twice – once when they enter the air from the water and again as they enter the eyes; a clearer image is now focused on the retina. Due to imperfect correction, however, the retinal image is larger. Objects may now appear approximately 25% larger because of the larger-than-normal retinal image.

The visual distortions caused by the mask vary considerably with the viewing distance. For example, at distances of less than four feet (1.2 m), objects appear closer than they actually are. However, overestimation occurs at distances greater than four feet, and this degree of error increases in turbid or muddy water. Other perceptual distortions are also apparent. Stationary objects appear to move when the head is turned from side to side.

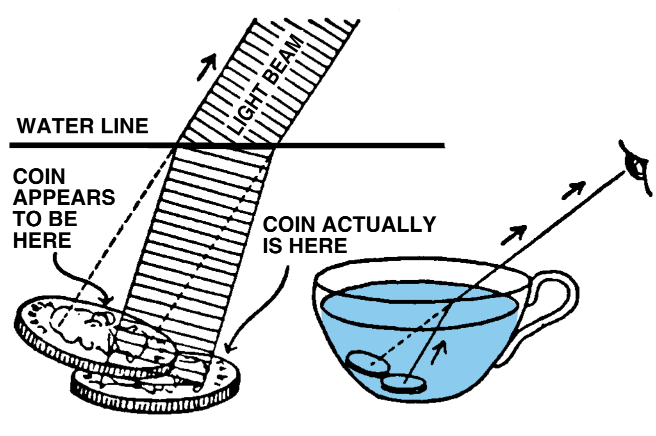
Turbidity is another factor affecting underwater visibility. Turbidity refers to the clarity of the water, and depends on the quantity of particulates in suspension. Muddy water is more turbid than clear water. Turbidity can cause overestimation of the distance of an underwater object.

It is important to remember that underwater distance perception is very likely to be inaccurate and that errors of both underestimation and overestimation may occur. As a rule of thumb, the closer the object, the more likely it will appear to be closer than it really is. Additionally the more turbid the water, the more likely it will appear farther than it really is.

2.10.1 Colors

Water absorbs light according to its wavelength. The deeper the light penetrates the water, the more light wavelengths are absorbed. Absorption begins at the red end of the spectrum. Orange is the next color to be lost, followed by yellow, and then green. In very deep water, the only colors visible are blue and violet. Turbidity affects the ability to see colors because the suspended particles diffuse and scatter light. Turbid water gives greatest transparency to wavelengths in the green range. Thus, very clear water is blue, while turbid water is usually green.

Three feet of distilled water absorbs twelve percent of the red, but only one percent of the blue rays. Therefore at 65 ft. (19.8 m), even in very clear water, red is not visible, and the intensity of yellow has decreased by about 95 percent. At the



Light rays are refracted as they enter the water. The coin is actually closer than it appears.

FIGURE 2.9
Refraction

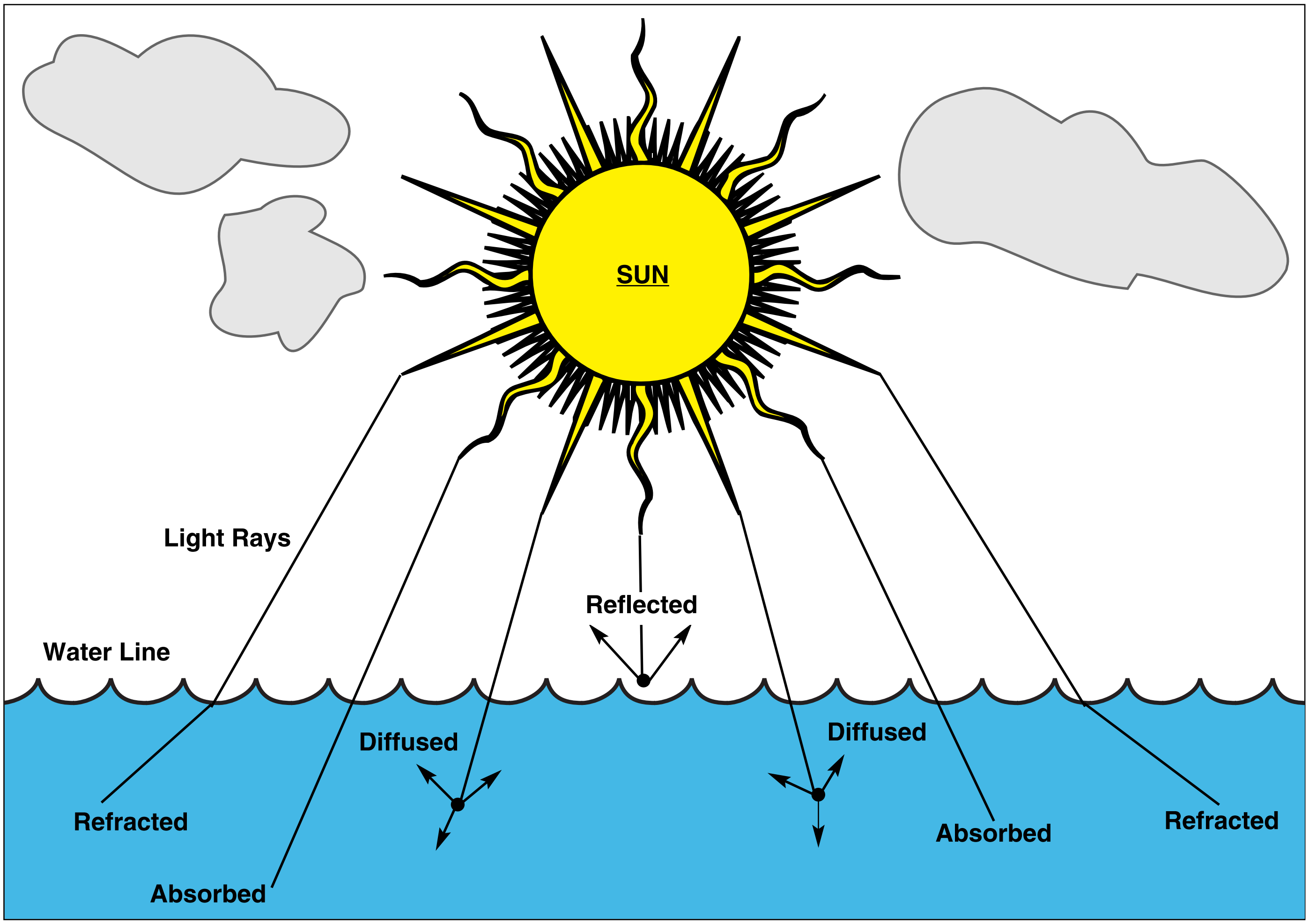


FIGURE 2.10
Sunlight In Air And Water

same depth, however, blue appears with 40 to 50 percent of its initial surface intensity. Some sunlight may penetrate to as deep as 2,000 ft. (610 m) (Kinney 1985).

Turbidity also affects the ability to see colors because the suspended particles diffuse and scatter light.

In general, as depth increases, the ability to discern colors decreases, until visible objects are distinguishable only by differences in brightness. At deeper depths, contrast becomes the most important factor in visibility. Fluorescent paint does aid visibility (see Table 2.8).

2.11 SOUND

Although light and sound both travel in waves, the nature of these two waves is different. Light waves are electromagnetic. Sound is produced by pressure waves triggered by vibration. As the medium containing the pressure wave comes into contact with another medium, a sympathetic vibration occurs. This transfers the wave pattern to the second medium. As an example, a sound is produced and the disturbance travels through the air as a pressure wave striking our eardrums. This sets off a sympathetic vibration in the eardrums. The inner ear turns this mechanical vibration of the eardrum into a nerve impulse. The impulses are sent to our brain for interpretation.

The more dense the medium through which sound travels, the faster the speed of sound. In dense media, molecules are packed close together, allowing easier transmission of the wave motion. The speed of sound through air is 1,125 ft. (343 m) per second; the speed of sound through seawater is 5,023 ft. (1,531 m) per second; and the speed of sound through steel is 16,600 ft. (5,060 m) per second. The speed of transmission of sound in water depends on the temperature of the water (colder water is denser, thereby allowing it to transmit sound faster) and salinity (seawater allows sound to travel faster than freshwater, again because it is more dense).

Because the speed of sound depends on the density of the medium it travels through, interesting acoustical effects occur in water that has several temperature layers (known as thermoclines). The density of water varies according to its temperature. When sound waves transfer from water of one temperature/density to another, as when they

TABLE 2.8
Colors That Give Best Visibility Against a Water Background

Water Condition	Natural Illumination	Incandescent Illumination	Mercury Light
Murky, turbid water of low visibility (rivers, harbors, etc.)	Fluorescent yellow, orange, and red	Yellow, orange, red, and white (no advantage in fluorescent paint)	Fluorescent yellow-green and yellow-orange
	Regular yellow, orange, and white		Regular yellow, white
Moderately turbid water (sounds, bays, coastal water)	Any fluorescence in the yellows, oranges, or reds	Any fluorescence in the yellows, oranges, or reds	Fluorescent yellow-green or yellow-orange
	Regular paint of yellow, orange, and white	Regular paint of yellow, orange, and white	Regular yellow, white
Clear water (southern water, deep water offshore, etc. See note.)	Fluorescent paint	Fluorescent paint	Fluorescent paint

NOTE: With any type of illumination, fluorescent paints are superior.
a. With long viewing distances, fluorescent green and yellow-green are excellent.
b. With short viewing distances, fluorescent orange is excellent.

encounter a thermocline, substantial energy is lost. This tends to isolate sound within water of a consistent temperature. Interestingly, a diver who is not in the same thermocline range as the source of a sound often cannot hear that sound, even though it is coming from only a few feet away.

Hearing under water is affected in important ways. It is almost impossible to determine from which direction a sound originates. On land, sound reaches one ear before the other; thus, the direction of the source can be determined. Under water, sound travels so quickly it reaches both ears without an appreciable interval. The sound seems to originate from all directions. Sound travels faster, seems non-directional, and is more easily heard under water.

NOTES

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