

for depressions. On the other hand, the more generally practical use of increased pressures is in connection with the displacement of water while diving and with water and solids (such as sand and mud) in caisson work. In these cases the weight of the air becomes a constant and usually negligible factor upon which is superimposed the pressure exerted by the column of fluid above the free surface in such an enclosed system. In diving a uniform, or arithmetic, increment in pressure per unit depth is added on descent. Due to the relatively moderate depths that can be withstood in diving, and the nearly incompressible nature of fluids, this increment is practically constant. In sea water it is about 1 atm. for every 33 feet, or roughly $\frac{1}{2}$ p.s.i. per foot.

4. Partial Pressures

The pressure changes described above are shared by the constituent gases of the atmosphere in the ratio of their occurrence (Dalton's law) in air, or as modified in any gas space, such as the alveoli of the lungs. This situation follows from the kinetic concept of gases, since each species of molecule exerts its pressure throughout any definable space as if it were there alone. Thus, $P(\text{air}) = p'N_2 + p''O_2 + p'''H_2O + p'''' \text{inert gas, etc.}$, where $P(\text{air})$ is the barometric pressure and p' , p'' , and so forth, are the partial pressures of the several gases of the air. (Certain deviations from this simple relation may be expected under high positive pressures due to interference of different molecular species with each other.) Any change in the sum of these component pressures must be shared proportionately by the constituents. Thus, a curve describing the partial pressure of oxygen in air at different altitudes would be the line drawn about one fifth the distance from the ordinate axis for air pressure at any altitude in Figure 1. It will be seen later that the partial pressure of gases is most important in the physiology of gaseous exchange and actually forms the nucleus of all rational approach to the problems of the effects of abnormal atmospheric pressures upon living organisms.

The pressure-volume relationship of gases at constant temperature at any altitude is determined from Boyle's law ($PV = \text{constant}$) by the relation $P_1V_1 = P_2V_2$. By suitable arrangement the volume assumed by wet air at constant temperature at any altitude is calculable from the formula^{10,11}

$$V_2(\text{wet}) = V_1(\text{wet}) \frac{P_1 - p_{H_2O}}{P_2 - p_{H_2O}}$$

where the symbols have the following significance: V_1 = the initial volume at ground level, V_2 = the resulting volume at altitude of the same weight of gas contained in V_1 , P_1 = the initial pressure at ground level, P_2 = the resulting

¹⁰ C. H. Fugitt, U. S. Naval Air Training Bases, Pensacola, Fla., Project X-579, Report No. 1 (1945).

¹¹ Aero Medical Laboratory Staff, AAF Manual No. 25-2 (March, 1945).

pressure at altitude, and p_{H_2O} = the vapor pressure of water at the specified temperature.

5. Solubility or Absorption

To a greater or lesser extent all gases are soluble in other substances and, barring for the moment chemical reaction with the solvent, the concentration of gas dissolved is proportional to its pressure (partial pressure) above the interface of gas and solvent (Henry's law). The rate of solution depends on the difference between the frequency at which the gas molecules impinge on the surface and that at which the gas in solution tends to escape from solution, and on the nature of the solvent.

These facts are equally true of gases dissolved in body substances. The living organism exists in a state of dynamic equilibrium with the gases that impinge upon it. The diffusion of gases through the outer skin of the body is of a low order as shown by Behnke and Willmon.¹² As far as practical considerations require the gaseous exchange of the body is through the lung surface, where solution is effected readily.

It has been shown by Haldane¹⁴ and others that the volume of oxygen passed by a human lung-blood interface per mm. Hg differential pressure (across the interface) varies from 25 to 56 ml. per minute, depending on the state of exercise. For carbon dioxide the value is at least 20 to 30 times as great. This ratio is fantastically equal but unrelated to that of their solubilities in water at equilibrium, and actually depends on factors affecting diffusivity.¹⁵

The greater solubility in water of carbon dioxide, as compared with oxygen, is due mainly to its tendency to form a hydrate with water and to ionize in solution. This tendency acts to decrease the partial pressure of carbon dioxide in the water and thus to maintain a greater difference in concentration across the interface than can be effected with inert gases.

As implied above, the solution of gases in the body is primarily a solution in water, and after this initial action further distribution depends on the diffusion, circulation, partition, and chemical combination, if any, in other body substances. Other than the blood and watery fluids the most important gas reservoir of the body is its fat, since fat occurs in a relatively large quantity. Gas dissolved in fat has been shown by Behnke,¹⁶ Gersh,¹⁷ and others to be a major source of difficulty in caisson disease. Table 2 shows the solubilities of several atmospheric gases in water and olive oil and the ratio of these solubilities, indicating that at

¹² J. Sendroy, Jr., R. T. Dillon, and D. D. Van Slyke, *J. Biol. Chem.*, **105**, 597 (1934).

¹³ A. R. Behnke and T. L. Willmon, *Am. J. Physiol.*, **131**, 627 (1940-41).

¹⁴ J. S. Haldane and J. G. Priestley, *Respiration*. Yale Univ. Press, New Haven, 1935.

¹⁵ J. Verduin, *Science*, **121**, 215 (1955).

¹⁶ A. R. Behnke, *Bull. New York Acad. Med.*, **18**, 561 (1942).

¹⁷ J. Gersh and G. E. Hawkinson, *Naval Med. Research Inst., Research Project X-28*, Report No. 12 (1944).