

THE PHYSICS OF DIVING

Des Gorman

PRESSURE

A pressure is a force applied over an area and some units of pressure reflect this (e.g. kg/cm² and lbs/in²). There are numerous units of pressure used and they are usually referred to a standard of one atmosphere of pressure (taken to be the average pressure exerted by the earth's gaseous atmosphere at sea-level). Pressure unit-conversions are listed in TABLE 1. The unit, bar, will be used where possible in this text.

A diver is not only subject to the atmospheric pressure but also to that exerted by the water. This **water pressure** is a product of depth and water density (EQUATION ONE).

$$\text{Water pressure (PH}_2\text{O)} = \text{Depth (m)} \times \text{density (Gm}^{-3}\text{)} \quad \text{Eq. 1.}$$

As can be seen from TABLE 2.1, approximately 10 MSW, 33fsw and 34 feet fresh water are equivalent to one atmosphere or one bar of pressure. The total ambient pressure, usually called the **absolute pressure** acting on a diver is a sum of the atmospheric pressure and the water pressure (EQUATION TWO).

$$\text{Absolute pressure (Pabs)} = \text{Atmospheric pressure} + \text{PH}_2\text{O} \quad \text{Eq. 2.}$$

As the average atmospheric pressure is 1 bar, and each 10 MSW or 33fsw constitute another bar of pressure, the absolute pressure can be calculated as follows (EQUATIONS THREE and FOUR):

$$\text{Pabs (bars)} = \frac{\text{Depth (msw)}}{10} + 1 \quad \text{Eq. 3.}$$

$$\text{Pabs (bars)} = \frac{\text{Depth (fsw)}}{33} + 1 \quad \text{Eq. 4.}$$

The absolute pressure can also be easily converted back to depth using these calculations (EQUATIONS FIVE and SIX):

$$\text{Depth (msw)} = (\text{Pabs [bars]} - 1) \times 10 \quad \text{Eq. 5.}$$

$$\text{Depth (fsw)} = (\text{Pabs [bars]} - 1) \times 33 \quad \text{Eq. 6.}$$

The relationship between the absolute pressure acting on a diver and their depth in the ocean is shown in FIGURE 1. With increasing altitude above sea-level, atmospheric pressure decreases; the relationship between atmospheric pressure and altitude is shown in FIGURE 2.

Pressure and Liquids. A pressure applied to a liquid will be transmitted "equally" throughout the liquid (and in this context the body tissues should be considered to be liquid) and will

cause only a small decrease in liquid volume (as liquids are difficult to compress in comparison to gases). The compression will vary according to the space within and between the liquid molecules. This transmission of pressure is described by **Pascal's Principle**, and is the basis of hydraulics. It also explains why the middle-ear and paranasal sinuses are subjected to ambient pressure (absolute pressure), as the pressure acting on the diver is transmitted through the blood and into the mucosal linings of these spaces. In addition, Pascal's Principle explains why a victim of a lethal **underwater blast** may show little superficial evidence of trauma. The blast wave will transmit through the water and body tissues (just more liquid) until they reach a liquid-gas interface (e.g. lungs and gut) where extensive damage may occur.

Pressure and Gases. A pressure applied to a gas causes significant changes in gas volume and temperature. In the conventional diving range this is reasonably described by the **Ideal Gas Law**. This is the most used law in diving medicine along with two derivatives, **Boyle's Law** and **Charles' Law**.

Ideal Gas Law. The Ideal Gas Law states that the product of absolute gas pressure (P) and volume (V) must equal the product of the number of moles of gas (N), the gas constant (R) and the absolute temperature (T) [EQUATION SEVEN].

$$P \times V = N \times R \times T \qquad \text{Eq. 7.}$$

A mole of gas is the molecular weight of the gas in grams. For example, 2 grams of hydrogen (H₂) is one mole of this gas, as is 4 grams of helium (He), 28 grams of nitrogen (N₂) and 32 grams of oxygen (O₂). Regardless of which gas is involved, one mole of any gas will contain exactly the same number of molecules (Avigadro's number, 6.023 x 10²³) and at 0°C and 1 bar will always occupy a volume of 22.4 litres.

The gas constant has been found by "experiment" to be 0.082.

Absolute temperature is in the kelvin scale. Each degree celsius (°C) is equivalent to one kelvin. However, the kelvin scale begins at -273°C, such that water will freeze at 273k (0°C) and body temperature is 310k (37°C).

If gas is neither added to nor subtracted from a system then the number of moles of gas will not change, and as the gas constant is just that, constant, EQUATION SEVEN can be simplified to (EQUATION EIGHT):

$$\frac{P \times V}{T} \text{ is constant} \qquad \text{Eq. 8.}$$

This can also be related to two sets of gas conditions. If a closed gas system is at a specific pressure (P₁), it will have a corresponding volume (V₁) and temperature (T₁). If any of these parameters are changed, then the other two must change (EQUATION NINE).

$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2} \qquad \text{Eq. 9.}$$

Simplification of EQUATION NINE yields both Boyle's and Charles' Law.

Boyle's Law. If the mass of gas is not changed and its temperature is kept constant, then the product of gas pressure and volume will be constant (EQUATION TEN). That is, if the pressure is doubled, the volume will halve. If the pressure is increased 10 times, the volume will be decreased 10 times. Conversely, if the pressure is reduced 6 times, gas volumes will increase 6 times.

$$P_1 \times V_1 = P_2 \times V_2 \quad \text{Eq. 10.}$$

The relationship between the gas pressure and volume for a fixed mass of isothermic gas is predicted by Boyle's Law and is displayed in FIGURE 3. This shows why most damage done by expanding gases during the decompression of divers (barotrauma of ascent) occurs close to the surface (the rate of gas volume change is greatest) and why modern treatment regimens for decompression illness are largely restricted to low pressures (as most of the bubble volume reduction will occur with the first trebling of pressure and 100% oxygen can be breathed for reasonable periods with an acceptable incidence of overt oxygen toxicity).

Charles' Law. If the mass of gas is not changed and its volume is kept constant, then the gas pressure will vary linearly with its temperature (EQUATION ELEVEN).

$$\frac{P}{T} \text{ is constant; } \frac{P_1}{T_1} = \frac{P_2}{T_2} \quad \text{Eq. 11.}$$

When a gas cylinder is compressed to a greater pressure, the cylinder volume does not change but the gas temperature increases. Although this may appear to be an example of Charles' law, it is not as gas molecules are being added to the cylinder. Much of the heat is due to the work that the compressor has to perform to add and compress this additional gas. This is called **adiabatic heating**. However, when the cylinder is fully compressed and the valve closed, it will cool and the gas pressure will fall. As gas is neither being added or subtracted this is an example of Charles' law.

Non-ideal Gas Law. The Ideal Gas Law assumes that gas molecules exert no influence on each other and hence that gases are perfectly compressible. This is a reasonable assumption for conventional diving depths, but as gas molecules are compressed closer together they begin to exert significant intermolecular forces and the gas becomes less compressible. For example, at 600 MSW (61 Bars) gases are about 8% less compressible than would be predicted by the Ideal Gas Law. In this circumstance **Van der Waal's Equation** gives a more accurate description (EQUATION TWELVE).

$$\left(P + \frac{N^2 a}{V^2} \right) \times (V - Nb) = NRT \quad \text{Eq. 12.}$$

Van der Waal's Equation is obviously a derivation of the ideal gas law and **a** and **b** are empirical constants for each gas which are listed in handbooks of chemistry and physics. Van der Waal's Equation is seldom used as the Ideal Gas Law error is usually small and calculating gas supplies

needed for divers to conduct a dive using the Ideal Gas Law will over-estimate their requirements.

Respiratory Gases. Inspired gases will be both heated and saturated with water vapour in the diver's upper airways. **Vapour pressures** are influenced by temperature and are not predicted by the Ideal Gas Law. Hence, it is necessary to allow for this in calculating diver's needs as gas leaving a cylinder will be dry and relatively cold. These conditions are referred to as BTPS (body temperature and pressure, saturated with water vapour) and STPD (standard temperature and pressure, dry). Standard temperature is 273k and pressure is 1 bar. This conversion can be made as follows (EQUATION THIRTEEN):

$$V_{\text{STPD}} = V_{\text{BTPS}} \times \frac{(P_{\text{abs}} - 47)}{760} \times \frac{273 + 0}{273 + 37} \quad \text{Eq. 13.}$$

This is based on a water vapour pressure of 47mmHg at body temperature (310k). Again, this conversion is usually not made when calculating diver's gas volume needs as ignoring water vapour and temperature changes will overestimate their gas requirements.

It is nevertheless important that when divers' lung function is tested on a spirometer that the BTPS scale (right hand) and not the STPD scale (left-hand) is used.

GAS PARTIAL PRESSURE

The partial pressure of a gas is the pressure that gas would exert if it occupied the space by itself. Such partial pressures determine gas diffusion rates and solubility; only the flammability of oxygen is concentration and not partial pressure dependent. The partial pressure of a gas in a mixture of gases is determined by its fractional concentration and is described by Dalton's law.

Dalton's law. This is expressed in two ways. Firstly, the total pressure exerted by a mixture of gases (P_{abs}) is equal to the sum of the partial pressures (P_{gas}) of each of the component gases. For example, if a mixture contained 3 different gases, then (EQUATION FOURTEEN):

$$P_{\text{abs}} = P_{\text{gas A}} + P_{\text{gas B}} + P_{\text{gas C}} \quad \text{Eq. 14.}$$

Secondly, the partial pressure of an individual gas is equal to the product of its fractional concentration (F_{gas}) [e.g. in a 20% O_2 and 80% N_2 mixture the fractional concentration of O_2 is 0.2 and N_2 is 0.8] and the total gas pressure (P_{abs}) [EQUATION FIFTEEN].

$$P_{\text{gas A}} = F_{\text{gas A}} \times P_{\text{abs}} \quad \text{Eq. 15.}$$

SOLUBILITY OF GASES IN LIQUIDS

The volume of gas that can be dissolved in a liquid is directly related to the partial pressure of gas to which the liquid is exposed and is described by Henry's law.

Henry's law. The volume of gas dissolved in a liquid (Q; mls of gas STPD/ml of liquid) is equal to the product of Henry's constant (K; available in handbooks of chemistry and physics) and the partial pressure of the gas (EQUATION SIXTEEN).

$$Q = K \times P_{\text{gas}}$$

Eq. 16.

It follows that if the gas partial pressure is doubled then the amount of gas in solution at equilibrium will double (e.g. during descent and at depth during a dive), and conversely, if the gas partial pressure is halved then half of the gas will leave solution (e.g. during or after ascent). If gas coming out of solution can not be expired by a diver, bubbles may form.

The solubility of a gas in a liquid decreases as the temperature of the liquid increases (FIGURE 4). It follows that if a diver becomes cold during a dive then more gas will go into solution and then when they subsequently re-warm that this gas will leave solution, perhaps to bubble. Gas solubilities in liquids also decrease slightly in the presence of electrolytes and with increasing ambient pressure.

Nitrogen's greater solubility in lipids than He, largely explains why only the former is narcotic (as it dissolves into the cell membranes of neurons), why N₂ more often affects the nervous system than He (as it dissolves into myelin) and both why adipose tissue acts as a depot for N₂ and why increasing adiposity is a recognised risk-factor for decompression illness.

GAS DIFFUSION

Gas diffusion is the net movement of gas molecules from an area of higher to an area of lower partial pressure.

Within a liquid, the diffusion rate of a gas is influenced by the gas's molecular weight (increasing weight decreases the rate), the viscosity of the liquid (increasing viscosity also decreases the rate) and the temperature of the liquid (increasing temperature increases the rate not only because it increases the gas energy but also because it reduces liquid viscosity). This effect of temperature probably explains why the risk of developing decompression illness after an O₂He dive increases if the diver's body temperature is increased; as He is essentially insoluble a negligible change in solubility will occur but He diffusion into tissues will be accelerated.

The rate of diffusion of a gas through another gas is determined by the molecular weights of the 2 gases, the ambient temperature (rates increase with increasing temperature) and ambient pressure. The diffusion rate varies inversely with pressure such that at 300 MSW (31 bars) the diffusion of O₂ through alveolar gas will be 31 times slower.

GAS FLUX

The rate of uptake and elimination of gases into and from tissues, bubbles and other gas spaces (i.e. the gas flux) is a product of the gases' diffusion and solubility co-efficients. This will be discussed further in the section on Decompression Theory (CHAPTER EIGHT).

GAS DENSITY AND VISCOSITY

The density of a gas is expressed in grams/litre/bar and at a specific temperature. It follows that gas density increases with increasing pressure. Indeed, gas density and pressure vary linearly (i.e. if pressure is increased 10 times, gas density also increases 10 times). Gas densities decrease with increasing temperature.

Gas viscosity is a measure of the "stickiness" of a gas and (opposite to liquids) increases as a linear function of gas temperature. Some gases (e.g. O₂ and N₂) show a slight increase in viscosity with increasing gas pressure, while others (e.g. He) do not change.

Work of breathing is related to both gas density and viscosity. Although He is more viscous than N₂ at atmospheric pressure, nitrogen's greater density requires more work for it to be breathed. In diving, increasing gas pressure is accompanied by similarly increasing density and only negligible or no change in viscosity such that breathing becomes essentially gas density limited. This is a major advantage for He over N₂ as an inert gas in diving, as is the avoidance of narcosis, an increased O₂ tolerance in the brain and typically only being associated with mild episodes of decompression illness. The disadvantages in using He include its cost, thermal conductivity and capacity (see below), its inability to ameliorate the High Pressure Neurological Syndrome and its distortion of speech (because of an increased speed of sound conduction).

THERMAL CAPACITY AND CONDUCTIVITY OF GASES

The **thermal capacity** of a gas is the amount of heat energy needed to raise the temperature of the gas by one kelvin (1°C); that is, it is a measure of how much heat a gas can store. It is measured by either holding volume or pressure constant.

The **thermal conductivity** of a gas is the rate at which heat is transferred through the gas.

Helium has 4 to 5 times the thermal capacity and 6 times the thermal conductivity of N₂. Consequently, diving with He has a much greater associated risk of both hypothermia and hyperthermia.

BUBBLES AND SURFACE TENSION PRESSURE

This is discussed in the section on Decompression Theory.

BUOYANCY

Archimedes Principle is that an object wholly or partly immersed in water is buoyed up by a force equal to the weight of water displaced by the object. This buoyant force then is a product of object (displaced water) volume and water density (EQUATIONS SEVENTEEN and EIGHTEEN).

$$\text{Nett weight} = \text{Object weight} - \text{buoyant force} \qquad \text{Eq. 17.}$$

$$\text{Buoyant force} = \text{Object volume (displaced water)} \\ \times \text{water density} \qquad \text{Eq. 18.}$$

SUGGESTED READING

1. Edmonds CW, Lowry C, Pennefather JW. Diving and subaquatic medicine, 4th edition.

FIGURE 1: The relationship between the depth of a diver in the ocean and the absolute pressure acting on the diver.

Depth Meters	Depth Feet	Air Pressure	Water Pressure	Atmospheric Absolute Pressure
Surface	Surface	1atm	Nil	1 ATA
10	33	1atm	1 atm	2 ATA
20	66	1atm	2 atm	3 ATA
30	99	1atm	3 atm	4 ATA
40	132	1atm	4 atm	5 ATA
90	297	1atm	9 atm	10 ATA

FIGURE 3: The relationship between the absolute pressure (horizontal axis in bars) and the relative volume of a fixed volume of isothermic gas (vertical axis in litres assuming a relative volume of 1.0 at 1 bar).

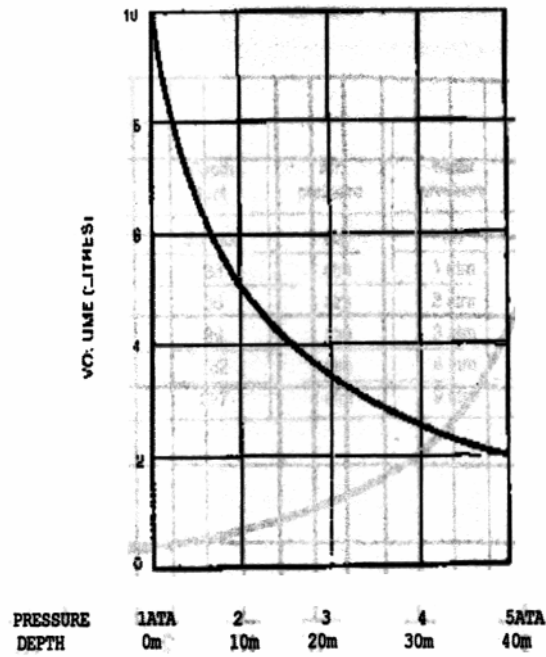


TABLE 1: PRESSURE UNITS

One atmosphere (At) = 10.08 (10) metres sea water (msw)
= 33.07 (33) feet sea water (fsw)
= 33.9 (34) feet fresh water
= 1.033 kg/cm ²
= 14.696 (14.7) Lbs/in ² (psi)
= 1.013 (1) bars
= 1013 millibars
= 101.3 kilopascals (kPa)
= 760 millimetres mercury (mmHg)
= 760 torr