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Buoyancy of gas-filled bladders at great depth

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ABSTRACT

At high hydrostatic pressures exceeding 20 MPa or 200 bar, equivalent to depths exceeding *ca.* 2000 m, the behaviour of gases deviates significantly from the predictions of standard equations such as Boyle's Law, the Ideal Gas Law and Van der Waals equation. The predictions of these equations are compared with experimental data for nitrogen, oxygen and air at 0 °C and 15 °C, at pressures up to 1100 bar (110 MPa) equivalent to full ocean depth of *ca.* 11000 m. Owing to reduced compressibility of gases at high pressures, gas-filled bladders at full ocean depth have a density of 847 kg m^{-3} for Oxygen, 622 kg m^{-3} for Nitrogen and 660 kg m^{-3} for air providing potentially useful buoyancy comparable with that available from man-made materials. This helps explain why some of the deepest-living fishes at *ca.* 7000 m depth (700 bar or 70 MPa) have gas-filled swim bladders. A table is provided of the density and buoyancy of oxygen, nitrogen and air at 0 °C and 15 °C from 100 to 1100 bar.

1. Introduction

Air or gas-filled bladders are widely used both in natural and human-engineered underwater systems to provide buoyancy. They have the advantage over rigid structures, such as hollow spherical fishing floats or buoyancy shells of certain cephalopods (Nautilus and cuttlefish), that weight is minimal, comprising only a gas-impermeable enclosing membrane which is supported by the internal gas pressure in equilibrium with the pressure of the surrounding water. The main disadvantage is that, unlike rigid floats, the volume of a gas bladder changes with depth and pressure in accordance with Boyle's Law. For a given quantity of gas:

$$PV = P_0 \ \ V_0 \tag{1}$$

Where P is pressure and V is the volume at a given depth and P_0 and V_0 are corresponding reference values usually at the sea surface, pressure = 1 atm. Boyle's Law is used for the calculation of volumes of air necessary in buoyancy bags for diving and underwater engineering. During descent, volume and hence buoyancy decrease, therefore to avoid sinking air must be added. Conversely, during ascent, to avoid expansion and uncontrolled acceleration to surface, air must be removed by relief valves and bottom-anchored tethers are often used to ensure a safe ascent (IMCA, 2007). Fishes maintain buoyancy close to neutral to enhance manoeuvrability and energetic efficiency of swimming (Alexander, 1966). In many species buoyancy is regulated by secreting gas into the swim bladder from a gas gland or removing gas through a pneumatic duct or a blood capillary network (Pelster, 1997).

At shallow depths the weight of gas in the buoyancy bladder is negligible and can be safely ignored, but in the deep sea as pressure increases the density of gas increases to the point that its weight must be taken into account. According to the Ideal Gas Law:

$$PV = nRT \tag{2}$$

Where n is the number of moles of gas, R is the universal gas constant (8.314 J mol^{-1}. K^{-1}) and T is the absolute temperature (K). This equation can be modified to solve for density so that:

$$\rho_{PT} = MP/RT \tag{3}$$

Where ρ_{PT} is the density of gas at pressure P and temperature T, and M is the molecular weight of the gas. Density increases linearly with pressure so that at abyssal depths of 5000 m the density of oxygen would be 734 kg m^{-3}, providing only one quarter of the buoyancy compared with the same volume at the surface. Marshall (1960) calculated that at 7200 m depth the density of oxygen becomes equal to that of seawater (1028 kg m^{-3}), providing no buoyancy. This is close to the maximum depth at which fish with swim bladders have been observed (Pelster, 1997; Linley et al., 2017) and raises the question as to why fish would maintain swim bladder function in the face of negligible or zero benefit in terms of buoyancy. Calculating from published tables of oxygen compressibility, Alexander (1966) however pointed out that at the elevated pressures in the deep sea, oxygen does not behave as an ideal gas and the true density of oxygen at 5000 m is about 600 kg m^{-3} and at 7000 m about 700 kg m^{-3} providing appreciable buoyancy at these great depths.

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The Ideal Gas Law assumes that gas molecules have no volume and that there are no forces of attraction or repulsion between them. This is acceptable at low pressures but at high pressures real gases increasingly resist compression as the molecules come closer together. The Van der Waals equation is a modification of the Ideal Gas Law that attempts to take account of these effects:

$$\left(P + \frac{n^2\ a}{V^2}\right)(V - nb) = nRT \tag{4}$$

Where a and b are Van der Waals' constants that represent the putative molecular forces and molecular volume respectively. Using this equation Priede (2017) estimated the density of oxygen to be 574 kg m^{-3} at 5000 m and 643 kg m^{-3} at 7000 m depth, closely mirroring the calculations of Alexander (1966) and confirming that an oxygen-filled bladder can provide buoyancy at great depths beyond the limit predicted by the Ideal Gas Law. The Van der Waals equation however has limitations and numerous modifications have been proposed to provide more accurate predictions; for example Myers (1948) fits an equation with 3 virial coefficients for oxygen at pressures up to 340 bar.

Apart from the calculations by Alexander (1966) and Priede (2017) there have been no attempts to accurately quantify the buoyancy potential of gas-filled bladders at great ocean depths. This is despite the fact that experimental data on compression of gases have been available since the pioneering studies by Émile Amagat in the late 19th century showed that compressibility of gases decreases at high pressure (Amagat, 1893).

The present paper compiles information on the density of oxygen, nitrogen and air, derived from these published experimental data and compares these results with the standard gas equations in order to evaluate the potential performance of gas-filled buoyancy bladders. Pressure here is generally expressed in bar, which is defined as 0.1 MPa (Megapascals) and is equivalent to 0.9869 atm or approximately 10 m depth in seawater.

2. Methods

Since equations that describe behaviour of real gases at high pressure are either complex or subject to systematic errors, for practical purposes the calculations here are based on published experimental data. Following the work of Amagat (1893), others have measured compressibility of a wide range of gases, including mixtures, at an increasing range of temperatures and pressures. Hagenbach and Comings (1953) compare the results for Nitrogen at 50 °C at 50, 100 and 200 bar measured by five different laboratories and the average deviation between them is 0.236%. In experiments by Bartlett et al. (1928) the maximum deviation from the results of Amagat (1893) for nitrogen at pressures of 100–1000 bar was 0.60% with a mean of 0.30%. Results of gas compression experiments therefore are consistent within 1% and hence provide a good basis for prediction of gas behaviour in the deep sea. In these experiments samples of gas were compressed inside a tubular pipette of known volume with various methods of measuring the pressure. Bartlett et al. (1928) speculate that inaccuracies are caused by small differences in the purity of gas samples, uncertainties in manometer calibration, solubility of gases in the test chamber walls, temperature variations and other issues which would all be recognised in practical engineering design of a buoyancy system.

Most published gas compression tables do not give densities of gases. Since the exact mass of gas in the experiments is generally unknown, the results are usually presented in the form of the *Compressibility Factor* as a function of temperature and pressure. The *Compressibility Factor* is defined as the ratio of the volume of a real gas over the volume predicted by the Ideal Gas Law. Two versions are prevalent in the literature. In older tables including Amagat (1893), Washburn (1928) and Bartlett et al. (1928) the *Compressibility Factor* (C), sometimes shown as the product (PV), is defined in terms of Boyle's Law:

$$C = PV/P_0V_0 \tag{5}$$

Bartlett et al. (1928) show that density ρ at pressure P and temperature T is then given by:

$$\rho_{PT} = \rho_0 P\ (P_0V_0/PV) \tag{6}$$

If P_0 is the reference pressure, 1 bar, and V_0 is the reference volume, 1 m^3 at 0 °C then the equation becomes:

$$\rho_{PT} = \rho_0 P\ (1/C_{PT}) \tag{7}$$

The value of compressibility factor C_{PT} for the pressure and temperature of interest can be looked up in published tables. ρ_0 is the density of the gas at 0 °C and 1 bar. Here the values of ρ_0 were taken as 1.25 kg m^{-3} for Nitrogen, 1.428 kg m^{-3} for Oxygen and 1.286 kg m^{-3} for air.

In more recent tables such as Hagenbach and Comings (1953) and standard chemical engineering handbooks (Green and Perry, 2007) the *Compressibility Factor* (Z) is defined in terms of the Ideal Gas Law:

$$Z = PV/nRT \tag{8}$$

From which density is calculated as:

$$\rho_{PT} = MP/ZRT \tag{9}$$

Where M is the molecular weight of the gas, and T is absolute temperature (K). The values of C and Z from the same original data source are different in the two kinds of tables and great care must be taken to ensure that the correct compressibility factor is used.

The tables in Green and Perry (2007) only cover the pressure range up to 500 bar for industrial applications and the temperature range 75–1000 K at 50 K increments giving values at −23 °C and 27 °C which have no relevance for studies in the deep sea. For temperatures and pressures most appropriate for the deep sea, it is preferable to revert to the original work of Amagat (1893) which gives values of compressibility (C) up to 1000 atm and additional data at higher pressures by the same author tabulated in Washburn (1928). These two publications give data in 50 atm pressure increments from 100 atm up to 1100 atm; equivalent to full ocean depth *ca.* 11000 m (Jamieson, 2015). The original tables give values for nitrogen at 0 °C and 16.0 °C, oxygen at 0 °C and 15.65 °C and air at 0 °C and 15.70 °C. The C values at the higher temperature were normalised to 15 °C by linear interpolation prior to calculation of densities of these three gases at 0 ° and 15 °C using Eq. (7). These two temperatures span most of the likely range of temperatures encountered in the deep sea, from a minimum of −1.0 °C in the Arctic Ocean to a maximum of 15 °C in the deep basins of the Eastern Mediterranean (Priede, 2017). As no values are given for oxygen at 1100 atm these were filled in by linear interpolations between the C values for 1000 and 1200 atm. The final results were corrected to pressure in bar by linear interpolation assuming 1 bar = 0.9869 atm. For purposes of comparison, the densities of Nitrogen and Oxygen were also calculated in the conventional way using the Ideal Gas Law (2) and Van Der Waals Eq. (4). The values of the Van der Waals constants were taken as a = 0.1370 J m^3 mole^{-2}, b = 0.0387 L.mole^{-1} for Nitrogen and a = 0.1383 J m^3 mole^{-2}, b = 0.03186 L.mole^{-1} for Oxygen. The Van Der Waals calculations were done with the aid of Calistry (2017).

The upward buoyancy force (B) in Newtons (N) is calculated by Archimedes' principle as:

$$B = (\rho_{seawater} - \rho_{PT}) \times g \tag{10}$$

Where $\rho_{seawater}$ is the density of seawater and g is the earth's gravitational acceleration, 9.81 m s^{-2}. It is convenient to use the density difference

$$\delta\rho = \rho_{seawater} - \rho_{PT} \tag{11}$$

This is the payload (kg.m^{-3}) that can be supported by the gas

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