SCUBA Diving and Gas Laws[1]

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Students are often fascinated by extreme sports such as SCUBA (Self-Contained Underwater Breathing Apparatus) diving. This interest can be harnessed to teach an exciting lesson on gas laws and their importance to SCUBA diving. **Note:** ***SCUBA diving is a sport filled with many inherent dangers and requires specialized training and equipment. Do not attempt any diving activity without proper training and certification***.

SCUBA diving basics

The dry air we breathe every day is composed of 21% oxygen, 78% nitrogen, and <1% other gases. Its average pressure at sea level is 1 atm (14.7 psi). For SCUBA, this air is compressed into a SCUBA cylinder or "tank." SCUBA tanks can be made of steel or aluminum; each of these materials has pros and cons that impact the diver's decision on which type to use.

Figure 1 © J. Robert Patrick

The compressed air in the tank is delivered to the diver through a regulator, which reduces the pressure from the tank to match the ambient pressure. At the surface, ambient pressure is 1 atm and it increases by 1 atm for every 10 m in depth through which a diver descends. **Note:** *Other gas mixes such as nitrox (an oxygen/nitrogen mixture with a greater amount of oxygen than air), heliox (a helium and oxygen mixture), and trimix (a mixture of oxygen, nitrogen, and helium) or even pure oxygen are also used for technical diving, but those mixes go beyond the scope of this discussion.*

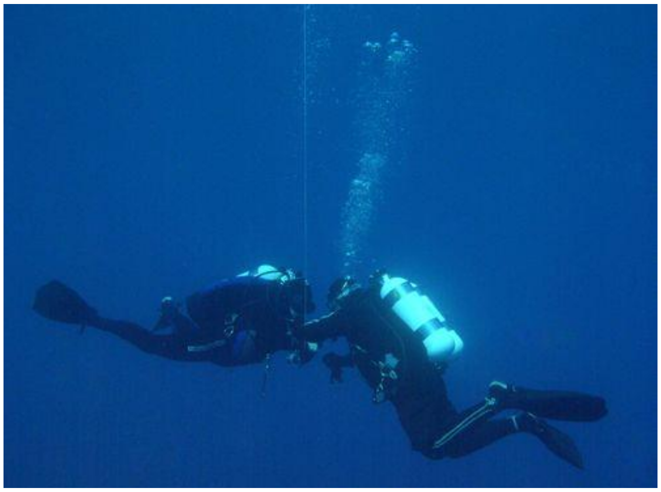

Figure 2 © John Simone

Boyle's law: $P_1V_1 = P_2V_2$

A fundamental rule of SCUBA diving is to "never hold your breath." A look at Boyle's law explains why this rule exists. When a diver inhales air from a SCUBA tank, the air that enters the diver's lungs is at ambient pressure. If a diver inhales from the tank on the surface, the pressure in her lungs will be at 1 atm. If she inhales air from her tank at a depth of 30 m (~99 ft), the pressure in her lungs will be 4 atm (30 m / 10 m/atm = 3 atm from the water plus 1 atm from the air at the surface = 4 atm). Assuming the diver's lung volume is 1 L, we can complete the left side of the equation for Boyle's law. If a diver at 30 m has 1 L (V_1) of air at a pressure of 4 atm (P_1) in her lungs and ascends to the surface (P_2) while holding her breath, the following equation applies:

$$4 \text{ atm} \times 1 \text{ L} = 1 \text{ atm} \times V_2$$

Solving for V_2, we see that the diver's lung volume would increase to 4 times its typical volume. This increase will result in severe damage to the lungs, which can be fatal. The increase of volume with a

[1] https://www.carolina.com/teacher-resources/Interactive/scuba-diving-and-gas-laws/tr29802.tr

decrease in pressure can also be seen in the gas bubbles exhaled by a diver as she rises to the surface. The exhaled air bubbles are small at depth and increase in size as they travel towards the surface. See Figure 2.

After working through this example, students often ask why free divers are able to dive to such extreme depths. Free divers fill their lungs at the surface with air at ambient pressure (P_1) then descend while holding their breath. The pressure change has the opposite impact on the volume of their lungs. A free diver diving to a depth of 30 m would have his lungs shrink to ¼ of their initial volume, which can be determined using the following equation:

$$1 \text{ atm} \times 1 \text{ L} = 4 \text{ atm} \times V_2$$

SCUBA instructors sometimes demonstrate this principle to their students by bringing a foam cup along on a dive. As the pressure increases with depth, the gas bubbles trapped in the foam decrease in volume, shrinking the cup.

Boyle's law also has implications on the amount of air used from the tank with each breath. At 10 m (2 atm) twice as many oxygen and nitrogen molecules are inhaled with each breath. Deeper dives require closer monitoring of a diver's air supply because the diver uses his supply more rapidly. Another question students often ask in this discussion is, "How is the SCUBA tank impacted by these changes in pressure?" Because the tank is a rigid container, its volume is not altered with the change in external pressure nor is the gas it contains affected.

Gay-Lussac's law: $P_1 / T_1 = P_2 / T_2$

In SCUBA diving, Gay-Lussac's law (sometimes referred to as Amontons' law of pressure-temperature) is most important in relation to the amount of breathable air in a tank. The pressure of an "empty" tank is low (around 500 psi), and the temperature is equal to the ambient temperature. SCUBA tanks made out of aluminum typically have a rated fill pressure of 3,000 psi.

A SCUBA tank is a rigid container, therefore its volume is held constant. When a tank is filled, additional oxygen and nitrogen molecules are added to the tank and the pressure and temperature increase. If a tank is filled rapidly to 3,000 psi (P_1), its temperature can rise to as much as 150° F (65.6° C). Since all gas laws use absolute temperatures, this temperature needs to be converted.

Most students know they can convert a Celsius temperature to an absolute temperature of Kelvin by adding 273. However, they are not likely to be aware that they can add 460 to a Fahrenheit temperature to convert it to a Rankine temperature, which is based on the Fahrenheit scale but with zero representing absolute zero. As the tank cools to ambient temperature (T_2) after the rapid fill, the gas pressure in the tank will also decrease. Assuming the ambient temperature is 70° F (21° C), the following equations can be used to determine the pressure at the lower temperature:

Using the Kelvin scale:

$T_1 = 65.6 + 273 = 338.6$ K

$T_2 = 21 + 273 = 294$ K

3,000 psi / 338.6 K = P_2 / 294 K

P_2 = 2,604 psi

Using the Rankine scale:

$T_1 = 150 + 460 = 610$ R

$T_2 = 70 + 460 = 530$ R

3,000 psi / 610 R = P_2 / 530 R

P_2 = 2,606 psi

Charles's law: $V_1 / T_1 = V_2 / T_2$

Charles's law is seldom relevant to diver safety; however, the implications of this law are responsible for an interesting phenomenon for divers using dry suits. A dry suit is a watertight garment worn by divers (typically over warm clothing) that serves to keep the diver warm by trapping a layer of air between the diver and the suit. Dry suits are usually worn in cold air and/or water temperatures.

During the dive, divers can add and remove air from their dry suits through their regulators. This allows them to adjust for changes in their suits' gas volumes due to pressure changes during assent and descent. If the air temperature is colder than the water temperature when the divers emerge at the end of the dive, they can become "vacuum sealed" in their suits due to the decrease in their suits' gas volumes. Divers can add air to the suits from their tanks, or unzip their suits, to release the "squeeze."

Dalton's law: $P_{Total} = P_1 + P_2 + P_3 \ldots$

Also known as Dalton's law of partial pressures, this law states that the total pressure of a gas mixture is equal to the sum of the partial pressures of its component gases. As mentioned earlier, dry air is a mixture composed of 21% oxygen and 78% nitrogen. Both of these gases can have negative impacts on a diver at high pressures. Low partial pressures of oxygen are also dangerous but are only an issue for technical diving, which is beyond the scope of this discussion.

Oxygen can become toxic to a diver when the partial pressure of the oxygen breathed is above 1.6 atm. Symptoms of oxygen toxicity can include changes in vision, dizziness/vertigo, and seizures, all of which can be problematic for a diver and can lead to death. To calculate at what depth a diver might begin to experience symptoms of oxygen toxicity when diving with compressed air, we need to first calculate at what air pressure the partial pressure of oxygen would be equal to 1.6 atm or greater.

At 1 atm of total pressure for air, oxygen would have a partial pressure of 0.21 atm. Therefore, the total pressure of the air would be 7.6 atm (1.6/0.21 atm) for the partial pressure of oxygen to be at 1.6 atm or greater. Remember that for each 10 m of depth the pressure increases by 1 atm, but the pressure at the surface is 1 atm, so the partial pressure of oxygen in air would be 1.6 atm at 66 m (216 ft).

Nitrogen narcosis can result from a diver's exposure to high partial pressures of nitrogen during her dive. Symptoms of nitrogen narcosis most closely resemble those of alcohol intoxication. These symptoms appear more gradually than those of oxygen toxicity but also increase with depth.

Henry's law

Henry's law states that the solubility of a liquid is directly proportional to the partial pressure of the gas above the liquid. The implication of this law for SCUBA diving is that as depth increases (and therefore pressure) the amount of a gas dissolved in the diver's blood will also increase. Oxygen is consumed by the body's physiological processes, but nitrogen is physiologically inert. The longer that a diver remains at depth, the more nitrogen is dissolved in his blood.

During long dives a considerable amount of nitrogen can be dissolved in the diver's bloodstream. When the diver ascends the partial pressure of nitrogen drops, and due to Henry's law the dissolved nitrogen begins to come out of solution. Nitrogen bubbles form in the diver's bloodstream, which can lead to decompression sickness (DCS).

The symptoms of DCS and their severity depend on where in the diver's body the bubbles migrate and can range from soreness in the joints or blisters under the skin to death. Treatment for DCS typically involves several sessions in a hyperbaric oxygen chamber. In their training, divers are taught to stay within dive time and depth limits to minimize their risk of DCS and to ascend slowly from every dive.