

Oxygen is a chemical element with symbol O and atomic number 8. It is a member of the chalcogen group on the periodic table and is a highly reactive nonmetal and oxidizing agent that readily forms compounds (notably oxides) with most elements. By mass, oxygen is the third-most abundant element in the universe, after hydrogen and helium. At standard temperature and pressure, two atoms of the element bind to form dioxygen, a colorless and odorless diatomic gas with the formula O₂. Diatomic oxygen gas constitutes 20.8% of the Earth's atmosphere. However, monitoring of atmospheric oxygen levels show a global downward trend, because of fossil-fuel burning. Oxygen is the most abundant element by mass in the Earth's crust as part of oxide compounds such as silicon dioxide, making up almost half of the crust's mass.

Many major classes of organic molecules in living organisms, such as proteins, nucleic acids, carbohydrates, and fats, contain oxygen, as do the major inorganic compounds that are constituents of animal shells, teeth, and bone. Most of the mass of living organisms is oxygen as it is a part of water, the major constituent of lifeforms. Oxygen is used in cellular respiration and released by photosynthesis, which uses the energy of sunlight to produce oxygen from water. It is too chemically reactive to remain a free element in air without being continuously replenished by the photosynthetic action of living organisms. Another form (allotrope) of oxygen, ozone (O₃), strongly absorbs UVB radiation and consequently the high-altitude ozone layer helps protect the biosphere from ultraviolet radiation, but is a pollutant near the surface where it is a by-product of smog. At even higher low earth orbit altitudes, sufficient atomic oxygen is present to cause erosion for spacecraft.

In the late 17th century, Robert Boyle proved that air is necessary for combustion. English chemist John Mayow (1641–1679) refined this work by showing that fire requires only a part of air that he called spiritus nitroaereus or just nitroaereus. In one experiment he found that placing either a mouse or a lit candle in a closed container over water caused the water to rise and replace one-fourteenth of the air's volume before extinguishing the subjects. From this he surmised that nitroaereus is consumed in both respiration and combustion.

In the meantime, on August 1, 1774, an experiment conducted by the British clergyman Joseph Priestley focused sunlight on mercuric oxide (HgO) inside a glass tube, which liberated a gas he named "dephlogisticated air". He noted that candles burned brighter in the gas and that a mouse was more active and lived longer while breathing it. After breathing the gas himself, he wrote: "The feeling of it to my lungs was not sensibly different from that of common air, but I fancied that my breast felt peculiarly light and easy for some time afterwards." Priestley published his findings in 1775 in a paper titled "An Account of Further Discoveries in Air" which was included in the second volume of his book titled *Experiments and Observations on Different Kinds of Air*. Because he published his findings first, Priestley is usually given priority in the discovery.

One of the first known experiments on the relationship between combustion and air was conducted by the 2nd century BCE Greek writer on mechanics, Philo of Byzantium. In his work *Pneumatica*, Philo observed that inverting a vessel over a burning candle and surrounding the vessel's neck with water resulted in some water rising into the neck. Philo incorrectly surmised that parts of the air in the vessel were converted into the classical element fire and thus were able to escape through pores in the glass. Many centuries later Leonardo da Vinci built on Philo's work by observing that a portion of air is consumed during combustion and respiration.

Highly concentrated sources of oxygen promote rapid combustion. Fire and explosion hazards exist when concentrated oxidants and fuels are brought into close proximity; an ignition event, such as heat or a spark, is needed to trigger combustion. Oxygen is the oxidant, not the fuel, but nevertheless the source of most of the chemical energy released in combustion. Combustion hazards also apply to compounds of oxygen with a high oxidative potential, such as peroxides, chlorates, nitrates, perchlorates, and dichromates because they can donate oxygen to a fire.

Concentrated O₂ will allow combustion to proceed rapidly and energetically. Steel pipes and storage vessels used to store and transmit both gaseous and liquid oxygen will act as a fuel; and therefore the design and manufacture of O₂ systems requires special training to ensure that ignition sources are minimized. The fire that killed the Apollo 1 crew in a launch pad test spread so rapidly because the capsule was pressurized with pure O₂ but at slightly more than atmospheric pressure, instead of the 1/3 normal pressure that would be used in a mission.[k]

Oxygen is present in the atmosphere in trace quantities in the form of carbon dioxide (CO₂). The Earth's crustal rock is composed in large part of oxides of silicon (silica SiO₂, as found in granite and quartz), aluminium (aluminium oxide Al₂O₃, in bauxite and corundum), iron (iron(III) oxide Fe₂O₃, in hematite and rust), and calcium carbonate (in limestone). The rest of the Earth's crust is also made of oxygen compounds, in particular various complex silicates (in silicate minerals). The Earth's mantle, of much larger mass than the crust, is largely composed of silicates of magnesium and iron.

John Dalton's original atomic hypothesis assumed that all elements were monatomic and that the atoms in compounds would normally have the simplest atomic ratios with respect to one another. For example, Dalton assumed that water's formula was HO, giving the atomic mass of oxygen as 8 times that of hydrogen, instead of the modern value of about 16. In 1805, Joseph Louis Gay-Lussac and Alexander von Humboldt showed that water is formed of two volumes of hydrogen and one volume of oxygen; and by 1811 Amedeo Avogadro had arrived at the correct interpretation of water's composition, based on what is now called Avogadro's law and the assumption of diatomic elemental molecules.[a]

Highly combustible materials that leave little residue, such as wood or coal, were thought to be made mostly of phlogiston; whereas non-combustible substances that corrode, such as iron, contained very little. Air did not play a role in phlogiston theory, nor were any initial quantitative experiments conducted to test the idea; instead, it was based on observations of what happens when something burns, that most common objects appear to become lighter and seem to lose something in the process. The fact that a substance like wood gains overall weight in burning was hidden by the buoyancy of the gaseous combustion products. Indeed, one of the first clues that the phlogiston theory was incorrect was that metals, too, gain weight in rusting (when they were supposedly losing phlogiston).

In this dioxygen, the two oxygen atoms are chemically bonded to each other. The bond can be variously described based on level of theory, but is reasonably and simply described as a covalent double bond that results from the filling of molecular orbitals formed from the atomic orbitals of the individual oxygen atoms, the filling of which results in a bond order of two. More specifically, the double bond is the result of sequential, low-to-high energy, or Aufbau, filling of orbitals, and the resulting cancellation of contributions from the 2s electrons, after sequential filling of the low σ and σ^* orbitals; σ overlap of the two atomic 2p orbitals that lie along the O-O molecular axis and π overlap of two pairs of atomic 2p orbitals perpendicular to the O-O molecular axis, and then cancellation of contributions from the remaining two of the six 2p electrons after their partial filling of the lowest π and π^* orbitals.

Oxygen was discovered independently by Carl Wilhelm Scheele, in Uppsala, in 1773 or earlier, and Joseph Priestley in Wiltshire, in 1774, but Priestley is often given priority because his work was published first. The name oxygen was coined in 1777 by Antoine Lavoisier, whose experiments with oxygen helped to discredit the then-popular phlogiston theory of combustion and corrosion. Its name derives from the Greek roots $\sigma\chi\acute{o}\varsigma$ oxy's, "acid", literally "sharp", referring to the sour taste of acids and $\gamma\epsilon\nu\acute{o}\varsigma$ -genes, "producer", literally "begetter", because at the time of naming, it was mistakenly thought that all acids required oxygen in their composition. Common uses of oxygen includes the production cycle of steel, plastics and textiles, brazing, welding and cutting of steels and other metals, rocket propellant, in oxygen therapy and life support systems in aircraft, submarines, spaceflight and diving.

This combination of cancellations and σ and π overlaps results in dioxygen's double bond character and reactivity, and a triplet electronic ground state. An electron configuration with two unpaired

electrons as found in dioxygen (see the filled π^* orbitals in the diagram), orbitals that are of equal energy—i.e., degenerate—is a configuration termed a spin triplet state. Hence, the ground state of the O₂ molecule is referred to as triplet oxygen.[b] The highest energy, partially filled orbitals are antibonding, and so their filling weakens the bond order from three to two. Because of its unpaired electrons, triplet oxygen reacts only slowly with most organic molecules, which have paired electron spins; this prevents spontaneous combustion.

In one experiment, Lavoisier observed that there was no overall increase in weight when tin and air were heated in a closed container. He noted that air rushed in when he opened the container, which indicated that part of the trapped air had been consumed. He also noted that the tin had increased in weight and that increase was the same as the weight of the air that rushed back in. This and other experiments on combustion were documented in his book *Sur la combustion en général*, which was published in 1777. In that work, he proved that air is a mixture of two gases; 'vital air', which is essential to combustion and respiration, and azote (Gk. ■ζωτον "lifeless"), which did not support either. Azote later became nitrogen in English, although it has kept the name in French and several other European languages.

Trioxxygen (O₃) is usually known as ozone and is a very reactive allotrope of oxygen that is damaging to lung tissue. Ozone is produced in the upper atmosphere when O₂ combines with atomic oxygen made by the splitting of O₂ by ultraviolet (UV) radiation. Since ozone absorbs strongly in the UV region of the spectrum, the ozone layer of the upper atmosphere functions as a protective radiation shield for the planet. Near the Earth's surface, it is a pollutant formed as a by-product of automobile exhaust. The metastable molecule tetraoxygen (O₄) was discovered in 2001, and was assumed to exist in one of the six phases of solid oxygen. It was proven in 2006 that this phase, created by pressurizing O₂ to 20 GPa, is in fact a rhombohedral O₈ cluster. This cluster has the potential to be a much more powerful oxidizer than either O₂ or O₃ and may therefore be used in rocket fuel. A metallic phase was discovered in 1990 when solid oxygen is subjected to a pressure of above 96 GPa and it was shown in 1998 that at very low temperatures, this phase becomes superconducting.

The common allotrope of elemental oxygen on Earth is called dioxygen, O₂. It is the form that is a major part of the Earth's atmosphere (see Occurrence). O₂ has a bond length of 121 pm and a bond energy of 498 kJ·mol^{−1}, which is smaller than the energy of other double bonds or pairs of single bonds in the biosphere and responsible for the exothermic reaction of O₂ with any organic molecule. Due to its energy content, O₂ is used by complex forms of life, such as animals, in cellular respiration (see Biological role). Other aspects of O₂ are covered in the remainder of this article.

In 1891 Scottish chemist James Dewar was able to produce enough liquid oxygen to study. The first commercially viable process for producing liquid oxygen was independently developed in 1895 by German engineer Carl von Linde and British engineer William Hampson. Both men lowered the temperature of air until it liquefied and then distilled the component gases by boiling them off one at a time and capturing them. Later, in 1901, oxyacetylene welding was demonstrated for the first time by burning a mixture of acetylene and compressed O₂. This method of welding and cutting metal later became common.

Oxygen is more soluble in water than nitrogen is. Water in equilibrium with air contains approximately 1 molecule of dissolved O₂ for every 2 molecules of N₂, compared to an atmospheric ratio of approximately 1:4. The solubility of oxygen in water is temperature-dependent, and about twice as much (14.6 mg·L^{−1}) dissolves at 0 °C than at 20 °C (7.6 mg·L^{−1}). At 25 °C and 1 standard atmosphere (101.3 kPa) of air, freshwater contains about 6.04 milliliters (mL) of oxygen per liter, whereas seawater contains about 4.95 mL per liter. At 5 °C the solubility increases to 9.0 mL (50% more than at 25 °C) per liter for water and 7.2 mL (45% more) per liter for sea water.

Oxygen is the most abundant chemical element by mass in the Earth's biosphere, air, sea and land. Oxygen is the third most abundant chemical element in the universe, after hydrogen and helium. About 0.9% of the Sun's mass is oxygen. Oxygen constitutes 49.2% of the Earth's crust by mass and is the major component of the world's oceans (88.8% by mass). Oxygen gas is the second most common component of the Earth's atmosphere, taking up 20.8% of its volume and 23.1% of its mass (some 1015 tonnes).[d] Earth is unusual among the planets of the Solar System in having such a high concentration of oxygen gas in its atmosphere: Mars (with 0.1% O₂ by volume) and Venus have far lower concentrations. The O₂ surrounding these other planets is produced solely by ultraviolet radiation impacting oxygen-containing molecules such as carbon dioxide.

By the late 19th century scientists realized that air could be liquefied, and its components isolated, by compressing and cooling it. Using a cascade method, Swiss chemist and physicist Raoul Pierre Pictet evaporated liquid sulfur dioxide in order to liquefy carbon dioxide, which in turn was evaporated to cool oxygen gas enough to liquefy it. He sent a telegram on December 22, 1877 to the French Academy of Sciences in Paris announcing his discovery of liquid oxygen. Just two days later, French physicist Louis Paul Cailletet announced his own method of liquefying molecular oxygen. Only a few drops of the liquid were produced in either case so no meaningful analysis could be conducted. Oxygen was liquified in stable state for the first time on March 29, 1883 by Polish scientists from Jagiellonian University, Zygmunt Wróblewski and Karol Olszewski.

Planetary geologists have measured different abundances of oxygen isotopes in samples from the Earth, the Moon, Mars, and meteorites, but were long unable to obtain reference values for the isotope ratios in the Sun, believed to be the same as those of the primordial solar nebula. Analysis of a silicon wafer exposed to the solar wind in space and returned by the crashed Genesis spacecraft has shown that the Sun has a higher proportion of oxygen-16 than does the Earth. The measurement implies that an unknown process depleted oxygen-16 from the Sun's disk of protoplanetary material prior to the coalescence of dust grains that formed the Earth.

Singlet oxygen is a name given to several higher-energy species of molecular O₂ in which all the electron spins are paired. It is much more reactive towards common organic molecules than is molecular oxygen per se. In nature, singlet oxygen is commonly formed from water during photosynthesis, using the energy of sunlight. It is also produced in the troposphere by the photolysis of ozone by light of short wavelength, and by the immune system as a source of active oxygen. Carotenoids in photosynthetic organisms (and possibly also in animals) play a major role in absorbing energy from singlet oxygen and converting it to the unexcited ground state before it can cause harm to tissues.

Paleoclimatologists measure the ratio of oxygen-18 and oxygen-16 in the shells and skeletons of marine organisms to determine what the climate was like millions of years ago (see oxygen isotope ratio cycle). Seawater molecules that contain the lighter isotope, oxygen-16, evaporate at a slightly faster rate than water molecules containing the 12% heavier oxygen-18; this disparity increases at lower temperatures. During periods of lower global temperatures, snow and rain from that evaporated water tends to be higher in oxygen-16, and the seawater left behind tends to be higher in oxygen-18. Marine organisms then incorporate more oxygen-18 into their skeletons and shells than they would in a warmer climate. Paleoclimatologists also directly measure this ratio in the water molecules of ice core samples that are up to several hundreds of thousands of years old.

Oxygen presents two spectrophotometric absorption bands peaking at the wavelengths 687 and 760 nm. Some remote sensing scientists have proposed using the measurement of the radiance coming from vegetation canopies in those bands to characterize plant health status from a satellite platform. This approach exploits the fact that in those bands it is possible to discriminate the vegetation's reflectance from its fluorescence, which is much weaker. The measurement is technically difficult owing to the low signal-to-noise ratio and the physical structure of vegetation; but it has been proposed as a

possible method of monitoring the carbon cycle from satellites on a global scale.

In the triplet form, O₂ molecules are paramagnetic. That is, they impart magnetic character to oxygen when it is in the presence of a magnetic field, because of the spin magnetic moments of the unpaired electrons in the molecule, and the negative exchange energy between neighboring O₂ molecules. Liquid oxygen is attracted to a magnet to a sufficient extent that, in laboratory demonstrations, a bridge of liquid oxygen may be supported against its own weight between the poles of a powerful magnet.[c]

Reactive oxygen species, such as superoxide ion (O₂⁻) and hydrogen peroxide (H₂O₂), are dangerous by-products of oxygen use in organisms. Parts of the immune system of higher organisms create peroxide, superoxide, and singlet oxygen to destroy invading microbes. Reactive oxygen species also play an important role in the hypersensitive response of plants against pathogen attack. Oxygen is toxic to obligately anaerobic organisms, which were the dominant form of early life on Earth until O₂ began to accumulate in the atmosphere about 2.5 billion years ago during the Great Oxygenation Event, about a billion years after the first appearance of these organisms.

Oxygen condenses at 90.20 K (−182.95 °C, −297.31 °F), and freezes at 54.36 K (−218.79 °C, −361.82 °F). Both liquid and solid O₂ are clear substances with a light sky-blue color caused by absorption in the red (in contrast with the blue color of the sky, which is due to Rayleigh scattering of blue light). High-purity liquid O₂ is usually obtained by the fractional distillation of liquefied air. Liquid oxygen may also be produced by condensation out of air, using liquid nitrogen as a coolant. It is a highly reactive substance and must be segregated from combustible materials.

Free oxygen also occurs in solution in the world's water bodies. The increased solubility of O₂ at lower temperatures (see Physical properties) has important implications for ocean life, as polar oceans support a much higher density of life due to their higher oxygen content. Water polluted with plant nutrients such as nitrates or phosphates may stimulate growth of algae by a process called eutrophication and the decay of these organisms and other biomaterials may reduce amounts of O₂ in eutrophic water bodies. Scientists assess this aspect of water quality by measuring the water's biochemical oxygen demand, or the amount of O₂ needed to restore it to a normal concentration.

Free oxygen gas was almost nonexistent in Earth's atmosphere before photosynthetic archaea and bacteria evolved, probably about 3.5 billion years ago. Free oxygen first appeared in significant quantities during the Paleoproterozoic eon (between 3.0 and 2.3 billion years ago). For the first billion years, any free oxygen produced by these organisms combined with dissolved iron in the oceans to form banded iron formations. When such oxygen sinks became saturated, free oxygen began to outgas from the oceans 3–2.7 billion years ago, reaching 10% of its present level around 1.7 billion years ago.

The unusually high concentration of oxygen gas on Earth is the result of the oxygen cycle. This biogeochemical cycle describes the movement of oxygen within and between its three main reservoirs on Earth: the atmosphere, the biosphere, and the lithosphere. The main driving factor of the oxygen cycle is photosynthesis, which is responsible for modern Earth's atmosphere. Photosynthesis releases oxygen into the atmosphere, while respiration and decay remove it from the atmosphere. In the present equilibrium, production and consumption occur at the same rate of roughly 1/2000th of the entire atmospheric oxygen per year.

The other major method of producing O₂ gas involves passing a stream of clean, dry air through one bed of a pair of identical zeolite molecular sieves, which absorbs the nitrogen and delivers a gas stream that is 90% to 93% O₂. Simultaneously, nitrogen gas is released from the other nitrogen-saturated zeolite bed, by reducing the chamber operating pressure and diverting part of the oxygen gas from the producer bed through it, in the reverse direction of flow. After a set cycle time the operation of the two beds is interchanged, thereby allowing for a continuous supply of gaseous oxygen to be pumped through a pipeline. This is known as pressure swing adsorption. Oxygen gas is increasingly obtained by

these non-cryogenic technologies (see also the related vacuum swing adsorption).

Oxygen gas can also be produced through electrolysis of water into molecular oxygen and hydrogen. DC electricity must be used: if AC is used, the gases in each limb consist of hydrogen and oxygen in the explosive ratio 2:1. Contrary to popular belief, the 2:1 ratio observed in the DC electrolysis of acidified water does not prove that the empirical formula of water is H₂O unless certain assumptions are made about the molecular formulae of hydrogen and oxygen themselves. A similar method is the electrocatalytic O₂ evolution from oxides and oxoacids. Chemical catalysts can be used as well, such as in chemical oxygen generators or oxygen candles that are used as part of the life-support equipment on submarines, and are still part of standard equipment on commercial airliners in case of depressurization emergencies. Another air separation technology involves forcing air to dissolve through ceramic membranes based on zirconium dioxide by either high pressure or an electric current, to produce nearly pure O₂ gas.

Oxygen, as a supposed mild euphoric, has a history of recreational use in oxygen bars and in sports. Oxygen bars are establishments, found in Japan, California, and Las Vegas, Nevada since the late 1990s that offer higher than normal O₂ exposure for a fee. Professional athletes, especially in American football, also sometimes go off field between plays to wear oxygen masks in order to get a "boost" in performance. The pharmacological effect is doubtful; a placebo effect is a more likely explanation. Available studies support a performance boost from enriched O₂ mixtures only if they are breathed during aerobic exercise.

Hyperbaric (high-pressure) medicine uses special oxygen chambers to increase the partial pressure of O₂ around the patient and, when needed, the medical staff. Carbon monoxide poisoning, gas gangrene, and decompression sickness (the 'bends') are sometimes treated using these devices. Increased O₂ concentration in the lungs helps to displace carbon monoxide from the heme group of hemoglobin. Oxygen gas is poisonous to the anaerobic bacteria that cause gas gangrene, so increasing its partial pressure helps kill them. Decompression sickness occurs in divers who decompress too quickly after a dive, resulting in bubbles of inert gas, mostly nitrogen and helium, forming in their blood. Increasing the pressure of O₂ as soon as possible is part of the treatment.

Uptake of O₂ from the air is the essential purpose of respiration, so oxygen supplementation is used in medicine. Treatment not only increases oxygen levels in the patient's blood, but has the secondary effect of decreasing resistance to blood flow in many types of diseased lungs, easing work load on the heart. Oxygen therapy is used to treat emphysema, pneumonia, some heart disorders (congestive heart failure), some disorders that cause increased pulmonary artery pressure, and any disease that impairs the body's ability to take up and use gaseous oxygen.

Due to its electronegativity, oxygen forms chemical bonds with almost all other elements to give corresponding oxides. The surface of most metals, such as aluminium and titanium, are oxidized in the presence of air and become coated with a thin film of oxide that passivates the metal and slows further corrosion. Many oxides of the transition metals are non-stoichiometric compounds, with slightly less metal than the chemical formula would show. For example, the mineral FeO (wüstite) is written as Fe_{1-x}O, where x is usually around 0.05.

People who climb mountains or fly in non-pressurized fixed-wing aircraft sometimes have supplemental O₂ supplies.[h] Passengers traveling in (pressurized) commercial airplanes have an emergency supply of O₂ automatically supplied to them in case of cabin depressurization. Sudden cabin pressure loss activates chemical oxygen generators above each seat, causing oxygen masks to drop. Pulling on the masks "to start the flow of oxygen" as cabin safety instructions dictate, forces iron filings into the sodium chlorate inside the canister. A steady stream of oxygen gas is then produced by the exothermic reaction.

Oxygen storage methods include high pressure oxygen tanks, cryogenics and chemical compounds. For reasons of economy, oxygen is often transported in bulk as a liquid in specially insulated tankers, since one liter of liquefied oxygen is equivalent to 840 liters of gaseous oxygen at atmospheric pressure and 20 °C (68 °F). Such tankers are used to refill bulk liquid oxygen storage containers, which stand outside hospitals and other institutions with a need for large volumes of pure oxygen gas. Liquid oxygen is passed through heat exchangers, which convert the cryogenic liquid into gas before it enters the building. Oxygen is also stored and shipped in smaller cylinders containing the compressed gas; a form that is useful in certain portable medical applications and oxy-fuel welding and cutting.

Among the most important classes of organic compounds that contain oxygen are (where "R" is an organic group): alcohols (R-OH); ethers (R-O-R); ketones (R-CO-R); aldehydes (R-CO-H); carboxylic acids (R-COOH); esters (R-COO-R); acid anhydrides (R-CO-O-CO-R); and amides (R-C(O)-NR₂). There are many important organic solvents that contain oxygen, including: acetone, methanol, ethanol, isopropanol, furan, THF, diethyl ether, dioxane, ethyl acetate, DMF, DMSO, acetic acid, and formic acid. Acetone ((CH₃)₂CO) and phenol (C₆H₅OH) are used as feeder materials in the synthesis of many different substances. Other important organic compounds that contain oxygen are: glycerol, formaldehyde, glutaraldehyde, citric acid, acetic anhydride, and acetamide. Epoxides are ethers in which the oxygen atom is part of a ring of three atoms.

The element is found in almost all biomolecules that are important to (or generated by) life. Only a few common complex biomolecules, such as squalene and the carotenes, contain no oxygen. Of the organic compounds with biological relevance, carbohydrates contain the largest proportion by mass of oxygen. All fats, fatty acids, amino acids, and proteins contain oxygen (due to the presence of carbonyl groups in these acids and their ester residues). Oxygen also occurs in phosphate (PO₃⁻⁴) groups in the biologically important energy-carrying molecules ATP and ADP, in the backbone and the purines (except adenine) and pyrimidines of RNA and DNA, and in bones as calcium phosphate and hydroxylapatite.

Oxygen toxicity to the lungs and central nervous system can also occur in deep scuba diving and surface supplied diving. Prolonged breathing of an air mixture with an O₂ partial pressure more than 60 kPa can eventually lead to permanent pulmonary fibrosis. Exposure to a O₂ partial pressures greater than 160 kPa (about 1.6 atm) may lead to convulsions (normally fatal for divers). Acute oxygen toxicity (causing seizures, its most feared effect for divers) can occur by breathing an air mixture with 21% O₂ at 66 m or more of depth; the same thing can occur by breathing 100% O₂ at only 6 m.

Breathing pure O₂ in space applications, such as in some modern space suits, or in early spacecraft such as Apollo, causes no damage due to the low total pressures used. In the case of spacesuits, the O₂ partial pressure in the breathing gas is, in general, about 30 kPa (1.4 times normal), and the resulting O₂ partial pressure in the astronaut's arterial blood is only marginally more than normal sea-level O₂ partial pressure (for more information on this, see space suit and arterial blood gas).

Oxygen gas (O₂) can be toxic at elevated partial pressures, leading to convulsions and other health problems.[j] Oxygen toxicity usually begins to occur at partial pressures more than 50 kilopascals (kPa), equal to about 50% oxygen composition at standard pressure or 2.5 times the normal sea-level O₂ partial pressure of about 21 kPa. This is not a problem except for patients on mechanical ventilators, since gas supplied through oxygen masks in medical applications is typically composed of only 30%–50% O₂ by volume (about 30 kPa at standard pressure). (although this figure also is subject to wide variation, depending on type of mask).