Oxygen

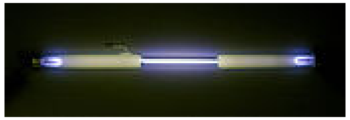
From Wikipedia, the free encyclopedia

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 oxides with most elements as well as other compounds.[3] By mass, oxygen is the third-most abundant element in the universe, after hydrogen and helium.[4] At standard temperature and pressure, two atoms of the element bind to form dioxygen, a colorless and odorless diatomic gas with the formula O_2. This is an important part of the atmosphere and diatomic oxygen gas constitutes 20.8% of the Earth's atmosphere.[5] Additionally, as oxides the element makes up almost half of the Earth's crust.[6]

Dioxygen is used in cellular respiration and many major classes of organic molecules in living organisms contain oxygen, such as proteins, nucleic acids, carbohydrates, and fats, as do the major constituent inorganic compounds of animal shells, teeth, and bone. Most of the mass of living organisms is oxygen as a component of water, the major constituent of lifeforms. Conversely, oxygen is continuously replenished by photosynthesis, which uses the energy of sunlight to produce oxygen from water and carbon dioxide. Oxygen 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 (O_3), strongly absorbs ultraviolet UVB radiation and the high-altitude ozone layer helps protect the biosphere from ultraviolet radiation. But ozone is a pollutant near the surface where it is a by-product of smog. At low earth orbit altitudes, sufficient atomic oxygen is present to cause corrosion of spacecraft.[7]

Common use of oxygen includes residential heating, internal combustion engines, production of steel, plastics and textiles, brazing, welding and cutting of steels and other metals, rocket propellant, oxygen therapy, and life support systems in aircraft, submarines, spaceflight and diving.

Oxygen, $_8$O



Spectral lines of oxygen

General properties	
Name, symbol	oxygen, O
Pronunciation	/ˈɒksɪdʒən/ ***ok**-si-jən*
Appearance	gas: colorless liquid: pale blue
Oxygen in the periodic table	
Atomic number (*Z*)	8
Group, block	group 16 (chalcogens), p-block
Period	period 2
Element category	□ diatomic nonmetal
Standard atomic weight (A_r)	15.999[1] (15.99903–15.99977)[2]
Electron configuration	[He] $2s^2$ $2p^4$
per shell	2, 6
Physical properties	
Phase	gas
Melting point	54.36 K (−218.79 °C, −361.82 °F)

Characteristics

Properties and molecular structure

At standard temperature and pressure, oxygen is a colorless, odorless, and tasteless gas with the molecular formula O_2, referred to as dioxygen.[28]

As *dioxygen*, two oxygen atoms are chemically bound 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.[27]

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 is found in dioxygen orbitals (see the filled π* orbitals in the diagram) that are of equal energy—i.e., degenerate—is a configuration termed a spin triplet state. Hence, the ground state of the O_2 molecule is referred to as triplet oxygen.[29][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.[30]

In the triplet form, O_2 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 O_2 molecules.[23] Liquid oxygen is so magnetic that, in laboratory demonstrations, a bridge of liquid oxygen may be supported against its own weight between the poles of a powerful magnet.[31][c]

Singlet oxygen is a name given to several higher-energy species of molecular O_2 in which all the electron spins are paired. It is much more reactive with common organic molecules than is molecular oxygen per se. In nature, singlet oxygen is commonly

Boiling point	90.188 K (−182.962 °C, −297.332 °F)
Density at stp (0 °C and 101.325 kPa)	1.429 g/L
when liquid, at b.p.	1.141 g/cm^3
Triple point	54.361 K, 0.1463 kPa
Critical point	154.581 K, 5.043 MPa
Heat of fusion	(O_2) 0.444 kJ/mol
Heat of vaporization	(O_2) 6.82 kJ/mol
Molar heat capacity	(O_2) 29.378 J/(mol·K)

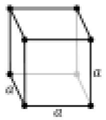
Vapor pressure

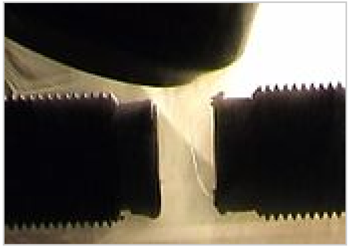
P (Pa)	**1**	**10**	**100**	**1 k**	**10 k**	**100 k**
at T (K)				61	73	90

Atomic properties

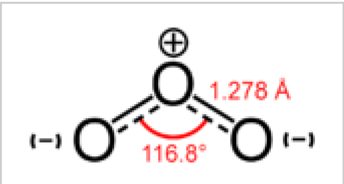
Oxidation states	2, 1, −1, **−2**
Electronegativity	Pauling scale: 3.44
Ionization energies	1st: 1313.9 kJ/mol 2nd: 3388.3 kJ/mol 3rd: 5300.5 kJ/mol (more)
Covalent radius	66±2 pm
Van der Waals radius	152 pm

Miscellanea

Crystal structure	cubic 
Speed of sound	330 m/s (gas, at 27 °C)



A trickle of liquid oxygen is deflected by a magnetic field, illustrating its paramagnetic property



Ozone is a rare gas on Earth found mostly in the stratosphere.

formed from water during photosynthesis, using the energy of sunlight.[32] It is also produced in the troposphere by the photolysis of ozone by light of short wavelength,[33] and by the immune system as a source of active oxygen.[34] Carotenoids in photosynthetic organisms (and possibly 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.[35]

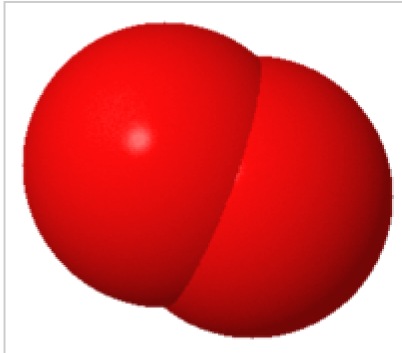
Thermal conductivity	26.58×10^{-3} W/(m·K)
Magnetic ordering	paramagnetic
CAS Number	7782-44-7
History	
Discovery	Carl Wilhelm Scheele (1771)
Named by	Antoine Lavoisier (1777)

Most stable isotopes of oxygen

iso	NA	half-life	DM	DE (MeV)	DP
^{16}O	99.76%	is stable with 8 neutrons			
^{17}O	0.04%	is stable with 9 neutrons			
^{18}O	0.20%	is stable with 10 neutrons			

Allotropes

The common allotrope of elemental oxygen on Earth is called dioxygen, O_2, the major part of the Earth's atmospheric oxygen (see Occurrence). O_2 has a bond length of 121 pm and a bond energy of 498 kJ·mol^{-1},[36] 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 O_2 with any organic molecule.[30][37] Due to its energy content, O_2 is used by complex forms of life, such as animals, in cellular respiration (see Biological role). Other aspects of O_2 are covered in the remainder of this article.

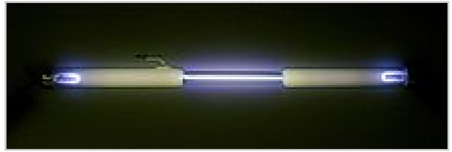


Space-filling model representation of dioxygen (O_2) molecule

Trioxygen (O_3) is usually known as ozone and is a very reactive allotrope of oxygen that is damaging to lung tissue.[38] Ozone is produced in the upper atmosphere when O_2 combines with atomic oxygen made by the splitting of O_2 by ultraviolet (UV) radiation.[8] 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.[8] Near the Earth's surface, it is a pollutant formed as a by-product of automobile exhaust.[38] The metastable molecule tetraoxygen (O_4) was discovered in 2001,[39][40] 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 O_2 to 20 GPa, is in fact a rhombohedral O_8 cluster.[41] This cluster has the potential to be a much more powerful oxidizer than either O_2 or O_3 and may therefore be used in rocket fuel.[39][40] A metallic phase was discovered in 1990 when solid oxygen is subjected to a pressure of above 96 GPa[42] and it was shown in 1998 that at very low temperatures, this phase becomes superconducting.[43]

Physical properties

Oxygen dissolves more readily in water than nitrogen, and in freshwater more readily than seawater. Water in equilibrium with air contains approximately 1 molecule of dissolved O_2 for every 2 molecules of N_2 (1:2), compared with 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}).[14][44] At 25 °C and 1 standard atmosphere (101.3 kPa) of air, freshwater contains about 6.04 milliliters (mL) of oxygen per liter, and seawater contains about 4.95 mL per liter.[45] 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 discharge (spectrum) tube. The green color is similar to the color of an "aurora borealis"

Oxygen condenses at 90.20 K (−182.95 °C, −297.31 °F), and freezes at 54.36 K (−218.79 °C, −361.82 °F).[46] Both liquid and solid O_2 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 O_2 is usually obtained by the fractional distillation of liquefied air.[47] Liquid oxygen may also be condensed from air using liquid nitrogen as a coolant.[48]

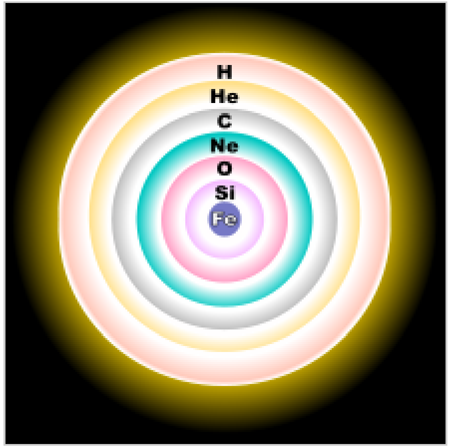
Oxygen is a highly reactive substance and must be segregated from combustible materials.[48]

The spectroscopy of molecular oxygen is associated with the atmospheric processes of aurora, airglow and nightglow.[49] The absorption in the Herzberg continuum and Schumann–Runge bands in the ultraviolet produces atomic oxygen that is important in the chemistry of the middle atmosphere.[50] Excited state singlet molecular oxygen is responsible for red chemiluminescence in solution.[51]

Isotopes and stellar origin

Naturally occurring oxygen is composed of three stable isotopes, ^{16}O, ^{17}O, and ^{18}O, with ^{16}O being the most abundant (99.762% natural abundance).[52]

Most ^{16}O is synthesized at the end of the helium fusion process in massive stars but some is made in the neon burning process.[53] ^{17}O is primarily made by the burning of hydrogen into helium during the CNO cycle, making it a common isotope in the hydrogen burning zones of stars.[53] Most ^{18}O is produced when ^{14}N (made abundant from CNO burning) captures a ^{4}He nucleus, making ^{18}O common in the helium-rich zones of evolved, massive stars.[53]



Late in a massive star's life, ^{16}O concentrates in the O-shell, ^{17}O in the H-shell and ^{18}O in the He-shell.

Fourteen radioisotopes have been characterized. The most stable are ^{15}O with a half-life of 122.24 seconds and ^{14}O with a half-life of 70.606 seconds.[52] All of the remaining radioactive isotopes have half-lives that are less than 27 s and the majority of these have half-lives that are less than 83 milliseconds.[52] The most common decay mode of the isotopes lighter than ^{16}O is β^+ decay[54][55][56] to yield nitrogen, and the most common mode for the isotopes heavier than ^{18}O is beta decay to yield fluorine.[52]

Occurrence

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.[4] About 0.9% of the Sun's mass is oxygen.[5] Oxygen constitutes 49.2% of the Earth's crust by mass[6] as part of oxide compounds such as silicon dioxide and is the most abundant element by mass in the Earth's crust. It is also the major component of the world's oceans (88.8% by mass).[5] 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 10^{15} tonnes).[5][58][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% O_2 by volume) and Venus have much less. The O_2 surrounding those planets is produced solely by ultraviolet radiation on oxygen-containing molecules such as carbon dioxide.

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, decay, and combustion 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.

Free oxygen also occurs in solution in the world's water bodies. The increased solubility of O_2 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.[59] 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 the O_2 content in eutrophic water bodies. Scientists assess this aspect of water quality by measuring the water's biochemical oxygen demand, or the amount of O_2 needed to restore it to a normal concentration.[60]

Analysis

Paleoclimatologists measure the ratio of oxygen-18 and oxygen-16 in the shells and skeletons of marine organisms to determine the climate 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, and this disparity increases at lower temperatures.[61] 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.[61] Paleoclimatologists also directly measure this ratio in the water molecules of ice core samples as old as hundreds of thousands of years.

Planetary geologists have measured the relative quantities 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.[62]

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.[63] 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.

Source

- Wikipedia: Oxygen (https://en.wikipedia.org/wiki/Oxygen)