Carbon

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Carbon (from Latin: *carbo* "coal") is a chemical element with symbol **C** and atomic number 6. It is nonmetallic and tetravalent—making four electrons available to form covalent chemical bonds. Three isotopes occur naturally, ^{12}C and ^{13}C being stable while ^{14}C is a radioactive isotope, decaying with a half-life of about 5,730 years.[14] Carbon is one of the few elements known since antiquity.[15]

Carbon is the 15th most abundant element in the Earth's crust, and the fourth most abundant element in the universe by mass after hydrogen, helium, and oxygen. Carbon's abundance, its unique diversity of organic compounds, and its unusual ability to form polymers at the temperatures commonly encountered on Earth enables this element to serve as a common element of all known life. It is the second most abundant element in the human body by mass (about 18.5%) after oxygen.[16]

The atoms of carbon can be bonded together in different ways, termed allotropes of carbon. The best known are graphite, diamond, and amorphous carbon.[17] The physical properties of carbon vary widely with the allotropic form. For example, graphite is opaque and black while diamond is highly transparent. Graphite is soft enough to form a streak on paper (hence its name, from the Greek verb "γράφειν" which means "to write"), while diamond is the hardest naturally occurring material known. Graphite is a good electrical conductor while diamond has a low electrical conductivity. Under normal conditions, diamond, carbon nanotubes, and graphene have the highest thermal conductivities of all known materials. All carbon allotropes are solids under normal conditions, with graphite being the most thermodynamically stable form. They are chemically resistant and require high temperature to react even with oxygen.

Carbon, $_{6}C$

Graphite (left) and diamond (right), the two most well-known allotropes of carbon

Spectral lines of carbon

General properties

Name, symbol	carbon, C
Pronunciation	/ˈkɑːrbən/ *KAR-bən*
Allotropes	graphite, diamond
Appearance	graphite: black diamond: clear

Carbon in the periodic table

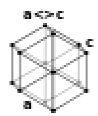
Atomic number (*Z*)	6
Group, block	group 14 (carbon group), p-block
Period	period 2
Element category	□ polyatomic nonmetal, sometimes considered a metalloid
Standard atomic weight (A_r)	12.011[1] (12.0096–12.0116)[2]
Electron	[He] $2s^2$ $2p^2$

The most common oxidation state of carbon in inorganic compounds is +4, while +2 is found in carbon monoxide and transition metal carbonyl complexes. The largest sources of inorganic carbon are limestones, dolomites and carbon dioxide, but significant quantities occur in organic deposits of coal, peat, oil, and methane clathrates. Carbon forms a vast number of compounds, more than any other element, with almost ten million compounds described to date,[18] and yet that number is but a fraction of the number of theoretically possible compounds under standard conditions. For this reason, carbon has often been referred to as the "king of the elements".[19]

Characteristics

The allotropes of carbon include graphite, one of the softest known substances, and diamond, the hardest naturally occurring substance. It bonds readily with other small atoms including other carbon atoms, and is capable of forming multiple stable covalent bonds with suitable, multivalent atoms. Carbon is known to form almost ten million different compounds, a large majority of all chemical compounds.[18] Carbon also has the highest sublimation point of all elements. At atmospheric pressure it has no melting point as its triple point is at 10.8 ± 0.2 MPa and 4,600 ± 300 K (~4,330 °C or 7,820 °F),[4][5] so it sublimes at about 3,900 K.[20][21] Graphite is much more reactive than diamond at standard conditions, despite being more thermodynamically stable, as its delocalised pi system is much more vulnerable to attack. For example, graphite can be oxidised by hot concentrated nitric acid at standard conditions to mellitic acid, $C_6(CO_2H)_6$, which preserves the hexagonal units of graphite while breaking up the larger structure.[22]

Carbon sublimes in a carbon arc which has a temperature of about 5,800 K (5,530 °C; 9,980 °F). Thus, irrespective of its allotropic form, carbon remains solid at higher temperatures than the highest melting point metals such as tungsten or rhenium. Although

configuration	
per shell	2, 4
Physical properties	
Phase	solid
Sublimation point	3915 K (3642 °C, 6588 °F)
Density near r.t.	amorphous: 1.8–2.1 g/cm^3[3] graphite: 2.267 g/cm^3 diamond: 3.515 g/cm^3
Triple point	4600 K, 10,800 kPa[4][5]
Heat of fusion	graphite: 117 kJ/mol
Molar heat capacity	graphite: 8.517 J/(mol·K) diamond: 6.155 J/(mol·K)
Atomic properties	
Oxidation states	**+4**, +3,[6] +2, +1,[7] 0, −1, −2, −3, **−4**[8] (a mildly acidic oxide)
Electronegativity	Pauling scale: 2.55
Ionization energies	1st: 1086.5 kJ/mol 2nd: 2352.6 kJ/mol 3rd: 4620.5 kJ/mol (more)
Covalent radius	sp^3: 77 pm sp^2: 73 pm sp: 69 pm
Van der Waals radius	170 pm
Miscellanea	
Crystal structure	graphite: simple hexagonal (black) 
Crystal structure	diamond: face-centered diamond-cubic

thermodynamically prone to oxidation, carbon resists oxidation more effectively than elements such as iron and copper that are weaker reducing agents at room temperature.

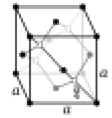
Carbon is the sixth element, with a ground-state electron configuration of $1s^2 2s^2 2p^2$, of which the four outer electrons are valence electrons. Its first four ionisation energies, 1086.5, 2352.6, 4620.5 and 6222.7 kJ/mol, are much higher than those of the heavier group 14 elements. The electronegativity of carbon is 2.5, significantly higher than the heavier group 14 elements (1.8–1.9), but close to most of the nearby nonmetals as well as some of the second- and third-row transition metals. Carbon's covalent radii are normally taken as 77.2 pm (C–C), 66.7 pm (C=C) and 60.3 pm (C≡C), although these may vary depending on coordination number and what the carbon is bonded to. In general, covalent radius decreases with lower coordination number and higher bond order.[23]

Carbon compounds form the basis of all known life on Earth, and the carbon-nitrogen cycle provides some of the energy produced by the Sun and other stars. Although it forms an extraordinary variety of compounds, most forms of carbon are comparatively unreactive under normal conditions. At standard temperature and pressure, it resists all but the strongest oxidizers. It does not react with sulfuric acid, hydrochloric acid, chlorine or any alkalis. At elevated temperatures, carbon reacts with oxygen to form carbon oxides, and will rob oxygen from metal oxides to leave the elemental metal. This exothermic reaction is used in the iron and steel industry to smelt iron and to control the carbon content of steel:

$$Fe_3O_4 + 4\ C_{(s)} \rightarrow 3\ Fe_{(s)} + 4\ CO_{(g)}$$

with sulfur to form carbon disulfide and with steam in the coal-gas reaction:

$$C_{(s)} + H_2O_{(g)} \rightarrow CO_{(g)} + H_{2(g)}.$$



Speed of sound thin rod	diamond: 18,350 m/s (at 20 °C)
Thermal expansion	diamond: 0.8 µm/(m·K) (at 25 °C)[9]
Thermal conductivity	graphite: 119–165 W/(m·K) diamond: 900–2300 W/(m·K)
Electrical resistivity	graphite: 7.837 µΩ·m[10]
Magnetic ordering	diamagnetic[11]
Young's modulus	diamond: 1050 GPa[9]
Shear modulus	diamond: 478 GPa[9]
Bulk modulus	diamond: 442 GPa[9]
Poisson ratio	diamond: 0.1[9]
Mohs hardness	graphite: 1–2 diamond: 10
CAS Number	7440-44-0
	History
Discovery	Egyptians and Sumerians[12] (3750 BCE)
Recognized as an element by	Antoine Lavoisier[13] (1789)

Most stable isotopes of carbon

iso	NA	half-life	DM	DE (MeV)	DP
^{11}C	syn	20 min	β+	0.96	^{11}B
^{12}C	98.9%	is stable with 6 neutrons			
^{13}C	1.1%	is stable with 7 neutrons			
^{14}C	trace	5730 y	β−	0.156	^{14}N

Carbon combines with some metals at high temperatures to form metallic carbides, such as the iron carbide cementite in steel, and tungsten carbide, widely used as an abrasive and for making hard tips for cutting tools.

The system of carbon allotropes spans a range of extremes:

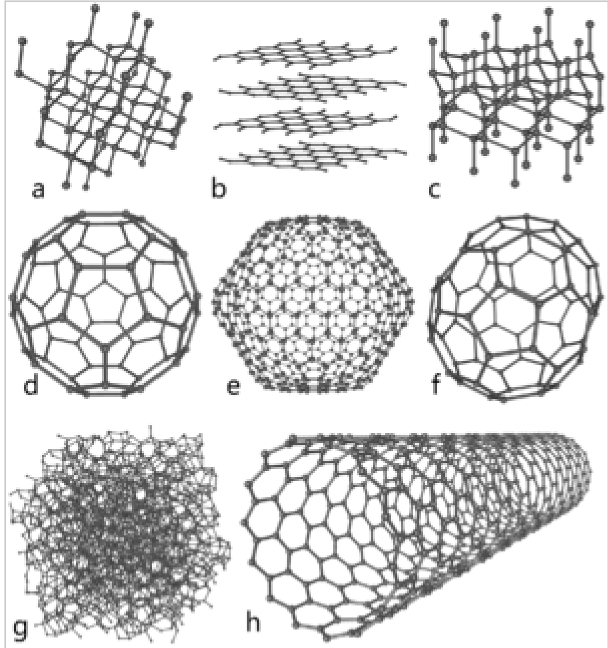
Allotropes

Atomic carbon is a very short-lived species and, therefore, carbon is stabilized in various multi-atomic structures with different molecular configurations called allotropes. The three relatively well-known allotropes of carbon are amorphous carbon, graphite, and diamond. Once considered exotic, fullerenes are nowadays commonly synthesized and used in research; they include buckyballs,[29][30] carbon nanotubes,[31] carbon nanobuds[32] and nanofibers.[33][34] Several other exotic allotropes have also been discovered, such as lonsdaleite (questionable),[35] glassy carbon,[36] carbon nanofoam[37] and linear acetylenic carbon (carbyne).[38]

As of 2009, graphene appears to be the strongest material ever tested.[39] The process of separating it from graphite will require some further technological development before it is economical for industrial processes.[40] If successful, graphene could be used in the construction of an Earth to Space Elevator. It could also be used to safely store hydrogen for use in a hydrogen based engine in cars.[41]

The amorphous form is an assortment of carbon atoms in a non-crystalline, irregular, glassy state, not held in a crystalline macrostructure. It is present as a powder, and is the main constituent of substances such as charcoal, lampblack (soot) and activated carbon. At normal pressures, carbon takes the form of graphite, in which each atom is bonded trigonally to three others in a plane composed of fused hexagonal rings, just like those in aromatic hydrocarbons.[42] The resulting network is 2-dimensional, and the resulting flat sheets are stacked and loosely bonded through weak van der Waals forces. This gives graphite its softness and its cleaving properties (the sheets slip easily past one another). Because of the delocalization of one of the outer electrons of each atom to form a π-cloud, graphite conducts electricity, but only in the plane of each covalently bonded sheet. This results in a lower bulk electrical conductivity for carbon than for most metals. The delocalization also accounts for the energetic stability of graphite over diamond at room temperature.

At very high pressures, carbon forms the more compact allotrope, diamond, having nearly twice the density of graphite. Here, each atom is bonded tetrahedrally to four others, forming a 3-dimensional network of puckered six-membered rings of atoms. Diamond has the same cubic structure as silicon and germanium, and because of the strength of the carbon-carbon bonds, it is the hardest naturally occurring substance measured by resistance to scratching. Contrary to the popular belief that "*diamonds are forever*", they are thermodynamically unstable under normal conditions and transform into graphite.[17] Due to a high activation energy barrier, the transition into graphite is so slow at normal temperature that it is unnoticeable. Under some conditions, carbon crystallizes as lonsdaleite, a hexagonal crystal lattice with all atoms covalently bonded and properties similar to those of diamond.[35]



Some allotropes of carbon: a) diamond; b) graphite; c) lonsdaleite; d–f) fullerenes (C_{60}, C_{540}, C_{70}); g) amorphous carbon; h) carbon nanotube.

Fullerenes are a synthetic crystalline formation with a graphite-like structure, but in place of hexagons, fullerenes are formed of pentagons (or even heptagons) of carbon atoms. The missing (or additional) atoms warp the sheets into spheres, ellipses, or cylinders. The properties of fullerenes (split into buckyballs, buckytubes, and nanobuds) have not yet been fully analyzed and represent an intense area of research in nanomaterials. The names "*fullerene*" and "*buckyball*" are given after Richard Buckminster Fuller, popularizer of geodesic domes, which resemble the structure of fullerenes. The buckyballs are fairly large molecules formed completely of carbon bonded trigonally, forming spheroids (the best-known and simplest is the soccerball-shaped C_{60} buckminsterfullerene).[29] Carbon nanotubes are structurally similar to buckyballs, except that each atom is bonded trigonally in a curved sheet that forms a hollow cylinder.[30][31] Nanobuds were first reported in 2007 and are hybrid bucky tube/buckyball materials (buckyballs are covalently bonded to the outer wall of a nanotube) that combine the properties of both in a single structure.[32]

Of the other discovered allotropes, carbon nanofoam is a ferromagnetic allotrope discovered in 1997. It consists of a low-density cluster-assembly of carbon atoms strung together in a loose three-dimensional web, in which the atoms are bonded trigonally in six- and seven-membered rings. It is among the lightest known solids, with a density of about 2 kg/m^3.[43] Similarly, glassy carbon contains a high proportion of closed porosity,[36] but contrary to normal graphite, the graphitic layers are not stacked like pages in a book, but have a more random arrangement. Linear acetylenic carbon[38] has the chemical structure[38] -(C:::C)$_n$-. Carbon in this modification is linear with *sp* orbital hybridization, and is a polymer with alternating single and triple bonds. This carbyne is of considerable interest to nanotechnology as its Young's modulus is forty times that of the hardest known material – diamond.[44]

In 2015, a team at the North Carolina State University announced the development of another allotrope they have dubbed Q-carbon, created by a high energy low duration laser pulse on amorphous carbon dust. Q-carbon is reported to exhibit ferromagetism, fluorescence, and a hardness superior to diamonds.[45]

Occurrence

Carbon is the fourth most abundant chemical element in the universe by mass after hydrogen, helium, and oxygen. Carbon is abundant in the Sun, stars, comets, and in the atmospheres of most planets.[46] Some meteorites contain microscopic diamonds that were formed when the solar system was still a protoplanetary disk. Microscopic diamonds may also be formed by the intense pressure and high temperature at the sites of meteorite impacts.[47]

Graphite ore. Penny is included for scale.

In 2014 NASA announced a greatly upgraded database (http://www.astrochem.org/pahdb/) for tracking polycyclic aromatic hydrocarbons (PAHs) in the universe. More than 20% of the carbon in the universe may be associated with PAHs, complex compounds of carbon and hydrogen without oxygen.[48] These compounds figure in the PAH world hypothesis where they are hypothesized to have a role in abiogenesis and formation of life. PAHs seem to have been formed "a couple of billion years" after the Big Bang, are widespread throughout the universe, and are associated with new stars and exoplanets.[46]

Raw diamond crystal.

It has been estimated that the solid earth as a whole contains 730 ppm of carbon, with 2000 ppm in the core and 120 ppm in the combined mantle and crust.[49] Since the mass of the earth is 5.972×10^{24} kg, this would imply 4360 million gigatonnes of carbon. This is much more than the amount of carbon in the oceans or atmosphere (below).

In combination with oxygen in carbon dioxide, carbon is found in the Earth's atmosphere (approximately 810 gigatonnes of carbon) and dissolved in all water bodies (approximately 36,000 gigatonnes of carbon). Around 1,900 gigatonnes of carbon are present in the biosphere. Hydrocarbons (such as coal, petroleum, and natural gas) contain carbon as well. Coal "reserves" (not "resources") amount to around 900 gigatonnes with perhaps 18 000 Gt of resources.[50] Oil reserves are around 150 gigatonnes. Proven sources of natural gas are about 175 10^{12} cubic metres (containing about 105 gigatonnes of carbon), but studies estimate another 900 10^{12} cubic metres of "unconventional" deposits such as shale gas, representing about 540 gigatonnes of carbon.[51]

Carbon is also found in methane hydrates in polar regions and under the seas. Various estimates put this carbon between 500, 2500 Gt,[52] or 3000 Gt.[53]

In the past, quantities of hydrocarbons were greater. According to one source, in the period from 1751 to 2008 about 347 gigatonnes of carbon were released as carbon dioxide to the atmosphere from burning of fossil fuels.[54] Another source puts the amount added to the atmosphere for the period since 1750 at 879 Gt, and the total going to the atmosphere, sea, and land (such as peat bogs) at almost 2000 Gt.[55]

Carbon is a constituent (about 12% by mass) of the very large masses of carbonate rock (limestone, dolomite, marble and so on). Coal is very rich in carbon (anthracite contains 92–98%)[56] and is the largest commercial source of mineral carbon, accounting for 4,000 gigatonnes or 80% of fossil fuel.[57]

As for individual carbon allotropes, graphite is found in large quantities in the United States (mostly in New York and Texas), Russia, Mexico, Greenland, and India. Natural diamonds occur in the rock kimberlite, found in ancient volcanic "necks", or "pipes". Most diamond deposits are in Africa, notably in South Africa, Namibia, Botswana, the Republic of the Congo, and Sierra Leone. Diamond deposits have also been found in Arkansas, Canada, the Russian Arctic, Brazil, and in Northern and Western Australia. Diamonds are now also being recovered from the ocean floor off the Cape of Good Hope. Diamonds are found naturally, but about 30% of all industrial diamonds used in the U.S. are now manufactured.

Carbon-14 is formed in upper layers of the troposphere and the stratosphere at altitudes of 9–15 km by a reaction that is precipitated by cosmic rays.[58] Thermal neutrons are produced that collide with the nuclei of nitrogen-14, forming carbon-14 and a proton. As such, 1.2×10^{10}% of atmospheric carbon dioxide contains carbon-14.[23]

Carbon-rich asteroids are relatively preponderant in the outer parts of the asteroid belt in our solar system. These asteroids have not yet been directly sampled by scientists. The asteroids can be used in hypothetical space-based carbon mining, which may be possible in the future, but is currently technologically impossible.[59]

Isotopes

Isotopes of carbon are atomic nuclei that contain six protons plus a number of neutrons (varying from 2 to 16). Carbon has two stable, naturally occurring isotopes.[14] The isotope carbon-12 (^{12}C) forms 98.93% of the carbon on Earth, while carbon-13 (^{13}C) forms the remaining 1.07%.[14] The concentration of ^{12}C is further increased in biological materials because biochemical reactions discriminate against ^{13}C.[60] In 1961, the International Union of Pure and Applied Chemistry (IUPAC) adopted the isotope carbon-12 as the basis for atomic weights.[61] Identification of carbon in nuclear magnetic resonance (NMR) experiments is done with the isotope ^{13}C.

Carbon-14 (^{14}C) is a naturally occurring radioisotope, created in the upper atmosphere (lower stratosphere and upper troposphere) by interaction of nitrogen with cosmic rays.[62] It is found in trace amounts on Earth of up to 1 part per trillion (0.0000000001%), mostly confined to the atmosphere and superficial deposits, particularly of peat and other organic materials.[63] This isotope decays by 0.158 MeV β^- emission. Because of its relatively short half-life of 5730 years, ^{14}C is virtually absent in ancient rocks. The amount of ^{14}C in the atmosphere and in living organisms is almost constant, but decreases predictably in their bodies after death. This principle is used in radiocarbon dating, invented in 1949, which has been used extensively to determine the age of carbonaceous materials with ages up to about 40,000 years.[64][65]

There are 15 known isotopes of carbon and the shortest-lived of these is ^{8}C which decays through proton emission and alpha decay and has a half-life of 1.98739×10^{-21} s.[66] The exotic ^{19}C exhibits a nuclear halo, which means its radius is appreciably larger than would be expected if the nucleus were a sphere of constant density.[67]

Formation in stars

Formation of the carbon atomic nucleus requires a nearly simultaneous triple collision of alpha particles (helium nuclei) within the core of a giant or supergiant star which is known as the triple-alpha process, as the products of further nuclear fusion reactions of helium with hydrogen or another helium nucleus produce lithium-5 and beryllium-8 respectively, both of which are highly unstable and decay almost instantly back into smaller nuclei.[68] This happens in conditions of temperatures over 100 megakelvin and helium concentration that the rapid expansion and cooling of the early universe prohibited, and therefore no significant carbon was created during the Big Bang.

According to current physical cosmology theory, carbon is formed in the interiors of stars in the horizontal branch by the collision and transformation of three helium nuclei.[69] When those stars die as supernova, the carbon is scattered into space as dust. This dust becomes component material for the formation of second or third-generation star systems with accreted planets.[46][70] The Solar System is one such star system with an abundance of carbon, enabling the existence of life as we know it.

The CNO cycle is an additional fusion mechanisms that powers stars, wherein carbon operates as a catalyst.

Rotational transitions of various isotopic forms of carbon monoxide (for example, ^{12}CO, ^{13}CO, and ^{18}CO) are detectable in the submillimeter wavelength range, and are used in the study of newly forming stars in molecular clouds.[71]

Carbon cycle

Under terrestrial conditions, conversion of one element to another is very rare. Therefore, the amount of carbon on Earth is effectively constant. Thus, processes that use carbon must obtain it from somewhere and dispose of it somewhere else. The paths of carbon in the environment form the carbon cycle. For example, photosynthetic plants draw carbon dioxide from the atmosphere (or seawater) and build it into biomass, as in the Calvin cycle, a process of carbon fixation. Some of this biomass is eaten by animals, while some carbon is exhaled by animals as carbon dioxide. The carbon cycle is considerably more complicated than this short loop; for example, some carbon dioxide is dissolved in the oceans; if bacteria do not consume it, dead plant or animal matter may become petroleum or coal, which releases carbon when burned.[72][73]

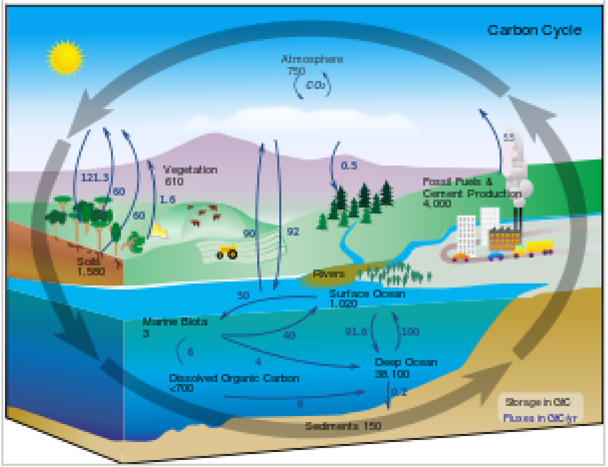


Diagram of the carbon cycle. The black numbers indicate how much carbon is stored in various reservoirs, in billions tonnes ("GtC" stands for gigatonnes of carbon; figures are circa 2004). The purple numbers indicate how much carbon moves between reservoirs each year. The sediments, as defined in this diagram, do not include the ≈70 million GtC of carbonate rock and kerogen.

Compounds

Organic compounds

Carbon can form very long chains of interconnecting C-C bonds, a property that is called catenation. Carbon-carbon bonds are strong and stable. Through catenation, carbon forms a countless number of compounds. A tally of unique compounds shows that more contain carbon that those that do not. A similar claim can be made for hydrogen because most organic compounds also contain hydrogen.

The simplest form of an organic molecule is the hydrocarbon—a large family of organic molecules that are composed of hydrogen atoms bonded to a chain of carbon atoms. Chain length, side chains and functional groups all affect the properties of organic molecules.

Carbon occurs in all known organic life and is the basis of organic chemistry. When united with hydrogen, it forms various hydrocarbons that are important to industry as refrigerants, lubricants, solvents, as chemical feedstock for the manufacture of plastics and petrochemicals, and as fossil fuels.

When combined with oxygen and hydrogen, carbon can form many groups of important biological compounds including sugars, lignans, chitins, alcohols, fats, and aromatic esters, carotenoids and terpenes. With nitrogen it forms alkaloids, and with the addition of sulfur also it forms antibiotics, amino acids, and rubber products. With the addition of phosphorus to these other elements, it forms DNA and RNA, the chemical-code carriers of life, and adenosine triphosphate (ATP), the most important energy-transfer molecule in all living cells.

Inorganic compounds

Commonly carbon-containing compounds which are associated with minerals or which do not contain hydrogen or fluorine, are treated separately from classical organic compounds; the definition is not rigid (see reference articles above). Among these are the simple oxides of carbon. The most prominent oxide is carbon dioxide (CO_2). This was once the principal constituent of the paleoatmosphere, but is a minor component of the Earth's atmosphere today.[74] Dissolved in water, it forms carbonic acid

(H_2CO_3), but as most compounds with multiple single-bonded oxygens on a single carbon it is unstable.[75] Through this intermediate, though, resonance-stabilized carbonate ions are produced. Some important minerals are carbonates, notably calcite. Carbon disulfide (CS_2) is similar.[22]

The other common oxide is carbon monoxide (CO). It is formed by incomplete combustion, and is a colorless, odorless gas. The molecules each contain a triple bond and are fairly polar, resulting in a tendency to bind permanently to hemoglobin molecules, displacing oxygen, which has a lower binding affinity.[76][77] Cyanide (CN^-), has a similar structure, but behaves much like a halide ion (pseudohalogen). For example, it can form the nitride cyanogen molecule ($(CN)_2$), similar to diatomic halides. Other uncommon oxides are carbon suboxide (C_3O_2),[78] the unstable dicarbon monoxide (C_2O),[79][80] carbon trioxide (CO_3),[81][82] cyclopentanepentone (C_5O_5),[83] cyclohexanehexone (C_6O_6),[83] and mellitic anhydride ($C_{12}O_9$).

With reactive metals, such as tungsten, carbon forms either carbides (C^{4-}) or acetylides (C_2^{2-}) to form alloys with high melting points. These anions are also associated with methane and acetylene, both very weak acids. With an electronegativity of 2.5,[84] carbon prefers to form covalent bonds. A few carbides are covalent lattices, like carborundum (SiC), which resembles diamond. Nevertheless, even the most polar and salt-like of carbides are not completely ionic compounds.[85]

Organometallic compounds

Organometallic compounds by definition contain at least one carbon-metal bond. A wide range of such compounds exist; major classes include simple alkyl-metal compounds (for example, tetraethyllead), η^2-alkene compounds (for example, Zeise's salt), and η^3-allyl compounds (for example, allylpalladium chloride dimer); metallocenes containing cyclopentadienyl ligands (for example, ferrocene); and transition metal carbene complexes. Many metal carbonyls exist (for example, tetracarbonylnickel); some workers consider the carbon monoxide ligand to be purely inorganic, and not organometallic.

While carbon is understood to exclusively form four bonds, an interesting compound containing an octahedral hexacoordinated carbon atom has been reported. The cation of the compound is $[(Ph_3PAu)_6C]^{2+}$. This phenomenon has been attributed to the aurophilicity of the gold ligands.[86]

Source

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