Fluorine

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Fluorine is a chemical element with symbol **F** and atomic number 9. It is the lightest halogen and exists as a highly toxic pale yellow diatomic gas at standard conditions. As the most electronegative element, it is extremely reactive: almost all other elements, including some noble gases, form compounds with fluorine.

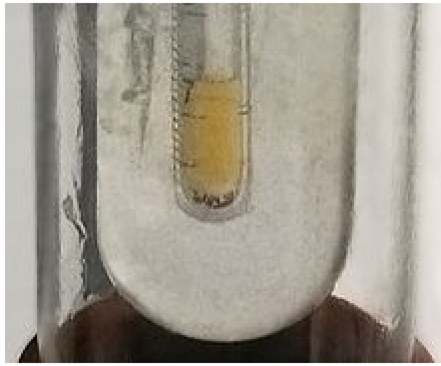
Among the elements, fluorine ranks 24th in universal abundance and 13th in terrestrial abundance. Fluorite, the primary mineral source of fluorine, was first described in 1529; as it was added to metal ores to lower their melting points for smelting, the Latin verb *fluo* meaning "flow" became associated with it. Proposed as an element in 1810, fluorine proved difficult and dangerous to separate from its compounds, and several early experimenters died or sustained injuries from their attempts. Only in 1886 did French chemist Henri Moissan isolate elemental fluorine using low-temperature electrolysis, a process still employed for modern production. Industrial production of fluorine gas for uranium enrichment, its largest application, began during the Manhattan Project in World War II.

Owing to the expense of refining pure fluorine, most commercial applications use fluorine compounds, with about half of mined fluorite used in steelmaking. The rest of the fluorite is converted into corrosive hydrogen fluoride en route to various organic fluorides, or into cryolite which plays a key role in aluminium refining. Organic fluorides have very high chemical and thermal stability; their major uses are as refrigerants, electrical insulation and cookware, the last as PTFE (Teflon). Pharmaceuticals such as atorvastatin and fluoxetine also contain fluorine, and the fluoride ion inhibits dental cavities, and so finds use in toothpaste and water fluoridation. Global fluorochemical sales amount to more than US$15 billion a year.

Fluorocarbon gases are generally greenhouse gases with global-warming potentials 100 to 20,000 times that of carbon dioxide. Organofluorine compounds persist in the environment due to the strength of the carbon–fluorine bond. Fluorine has no known metabolic role in mammals; a few plants synthesize organofluorine poisons that deter herbivores.

Characteristics

Fluorine, $_9$F



Liquid fluorine at cryogenic temperatures

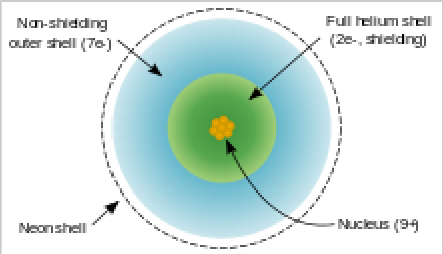
General properties

Name, symbol	fluorine, F
Pronunciation	/ˈflʊəriːn/, /ˈflʊərɪn/, /ˈflɔəriːn/ FLOOR-*een*, FLOOR-*in*, FLOHR-*een*
Allotropes	alpha, beta
Appearance	gas: very pale yellow liquid: bright yellow solid: alpha is opaque, beta is transparent

Fluorine in the periodic table

Atomic number (*Z*)	9
Group, block	group 17 (halogens), p-block
Period	

Electron configuration



Simplified structure of the fluorine atom

Fluorine atoms have nine electrons, one fewer than neon, and electron configuration $1s^2 2s^2 2p^5$: two electrons in a filled inner shell and seven in an outer shell requiring one more to be filled. The outer electrons are ineffective at nuclear shielding, and experience a high effective nuclear charge of 9 − 2 = **7**; this affects the atom's physical properties.[2]

Fluorine's first ionization energy is third-highest among all elements, behind helium and neon,[13] which complicates the removal of electrons from neutral fluorine atoms. It also has a high electron affinity, second only to chlorine,[14] and tends to capture an electron to become isoelectronic with the noble gas neon;[2] it has the highest electronegativity of any element.[15] Fluorine atoms have a small covalent radius of around 60 picometers, similar to those of its period neighbors oxygen and neon.[16][17][note 1]

Reactivity

The bond energy of difluorine is much lower than that of either Cl_2 or Br_2 and similar to the easily cleaved peroxide bond; this, along with high electronegativity, accounts for fluorine's easy dissociation, high reactivity, and strong bonds to non-fluorine atoms.[18][19] Conversely, bonds to other atoms are very strong because of fluorine's high electronegativity. Unreactive substances like powdered steel, glass fragments, and asbestos fibers react quickly with cold fluorine gas; wood and water spontaneously combust under a fluorine jet.[4][20]

Reactions of elemental fluorine with metals require varying conditions. Alkali metals cause explosions and alkaline earth metals display vigorous activity in bulk; to prevent passivation from the formation of metal fluoride layers, most other metals such as aluminium and iron must be powdered,[18] and noble metals require pure fluorine gas at 300–450 °C (575–850 °F).[21] Some solid nonmetals (sulfur, phosphorus) react

	period 2
Element category	□ diatomic nonmetal
Standard atomic weight (±) (A_r)	18.998403163(6)[1]
Electron configuration	[He] $2s^2$ $2p^5$[2]
per shell	2, 7
Physical properties	
Phase	gas
Melting point	53.48 K (−219.67 °C, −363.41 °F)[3]
Boiling point	85.03 K (−188.11 °C, −306.60 °F)[3]
Density at stp (0 °C and 101.325 kPa)	1.696 g/L[4]
when liquid, at b.p.	1.505 g/cm^3[5]
Triple point	53.48 K, 90 kPa[3]
Critical point	144.41 K, 5.1724 MPa[3]
Heat of vaporization	6.51 kJ/mol[4]
Molar heat capacity	C_p: 31 J/(mol·K)[5] (at 21.1 °C) C_v: 23 J/(mol·K)[5] (at 21.1 °C)

Vapor pressure

P (Pa)	**1**	**10**	**100**	**1 k**	**10 k**	**100 k**
at T (K)	38	44	50	58	69	85

Atomic properties

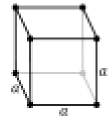
vigorously in liquid air temperature fluorine.[22] Hydrogen sulfide[22] and sulfur dioxide[23] combine readily with fluorine, the latter sometimes explosively; sulfuric acid exhibits much less activity, requiring elevated temperatures.[24]

Hydrogen, like some of the alkali metals, reacts explosively with fluorine.[25] Carbon, as lamp black, reacts at room temperature to yield fluoromethane. Graphite combines with fluorine above 400 °C (750 °F) to produce non-stoichiometric carbon monofluoride; higher temperatures generate gaseous fluorocarbons, sometimes with explosions.[26] Carbon dioxide and carbon monoxide react at or just above room temperature,[27] whereas paraffins and other organic chemicals generate strong reactions:[28] even fully substituted haloalkanes such as carbon tetrachloride, normally incombustible, may explode.[29] Although nitrogen trifluoride is stable, nitrogen requires an electric discharge at elevated temperatures for reaction with fluorine to occur, due to the very strong triple bond in elemental nitrogen;[30] ammonia may react explosively.[31][32] Oxygen does not combine with fluorine under ambient conditions, but can be made to react using electric discharge at low temperatures and pressures; the products tend to disintegrate into their constituent elements when heated.[33][34][35] Heavier halogens[36] react readily with fluorine as does the noble gas radon;[37] of the other noble gases, only xenon and krypton react, and only under special conditions.[38]

Phases

At room temperature, fluorine is a gas of diatomic molecules,[4] pale yellow when pure (sometimes described as yellow-green).[39] It has a characteristic pungent odor detectable at 20 ppb.[40] Fluorine condenses into a bright yellow liquid at −188 °C (−306 °F), a transition temperature similar to those of oxygen and nitrogen.[41]

Fluorine has two solid forms, α- and β-fluorine. The latter crystallizes at −220 °C (−364 °F) and is transparent and soft, with the same disordered cubic structure of freshly crystallized solid oxygen,[41][note 2] unlike the orthorhombic systems of other solid halogens.[45][46] Further cooling to −228 °C (−378 °F) induces a phase transition

Oxidation states	**−1** (oxidizes oxygen)
Electronegativity	Pauling scale: 3.98[2]
Ionization energies	1st: 1681 kJ/mol 2nd: 3374 kJ/mol 3rd: 6147 kJ/mol (more)[6]
Covalent radius	64 pm[7]
Van der Waals radius	135 pm[8]
Miscellanea	
Crystal structure	cubic 
Thermal conductivity	0.02591 W/(m·K)[9]
Magnetic ordering	diamagnetic (-1.2×10^{-4})[10][11]
CAS Number	7782-41-4[2]
History	
Naming	after the mineral fluorite, itself named after Latin *fluo* (to flow, in smelting)
Discovery	André-Marie Ampère (1810)
First isolation	Henri Moissan[2] (June 26, 1886)
Named by	Humphry Davy
Most stable isotopes of fluorine[12]	

into opaque and hard α-fluorine, which has a monoclinic structure with dense, angled layers of molecules. The transition from β- to α-fluorine is more exothermic than the condensation of fluorine, and can be violent.[45][46][note 3]

iso	NA	half-life	DM	DE (MeV)	DP
^{18}F	trace	109.77 min	β^+ (96.9%)	0.634	^{18}O
			ε (3.1%)	1.656	^{18}O
^{19}F	100%	is stable with 10 neutrons			

Isotopes

Only one isotope of fluorine occurs naturally in abundance, the stable isotope ^{19}F.[47] It has a high magnetogyric ratio[note 4] and exceptional sensitivity to magnetic fields; because it is also the only stable isotope, it is used in magnetic resonance imaging.[49] Seventeen radioisotopes with mass numbers from 14 to 31 have been synthesized, of which ^{18}F is the most stable with a half-life of 109.77 minutes. Other radioisotopes have half-lives less than 70 seconds; most decay in less than half a second.[50] The isotopes ^{17}F and ^{18}F undergo β^+ decay, lighter isotopes decay by electron capture, and those heavier than ^{19}F undergo β^- decay or neutron emission.[50] One metastable isomer of fluorine is known, ^{18m}F, with a half-life of 234 nanoseconds.[51]

Source

- Wikipedia: Fluorine