

Iodine

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Iodine is a chemical element with symbol **I** and atomic number 53. The heaviest of the stable halogens, it exists as a lustrous, purple-black metallic solid at a standard conditions that sublimes readily to form a violet gas. The elemental form was discovered by the French chemist Bernard Courtois in 1811. It was named two years later by Joseph-Louis Gay-Lussac from this property, after the Greek *ἰωδης* "violet-coloured".

Iodine occurs in many oxidation states, including iodide (I^-), iodate (IO_3^-), and the various periodate anions. It is the least abundant of the stable halogens, being the sixty-first most abundant element. It is even less abundant than the so-called rare earths. It is the heaviest essential element. Iodine is found in the thyroid hormones. Iodine deficiency affects about two billion people and is the leading preventable cause of intellectual disabilities.

The dominant producers of iodine today are Chile and Japan. Iodine and its compounds are primarily used in nutrition. Due to its high atomic number and ease of attachment to organic compounds, it has also found favour as a non-toxic radiocontrast material. Because of the specificity of its uptake by the human body, radioactive isotopes of iodine can also be used to treat thyroid cancer. Iodine is also used as a catalyst in the industrial production of acetic acid and some polymers.

Properties

Iodine is the fourth halogen, being a member of group 17 in the periodic table, below fluorine, chlorine, and bromine; it is the heaviest stable member of its group. (The scarce and fugitive fifth halogen, the radioactive astatine, is not well-studied due to its expense and inaccessibility in large quantities, but appears to show various unusual properties due to relativistic effects.) Iodine has an electron configuration of $[\text{Kr}]4\text{d}^{10}5\text{s}^25\text{p}^5$, with the seven electrons in the fifth and outermost shell being its valence electrons. Like the other halogens, it is one electron short of a full octet and is hence a strong oxidising agent, reacting with many elements in order to complete its outer shell, although in keeping with periodic trends, it is the weakest oxidising agent among the stable halogens: it has the lowest electronegativity among them, just 2.66

Iodine, $_{53}\text{I}$



General properties

Name, symbol	iodine, I
Appearance	lustrous metallic gray, violet as a gas

Iodine in the periodic table

Atomic number (<i>Z</i>)	53
Group, block	group 17 (halogens), p-block
Period	period 5
Element category	☐ diatomic nonmetal
Standard atomic weight (\pm) (<i>A</i> _r)	126.90447(3) ^[1]
Electron configuration	[Kr] 4d ¹⁰ 5s ² 5p ⁵
 per shell	2, 8, 18, 18, 7

Physical properties

Phase	solid
Melting point	386.85 K (113.7 °C, 236.66 °F)
Boiling point	457.4 K (184.3 °C, 363.7 °F)



Violet iodine vapour in a round-bottomed flask

on the Pauling scale (compare fluorine, chlorine, and bromine at 3.98, 3.16, and 2.96 respectively; astatine continues the trend with an electronegativity of 2.2). Elemental iodine hence forms diatomic molecules with chemical formula I_2 , where two iodine atoms share a pair of electrons in order to each achieve a stable octet for themselves; at high temperatures, these diatomic molecules reversibly dissociate a pair of iodine atoms. Similarly, the iodide anion, I^- , is the strongest reducing agent among the stable halogens, being the most easily oxidised back to diatomic I_2 .^[14] (Astatine goes further, being indeed unstable as At^- and readily oxidised to At^0 or At^+ , although the existence of At_2 is not settled.)^[15]

The halogens darken in colour as the group is descended: fluorine is a very pale yellow gas, chlorine is greenish-yellow, and bromine is a reddish-brown volatile liquid. Iodine conforms to the prevailing trend, being a shiny black crystalline solid that melts at 114 °C and boils at 183 °C to form a violet gas. This trend occurs because the wavelengths of visible light absorbed by the halogens increase down the group (though astatine may not conform to it, depending on how metallic it turns out to be).^[14] Specifically, the violet colour of iodine gas results from the electron transition between the highest occupied antibonding π_g molecular orbital and the lowest vacant antibonding σ_u molecular orbital.^[16]

Elemental iodine is slightly soluble in water, with one gram dissolving in 3450 ml at 20 °C and 1280 ml at 50 °C; potassium iodide may be added to increase solubility via formation of triiodide ions, among other polyiodides.^[16] Nonpolar solvents such as hexane and carbon tetrachloride provide a higher solubility.^[17] Polar solutions, such as aqueous solutions, are brown, reflecting the role of these solvents as Lewis bases; on the other hand, nonpolar solutions are violet, the color of iodine vapour.^[16] Charge-transfer complexes form when iodine is dissolved in polar solvents, hence changing the colour. Iodine is violet when dissolved in carbon tetrachloride and saturated hydrocarbons but deep brown in alcohols and amines, solvents that form charge-transfer adducts.^[18]

Density near r.t.	4.933 g/cm ³
Triple point	386.65 K, 12.1 kPa
Critical point	819 K, 11.7 MPa
Heat of fusion	(I_2) 15.52 kJ/mol
Heat of vaporisation	(I_2) 41.57 kJ/mol
Molar heat capacity	(I_2) 54.44 J/(mol·K)

Vapour pressure (rhombic)

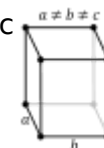
P (Pa)	1	10	100	1 k	10 k	100 k
at T (K)	260	282	309	342	381	457

Atomic properties

Oxidation states	7, 6, 5, 4, 3, 1, −1 (a strongly acidic oxide)
Electronegativity	Pauling scale: 2.66
Ionisation energies	1st: 1008.4 kJ/mol 2nd: 1845.9 kJ/mol 3rd: 3180 kJ/mol
Atomic radius	empirical: 140 pm
Covalent radius	139±3 pm
Van der Waals radius	198 pm

Miscellanea

Crystal structure	orthorhombic
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Thermal conductivity	0.449 W/(m·K)
Electrical resistivity	$1.3 \times 10^7 \Omega \cdot m$ (at 0 °C)

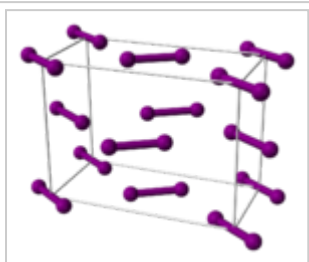
The melting and boiling points of iodine are the highest among the halogens, conforming to the increasing trend down the group, since iodine has the largest electron cloud among them that is the most easily polarised, resulting in its molecules having the strongest van der Waals interactions among the halogens. Similarly, iodine is the least volatile of the halogens.^[14] Because it has the largest atomic radius among the halogens, iodine has the lowest first ionisation energy, lowest electron affinity, lowest electronegativity and lowest reactivity of the halogens.^[14]

The interhalogen bond in diiodine is the weakest of all the halogens. As such, 1% of a sample of gaseous iodine at atmospheric pressure is dissociated into iodine atoms at 575 °C. Temperatures greater than 750 °C are required for fluorine, chlorine, and bromine to dissociate to a similar extent. Most bonds to iodine are weaker than the analogous bonds to the lighter halogens.^[14] Gaseous iodine is composed of I₂ molecules with an I-I bond length of 266.6 pm. The I-I bond is one of the longest single bonds known. It is even longer (271.5 pm) in solid orthorhombic crystalline iodine, which has the same crystal structure as chlorine and bromine. (The record is held by iodine's neighbour xenon: the Xe-Xe bond length is 308.71 pm.)^[19] As such, within the iodine molecule, significant electronic interactions occur with the two next-nearest neighbours of each atom, and these interactions give rise, in bulk iodine, to a shiny appearance and semiconducting properties.^[14] Iodine is a two-dimensional semiconductor with a band gap of 1.3 eV (125 kJ/mol): it is a semiconductor in the plane of its crystalline layers and an insulator in the perpendicular direction.^[14]

Isotopes

Of the thirty-seven known isotopes of iodine, only one occurs in nature, iodine-127. The others are radioactive and have half-lives too short to be primordial. As such, iodine is monoisotopic and its atomic weight is known to great precision, as it is a constant of nature.^[14]

The longest-lived of the radioactive isotopes of iodine is iodine-129, which has a half-life of 15.7 million years, decaying via beta decay to stable xenon-129.^[20] Some iodine-129 was formed along with iodine-127 before the formation of the Solar System, but it has by now completely decayed away, making it an extinct radionuclide that is nevertheless still useful in dating the history of the early Solar System or very old groundwaters, due to its mobility in the environment. Its former presence may be determined from an excess of its daughter xenon-129.^{[21][22][23][24][25]} Traces of iodine-129 still exist today, as it is also a cosmogenic nuclide, formed from cosmic ray



Structure of solid iodine

spallation of atmospheric xenon: these traces make up 10^{-14} to 10^{-10} of all natural iodine. It also occurs from open-air nuclear testing, and is not hazardous because of its incredibly long half-life, the longest of all fission products. At the peak of thermonuclear testing in the 1960s and 1970s, iodine-129 still only made up about 10^{-7} of all natural iodine.^[26] Excited states of iodine-127 and iodine-129 are often used in Mössbauer spectroscopy.^[14]

The other iodine radioisotopes have much shorter half-lives, no longer than days.^[20] Some of them have medical applications involving the thyroid gland, where the iodine that enters the body is stored and concentrated. Iodine-123 has a half-life of thirteen hours and decays by electron capture to tellurium-123, emitting gamma radiation; it is used in nuclear medicine imaging, including single photon emission computed tomography (SPECT) and X-ray computed tomography (X-Ray CT) scans.^[27] Iodine-125 has a half-life of fifty-nine days, decaying by electron capture to tellurium-125 and emitting low-energy gamma radiation; the second-longest-lived iodine radioisotope, it has uses in biological assays, nuclear medicine imaging and in radiation therapy as brachytherapy to treat a number of conditions, including prostate cancer, uveal melanomas, and brain tumours.^[28] Finally, iodine-131, with a half-life of eight days, beta decays to an excited state of stable xenon-131 that then converts to the ground state by emitting gamma radiation. It is a common fission product and thus is present in high levels in radioactive fallout. It may then be absorbed through contaminated food, and will also accumulate in the thyroid. As it decays, it may cause damage to the thyroid. The primary risk from exposure to high levels of iodine-131 is the chance occurrence of radiogenic thyroid cancer in later life. Other risks include the possibility of non-cancerous growths and thyroiditis.^[29] The usual means of protection against the negative effects of iodine-131 is by saturating the thyroid gland with stable iodine-127 in the form of potassium iodide tablets, taken daily for optimal prophylaxis.^[30] However, iodine-131 may also be used for medicinal purposes in radiation therapy for this very reason, when tissue destruction is desired after iodine uptake by the tissue.^[31] Iodine-131 is also used as a radioactive tracer.^{[32][33][34][35]}

Source

- Wikipedia: Iodine