

Caesium

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Caesium or **cesium**^[note 1] is a chemical element with symbol **Cs** and atomic number 55. It is a soft, silvery-gold alkali metal with a melting point of 28.5 °C (83.3 °F), which makes it one of only five elemental metals that are liquid at or near room temperature.^[note 2] Caesium is an alkali metal and has physical and chemical properties similar to those of rubidium and potassium. The metal is extremely reactive and pyrophoric, reacting with water even at −116 °C (−177 °F). It is the least electronegative element. It has only one stable isotope, caesium-133. Caesium is mined mostly from pollucite, while the radioisotopes, especially caesium-137, a fission product, are extracted from waste produced by nuclear reactors.

The German chemist Robert Bunsen and physicist Gustav Kirchhoff discovered caesium in 1860 by the newly developed method of flame spectroscopy. The first small-scale applications for caesium were as a "getter" in vacuum tubes and in photoelectric cells. In 1967, acting on Einstein's proof that the speed of light is the most constant dimension in the universe, the International System of Units used two specific wave counts from an emission spectrum of caesium-133 to co-define the second and the metre. Since then, caesium has been widely used in highly accurate atomic clocks.

Since the 1990s, the largest application of the element has been as caesium formate for drilling fluids, but it has a range of applications in the production of electricity, in electronics, and in chemistry. The radioactive isotope caesium-137 has a half-life of about 30 years and is used in medical applications, industrial gauges, and hydrology. Although the element is only mildly toxic, the metal is a hazardous material and the radioisotopes would present a significant health and ecological hazard in the environment.

Characteristics

Physical properties

Caesium, ⁵⁵Cs



Spectral lines of caesium

General properties

Name, symbol	caesium, Cs
Alternative name	cesium
Appearance	pale gold

Caesium in the periodic table

Atomic number (<i>Z</i>)	55
Group, block	group 1 (alkali metals), s-block
Period	period 6
Element category	☐ alkali metal
Standard atomic weight (<i>±</i>) (<i>A</i> _r)	132.90545196(6) ^[1]
Electron configuration	[Xe] 6s ¹
per shell	2, 8, 18, 18, 8, 1



High-purity caesium-133 stored in argon.

Caesium is the softest element (it has a hardness of 0.2 Mohs). It is very ductile, pale metal, which darkens in the presence of trace amounts of oxygen.^{[10][11][12]} It has a melting point of 28.4 °C (83.1 °F), making it one of the few elemental metals that are liquid near room temperature. Mercury is the only elemental metal with a known melting point lower than caesium.^{[note 3][14]} In addition, the metal has a rather low boiling point, 641 °C (1,186 °F), the lowest of all metals other than mercury.^[15] Its compounds burn with a blue^{[16][17]} or violet^[17] colour.

Caesium forms alloys with the other alkali metals, gold, and mercury (amalgams). At temperatures below 650 °C (1,202 °F), it does not alloy with cobalt, iron, molybdenum, nickel, platinum, tantalum, or tungsten. It forms well-defined intermetallic compounds with antimony, gallium, indium, and thorium, which are photosensitive.^[10] It mixes with all the other alkali metals (except lithium); the alloy with a molar distribution of 41% caesium, 47% potassium, and 12% sodium has the lowest melting point of any known metal alloy, at −78 °C (−108 °F).^{[14][18]} A few amalgams have been studied: CsHg₂ is black with a purple metallic lustre, while CsHg is golden-coloured, also with a metallic lustre.^[19]

Chemical properties

Caesium metal is highly reactive and very pyrophoric. It ignites spontaneously in air, and reacts explosively with water even at low temperatures, more so than the other alkali metals (first group of the periodic table).^[10] It reacts with solid water at temperatures as low as −116 °C (−177 °F).^[14] Because of this high reactivity, caesium metal is classified as a hazardous material. It is stored and shipped in dry, saturated hydrocarbons such as mineral oil. It can be handled only under inert gas, such as argon. However, a caesium-water explosion is often less powerful than a sodium-water explosion with a similar amount of sodium. This is because caesium explodes instantly upon contact with water, leaving little time for hydrogen to

Physical properties

Phase	solid
Melting point	301.7 K (28.5 °C, 83.3 °F)
Boiling point	944 K (671 °C, 1240 °F)
Density near r.t.	1.93 g/cm ³
when liquid, at m.p.	1.843 g/cm ³
Critical point	1938 K, 9.4 MPa ^[2]
Heat of fusion	2.09 kJ/mol
Heat of vaporization	63.9 kJ/mol
Molar heat capacity	32.210 J/(mol·K)

Vapour pressure

P (Pa)	1	10	100	1 k	10 k	100 k
at T (K)	418	469	534	623	750	940

Atomic properties

Oxidation states	+1 , −1 ^[3] (a strongly basic oxide)
Electronegativity	Pauling scale: 0.79
Ionization energies	1st: 375.7 kJ/mol 2nd: 2234.3 kJ/mol 3rd: 3400 kJ/mol
Atomic radius	empirical: 265 pm
Covalent radius	244±11 pm
Van der Waals radius	343 pm

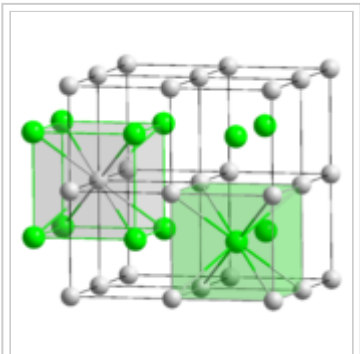
Miscellanea

Crystal structure	body-centred cubic (bcc)
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accumulate.^[20] Caesium can be stored in vacuum-sealed borosilicate glass ampoules. In quantities of more than about 100 grams (3.5 oz), caesium is shipped in hermetically sealed, stainless steel containers.^[10]

The chemistry of caesium is similar to that of other alkali metals, in particular rubidium, the element above caesium in the periodic table.^[21] As expected for an alkali metal, the only common oxidation state is +1.^[note 4] Some small differences arise from the fact that it has a higher atomic mass and is more electropositive than other (nonradioactive) alkali metals.^[23] Caesium is the most electropositive chemical element.^{[note 5][14]} The caesium ion is also larger and less "hard" than those of the lighter alkali metals.

Compounds



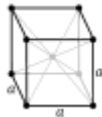
Ball-and-stick model of the cubic coordination of Cs and Cl in CsCl

Most caesium compounds contain the element as the cation Cs⁺, which binds ionically to a wide variety of anions. One noteworthy exception is the caeside anion (Cs[−]),^[3] and others are the several suboxides (see section on oxides below).

Salts of Cs⁺ are usually colourless unless the anion itself is coloured. Many of the simple salts are hygroscopic, but less so than the corresponding salts of lighter alkali metals. The phosphate,^[25] acetate, carbonate, halides, oxide, nitrate, and sulfate salts are water-soluble. Double salts are often less soluble, and the low solubility of caesium aluminium sulfate is exploited in refining Cs from ores. The double salt

with antimony (such as CsSbCl₄), bismuth, cadmium, copper, iron, and lead are also poorly soluble.^[10]

Caesium hydroxide (CsOH) is hygroscopic and strongly basic.^[21] It rapidly etches the surface of semiconductors such as silicon.^[26] CsOH has been previously regarded by chemists as the "strongest base", reflecting the relatively weak attraction between



Thermal expansion	97 μm/(m·K) (at 25 °C)
Thermal conductivity	35.9 W/(m·K)
Electrical resistivity	205 nΩ·m (at 20 °C)
Magnetic ordering	paramagnetic ^[4]
Young's modulus	1.7 GPa
Bulk modulus	1.6 GPa
Mohs hardness	0.2
Brinell hardness	0.14 MPa
CAS Number	7440-46-2

History

Naming	from Latin <i>caesius</i> , sky blue, for its spectral colours
Discovery	Robert Bunsen and Gustav Kirchhoff (1860)
First isolation	Carl Setterberg (1882)

Most stable isotopes of caesium

iso	NA	half-life	DM	DE (MeV)	DP
¹³³Cs	100%	is stable with 78 neutrons			
¹³⁴Cs	syn	2.0648 y	ε	1.229	¹³⁴ Xe
			β [−]	2.059	¹³⁴ Ba
¹³⁵Cs	trace	2.3×10 ⁶ y	β [−]	0.269	¹³⁵ Ba
¹³⁷Cs	syn	30.17 y ^[5]	β [−]	1.174	¹³⁷ Ba

the large Cs^+ ion and OH^- ; ^[16] it is indeed the strongest Arrhenius base, but a number of compounds that do not dissolve in water, such as *n*-butyllithium and sodium amide, ^[21] are more basic.

A stoichiometric mixture of caesium and gold will react to form yellow caesium auride (Cs^+Au^-) upon heating. The auride anion here behaves as a pseudohalogen. The compound reacts violently with water, yielding caesium hydroxide, metallic gold, and hydrogen gas; in liquid ammonia it can be reacted with a caesium-specific ion exchange resin to produce tetramethylammonium auride. The analogous platinum compound, red caesium platinide (Cs_2Pt), contains the platinide ion that behaves as a pseudochalcogen. ^[27]

Complexes

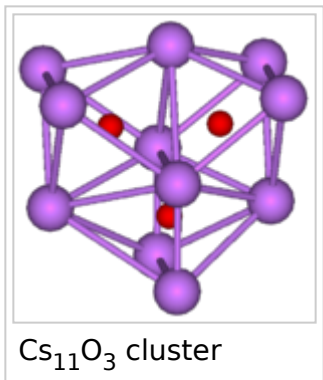
Like all metal cations, Cs^+ forms complexes with Lewis bases in solution. Because of its large size, Cs^+ usually adopts coordination numbers greater than 6, the number typical for the smaller alkali metal cations. This difference is apparent in the 8-coordination of CsCl . This high coordination number and softness (tendency to form covalent bonds) are properties exploited in separating Cs^+ from other cations in the remediation of nuclear wastes, where $^{137}\text{Cs}^+$ must be separated from large amounts of nonradioactive K^+ . ^[28]

Halides

Caesium fluoride (CsF) is a hygroscopic white solid that is widely used in organofluorine chemistry as a source of fluoride anions. ^[30] Caesium fluoride has the halite structure, which means that the Cs^+ and F^- pack in a cubic closest packed array as do Na^+ and Cl^- in sodium chloride. ^[21] Notably, caesium and fluorine have the lowest and highest electronegativities, respectively, among all the known elements.

Caesium chloride (CsCl) crystallizes in the simple cubic crystal system. Also called the "caesium chloride structure", ^[23] this structural motif is composed of a primitive cubic lattice with a two-atom basis, each with an eightfold coordination; the chloride atoms lie upon the lattice points at the edges of the cube, while the caesium atoms lie in the holes in the centre of the cubes. This structure is shared with CsBr and CsI , and many other compounds that do not contain Cs. In contrast, most other alkaline halides have the sodium chloride (NaCl) structure. ^[23] The CsCl structure is preferred because Cs^+ has an ionic radius of 174 pm and Cl^- 181 pm. ^[31]

Oxides



More so than the other alkali metals, caesium forms numerous binary compounds with oxygen. When caesium burns in air, the superoxide CsO₂ is the main product.^[32] The "normal" caesium oxide (Cs₂O) forms yellow-orange hexagonal crystals,^[33] and is the only oxide of the anti-CdCl₂ type.^[34] It vaporizes at 250 °C (482 °F), and decomposes to caesium metal and the peroxide Cs₂O₂ at temperatures above 400 °C (752 °F).^[35] In addition to the superoxide and the ozonide CsO₃,^{[36][37]} several brightly coloured suboxides have also been studied.^[38] These include Cs₇O, Cs₄O, Cs₁₁O₃, Cs₃O (dark-green^[39]), CsO, Cs₃O₂,^[40] as well as Cs₇O₂.^{[41][42]} The latter may be heated in a vacuum to generate Cs₂O.^[34] Binary compounds with sulfur, selenium, and tellurium also

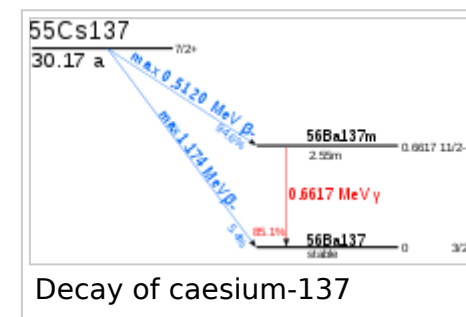
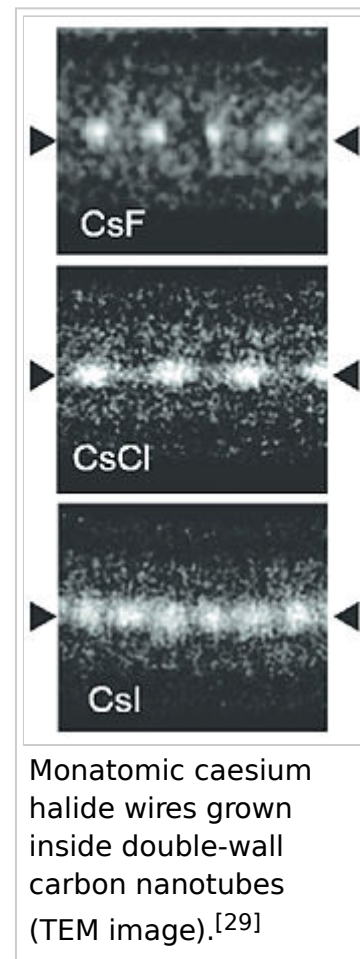
exist.^[10]

Isotopes

Caesium has 39 known isotopes, ranging in mass number (i.e. number of nucleons in the nucleus) from 112 to 151. Several of these are synthesized from lighter elements by the slow neutron capture process (S-process) inside old stars^[43] and by the R-process in supernova explosions.^[44] The only stable caesium isotope is ¹³³Cs, with 78 neutrons. Although it has a large nuclear spin ($\frac{7}{2}+$), nuclear magnetic resonance studies can use this isotope at a resonating frequency of 11.7 MHz.^[45]

The radioactive ¹³⁵Cs has a very long half-life of about 2.3 million years, longest of all radioactive isotopes of caesium. ¹³⁷Cs and ¹³⁴Cs have half-lives of 30 and two years, respectively. ¹³⁷Cs decomposes to a short-lived ^{137m}Ba by beta decay, and then to nonradioactive barium, while ¹³⁴Cs transforms into ¹³⁴Ba directly. The isotopes with mass numbers of 129, 131, 132 and 136, have half-lives between a day and two weeks, while most of the other isotopes have half-lives from a few seconds to fractions of a second. At least 21 metastable nuclear isomers exist. Other than ^{134m}Cs (with a half-life of just under 3 hours), all are very unstable and decay with half-lives of a few minutes or less.^{[46][47]}

The isotope ¹³⁵Cs is one of the long-lived fission products of uranium produced in nuclear reactors.^[48] However, this fission product yield is reduced in most reactors because the predecessor, ¹³⁵Xe, is a potent neutron poison and frequently transmutes to stable ¹³⁶Xe before it can decay to ¹³⁵Cs.^{[49][50]}



The beta decay from ^{137}Cs to $^{137\text{m}}\text{Ba}$ is a strong emission of gamma radiation.^[51] ^{137}Cs and ^{90}Sr are the principal medium-lived products of nuclear fission, and the prime sources of radioactivity from spent nuclear fuel after several years of cooling, lasting several hundred years.^[52] Those two isotopes are the largest source of residual radioactivity in the area of the Chernobyl disaster.^[53] Because of the low capture rate, disposing of ^{137}Cs through neutron capture is not feasible and the only current solution is allow to decay over time.^[54]

Almost all caesium produced from nuclear fission comes from the beta decay of originally more neutron-rich fission products, passing through various isotopes of iodine and xenon.^[55] Because iodine and xenon are volatile and can diffuse through nuclear fuel or air, radioactive caesium is often created far from the original site of fission.^[56] With nuclear weapons testing in the 1950s through the 1980s, ^{137}Cs was released into the atmosphere and returned to the surface of the earth as a component of radioactive fallout. It is a ready marker of the movement of soil and sediment from those times.^[10]

Occurrence

Caesium is a relatively rare element estimated to average 3 parts per million in the Earth's crust.^[57] It is the 45th most abundant element and the 36th among the metals. Nevertheless, it is more abundant than such elements as antimony, cadmium, tin, and tungsten, and two orders of magnitude more abundant than mercury and silver; it is 3.3% as abundant as rubidium, with which it is closely associated, chemically.^[10]

Due to its large ionic radius, caesium is one of the "incompatible elements".^[58] During magma crystallization, caesium is concentrated in the liquid phase and crystallizes last. Therefore, the largest deposits of caesium are zone pegmatite ore bodies formed by this enrichment process. Because caesium does not substitute for potassium as readily as rubidium, the alkali evaporite minerals sylvite (KCl) and carnallite ($\text{KMgCl}_3 \cdot 6\text{H}_2\text{O}$) may contain only 0.002% caesium. Consequently, Cs is found in few minerals. Percentage amounts of caesium may be found in beryl ($\text{Be}_3\text{Al}_2(\text{SiO}_3)_6$) and avogadrite ($(\text{K,Cs})\text{BF}_4$), up to 15 wt% Cs_2O in the closely related mineral pezzottaite ($\text{Cs}(\text{Be}_2\text{Li})\text{Al}_2\text{Si}_6\text{O}_{18}$), up to 8.4 wt% Cs_2O in the rare mineral londonite

$(\text{Cs,K})\text{Al}_4\text{Be}_4(\text{B,Be})_{12}\text{O}_{28}$, and less in the more widespread rhodizite.^[10] The only economically important ore for caesium is pollucite $\text{Cs}(\text{AlSi}_2\text{O}_6)$, which is found in a few places around the world in zoned pegmatites, associated with the more commercially important lithium minerals, lepidolite and petalite. Within the pegmatites, the large grain size and the strong separation of the minerals results in high-grade ore for mining.^[59]



Pollucite, a caesium mineral

One of the world's most significant and richest sources of caesium is the Tanco Mine at Bernic Lake in Manitoba, Canada, estimated to contain 350,000 metric tons of pollucite ore, representing more than two-thirds of the world's reserve base.^{[59][60]} Although the stoichiometric content of caesium in pollucite is 42.6%, pure pollucite samples from this deposit contain only about 34% caesium, while the average content is 24 wt%.^[60] Commercial pollucite contains more than 19% caesium.^[61] The Bikita pegmatite deposit in Zimbabwe is mined for its petalite, but it also contains a significant amount of pollucite. Another notable source of pollucite is in the Karibib Desert, Namibia.^[60] At the present rate of world mine production of 5 to 10 metric tons per year, reserves will last for thousands of years.^[10]

Source

- Wikipedia: Cesium (<https://en.wikipedia.org/wiki/Cesium>)