

Iron

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Iron is a chemical element with symbol **Fe** (from Latin: *ferrum*) and atomic number 26. It is a metal in the first transition series. It is by mass the most common element on Earth, forming much of Earth's outer and inner core. It is the fourth most common element in the Earth's crust. Its abundance in rocky planets like Earth is due to its abundant production by fusion in high-mass stars, where it is the last element to be produced with release of energy before the violent collapse of a supernova, which scatters the iron into space.

Like the other group 8 elements, ruthenium and osmium, iron exists in a wide range of oxidation states, -2 to $+6$, although $+2$ and $+3$ are the most common. Elemental iron occurs in meteoroids and other low oxygen environments, but is reactive to oxygen and water. Fresh iron surfaces appear lustrous silvery-gray, but oxidize in normal air to give hydrated iron oxides, commonly known as rust. Unlike the metals that form passivating oxide layers, iron oxides occupy more volume than the metal and thus flake off, exposing fresh surfaces for corrosion.

Iron metal has been used since ancient times, although copper alloys, which have lower melting temperatures, were used even earlier in human history. Pure iron is relatively soft, but is unobtainable by smelting because it is significantly hardened and strengthened by impurities, in particular carbon, from the smelting process. A certain proportion of carbon (between 0.002% and 2.1%) produces steel, which may be up to 1000 times harder than pure iron. Crude iron metal is produced in blast furnaces, where ore is reduced by coke to pig iron, which has a high carbon content. Further refinement with oxygen reduces the carbon content to the correct proportion to make steel. Steels and iron alloys formed with other metals (alloy steels) are by far the most common industrial metals because they have a great range of desirable properties and iron-bearing rock is abundant.

Iron, $_{26}\text{Fe}$



Spectral lines of iron

General properties

Name, symbol	iron, Fe
Appearance	lustrous metallic with a grayish tinge

Iron in the periodic table

Atomic number (<i>Z</i>)	26
Group, block	group 8, d-block
Period	period 4
Element category	▢ transition metal
Standard atomic weight (\pm) (<i>A</i> _r)	55.845(2) ^[1]
Electron configuration	[Ar] 3d ⁶ 4s ²
per shell	2, 8, 14, 2

Physical properties

Phase	solid
Melting point	1811 K (1538 °C, 2800 °F)
Boiling point	3134 K (2862 °C, 5182 °F)

Iron chemical compounds have many uses. Iron oxide mixed with aluminium powder can be ignited to create a thermite reaction, used in welding and purifying ores. Iron forms binary compounds with the halogens and the chalcogens. Among its organometallic compounds is ferrocene, the first sandwich compound discovered.

Iron plays an important role in biology, forming complexes with molecular oxygen in hemoglobin and myoglobin; these two compounds are common oxygen transport proteins in vertebrates. Iron is also the metal at the active site of many important redox enzymes dealing with cellular respiration and oxidation and reduction in plants and animals. A human male of average height has about 4 grams of iron in his body, a female about 3.5 grams. This iron is distributed throughout the body in hemoglobin, tissues, muscles, bone marrow, blood proteins, enzymes, ferritin, hemosiderin, and transport in plasma.^[4]

Characteristics

Mechanical properties

The mechanical properties of iron and its alloys can be evaluated using a variety of tests, including the Brinell test, Rockwell test and the Vickers hardness test. The data on iron is so consistent that it is often used to calibrate measurements or to compare tests.^{[6][7]} However, the mechanical properties of iron are significantly affected by the sample's purity: pure, single crystals of iron are actually softer than aluminium,^[5] and the purest industrially produced iron (99.99%) has a hardness of 20–30 Brinell.^[8] An increase in the carbon content will cause a significant increase in the hardness and tensile strength of iron. Maximum hardness of 65 R_c is achieved with a 0.6% carbon content, although the alloy has low tensile strength.^[9] Because of the softness of iron, it is much easier to work with than its heavier congeners ruthenium and osmium.^[10]

Density near r.t.	7.874 g/cm ³
when liquid, at m.p.	6.98 g/cm ³
Heat of fusion	13.81 kJ/mol
Heat of vaporization	340 kJ/mol
Molar heat capacity	25.10 J/(mol·K)

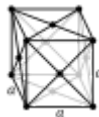
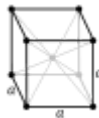
Vapor pressure						
P (Pa)	1	10	100	1 k	10 k	100 k
at T (K)	1728	1890	2091	2346	2679	3132

Atomic properties	
Oxidation states	−4, −2, −1, +1, ^[2] +2, +3 , +4, +5, ^[3] +6 (an amphoteric oxide)
Electronegativity	Pauling scale: 1.83
Ionization energies	1st: 762.5 kJ/mol 2nd: 1561.9 kJ/mol 3rd: 2957 kJ/mol (more)
Atomic radius	empirical: 126 pm
Covalent radius	Low spin: 132±3 pm High spin: 152±6 pm

Miscellanea	
Crystal structure	body-centered cubic (bcc) a=286.65 pm

Crystal structure	face-centered cubic (fcc) between 1185–1667 K
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Speed of sound thin rod	5120 m/s (at r.t.) (electrolytic)
Thermal expansion	11.8 μm/(m·K) (at 25 °C)



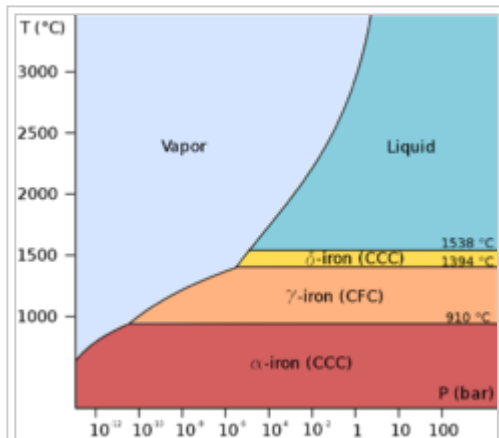
Because of its significance for planetary cores, the physical properties of iron at high pressures and temperatures have also been studied extensively. The form of iron that is stable under standard conditions can be subjected to pressures up to ca. 15 GPa before transforming into a high-pressure form, as described in the next section.

Phase diagram and allotropes

Iron represents an example of allotropy in a metal. There are at least four allotropic forms of iron, known as α , γ , δ , and ϵ ; at very high pressures and temperatures, some controversial experimental evidence exists for a stable β phase.^[11]

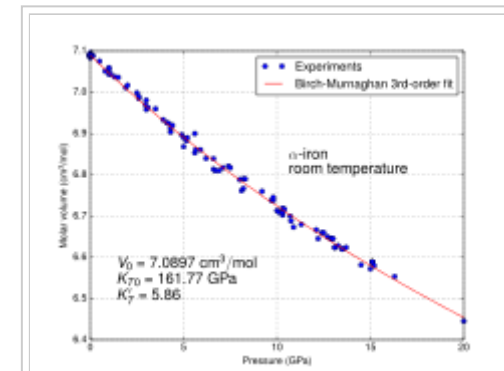
As molten iron cools past its freezing point of 1538 °C, it crystallizes into its δ allotrope, which has a body-centered cubic (bcc) crystal structure. As it cools further to 1394 °C, it changes to its γ -iron allotrope, a face-centered cubic (fcc) crystal structure, or austenite. At 912 °C and below, the crystal structure again becomes the bcc α -iron allotrope, or ferrite. Finally, at 770 °C (the Curie point, T_C) iron's magnetic ordering changes from paramagnetic to ferromagnetic. As the iron passes through the Curie temperature there is no change in crystalline structure, but there is a change in "domain structure", where each domain contains iron atoms with a particular electronic spin. In unmagnetized iron, all the electronic spins of the atoms within one domain have the same axis orientation; however, the electrons of neighboring domains have other orientations with the result of mutual cancellation and no magnetic field. In magnetized iron, the electronic spins of the domains are aligned and the magnetic effects are reinforced. Although each domain contains billions of atoms, they are very small, about 10 micrometres across.^[12] This happens because the two unpaired electrons on each iron atom are in the d_{z^2} and $d_{x^2 - y^2}$ orbitals, which do not point directly at the nearest neighbors in the body-centered cubic lattice and therefore do not participate in metallic bonding; thus, they can interact magnetically with each other so that their spins align.^[13]

Thermal conductivity	80.4 W/(m·K)				
Electrical resistivity	96.1 nΩ·m (at 20 °C)				
Curie point	1043 K				
Magnetic ordering	ferromagnetic				
Young's modulus	211 GPa				
Shear modulus	82 GPa				
Bulk modulus	170 GPa				
Poisson ratio	0.29				
Mohs hardness	4				
Vickers hardness	608 MPa				
Brinell hardness	200–1180 MPa				
CAS Number	7439-89-6				
History					
Discovery	before 5000 BC				
Most stable isotopes of iron					
iso	NA	half-life	DM	DE (MeV)	DP
54Fe	5.85%	is stable with 28 neutrons			
55Fe	syn	2.73 y	ε	0.231	55Mn
56Fe	91.75%	is stable with 30 neutrons			
57Fe	2.12%	is stable with 31 neutrons			
58Fe	0.28%	is stable with 32 neutrons			
59Fe	syn	44.503 d	β−	1.565	59Co
60Fe	syn	2.6×10 ⁶ y	β−	3.978	60Co



Low-pressure phase diagram of pure iron

At pressures above approximately 10 GPa and temperatures of a few hundred kelvin or less, α -iron changes into a hexagonal close-packed (hcp) structure, which is also known as ϵ -iron; the higher-temperature γ -phase also changes into ϵ -iron, but does so at higher pressure. The β -phase, if it exists, would appear at pressures of at least 50 GPa and temperatures of at least 1500 K and have an orthorhombic or a double hcp structure.^[11] These high-pressure phases of iron are important as endmember models for the solid parts of planetary cores. The inner core of the Earth is generally presumed to be an iron-nickel alloy with ϵ (or β) structure.^[14] Somewhat confusingly, the term " β -iron" is sometimes also used to refer to α -iron above its Curie point, when it changes from being ferromagnetic to paramagnetic, even though its crystal structure has not changed.^[13]



Molar volume vs. pressure for α iron at room temperature

The melting point of iron is experimentally well defined for pressures less than 50 GPa. For greater pressures, studies put the γ - ϵ -liquid triple point at pressures that differ by tens of gigapascals and 1000 K in the melting point. Generally speaking, molecular dynamics computer simulations of iron melting and shock wave experiments suggest higher melting points and a much steeper slope of the melting curve than static experiments carried out in diamond anvil cells.^[15] The melting and boiling points of iron, along with its enthalpy of atomization, are lower than those of the earlier 3d elements from scandium to chromium, showing the lessened contribution of the 3d electrons to metallic bonding as they are attracted more and more into the inert core by the nucleus;^[16] however, they are higher than the values for the previous element manganese because that element has a half-filled 3d subshell and consequently its d-electrons are not easily delocalized. This same trend appears for ruthenium but not osmium.^[10]

Isotopes

Naturally occurring iron consists of four stable isotopes: 5.845% of ^{54}Fe , 91.754% of ^{56}Fe , 2.119% of ^{57}Fe and 0.282% of ^{58}Fe . Of these stable isotopes, only ^{57}Fe has a nuclear spin ($-\frac{1}{2}$). The nuclide ^{54}Fe theoretically can undergo double electron capture to ^{54}Cr , but the process has never been observed and only a lower limit on the half-life of 3.1×10^{22} years has been established.^[17]

^{60}Fe is an extinct radionuclide of long half-life (2.6 million years).^[18] It is not found on Earth, but its ultimate decay product is its granddaughter, the stable nuclide ^{60}Ni .^[17] Much of the past work on isotopic composition of iron has focused on the nucleosynthesis of ^{60}Fe through studies of meteorites and ore formation. In the last decade, advances in mass spectrometry have allowed the detection and quantification of minute, naturally occurring variations in the ratios of the stable isotopes of iron. Much of this work is driven by the Earth and planetary science communities, although applications to biological and industrial systems are emerging.^[19]

In phases of the meteorites *Semarkona* and *Chervony Kut*, a correlation between the concentration of ^{60}Ni , the granddaughter of ^{60}Fe , and the abundance of the stable iron isotopes provided evidence for the existence of ^{60}Fe at the time of formation of the Solar System. Possibly the energy released by the decay of ^{60}Fe , along with that released by ^{26}Al , contributed to the remelting and differentiation of asteroids after their formation 4.6 billion years ago. The abundance of ^{60}Ni present in extraterrestrial material may bring further insight into the origin and early history of the Solar System.^[20]

The most abundant iron isotope ^{56}Fe is of particular interest to nuclear scientists because it represents the most common endpoint of nucleosynthesis.^[21] Since ^{56}Ni (14 alpha particles) is easily produced from lighter nuclei in the alpha process in nuclear reactions in supernovae (see silicon burning process), it is the endpoint of fusion chains inside extremely massive stars, since addition of another alpha particle, resulting in ^{60}Zn , requires a great deal more energy. This ^{56}Ni , which has a half-life of about 6 days, is created in quantity in these stars, but soon decays by two successive positron emissions within supernova decay products in the supernova remnant gas cloud, first to radioactive ^{56}Co , and then to stable ^{56}Fe . As such, iron is the most abundant element in the core of red giants, and is the most abundant metal in iron meteorites and in the dense metal cores of planets such as Earth.^[22] It is also very common in the universe, relative to other stable metals of approximately the same atomic weight.^{[22][23]} Iron is the sixth most abundant element in the Universe, and the most common refractory element.^[24]

Although a further tiny energy gain could be extracted by synthesizing ^{62}Ni , which has a marginally higher binding energy than ^{56}Fe , conditions in stars are unsuitable for this process. Element production in supernovas and distribution on Earth greatly favor iron over nickel, and in any case, ^{56}Fe still has a lower mass per nucleon than ^{62}Ni due to its higher fraction of lighter protons.^[25] Hence, elements heavier than iron require a supernova for their formation, involving rapid neutron capture by starting ^{56}Fe nuclei.^[22]

In the far future of the universe, assuming that proton decay does not occur, cold fusion occurring via quantum tunnelling would cause the light nuclei in ordinary matter to fuse into ^{56}Fe nuclei. Fission and alpha-particle emission would then make heavy nuclei decay into iron, converting all stellar-mass objects to cold spheres of pure iron.^[26]

Occurrence

Metallic or native iron is rarely found on the surface of the Earth because it tends to oxidize, but its oxides are pervasive and represent the primary ores. While it makes up about 5% of the Earth's crust, both the Earth's inner and outer core are believed to consist largely of an iron-nickel alloy constituting 35% of the mass of the Earth as a whole. Iron is consequently the most abundant element on Earth, but only the fourth most abundant element in the Earth's crust, after oxygen, silicon, and aluminium.^{[27][28]} Most of the iron in the crust is found combined with oxygen as iron oxide minerals such as hematite (Fe_2O_3), magnetite (Fe_3O_4), and siderite (FeCO_3). Many igneous rocks also contain the sulfide minerals pyrrhotite and pentlandite.^{[29][30]}

Ferropericlase (Mg,FeO), a solid solution of periclase (MgO) and wüstite (FeO), makes up about 20% of the volume of the lower mantle of the Earth, which makes it the second most abundant mineral phase in that region after silicate perovskite (Mg,FeSiO_3); it also is the major host for iron in the lower mantle.^[31] At the bottom of the transition zone of the mantle, the reaction $\gamma\text{-(Mg,Fe)}_2[\text{SiO}_4] \leftrightarrow (\text{Mg,Fe})[\text{SiO}_3] + (\text{Mg,Fe})\text{O}$ transforms γ -olivine into a mixture of perovskite and ferropericlase and vice versa. In the literature, this mineral phase of the lower mantle is also often called magnesiowüstite.^[32] Silicate perovskite may form up to 93% of the lower mantle,^[33] and the magnesium iron form, $(\text{Mg,Fe})\text{SiO}_3$, is considered to be the most abundant mineral in the Earth, making up 38% of its volume.^[34]

Large deposits of iron are found in banded iron formations. These geological formations are a type of rock consisting of repeated thin layers of iron oxides alternating with bands of iron-poor shale and chert. The banded iron formations were laid down in the time between 3,700 million years ago and 1,800 million years ago.^{[35][36]}

The mentioned iron compounds have been used as pigments (compare ochre) since historical time and contribute as well to the color of various geological formations, e.g. the Bundsandstein (British Bunter, colored sandstein).^[37] In the case of the Eisensandstein (a jurassic 'iron sandstone', e.g. from Donzdorf) in Germany^[38] and Bath stone in the UK, iron pigments contribute to the yellowish color of large amounts of historical buildings and sculptures.^[39] The proverbial red color of the surface of Mars is derived from an iron oxide-rich regolith.^[40]



Iron meteorites, similar in composition to the Earth's inner- and outer core



Ochre path in the Roussillon

Significant amounts of iron occur in the iron sulfide mineral pyrite (FeS_2), but it is difficult to extract iron from it and it is therefore not used. In fact, iron is so common that production generally focuses only on ores with very high quantities of it. During weathering, iron tends to leach from sulfide deposits as the sulfate and from silicate deposits as the bicarbonate. Both of these are oxidized in aqueous solution and precipitate in even mildly elevated pH as iron(III) oxide.^[41]

About 1 in 20 meteorites consist of the unique iron-nickel minerals taenite (35–80% iron) and kamacite (90–95% iron). Although rare, iron meteorites are the main form of natural metallic iron on the Earth's surface.^[42] According to the International Resource Panel's Metal Stocks in Society report, the global stock of iron in use in society is 2200 kg per capita. Much of this is in more-developed countries (7000–14000 kg per capita) rather than less-developed countries (2000 kg per capita).^[43]

Source

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