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Extraction of aluminium from bauxite pdf

Chemical element with atomic number 13 Chemical element, symbol Al and atomic number 13Aluminium, 13AlAluminiumPronunciationaluminium: /ˌæl(j)ʊˈmɪniəm/ (listen)(AL-(y)uu-MIN-ee-əm)aluminum: /əˈljuːmɪnəm/ (listen)(ə-LEW-mi-nəm)Alternative namealuminum (U.S., Canada)Appearancesilvery gray metallicStandard atomic weight Ar, std(Al)26.9815384(3)[1]Aluminium in the periodic table Hydrogen Helium Lithium Beryllium Boron Carbon Nitrogen Oxygen Fluorine Neon Sodium Magnesium Aluminium Silicon Phosphorus Sulfur Chlorine Argon Potassium Calcium Scandium Titanium Vanadium Chromium Manganese Iron Cobalt Nickel Copper Zinc Gallium Germanium Arsenic Selenium Bromine Krypton Rubidium Strontium Yttrium Zirconium Niobium Molybdenum Technetium Ruthenium Rhodium Palladium Silver Cadmium Indium Tin Antimony Tellurium Iodine Xenon Caesium Barium Lanthanum Cerium Praseodymium Neodymium Promethium Samarium Europium Gadolinium Terbium Dysprosium Holmium Erbium Thulium Ytterbium Lutetium Hafnium Tantalum Tungsten Rhenium Osmium Iridium Platinum Gold Mercury (element) Thallium Lead Bismuth Polonium Astatine Radon Francium Radium Actinium Thorium Protactinium Uranium Neptunium Plutonium Americium Curium Berkelium Californium Einsteinium Fermium Mendelevium Nobelium Lawrencium Rutherfordium Dubnium Seaborgium Bohrium Hassium Meitnerium Darmstadtium Roentgenium Copernicium Nihonium Flerovium Moscovium Livermorium Tennessine Oganesson B↑Al↓Ga magnesium ← aluminium → silicon Atomic number (Z)13Groupgroup 13 (boron group)Periodperiod 3Block p-blockElectron configuration[Ne] 3s2 3p1Electrons per shell2, 8, 3Physical propertiesPhase at STPsolidMelting point933.47 K (660.32 °C, 1220.58 °F) Boiling point2743 K (2470 °C, 4478 °F) Density (near r.t.)2.70 g/cm3when liquid (at m.p.)2.375 g/cm3 Heat of fusion10.71 kJ/mol Heat of vaporization284 kJ/mol Molar heat capacity24.20 J/(mol·K) Vapor pressure P (Pa) 1 10 100 1 k 10 k 100 k at T (K) 1482 1632 1817 2054 2364 2790 Atomic propertiesOxidation states−2, −1, +1,[2] +2,[3] +3 (an amphoteric oxide)ElectronegativityPauling scale: 1.61 Ionization energies1st: 577.5 kJ/mol 2nd: 1816.7 kJ/mol 3rd: 2744.8 kJ/mol (more) Atomic radiusempirical: 143 pm Covalent radius121±4 pm Van der Waals radius184 pm Spectral lines of aluminiumOther propertiesNatural occurrenceprimordialCrystal structure face-centered cubic (fcc)Speed of sound thin rod(rolled) 5000 m/s (at r.t.) Thermal expansion23.1 µm/(m·K) (at 25 °C) Thermal conductivity237 W/(m·K) Electrical resistivity26.5 nΩ·m (at 20 °C) Magnetic orderingparamagnetic[4] Molar magnetic susceptibility+16.5×10−6 cm3/molYoung's modulus70 GPa Shear modulus26 GPa Bulk modulus76 GPa Poisson ratio0.35 Mohs hardness2.75 Vickers hardness160–350 MPa Brinell hardness160–550 MPa CAS Number7429-90-5 HistoryNamingfrom alumine, obsolete name for aluminaPredictionAntoine Lavoisier (1782)DiscoveryHans Christian Ørsted (1824)Named byHumphry Davy (1812[a])Main isotopes of aluminium Isotope Abundance Half-life (t1/2) Decay mode Product 26Al trace 7.17×105 y β+ 26Mg ε 26Mg γ – 27Al 100% stable Category: Aluminiumviewtalkedit | references Aluminium (aluminum in American and Canadian English) is a chemical element with the symbol Al and atomic number 13. Aluminium has a density lower than those of other common metals, at approximately one third that of steel. It has a great affinity towards oxygen, and forms a protective layer of oxide on the surface when exposed to air. Aluminium visually resembles silver, both in its color and in its great ability to reflect light. It is soft, non-magnetic and ductile. It has one stable isotope, 27Al; this isotope is very common, making aluminium the twelfth most common element in the Universe. The radioactivity of 26Al is used in radiodating. Chemically, aluminium is a weak metal in the boron group; as it is common for the group, aluminium forms compounds primarily in the +3 oxidation state. The aluminium cation Al3+ is small and highly charged; as such, it is polarizing, and bonds aluminium forms tend towards covalency. The strong affinity towards oxygen leads to aluminium's common association with oxygen in nature in the form of oxides; for this reason, aluminium is found on Earth primarily in rocks in the crust, where it is the third most abundant element after oxygen and silicon, rather than in the mantle, and virtually never as the free metal. The discovery of aluminium was announced in 1825 by Danish physicist Hans Christian Ørsted. The first industrial production of aluminium was initiated by French chemist Henri Étienne Sainte-Claire Deville in 1856. Aluminium became much more available to the public with the Hall–Héroult process developed independently by French engineer Paul Héroult and American engineer Charles Martin Hall in 1886, and the mass production of aluminium led to its extensive use in industry and everyday life. In World Wars I and II, aluminium was a crucial strategic resource for aviation. In 1954, aluminium became the most produced non-ferrous metal, surpassing copper. In the 21st century, most aluminium was consumed in transportation, engineering, construction, and packaging in the United States, Western Europe, and Japan. Despite its prevalence in the environment, no living organism is known to use aluminium salts metabolically, but aluminium is well tolerated by plants and animals. Because of the abundance of these salts, the potential for a biological role for them is of continuing interest, and studies continue. Physical characteristics Isotopes Main article: Isotopes of aluminium Of aluminium isotopes, only 27Al is stable. This situation is common for elements with an odd atomic number.[b] It is the only primordial aluminium isotope, i.e. the only one that has existed on Earth in its current form since the formation of the planet. Nearly all aluminium on Earth is present as this isotope, which makes it a mononuclidic element and means that its standard atomic weight is virtually the same as that of the isotope. This makes aluminium very useful in nuclear magnetic resonance (NMR), as its single stable isotope has a high NMR sensitivity.[6] The standard atomic weight of aluminium is low in comparison with many other metals,[c] which has consequences for the element's properties (see below). All other isotopes of aluminium are radioactive. The most stable of these is 26Al: while it was present along with stable 27Al in the interstellar medium from which the Solar System formed, having been produced by stellar nucleosynthesis as well, its half-life is only 717,000 years and therefore a detectable amount has not survived since the formation of the planet.[7] However, minute traces of 26Al are produced from argon in the atmosphere by spallation caused by cosmic ray protons. The ratio of 26Al to 10Be has been used for radiodating of geological processes over 105 to 106 year time scales, in particular transport, deposition, sediment storage, burial times, and erosion.[8] Most meteorite scientists believe that the energy released by the decay of 26Al was responsible for the melting and differentiation of some asteroids after their formation 4.55 billion years ago.[9] The remaining isotopes of aluminium, with mass numbers ranging from 22 to 43, all have half-lives well under an hour. Three metastable states are known, all with half-lives under a minute.[5] Electron shell An aluminium atom has 13 electrons, arranged in an electron configuration of [Ne] 3s2 3p1,[10] with three electrons beyond a stable noble gas configuration. Accordingly, the combined first three ionization energies of aluminium are far lower than the fourth ionization energy alone.[11] Such an electron configuration is shared with the other well-characterized members of its group, boron, gallium, indium, and thallium; it is also expected for nihonium. Aluminium can relatively easily surrender its three outermost electrons in many chemical reactions (see below). The electronegativity of aluminium is 1.61 (Pauling scale).[12] High-resolution STEM-HAADF micrograph of Al atoms viewed along the [001] zone axis. A free aluminium atom has a radius of 143 pm.[13] With the three outermost electrons removed, the radius shrinks to 39 pm for a 4-coordinated atom or 53.5 pm for a 6-coordinated atom.[13] At standard temperature and pressure, aluminium atoms (when not affected by atoms of other elements) form a face-centered cubic crystal system bound by metallic bonding provided by atoms' outermost electrons; hence aluminium (at these conditions) is a metal.[14] This crystal system is shared by many other metals, such as lead and copper; the size of a unit cell of aluminium is comparable to that of those other metals.[14] The system, however, is not shared by the other members of its group; boron has ionization energies too high to allow metallization, thallium has a hexagonal close-packed structure, and gallium and indium have unusual structures that are not close-packed like those of aluminium and thallium. The few electrons are available for metallic bonding in aluminium metal are a probably cause for it being soft with a low melting point and low electrical resistivity.[15] Bulk Aluminium metal has an appearance ranging from silvery white to dull gray, depending on the surface roughness.[d] A fresh film of aluminium serves as a good reflector (approximately 92%) of visible light and an excellent reflector (as much as 98%) of medium and far infrared radiation.[citation needed] Aluminium mirrors are the most reflective of all metal mirrors for the near ultraviolet and far infrared light, and one of the most reflective in the visible spectrum, nearly on par with silver, and the two therefore look similar. Aluminium is also good at reflecting solar radiation, although prolonged exposure to sunlight in air add wear to the surface of the metal; this may be prevented if aluminium is anodized, which adds a protective layer of oxide on the surface.[citation needed] The density of aluminium is 2.70 g/cm3, about 1/3 that of steel, much lower than other commonly encountered metals, making aluminium parts easily identifiable through their lightness.[18] Aluminium's low density compared to most other metals arises from the fact that its nuclei are much lighter, while difference in the unit cell size does not compensate for this difference. The only lighter metals are the metals of groups 1 and 2, which apart from beryllium and magnesium are too reactive for structural use (and beryllium is very toxic).[19] Aluminium is not as strong or stiff as steel, but the low density makes up for this in the aerospace industry and for many other applications where light weight and relatively high strength are crucial.[20] Pure aluminium is quite soft and lacking in strength. In most applications various aluminium alloys are used instead because of their higher strength and hardness.[21] The yield strength of pure aluminium is 7–11 MPa, while aluminium alloys have yield strengths ranging from 200 MPa to 600 MPa.[22] Aluminium is ductile, with a percent elongation of 50-70%,[23] and malleable allowing it to be easily drawn and extruded.[24] It is also easily machined and cast.[24] Aluminium is an excellent thermal and electrical conductor, having around 60% the conductivity of copper, both thermal and electrical, while having only 30% of copper's density.[25] Aluminium is capable of superconductivity, with a superconducting critical temperature of 1.2 kelvin and a critical magnetic field of about 100 gauss (10 milliteslas).[26] It is paramagnetic and thus essentially unaffected by static magnetic fields.[27] The high electrical conductivity, however, means that it is strongly affected by alternating magnetic fields through the induction of eddy currents.[28] Chemistry Main article: Compounds of aluminium Aluminium combines characteristics of pre- and post-transition metals. Since it has few available electrons for metallic bonding, like its heavier group 13 congeners, it has the characteristic physical properties of a post-transition metal, with longer-than-expected interatomic distances.[15] Furthermore, as Al3+ is a small and highly charged cation, it is strongly polarizing and bonding in aluminium compounds tends towards covalency;[29] this behavior is similar to that of beryllium (Be2+), and the two display an example of a diagonal relationship.[30] The underlying core under aluminium's valence shell is that of the preceding noble gas, whereas those of its heavier congeners gallium, indium, thallium, and nihonium also include a filled d-subshell and in some cases a filled f-subshell. Hence, the inner electrons of aluminium shield the valence electrons almost completely, unlike those of aluminium's heavier congeners. As such, aluminium is the most electropositive metal in its group, and its hydroxide is in fact more basic than that of gallium.[29][e] Aluminium also bears minor similarities to the metalloid boron in the same group: AlX3 compounds are valence isoelectronic to BX3 compounds (they have the same valence electronic structure), and both behave as Lewis acids and readily form adducts.[31] Additionally, one of the main motifs of boron chemistry is regular icosahedral structures, and aluminium forms an important part of many icosahedral quasicrystal alloys, including the Al–Zn–Mg class.[32] Aluminium has a high chemical affinity to oxygen, which renders it suitable for use as a reducing agent in the thermite reaction. A fine powder of aluminium metal reacts explosively on contact with liquid oxygen; under normal conditions, however, aluminium forms a thin oxide layer (~5 nm at room temperature)[33] that protects the metal from further corrosion by oxygen, water, or dilute acid, a process termed passivation.[29][34] Because of its general resistance to corrosion, aluminium is one of the few metals that retains silvery reflectance in finely powdered form, making it an important component of silver-colored paints.[35] Aluminium is not attacked by oxidizing acids because of its passivation. This allows aluminium to be used to store reagents such as nitric acid, concentrated sulfuric acid, and some organic acids.[36] In hot concentrated hydrochloric acid, aluminium reacts with water with evolution of hydrogen, and in aqueous sodium hydroxide or potassium hydroxide at room temperature to form aluminates—protective passivation under these conditions is negligible.[37] Aqua regia also dissolves aluminium.[36] Aluminium is corroded by dissolved chlorides, such as common sodium chloride, which is why household plumbing is never made from aluminium.[37] The oxide layer on aluminium is also destroyed by contact with mercury due to amalgamation or with salts of some electropositive metals.[29] As such, the strongest aluminium alloys are less corrosion-resistant due to galvanic reactions with alloyed copper,[22] and aluminium's corrosion resistance is greatly reduced by aqueous salts, particularly in the presence of dissimilar metals.[15] Aluminium reacts with most nonmetals upon heating, forming compounds such as aluminium nitride (AlN), aluminium sulfide (Al2S3), and the aluminium halides (AlX3). It also forms a wide range of intermetallic compounds involving metals from every group on the periodic table.[29] Inorganic compounds The vast majority of compounds, including all aluminium-containing minerals and all commercially significant aluminium compounds, feature aluminium in the oxidation state 3+. The coordination number of such compounds varies, but generally Al3+ is either six- or four-coordinate. Almost all compounds of aluminium(III) are colorless. [29] Aluminium hydrolysis as a function of pH. Coordinated water molecules are omitted. (Data from Baes and Mesmer)[38] In aqueous solution, Al3+ exists as the hexaaqua cation [Al(H2O)6]3+, which has an approximate Ka of 10−5.[6] Such solutions are acidic as this cation can act as a proton donor and progressively hydrolyze until a precipitate of aluminium hydroxide, Al(OH)3, forms. This is useful for clarification of water, as the precipitate nucleates on suspended particles in the water, hence removing them. Increasing the pH even further leads to the hydroxide dissolving again as aluminate, [Al(H2O)2(OH)4]−, is formed. Aluminium hydroxide forms both salts and aluminates and dissolves in acid and alkali, as well as on fusion with acidic and basic oxides.[29] This behavior of Al(OH)3 is termed amphoterism and is characteristic of weakly basic cations that form insoluble hydroxides and whose hydrated species can also donate their protons. One effect of this is that aluminium salts with weak acids are hydrolyzed in water to the aquated hydroxide and the corresponding nonmetal hydride: for example, aluminium sulfide yields hydrogen sulfide. However, some salts like aluminium carbonate exist in aqueous solution but are unstable as such; and only incomplete hydrolysis takes place for salts with strong acids, such as the halides, nitrate, and sulfate. For similar reasons, anhydrous aluminium salts cannot be made by heating their "hydrates": hydrated aluminium chloride is in fact not AlCl3·6H2O but [Al(H2O)6]Cl3, and the Al–O bonds are so strong that heating is not sufficient to break them and form Al–Cl bonds instead:[29] 2[Al(H2O)6]Cl3 heat→ Al2O3 + 6 HCl + 9 H2O All four trihalides are well known. Unlike the structures of the three heavier trihalides, aluminium fluoride (AlF3) features six-coordinate aluminium, which explains its involatility and insolubility as well as high heat of formation. Each aluminium atom is surrounded by six fluorine atoms in a distorted octahedral arrangement, with each fluorine atom being shared between the corners of two octahedra. Such {AlF6} units also exist in complex fluorides such as cryolite, Na3AlF6.[f] AlF3 melts at 1,290 °C (2,354 °F) and is made by reaction of aluminium oxide with hydrogen fluoride gas at 700 °C (1,300 °F).[39] With heavier halides, the coordination numbers are lower. The other trihalides are dimeric or polymeric with tetrahedral four-coordinate aluminium centers.[g] Aluminium trichloride (AlCl3) has a layered polymeric structure below its melting point of 192.4 °C (378 °F) but transforms on melting to Al2Cl6 dimers. At higher temperatures those increasingly dissociate into trigonal planar AlCl3 monomers similar to the structure of BCl3. Aluminium tribromide and aluminium triiodide form Al2X6 dimers in all three phases and hence do not show such significant changes of properties upon phase change.[39] These materials are prepared by treating aluminium metal with the halogen. The aluminium trihalides form many addition compounds or complexes; their Lewis acidic nature makes them useful as catalysts for the Friedel–Crafts reactions. Aluminium trichloride has major industrial uses involving this reaction, such as in the manufacture of anthraquinones and styrene; it is also often used as the precursor for many other aluminium compounds and as a reagent for converting nonmetal fluorides into the corresponding chlorides (a transhalogenation reaction).[39] Aluminium forms one stable oxide with the chemical formula Al2O3, commonly called alumina.[40] It can be found in nature in the mineral corundum, α-alumina;[41] there is also a γ-alumina phase.[6] Its crystalline form, corundum, is very hard (Mohs hardness 9), has a high melting point of 2,045 °C (3,713 °F), has very low volatility, is chemically inert, and a good electrical insulator, it is often used in abrasives (such as toothpaste), as a refractory material, and in ceramics, as well as being the starting material for the electrolytic production of aluminium metal. Sapphire and ruby are impure corundum contaminated with trace amounts of other metals.[6] The two main oxide-hydroxides, AlO(OH), are boehmite and diaspore. There are three main trihydroxides: bayerite, gibbsite, and nordstrandite, which differ in their crystalline structure (polymorphs). Many other intermediate and related structures are also known.[6] Most are produced from ores by a variety of wet processes using acid and base. Heating the hydroxides leads to formation of corundum. These materials are of central importance to the production of aluminium and are themselves extremely useful. Some mixed oxide phases are also very useful, such as spinel (MgAl2O4), Na-β-alumina (NaAl11O17), and tricalcium aluminate (Ca3Al2O6, an important mineral phase in Portland cement).[6] The only stable chalcogenides under normal conditions are aluminium sulfide (Al2S3), selenide (Al2Se3), and telluride (Al2Te3). All three are prepared by direct reaction of their elements at about 1,000 °C (1,800 °F) and quickly hydrolyze completely in water to yield aluminium hydroxide and the respective hydrogen chalcogenide. As aluminium is a small atom relative to these chalcogens, these have four-coordinate tetrahedral aluminium with various polymorphs having structures related to wurtzite, with two-thirds of the possible metal sites occupied either in an orderly (α) or random (β) fashion; the sulfide also has a γ form related to γ-alumina, and an unusual high-temperature hexagonal form where half the aluminium atoms have tetrahedral four-coordination and the other half have trigonal bipyramidal five-coordination. [42] Four pnictides – aluminium nitride (AlN), aluminium phosphide (AlP), aluminium arsenide (AlAs), and aluminium antimonide (AlSb) – are known. They are all III-V semiconductors isoelectronic to silicon and germanium, all of which but AlN have the zinc blende structure. All four can be made by high-temperature (and possibly high-pressure) direct reaction of their component elements.[42] Aluminium alloys well with most other metals (with the exception of most alkali metals and group 13 metals) and over 150 intermetallics with other metals are known. Preparation involves heating fixed metals together in certain proportion, followed by gradual cooling and annealing. Bonding in them is predominantly metallic and the crystal structure primarily depends on efficiency of packing.[43] There are few compounds with lower oxidation states. A few aluminium(I) compounds exist: AlF, AlCl, AlBr, and AlI exist in the gaseous phase when the respective trihalide is heated with aluminium, and at cryogenic temperatures.[39] A stable derivative of aluminium monoiodide is the cyclic adduct formed with triethylamine, Al4I4(NEt3)4. Al2O and Al2S also exist but are very unstable.[44] Very simple aluminium(II) compounds are invoked or observed in the reactions of Al metal with oxidants. For example, aluminium monoxide, AlO, has been detected in the gas phase after explosion[45] and in stellar absorption spectra.[46] More thoroughly investigated are compounds of the formula R4Al2 which contain an Al–Al bond and where R is a large organic ligand.[47] Organoaluminium compounds and related hydrides Main article: Organoaluminium compound Structure of trimethylaluminium, a compound that features five-coordinate carbon. A variety of compounds of empirical formula AlR3 and AlR1.5Cl1.5 exist.[48] The aluminium trialkyls and triaryls are reactive, volatile, and colorless liquids or low-melting solids. They catch fire spontaneously in air and react with water, thus necessitating precautions when handling them. They often form dimers, unlike their boron analogues, but this tendency diminishes for branched-chain alkyls (e.g. Pri, Bui, Me3CCH2); for example, triisobutylaluminium exists as an equilibrium mixture of the monomer and dimer.[49][50] These dimers, such as trimethylaluminium (Al2Me6), usually feature tetrahedral Al centers formed by dimerization with some alkyl group bridging between both aluminium atoms. They are hard acids and react readily with ligands, forming adducts. In industry, they are mostly used in alkene insertion reactions, as discovered by Karl Ziegler, most importantly in "growth reactions" that form long-chain unbranched primary alkenes and alcohols, and in the low-pressure polymerization of ethene and propene. There are also some heterocyclic and cluster organoaluminium compounds involving Al–N bonds.[49] The industrially most important aluminium hydride is lithium aluminium hydride (LiAlH4), which is used in as a reducing agent in organic chemistry. It can be produced from lithium hydride and aluminium trichloride.[51] The simplest hydride, aluminium hydride or alane, is not as important. It is a polymer with the formula (AlH3)n, in contrast to the corresponding boron hydride that is a dimer with the formula (BH3)2.[51] Natural occurrence See also: List of countries by bauxite production Space Aluminium's per-particle abundance in the Solar System is 3.15 ppm (parts per million).[52][h] It is the twelfth most abundant of all elements and third most abundant among the elements that have odd atomic numbers, after hydrogen and nitrogen.[52] The only stable isotope of aluminium, 27Al, is the eighteenth most abundant nucleus in the Universe. It is created almost entirely after fusion of carbon in massive stars that will later become Type II supernovas: this fusion creates 26Mg, which, upon capturing free protons and neutrons becomes aluminium. Some smaller quantities of 27Al are created in hydrogen burning shells of evolved stars, where 26Mg can capture free protons.[53] Essentially all aluminium now in existence is 27Al. 26Al was present in the early Solar System with abundance of 0.005% relative to 27Al but its half-life of 728,000 years is too short for any original nuclei to survive; 26Al is therefore extinct.[53] Unlike for 27Al, hydrogen burning is the primary source of 26Al, with the nuclide emerging after a nucleus of 25Mg catches a free proton. However, the trace quantities of 26Al that do exist are the most common gamma ray emitter in the interstellar gas;[53] if the original 26Al were still present, gamma ray maps of the Milky Way would be brighter.[53] Earth Bauxite, a major aluminium ore. The red-brown color is due to the presence of iron oxide minerals. Overall, the Earth is about 1.59% aluminium by mass (seventh in abundance by mass).[54] Aluminium occurs in greater proportion in the Earth's crust than in the Universe at large, because aluminium easily forms the oxide and becomes bound into rocks and stays in the Earth's crust, while less reactive metals sink to the core.[53] In the Earth's crust, aluminium is the most abundant metallic element (8.23% by mass[23]) and the third most abundant of all elements (after oxygen and silicon).[55] A large number of silicates in the Earth's crust contain aluminium.[56] In contrast, the Earth's mantle is only 2.38% aluminium by mass.[57] Aluminium also occurs in seawater at a concentration of 2 µg/kg.[23] Because of its strong affinity for oxygen, aluminium is almost never found in the elemental state; instead it is found in oxides or silicates. Feldspars, the most common group of minerals in the Earth's crust, are aluminosilicates. Aluminium also occurs in the minerals beryl, cryolite, garnet, spinel, and turquoise.[58] Impurities in Al2O3, such as chromium and iron, yield the gemstones ruby and sapphire, respectively.[59] Native aluminium metal is extremely rare and can only be found as a minor phase in low oxygen fugacity environments, such as the interiors of certain volcanoes.[60] Native aluminium has been reported in cold seeps in the northeastern continental slope of the South China Sea. It is possible that these deposits resulted from bacterial reduction of tetrahydroxoaluminate Al(OH)4−.[61] Although aluminium is a common and widespread element, not all aluminium minerals are economically viable sources of the metal. Almost all metallic aluminium is produced from the ore bauxite (AlOx(OH)3–2x). Bauxite occurs as a weathering product of low iron and silica bedrock in tropical climatic conditions.[62] In 2017, most bauxite was mined in Australia, China, Guinea, and India.[63] History Main article: History of aluminium Friedrich Wöhler, the chemist who first thoroughly described metallic elemental aluminium The history of aluminium has been shaped by usage of alum. The first written record of alum, made by Greek historian Herodotus, dates back to the 5th century BCE.[64] The ancients are known to have used alum as a dyeing mordant and for city defense.[64] After the Crusades, alum, an indispensable good in the European fabric industry,[65] was a subject of international commerce;[66] it was imported to Europe from the eastern Mediterranean until the mid-15th century.[67] The nature of alum remained unknown. Around 1530, Swiss physician Paracelsus suggested alum was a salt of an earth of alum.[68] In 1595, German doctor and chemist Andreas Libavius experimentally confirmed this.[69] In 1722, German chemist Friedrich Hoffmann announced his belief that the base of alum was a distinct earth.[70] In 1754, German chemist Andreas Sigismund Marggraf synthesized alumina by boiling clay in sulfuric acid and subsequently adding potash.[70] Attempts to produce aluminium metal date back to 1760.[71] The first successful attempt, however, was completed in 1824 by Danish physicist and chemist Hans Christian Ørsted. He reacted anhydrous aluminium chloride with potassium amalgam, yielding a lump of metal looking similar to tin.[72][73][74] He presented his results and demonstrated a sample of the new metal in 1825.[75][76] In 1827, German chemist Friedrich Wöhler repeated Ørsted's experiments but did not identify any aluminium.[77] (The reason for this inconsistency was only discovered in 1921.)[78] He conducted a similar experiment in the same year by mixing anhydrous aluminium chloride with potassium and produced a powder of aluminium.[74] In 1845, he was able to produce small pieces of the metal and described some physical properties of this metal.[78] For many years thereafter, Wöhler was credited as the discoverer of aluminium.[79] The statue of Anteros in Piccadilly Circus, London, was made in 1893 and is one of the first statues cast in aluminium. As Wöhler's method could not yield great quantities of aluminium, the metal remained rare; its cost exceeded that of gold.[77] The first industrial production of aluminium was established in 1856 by French chemist Henri Etienne Sainte-Claire Deville and companions.[80] Deville had discovered that aluminium trichloride could be reduced by sodium, which was more convenient and less expensive than potassium, which Wöhler had used.[81] Even then, aluminium was still not of great purity and produced aluminium differed in properties by sample.[82] The first industrial large-scale production method was independently developed in 1886 by French engineer Paul Héroult and American engineer Charles Martin Hall; it is now known as the Hall–Héroult process.[83] The Hall–Héroult process converts alumina into metal. Austrian chemist Carl Joseph Bayer discovered a way of purifying bauxite to yield alumina, now known as the Bayer process, in 1889.[84] Modern production of the aluminium metal is based on the Bayer and Hall–Héroult processes.[85] Prices of aluminium dropped and aluminium became widely used in jewelry, everyday items, eyeglass frames, optical instruments, tableware, and foil in the 1890s and early 20th century. Aluminium's ability to form hard yet light alloys with other metals provided the metal with many uses at the time.[86] During World War I, major governments demanded large shipments of aluminium for light strong airframes;[87] during World War II, demand by major governments for aviation was even higher.[88][89][90] By the mid-20th century, aluminium had become a part of everyday life and an essential component of housewares.[91] In 1954, production of aluminium surpassed that of copper,[i] historically second in production only to iron,[94] making it the most produced non-ferrous metal. During the mid-20th century, aluminium emerged as a civil engineering material, with building applications in both basic construction and interior finish work,[95] and increasingly being used in military engineering, for both airplanes and land armor vehicle engines.[96] Earth's first artificial satellite, launched in 1957, consisted of two separate aluminium semi-spheres joined together and all subsequent space vehicles have used aluminium to some extent.[85] The aluminium can was invented in 1956 and employed as a storage for drinks in 1958.[97] World production of aluminium since 1900 Throughout the 20th century, the production of aluminium rose rapidly: while the world production of aluminium in 1900 was 6,800 metric tons, the annual production first exceeded 100,000 metric tons in 1916; 1,000,000 tons in 1941; 10,000,000 tons in 1971.[92] In the 1970s, the increased demand for aluminium made it an exchange commodity; it entered the London Metal Exchange, the oldest industrial metal exchange in the world, in 1978.[85] The output continued to grow: the annual production of aluminium exceeded 50,000,000 metric tons in 2013.[92] The real price for aluminium declined from $14,000 per metric ton in 1900 to $2,340 in 1948 (in 1998 United States dollars).[92] Extraction and processing costs were lowered over technological progress and the scale of the economies. However, the need to exploit lower-grade poorer quality deposits and the use of fast increasing input costs (above all, energy) increased the net cost of aluminium;[98] the real price began to grow in the 1970s with the rise of energy cost.[99] Production moved from the industrialized countries to countries where production was cheaper.[100] Production costs in the late 20th century changed because of advances in technology, lower energy prices, exchange rates of the United States dollar, and alumina prices.[101] The BRIC countries' combined share in primary production and primary consumption grew substantially in the first decade of the 21st century.[102] China is accumulating an especially large share of the world's production thanks to an abundance of resources, cheap energy, and governmental stimuli;[103] it also increased its consumption share from 2% in 1972 to 40% in 2010.[104] In the United States, Western Europe, and Japan, most aluminium was consumed in transportation, engineering, construction, and packaging.[105] Etymology The names aluminium and aluminum are derived from the word alumine, an obsolete term for alumina,[j] a naturally occurring oxide of aluminium.[107] Alumine was borrowed from French, which in turn derived it from alumen, the classical Latin name for alum, the mineral from which it was collected.[108] The Latin word alumen stems from the Proto-Indo-European root *alu- meaning "bitter" or "beer".[109] 1897 American advertisement featuring the aluminum spelling Coinage British chemist Humphry Davy, who performed a number of experiments aimed to isolate the metal, is credited as the person who named the element. The first name proposed for the metal to be isolated from alum was alumium, which Davy suggested in an 1808 article on his electrochemical research, published in Philosophical Transactions of the Royal Society.[110] It appeared that the name was coined from the English word alum and the Latin suffix -ium; however, it was customary at the time that the elements should have names originating in the Latin language, and as such, this name was not adopted universally. This name was criticized by contemporary chemists from France, Germany, and Sweden, who insisted the metal should be named for the oxide, alumina, from which it would be isolated.[111] The English word name alum does not directly reference the Latin language, whereas alumine/alumina easily references the Latin word alumen (upon declension, alumen changes to alumin-). One example was a writing in French by Swedish chemist Jöns Jacob Berzelius titled Essai sur la Nomenclature chimique, published in July 1811; in this essay, among other things, Berzelius used the name aluminium for the element that would be synthesized from alum.[112][k] (Another article in the same journal issue also refers to the metal whose oxide forms the basis of sapphire as to aluminium.)[114] A January 1811 summary of one of Davy's lectures at the Royal Society mentioned the name aluminium as a possibility.[115] The following year, Davy published a chemistry textbook in which he used the spelling aluminum.[116] Both spellings have coexisted since; however, their usage has split by region: aluminum is the primary spelling in the United States and Canada while aluminium is in the rest of the English-speaking world.[117] Spelling In 1812, British scientist Thomas Young[118] wrote an anonymous review of Davy's book, in which he proposed the name aluminium instead of aluminum, which he felt had a "less classical sound".[119] This name did catch on: while the -um spelling was occasionally used in Britain, the American scientific language used -ium from the start.[120] Most scientists used -ium throughout the world in the 19th century,[117] and it was entrenched in many other European languages, such as French, German, or Dutch.[l] In 1828, American lexicographer Noah Webster used exclusively the aluminum spelling in his American Dictionary of the English Language.[121] In the 1830s, the -um spelling started to gain usage in the United States; by the 1860s, it had become the more common spelling there outside science.[120] In 1892, Hall used the -um spelling in his advertising handbill for his new electrolytic method of producing the metal, despite his constant use of the -ium spelling in all the patents he filed between 1886 and 1903. It remains unknown whether this spelling was introduced by mistake or intentionally; however, Hall preferred aluminum since its introduction because it resembled platinum, the name of a prestigious metal.[122] By 1890, both spellings had been common in the U.S. overall, the -ium spelling being slightly more common; by 1895, the situation had reversed; by 1900, aluminum had become twice as common as aluminium; during the following decade, the -um spelling dominated American usage. In 1925, the American Chemical Society adopted this spelling.[117] The International Union of Pure and Applied Chemistry (IUPAC) adopted aluminium as the standard international name for the element in 1990.[123] In 1993, they recognized aluminum as an acceptable variant;[123] the most recent 2005 edition of the IUPAC nomenclature of inorganic chemistry acknowledges this spelling as well.[124] IUPAC official publications use the -ium spelling as primary but list both where appropriate.[m] Production and refinement See also: List of countries by primary aluminium production The production of aluminium starts with the extraction of bauxite rock from the ground. The bauxite is processed and transformed using the Bayer process into alumina, which is then processed using the Hall–Héroult process, resulting in the final aluminium metal. Aluminium production is highly energy-consuming, and so the producers tend to locate smelters in places where electric power is both plentiful and inexpensive.[126] As of 2019, the world's largest smelters of aluminium are located in China, India, Russia, Canada, and the United Arab Emirates,[127] while China is by far the top producer of aluminium with a world share of fifty-five percent. According to the International Resource Panel's Metal Stocks in Society report, the global per capita stock of aluminium in use in society (i.e. in cars, buildings, electronics, etc.) is 80 kg (180 lb). Much of this is in more-developed countries (350–500 kg (770–1,100 lb) per capita) rather than less-developed countries (35 kg (77 lb) per capita).[128] Bayer process Main article: Bayer process See also: List of countries by bauxite production Bauxite is converted to alumina by the Bayer process. Bauxite is blended for uniform composition and then is ground. The resulting slurry is mixed with a hot solution of sodium hydroxide; the mixture is then treated in a digester vessel at a pressure well above atmospheric, dissolving the aluminium hydroxide in bauxite while converting impurities into relatively insoluble compounds:[129] Al(OH)3 + Na+ + OH− → Na+ + [Al(OH)4]− After this reaction, the slurry is at a temperature above its atmospheric boiling point. It is cooled by removing steam as pressure is reduced. The bauxite residue is separated from the solution and discarded. The solution, free of solids, is seeded with small crystals of aluminium hydroxide; this causes decomposition of the [Al(OH)4]− ions to aluminium hydroxide. After about half of aluminium has precipitated, the mixture is sent to classifiers. Small crystals of aluminium hydroxide are collected to serve as seeding agents; coarse particles are converted to alumina by heating; the excess solution is removed by evaporation, (if needed) purified, and recycled.[129] Hall–Héroult process Extrusion billets of aluminium Main articles: Hall–Héroult process and Aluminium smelting See also: List of countries by aluminium oxide production The conversion of alumina to aluminium metal is achieved by the Hall–Héroult process. In this energy-intensive process, a solution of alumina in a molten (950 and 980 °C (1,740 and 1,800 °F)) mixture of cryolite (Na3AlF6) with calcium fluoride is electrolyzed to produce metallic aluminium. The liquid aluminium metal sinks to the bottom of the solution and is tapped off, and usually cast into large blocks called aluminium billets for further processing.[36] Anodes of the electrolysis cell are made of carbon—the most resistant material against fluoride corrosion—and either bake at the process or are prebaked. The former, also called Söderberg anodes, are less power-efficient and fumes released during baking are costly to collect, which is why they are being replaced by prebaked anodes even though they save the power, energy, and labor to prebake the cathodes. Carbon for anodes should be preferably pure so that neither aluminium nor the electrolyte is contaminated with ash. Despite carbon's resistivity against corrosion, it is still consumed at a rate of 0.4–0.5 kg per each kilogram of produced aluminium. Cathodes are made of anthracite; high purity for them is not required because impurities leach only very slowly. The cathode is consumed at a rate of 0.02–0.04 kg per each kilogram of produced aluminium. A cell is usually terminated after 2–6 years following a failure of the cathode.[36] The Hall–Heroult process produces aluminium with a purity of above 99%. Further purification can be done by the Hoopes process. This process involves the electrolysis of molten aluminium with a sodium, barium, and aluminium fluoride electrolyte. The resulting aluminium has a purity of 99.99%.[36][130] Electric power represents about 20 to 40% of the cost of producing aluminium, depending on the location of the smelter. Aluminium production consumes roughly 5% of electricity generated in the United States.[123] Because of this, alternatives to the Hall–Héroult process have been researched, but none has turned out to be economically feasible.[36] Recycling Common bins for recyclable waste along with a bin for unrecyclable waste. The bin with a yellow top is labeled "aluminum". Rhodes, Greece. Main article: Aluminium recycling Recovery of the metal through recycling has become an important task of the aluminium industry. Recycling was a low-profile activity until the late 1960s, when the growing use of aluminium beverage cans brought it to public awareness.[131] Recycling involves melting the scrap, a process that requires only 5% of the energy used to produce aluminium from ore, though a significant part (up to 15% of the input material) is lost as dross (ash-like oxide).[132] An aluminium stack melter produces significantly less dross, with values reported below 1%.[133] White dross from primary aluminium production and from secondary recycling operations still contains useful quantities of aluminium that can be extracted industrially. The process produces aluminium billets, together with a highly complex waste material. This waste is difficult to manage. It reacts with water, releasing a mixture of gases (including, among others, hydrogen, acetylene, and ammonia), which spontaneously ignites on contact with air;[134] contact with damp air results in the release of copious quantities of ammonia gas. Despite these difficulties, the waste is used as a filler in asphalt and concrete.[135] Applications Aluminium-bodied Austin A40 Sports (c. 1951) Metal See also: Aluminium alloy

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