Calcium

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Calcium is a chemical element with symbol **Ca** and atomic number 20. Calcium is a soft gray Group 2 alkaline earth metal, fifth-most-abundant element by mass in the Earth's crust. The ion Ca^{2+} is also the fifth-most-abundant dissolved ion in seawater by both molarity and mass, after sodium, chloride, magnesium, and sulfate.[5] Free calcium metal is too reactive to occur in nature. Calcium is produced in supernova nucleosynthesis.

Calcium is essential for living organisms, particularly in cell physiology where movement of the calcium ion into and out of the cytoplasm functions as a signal for many cellular processes. As a major material used in mineralization of bone, teeth and shells, calcium is the most abundant metal by mass in many animals.

Contents

Calcium, $_{20}$Ca

Spectral lines of calcium

General properties

Name, symbol	calcium, Ca
Pronunciation	/ˈkælsiəm/ ***KAL-see-əm***
Appearance	dull gray, silver; with a pale yellow tint[1]

Calcium in the periodic table

Atomic number (*Z*)	20
Group, block	group 2 (alkaline earth metals), s-block
Period	period 4
Element category	alkaline earth metal
Standard atomic weight (±) (A_r)	40.078(4)[2]
Electron configuration	[Ar] $4s^2$
per shell	2, 8, 8, 2

Physical properties

Phase	solid
Melting point	1115 K (842 °C, 1548 °F)
Boiling point	1757 K (1484 °C, 2703 °F)

Uses

Calcium metal find use as a deoxidizer, desulfurizer, and decarbonizer for production of some ferrous and nonferrous alloys. In steelmaking and production of iron, Ca reacts with oxygen and sulfur-containing impurities.[6] Almost all applications are associated with calcium compounds and salts.[7]

- Calcium carbonate ($CaCO_3$) is used in manufacturing cement and mortar, lime, limestone (usually used in the steel industry) and aids in production in the glass industry. It also has chemical and optical uses as mineral specimens in toothpastes, for example.
- Calcium hydroxide solution ($Ca(OH)_2$) (also known as limewater) is used to detect the presence of carbon dioxide in a gas sample bubbled through a solution. The solution turns cloudy where CO_2 is present.
- Calcium arsenate ($Ca_3(AsO_4)_2$) is used in insecticides.
- Calcium carbide (CaC_2) is used to make acetylene gas (for use in acetylene torches for welding) and various plastics.
- Calcium chloride ($CaCl_2$) is used in ice removal and dust control on dirt roads, as a conditioner for concrete, as an additive in canned tomatoes, and to provide body for automobile tires.
- Calcium citrate ($Ca_3(C_6H_5O_7)_2$) is used as a food preservative.
- Calcium cyclamate ($Ca(C_6H_{11}NHSO_3)_2$) is used as a sweetening agent in several countries. In the United States, it has been outlawed as a suspected carcinogen.[8]
- Calcium gluconate ($Ca(C_6H_{11}O_7)_2$) is used as a food additive and in vitamin pills.
- Calcium hypochlorite ($Ca(OCl)_2$) is used as a swimming pool disinfectant, as a bleaching agent, as an ingredient in deodorant, and in algaecide and fungicide.
- Calcium permanganate ($Ca(MnO_4)_2$) is used in liquid rocket propellant, textile production, as a water sterilizing agent and in dental procedures.
- Calcium phosphate ($Ca_3(PO_4)_2$) is used as a supplement for animal feed, fertilizer, in commercial production for dough and yeast products, in the manufacture of glass, and in dental products.

Density near r.t.	1.55 g/cm^3
when liquid, at m.p.	1.378 g/cm^3
Heat of fusion	8.54 kJ/mol
Heat of vaporization	154.7 kJ/mol
Molar heat capacity	25.929 J/(mol·K)

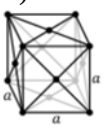
Vapor pressure

P (Pa)	**1**	**10**	**100**	**1 k**	**10 k**	**100 k**
at T (K)	864	956	1071	1227	1443	1755

Atomic properties

Oxidation states	**+2**, +1[3] (a strongly basic oxide)
Electronegativity	Pauling scale: 1.00
Ionization energies	1st: 589.8 kJ/mol 2nd: 1145.4 kJ/mol 3rd: 4912.4 kJ/mol (more)
Atomic radius	empirical: 197 pm
Covalent radius	176±10 pm
Van der Waals radius	231 pm

Miscellanea

Crystal structure	face-centered cubic (fcc) 
Speed of sound thin rod	3810 m/s (at 20 °C)
Thermal expansion	22.3 µm/(m·K) (at 25 °C)
Thermal conductivity	201 W/(m·K)
Electrical resistivity	33.6 nΩ·m (at 20 °C)
Magnetic ordering	diamagnetic
Magnetic susceptibility (χ_{mol})	+40.0·10^{-6} cm^3/mol[4]
Young's modulus	20 GPa
Shear modulus	7.4 GPa
Bulk modulus	17 GPa
Poisson ratio	0.31
Mohs hardness	1.75

- Calcium phosphide (Ca_3P_2) is used in fireworks, rodenticide, torpedoes, and flares.
- Calcium stearate ($Ca(C_{18}H_{35}O_2)_2$) is used in the manufacture of wax crayons, cements, certain kinds of plastics, and cosmetics, as a food additive, in the production of water resistant materials, and in the production of paints.
- Calcium sulfate ($CaSO_4 \cdot 2H_2O$) (gypsum) is used as common blackboard chalk, as well as, in its hemihydrate form, Plaster of Paris.
- Calcium tungstate ($CaWO_4$) is used in luminous paints, fluorescent lights, and in X-ray studies.
- Hydroxylapatite ($Ca_5(PO_4)_3(OH)$, but is usually written $Ca_{10}(PO_4)_6(OH)_2$) makes up seventy percent of bone. Carbonated-calcium deficient hydroxylapatite is the main mineral of which dental enamel and dentin are comprised.

Brinell hardness	170–416 MPa
CAS Number	7440-70-2
	History
Discovery and first isolation	Humphry Davy (1808)

Most stable isotopes of calcium

iso	NA	half-life	DM	DE (MeV)	DP
^{40}Ca	96.941%	is stable with 20 neutrons			
^{41}Ca	trace	1.03×10^5 y	ε	–	^{41}K
^{42}Ca	0.647%	is stable with 22 neutrons			
^{43}Ca	0.135%	is stable with 23 neutrons			
^{44}Ca	2.086%	is stable with 24 neutrons			
^{45}Ca	syn	162.7 d	β^-	0.258	^{45}Sc
^{46}Ca	0.004%	is stable with 26 neutrons			
^{47}Ca	syn	4.536 d	β^-	0.694, 1.99	^{47}Sc
			γ	1.297	–
^{48}Ca	0.187%	4.3×10^{19} y	$\beta^-\beta^-$	4.274	^{48}Ti

Food

In solution, the calcium ion varies remarkably to the human taste, being reported as mildly salty, sour, "mineral-like", or even "soothing." It is apparent that many animals can taste, or develop a taste, for calcium, and use this sense to detect the mineral in salt licks or other sources.[9] In human nutrition, soluble calcium salts may be added to tart juices without much effect to the average palate.

Recommended adequate intake by the IOM for calcium:[10][11]

Age	Calcium (mg/day)
0–6 months	200
7–12 months	260
1–3 years	700
4–8 years	1000
9–18 years	1300
19–50 years	1000
51–70 years (male)	1000
51–70 years (female)	1200
71+ years	1200

Calcium is an important component of a healthy diet and a mineral necessary for life. The National Osteoporosis Foundation states, "Calcium plays an important role in building stronger, denser bones early in life and keeping bones strong and healthy later in life." Approximately 99 percent of the calcium in the human body is in the bones and teeth.[12] The rest of the calcium in the body has other important uses, such as some exocytosis, especially neurotransmitter release, and muscle contraction. Intracellular calcium overload may lead some kind of cells to oxidative stress and apoptosis, and produces several diseases.[13] In the electrical conduction system of the heart, calcium replaces sodium as the mineral that depolarizes the cell, proliferating the action potential. In cardiac muscle, sodium influx commences an action potential, but during potassium efflux, the cardiac myocyte experiences calcium influx, prolonging the action potential and creating a plateau phase of dynamic equilibrium. Long-term calcium deficiency can lead to rickets and poor blood clotting; in menopausal women, deficiency can lead to osteoporosis, a condition in which the bone deteriorates and fractures more readily. While a lifelong deficit can affect bone and tooth formation, over-retention can cause hypercalcemia (elevated levels of calcium in the blood), impaired kidney function, and decreased absorption of

other minerals.[14][15] Several sources suggest a correlation between high calcium intake (2000 mg per day, or twice the U.S. recommended daily allowance, equivalent to six or more glasses of milk per day) and prostate cancer.[16] Vitamin D is needed to absorb calcium.

Dairy products, such as milk and cheese, are a well-known source of calcium. Some individuals are allergic to dairy products and even more people, in particular those of non Indo-European descent, are lactose-intolerant, leaving them unable to consume non-fermented dairy products in quantities larger than about half a liter per serving. Others, such as ethical and/or health vegans, voluntarily avoid dairy products.

Many good vegetable sources of calcium exist:

- seaweeds such as kelp, wakame, and hijiki;
- Nuts and seeds such almonds, hazelnuts, sesame, and pistachio;
- Blackstrap molasses;
- beans (especially soy beans);
- figs, quinoa, okra; rutabaga; broccoli; dandelion leaves; and kale.
- Several foods and drinks, such as orange juice, soy milk, tofu, breakfast cereals, and breads are often fortified with calcium.[17]

Numerous vegetables, notably spinach, chard, and rhubarb have a high calcium content, but they may also contain varying amounts of oxalic acid that binds calcium and reduces its absorption. The same problem may affect the absorption of calcium from amaranth, collard greens, and chicory greens. This process may also be related to the generation of calcium oxalate.

An overlooked source of calcium is eggshell, which can be ground into a powder and mixed into food or a glass of water.[18][19][20]

The calcium content of most foods can be found in the USDA National Nutrient Database.[21]

Dietary supplements

500 milligram calcium supplements made from calcium carbonate

Calcium supplements are used to prevent and to treat calcium deficiencies. However, it has been found that the taking of calcium supplements by people with a history of stroke or with white matter lesions greatly increased their chances of developing dementia.[22][23] The Office of Dietary Supplements (National Institutes of Health) recommends that no more than 600 mg of supplement should be taken at a time because the percent of calcium absorbed decreases as the amount of calcium in the supplement increases.[10] It is therefore recommended to spread doses throughout the day.[24] Recommended daily calcium intake for adults ranges from 1000 to 1300 mg.[24] Calcium supplements may have side effects such as bloating and constipation in some people. It is suggested that taking the supplements with food may aid in nullifying these side effects.[24]

Vitamin D is added to some calcium supplements. Proper vitamin D status is important because vitamin D is converted to a hormone in the body, which then induces the synthesis of intestinal proteins responsible for calcium absorption.[25]

- The absorption of calcium from most food and commonly used dietary supplements is very similar.[26] This is contrary to what many calcium supplement manufacturers claim in their promotional materials.
- Milk is an excellent source of dietary calcium for those whose bodies tolerate it because it has a high concentration of calcium and the calcium in milk is excellently absorbed.[26]
- Soymilk and other vegetable milks are usually sold with calcium added so that their calcium concentration is as high as in milk.
- Also different kind of juices boosted with calcium are widely available.
- Calcium carbonate is the most common and least expensive calcium supplement. It should be taken with food, and depends on low pH levels (acidic) for proper absorption in the intestine.[27] Some studies suggests that the absorption of calcium from calcium carbonate is similar to the absorption of calcium from milk.[28][29]
- Antacids frequently contain calcium carbonate, and are a commonly used, inexpensive calcium supplement.
- Coral calcium is a salt of calcium derived from fossilized coral reefs. Coral calcium is composed of calcium carbonate and trace minerals.
- Calcium citrate can be taken without food and is the supplement of choice for individuals with achlorhydria or who are taking histamine-2 blockers or proton-pump inhibitors.[30] Calcium citrate is about 21% elemental calcium. 1000 mg will provide 210 mg of calcium. It is more expensive than calcium carbonate and more of it must be taken to get the same amount of calcium.
- Calcium phosphate costs more than calcium carbonate, but less than calcium citrate. Microcrystalline Hydroxyapatite (MH) is one of several forms of calcium phosphate used as a dietary supplement. Hydroxyapatite is about 40% calcium.
- Calcium lactate has similar absorption as calcium carbonate,[31] but is more expensive. Calcium lactate and calcium gluconate are less concentrated forms of calcium and are not practical oral supplements.[30]

Bone health

Dietary calcium supplement is generally not necessary for maintaining bone mineral density, and carries risks that outweigh any benefits.[32] Calcium intake is not significantly associated with hip fracture risk in either men nor women.[33] The U.S. Preventive Service Task Force therefore recommends against a daily supplement of calcium or Vitamin D.[34]

Cardiovascular impact

A study investigating the effects of personal calcium supplement on cardiovascular risk in the Women's Health Initiative Calcium/Vitamin D Supplementation Study (WHI CaD Study) found a modestly increased risk of cardiovascular events, particularly myocardial infarction in postmenopausal women. A broad recommendation of calcium/vitamin D supplements is therefore not warranted.[35] In contrast, the authors of a 2013 literature review concluded that the benefits of calcium supplementation in some studies, such as on bone health, appear to outweigh any risk calcium supplementation may pose to the cardiovascular health.[36]

Cancer

Overall, there is no clear evidence of the effect of calcium supplements in cancer prevention: some studies suggest it might decrease the risk, but others suggest it might increase the risk. Consequently, the National Cancer Institute, part of the National Institutes of Health, does not recommend the use of calcium supplements

for this purpose.[37]

There is weak evidence calcium supplementation might have a preventative effect against developing colorectal adenomatous polyps, but the evidence is not sufficient to recommend such supplementation.[38]

Side effects

Compared with other metals, the calcium ion and most calcium compounds have low toxicity. This is not surprising, given the very high natural abundance of calcium compounds in the environment and in organisms.

Calcium poses few serious environmental problems.

High calcium intakes or high calcium absorption were previously thought to contribute to the development of kidney stones. However, a high calcium intake has been associated with a lower risk for kidney stones in more recent research.[39][40][41]

Acute calcium poisoning is rare, and difficult to achieve without administering calcium intravenously. For example, the oral median lethal dose (LD^{50}) for rats for calcium carbonate and calcium chloride are 6.45[42] and 1.4 g/kg,[43] respectively.

Calcium metal is hazardous because of its sometimes-violent reactions with water and acids. Calcium metal is found in some drain cleaners, where it functions to generate heat and calcium hydroxide that saponifies the fats and liquefies the proteins (e.g., hair) that block drains. When swallowed, calcium metal has the same effect on the mouth, esophagus, and stomach, and can be fatal.[44]

Excessive consumption of calcium carbonate antacids/dietary supplements (such as Tums) over a period of weeks or months can cause milk-alkali syndrome, with symptoms ranging from hypercalcemia to potentially fatal renal failure. What constitutes “excessive” consumption is not well known and, it is presumed, varies a great deal from person to person. Persons consuming more than 10 grams/day of $CaCO_3$ (=4 g Ca) are at risk of developing milk-alkali syndrome,[45] but the condition has been reported in at least one person consuming only 2.5 grams/day of $CaCO_3$ (=1 g Ca), an amount usually considered moderate and safe.[46]

Oral calcium supplements diminish the absorption of thyroxine when taken within four to six hours of each other.[47] Thus, people taking both calcium and thyroxine run the risk of inadequate thyroid hormone replacement and thence hypothyroidism if they take them simultaneously or near-simultaneously.[48]

Although some studies have suggested that excessive intake of calcium in the diet or as supplements could be associated with increased cardiovascular mortality,[49][50] other studies found no risk,[51] leading a review to conclude that any risk could only be ascertained with specific further research.[52]

Notable characteristics

Calcium is reactive and relatively soft for a metal. Although harder than lead, it can be cut with a knife with difficulty. It is a silvery metallic element that can be extracted by electrolysis from a fused salt like calcium chloride.[53] When exposed to the air, it rapidly forms a gray-white coating of calcium oxide and calcium nitride. In bulk form (typically as chips or "turnings"), the metal is somewhat difficult to ignite, more difficult even than magnesium chips; but, when lit, the metal burns in air with a brilliant high-intensity orange-red light.

Flame test. Brick-red color originates from calcium.

Calcium metal reacts with water, producing hydrogen gas at a moderate rate without generating much heat, making it useful for generating hydrogen.[54] In powdered form, however, the reaction with water is extremely rapid, as the increased surface area of the powder accelerates the reaction. Part of the reason for the slowness of the calcium–water reaction is a partial passivation (chemically protective coating) of insoluble white calcium hydroxide; in acidic solutions, where this compound is more soluble, calcium reacts vigorously.

With a density of 1.54 g/cm^3,[55] calcium is the lightest of the alkaline earth metals; magnesium (specific gravity 1.74) and beryllium (1.84) are denser though lighter in atomic mass. From strontium onward, the alkali earth metals become denser with increasing atomic mass. Calcium has two allotropes.[56]

Calcium metal has a higher electrical resistivity than copper or aluminium, yet weight-for-weight, due to its much lower density, it is a better conductor than either. Its use as such in terrestrial applications is usually limited by its high reactivity with air; however, it has potential for use as wiring in off-world applications.[57]

Calcium is the fifth-most-abundant element by mass in the human body, where it is an important cellular ionic messenger with many functions. Calcium also serves as a structural element in bone. It is the relatively high atomic number of calcium that causes bone to be radio-opaque. Of the human body's solid components after drying and burning of organics (as for example, after cremation), about a third of the total "mineral" mass remaining is the approximately one kilogram of calcium that composes the average skeleton (the remainder being mostly phosphorus and oxygen).

H and K lines

Visible spectra of many stars, including the Sun, exhibit strong emission lines of singly ionized calcium. Prominent among these are the H-line at 3968.5 Å and the K line at 3933.7 Å of singly ionized calcium, or Ca II. In the Sun or other stars with low temperatures, the prominence of the H and K lines in the visible spectra can be an indication of strong magnetic activity in the chromosphere. Periodic variations of these active regions can indicate the rotation periods of these stars.[58]

Compounds

Calcium chemistry is almost exclusively that of Ca^{2+} salts.[59] Ca^{2+} is a "hard cation", that is, it characteristically favors oxide ligands. Hence the abundance of carbonates, nitrates, phosphates, and sulfates in the mineral kingdom. Many of these species crystallize with water. Because it is generally nontoxic and abundant, calcium is found in many foods and useful materials. Most calcium salts are colorless. As with magnesium salts and other alkaline earth metal salts, the halides are soluble in water.

Combined with phosphate, calcium forms hydroxylapatite ($Ca_5(PO_4)_3(OH)$), the mineral portion of animal bones, teeth, and some corals.[60] Large-scale chemical processes are involved in the conversion of calcium phosphate minerals into fertilizer.

Calcium is the main problematic ion in hard water: it forms insoluble deposits of calcium carbonate that are problematic in plumbing. It also reacts with soap to form soap scum. Calcium carbonate occurs naturally as limestone and chalk. When water percolates through limestone or other calcium-containing rocks, it partially dissolves the rock. The slow re-precipitation of minerals derived from dissolved calcium leads to formation of stalactites and stalagmites.

When heated above 825 °C, calcium carbonate converts calcium oxide (CaO), also known as quicklime:

$$CaCO_3 \rightarrow CaO + CO_2$$

When added to water, quicklime vigorously reacts (hence its name) to form calcium hydroxide:

$$CaO + H_2O \rightarrow Ca(OH)_2$$

Also known as slaked lime, $Ca(OH)_2$ is an inexpensive base material used throughout the chemical industry. When mixed with sand, it hardens into a mortar and is turned into plaster by carbon dioxide uptake. Mixed with other compounds, lime forms an important part of Portland cement.

Combined with sulfate, calcium forms the mineral gypsum. When heated to about 300 °F (150 °C), it undergoes partial dehydration:[61]

$$CaSO_4 \cdot 2H_2O \rightarrow CaSO_4 \cdot 0.5H_2O + 1.5\ H_2O$$

The resulting powder, when mixed with water, forms a stiff but workable paste that hardens to give Plaster of Paris.

Organocalcium compounds, those containing Ca-C bonds are known, but generally of specialized interest in the research laboratory. One major exception is calcium carbide, which arises from heating calcium compounds with coal or other carbon-rich reducing agents.

$$CaO + 3\ C \rightarrow CaC_2 + CO$$

It was historically important precursor to acetylene.

$$CaC_2 + 2\ H_2O \rightarrow C_2H_2 + Ca(OH)_2$$

Other important calcium compounds are calcium nitrate used in fertilizers, calcium chloride used as for deicing roads, calcium cyanamide, and calcium hypochlorite, used for bleaching.

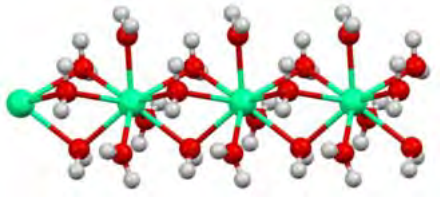
Focusing on chemical structure, Ca^{2+} is a relatively large ion that tends to adopt a high coordination number. In CaF_2, the mineral fluorite, each Ca^{2+} ion is surrounded by eight F^- ligands.

Isotopes

Calcium has five stable isotopes (^{40}Ca, ^{42}Ca, ^{43}Ca, ^{44}Ca and ^{46}Ca), plus one more (^{48}Ca) that has such a long half-life, it can also be considered stable for all practical purposes. The 20% range in relative mass among naturally occurring calcium isotopes is greater than for any element other than hydrogen and helium. Calcium also has a cosmogenic isotope, radioactive ^{41}Ca, which has a half-life of 103,000 years. Unlike cosmogenic isotopes produced in the atmosphere, ^{41}Ca is produced by neutron activation of ^{40}Ca, primarily in the top

metre of the soil column, where the cosmogenic neutron flux is still sufficiently strong. ^{41}Ca has received much attention in stellar studies because it decays to ^{41}K, a critical indicator of solar-system anomalies.

Ninety-seven percent of naturally occurring calcium is in the form of ^{40}Ca, one of the daughter products of ^{40}K decay, along with ^{40}Ar. While K–Ar dating has been used extensively in the geological sciences, the prevalence of ^{40}Ca in nature has impeded its use in dating. Techniques using mass spectrometry and a double spike isotope dilution have been used for K-Ca age dating.



Structure of the polymeric $[Ca(H_2O)_6]^{2+}$ center in hydrated calcium chloride, illustrating the high coordination number typical for calcium complexes.

^{40}Ca has a nucleus of 20 protons and 20 neutrons and is the heaviest stable isotope of any element that has equal numbers of protons and neutrons. In supernova explosions, calcium is formed from the reaction of carbon with various numbers of alpha particles (helium nuclei), until the most common calcium isotope (containing 10 helium nuclei) has been synthesized.

Isotope fractionation

As with the isotopes of other elements, a variety of processes fractionate, or alter the relative abundance of, calcium isotopes.[62] The best studied of these processes is the mass-dependent fractionation of calcium isotopes that accompanies the precipitation of calcium minerals, such as calcite, aragonite and apatite, from solution. Isotopically light calcium is preferentially incorporated into minerals, leaving the solution from which the mineral precipitated enriched in isotopically heavy calcium. At room temperature the magnitude of this fractionation is roughly 0.25‰ (0.025%) per atomic mass unit (AMU). Mass-dependent differences in calcium isotope composition conventionally are expressed by the ratio of two isotopes (usually $^{44}Ca/^{40}Ca$) in a sample compared to the same ratio in a standard reference material. $^{44}Ca/^{40}Ca$ varies by about 1% among common earth materials.[63]

Calcium isotope fractionation during mineral formation has led to several applications of calcium isotopes. In particular, the 1997 observation by Skulan and DePaolo[64] that calcium minerals are isotopically lighter than the solutions from which the minerals precipitate is the basis of analogous applications in medicine and in paleooceanography. In animals with skeletons mineralized with calcium, the calcium isotopic composition of soft tissues reflects the relative rate of formation and dissolution of skeletal mineral. In humans, changes in the calcium isotopic composition of urine have been shown to be related to changes in bone mineral balance. When the rate of bone formation exceeds the rate of bone resorption, the ration $^{44}Ca/^{40}Ca$ in soft tissue rises. Soft tissue $^{44}Ca/^{40}Ca$ falls when bone resorption exceeds bone formation. Because of this relationship, calcium isotopic measurements of urine or blood may be useful in the early detection of metabolic bone diseases like osteoporosis.[65]

A similar system exists in the ocean, where $^{44}Ca/^{40}Ca$ in seawater tends to rise when the rate of removal of Ca^{2+} from seawater by mineral precipitation exceeds the input of new calcium into the ocean, and fall when calcium input exceeds mineral precipitation. It follows that rising $^{44}Ca/^{40}Ca$ corresponds to falling seawater Ca^{2+} concentration, and falling $^{44}Ca/^{40}Ca$ corresponds to rising seawater Ca^{2+} concentration. In 1997 Skulan and DePaolo presented the first evidence of change in seawater $^{44}Ca/^{40}Ca$ over geologic time, along with a theoretical explanation of these changes. More recent papers have confirmed this observation, demonstrating that seawater Ca^{2+} concentration is not constant, and that the ocean probably never is in “steady state” with

respect to its calcium input and output.[66][67] This has important climatological implications, as the marine calcium cycle is closely tied to the carbon cycle (see below).

Geochemical cycling

Calcium provides an important link between tectonics, climate, and the carbon cycle. In the simplest terms, uplift of mountains exposes calcium-bearing rocks to chemical weathering and releases Ca^{2+} into surface water. This Ca^{2+} eventually is transported to the ocean where it reacts with dissolved CO_2 to form limestone. Some of this limestone settles to the sea floor where it is incorporated into new rocks. Dissolved CO_2, along with carbonate and bicarbonate ions, are termed "dissolved inorganic carbon" (DIC).

The actual reaction is more complicated and involves the bicarbonate ion (HCO_3^-) that forms when CO_2 reacts with water at seawater pH:

$$Ca^{2+} + 2HCO_3^- \rightarrow CaCO_3 \text{ (limestone)} + CO_2 + H_2O$$

Note that at seawater pH, most of the CO_2 is immediately converted back into HCO_3^-. The reaction results in a net transport of one molecule of CO_2 from the ocean/atmosphere into the lithosphere.[68]

Travertine terraces Pamukkale, Turkey

The result is that each Ca^{2+} ion released by chemical weathering ultimately removes one CO_2 molecule from the surficial system (atmosphere, ocean, soils and living organisms), storing it in carbonate rocks where it is likely to stay for hundreds of millions of years. The weathering of calcium from rocks thus scrubs CO_2 from the ocean and atmosphere, exerting a strong long-term effect on climate.[69] Analogous cycles involving magnesium, and to a much smaller extent strontium and barium, have the same effect.

As the weathering of limestone ($CaCO_3$) liberates equimolar amounts of Ca^{2+} and CO_2, it has no net effect on the CO_2 content of the atmosphere and ocean. The weathering of silicate rocks like granite, on the other hand, is a net CO_2 sink because it produces abundant Ca^{2+} but very little CO_2.

History

Lime as building material was used since prehistoric times going as far back as 7000 to 14000 BC.[70] Significant statues made from lime plaster date back into the 7 millennia BC were found in 'Ain Ghazal.[71] The first dated lime kiln dates back to 2500 BC and was found in Khafajah mesopotamia.[72][73] Calcium (from Latin *calx*, genitive *calcis*, meaning "lime")[74] was known as early as the first century when the Ancient Romans prepared lime as calcium oxide. Literature dating back to 975 AD notes that plaster of paris (calcium sulfate), is useful for setting broken bones. It was not isolated until 1808 in England when Sir Humphry Davy electrolyzed a mixture of lime and mercuric oxide.[75] Calcium metal was not available in large scale until the beginning of the 20th century.

Occurrence

Calcium is not naturally found in its elemental state. Calcium occurs most commonly in sedimentary rocks in the minerals calcite, dolomite, and gypsum. It also occurs in igneous and metamorphic rocks chiefly in the silicate minerals: plagioclases, amphiboles, pyroxenes, and garnets.


'Ain Ghazal figure

Cell signaling

The release of calcium ions from the sarcoplasmic reticulum into the cytoplasm is an essential intracellular signal, important in many cellular functions and processes, including muscle contraction, neuronal transmission as in an excitatory synapse, cellular motility (including the movement of flagella and cilia), fertilisation, cell growth or proliferation, learning, memory (as with synaptic plasticity), and secretion of saliva.[76] Calcium signalling can be studied by loading a cell's cytoplasm with a calcium-sensitive fluorescent dye such as Fura-2.[77][78] Many of these dyes were developed by Roger Y. Tsien.[79]

See also

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External links

- WebElements.com – Calcium (http://www.webelements.com/webelements/elements/text/Ca/index.html)
- Calcium (http://www.periodicvideos.com/videos/020.htm) at *The Periodic Table of Videos* (University of Nottingham)
- UK Food Standards Agency: Calcium (https://web.archive.org/web/20101007105207/http://www.eatwell.gov.uk/healthydiet/nutritionessentials/vitaminsandminerals/calcium/)
- Nutrition fact sheet from the National Institutes of Health (http://ods.od.nih.gov/factsheets/calcium.asp)

Categories: Alkaline earth metals | Calcium | Chemical elements | Dietary minerals | Dietary supplements | Reducing agents | Sodium channel blockers | World Health Organization essential medicines

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