Calcium is the 20th element in the periodic table. It is a group 2 metal, also known as an alkaline-earth metal, and no populated d-orbital electrons. Calcium is the fifth most abundant element by mass (3.4%) in both the Earth's crust and in seawater. All living organisms require calcium for survival. Calcium is a silver-gray metal which takes its name from the Latin word calx, which means lime. It is the fifth most abundant element in the earth's crust and is widely distributed as limestone (CaCO3), quicklime (CaO) and calcium fluoride.

### General Properties of Calcium

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Ca</th>
</tr>
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<tbody>
<tr>
<td>Color</td>
<td>dull gray or silver</td>
</tr>
<tr>
<td>Atomic Number</td>
<td>20</td>
</tr>
<tr>
<td>Category</td>
<td>alkaline earth metal</td>
</tr>
<tr>
<td>Atomic Weight</td>
<td>40.078 g•mol⁻¹</td>
</tr>
<tr>
<td>Group,Period,Block</td>
<td>2,4,s</td>
</tr>
<tr>
<td>Electron Configuration</td>
<td>[Ar]4s²</td>
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<tr>
<td>Valence Electrons</td>
<td>2</td>
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<tr>
<td>Phase (room temperature)</td>
<td>solid</td>
</tr>
<tr>
<td>Melting Point</td>
<td>1115 K, 842°C</td>
</tr>
<tr>
<td>Boiling Point</td>
<td>1757 K, 1484 °C</td>
</tr>
<tr>
<td>Atomic Radius</td>
<td>197 pm</td>
</tr>
<tr>
<td>Oxidation States</td>
<td>2</td>
</tr>
<tr>
<td>Density at room temp</td>
<td>1.55 g•cm⁻³</td>
</tr>
<tr>
<td>Electronegativity</td>
<td>1.00 (Pauling)</td>
</tr>
<tr>
<td>First ionization energy</td>
<td>589.8 kJ•mol⁻¹</td>
</tr>
<tr>
<td>Number of stable isotopes</td>
<td>4 (2 more are fairly stable)</td>
</tr>
</tbody>
</table>
Discovery and Properties of Calcium

In 1808, British chemist Sir Humphry Davy first isolated elemental calcium using electrolysis. Calcium is the lightest of all metals, with a density 1.55 g/cm\(^3\). It reacts with both air and water, usually in reactions involving calcium carbonate, but this reaction is quite slow because calcium hydroxide, Ca(OH)\(_2\), is not very soluble in water.

Reactions

Reaction of Calcium with Halides

Calcium forms salts with halides, such as CaCl\(_2\) or CaF\(_2\). They have a variety of uses, but the most usage most familiar to chemistry students is the use of calcium chloride as a desiccant (drying agent).

\[ CaCl_2 + 2 H_2O(l) \rightarrow CaCl_2 \cdot 2H_2O \]

Reaction of Calcium with Carbonates

Calcium carbonate is important in the formation of cave stalactites and stalagmites. This reaction allows calcium carbonate to be dissolved into solution as calcium bicarbonate.

\[ CaCO_3(s) + CO_2(l) + H_2O(l) \rightarrow Ca(HCO_3)_2 \]

The reverse reaction then allows the solution to become solid calcium carbonate once again, forming spikes of limestone in caves as the calcium bicarbonate solution drips vertically for several millennia.
**Reaction of Calcium with Water**

Calcium metal is fairly reactive and combines with water at room temperature to produce hydrogen gas and calcium hydroxide.

\[ \text{Ca(s)} + 2\text{H}_2\text{O(g)} \rightarrow \text{Ca(OH)}_2(\text{aq}) + \text{H}_2(\text{g}) \]

Product will reveal hydrogen bubbles on calcium metal's surface.

**Reaction of Calcium with Acid**

Calcium dissolves in acid to form dissociated ions of Ca and Cl along with hydrogen gas.

\[ \text{Ca(s)} + 2\text{HCl(aq)} \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) + \text{H}_2(\text{g}) \]

**Reaction of Calcium with Oxygen**

Calcium metal slowly oxidizes in air, becoming encrusted with white \((\text{CaO})\) and \((\text{CaCO}_3)\), which protect from attack by air. When ignited, Calcium burns to give calcium oxide.

\[ 2\text{Ca (s)} + \text{O}_2 (g) \rightarrow 2\text{CaO(s)} \]

**Calcium in living organisms**

Because calcium is essential for life, it can be found in all organisms, living or dead. Shells of aquatic organisms, snail shells, and egg shells are all composed of mostly calcium carbonate, which can be dissolved in acid. Besides skeletal functions, the Ca\(^{2+}\) ion in animals and many organisms also plays an essential role in signal transduction pathways, neurotransmission, muscle function, fertilization, and enzymatic function. In plants, calcium is also important in the cell wall, membrane, and vacuole.

One of the most important calcium deposits is in coral reefs, which are comprised of mostly calcium carbonate. Coral secrete calcium carbonate over the period of their life, then die to allow new coral to build on top of their calcium carbonate structure. Over massive amounts of time, these calcium deposits grow into gigantic reefs, some of which can be seen from space (like the Great Barrier Reef in Australia). With the waters rich in sunlight and minerals like calcium, photosynthesis in sea plants is highly favored, allowing fish and other marine life to flourish in these regions.

Human bones are made up of mostly calcium phosphate \((\text{Ca}_3(\text{PO}_4)_2)\). Cow milk also contains a large amount of calcium phosphate, which is why human culture encourages children and those particularly susceptible to osteoporosis to drink milk.

**Uses of Calcium**

The first known uses of calcium were by the Romans in the first century to make calcium oxide. Other written documentation around 975 C.E. suggests that plaster of Paris and Calcium Sulfate were medically useful. Calcium is
available in a wide variety of forms, from limestone and chalk (calcium carbonate) to marble (calcite) and pearls. It also has many mineral forms (see link provided in Outside Links section). Calcium carbonate, in moderate amounts, can also be used as an antacid or a calcium supplement. Calcium nitrate is also a common fertilizer. Pure $\text{CaCO}_3$ can be extracted from limestone in a series of three reactions:

- Calcination: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
- Slaking: $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$
- Carbonation: $\text{Ca(OH)}_2 + \text{CO}_2 \rightarrow \text{CaCO}_3 + \text{H}_2\text{O}$

![Fig. 2: A Blue Starfish resting on hard Acropora coral. Lighthouse, Ribbon Reefs, Great Barrier Reef. Image used with permission from Wikipedia.](image)

**Hard water**

Hard water, as opposed to soft water, has a high mineral content of calcium sulfate ($\text{CaSO}_4$) or calcium carbonate ($\text{CaCO}_3$). It also includes magnesium ions ($\text{Mg}^{2+}$) and sometimes iron, aluminum, and manganese. When left to evaporate, white calcium minerals can be seen on sinks, showers, etc.

**Calcium oxide (quicklime)**

Calcium oxide, often referred to as quicklime, has many commercial functions, some of which include making mortar and pottery, food, construction, agriculture, pollution control, and in medicine. It also heats quite readily with water in this reaction:

$$\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 \quad \Delta H = -63.7\, \text{kJ/mol}$$
Outside links

- http://en.wikipedia.org/wiki/Category:Calcium_minerals
- Coca Cola and Egg Shell Experiment

- Egg and Vinegar Experiment
Examples of calcium forms, with pictures

References

Edit section

5. Fig. 1: http://www.green-planet-solar-energy…he-element.gif
6. Fig. 2: http://chemwiki.ucdavis.edu/@api/dek…G?size=webview
7. Fig. 3:http://waterlogicsonline.com/wp-cont…r_and_drop.jpg

Problems
1. Chalk (calcium carbonate) is a common compound of calcium. When an acid such as acetic acid is added to chalk, carbon dioxide is formed. Write and balance this reaction. What are the products of this reaction?

2. How can pure calcium carbonate be produced? What are the reactions called?

3. Which reaction is responsible for the formation of stalactites and stalagmites? Under what conditions?

4. Which compound of calcium is found in cow's milk, and how does this relate to overall bone health?

5. A sample of hard water has 156.9 ppm Ca\(^{2+}\). An ion exchange column removes all Ca\(^{2+}\) and replaces it with H\(_3\)O\(^+\). What is the final pH of the water as it leaves the column, assuming it began at pH of 7 and no other ions were produced?

**Solutions**

1) \(2\text{CaCO}_3 + \text{HC}_2\text{H}_3\text{O}_2 \rightarrow \text{Ca(C}_2\text{H}_3\text{O}_2)_2 + \text{CO}_2 + \text{H}_2\text{O}\)

   Calcium acetate, carbon dioxide, and water.

2) Purification of limestone through calcination, slaking, and carbonation.
   \(\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2\)
   \(\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2\)
   \(\text{Ca(OH)}_2 + \text{CO}_2 \rightarrow \text{CaCO}_3 + \text{H}_2\text{O}\)

3) \(\text{CaCO}_3 + \text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{Ca(HCO}_3)_2\)

   Under acidic conditions.

4) Calcium phosphate (Ca\(_3\)(PO\(_4\))\(_2\)). Bones are made from a mineral of this called hydroxylapatite with the formula Ca\(_{10}\)(PO\(_4\))\(_6\)(OH)\(_2\). More calcium phosphate means more mineral for bone growth, ensuring that enough calcium is available for the body to both make bones and have enough Ca\(^{2+}\) ions for other important signaling processes.

5) \(156.9 \text{ mg } \text{Ca}^{2+} \times \left(1 \text{ mol } \text{Ca}^{2+}\right) / \left(40080 \text{ mg } \text{Ca}^{2+}\right) \times \left(2 \text{ mol } \text{H}_3\text{O}^+\right) / \left(1 \text{ mol } \text{Ca}^{2+}\right) = 7.83 \times 10^{-3} \text{ M } \text{H}_3\text{O}^+\)

   \(-\log(7.83 \times 10^{-3}) = \text{pH } 2.11\)
Contributors

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