The elements in the group include beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra).

### Introduction

Group 2 contains soft, silver metals that are less metallic in character than the Group 1 elements. Although many characteristics are common throughout the group, the heavier metals such as Ca, Sr, Ba, and Ra are almost as reactive as the Group 1 Alkali Metals. All the elements in Group 2 have two electrons in their valence shells, giving them an oxidation state of +2. This enables the metals to easily lose electrons, which increases their stability and allows them to form compounds via ionic bonds. The following diagram shows the location of these metals in the Periodic Table:

![Periodic Table of Elements](image)

The table below gives a detailed account of the descriptive chemistry of each of the individual elements. Notice an increase down the group in atomic number, mass, and atomic radius, and a decrease down the group for ionization energy. These common periodic trends are consistent across the whole periodic table.

<table>
<thead>
<tr>
<th></th>
<th>Atomic #</th>
<th>Mass (g)</th>
<th>Oxidation State(s)</th>
<th>Atomic Radius (pm)</th>
<th>Melting Point (K)</th>
<th>Boiling Point (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Be</td>
<td>4</td>
<td>9.012</td>
<td>+2</td>
<td>192</td>
<td>1560</td>
<td>2746</td>
</tr>
<tr>
<td>Mg</td>
<td>12</td>
<td>24.31</td>
<td>+1, +2</td>
<td>137</td>
<td>923</td>
<td>1363</td>
</tr>
<tr>
<td>Ca</td>
<td>20</td>
<td>40.08</td>
<td>+2</td>
<td>143</td>
<td>1115</td>
<td>1750</td>
</tr>
<tr>
<td>Sr</td>
<td>38</td>
<td>87.62</td>
<td>+2</td>
<td>246</td>
<td>1042</td>
<td>1652</td>
</tr>
</tbody>
</table>
Alkaline Earth Metal Reactions

The reactions of the alkaline earth metals differ from those of the Group 1 metals. Radium is radioactive and is not considered in this section.

- **Reactions with Hydrogen**: All the alkaline earth metals react with hydrogen to create metallic hydrides. Below is an example of a reaction of this type:

  \[ Ca_{(s)} + H_{2\,(g)} \rightarrow CaH_{2\,(s)} \]

- **Reactions with Oxygen**: The alkaline earth metals react with oxygen to produce metal oxides. An oxide is a compound containing oxygen in a -2 oxidation state. The following is an example reaction of an alkaline earth metal with oxygen (beryllium does not react with oxygen, but the other metals react in this manner):

  \[ Sr_{(s)} + O_{2\,(g)} \rightarrow SrO_{2\,(s)} \]

- **Reactions with Nitrogen**: These reactions cannot occur in ordinary condition; extremely high temperatures are required. A theoretical reaction of an alkaline earth metal with nitrogen would proceed in the following manner:

  \[ 3Mg_{(s)} + N_{(g)} \rightarrow Mg_3N_{2\,(s)} \]

- **Reactions with Halogens**: Alkaline earth metals react with halogens to form metal halides. A halide is a compound containing an ionic halogen. A reaction of this type between magnesium and chlorine is given below:

  \[ Mg_{(s)} + Cl_{2\,(g)} \rightarrow MgCl_{2\,(s)} \]

- **Reactions with Water**: Beryllium does not react with water; however, magnesium, calcium, strontium, and barium do react to form metal hydroxides and hydrogen gas. The reaction of barium and water is illustrated in the following equation:

  \[ Ba_{(s)} + 2H_2O_{(l)} \rightarrow Ba(OH)_{2\,(aq)} + H_{2\,(g)} \]

Properties of Individual Alkaline Earth Metals
Beryllium (Be)

- Atomic Number = 4 Mass = 9.012 g mol\(^{-1}\) Electron Configuration = \(1s^22s^2\) Density = 1.85 g cm\(^{-3}\)

Beryllium was first identified in 1798 by Louis-Nicolas Vauquelin, while performing a chemical analysis on aluminum silicates. The element was originally named glucinium, and it was first isolated in 1828 by Antoine Bussy and Friedrich Wohler. In 1898, Paul Lebeau was able to produce the first pure samples of Beryllium by electrolyzing molten beryllium fluoride and sodium fluoride. It was later renamed beryllium (from the Greek word beryl meaning “to become pale”) after a pale, beryllium-containing gemstone called a beryl.

Beryllium is the very first element in Group 2, and has the highest melting point (1560 K) of any element in the group. It is very rare on Earth as well as in the universe and is not considered important for plant or animal life. In nature, it can only be found in compounds with other elements. In solutions, it remains in elemental form only for pH values below 5.5. Beryllium is extremely light with a high ionization energy, and it is used primarily to strengthen alloys.

Applications:

Because beryllium is relatively light and has a wide temperature range, it has many mechanical uses. It can be used in aircraft production in nozzles of liquid-fueled spacecrafts, and mirrors in meteorological satellites. The famous Spitzer Space Telescope’s optics are composed entirely of beryllium.

One of the most important applications of beryllium is the production of radiation windows. As beryllium is almost transparent to x-rays, it can be used in windows for x-ray tubes. The minimal absorption by Beryllium greatly reduces heating effects due to intense radiation.

Beryllium Isotopes:

Beryllium is a monoisotopic element—it has only one stable isotope, \(^{9}\text{Be}\). Another notable isotope is cosmogenic \(^{10}\text{Be}\), which is produced by cosmic ray spallation of oxygen and nitrogen. This isotope has a relatively long half-life of 1.51 million years, and is useful in examining soil erosion and formation, as well as the age of ice cores.

Beryllium Compounds:

Beryllium forms compounds with most non-metals. The most common compound is beryllium oxide (\(\text{BeO}\)) which does not react with water and dissolves in strongly basic solutions. Because of its high melting point, \(\text{BeO}\) is a good heat conductor in electrical insulators. It is also an amphoteric oxide, meaning it can react with both strong acids and bases.
Cation: \(\text{H}_2\text{O}_{(l)} + \text{BeO}_{(s)} + 2\text{H}_3\text{O}^+_{(aq)} \rightarrow [\text{Be(H}_2\text{O)}_4]^2+_{(aq)}\)  
Anion: \(\text{H}_2\text{O}_{(l)} + \text{BeO}_{(s)} + 2\text{OH}^-_{(aq)} \rightarrow [\text{Be(OH)}_4]^{2-}_{(aq)}\)

**Magnesium (Mg)**

- Atomic Number = 12  
- Mass = 24.31 g mol\(^{-1}\)  
- Electron Configuration = [Ne]3s\(^2\)  
- Density = 1.738 g cm\(^{-3}\)

Magnesium was first discovered in 1808 by Sir Humphry Davy in England by the electrolysis of magnesia and mercury oxide. Antoine Bussy was the first to produce it in consistent form in 1831.

It is the 8th most abundant element in the Earth’s crust, constituting 2% by mass. It is also the 11th most common element in the human body: fifty percent of magnesium ions are found in bones, and it is a required catalyst for over three hundred different enzymes. Magnesium has a melting point of 923 K and reacts with water at room temperature, although extremely slowly. It is also highly flammable and extremely difficult to extinguish once ignited. As a precaution, when burning or lighting pure magnesium, UV-protected goggles should be worn, as the bright white light produced can permanently damage the retina.

Magnesium can be found in over 100 different minerals, but most commercial magnesium is extracted from dolomite and olivine. The Mg\(^{2+}\) ion is extremely common in seawater and can be filtered and then electrolyzed to produce pure magnesium.

**Applications:**

- In its elemental form, magnesium is used for structural purposes in car engines, pencil sharpeners, and many electronic devices such as laptops and cell phones. Due to its bright white flame color, magnesium is also often used in fireworks.
- In a biological sense, magnesium is vital to the body’s health: the Mg\(^{2+}\) ion is a component of every cell type. Magnesium can be obtained by eating foods rich in magnesium, such as nuts and certain vegetables, or by eating supplementary diet pills. Chlorophyll, the pigment that absorbs light in plants, interacts heavily with magnesium and is necessary for photosynthesis.

**Magnesium Isotopes:**

The three stable isotopes of magnesium are \(^{24}\text{Mg}\), \(^{25}\text{Mg}\), and \(^{26}\text{Mg}\). The lightest isotope, \(^{24}\text{Mg}\), composes about 74% of Mg, whereas \(^{26}\text{Mg}\) is associated with meteorites in the solar system.

\(^{28}\text{Mg}\) is the only known radioactive isotope of Magnesium and it has a half-life of around 21 hours.

**Magnesium Compounds:**

Magnesium ions are essential for all life on Earth, and can be found mainly in seawater and the mineral carnallite. Some examples of magnesium compounds include: magnesium carbonate (MgCO\(_3\)), a white...
powder used by athletes and gymnasts to dry their hands for a firm grip; and magnesium hydroxide (Mg(OH)$_2$), or milk of magnesia, used as a common component of laxatives.

**Calcium (Ca)**

- Atomic Number = 20 Mass = 40.08 g mol$^{-1}$ Electron Configuration = [Ar]$4s^2$ Density = 1.55 g cm$^{-3}$

Calcium was isolated in 1808 by Sir Humphry Davy by the electrolysis of lime and mercuric oxide. In nature, it is only found in combination with other elements. It is the 5th most abundant element in the Earth’s crust, and is essential for living organisms. Calcium, in the presence of Vitamin D, is well known for its role in building stronger, denser bones early in the lives of humans and other animals. Calcium can be found in leafy green vegetables as well as in milk, cheese, and other dairy products. Calcium has a melting point of 1115 K and gives off a red flame when ignited. Calcium was not readily available until the early 20th Century.

**Applications:**

Calcium is an important component in cement and mortars, and thus is necessary for construction. It is also used to aid cheese production.

**Calcium Isotopes:**

The four stable isotopes of calcium are $^{40}$Ca, $^{42}$Ca, $^{43}$Ca, $^{44}$Ca. The most abundant isotope, $^{40}$Ca, composes about 97% of naturally occurring calcium. $^{41}$Ca is the only radioactive isotope of calcium with a half life of 103,000 years.

**Calcium Compounds:**

The most common calcium compound is calcium carbonate ($\text{CaCO}_3$). Calcium carbonate is a component of shells in living organisms, and is used as a commercial antacid. It is also the main component of limestone. As shown below, three steps are required to obtain pure $\text{CaCO}_3$ from limestone: calcination, slaking, and carbonation.

$$\text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g) \text{ (calcination)}$$

$$\text{CaO}(s) + \text{H}_2\text{O}(l) \rightarrow \text{Ca(OH)}_2(s) \text{ (slaking)}$$

$$\text{Ca(OH)}_2 + \text{CO}_2(g) \rightarrow \text{CaCO}_3(s) + \text{H}_2\text{O}(l) \text{ (carbonation)}$$

Another important calcium compound is calcium hydroxide ($\text{Ca(OH)}_2$). Often referred to as 'slack lime', it can be refined to form cement. The formation of calcium hydroxide is given below:

$$\text{[CaO}_2\text{(s)}+\text{H}_2\text{O}_2\text{(l)}] \rightarrow \text{Ca(OH)}_2$$

**Strontium (Sr)**

- Atomic Number = 38 Mass = 87.62 g mol$^{-1}$ Electron Configuration = [Kr]$5s^2$ Density = 2.64 g cm$^{-3}$
Strontium was first discovered in 1790 by Adair Crawford in Scotland and is named after the village it was discovered in, Strontian. In nature, it is only found in combination with other elements as it is extremely reactive. It is the 15th most abundant element on Earth and is commonly found in the form of the mineral celestite. Strontium metal is a slightly softer than calcium and has a melting point of 1042 K.

**Applications:**

In its pure form, Strontium is used in alloys. It can also be used in fireworks as it produces a scarlet flame color. Strontium ranelate ($\text{C}_{12}\text{H}_6\text{N}_2\text{O}_8\text{Sr}_2$) is used to treat sufferers of osteoporosis and strontium chloride ($\text{SrCl}_2$) is used to make toothpaste for sensitive teeth.

**Strontium Isotopes:**

Strontium has four stable isotopes: $^{84}\text{Sr}$, $^{86}\text{Sr}$, $^{87}\text{Sr}$, and $^{88}\text{Sr}$. About 82% of naturally occurring strontium comes in the form of $^{88}\text{Sr}$.

**Strontium Compounds:**

Some applications of strontium compounds include strontium carbonate ($\text{SrCO}_3$), strontium sulfate ($\text{SrSO}_4$), and strontium nitrate ($\text{Sr(NO}_3)_2$), which can be used as a red flame in fireworks.

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**Radium (Ra)**

- Atomic Number = 88  Mass = 226.0 g mol$^{-1}$  Electron Configuration = $[\text{Rn}]7s^2$
- Density = 5.5 g cm$^{-3}$
Radium was first discovered in 1898 by Marie Sklodowska-Curie, and her husband, Pierre Curie, in a pitchblende uranium ore in North Bohemia in the Czech Republic; however, it was not isolated as a pure metal until 1902.

Radium is the heaviest and most radioactive of the alkaline earth metals and it reacts explosively with water. Radium appears pure white but when exposed to air it immediately oxidizes and turns black. Because radium is a decay product of uranium, it can be found in trace amounts in all uranium ores. The exposure or inhalation of radium can cause great harm in the form of cancer and other disorders.

Radium Isotopes:

There are 25 isotopes of radium that are known to exist, but only 4 are found in nature. However, none are stable. The isotope which has the longest half life is $^{226}\text{Ra}$, which is produced by the decay of Uranium.

The four most stable isotopes are $^{223}\text{Ra}$, $^{224}\text{Ra}$, $^{226}\text{Ra}$, and $^{228}\text{Ra}$. The three most abundant: $^{223}\text{Ra}$, $^{224}\text{Ra}$ and $^{226}\text{Ra}$ decay by emitting alpha particles, whereas $^{228}\text{Ra}$ decays emitting a beta particle. Most radium isotopes have relatively short half-lives.

Radium Compounds:

Radium compounds are extremely rare in nature because of its short half-life and intense radioactivity. As such, radium compounds are found almost entirely in uranium and thorium ores. All known radium compounds have a crimson colored flame.

The most important compound of radium is radium chloride ($\text{RaCl}_2$). Previously, it had only been found in a mixture with barium chloride, but $\text{RaCl}_2$ appeared to be less soluble than barium chloride, the mixture could continually be treated to form a precipitate. This procedure was repeated several times until the radioactivity of the precipitate no longer increased, as radium chloride could be electrolyzed using a mercury cathode to produce pure radium. Currently, $\text{RaCl}_2$ is still used to separate radium from barium and it is also used to produce Radon gas, which can be used to treat cancer.

References

- Cowen, James A. The Biological Chemistry of Magnesium. VCH, 1995
Problems

1. Which of the following is NOT an alkaline earth metal?
   1. (a) Ba (b) K (c) Mg (d) Be (e) Ra

2. True or False: Alkaline earth metals do not react vigorously with water.

3. What alkaline metal is a main component in our bones?

4. Which group 2 metal has the largest ionization energy?

5. Which element has the lowest melting point?
   1. (a) Ba (b) Ra (c) Ca (d) Mg

6. Why is the term alkaline used to describe the Group 2 elements?

7. Which Group 2 element has a scarlet flame color?

8. What is the oxidation state of all Group 2 elements?

9. Which Group 2 element is the most radioactive?

10. Which substance acts as the reducing agent in the reaction below?

\[ \text{Be}_{(s)} + \text{F}_2 \rightarrow \text{BeF} \]

Answers

1. B (Potassium)

2. True.

3. Calcium (Ca)

4. Be

5. Ra

6. The name "Alkaline" is from its slight solubility to water, while "Earth" is derived from its inability to decompose when exposed to heat.

7. Strontium

8. +2

9. Radium

10. Be