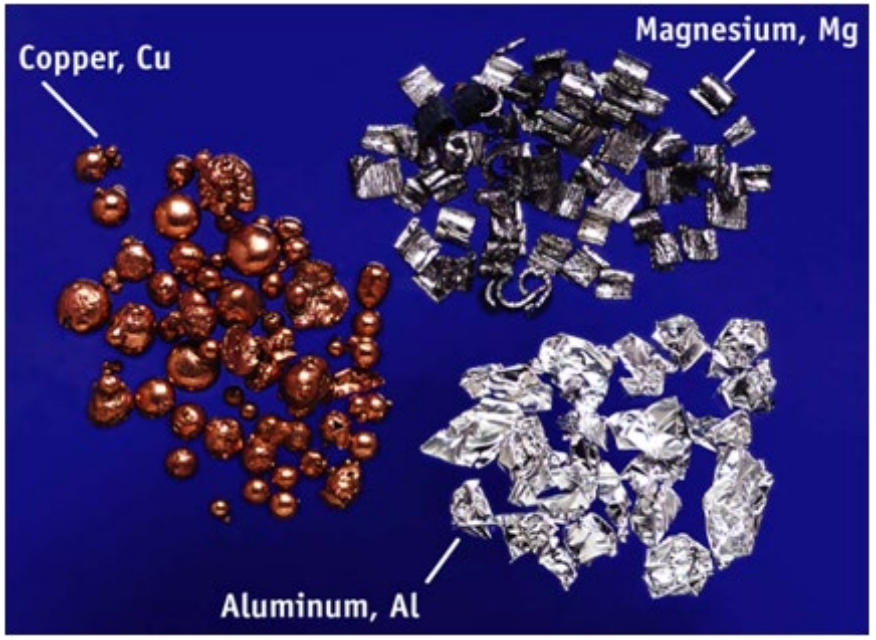
METALS



CHEMISTRY 1000

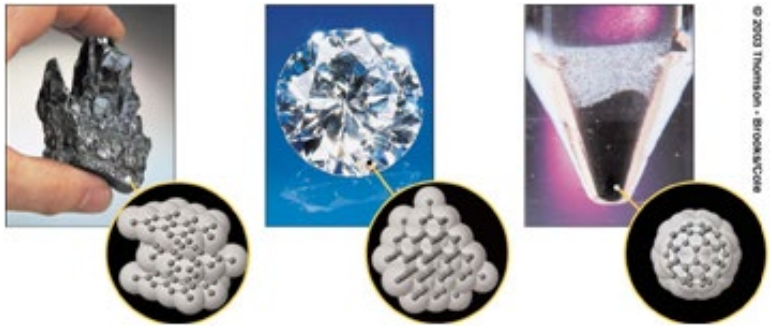
NONMETALS

Topic #2: The Chemical Alphabet

Fall 2020

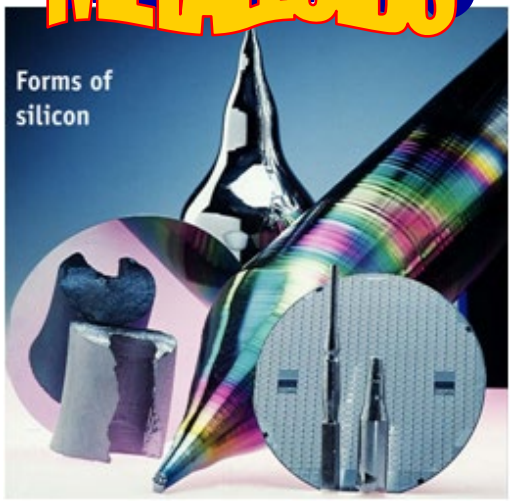
Dr. Susan Findlay

See Exercises 6.1 to 6.5 and 7.2



Forms of Carbon

METALLOIDS



The Alkaline Earth Metals (Group 2)

Beryllium
9.0122
Be
4

Magnesium
24.3050
Mg
12

Calcium
40.078
Ca
20

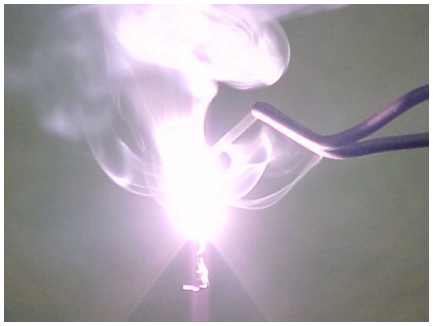
Strontium
87.62
Sr
38

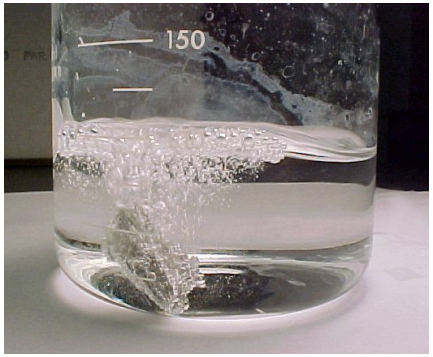
Barium
137.327
Ba
56

Radium
226.025
Ra
88

- What is an alkaline earth metal?
  - Any element in Group 2
  - Form oxides and hydroxides that are "earths" (insoluble in water and heat stable).
  - Harder, denser, higher boiling and higher melting than the alkali metals.

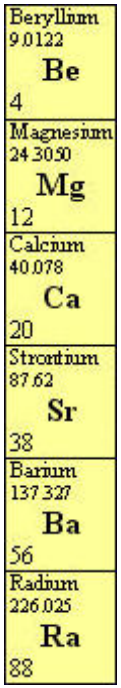
	Melting Point	**Boiling Point**	**Density** (at 20 °C)
Lithium	*180.5 °C*	*1347 °C*	*0.534 g/cm³*
Beryllium	1278 °C	>3000 °C	1.85 g/cm³
Magnesium	648.8 °C	1090 °C	1.74 g/cm³
Calcium	839 °C	1484 °C	1.55 g/cm³
Strontium	769 °C	1384 °C	2.54 g/cm³
Barium	729 °C	1637 °C	3.60 g/cm³
Cesium	*28.4 °C*	*678.5 °C*	*1.873 g/cm³*





Images from http://www.uncp.edu/home/mcclurem/ptable

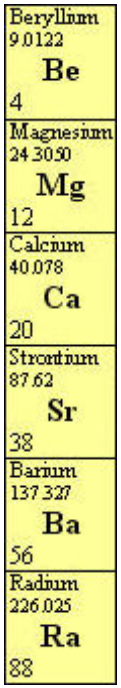
The Alkaline Earth Metals (Group 2)



- What is an alkaline earth metal?
  - Only forms one cation (+2) and no anions
  - Has two valence electrons (electron configuration $[N.G.]ns^2$) and relatively low first and second ionization energies.

	First Ionization Energy (kJ/mol)	**Second Ionization Energy** (kJ/mol)	**Standard Reduction Potential** (V = J/C)
Lithium	*520.2*	*7298*	*-3.040*
Beryllium	899.4	1757	-1.85
Magnesium	737.7	1451	-2.356
Calcium	589.7	1145	-2.84
Strontium	549.5	1064	-2.89
Barium	502.8	965	-2.92
Cesium	*375.7*	*2234*	*-2.923*

The Alkaline Earth Metals (Group 2)



- What is an alkaline earth metal?
  - Most are excellent **reducing agents** (good at losing electrons so that <u>other</u> elements can be reduced). Beryllium is the exception.

	First Ionization Energy (kJ/mol)	**Second Ionization Energy** (kJ/mol)	**Standard Reduction Potential** (V = J/C)
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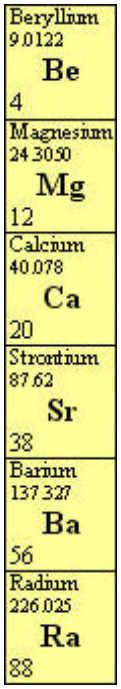
The Alkaline Earth Metals (Group 2)

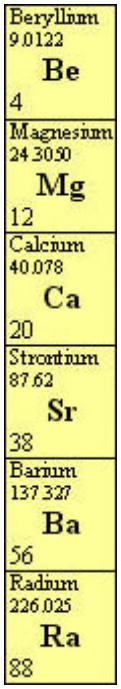


- How can I distinguish between the alkaline earth metals?
  - Flame test. Some of the alkaline earth metal cations give positive flame tests. Complete the table below after doing the Metals Lab.
  - Reactivity with water. Most of the alkaline earth metals react exothermically with water to give the corresponding hydroxide and hydrogen gas. You tested magnesium and calcium in the Metals Lab.

	Flame Colour	**Strength of Reaction with Water**
Beryllium	None	No reaction
Magnesium	None *(Magnesium metal heated in a flame burns bright white, but this is a combustion reaction not a flame test.)*	
Calcium		
Strontium		
Barium		

The Alkaline Earth Metals (Group 2)

- What are some other reactions of the alkaline earth metals?
  - Reaction with oxygen? The alkaline earth metals react with oxygen in the air to give the corresponding oxide:



  - Reaction with nitrogen? When burned in air, alkaline earth metals will react with nitrogen *(as well as with oxygen)* to give the corresponding nitride:

This is different from the alkali metals, of whom only lithium reacts with N_2. Lithium and magnesium have a **diagonal relationship**, meaning that many of lithium's properties are more similar to those of magnesium than of the elements in its own group.

The Alkaline Earth Metals (Group 2)

- What are some other reactions of the alkaline earth metals?
  - Reaction with halogens? The alkaline earth metals react with halogens to give the corresponding halides:

Beryllium 9.0122 **Be** 4
Magnesium 24.3050 **Mg** 12
Calcium 40.078 **Ca** 20
Strontium 87.62 **Sr** 38
Barium 137.327 **Ba** 56
Radium 226.025 **Ra** 88

  - Reactivity of alkaline earth metals increases as size increases because the valence electrons are farther away from the nucleus and therefore easier to remove *(recall ionization energy trends!)*.
  - It's important to note that beryllium is significantly less reactive than all the other alkaline earth metals. It only burns if powdered and heated, and it only reacts with halogens on heating.

The Alkaline Earth Metals (Group 2)

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- The alkaline earth metals react (quite violently!) with acid to produce hydrogen gas and the corresponding halide salt.

  What volume of hydrogen gas is produced if 2.50 g of magnesium metal is dissolved in excess hydrochloric acid in a fumehood with a temperature of 25 °C and a pressure of 1 bar (100 kPa)?

The Alkaline Earth Metals (Group 2)

- What makes beryllium so special?
  - Originally named "glucinium" because some of its salts taste sweet, beryllium and its simpler salts are actually highly toxic, causing cancer or lung damage. The human body cannot excrete beryllium.
  - Now named "beryllium" for one of its most abundant forms in nature – beryl($Be_3Al_2Si_6O_{18}$), a mineral which can be quite valuable when contaminated with the right impurities: chromium-containing beryl is emerald; iron-containing beryl is aquamarine.

  - Beryllium is much smaller than the other alkaline earth metals, so its valence electrons are more strongly attracted to its nucleus. Free Be^{2+} is rarely formed because the +2 charge would be distributed over a very small volume – giving Be^{2+} an unusually high **charge density**. Thus, while $MgCl_2$ and $CaCl_2$ are ionic compounds, $BeCl_2$ is not (its electrons are covalently bonded – shared between two atoms).
  - As a result, beryllium actually behaves more like aluminum than magnesium! This **diagonal relationship** can also be seen between lithium/magnesium and boron/silicon.

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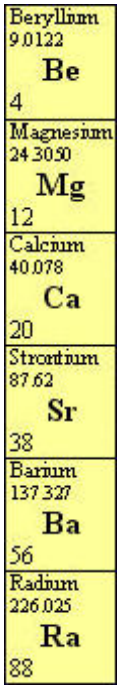
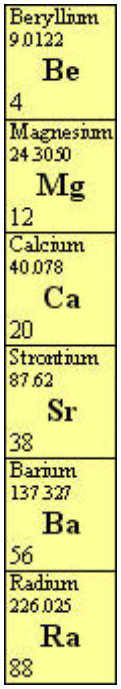


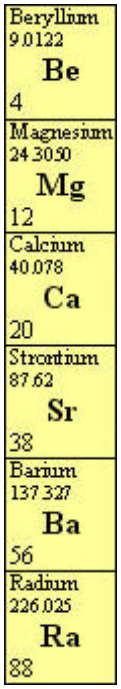
- Diagonal relationship?
  - What happens when we add BeO, MgO or Al_2O_3 to water?

The Alkaline Earth Metals (Group 2)



- Diagonal relationship?
  - What happens when we add BeO, MgO or Al_2O_3 to aqueous acid?

The Alkaline Earth Metals (Group 2)



- Diagonal relationship?
  - What happens when we add BeO, MgO or Al_2O_3 to aqueous base?

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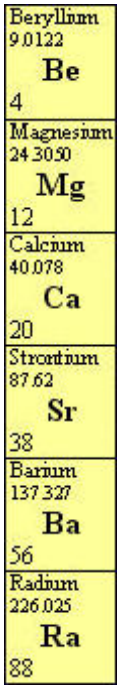
- Are there any really useful alkaline earth metal compounds?
  - Limestone (mostly $CaCO_3$ but also some $\mathrm{Mg}CO_3$) and lime (CaO) are of particular economic importance:
    - Limestone is spread on fields to neutralize acidic compounds in soil and provide Ca^{2+} and Mg^{2+}.
    - Limestone can also be thermally decomposed to make lime:
    - Historically, lime was mixed with sand and water to make mortar (used to bind stone or bricks):

The Alkaline Earth Metals (Group 2)



- Portland cement is the basis of modern mortars and concrete. Lime and shales (mixtures of aluminosilicates) are ground together then heated to 1500 °C. Carbon dioxide is released, and the residue cools to lumps called clinker. The clinker is ground to a fine powder, stabilized by adding calcium sulfate ($CaSO_4$) and sold as **Portland cement** (26% Ca_2SiO_4, 51% Ca_3SiO_5, 11% $Ca_3Al_2O_6$).

  When water is added to the Portland cement, a number of complex hydration reactions set in. To give an idea of the chemistry involved, the following is an idealized reaction equation for this process:

  The main product of this hydration is called **tobermorite gel** (a hydrated calcium silicate) which forms strong crystals that bond strongly to the sand and gravel added to give the concrete volume and strength.

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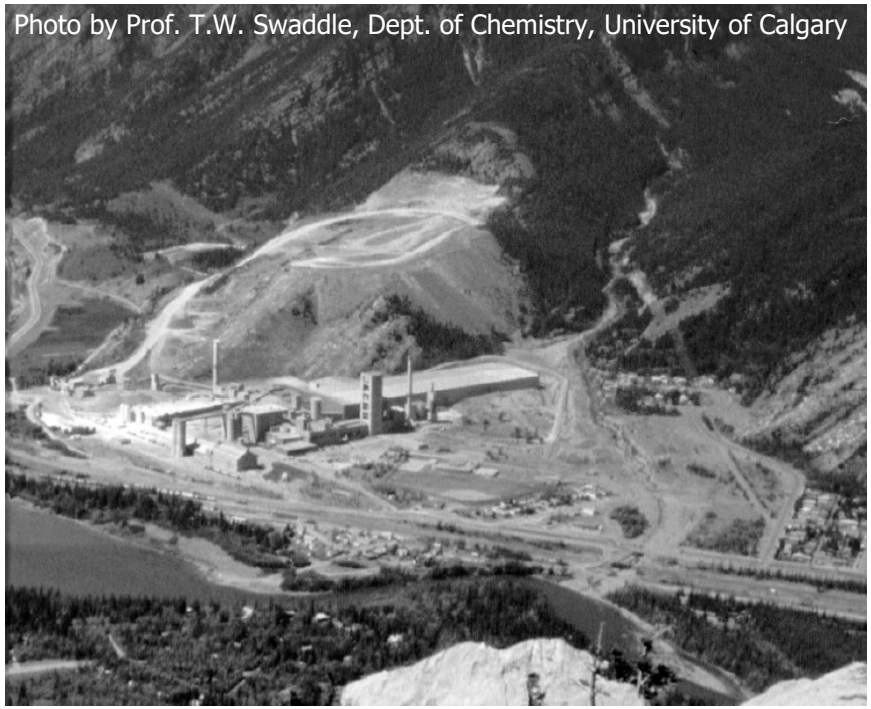
- The main advantage of cement over lime mortar is that is doesn't need external carbon dioxide to set.

  The photo below is of the Lafarge Cement Plant at Exshaw (just east of Banff National Park). The limestone is harvested directly behind the plant; the shales come from Seebe, 7km downstream along the Bow river.



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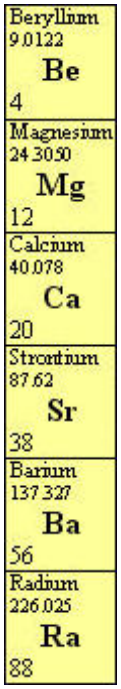

- Calcium carbonate $(CaCO_3)$ – like all other carbonates and sulfates of alkaline earth metals – is insoluble in neutral water.
  Why is that?

- These carbonates are, however, soluble in acid, reacting to produce carbon dioxide:

This reaction can be generalized to all carbonate salts.

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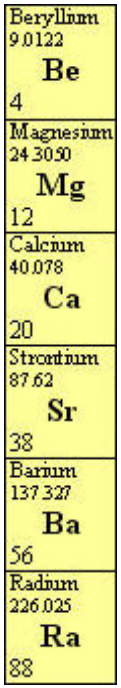
- For this reason, while barium sulfate $(BaSO_4)$ can be used as a contrast agent for x-rays, it must be completely free of barium chloride $(BaCl_2)$ or barium carbonate $(BaCO_3)$.
- The dissolved barium cation $\left(Ba^{2+}_{(aq)}\right)$ is poisonous. In fact, barium carbonate is used as a rat poison. Why would this work?

The Alkaline Earth Metals (Group 2)



- Calcium and magnesium cations are the main culprits in "hard water". Tap water is typically acidic due to its dissolved CO_2:

Thus, carbonates that would otherwise be insoluble (primarily $MgCO_3$ and $CaCO_3$) dissolve in the water:

When this water is heated, the solubility of the carbon dioxide decreases and the reaction above is forced to proceed "in reverse" to restore equilibrium. This leaves "hard" rings of carbonates in your pots, kettles, bathtubs, etc. The cations in hard water also react with soaps (anions with a -1 charge at one end and a long 'greasy' 'tail'), making them precipitate. As such, more soap is needed in "hard water" than in "soft water".

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- Water is typically "softened" by ion-exchange. The water is passed through a system that replaces the calcium and magnesium cations with sodium cations. Since sodium carbonate (like all sodium salts) is soluble in water, regardless of pH, and low concentrations of sodium cations don't make soap precipitate, it doesn't lead to the same problems.

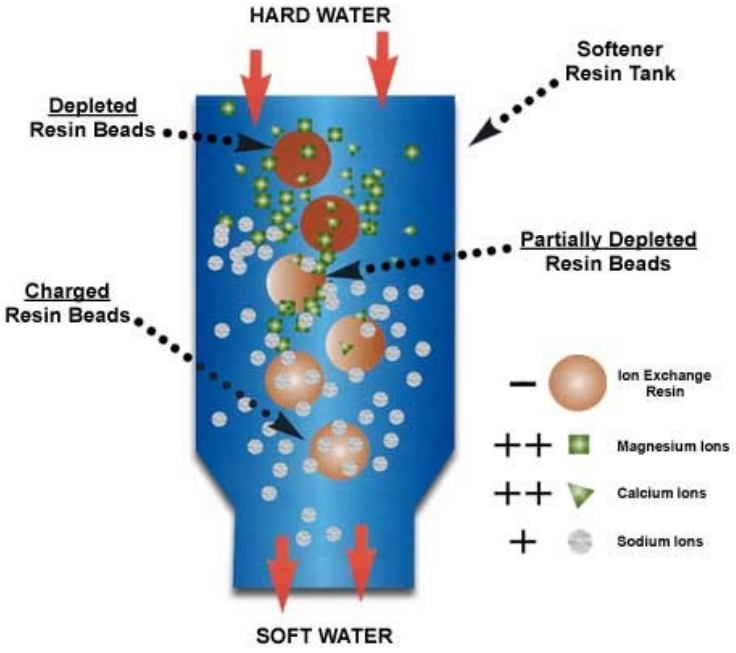

Image from http://www.systemsaver.com/morton-website/education/how-softeners-work/how-softeners-work.html