This page explores the trends in some atomic and physical properties of the Group 2 elements: beryllium, magnesium, calcium, strontium and barium. Sections below cover the trends in atomic radius, first ionization energy, electronegativity, and physical properties.

### Trends in atomic radius

The atomic radius increases down the group. Notice that beryllium has a particularly small radius compared with the rest of the group.

The radius of an atom is governed by

1. the number of layers of electrons around the nucleus
2. the attraction the outer electrons feel from the nucleus.

Compare the electron configurations of beryllium and magnesium:

- Be: $1s^22s^2$
- Mg: $1s^22s^22p^63s^2$

In each case, the two outer electrons experience a net charge of +2 from the nucleus. The positive charge on the nucleus is balanced by the negativeness of the inner electrons.
This is true for all the atoms in Group 2. The only factor influencing the size of the atom is therefore the number of layers of inner electrons that fit around the atom. More layers of electrons take up more space, due to electron-electron repulsion, so atoms are larger down the group.

Trends in first ionization energy

The first ionization energy is the energy needed to remove the most loosely held electron from each of one mole of gaseous atoms to make one mole of singly charged gaseous ions. In other words, for 1 mole of reactant in the process shown below:

\[ X(g) \rightarrow X^+ (g) + e^- \]

Notice that first ionization energy decreases down the group. Ionization energy is governed by three factors:

- the charge on the nucleus,
- the amount of screening by the inner electrons,
- the distance between the outer electrons and the nucleus.

The increase in nuclear charge down the group is exactly offset by the increase in the number of inner electrons. As mentioned before, the outer electrons feel a net charge of +2 from the center. However, down the group, the distance between the nucleus and the outer electrons increases and so they become easier to remove—the ionization energy decreases.
Electronegativity

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. It is usually measured on the Pauling scale, on which the most electronegative element (fluorine) is assigned an electronegativity of 4.0.

Each of these elements has a low electronegativity on the Pauling scale, and the electronegativities clearly decrease down the group. The atoms are decreasingly effective at attracting bonding pairs of electrons. Imagine a bond between a magnesium atom and a chlorine atom. The bond can be considered covalent at first, consisting of a pair of shared electrons. The electron pair is pulled toward the chlorine atom because there is a much greater net pull from the chlorine nucleus than from that of magnesium.

At this point, the electrons are so close to the chlorine that there is essentially a transfer of an electron to the chlorine; in other words, ions are formed. The strong attractive force from the chlorine nucleus makes chlorine much more electronegative than magnesium. This result can be compared with a beryllium-chlorine bond. The net pull from each end of the bond is the same as before, but the beryllium atom is smaller than a magnesium atom; consequently, the electron pair is closer to the net +2 charge from the beryllium end, and is more strongly attracted.
In this case, the electron pair is not close enough to the chlorine for an ionic bond to be formed. Because of its small size, beryllium forms covalent bonds, not ionic bonds. The attraction between the beryllium nucleus and a bonding pair is always too great for ionic bonding.

Summary

As the metal atoms get larger, any bonding pair is pulled progressively farther from the metal nucleus, and so is less strongly attracted toward the metal. In other words, down the group, the elements become less electronegative. Moving down the group, the bonds formed between these elements and other elements such as chlorine become more ionic. The bonding pair is increasingly attracted away from the Group 2 element toward the more electronegative element.

Trends in Melting Points, Boiling Points, and Atomization Energies

- **Melting points** (purple bars): Aside from an irregularity at magnesium, the melting point of each element decreases down the group.
- **Boiling points** (cyan bars): There is no meaningful trend in boiling points for the Group 2 elements.
- **Atomization energy** (orange bars): The atomization energy is the energy needed to produce 1 mole of atoms in the gas phase from the element in its standard state (the expected state at approximately room temperature and pressure). As with boiling points, there is no simple pattern for atomization energies for the Group 2 elements. The chart looks similar, but not identical to, the boiling point chart.
Explaining the trend in atomization energy

Most explanations relate to the melting point. In this section, the most common explanation is given and refuted.

The Incorrect Explanation

Group 2 elements are held together by metallic bonds. The melting points decrease down the group because the metallic bonds are weaker for the larger elements. The deviation of magnesium requires an additional explanation.

The atoms in a metal are held together by the attraction between the nuclei and the delocalized electrons. As the atoms increase in size, the nuclei are further away from these delocalized electrons; therefore, the attractions decrease. The atoms are more easily separated to form a liquid, and then a gas.

Moving down the group, the atoms form different crystal structures in the solid phase. Beryllium and magnesium each have a hexagonal close-packed structure; calcium and strontium form face-centered cubic structures; the structure of barium is body-centered cubic. There is a change in crystal structure between magnesium and calcium, and it could be supposed that this change accounts for the behavior of magnesium.

A problem with this explanation

Despite the fact that the first four elements have two different structures, those structures are both 12-coordinated. Each atom is adjacent to 12 atoms. Therefore, the metallic bond should be similar in each case, because the orbitals are going to overlap and delocalize in the same manner. Any differences due exclusively to structural differences should be minor.
By contrast, barium is 8-coordinated (like the Group 1 metals). This is a less efficient packing system, which is intuitively expected to result in a much weaker metallic bond. Although the barium melting point is lower than that of strontium, it is not dramatically lower. It follows the general trend, suggesting that the major structural change does not make much difference. If a minor structural change between magnesium calcium makes a significant difference, then a major one at barium should make an even more dramatic difference. This is obviously not the case; therefore, the change in structure between magnesium and calcium does not explain the deviation.

The strength of the metallic bonds

Melting points are not accurate indicators of the strength of the metallic bonds. When a metal melts, the bonds are not completely broken—only loosened enough for the atoms to move around. Metallic bonds are still present in the molten metal, breaking only when the metal boils.

Both boiling point and atomization energy are much better indicators of metallic bond strength. Both of these parameters consider the transition to free atoms in the gas state with the metallic bond completely broken.

Cotton and Wilkinson, in their classic degree level book *Advanced Inorganic Chemistry*, support this conclusion: "The strength of binding between the atoms in metals can conveniently be measured by the energies of atomization of the metallic elements." (Third edition, page 68.)

Referring to the atomization energy chart above, magnesium has the lowest value, but there is no obvious trend in atomization energies. Neither is the explanation about weaker metallic bonds down the group accurate.

The figures for Group 1 show that the trends for all the properties discussed (melting point, boiling point and atomization energy) are consistent down the group. Something unique to Group 2 causes complications in simple trends.

A final comment

There is no convincing explanation for the low melting point of magnesium, or the lack of any pattern with the other two properties. A. G. Sharpe, in his degree level book *Inorganic Chemistry* admits that there is no easy explanation for the variations in the physical data in Group 2.

Contributors

Jim Clark ([Chemguide.co.uk](http://Chemguide.co.uk))