Group II - *TheAlkaline Earth Metals*

The elements of Group II, the Alkaline Earth Metals, are:

	symbol	electron configuration
beryllium	Be	$[He]2s^2$
magnesium	Mg	$[Ne]3s^2$
calcium	Ca	$[Ar]4s^2$
strontium	Sr	$[Kr]5s^2$
barium	Ba	$[Xe]6s^2$
radium	Ra	$[Rn]7s^2$

The last element, radium, is radioactive and will not be considered here.

Appearance

The Group II elements are all metals with a shiny, silvery-white colour.

General Reactivity

The alkaline earth metals are high in the reactivity series of metals, but not as high as the alkali metals of Group I.

Occurrence and Extraction

These elements are all found in the earth's crust, but not in the elemental form as they are so reactive. Instead, they are widely distributed in rock structures. The main minerals in which magnesium is found are carnellite, magnesite and dolomite. Calcium is found in chalk, limestone, gypsum and anhydrite. Magnesium is the eighth most abundant element in the earth's crust, and calcium is the fifth.

Only magnesium of the elements in this Group is produced on a large scale. It is extracted from sea-water by the addition of calcium hydroxide, which precipitates out the less soluble magnesium hydroxide. This hydroxide is then converted to the chloride, which is electrolysed in a Downs cell to extract magnesium metal.

Physical Properties

The metals of Group II are harder and denser than sodium and potassium, and have higher melting points. These properties are due largely to the presence of 2 valence electrons on each atom, which leads to stronger metallic bonding than occurs in Group 1.

Three of these elements give characteristic colours when heated in a flame:

Mg	brilliant white	Ca	brick-red	Sr	crimson	Ba	apple green

Atomic and ionic radii increase smoothly down the Group. The ionic radii are all much smaller than the corresponding atomic radii. This is because the atom contains 2 electrons in an s level relatively far from the nucleus, and it is these electrons which are removed to form the ion. Remaining electrons are thus in levels closer to the nucleus, and in addition the increased effective nuclear charge attracts the electrons towards the nucleus and decreases the size of the ion.

Chemical Properties

The chemical properties of Group II elements are dominated by the strong reducing power of the metals. The elements become increasingly electropositive on descending the Group.

Once started, the reactions with oxygen and chlorine are vigorous:

$2Mg(s) + O_{2(}g) \rightarrow 2MgO(s)$

$Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$

All the metals except beryllium form oxides in air at room temperature which dulls the surface of the metal. Beryllium is so reactive it is stored under oil.

All the metals except beryllium reduce water and dilute acids to hydrogen:

$Mg(s) + 2H^+(aq) \rightarrow Mg(aq) + H_2(g)$

Magnesium reacts only slowly with water unless the water is boiling, but calcium reacts rapidly even at room temperature, and forms a cloudy white suspension of sparingly soluble calcium hydroxide.

Calcium, strontium and barium can reduce hydrogen gas when heated, forming the hydride:

$Ca(s) + H_2(g) \rightarrow CaH_2(s)$

The hot metals are also sufficiently strong reducing agents to reduce nitrogen gas and form nitrides:

$3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

Magnesium can reduce, and burn in, carbon dioxide:

$2Mg(s) + CO_2(g) \rightarrow 2MgO(s) + C(s)$

This means that magnesium fires cannot be extinguished using carbon dioxide fire exting u is hers.

Oxides

The oxides of alkaline earth metals have the general formula MO and are basic. They are normally prepared by heating the hydroxide or carbonate to release carbon dioxide gas. They have high lattice enthalpies and melting points. Peroxides, MO_2, are known for all these elements except beryllium, as the Be^{2+} cation is too small to accommodate the peroxide anion.

Hydroxides

Calcium, strontium and barium oxides react with water to form hydroxides:

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

Calcium hydroxide is known as slaked lime. It is sparingly soluble in water and the resulting mildly alkaline solution is known as lime water which is used to test for the acidic gas carbon dioxide.

Halides

The Group II halides are normally found in the hydrated form. They are all ionic except beryllium chloride. Anhydrous calcium chloride has such a strong affinity for water it is used as a drying agent.

Oxidation States and Ionisation Energies

In all their compounds these metals have an oxidation number of +2 and, with few exceptions, their compounds are ionic. The reason for this can be seen by examination of the electron configuration, which always has 2 electrons in an outer quantum level. These electrons are relatively easy to remove, but removing the third electron is much more difficult, as it is close to the nucleus and in a filled quantum shell. This results in the formation of $M2^+$ The ionisation energies reflect this electron arrangement. The first two ionisation energies are relatively low, and the third very much higher.

Industrial Information

Magnesium is the only Group II element used on a large scale. It is used in flares, tracer bullets and incendiary bombs as it burns with a brilliant white light. It is also alloyed with aluminium to produce a low-density, strong material used in aircraft. Magnesium oxide has such a high melting point it is used to line furnaces.

Further Information

For further information look up the individual elements.

Data

	Atomic Number	Relative Atomic Mass	Melting Point/K	Density/kg m^{-3}
Be	4	9.012	1551	1847.7
Mg	12	24.31	922	1738
Ca	20	40.08	1112	1550
Sr	38	87.62	1042	2540
Ba	56	137.33	1002	3594

Ionisation Energies/kJ mol^{-1}

	1st	2nd	3rd
Be	899.4	1757.1	14848
Mg	737.7	1450.7	7732.6
Ca	589.7	1145	4910
Sr	549.5	1064.2	4210
Ba	502.8	965.1	3600

	Atomic Radius/nm	Ionic Radius/nm (M^{2+})	Standard Electrode Potentials/V
Be	0.113	0.034	-1.85
Mg	0.160	0.078	-2.36
Ca	0.197	0.106	-2.87
Sr	0.215	0.127	-2.89
Ba	0.217	0.143	-2.90