

THE PERIODIC TABLE

Group 2

The physical properties of the elements.

Elements of group 2 exist as solids at room temperature due to the strength of metallic bonding. The delocalised electrons in the structures give each element a silvery sheen and causes them to be good electrical conductors.

They differ from other metals of the periodic table in a number of ways:

- i) They are soft.
- ii) They have low melting points, boiling points and low densities.

On descending the group, the atomic radii become bigger but the number of delocalised electrons remain the same.

The metallic bonds then become weaker as delocalised electrons become more thinly spread as metallic radius increases.

This is why the hardness, melting and boiling points decrease from top to bottom.

Densities increase down a group due to the mass of the nuclei increasing faster than the atomic radii and therefore the atomic volumes.

Flame colours

When group 1 or 2 metals or their compounds are placed in a flame, their electrons are pushed to a higher energy level. As the electrons drop back to a lower level again they give out energy in the form of light. The light is characteristic of each particular metal and gives the flame a certain colour.

The energy emitted is only of certain allowed frequencies which correspond to certain electronic transitions. For groups 1 and 2 this happens to be in the visible region.

Lithium	red	calcium	* brick red
Sodium	yellow	strontium	crimson / red
Potassium	lilac	barium	apple green

These can be used in analysis to detect the presence of these elements.

* Note – Calcium gives a red flame if there aren't any sodium impurities present.

When viewed through a diffraction grating the colours are seen as bands of coloured lines. Each element has a specific set of lines and so can be identified in the **emission spectrum** even if other elements are present.

Light coming from a gas in the laboratory or from a distant star can be analysed in this way.

First ionization energy

Element	First Ionisation Energy /kJmol ⁻¹	
Beryllium	900	0.112
Magnesium	740	0.160
Calcium	590	0.197
Strontium	550	0.215
Barium	500	0.217

The first Ionization energies drop down the group as the outer electron is further away from the nucleus and the inner shielding increases. The drop in ionization energy down the group means that the metals become more reactive down the group.

Successive ionisation energies.

The second ionisation energy of group 2 elements is higher than the first as the second electron is removed from an already positive ion. There is greater attraction to as the effective nuclear charge attracting the outer electron has increased.

The third ionisation energy of group 2 elements is much higher than the second. The third electron must be removed from an doubly positively charged ion and from a stable full energy level closer to the nucleus.

Group 2 Element	configuration	I.E.s in kJ mol ⁻¹		
		first	second	Third
Mg at.no. 12	[Ne] 3s ²	736	1450	7733
Ca at.no. 20	[Ar] 4s ²	590	1150	4912
Sr at.no. 38	[Kr] 5s ²	548	1060	4210
Ba at.no. 56	[Xe] 6s ²	502	966	

Reactions of group 2 elements

Reactions with oxygen

Group 2 elements form the oxide of formula MO when heated in oxygen.

For example magnesium;



On prolonged heating in oxygen, Sr and Ba can go on to form the peroxide (containing the O_2^{2-} ion, which is stabilised by large cations):



Reactions with chlorine

All the group 2 metals react when heated in chlorine to form chlorides.

Magnesium for example:

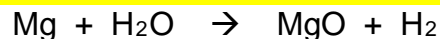


Reactions with water

Group 2 elements are less reactive than the corresponding group 1 element.

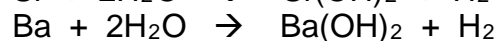
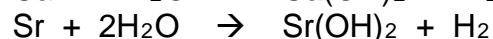
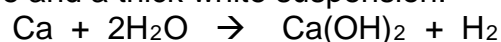
Beryllium does not react even in steam.

Magnesium burns when heated in steam to form magnesium oxide and hydrogen.



Calcium, strontium and barium all react with cold water with increasing vigour to form the metal hydroxide and hydrogen.

Calcium hydroxide is only sparingly soluble in water, so this reaction produces effervescence and a thick white suspension.



Reactions of the oxides of group 2 elements

Reactions with water

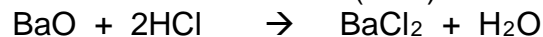
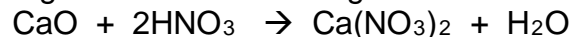
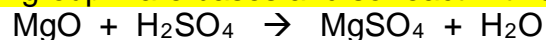
BeO is not attacked by water.

The oxides of the other group 2 elements are ionic and react with water to form hydroxides.



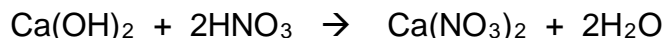
Reactions with dilute acid

The oxides of group 2 are bases and so react with acids to form salt and water.



Reactions of the hydroxides of group 2 elements

Hydroxides of group 2 are also bases and so react with acids to form salts and water;



Solubility of group 2 sulphates and hydroxides

The table below shows the trend in the solubilities of group 2 sulphates.

Sulphate	description	Solubility mol/100g water
Magnesium	Soluble	3600×10^{-4}
Calcium	Slightly soluble	11×10^{-4}
Strontium	Insoluble	0.62×10^{-4}
Barium	Insoluble	0.009×10^{-4}

It can be seen that the sulphates become less soluble down the group.

Magnesium sulphate is very soluble, barium sulphate is insoluble and is part of the test for sulphates.

The hydroxides show the reverse of this trend with the compounds in general becoming more soluble down the group.

Calcium hydroxide is only slightly soluble in limewater but barium hydroxide is a very soluble alkali which can be used in titrations.

Hydroxide	Solubility mol/100g water
Magnesium	0.2×10^{-4}
Calcium	16×10^{-4}
Strontium	330×10^{-4}
Barium	240×10^{-4}

Thermal stability Carbonates and Nitrates

Decomposition of certain carbonates and nitrates of groups 1 and 2 occurs when the compounds are heated.

All group 1 carbonates are thermally stable up to Bunsen temperatures.

Lithium and group 2 carbonates decompose when heated to form the metal oxide and carbon dioxide.



The carbonates become less stable on ascending the group.

Carbonate	Decomposition temperature /°C
Beryllium	100
Magnesium	540
Calcium	900
Strontium	1290
Barium	1360

The decomposition is a result of the polarising power of the cation.

As polarising power of the cation increases, compounds become more covalent in character and less stable to heat.

Going down the group the cations become larger, so their polarising power decreases.

The smallest ion, Mg^{2+} , has the **highest charge density** (is the most polarising), and forms the compound with the small anion most readily (i.e. at the lowest temperature). A large anion like CO_3^{2-} or NO_3^- is **polarisable**, and is decomposed most readily by a polarising cation.

Group 1 compounds are more stable because the cation has only one positive charge and so it is less polarising.

It is the polarising power of the cation that causes decomposition of the nitrates also.

All the nitrates decompose.

Group 1 nitrates, except for lithium break down to form metal nitrite and oxygen.

Lithium and group 2 elements, having a more polarising cation break down to a greater extent into metal oxide, nitrogen dioxide and oxygen.

Li	Be
Na	Mg
K	Ca
Rb	Sr
Cs	Ba

