3.1.2 Group 2

Atomic radius
Atomic **radius increases** down the Group.
As one goes down the group the atoms have **more shells of electrons** making the atom bigger

Electronic Structure
Group 2 metals all have the outer shell s^2 electron configuration.

Melting points
Melting **points decrease** down the group. The **metallic bonding** weakens as the **atomic size increases**. The distance between the positive ions and delocalized electrons increases. Therefore the electrostatic attractive forces between the **positive ions** and the **delocalized electrons** weaken.

Ionisation energy

When the group 2 metals react, they lose their outer shell s^2 electrons in redox reactions to form 2^+ ions.
The energy to remove these electrons are the first and second ionisation energies.

The first ionisation energy is the energy needed to remove an electron from **each atom** in **one mole** of **gaseous atoms**

This is represented by the equation: $H_{(g)} \rightarrow H^+_{(g)} + e^-$

Always gaseous

The first and second ionisation energies decrease down the group.
The outermost electrons are held more weakly because they are **successively further from the nucleus** in additional shells
In addition, the outer shell electrons become **more shielded** from the attraction of the nucleus by the repulsive force of inner shell electrons

Second ionisation energy
The second ionisation energy is the enthalpy change when one mole of gaseous ions with a single positive charge forms one mole of gaseous ions with a double positive charge

This is represented by the equation: $Ti^+_{(g)} \rightarrow Ti^{2+}_{(g)} + e^-$

Group 2 reactions

Reactivity of group 2 metals **increases** down the group

The reactivity increases down the group. As the atomic radii increase there is more shielding. The nuclear attraction decreases and it is easier to remove outer electrons. Cations form more easily.

Reactions with oxygen.

The group 2 metals will burn in oxygen.
Mg burns with a bright white flame
$2Mg + O_2 \rightarrow 2MgO$

MgO is a white solid with a high melting point due to its ionic bonding

Mg will also react slowly with oxygen without a flame.
Mg ribbon will often have a thin layer of magnesium oxide on it formed by reaction with oxygen
$2Mg + O_2 \rightarrow 2MgO$
This needs to be cleaned off by emery paper before doing reactions with Mg ribbon
If testing for reaction rates with Mg and acid, an un-cleaned Mg ribbon would give a false result because both the Mg and MgO would react but at different rates.
$Mg + 2HCl \rightarrow MgCl_2 + H_2$
$MgO + 2HCl \rightarrow MgCl_2 + H_2O$

Reactions with water.

Magnesium reacts **in steam** to produce **magnesium oxide** and hydrogen. The Mg would burn with a bright white flame

$Mg\ (s) + H_2O\ (\mathbf{g}) \rightarrow \mathbf{MgO}\ (s) + H_2\ (g)$

Mg will also react with warm water, giving a different **magnesium hydroxide** product

$Mg + 2\ H_2O \rightarrow \mathbf{Mg(OH)_2} + H_2$

This is a much slower reaction than the reaction with steam and there is no flame

The other group 2 metals will react with **cold water** with increasing vigour down the group to form **hydroxides**

$Ca + 2\ H_2O\ (\mathbf{l}) \rightarrow \mathbf{Ca(OH)_2}\ (aq) + H_2\ (g)$

$Sr + 2\ H_2O\ (\mathbf{l}) \rightarrow \mathbf{Sr(OH)_2}\ (aq) + H_2\ (g)$

$Ba + 2\ H_2O\ (\mathbf{l}) \rightarrow \mathbf{Ba(OH)_2}\ (aq) + H_2\ (g)$

The hydroxides produced make the water alkaline

One would observe:

- fizzing, (more vigorous down group)
- the metal dissolving, (faster down group)
- the solution heating up (more down group)
- and with calcium a white precipitate appearing (less precipitate forms down group)

Reactions with acid

The group 2 metals will react with acids with increasing vigour down the group to form a salt and hydrogen

$Ca + 2HCl\ (aq) \rightarrow \mathbf{CaCl_2}\ (aq) + H_2\ (g)$

$Sr + 2\ HNO_3\ (\mathbf{aq}) \rightarrow \mathbf{Sr(NO_3)_2}\ (aq) + H_2\ (g)$

$Mg + H_2SO_4\ (\mathbf{aq}) \rightarrow MgSO_4\ (aq) + H_2\ (g)$

If barium metal is reacted with sulfuric acid it will only react slowly as the insoluble barium sulfate produced will cover the surface of the metal and act as a barrier to further attack.
$Ba + H_2SO_4 \rightarrow BaSO_4 + H_2$
The same effect will happen to a lesser extent with metals going up the group as the solubility increases.
The same effect does not happen with other acids like hydrochloric or nitric as they form soluble group 2 salts.

Action of water on oxides of elements in Group 2

The group 2 oxides react with water to form hydroxides of varying solubility

$CaO\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (aq)$ **pH 12**

Group 2 oxides are basic as the oxide ions accept H^+ ions to become hydroxide ions in these reactions

$MgO\ (s) + H_2O\ (l) \rightarrow Mg(OH)_2\ (s)$ **pH 9**
$Mg(OH)_2$ is only slightly soluble in water so fewer free OH^- ions are produced and so lower pH

Magnesium hydroxide is classed as partially soluble in water.

A suspension of magnesium hydroxide in water will appear slightly alkaline (pH 9) so some hydroxide ions must therefore have been produced by a very slight dissolving.

Magnesium hydroxide is used in medicine (in suspension as milk of magnesia) to neutralise excess acid in the stomach and to treat constipation.

$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$

It is safe to use as it so weakly alkaline.

Calcium hydroxide is reasonably soluble in water. It is used in agriculture to neutralise acidic soils.

If too much calcium hydroxide is added to the soil, excess will result in soils becoming too alkaline to sustain crop growth

An aqueous solution of calcium hydroxide is called lime water and can be used a test for carbon dioxide. The limewater turns cloudy as white calcium carbonate is produced.

$Ca(OH)_{2\ (aq)} + CO_{2\ (g)} \rightarrow CaCO_{3\ (s)} + H_2O_{(l)}$