9.2 Group II

Content

9.2.1 Similarities and trends in the properties of the Group II metals magnesium to barium and their compounds

Learning Outcomes

Candidates should be able to:

(a) describe the reactions of the elements with oxygen and water

(b) describe the behaviour of the oxides with water

(c) interpret and explain qualitatively the trend in the thermal stability of the nitrates in terms of the charge density of the cation and the polarisability of the large anion

(d) interpret, and make predictions from, the trends in physical and chemical properties of the elements and their compounds

9.2.1 Similarities and trends in the properties of Group II metals magnesium to barium and their compounds

Define

s-block element

An ***s*-block element** is any element in *groups I* and *II* of the Periodic Table.

- Group I elements are commonly known as *alkali metals* while Group II elements are commonly known as *alkali earth metals*.
- Group I elements have one valence electron (oxidation state +1) while Group II elements have two valence electrons (oxidation state +2).
- Outer shell configuration: ns^2
- Group II metals are reactive with low electronegativity. They form stable ions with an oxidation state of +2 since its two outer *s* electrons are easily lost.

Example

① Barium and magnesium are examples of Group II elements.

Group II elements	Atomic no.	Electronic configuration
Barium (Ba)	56	$[Xe]\mathbf{6s^2}$
Magnesium (Mg)	12	$1s^2 2s^2 2p^6 \mathbf{3s^2}$

Variations in the physical properties of alkali earth metals

Atomic and ionic radii

- Down the group, the *atomic* and *ionic radii* increase.
- Each element has progressively one more electron shell, thus the outer electrons are further from the nucleus.

Electronegativity

- Down the group, the electronegativity of the elements decreases.
- The atomic size of the elements increases down the group, thus the *nuclear charge effect* decreases. Additional electrons are not held as strongly by the nucleus.

Melting and boiling points

- Down the group, the melting and boiling points decrease.
- As the atomic size of the elements increases down the group, the attraction of the *ions* for the *electron cloud* decreases, thus weakening the metallic bonds.

Ionisation energy

- Down the group, the ionisation energy of the elements decreases.
- The atomic size of the elements increases down the group, thus the *screening effect* increases. Less energy is required to remove an electron.

Hardness

- Down the group, the hardness of the metals decreases.
- The strength of the metallic bonding decreases down the group.

Variations in the chemical properties of alkali earth metals

- All the alkali earth metals are strong *reducing agents* with low ionisation energy, low electronegativity and negative electrode potentials.
- The *reactivity* and *reducing* power of the elements increases down the group as it is easier to remove the outermost electrons.

Reaction with air or oxygen

- All the alkali earth metals burn brilliantly to form a mixture of *metal oxide* and *nitride*.
- Assuming X is the metal atom, the general equation is:

  $2X (s) + O_2 (g) \rightarrow 2XO (s)$

  $3X (s) + N_2 (g) \rightarrow X_3N_2 (s)$

- Going down the group, the rate of reaction and the proportion of nitride formed increases.

Example

① The equation for the reaction of calcium with air is as follows:

$2Ca\ (s) + O_2\ (g) \rightarrow 2CaO\ (s)$

$3Ca\ (s) + N_2\ (g) \rightarrow Ca_3N_2\ (s)$

Reaction with water

- All Group II elements (except beryllium) react with water to form *metal hydroxides* and *hydrogen* gas.
- The *reactivity* of Group II elements with water increases down the group.
- Magnesium reacts readily with steam to form magnesium oxide. It does not react with cold water.
- Calcium, strontium and barium react vigorously with cold water to form hydroxides.
- The *solubility* of Group II hydroxides increases down the group.
- Group II metals must be stored under *oil* to prevent reaction with oxygen and water vapour in the air.
- An exception will be beryllium and magnesium as they form protective oxide layers and protect the metals from corrosion.

Examples

① The equation for the reaction of magnesium with steam is as follows:

$Mg\ (s) + H_2O\ (l) \rightarrow MgO\ (s) + H_2\ (g)$

② The equation for the reaction of calcium with cold water is as follows:

$Ca\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (s) + H_2\ (g)$

Reaction with acids

- All Group II elements react with *acids* to form the corresponding *salts* and *hydrogen* gas.
- Beryllium does not react with acids at *room* temperature. It reacts slowly at *higher* temperatures.
- Magnesium, calcium, strontium and barium react vigorously with acids to form the corresponding ion and hydrogen gas.
- Group II metals should not be placed in *dilute acids* as explosions are likely.

Example

① The equation for the reaction of barium with acid is as follows:

$Ba (s) + 2H^+ (aq) \rightarrow Ba^{2+} (aq) + H_2 (g)$

Variations in the chemical properties of oxides of alkali earth metals

- With the exception of beryllium oxide, which is *amphoteric*, the other oxides are *basic* oxides.
- The oxides of all Group II elements are *colourless* ionic solids.
- Beryllium oxide is *amphoteric* as its bonding has some degree of *covalency*, contributing to its *acidic* nature.
- Group II oxides tend to be less soluble than Group I oxides due to the higher lattice energies.
- Going down the group, the *solubility* of the oxides increases as the lattice energy decreases.
  - Beryllium oxide is virtually insoluble.
  - Magnesium oxide is slightly soluble and forms an insoluble hydroxide.
  - The other oxides all dissolve rapidly to produce alkaline solutions.

    $XO (s) + H_2O (l) \rightarrow X^{2+} (g) + 2OH^- (aq)$

    where X is the metal atom
- The basic oxides react with acids to form *metallic salts* and *water*.

  $XO (s) + 2HCl (aq) \rightarrow XCl_2 (aq) + H_2O (l)$

  where X is the metal atom.
- Being amphoteric, beryllium oxide reacts with alkalis to form *beryllates*.

Examples

① The equation for the reaction of magnesium oxide with water is as follows:

$MgO (s) + H_2O (l) \rightarrow Mg(OH)_2 (aq)$

② The equation for the reaction of calcium oxide with dilute hydrochloric acid is as follows:

$CaO (s) + 2HCl (aq) \rightarrow CaCl_2 (aq) + H_2O (l)$

③ The equation for the reaction of beryllium oxide with sodium hydroxide is as follows:

$BeO (s) + 2NaOH (aq) + H_2O (l) \rightarrow Na_2Be(OH)_4 (aq)$

Relative thermal stabilities of the alkali earth salts

- The *thermal stability* of the salts increases down the group.
  - This is due to the increasing cation size, which decreases the *polarizing power* of the cation.
  - The cations of the elements lower in the group distort the anion clouds to a smaller extent than the cations of the elements higher in the group.
  - Since anions which have highly distorted electron clouds are more readily decomposed by heat, the thermal stability increases down the group.

- Group II salts are less stable than the Group I salts as the *polarizing effect* of the Group II ions are greater, thus distorting the anion cloud to a greater extent.

- Group II nitrates decompose on heating to form the corresponding *oxides*, *nitrogen dioxide gas* and *oxygen gas*.

- The oxides are more stable than the nitrates as it forms stronger lattices, and the smaller oxide anion is less polarisable.

Worked Examples

Example 1

State and explain whether magnesium nitrate or barium nitrate is more susceptible to thermal decomposition.

Solution:

The thermal stability of the nitrates increases down the group. This is due to the increasing cation size, which decreases the polarizing power of the cation. The cations of the elements lower in the group distort the anion clouds to a smaller extent than the cations of the elements higher in the group. Since anions which have highly distorted electron clouds are more readily decomposed by heat, the thermal stability increases down the group.

Thus magnesium nitrate is more susceptible to thermal decomposition. (ans)

☺ *Group II nitrates decompose on heating to form the corresponding oxides, nitrogen dioxide gas and oxygen gas.*

Example 2

Calcium is in Group II of the Periodic Table.

(a) Write a balanced equation to show the action of heat on calcium nitrate.

(b) Compare the thermal stability of calcium nitrate with barium nitrate.

Solution:

(a) $2Ca(NO_3)_2$ (s) $\rightarrow$ $2CaO$ (s) + $4NO_2$ (g) + O_2 (g) (ans)

(b) Calcium nitrate is expected to have a lower thermal stability than barium nitrate. It has a smaller cation than barium, which increases the polarizing action of the cation and the anion is distorted to a greater extent than in barium nitrate. Since anions which have highly distorted electron clouds are more readily decomposed by heat, calcium nitrate is more susceptible to heat. (ans)

Example 3

(a) Explain why magnesium nitride is formed in addition to magnesium oxide when magnesium is burned in air.

(b) Magnesium nitride contains about 73% of magnesium by mass.

Calculate the empirical formula of magnesium nitride.

(c) Write a balanced equation to show what happens when water is added to magnesium nitride.

Solution:

(a) Magnesium is very reactive. Thus nitrogen is also reduced by magnesium to form magnesium nitride at high temperatures, especially when the oxygen has been used up. This only occurs at high temperatures as the triple bond between the nitrogen atoms requires a large amount of energy for the reaction to occur. (ans)

(b) Assuming there is 100 g of magnesium nitride,

	Mg	**N**
Mass/g	73	27
Molar mass/g mol^{-1}	24.3	14
Amount/mol	3.00	1.93
Simplest ratio	3	2

Thus the empirical formula is Mg_3N_2. (ans)

(c) Mg_3N_2 (s) + $6H_2O$ (l) $\rightarrow$ $2Mg(OH)_2$ (s) + $2NH_3$ (g) (ans)

Notes: