GENERAL CHARACTERISTICS OF THE COMPOUNDS OF THE ALKALINE EARTH METALS

The prevalent valence state for elements in Group 2 is 2+, representing a dipositive oxidation state. Alkaline earth metals are known for their inclination to form ionic compounds, yet the ionic character of these compounds is notably lower when compared to alkali metals. This disparity arises from the elevated ionization energy resulting from an increased nuclear charge and a smaller atomic size.

As one progresses down the group, there is a noticeable augmentation in the tendency of these elements to form ionic compounds. This phenomenon explains why compounds such as oxides and others formed by beryllium and magnesium exhibit a more covalent nature when juxtaposed with those formed by the heavier and larger members, namely calcium, strontium, and barium.

Oxides and Hydroxides

Alkaline earth metals, possessing lower electro positivity compared to alkali metals, exhibit a gradual reaction with air or oxygen upon heating, forming monoxides denoted as MO. With the exception of beryllium oxide (BeO), all alkaline earth metals adopt a rock-salt structure.

Beryllium oxide (BeO) stands out for its essentially covalent nature. The enthalpies of formation associated with these oxides are notably high, contributing to their exceptional stability against heat. Furthermore, BeO displays amphoteric properties.

In contrast, the oxides of other alkaline earth metals are ionic and showcase basic characteristics. Upon contact with water, they undergo reactions yielding sparingly soluble hydroxides, as expressed by the general equation $MO + H_2O \rightarrow M(OH)_2$.

The solubility, thermal stability, and basic nature of these hydroxides exhibit an increasing trend with the ascending atomic number, progressing from $Mg\ (OH)_2$ to $Ba\ (OH)_2$. Notably, the hydroxides of alkaline earth metals, while following this trend, remain comparatively less basic and less stable when compared to those of alkali metals. Beryllium hydroxide, in particular, exhibits amphoteric behavior.

The reaction of beryllium hydroxide with hydrochloric acid and water results in the formation of the beryllate ion $[Be\ (OH)_4]_2$-: $Be\ (OH)_2 + 2OH \rightarrow [Be\ (OH)_4]^{2-}$.

Additionally, the reaction of beryllium hydroxide with hydrochloric acid and water leads to the formation of $[Be(H_2O)4]\ Cl_2$

Halides

(a) Alkali metals directly react with halogens to produce compounds of the form MX (where M represents the alkali metal and X denotes the halide ion).

(b) The ionic characteristics of MX compounds increase from LiCl to CsCl.

(c) LiCl exhibits covalent properties due to the polarization of the Cl^- ion by the small Li^+ ion, which causes it to hydrolyze with water, while the others are ionic and do not hydrolyze.

(d) Halides of potassium (K), rubidium (Rb), and cesium (Cs) react with additional halogens to form polyhalides.

$$KI + I_2 \rightarrow KI_3 \xrightarrow[\text{ionisation}]{\text{On}} K^+ + I_3^-$$

$$CSBr + Br_2 \rightarrow CsBr_3 \rightarrow Cs^+ + Br_3^-$$

(e) The alkali metals engage in a reaction with X (a halogen) to generate MX_2.

Ex. $(BeCl_2, MgCl_2, CaCl_2\ \text{etc.})$

(f) The ionic character of MX_2 increases as we move from one point to another.

(g) Barium undergoes combustion when it comes into contact with chlorine gas (Cl_2).

(h) The hydrolytic reactivity of these halides diminishes from one end to the other.

(i) Blank and blank exhibit covalent characteristics.

The order of ionic character is...
Please provide the specific elements or compounds for the blanks in statement (i) if you'd like a more precise rephrasing.

$$BeCl_2 < MgCl_2 < CaCl_2 < SrCl_2 < BaCl_2$$

Solubility in water

$$BeCl_2 > MgCl_2 > CaCl_2 > SrCl_2 > BaCl_2$$

Nature of Bicarbonates and carbonates

The nature of bicarbonates and carbonates in alkaline earth metals reveals distinct properties. Bicarbonates of these metals are only observed in solution. Upon heating, these bicarbonates undergo decomposition, releasing carbon dioxide.

For example, the reaction of calcium bicarbonate is represented as

$$Ca(HCO_3)_2 \xrightarrow{\Delta} CaCO_3 + CO_2 + H_2O$$

In contrast, carbonates of alkaline earth metals typically exist as solids. They exhibit solubility in water in the presence of carbon dioxide. Upon heating, carbonates undergo decomposition, yielding carbon monoxide and the respective metal oxide.

As illustrated by the equation $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$, the temperature required for this decomposition increases from magnesium to barium.

To prevent the decomposition of beryllium carbonate ($BeCO_3$), it is maintained in an atmosphere of carbon dioxide.

The decomposition temperatures for various carbonates are as follows:

$BeCO_3$: < 100°C
$MgCO_3$: 540°C
$CaCO_3$: 900°C
$SrCO_3$: 1290°C
$BaCO_3$: 1360°C

Beryllium carbonate ($BeCO_3$) exhibits unusually low thermal stability due to its existence as $[Be(H_2O_4)]^{+2}CO_3^{-2}$

Nature of Sulphates

The characteristics of sulphates within the alkaline earth metal group exhibit distinct trends. As one descends down the group, the solubility of sulphates diminishes. Calcium sulfate ($CaSO_4$) is sparingly soluble, whereas strontium sulfate ($SrSO_4$), barium sulfate ($BaSO_4$), and radium sulfate ($RaSO_4$) are virtually insoluble in water. The solubilities of beryllium sulfate ($BeSO_4$) and magnesium sulfate ($MgSO_4$) can be attributed to the high energy of solvation associated with the smaller Be^{2+} and Mg^{2+} ions.

Upon subjecting sulphates to heating, they undergo decomposition, resulting in the formation of the corresponding oxide (MO).

This thermal decomposition can be represented by the equation: $2MSO_4 \rightarrow 2MO + 2SO_2 + O_2$ where M represents an alkaline earth metal.

In summary, the solubility patterns and thermal behavior of alkaline earth metal sulphates are distinctive, with a general decrease in solubility and a decomposition process yielding the respective oxide, sulfur dioxide, and oxygen upon heating.

Nitrates

(a) Alkali metals generate nitrates of the form MNO_3 (M representing the alkali metal).

(b) The stability of nitrates increases from $LiNO_3$ to $CsNO_3$. $LiNO_3$ decomposes into lithium oxide and NO_2 when heated.

$$4LiNO_3 \xrightarrow[\text{Oxide}]{\triangle} 2Li_2O + 4NO_2 + O_2$$

(c) When other nitrates are heated, they decompose to yield nitrites and oxygen.

$$MNO_3 \xrightarrow{\triangle} \underset{\text{Nitrite}}{2MNO_2} + O_2$$