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Class 12
Chemistry

Chapter 5
Group-2 Elements
Handwritten Notes

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GROUP 2 ELEMENTS

Also called group IIA
Belongs in s-block (ns^2)
Alkaline Earth Metals
Be, Mg, Ca, Sr, Ba, Ra

General Trends of Physical Properties and Chemical Reactivity of Group 2 Elements.

All group 2 metals form ionic compounds in which they donate two outermost electrons to become an ion with +2 charge. They act as reducing agent as they get oxidized. Going down the group, the metals become more reactive.

- **Electronic Configuration:** The electronic configuration of first three elements are:
  $_4Be = 1s^2, 2s^2$
  $_{12}Mg = 1s^2, 2s^2, 2p^6, 3s^2$
  $_{20}Ca = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$

- **Ionization energy and chemical Reactivity:** Ionization energy is the minimum amount of energy required to remove one mole of electron from the isolated gaseous atom or ion. In group 2, ionization energy decreases from top to bottom. This is due to the increased sheilding effect and a larger distance between the outermost electrons and nucleus. The elements become more reactive going down the group as it gets easier for the atoms to lose two electrons and become 2+ ions.
  Reaction with dilute hydrochloric acid: $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_2 + H_{2(g)}$
  Reaction with oxygen, the metals get more reactive with oxygen down the group (Ba is so reactive that it must be stored in oil to prevent it from reacting with oxygen in the air)

Be > Mg > Ca > Sr > Ba

- **Atomic Radius:** The atoms of Coup 2 elements get larger going down the group. This is because of the increase in extra shell and sheilding effect, attraction between the nucleus and valence electrons decreases. Hence the atom size increases.
- **Melting point:** The melting point decreases down the group, as the outer electrons get further away from the nucleus. This means that the attraction between the nucleus and the delocalised electrons decreases hence strength of metalic bond decreases, causing a decrease in melting point.

Be < Mg < Ca < Sr < Ba

Trends in Chemical Reactivity with Oxygen, Water and Dilute Acids

Reactivity increases down the group as the atomic size increases and electrons are losely packed.

Be > [Ca > Sr > Ba > Mg]
1280 > 838 > 768 > 714 > 650
Melting point in C°

- **Reaction with Oxygen / Formation of Basic Oxide:** Group 2 metals burn in air and react more rapidly in pure oxygen. They form white solid oxides of the type MO. Strontium (Sr) and Barium (Ba) can also form peroxides (MO_2).
  $2Mg + O_2 \rightarrow 2MgO$. Magnesium burns with a bright white flame. The oxides of beryllium and magnesium are insoluble in water, while the solubility of the oxides increases down the group. The heavier oxides react with water to form soluble hydroxides. Most group 2 oxides are basic in nature, and their basic character increases down the group. However, beryllium oxide (BeO) is amphoteric, as it reacts with both acids and bases.
  $BeO + 2HCl \rightarrow BeCl_2 + H_2O$
  $BeO + 2NaOH \rightarrow Na_2BeO + H_2O$
  Reactivity of group 2 metals with oxygen increases down the group. Barium is highly reactive and is therefore stored under oil. Beryllium oxide is covalent in character due to the small size and high polarizing power of the Be^{2+} ion, while the oxides of the remaing group 2 are mainly ionic and basic.
  Some group 2 elements give characteristic flame colors. Flame color is due to 2+ ions. A nichrome wire, cleaned with concentrated HCl is dipped into salt sample and heated in non-luminuous bunsen flame. Calcium (Ca^{2+}) gives brick red, Strontium (Sr^{2+}) gives scarlet red and Barium (Ba^{2+}) gives apple green colored flames.
- **Reaction with water:** Beryllium does not react with water because a protective layer of BeO is formed on its surface, which prevents further reaction. Magnesium does not react with cold water, but it reacts with steam to form magnesium oxide and hydrogen gas. $Mg + H_2O_{(g)} \rightarrow MgO + H_2$

Calcium reacts more readily with water than magnesium. It reacts with cold water to form calcium hydroxide and hydrogen gas. $Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$. This reaction produces a cloudy white suspension because calcium hydroxide is only sligthly soluble in water. The solution becomes weakly alkaline and hydrogen gas is realesed at steady rate. As we go dawn the group, the reactivity with water increases, and hydrogen gas is evolved more rapidly. Except for beryllium and magnesium, all group 2 metals react with water according to the general equation: $M + 2H_2O \rightarrow M(OH)_2 + H_2$

- **Reaction with Dilute Acids:** Group 2 metals undergo redox reaction with dilute acids to form salts and hydrogen gas. The reactions become more vigorous and exothermic. The reaction of all group 2 elements with dilute HCl follows this general equation. $M + 2HCl \rightarrow MCl_2 + H_2$. (vigorous and exothermic more down the group)
- **Reaction of group 2 oxides with acids:** Group 2 oxides are basic oxides, so they react readily with acids to form salts and water. The reactions are examples of neutrallization reactions and are used to prepare many useful compounds.

$MgO + H_2SO_4 \longrightarrow MgSO_4 + H_2O$. Magnesium sulfate is used for short-term relief of constipation, also used as a soaking solution to relive minor sprains, bruises etc.

$CaO + 2HCl \longrightarrow CaCl_2 + H_2O$. Calcium chloride is used as a solidyfying agent in paint production, coagulant in the manufactoring of rubber.

Comparison of Reactivity of Group 1 and Group 2 Element

The reactivity of an element depends on its ionization energy. The lower the ionization energy, the higher the reactivity. Group 1 elements have lower ionization energy than Group 2 elements, so they are more reactive. Alkali metals react very vigorously with water, often producing hydrogen that may catch fire. $Na + 2H_2O \rightarrow 2NaOH + H_2$. Alkaline earth metals react more slowly with water, $Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$. Due to lower ionization energy, group 1 metals are stronger reducing agents than Group 2 metals.

Uses of Group 2 Compounds

Limestone is mainly calcium carbonate ($CaCO_3$) and is widely used as building material. It can be cut into blocks and joined using mortor. Marble, another form of calcium carbonate, is used for decorative tiles and buildings. Most calcium carbonate is used to make cement. In cement manufacture, limestone is heated in a lime kiln to form calcium oxide (quicklime) and carbon dioxide. $CaCO_3 \xrightarrow{heat} CaO + CO_2$. Calcium oxide is then heated with clay to produce cement, which is mixed with sand and stones to make concrete. The strength of concrete is increased by embedding iron rods in it.

Slaked lime (Calcium hydroxide $Ca(OH)_2$) is used in agriculture to neutralize acidic soils and increase soil pH. $Ca(OH)_2 + 2H^+ \longrightarrow Ca^{2+} + 2H_2O$

In medicine, barium sulfate ($BaSO_4$) is used in X-rays examinations of the stomach and intestines because it absorbs x-rays and is insoluble. Group 2 carbonates and hydroxides are used as anticides to neutralize excess stomach acid. For example, calcium carbonate and magnesium carbonate are used in antacid tablets, while milk of magnesia contains magnesium hydroxide. $CaCO_3 + 2HCl \longrightarrow CaCl_2 + CO_2 + H_2O$, $Mg(OH)_2 + 2HCl \longrightarrow MgCl_2 + 2H_2O$

Thermal Decomposition of Group 2 Carbonates and Nitrates

Thermal decomposition is the breakdown of a compound into two or more different substances by the application of heat. Carbonates and Nitrates of group 2 elements decompes when heated. Carbonates decompose to produce carbon dioxide and metal oxide. $MgCO_{3(s)} \xrightarrow{heat} MgO_{(s)} + CO_{2(g)}$. Carbonates require more heat to decompose when moving down the group 2. Group 2 nitrates on thermal decomposition produce a brown gas that is Nitrogen dioxide, along with metal oxide and oxygen. For example, $2Ca(NO_3)_{2(s)} \xrightarrow{heat} 2CaO_{(s)} + 4NO_{2(g)} + O_{2(g)}$. More temperature is required when going down the group.

- **Trend in thermal stabilities:** The thermal stability of Group 2 carbonates and nitrates increases down the group. This means more heat is required to decompose the compounds as we move from Mg to Ba. Compounds with more positive enthalpy of decomposition are more thermally stable. This trend is explained using ion polarization. The carbonate and nitrate ions are large and can be easily polarized by small, highly charged cations. Smaller group 2 cations, such as Mg^{2+}, have greater polarizing power

$MgCO_3$	540°C	(+117)
$CaCO_3$	900°C	(+176)
$SrCO_3$	1280°C	(+238)
$BaCO_3$	1360°C	(+268)

and distort the anions more strongly, that the bond formed has a considerable amount of covelent character. This weakens the bonds within the anion, making decomposition easier. As we go down the group, cation size increases, polarizing power decreases and thermal stability increases.
The order of polarization: $Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$
The order of thermal stability of carbonates: $BaCO_3 > SrCO_3 > CaCO_3 > MgCO_3$
Similar trend is seen for nitrates: $Ba(NO_3)_2 > Sr(NO_3)_2 > Ca(NO_3)_2 > Mg(NO_3)_2$

Trends in Solubility of the Group 2 Hydroxide and Sulphate

Beryllium is excluded because of its small size and high polarizing power give its compounds covelent character, so they decompose easily even at low temperatures

- Solubility of hydroxide of group II elements: The solubility of hydroxides increases as we move down the group, with barium hydroxide being highly soluble in water. Order: $Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$
- Change in hydration enthalpy down the group: Hydration enthalpy depends on ionic charge and size. For ions with the same charge, smaller ions have more negative (more exothermic) hydration enthalpy. Down Group 2, the ionic size increases, so hydration enthalpy decreases in the order: $Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$ This decrease is due to the increasing size of the cations, while the anion remains the same. (hydroxide anion)

(OH) Hydroxide is small anion and (SO_4^{-2}) Sulphate is large anion

- Change in lattice energy down the group: Lattice energy is higher when the ions forming the lattice are smaller. Down Group 2, the size of the cation increases, so lattice energy decreases in the order: $Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$. Since lattice energy is inversely proportional to the sum of ionic radii, the decrease down the group is large and mainly due to the increasing size of the metal cation, not the hydroxide ion.

Solubilities in water of group 2 OH	
$Mg(OH)_2$	2×10^{-5}
$Ca(OH)_2$	1.5×10^{-3}
$Sr(OH)_2$	3.4×10^{-3}
$Ba(OH)_2$	1.5×10^{-2}

- Difference in enthalpy change of solution of Group 2 hydroxide: The enthalpy of solution indicates how soluble a substance is in water. More negative (more exothermic) enthalpy of solution means greater solubility. Down Group 2, the lattice energy of hydroxides decreases greatly due to increasing cation size, while the hydration enthalpy decreases only slightly (becomes less exothermic). As a result, the enthalpy of solution becomes more negative down the group, so the solubility of Group 2 hydroxides increases down the group. A compound is soluble in water only if its enthalpy of solution is negative or slightly positive.
- Solubility of sulphates of group II elements: The solubility of group 2 sulphates decreases down the group, with barium sulphate being almost insoluble in water. This decrease occurs because the radius of the metal ion increases down the group. The order: $MgSO_4 > CaSO_4 > SrSO_4 > BaSO_4$. Down the group, both lattice dissociation enthalpy and hydration enthalpy decrease. However, the hydration enthalpy decreases more rapidly than the lattice dissociation enthalpy. As a result, the enthalpy of solution becomes more endothermic, leading to lower solubility.

Solubilities in water of group 2 SO_4	
$MgSO_4$	1.83
$CaSO_4$	4.66×10^{-2}
$SrSO_4$	7.11×10^{-4}
$BaSO_4$	9.43×10^{-6}

Complexes of the Alkaline Earth Metals

Group 2 elements mainly form ionic compounds containing M^{2+} ions. Compared to alkali metals, they have a greater tendency to form complexes with Lewis bases because they have a higher charge (+2) and smaller ionic size. This tendency is strongest for Be^{2+} and decreases down the group as the ionic radius increases. The chemistry of Be is dominated by its strong Lewis acid behaviour due to its small size and high charge density. In water, Be^{2+} forms the tetrahaderal complex $[Be(H_2O)_4]^{2+}$, making the solution acidic because Be^{2+} polarizes the water molecules. In the presence of strong bases, beryllium forms the hydroxo complex $[Be(OH)_4]^{2-}$, which explains the amphoteric nature of BeO. Beryllium also forms stable fluroide complexes such as $[BeF_4]^{2-}$ and $[BeF_3]^{+}$.
Heavier Group 2 metals form fewer complexes, but Mg^{2+} and Ca^{2+} still form important hydrated complexes. For example, in ageous solution, magnesium exists as the octahedral complex $[Mg(H_2O)_6]^{2+}$ with a coordination of six.

Extraction and Purification Process of Group 2 Elements and thier Compounds

Group 2 elements (Be, Mg, Ca, Sr, Ba, Ra) extraction depends on the nature of their ores, which are commonly found as carbonates, sulphates, orides, or sillicates.

- **Extraction:** 1) **Ore preparation:** Group 2 metals are obtained from ores such as carbonates and sulphates. The ore is first mined, crushed, and powdered to increase surface area and make extraction easier. 2) **Roasting:** Many Group 2 carbonates are thermally unstable and decompose on heating to form metal oxides and carbon dioxide. For example $MgCO_3 \xrightarrow{heat} MgO + CO_2$.
  3) **Reduction:** The metal oxide or halide is then reduced to obtain the metal. Magnesium is extracted by electrolysis of molten magnesium compounds. Calcium and Strontium can be obtained by reducing their halides with sodium or magnesium.
- **Electrolysis (Purification):** Pure Group 2 metals are often obtained or purified by electrolysis of molten chlorides or oxides. Electrolysis helps remove impurities and produces metals in high purity. Beryllium was first obtained by reducing its chloride. Radium chloride is obtained through several chemical separation steps due to its radioactive nature.
- **Examples:** 1) **Beryllium:** Extraction, obtained from beryl ore through chemical treatment involving acid digestion. Purification, Beryllium compounds are converted to beryllium fluroide, which is then reduced with magnesium to obtain pure beryllium metal. 2) **Magnesium:** Extraction, Commonly produced by electrolysis of molten magnesium chloride obtained from seawater or brine. Purification, Furthur purified by vacuum or fractional distillation.

SHORT QUESTIONS

Describe and explain the trend observed in the thermal stabilities of the carbonates of Group 2.
The thermal stability of carbonates of Group 2 increases down the group. On heating, Group 2 carbonates decompose into metal oxide and carbon dioxide. $MCO_3 \xrightarrow{heat} MO + CO_2$
Down the group, the ionic radius of M^{2+} increases. Larger metal ions have lower polarizing power, so they distort the carbonate ion less. As a result, the carbonate ion becomes more stable and requires higher temperature to decompose. $BeCO_3 < MgCO_3 < CaCO_3 < SrCO_3 < BaCO_3$.

Describe the use of Group 2 elements or its compounds in agriculture.
Group 2 elements and their compounds are widely used in agriculture to improve soil quality and plant growth. Calcium Carbonate ($CaCO_3$) is used as agricultural lime to neutralize acidic soils and improve soil fertility. Calcium Sulfate ($CaSO_4 \cdot 2H_2O$), gypsum, is used to improve soil structure and provide calcium and sulfur to plants. Magnesium compounds, such as $MgSO_4$, are used as fertilizers because magnesium is essential for chlorophyll formation. Calcium Nitrate [$Ca(NO_3)_2$] is used as a fertilizer to supply calcium and nitrogen for healthy plant growth.

Write an equation to represent the thermal decomposition of calcium nitrate.
$2Ca(NO_3)_2 \longrightarrow 2CaO + 4NO_2 + O_2$
On heating, calcium nitrate decomposes to form calcium oxide, nitrogen dioxide, and oxygen.

Why are the elements of group 2 called alkaline earth metals.
Group 2 elements are called alkaline earth metals because their oxides and hydroxides are alkaline (basic) in nature and form alkaline solutions in water. These oxides were historically called "earths" because they are insoluble in water and occur naturally in the Earth's crust. They are found mainly as minerals (carbonates and sulphates) in the earth. They show properties intermediate between alkali metals and other metals, hence the name alkaline earth metals.

How do Group 1 metals differ from Group 2 metals.
Group 1 metals have one valence electron and form +1 ions, while Group 2 metals have two valence electrons and form +2 ions. Group 1 metals are more reactive than Group 2 metals because they lose one electrons more easily. Compounds of Group 1 metals are mostly soluble in water, whereas many compounds of Group 2 metals are less soluble or insoluble. Group 1 metals are softer with low melting points, while Group 2 metals are harder and have higher melting points due to stronger metallic bonding.

Explain with the help of an equation the amphoteric nature of beryllium oxide.
Beryllium oxide (BeO) is amphoteric, meaning it reacts with both acids and bases.
With acids (act as a base), BeO reacts with acids to form beryllium salts and water.
$BeO + 2HCl \longrightarrow BeCl_2 + H_2O$.
With bases (act as an acid), BeO reacts with strong bases to form soluble beryllate complexes. $BeO + 2NaOH + H_2O \longrightarrow Na_2[Be(OH)_4]$

Explain qualitatively the variation in solubility of the hydroxides of the elements in group 2 down the group from magnesium to barium.
The solubility of Group 2 hydroxides increases down the group from magnesium to barium. Down the group, the lattice energy decreases rapidly due to the increase in cation size. Although the hydration enthalpy also decreases, it decreases by a smaller amount compared to lattice energy. As a result, the enthalpy of solution becomes more exothermic down the group, making the hydroxides more soluble. Hence, barium hydroxide is the most soluble and magnesium hydroxide is the least soluble. $Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$

CONCEPT ASSESSMENTS

What is the general trend in the melting points going down group 2.
The melting points of Group 2 generally decreases down the group from beryllium to barium. This is because, as we move down the group the atomic size increases, the metallic bonding becomes weaker, the charge density of M^{2+} ions decreases. As a result, less energy is required to break the metallic lattice, so melting points fall.
Explain why the atoms in Group 2, as in any other group, get larger with increasing atomic number.
Atoms in group 2 get larger as the atomic number increases because additional electron shells are added as we move down the group. Each new period introduces a new principal energy level, increasing atomic radius. The sheilding affect increases due to inner electrons repelling outer electrons from the nucleus. Although nuclear charge increases, the effect of increased sheilding is greater, so the outer electrons feel less effective nuclear attraction. Therefore, the outermost electrons are held less tightly, causing the atomic size to increase down the group 2, as in all the other groups.
Write a balanced chemical equation, including state symbols, for the reaction of:
i) strontium with oxygen ii) strontium oxide with hydrochloric acid
i) $2Sr_{(s)} + O_{2(g)} \longrightarrow 2SrO_{(s)}$
ii) $SrO_{(s)} + 2HCl_{(aq)} \longrightarrow SrCl_{2(aq)} + H_2O_{(l)}$
Barium reacts vigorously when added to water:
i) Write a balanced chemical equation, including state symbols, for the reaction of barium with water.
ii) State two observations that could be made during this reaction.
iii) Suggest the aproximate pH of the resulting solution.
iv) Will the reaction be more or less vigorous than the reaction of barium with water?
i) $Ba_{(s)} + 2H_2O_{(l)} \longrightarrow Ba(OH)_{2(aq)} + H_{2(g)}$
ii) Observations during the reaction are rapid effervescence due to the evolution of hydrogen gas. and solution becomes warms and clear as barium hydroxide dissolves.
(iii) The solution is strongly alkaline with pH of about 12-13
(iv) The reaction of barium with water is more vigorous than that of calcium. This because barium is lower down Group 2 has a larger atomic radius, and lower ionization energy, so it loses electrons more easily and reacts faster with water.
Describe what you would see when magnesium reacts with cold water and steam. Also write an equation for the reaction with steam.
Magnesium reacts very slowly with cold water, only few bubbles of hydrogen gas are seen. The reaction is slow because magnesium is protected by a thin layer of magnesium oxide on its surface. Magnesium reacts vigorously with steam, a bright white flame is observed as magnesium burns, hydrogen gas is realesed. $Mg_{(s)} + H_2O_{(g)} \longrightarrow MgO_{(s)} + H_{2(g)}$
Which one of three compounds listed will decompose at the lowest temperature:
i) Calcium carbonate, strontium carbonate, barium carbonate
ii) Barium nitrate, calcium nitrate, magnesium nitrate.
i) Calcium Carbonate ($CaCO_3$) will decompose at the lower temperature. Thermal stability of Group 2 carbonates increases down the group because the metal ion size increases and its polarizing power decreases. $BaCO_3 > SrCO_3 > CaCO_3$. So, $CaCO_3$ is the least stable and therefore decomposes most easily at lowest temperature.
(ii) Magnesium nitrate ($Mg(NO_3)_2$) will decompose at the lower temperature. For Group 2 nitrates, thermal stability increases down the group because the metal ion size increases and its polarizing power decreases. $Ba(NO_3)_2 > Ca(NO_3)_2 > Mg(NO_3)_2$. Since Mg^{2+} is the smallest ion and has the highest polarizing power, it destabilizes the nitrate ion most, so magnesium nitrate decomposes most easily.
Write a balanced chemical equation, including state symbols, for the thermal decomposition:
i) Strontium Carbonate ii) Barium Nitrate
i) $SrCO_{3(s)} \xrightarrow{\Delta} SrO_{(s)} + CO_{2(g)}$
ii) $2Ba(NO_3)_{2(s)} \xrightarrow{\Delta} 2BaO_{(s)} + 4NO_{2(g)} + O_{2(g)}$

Explain why magnesium sulfate is more soluble than barium sulfate by referring to the relatives values of the lattice energies and enthalpy changes of hydration.
Magnesium sulfate is more soluble and barium sulfate is less soluble because of the balance between lattice energy and enthalpy of hydration. Magnesium ions (Mg^{2+}) are smaller than barium ions (Ba^{2+}), so Mg^{2+} has a much higher (more negative) enthalpy of hydration. This means Mg^{2+} is strongly attracted to water molecules, releasing a large amount of energy when hydrated. This high hydration energy is sufficient to overome the lattice energy of magnesium sulfate, so it dissolves easily in water. Although magnesium sulfate also has a higher lattice energy than barium sulfate (due to the smaller Mg^{2+} ion), the hydration enthalpy decreases much more sharply down the group than the lattice energy. For barium sulfate, the hydration enthalpy of Ba^{2+} is much less negative and cannot compensate for the lattice energy, making the overall enthalpy of solution more endothermic. Therefore, $MgSO_4$ is soluble, while $BaSO_4$ is almost insoluble in water.
Why beryllium form complex compounds?
Beryllium forms complex compounds mainly because of its small size and high charge density (Be^{2+}). These properties make Be^{2+} highly polarizing and a strong Lewis acid, meaning it can attract electron pairs from Lewis bases (such as H_2O, OH^-, F^-, or NH_3). High charge-to-radius ratio Be^{2+} strongly attracts lone pair of electrons from donor atoms Forms tetrahaderal complexes to complete its octet, e.g. $[Be(H_2O)_4]^{2+}$, $[Be(OH)_4]^{2-}$, $[BeF_4]^{2-}$ Its amphoteric nature also comes from this complex formation (reacts with acids and bases). So, beryllium readily forms stable complexes, unlike the heavier group 2 metals, whose larger ions have lower charge density and weaker Lewis acidity. For example,
$Be^{2+} + 4H_2O \longrightarrow [Be(H_2O)_4]^{2+}$
Starting from calcium carbonate how would you extract pure calcium metal.
Calcium carbonate is first heated strongly to form calcium oxide and carbon dioxide.
$CaCO_{3(s)} \xrightarrow{\Delta} CaO_{(s)} + CO_{2(g)}$
The calcium oxide is then converted into calcium chloride by reacting it with HCl.
$CaO_{(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + H_2O_{(l)}$
Finally, molten calcium chloride is electrolyzed to obtain pure calcium metal at the cathode. It is done at high temperature (~850°C) because calcium must be molten.
$CaCl_2 \xrightarrow{\text{electrolysis}} Ca_{(s)} + Cl_{2(g)}$
This process gives pure calcium metal.