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The s-Block Elements Class 11 Notes | CBSE Chemistry Chapter 9

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The s-Block Elements Class 11 Notes | CBSE Chemistry Chapter 9

The s-Block Elements is Chapter 9 of CBSE Class 11 Chemistry — the chapter where the periodic table finally starts to feel predictable. It covers Group 1 (the alkali metals: Li, Na, K, Rb, Cs, Fr) and Group 2 (the alkaline earth metals: Be, Mg, Ca, Sr, Ba, Ra), whose outermost electron sits in an s-orbital. Once you see how one electronic configuration drives every property, the trends almost memorise themselves.

By the end of these notes you will be able to write electronic configurations, predict trends in atomic radii, ionization enthalpy and hydration enthalpy, explain the anomalous behaviour of Li and Be and the diagonal relationship, balance reactions with water, oxygen and hydrogen, and recall every important compound (NaOH, Na₂CO₃, NaHCO₃, CaO, CaCO₃, plaster of Paris) along with the biological roles of Na, K, Mg and Ca. This is a high-scoring chapter carrying roughly 6–8 marks in boards and a steady source of NEET/JEE inorganic questions.

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Key Concepts

1. What Are s-Block Elements?

The **s-block** consists of the elements in which the last electron enters the outermost **s-orbital**. Because an s-orbital holds at most two electrons, the block is just two groups wide.

- **Group 1 — Alkali metals:** Li, Na, K, Rb, Cs, Fr. General configuration **ns¹**.
- **Group 2 — Alkaline earth metals:** Be, Mg, Ca, Sr, Ba, Ra. General configuration **ns²**.

They are called “alkali” metals because their hydroxides are strong alkalis, and “alkaline earth” metals because their oxides are alkaline and were found in the earth’s crust. Hydrogen, despite being 1s¹, is *not* a true alkali metal — it is a non-metal placed separately.

2. Electronic Configuration

Every s-block element has a noble-gas core plus one or two outer s-electrons, which it readily loses to form +1 or +2 ions.

Element	Configuration	Common Ion
Li	[He] 2s ¹	Li ⁺
Na	[Ne] 3s ¹	Na ⁺
K	[Ar] 4s ¹	K ⁺
Be	[He] 2s ²	Be ²⁺
Mg	[Ne] 3s ²	Mg ²⁺
Ca	[Ar] 4s ²	Ca ²⁺

Key idea: Losing the loosely held ns electron(s) gives a stable noble-gas configuration, which is why these metals are so reactive and almost always found as +1 or +2 ions in nature.

3. Trends in Atomic and Ionic Radii

Down a group, a new shell is added at each step, so both atomic and ionic radii **increase** down Group 1 and Group 2.

- Alkali metals have the **largest atomic radii** in their respective periods.
 - For the same period, a Group 2 atom is **smaller** than the corresponding Group 1 atom, because Group 2 has a higher nuclear charge pulling the same shell inward.
 - Cations are always smaller than their parent atoms ($\text{Na}^+ < \text{Na}$) because a whole shell is lost and the remaining electrons feel a stronger effective nuclear charge.
-

4. Trends in Ionization Enthalpy

Ionization enthalpy is the energy needed to remove the outermost electron from a gaseous atom. Since the outer electron gets farther from the nucleus down a group, ionization enthalpy **decreases** down both groups.

- Alkali metals have the **lowest ionization enthalpies** in their periods — they lose one electron very easily.
 - Group 2 elements have **higher first ionization enthalpies** than Group 1 (smaller size, higher nuclear charge).
 - However, the **second ionization enthalpy** of Group 2 is far lower than that of Group 1, because removing a second electron from Group 1 means breaking into a stable noble-gas core — so Group 2 readily forms +2 ions.
-

5. Hydration Enthalpy

Hydration enthalpy is the energy released when one mole of gaseous ions dissolves in water and gets surrounded by water molecules. Smaller, more highly charged ions are hydrated more strongly.

- Hydration enthalpy **decreases** down a group as ionic size increases: $\text{Li}^+ > \text{Na}^+ > \text{K}^+ > \text{Rb}^+ > \text{Cs}^+$.
 - Group 2 ions (M^{2+}) have larger hydration enthalpies than Group 1 ions (M^+) of comparable size because of their higher charge.
 - Li^+ is the most heavily hydrated ion**, so hydrated Li^+ is effectively the largest — that is why lithium is the *least* mobile alkali metal ion in solution despite being the smallest atom.
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6. Physical Properties

s-Block metals are soft, light, silvery and excellent conductors, with low melting and boiling points that fall down the group as the metallic bond weakens.

- Softness:** alkali metals are so soft they can be cut with a knife; Group 2 metals are harder.
 - Density:** Li, Na and K are lighter than water; Group 2 metals are denser than Group 1.
 - Flame colour:** their loosely held electrons are easily excited, giving characteristic flame colours — Li (crimson red), Na (golden yellow), K (lilac/violet), Ca (brick red), Sr (crimson), Ba (apple green). Be and Mg do not impart colour (electrons too tightly bound).
-

7. Chemical Reactivity

Reactivity **increases** down both groups because ionization enthalpy falls. Group 1 metals are more reactive than Group 2 metals.

Reaction with Water

Alkali metals react vigorously with water to give a hydroxide and hydrogen gas:



The reaction gets more violent down the group (K catches fire, Cs explodes). Group 2 metals react less readily — Be does not react, Mg reacts only with hot water/steam, while Ca, Sr and Ba react with cold water:



Reaction with Oxygen

The product of burning depends on the metal — a favourite exam point:

- Lithium** forms mainly the **oxide** (Li_2O).
- Sodium** forms mainly the **peroxide** (Na_2O_2).
- K, Rb, Cs** form **superoxides** (e.g. KO_2).

- Group 2 metals form normal oxides (MO); Ba can form BaO₂.

Reaction with Hydrogen

Except Be, all form ionic (saline) hydrides on heating:



These hydrides contain the H⁻ ion and act as strong reducing agents.

8. Anomalous Behaviour of Lithium and Beryllium

The first member of each group differs from the rest because of its **very small size, high charge density (high polarising power), and absence of d-orbitals**.

Anomalies of Lithium

- Li is much harder, with the highest melting/boiling point of the alkali metals.
- LiCl is deliquescent and somewhat covalent; Li forms only the oxide (not peroxide/superoxide).
- Li₂CO₃, LiOH, LiF and Li₃PO₄ are sparingly soluble — unlike the very soluble salts of Na, K.
- Li reacts directly with N₂ to form Li₃N (other alkali metals do not).

Anomalies of Beryllium

- Be is harder, has a high melting point and forms predominantly **covalent** compounds (e.g. BeCl₂ is covalent and polymeric).
 - Be does not react with water; its oxide and hydroxide are **amphoteric**, while those of other Group 2 metals are basic.
 - Be forms complexes such as [BeF₄]²⁻ owing to its small size and high charge.
-

9. Diagonal Relationship

Because moving right increases charge density and moving down decreases it, the first element of a group resembles the **second element of the next group** placed diagonally. This is the **diagonal relationship**.

- Li resembles Mg:** both form nitrides with N₂, their carbonates decompose on heating, both have sparingly soluble fluorides and carbonates, and both form covalent organometallic compounds.
- Be resembles Al:** both are amphoteric, both form covalent halides that dimerise/polymerise, and both are passivated by concentrated HNO₃.

The cause is the **similar charge-to-radius ratio (polarising power)** of the diagonally placed elements.

10. Important Compounds of Sodium

Sodium Hydroxide (NaOH) — Caustic Soda

Made by the electrolysis of brine (Castner–Kellner process). It is a deliquescent white solid, strongly alkaline, used in soap, paper, and the petroleum industry.



Sodium Carbonate ($\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$) — Washing Soda

Made by the **Solvay (ammonia-soda) process**. Used in water softening, glass, soap and detergents. On heating, washing soda loses water of crystallisation to give the white anhydrous powder (soda ash).



Sodium Bicarbonate (NaHCO_3) — Baking Soda

A mild, non-toxic alkali used in baking (releases CO_2), as an antacid, and in fire extinguishers. It is a Solvay-process intermediate.

11. Important Compounds of Calcium

Calcium Oxide (CaO) — Quick Lime

Made by heating limestone: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$. Reacts vigorously with water (slaking) to give slaked lime, Ca(OH)_2 . Used in cement, mortar and as a drying agent.

Calcium Carbonate (CaCO_3) — Limestone/Marble

Occurs as marble, chalk and limestone. Used in cement manufacture, as a building stone, and an antacid. It decomposes on heating to give CaO.

Plaster of Paris — $\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$

Made by heating gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$) to about 393 K (120 °C):



When mixed with water it sets into a hard mass of gypsum, expanding slightly — which is why it is used for plaster casts (fractured bones), moulds, statues and blackboard chalk.

12. Biological Importance

s-Block ions are essential to life — a small but exam-frequent topic.

- **Sodium (Na^+):** the chief extracellular cation; controls fluid balance, blood pressure, and nerve-impulse transmission (the sodium–potassium pump).
- **Potassium (K^+):** the chief intracellular cation; needed for nerve transmission, muscle contraction and enzyme activation.
- **Magnesium (Mg^{2+}):** the central metal ion in **chlorophyll**; activates many enzymes and stabilises DNA/ATP.
- **Calcium (Ca^{2+}):** builds bones and teeth (as phosphate), and is vital for blood clotting, muscle contraction and nerve signalling.

Weightage in Board & Entrance Exams

Exam	Typical Weightage	Most-Tested Areas
CBSE Board (Class 11)	6–8 marks	Periodic trends, anomalies of Li/Be, diagonal relationship, important compounds
JEE Main	1–2 questions	Reactions with O_2 /water, oxides vs peroxides vs superoxides, properties of compounds
NEET	1–2 questions	Biological importance, hydration enthalpy, flame colours, plaster of Paris

[TABLE: Question-type split — VSA (1 mark): definitions, flame colours, formulae; SA (2–3 marks): trends, anomalies, diagonal relationship; LA (5 marks): preparation/uses of NaOH , Na_2CO_3 , plaster of Paris with equations.]

Important Definitions

Term	Definition
s-Block element	Element whose last electron enters the outermost s-orbital (Groups 1 and 2)
Alkali metals	Group 1 elements (ns^1) whose hydroxides are strong alkalis
Alkaline earth metals	Group 2 elements (ns^2) whose oxides are alkaline
Ionization enthalpy	Energy needed to remove the outermost electron from a gaseous atom
Hydration enthalpy	Energy released when one mole of gaseous ions is surrounded by water
Diagonal relationship	Resemblance of an element to the diagonally placed element of the next group (Li–Mg, Be–Al)
Superoxide	Compound containing the O_2^- ion, e.g. KO_2
Slaking of lime	Reaction of CaO with water to give $Ca(OH)_2$
Plaster of Paris	$CaSO_4 \cdot \frac{1}{2}H_2O$, made by partially dehydrating gypsum
Polarising power	Ability of a cation to distort an anion; high for small, highly charged ions

Solved Examples

Example 1

Write the products formed when Li, Na and K are each burnt in excess oxygen.

Answer: $Li \rightarrow Li_2O$ (oxide); $Na \rightarrow Na_2O_2$ (peroxide); $K \rightarrow KO_2$ (superoxide). The tendency to form larger anions increases as the cation gets larger.

Example 2

Why is lithium the strongest reducing agent among alkali metals in aqueous solution despite having the highest ionization enthalpy in the group?

Answer: Reducing power in solution depends on the standard electrode potential, which includes the very high **hydration enthalpy** of the small Li^+ ion. The large energy released on hydrating Li^+ outweighs its high ionization enthalpy, making Li the strongest reducing agent in water.

Example 3

Identify the gas evolved and the salt formed when sodium reacts with water. Write the equation.

Answer: Gas evolved is **hydrogen (H_2)** and the salt is sodium hydroxide. $2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2\uparrow$.

Example 4

Give two reasons for the diagonal relationship between lithium and magnesium.

Answer: (i) Both have nearly the same charge-to-radius ratio (polarising power). (ii) Both react directly with N_2 to form nitrides (Li_3N , Mg_3N_2) and both have sparingly soluble carbonates that decompose on heating.

Example 5

How is plaster of Paris prepared, and why must the temperature be controlled?

Answer: Gypsum is heated to $\sim 393\text{ K}$: $2(\text{CaSO}_4 \cdot 2\text{H}_2\text{O}) \rightarrow 2(\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}) + 3\text{H}_2\text{O}$. If heated above 393 K , all water is lost giving anhydrous CaSO_4 (dead burnt plaster), which no longer sets with water.

Example 6

Arrange Li^+ , Na^+ , K^+ in increasing order of (a) ionic radius and (b) hydration enthalpy.

Answer: (a) Ionic radius: $\text{Li}^+ < \text{Na}^+ < \text{K}^+$. (b) Hydration enthalpy: $\text{K}^+ < \text{Na}^+ < \text{Li}^+$ (smaller ion is hydrated more strongly).

Important Questions for Board Exams

1-Mark Questions (VSA)

1. Why are Group 1 elements called alkali metals?
2. Name the alkali metal that forms a nitride directly with nitrogen.
3. What is the flame colour imparted by potassium?
4. Write the formula of plaster of Paris.
5. Which Group 2 element forms an amphoteric oxide?

2–3-Mark Questions (SA)

1. Explain why ionization enthalpy decreases down Group 1 but the second ionization enthalpy of Group 2 is much lower than that of Group 1.
2. Discuss the anomalous behaviour of beryllium in Group 2 with two examples.
3. What is the diagonal relationship? Illustrate it with the Li–Mg pair.
4. Why is Li^+ the most heavily hydrated alkali metal ion? What is one consequence?

5-Mark Questions (LA)

1. Describe the manufacture of sodium carbonate by the Solvay process, giving the equations and two uses.
 2. How is plaster of Paris prepared from gypsum? Give the equation, its setting reaction and two uses.
 3. Compare Group 1 and Group 2 elements with respect to atomic radii, ionization enthalpy, reactivity with water, and nature of oxides.
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Quick Revision Points

- Group 1 = ns^1 (alkali metals); Group 2 = ns^2 (alkaline earth metals)
 - Down the group: atomic/ionic radii $\uparrow$, ionization enthalpy $\downarrow$, reactivity $\uparrow$, hydration enthalpy $\downarrow$
 - Burning in O_2 : Li $\rightarrow$ oxide, Na $\rightarrow$ peroxide, K/Rb/Cs $\rightarrow$ superoxide
 - Reaction with water: $2Na + 2H_2O \rightarrow 2NaOH + H_2$; Be no reaction, Mg with steam, Ca/Sr/Ba with cold water
 - Anomalies: Li (covalent LiCl, forms Li_3N , sparingly soluble salts); Be (covalent, amphoteric oxide, forms complexes)
 - Diagonal relationship: Li $\sim$ Mg, Be $\sim$ Al (same polarising power)
 - Key sodium compounds: NaOH (caustic soda), $Na_2CO_3 \cdot 10H_2O$ (washing soda), $NaHCO_3$ (baking soda)
 - Key calcium compounds: CaO (quick lime), $CaCO_3$ (limestone), $CaSO_4 \cdot \frac{1}{2}H_2O$ (plaster of Paris)
 - Biology: $Na^+/K^+ \rightarrow$ nerve impulses; $Mg^{2+} \rightarrow$ chlorophyll; $Ca^{2+} \rightarrow$ bones, clotting
 - Flame colours: Li crimson, Na golden yellow, K lilac, Ca brick red, Ba apple green
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Next Chapter: Chapter 8 — Hydrogen | Continue to Chapter 10 — The p-Block Elements



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