SENIOR FIVE TERM 2

TOPIC 1/3: Periodicity 1

Topic competency: The learner analyses the trends and periodic properties of elements, to explain and predict the reactivity and properties of elements in the Periodic Table.

The Periodic Table (PT)

The Periodic Table is a tabular arrangement such that vertical columns contain elements with electronically similar atoms. The horizontal rows of the Periodic Table are called **periods** while the vertical columns are called **groups**. There are 7 periods and 8 groups.

Some special groups

Group I (1A)

This consists of the elements, H, Li, Na, K, Rb, Cs, and Fr; except hydrogen, these elements are highly metallic in character and are known as **alkali metals**.

Group II (2A)

This consists of the elements, Be, Mg, Ca, Sr, Ba and Ra. All these elements are metals but their metallic character is less than that of group 1. These elements are known as **alkaline earth metals.**

Group VII (7B)

This consists of elements, F, Cl, Br, I and At. All these elements are nonmetals and are known as **halogens**.

Group VIII (0)

Consist of the elements, He, Ne, Ar, Kr, Xe and Rn. These elements are non-reactive and are known as the **noble elements or inert elements or zero element** and are all gaseous elements at room temperature.

Summary

The extreme left of the Periodic Table consists of metals with exception of hydrogen; the extreme right end consists of non-metals. The middle part of Periodic Table consists of semi-metals or metalloid.

Fig. 4.1. Simplified PT to show the variation of properties

Therefore, the metallic characters of the elements decrease across the period, i.e., (left - right).

The periodic table commonly used by Uganda National Examinations Board is given in figure 4.2;

1	2											3	4	5	6	7	8
1.0 H 1																1.0 H 1	4.0 He 2
6.9 Li 3	9.0 Be 4											10.8 B 5	12.0 C 6	14.0 N 7	16.0 O 8	19.0 F 9	20.2 Ne 10
23.0 Na 11	24.3 Mg 12											27.0 Al 13	28.1 Si 14	31.0 P 15	32.1 S 16	35.4 Cl 17	40.0 Ar 18
39.1 K 19	40.1 Ca 20	45.0 Sc 21	47.9 Ti 22	50.9 V 23	52.0 Cr 24	54.9 Mn 25	55.8 Fe 26	58.9 Co 27	58.7 Ni 28	63.5 Cu 29	65.7 Zn 30	69.7 Ga 31	72.6 Ge 32	74.9 As 33	79.0 Se 34	79.9 Br 35	83.8 Kr 36
85.5 Rb 37	87.6 Sr 38	88.9 Y 39	91.2 Zr 40	92.9 Nb 41	95.9 Mo 42	98.9 Tc 43	101 Ru 44	103 Rh 45	106 Pd 46	108 Ag 47	112 Cd 48	115 In 49	119 Sn 50	122 Sb 51	128 Te 52	127 I 53	131 Xe 54
133 Cs 55	137 Ba 56	139 La 57	178 Hf 72	181 Ta 73	184 W 74	186 Re 75	190 Os 76	192 Ir 77	195 Pt 78	197 Au 79	201 Hg 80	207 Ti 81	207 Pb 82	209 Bi 83	209 Po 84	210 At 85	222 Rn 86
223 Fr 87	226 Ra 88	227 Ac 89															
			139 La 57	140 Ce 58	141 Pr 59	144 Nd 60	147 Pm 61	150 Sm 62	152 Eu 63	157 Gd 64	159 Tb 65	162 Dy 66	165 Ho 67	167 Er 68	169 Tm 69	173 Yb 70	175 Lu 71
			227 AC 89	232 Th 90	231 Pa 91	238 U 92	237 Np 93	244 Pu 94	243 Am 95	247 Cm 96	247 Bk 97	251 Cf 98	254 Es 99	257 Fm 100	256 Md 101	254 No 102	260 Lw 103

Fig. 4.2 Periodic Table commonly used by Ugand National Examinations Board

Trends in physical properties of elements in the periodic table

The physical properties to be considered will include:

1. Electronegativity/Electropostivity.
2. Metallic character.
3. Atomic sizes/atomic radius.
4. Ionic size/ionic radius.
5. Ionization energy.
6. Electron affinity.
7. Melting points/boiling points.

Electronegativity

In a covalent bond between unlike atoms the shared pair of electrons will not be shared equally, because the atom with a greater attraction to electrons will pull the shared electron toward itself.

Electronegativity is the measure of an atom's ability to attract shared electrons in a chemical bond

The electronegativity of elements increases across each period, e.g.

Table 4.1 Variation of electronegativity across the second period of the Periodic Table

Elements of 2nd period:	**Li**	Be	B	C	N	O	F
Electronegativity (arbituary units):	1.0	1.5	2.0	2.5	3.0	3.5	4.0

NB: Fluorine, the most electronegative elements, is given an arbitrary value of 4 and the electronegativity of the atoms of other elements is related to it.

Reasons

Across a period, electronegativity increases because atomic size decrease while the nuclear charge increase thus pulling the electrons more strongly, table 4.1

Down a group, electronegativity decreases because atomic size and screening effects increase, the valence electrons become increasing far from the nucleus and hence experience less attraction, table 4.2.

Table 4.2 Variation of electronegativity down group 1 of the Periodic Table

Elements of group 1A:	Li	Na	K	Rb	Cs
Electronegativity (arbituary units):	1.0	0.9	0.8	0.8	0.7

Factors affecting electronegativity

(i) **Atomic size:** Electronegativity decreases as atomic size increases. Larger atoms have valence electrons farther from the nucleus, weakening the attraction for shared electrons.

(ii) **Nuclear charge:** A higher nuclear charge (more protons) increases electronegativity because it creates a stronger positive pull on electrons.

(iii) **Shielding effect:** This is the reduction of the nuclear pull on valence electrons due to inner-shell electrons. A greater shielding effect leads to lower electronegativity.

(iv) **Oxidation state:** Atoms with higher oxidation states have greater electronegativity. The increased positive charge on the atom pulls electrons more strongly.

(v) **Electronic configuration:** Atoms that are closer to a stable noble gas configuration, like those with nearly filled valence shells, have higher electronegativity.

(vi) **Hybridization:** The s-character in an atom's hybrid orbital affects its electronegativity. Orbitals with more s-character (e.g., sp>sp2>sp3 s) are closer to the nucleus, resulting in higher electronegativity.

Trial 4.1

(b) Explain what is meant by the term electronegativity

State the factors that determine the value of electronegativity of an element

(c) Explain how the following factors affect the value of electronegativity of an element

(i) atomic radius,

(ii) nuclear charge,

(iii) the screening effect of the inner electrons

Electropositivity

This is the measure of the ease of an atom of an element to lose valence electrons.

Across the period, electropositivity decreases due to an increase in the effective nuclear attraction on the valence electrons shells.

Down the group, electropositivity increases due to the decrease in effective nuclear attraction on the valence electrons as the atomic sizes and screening effect increase. Thus, group IA elements are the most electropositive elements.

Trial 4.2

Explain the variations in the electropositivity of the following elements

(i) *C, Ge and Sn*
(ii) *Mg, Al, P and Cl* *(8marks)*

Metallic character

Metallic character refers to how readily an element can lose electrons to form positive ions (cations).

Periodic Trends

Across a period (left → right): Metallic character decreases.

Reason: Ionization energy increases and atoms hold onto electrons more tightly.

Example: Sodium (Na, metallic) → Aluminum (Al, less metallic) → Silicon (Si, metalloid).

Down a group (top → bottom): Metallic character increases.

Reason: Atomic size increases, ionization energy decreases, so electrons are lost more easily.

Example: Lithium (Li, less metallic) → Cesium (Cs, highly metallic).

Atomic radii

Atomic radius is the distance between the center of the nucleus and furthest electron under the influence of a nuclear charge of an atom. Or **Atomic radius** is the distance from the nucleus of an atom to the outermost electron shell.

Periodic Trends

Across a period (left → right): Atomic radius decreases.

Reason: Nuclear charge increases, pulling electrons closer, while shielding remains nearly constant.

Example: Na > Mg > Al > Si > P > S > Cl.

Down a group (top → bottom): Atomic radius increases.

Reason: New electron shells are added, increasing distance from the nucleus despite higher nuclear charge.

Example: F < Cl < Br < I.

Variation of atomic radii in transition elements

From Sc to Mn: Atomic radii **decrease gradually**.

Reason: Increasing nuclear charge pulls electrons closer to the nucleus while shielding is minimal because d-electrons are not very effective at screening.

From Fe to Cu: Atomic radii remain **almost constant**.

Reason: Added d-electrons provide significant shielding, counteracting the increased nuclear charge. The attraction between the nucleus and outer 4s electrons is reduced.

At Zn (end of series): Atomic radius shows a **slight increase**.

Reason: Fully filled d-orbitals ($3d^{10}$) shield the 4s electrons strongly, reducing nuclear pull.

Trial 4.3

1. *(a) Explain what is meant by atomic radius* *(4marks)*

   (b) Describe and discuss the change in radii in the following series

   (i) Na, Mg, Al, Si, P, S and Cl *(7 marks)*

   (ii) F, Cl, Br and I *(4 marks)*

   (iii) Ti, V, Cr, Fe, Co *(4 marks)*

2. *(a) What is meant by the term '**atomic radius**'?* *(4marks)*

   (b) Explain how atomic radii vary

   (i) Along a short period of a Periodic Table

   (ii) Down the group in the periodic table *(8marks)*

3. *The atomic number and atomic radii of some transition metals are given in the table below*

Element	*Sc*	*Ti*	*V*	*Cr*	*Mn*	*Fe*	*Co*	*Ni*	*Cu*	*Zn*
Atomic number	*21*	*22*	*23*	*24*	*25*	*26*	*27*	*28*	*29*	*30*
Atomic radii (Å)	*1.44*	*1.32*	*1.22*	*1.17*	*1.17*	*1.16*	*1.16*	*1.15*	*1.17*	*1.25*

(i) Plot a graph of atomic radii versus atomic atomic number *(4marks)*

(ii) Explain the shape of your graph *(7marks)*

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Ionic radii

The variation of the ionic radii across the period and down the groups is similar to those of the atomic radii as demonstrated by examples in table 4.5 and 4.6 respectively.

Table 4.5 Variation of ionic radii across the second period of the Periodic Table

Ions	Li^+	Be^{2+}	B^{3+}	O^{2-}	F^-
Ionic radius(nm) x10	0.78	0.31	0.20	1.40	1.36

Table 4.6 Variation of ionic radii down group 1 of the Periodic Table

Ions	Li^+	Na^+	K^+	Rb^+	Cs^+
Radii (nm) x 10	0.060	0.095	0.133	0.148	0.165

NB: (a) Atomic radii are very much smaller than anionic radii but very much larger than cationic radii (table 4.8).

Table 4.7 Comparison atomic and ionic radii of cation and anion

Atoms	ions
Na 0.157nm	**Na^+** 0.095 nm (cation)
Cl 0.099n m	**Cl^-** 0.181 nm (anion)

(b) The radii of iso-electronic positive ions, (i.e., ions with the same number of electrons, e.g. both Li^+ and Be^{2+} have the number of electron), decrease with increasing positive charge. There is only a small change in ionic radii for iso-electronic negative ions with a change in negative charge e.g. Cl^- and S^{2-} in table 2.10

Table 4.8 Comparison of ionic radii of iso-electronic ions

Li^+ 0.060nm	S^{2-} 0.184nm
Be^{2+} 0.031nm	Cl^- 0.181nm

Trial 4.4

(a) *Explain how the sizes of ionic radii of group II ions affect the properties of their compounds compared to those formed by group I metals* *(8marks)*

(b) *Explain the following observation*

- (i) *In going from sodium (0.189nm) to chlorine (0.099nm) along period of the periodic table, there is steady decrease in the size of covalent radii.*
- (ii) *Although K^+ and Cl^- ion both have 18 electrons the ionic radius of K^+(0.132nm) while that of Cl^- is (0.181nm)*

Variation ionic radii of transition metals

(i) Since the transition metals show many oxidation states, the ionic radii of the ions of different oxidation states are different.

(ii) Generally, the ionic radii of different cations of the same element in different oxidation state decrease with the increase in oxidation state, e.g.

$$Cr^{2+} > Cr^{3+} > Cr^{4+} > Cr^{5+} > Cr^{6+}$$

This is due to increase in effective nuclear charge.

(iii) Ionic radii of cations of different elements in the same oxidation state generally decrease with the increase in nuclear charge or atomic number e.g.
$Fe^{2+} > Co^{2+} > Ni^{2+} > Cu^{2+}$

Ionization energy (I.E)

This is the energy required to remove one electron in the ground state completely from a gaseous atom or ion. The ionization energy thus, may be classified as first, second, third ionization energy, etc.

$A(g) \longrightarrow A^+(g) + e$ ΔH = 1st I.E

$A^+(g) \longrightarrow A^{2+}(g) + e$ ΔH = 2nd I.E

$A^{2+}(g) \longrightarrow A^{3+}(g) + e$ ΔH = 3rd I.E

Variation of the first ionization energy across the period

The first ionization energies of the first 20 elements of the PT. plotted against their atomic number are shown in Fig.4.2 below:

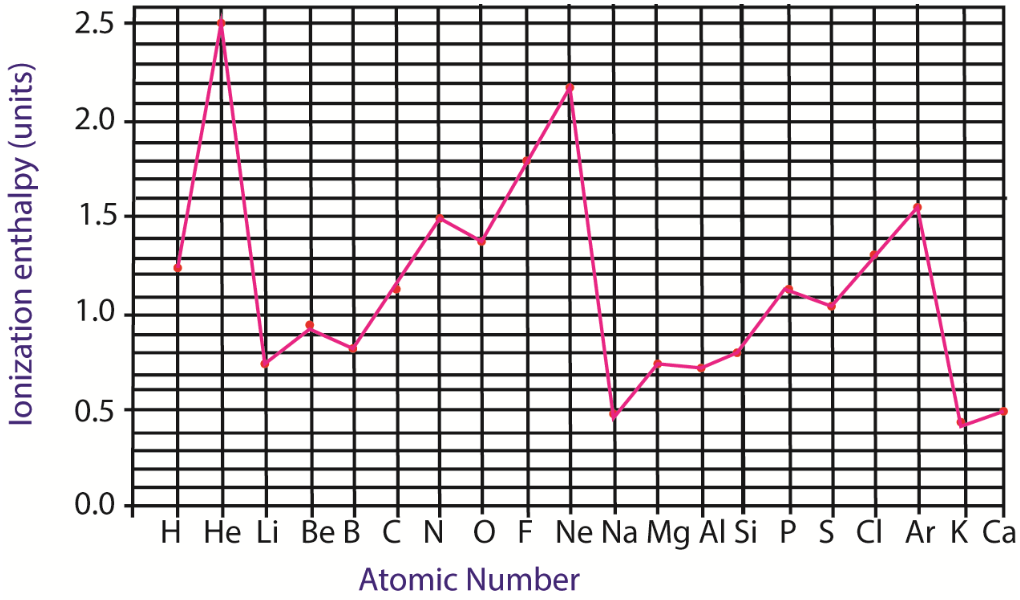


Fig. 4.2. The ionization enthalpies of the first 20 elements

From lithium to neon there is a pronounced increase in ionization energy with slight breaks occurring at positions occupied by boron and oxygen. An exactly similar trend occurs in the portion of the graph from sodium to argon.

Reason for two trends

1. The increase in the first ionization energy across the period is due to increase in the effective nuclear charge on the valence electron as the number of electrons on the outer shell increases.

2. Half filled (ns^1, np^3 nd^5) or completely filled ($ns^2np^6nd^{10}$) orbital are comparatively more stable and hence more energy is needed to remove an electron from them. This explains why the first ionization energy of beryllium (full orbital, $1s^2$ $2s^2$) and nitrogen (half full orbital, $1s^2$ $2s^2$ $2p^3$) are higher than that of boron and oxygen respectively.

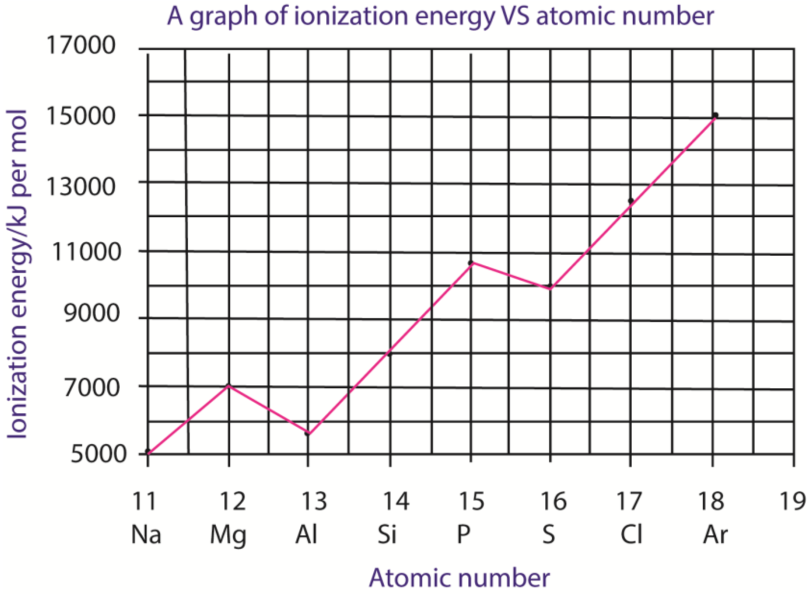
3. The sharp decrease in ionization energy from helium to lithium is explained by the fact that the outer electron in lithium is far from the nucleus than in helium. Also the removal of an electron from helium leads to a less stable half full orbital while removal of an electron from lithium leads to formation of a stable full orbital $1s^2$.

Variation of the first ionization energy down a group

For elements in the same group, the ionization energies decrease down the group, due to increase in the screening effect or decrease in the effective nuclear charge because electros are added to a new shell. For example the first ionization energy of lithium, sodium and potassium are 520, 494, and 418kJmol^{-1} respectively.

Trial 4.5

(a) the graph of first ionization energy against atomic number for the elements in the third short period of the Periodic Table is shown below



(i) What is meant by first ionization energy? Illustrate your answer using the sodium atom. (3 marks)

(ii) Explain the general trend in the first ionization energy of the elements shown in the graph (3marks)

(iii) The first ionization energies of Al and S are lower than expected. explain (5marks)

(b) Describe the trend in the first ionization energy down group I of the periodic Table and explain your answer (5marks)

(c) Show how the first ionization energies of the elements in a group are related to their reactivity. Illustrate your answer using the reaction of group I elements with water (4marks)

Trial 4.5

(a) Define the term ionization energy

(b) The first ionization energies of element in period III of the periodic Table and atomic numbers are given in the table below

	Na	Mg	Al	Si	P	S	Cl	Ar
Ionization Energies ($kJmol^{-1}$)	495	740	580	790	1060	1000	1255	1520
Atomic number	11	12	13	14	15	16	17	18

Plot a graph of ionization energy against atomic number and explain the shape of the graph (9 marks)

(c) The melting points of magnesium, silicon and Sulphur are 650^0C, 1423^0C and 120^0C respectively. Explain the difference in melting points of the elements (6 marks)

(d) (i) Name the type of bonding that exist in the hydrides of Na, P and S (1½ marks)

(e) Write equation to show the reaction, if any, of hydride in (i) with water (2½ marks)

Trial 4.6

The graph in figure 2 shows the variation of the first ionization energies of the elements in the first row of the periodic Tables

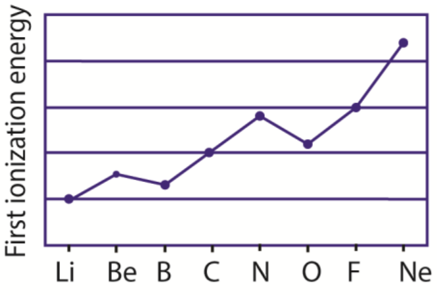


Fig. 2

Explain the following observations

(i) there is a general increase in first ionization energy from lithium to neon (3marks)

(ii) The first ionization energy of beryllium is higher than that of boron.

(iii) The first ionization energy of oxygen is lower than that of nitrogen.

Variation of successive Ionization energy of an atom

Very convincing evidence for arrangement of electrons into definite shells is available from tabulated values of successive ionization energies of atoms; those for the potassium atom are shown in figure 4.3. The ionization energies cover a very wide range of values and for the convenience of graphical plotting of their logarithmic values are used.

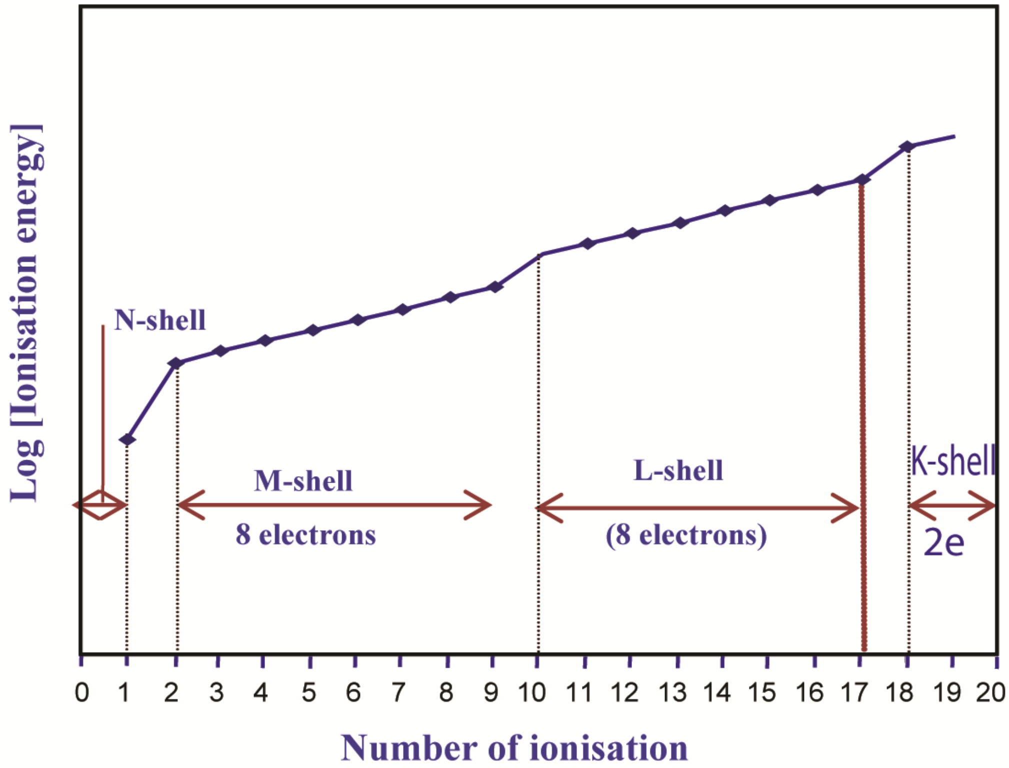


Fig. 4.3. Successive ionization energy of potassium atom

The clear-cut increases in ionization energy are observed when the 2nd, 10th and 18th electrons are involved, which indicate quite clearly that the nucleus of the potassium atom is surrounded by electron grouped into a number of shells. Successive ionization of electrons becomes difficult, since every time an electron is removed; the atom carries one more unit of positive charge.

However, the large increases in ionization energy for removal of the second 2nd, 10th and 18th electrons suggest that the 2nd electron is closer to the nucleus than the 1st electron; similarly the 10th and 18th electrons are nearer to the nucleus than the 9th and 17th electrons respectively. Thus, electrons of a potassium atom can be grouped into shells as follows:

Table 4.9 Arrangement of electrons in potassium atom

SHELL	K	L	M	N
NUCLEUS 19p, 20n	2	8	8	1

The electronic configuration of potassium is thus, written as $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$.

The properties of elements are closely linked to their ionization energies. The magnitude of the ionization energy is the measure of metallic character of the elements. The first ionization energies for metals are nearly all below 800kJ/mol whereas those of non-metal are nearly all above 2000kJmol^{-1}.

Trial 4.7

1. *Explain the following observation*

a) *Although ionization energy generally increase across a period in the periodic table, the ionization energy of boron is less than that of beryllium (5marks)*

b) *Both valence electron of magnesium occupy the same energy levels yet the second ionization energy of magnesium is greater than its ionization energy (3marks)*

c) *Calcium forms compound containing Ca^{2+} ions, but none containing Ca^{+} even though its first ionization energy is lower than the second ionization energy.*

2. (a) The first ionization energies of aluminium are 577, 1816, 2745, 11573kJ mol^{-1}.

Explain the trend in ionization energy. (3marks)

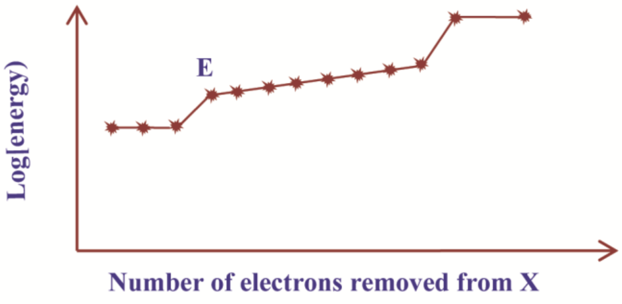
(b) Aluminium forms stable compound in which aluminium has an oxidation state of +3 but not +2 or +1. Explain this observation.

3. (a) *Write electronic configuration of strontium(1mark)*

(b) What is the principal oxidation state of strontium? (1mark)

(c) State whether the first ionization energy of strontium is greater or less than of barium. Explain your answers

4. The diagram below shows successive ionization energies for an element X, showing removal all electrons



(a) Giving reasons state

i. The group of element X

ii. The period of element X

iii. Identify element X (4marks)

(b) Explain the sudden increase in the energy required to remove electron E (2marks)

(c) Explain how the size of X will change as electrons are removed (1marks)

5. (a) the first eight ionization energies of an element B are shown below Ionization Energies/kJmol^{-1}

1st	2nd	3rd	4th	5th	6th	7th	8th
786	1580	3230	4360	16000	20000	23600	29100

(i) Explain what is meant by the term first ionization energy?

(ii) State the factors that determine the value of the first ionization energy and explain how they affect this values

(iii) To which group in the periodic table does element B belong explain your answer

6. The table below shows the ionization energies (in kJmole-1) of five elements lettered A, B, C, D and E

Elements	1st	2nd	3rd	4th
A	500	4600	6900	9500
B	740	1500	7700	10500
C	630	1600	3000	4800
D	900	1800	14800	21000
E	580	1800	2700	11600

(a) Which one of these elements is most likely to form an ion with a charge of +1? Give reason for your answer

(b) State

i. Two elements which belong to the same group in the periodic table

ii. The group to which you have stated in (b)(i) belong

(c) (i) write the formula of chloride of element E

(ii) Write equation for the reaction between the chloride of element E and water.

7. Explain the following

(i) The atomic radius decrease with the increasing atomic number in a period

(ii) Na-atom is larger than both Li and Mg

(iii) The second ionization energy is always higher than the first

(iv) The first ionization of Be is higher than that of B

Electron affinity

This is the energy involved in adding an electron to a *gaseous* atom. It is the energy change for the process.

$A(g) + e^- \rightarrow A^-(g)$ ΔH = electron affinity.

Electron affinities of elements **increase across a period** due to increase in the effective nuclear charge (or electronegativity) and **decreases down a group** due to the decrease in the effective nuclear charge on the valence shell. Halogens have the highest values of electron affinities. However, the electron affinity of fluorine is lower than that of chlorine. This is *due to the fact that the size of fluorine atoms is small, thus the incoming electron is repelled* by the nine electrons that are close to each other in the valence shell.

The successive electron affinities become more and more endothermic because the incoming electron is repelled by the electrons already in an atom; therefore, additional energy is required to add this electron.

Trial 4.8

1. *Explain the difference between electronegativity and electron affinity*
2. *The first electron affinity of group VII elements are given in table below*

Element	*F*	*Cl*	*Br*	*I*
1st electron affinity kJmol^{-1}	*-328*	*-349*	*-325*	*-295*

(i) *Explain what is meant by the term electron affinity*

(ii) *Explain the trend in the first ionization energy*

(iii) *The first and second electron affinity values for oxygen atom are -121kJmol^{-1} and +744kJmol^{-1} respectively. Explain why the second electron affinity is endothermic.*

Melting points of elements

The melting point of a substance or element is the measure of energy required to break down the regular arrangement of atoms, ions or molecules in the crystal (solid state). Therefore, the melting point indicates the strength of the forces knitting atoms, ions or molecules together in the crystal. The strength of those forces in elements varies according to the type of crystals formed and due to this; the melting points of elements in period do not change uniformly, (e.g table 4.10).

Table 4.10. The melting points of the elements of the third period of the Periodic Table

Elements	Na	Mg	Al	Si	P	S	Cl
Configuration	2.8.1	2.8.2	2.8.3	2.8.4	2.8.5	2.8.6	2.8.7
Melting points (ºC)	98	650	660	1423	44	120	-161
	Large increase		Large increase			Large increase	

From sodium to aluminium, the melting point increase due to increase in the strength of metallic bonds. The strength of metallic bonds increase as the number of electrons contributed to form the metallic bond increase. Sodium contains one loosely bound electron in the valence shell, and this electron is readily contributed to the formation of the metallic bond. Magnesium contains two electrons in the valence shell and both electrons are contributed to the formation of a metallic bond.

The metallic bond in magnesium is thus stronger than that in sodium, which explains the large difference in the melting points.

However, the increase in melting points from Mg to Al is not as sharp as expected, probably because only two electrons of aluminum are used in formation of the metallic bond.

Silicon has the highest melting point because it uses its four valence electrons to form an infinite three-dimensional assembly of atoms linked by a single covalent bond; thus, high temperatures are required to break these strong bonds.

In the structure, each silicon atom is tetrahedrally bonded to four silicon atoms with a crystal unit of five silicon atoms (ABCDE) in figure 4.11 (the structure similar to that of diamond). Such a structure is described as a three-dimensional giant structure. As the melting of silicon involves breaking of the strong covalent bonds, the melting point of silicon is thus very high.

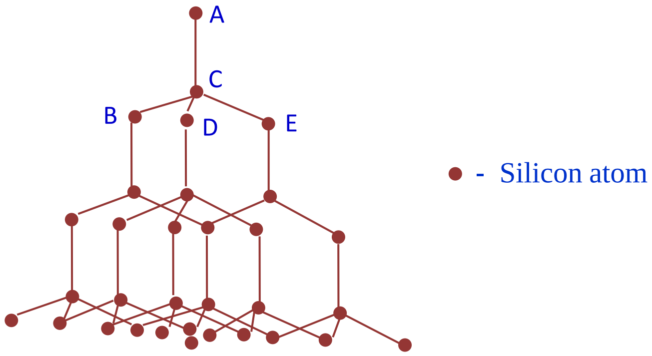


Fig. 4.11 Structure of silicon

The melting point of P, S, Cl is dependent on the size/mass of the molecules formed. The melting points of phosphorus and sulphur are relatively higher than expected because they exist as P_4 and S_8, rather than simples individual atoms; thus, the intermolecular forces are stronger. The boiling point of chlorine is very low because it exists as discrete diatomic molecule, Cl_2 held together by weak van der Waals forces.

Trial 4.8

The atomic numbers and the melting points of some elements in period 3 of the Periodic Table are shown below

Element	Na	Mg	Al	Si	P
Atomic number	11	12	13	14	15
Melting points/^{0}C	98	650	660	1410	44

(a) (i) Plot a graph of melting points against atomic number (3 marks)

(ii) Explain the shape of the graph in (i) (6marks)

(b) Describe and explain how the oxides of magnesium, aluminium and silicon react with

    (v) sodium hydroxide
    (vi) hydrochloric acid (9marks)

(c) state the type of bonding in the oxides of sodium and phosphorus (2marks)

Chemistry of period 3 elements

The chapter compares the chemistry of elements in the same period specifically period 3. Each element in a period provides a representative chemistry of the group into which it exists.

The elements of period 3 are Na, Mg, Al, Si, P, S, Cl.

Electronic structure

For each of these elements the total number of electrons in the outer most shell is equal to the group in which it exists.

Table 4.11. The electron configuration of period 3 elements

Elements	Na	Mg	Al	Si	P	S	Cl	Ar
valence configuration	$3s^1$	$3s^2$	$3s^2 3p^1$	$3s^2 3p^2$	$3s^2$3p3	$3s^2 3p^4$	$3s^2$3p5	$3s^2 3p^6$

Reactions of the period elements

Reaction period 3 elements with water

Reactivity with water decreases across the period because there is a decrease in electropositivity of these elements across the period.

i) Sodium reacts vigorously with cold water.

$$2Na\ (s) + 2H_2O\ (l) \longrightarrow 2NaOH\ (aq) + H_2\ (g)$$

ii) Mg does not react with cold water but with steam to form oxide

$$Mg\ (s) + 2H_2O\ (g) \longrightarrow MgO\ (aq) + H_2\ (g)$$

iii) Al react very slowly to form hydroxides:

$$2Al\ (s) + 6H_2O\ (l) \longrightarrow 2Al(OH)_3\ (s) + 3H_2\ (g)$$

iv) Si, P, S have no reaction with water.

v) Chlorine reacts with water to form chloric (I) acid and HCl.

$$Cl_2\ (g) + H_2O\ (l) \longrightarrow HOCl\ (aq) + HCl\ (aq)$$

Reaction period 3 elements with NaOH

i) Aluminium reacts with aqueous NaOH to produce hydrogen and a complex salt -sodium aluminate.

$2Al\ (s) + 2NaOH\ (aq) + 6H_2O\ (l) \rightarrow 2NaAl(OH)_4\ (aq) + 3H_2\ (g)$

Or $Al\ (s) + 2OH^-\ (aq) \rightarrow AlO^{2-}(aq) + H_2(g)$

This is because aluminium is amphoteric, i.e., possesses both metallic and nonmetallic properties.

ii) Chlorine reacts to form sodium chlorate (I), sodium chloride and water.

$Cl_2\ (g) + 2NaOH\ (aq) \rightarrow NaOCl\ (aq) + NaCl\ (aq) + H_2O\ (l)$

Sodium hypoochlorate (I) is used as a bleaching and antiseptic agent such as Jik (Fig. 6.1)

iii) Phosphorus reacts with hot concentrated NaOH to produce Phosphine

$P_4\ (s) + 3^-OH\ (aq) + 3H_2O\ (l) \rightarrow PH_3\ (g) + 3H_2PO_2^-\ (aq)$

iv) Na, Mg, Si, S have no reaction.

Fig.4.12 sodium chlorate 1 is used in Jik a house hold household detergent

Reaction of period 3 element with HCl

Reactivity decreases across the period due to decrease in metallic properties.

i) Na, Mg, Al react to form H_2 and metal salts.

$2Na\ (s) + 2HCl\ (aq) \rightarrow 2NaCl\ (aq) + H_2\ (g)$

$Mg\ (s) + 2HCl\ (aq) \rightarrow MgCl_2\ (aq) + H_2\ (g)$

$2Al\ (s) + 6HCl\ (aq) \rightarrow 2AlCl_3\ (aq) + 3H_2\ (g)$

ii) Si, P, S, Cl have no reaction.

Oxides of period 3 elements

Table 4.12: The oxide of Period 3 elements

LEMENTS	Na	Mg	Al	Si	P	S	Cl
OXIDES	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_3O_7
PRINCIPAL BONDING	Ionic	Ionic	Ionic-covalent	Covalent	Covalent	Covalent	Covalent
CHARACTER	Basic	Basic	Amphoteric	Acidic	Acidic	Acidic	Acidic
$\text{MP}\backslash{}^{0}\text{C}$	1193	3075	2300	1728	563	30	-91

Trial 4.9

(a) Complete the table below (4 ½ marks)

Element	Formula of oxides	Type of bonding in oxide	Structures of oxide
Al			
Si			
P			

(b) Write an equation for the reaction between the oxide of aluminium and sodium hydroxide (1½)

Physical properties of oxides of period 3 elements

i) Across the period there is a gradual charge in character of oxides from strongly basic to strongly acidic. Al_2O_3 is amphoteric oxide, i.e., possesses both basic and acid properties.

ii) Na_2O, MgO and Al_2O_3 have high melting points due to the strength of ionic bonding. SiO_2, although covalent has high melting point because the crystal of SiO_2 is a giant structure in which silicon atoms and oxygen atoms are bonded together by a single strong covalent bonds.

P_2O_5 (white solid) consists of P_4O_{10} molecules; these contain electric dipoles due to electron displacements in their bonds and are quite strongly attracted to each other. This is why its melting point is high compared with that of SO_3 which consists of simple SO_3 molecules.

Cl_2O_7 has a low melting point because the molecules are held together by the weak van der Waals forces.

Reactions of oxides of period 3 elements

Reaction oxides of period 3 elements with water

- Na_2O and MgO react with water to form hydroxides.

  Na_2O (s) + H_2O (l) ⟶ 2NaOH (aq)

- Al_2O_2 has no reaction with water because it is an insoluble oxide.
- SiO_2 has no reaction with water.
- P_2O_5, SO_2, Cl_2O_7 are acid anhydrides, i.e., react with water to form acid.

  P_2O_5 (s) + H_2O (l) ⟶ $2HPO_3$, phosphoric acid

  Then, HPO_3 + H_2O ⟶ H_3PO_4 phosphoric

  SO_3 (s) + H_2O (l) ⟶ H_2SO_4 (aq), sulphuric

  Cl_2O_7 (g) + H_2O (l) ⟶ $2HClO_4$(aq), perchloric acid

Reaction period 3 elements with NaOH

- Na_2O, MgO are basic, therefore, have no reaction with water.
- Al_2O_3 being amphoteric reacts with dilute NaOH to produce sodium aluminate and water.

  Al_2O_3 (g) + 2NaOH (aq) ⟶ $2NaAlO_2$ (aq) + H_2O (l)

- SiO_2 reacts to form sodium silicate and water.

  SiO_2 (s) + 2NaOH (aq) ⟶ Na_2SiO_3 (aq) + H_2O (l)

- P_2O_5, SO_3 and Cl_2O_7 react to form salts and water

  SO_3 (g) + 2NaOH (aq) ⟶ Na_2SO_4 (aq) + H_2O (l)

  P_2O_5 (s) + 2NaOH (aq) ⟶ $2NaPO_3$ (aq) + H_2O (l)

  Cl_2O_7 (g) + 2NaOH (aq) ⟶ $2NaClO_4$ (aq) + H_2O (l)

Trial 4.10

Write an ionic equation for the reaction between sodium hydroxide and (1 ½ mark each)

(a) Silicon (IV) oxide
(b) Lead (II) oxide
(c) Aluminium oxide

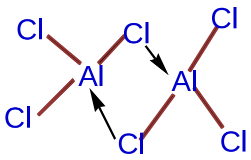
Chloridesof period 3 elements

Table 4.13: The chlorides of period 3 elements

ELEMENTS	Na	Mg	Al	Si	P	S	Cl
CHLORIDES	NaCl	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_5	S_2Cl_2	Cl_2
BONDING	Ionic	Ionic	Covalent	Covalent	Covalent	Covalent	Covalent
STATES	Solid	Solid	Solid	Liquid	Solid	Liquid	Gas
MP\\$^0$C	808	714	192	-68	160	-76	

Physical properties of period 3 chlorides

NaCl and $MgCl_2$ have high melting points due to the strength of the ionic bonding. $AlCl_3$ has fairly high melting point because in the solid state, it consists of Al_2Cl_6 molecules and not $AlCl_3$. These molecules are produced through dative bonding between Al and Cl in the Al_2Cl_6 molecules.



$SiCl_4$ and S_2Cl_2 are liquid and consist of respective simple $SiCl_4$ and S_2Cl_2 molecules and in the solid state the molecules are held by weak van der Waals' forces that explains their very low melting points. PCl_5 a pale yellow solid has a fairly high melting point because the solid undergoes partial ionization.

$$2PCl_5 \rightleftharpoons PCl_4^+ + PCl_6^-$$

Trial 4.11

Explain the following observations:

(a) Sodium chloride melts at 800^{0}C whereas aluminium chloride sublimes at 180^{0}C. (3marks)

Reactions of chlorides of period 3 element

Reactions of chlorides of period 3 element with water.

i) NaCl and $MgCl_2$ have no reaction with water but simply dissociate into ions.

$NaCl\ (s) + (aq) \longrightarrow Na^+\ (aq) + Cl^-\ (aq)$

$MgCl_2\ (s) + (aq) \longrightarrow Mg^{2+}\ (aq) + 2Cl^-\ (aq)$

ii) $AlCl_3$, $SiCl_4$, PCl_5, Cl_2 are hydrolyzed by water. However, the extent of the hydrolysis of these chlorides varies across the period.

$AlCl_3$ is slightly hydrolyzed liberating HCl and $Al(OH)_3$.

$AlCl_3 (s) + 3H_2O (l) \longrightarrow Al(OH)_3(s) + 3HCl (aq)$

However, when aluminium chloride is added to water, it reacts exothermically to give hydrated aluminium ions, $[Al(H_2O)_6]^{3+}$ and chloride ions, Cl^-. The energy needed to break the Al-Cl covalent bond is derived from the high enthalpy of hydration of small highly charged Al^{3+} ions.

$Al_2Cl_6 (s) + 12 H_2O (l) \longrightarrow 2[Al(H_2O)_6]^{3+}(Cl^-)_3$

covalent ionic

Hydrated aluminium chloride is readily soluble in water yielding

$[Al(H_2O)_6]^{3+} (aq) + 3Cl^- (aq)$ ions

$PCl_5 (s) + 4H_2O (l) \longrightarrow 5HCl (aq) + H_3PO_4 (aq)$ [phosphoric acid]

$SiCl_4$, PCl_5, and S_2Cl_2 are completely hydrolyzed liberating hydrogen chloride.

$SiCl_4 (l) + 3H_2O (l) \longrightarrow 4HCl (aq) + H_2SiO_3 (aq)$ [silicic (IV) acid]

$2S_2Cl_2 + 3H_2O (l) \longrightarrow 4HCl (aq) + H_2SO_3 (aq) + 3S (s)$

Hydrogen chloride under moist conditions appears as white fumes because these chlorides fume in moist air, their hydrolysis leads to formation of HCl. The hydrolysis of each of these chlorides occurs in stages. As an example, lets us consider the hydrolysis of $SiCl_4$.

Trial 4.12

The melting points of the chlorides of some elements are given below.

chlorides	*$MgCl_2$*	*$FeCl_3$*	*PCl_5*
Mp/^{0}C	*712*	*282*	*-112*
Type bonding			

(a) State the type of bonding that exists in each of the chloride in the table above.

(b) State what would be observed and write an equation for the reaction that takes place when water is added to each of the chloride in table above. (2½ marks each)

The hydrides of period 3 elements

The hydrides formed by period 3 elements are NaH, MgH_2, AlH_3, SiH_4, PH_3, H_2S, HCl.

Table 4.14: The properties of period 3 hydrides

ELEMENTS	NaH	MgH_2	AlH_3	SiH_4	PH_3	H_2S	HCl
BONDING	Ionic	Ionic	ionic	Covalent	Covalent	Covalent	Covalent
MP\\$^0$C	804				-134	-86	-144
bp\\0C	1413				-88	-60	-85

NaH and MgH_2 have high melting point and boiling points due to strength of ionic bonds. H_2S and HCl have relatively higher melting points and boiling points than SiH_4 and $PH_{3;}$ becase their molecules are polar molecules and are held together by hydrogen bonds whereas the molecules of SiH_4 and PH_3 are non polar and therefore, held together by the weak van der Waals forces.

Reactions of hydrides of period 3 elements

Reaction of hydrides of period 3 elements with water

Due to the high polarity of the bonds in the hydrides NaH and MgH_2, these hydrides react with water readily producing hydrogen and the metal hydroxide.

$NaH (s) + H_2O (l) \longrightarrow NaOH (aq) + H_2 (g)$

$MgH_2 (s) + 2H_2O (l) \longrightarrow Mg(OH)_2 (aq) + 2H_2 (g)$

AlH_3, SiH_4, and PH_3 have no reaction with water due to lack of polarity in their bonds.

H_2S and HCl dissociate in water producing acids.

$H_2S (g) + 2H_2O (l) \rightleftharpoons H_3O^+ (aq) + HS^- (aq) + H_2O (l) \rightleftharpoons 2H_3O^+ (aq) + S^{2-} (aq)$

$HCl (g) + H_2O (l) \longrightarrow H_3O^+ (aq) + Cl^- (aq)$

Reactions of hydrides of period 3 elements with hydroxides (NaOH)

HCl and H_2S react with hydroxides to form salts.

$HCl\ (g) + {}^{-}OH\ (aq) \longrightarrow Cl^{-}\ (aq) + H_2O\ (l)$

$H_2S\ (g) + {}^{-}OH\ (aq) \longrightarrow HS^{-}\ (aq) + H_2O\ (l)$

then $HS^{-}\ (aq) + {}^{-}OH\ (aq) \longrightarrow S^{2-}\ (aq) + H_2O\ (l)$

Reaction of hydrides of period 3 elements with HCl

NaH and MgH_2 react to form hydrogen and metal salts.

$NaH\ (s) + HCl\ (aq) \longrightarrow NaCl\ (aq) + H_2\ (g)$

$MgH_2\ (s) + 2HCl\ (aq) \longrightarrow MgCl_2 + 2H_2\ (g)$

Trial 4.13

The element contained in the third short period of the periodic Table, given in alphabetic order are; aluminium, argon, chlorine, magnesium, phosphorus, silicon, sodium and Sulphur.

(a) In the table below, write the formula of the hydrides formed by the elements listed, state the oxidation state (or valence) of the elements in these hydrides and classify the bonding in the a hydroxides as ionic or covalent (6marks)

Elements	Formula of hydride	Oxidation(or valence) od the elements in these hydrides	Type of bonding
Aluminium			
Chlorine			
Magnesium			
Phosporus			
Silcon			

(b) The hydrides formed by sodium and Sulphur were separately shaken with water

Write the equation to show the reaction that took place, if any with (3marks)

(i) sodium hydride

(ii) Sulphur hydrides (3marks)

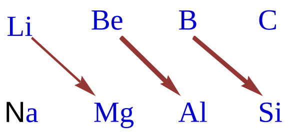
Chemistry of period II elements

Period 2 consists of the following elements.

Table 4.15. Period 2 elements

Element:	Li	Be	B	C	N	O	F	Ne
Electron Configuration	$2s^1$	$2s^2$	$2s^22p^1$	$2s^22p^2$	$2s^22p^3$	$2s^22p^4$	$2s^22p^6$.	$2s^22p^5$

The diagonal relationship



Li and Mg, Be and Al, B and Si have similarities in their chemical properties. Such similarities in chemical properties between the first element and the second element in the next higher group are called diagonal relationship. This kind of relation exists between some elements of period 2 and 3.

Thus: **diagonal relationship** is the similarity in chemical properties between elements in the second period with elements which are lying to their right in 3rd period or lying diagonally opposite to them.

The reason for diagonal relationship is that, the two diagonally related elements (e.g., Li and Mg) have very similar electropositivity. Since electropositivity increases from top to bottom in any periodic group and decreases from left to right across a particular period, the increase in electropositivity in going down one place in the periodic group, (e.g., magnesium is one place lower than lithium) is compensated for by the decrease that occurs in moving one step across a period from left to right (magnesium is one place to the right of lithium).

Alternatively, two diagonally related elements have the same polarizing power since on moving from left to right across a period, the cationic charge increase and cationic size decrease, the magnitude of polarizing power increase and moving down the group, however, cationic size increase reducing the polarizing power. Since these two variations of polarizing power of cations (along the period and other down the group) are opposite each other, they partially cancel (balance) each other when we move diagonally in periodic table.

Alternatively, two diagonally related elements have the same the electronegativity, since an increase in electronegativity one step across the period is canceled by a decrease in electronegativity one step down the group.

Similarities between Li and Mg

The following summaries show the ways in which lithium resembles Magnesium and, differ from other alkali metals.

Li and Mg		**Na, K, Rb and Cs**
1. Both form normal oxides when burn in oxygen. $4Li (s) + O_2 (g) \rightarrow 2Li_2O (s)$ $2Mg (s) + O_2 (g) \rightarrow 2MgO (s)$		Form peroxides, e.g., Na_2O_2 or hyper oxides, e.g., KO_2
2. Their carbonates, hydroxide and peroxide readily decompose to the oxides on heating. $Li_2CO_3 (s) \rightarrow Li_2O (s) + CO_2 (g)$ $MgCO_3 (s) \rightarrow MgO (s) + CO_2 (g)$		no similar decomposition
3. Their nitrates decompose on heating to give oxides, nitrogen dioxide and oxygen. $4LiNO_3 (s) \rightarrow 2Li_2O + 4NO_2 (g) + O_2 (g)$ $2Mg(NO_3)_2 (s) \rightarrow 2MgO + 4NO_2 (g) + O_2 (g)$		Nitrates decompose to give nitrites. $2KNO_3 \rightarrow 2KNO_2 + O_2$
4. Their carbonates and phosphates are insoluble in water.		Corresponding compounds are more soluble
3. Form strongly hydrated ions in solution		Corresponding ions are not hydrated in solution
6. They form carbides and nitride by direct combination with the element. $3Mg (s) + N_2 (g) \rightarrow (Mg^{2+})_3(N^{3-})_2 (s)$ $Mg (s) + 2C (s) \rightarrow Mg^{2+}(C\equiv C)^{2-} (s)$		No reaction with N_2 nor C
7. Halide (except fluorides) are soluble in organic solvents		Corresponding compounds much less soluble
8. Compounds have covalent character. Hydro carbonates known only in solution		Solid Hydro carbonates can be made

Trial 4.14

(i) Explain what is meant by diagonal relationship (3 ½ mark)

(ii) Name five properties in which lithium resemble magnesium but differ from the rest of group 1 metals. Give one reason for the resemblance between lithium and magnesium (10 marks)

Similarities between Be and Al.

The following summary shows how beryllium resembles aluminium, and differs from other group 2A metals.

Be and Al	**Mg, Ca, Sr and Ba**
1. Both are passive to nitric acid.	React with nitric acid to form salts
2. Both react with Na OH solution to involve hydrogen. $Be\ (s) + 2^{-}OH\ (aq) + 2H_2O\ (l) \rightarrow$ $Be(OH)_4^{2-}(aq) + H_2\ (g)$ beryllate ion $Al\ (s) + 2^{-}OH\ (aq) + 6H_2O\ (l) \rightarrow$ $2Al(OH)_4^{-}\ (aq) + 3H_2(g)$ aluminate ion	No reaction with alkalis
3. The oxides and hydroxides of beryllium and aluminium are amphoteric.	Oxides and hydroxides are basic
4. The chlorides are covalent polymeric solids (through dative bonding). When anhydrous $(BeCl_2)_X$ and $(AlCl_3)_X$ which readily dissolve in organic solvents. They are readily hydrolyzed by water, with evolution of HCl.	Chloride and hydroxides have high conductivities and boiling temperatures; they are not hydrolyzed and are ionic, e.g., $Ca^{2+}Cl^{-}_2$
5. Beryllium carbide, Be_2C, and aluminium carbide, Al_4C_3, give methane on treatment with water. $Be_2C(s) + 4H_2O(l) \rightarrow 2Be(OH)_2(s) + CH_4(g)$ $2Al(s) + 12H_2O(l) \rightarrow 4Al(OH)_3(s) + 3CH_4(g)$	CaC_2 give ethyne with water $CaC_2 + H_2O \rightarrow Ca(OH)_2 + HC{\equiv}CH$
6. Similar complexes of beryllium and aluminium have similar stabilities, e.g., BeF_4^{2-} and AlF_6^{3-}. The former is tetrahedral and the latter octahedral in shape, in agreement with simple theory of electrons their repulsion theory.	No similar compounds

Trial 4.15

1. Beryllium differs in some of its properties from the rest of the elements in the group

   (i) State two properties in which beryllium differs from the rest of the members of the group (2marks)

   (ii) Give reasons why beryllium shows different properties from the rest of the elements (2marks)

2. Explain the following observations

   Beryllium belongs to group II of the periodic Table and yet its chemistry and that of its compounds resemble that of aluminum

3. Beryllium, like aluminium can react with sodium hydroxide solution. Other group II elements do not.

   (i) Write ionic equation for the reaction of beryllium and aluminium with sodium hydroxide solution
   (ii) List **three** other properties in which beryllium show similarity to aluminium.
   (iii) Explain why beryllium behaves differently from other group II elements
   (iv) Name **two** other elements which have similar relationship like beryllium and aluminium.

Similarities between B and Si

The ways in which boron resembles silicon and differs from aluminium are summarized below.

B and Si		Al
1. Nonmetals with very similar properties; non-conducting		A metal with different physical properties; conducting
2. Compounds are covalent		Compounds are partially ionic
3. Their oxides (Be_2O_3 and SiO_2) are strongly acidic in character and form oxy-acid (H_3BO_3, H_2SiO_3) with water. These react with aqueous alkali to produce borates and silicates.		Oxide is amphoteric, with high melting temperature
4. Form covalent chlorides, BCl_3 and $SiCl_4$, which are readily hydrolyzed		$AlCl_3$ is less readily hydrolyzed
5. Form borides, e.g., MgB_2 and CaB_6, and silicides, e.g., Mg_2Si, with metals		No similar compounds.

Group 2A: The alkaline earth metals

The members of group 2A or the alkaline earth metals are Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra.). Radium is radioactive

All the members of this group are highly reactive and are essentially found as compounds in nature. Some physical properties of group 2A elements are given in a table below:

Table 4.16: Some properties of group 2A elements

Elements	Atomic number	Electronic configuration	M.p. ^{0}/C	B.pt. ^{0}/C	Atomic radii (nm)	Ionic radii (nm)	First Ionization energy /kJmol^{-1}	2nd Energies (kJ/mol)	Standard electrode potential / V
Be	4	2.2	1280	2477	0.089	0.031	899	1757	-1.85
Mg	12	2.8.2	650	1100	0.130	0.065	738	1451	-2.37
Ca	20	2.8.8.2	850	1487	0.174	0.099	590	1145	-2.87
Sr	38	2.8.18.8.2	768	1380	0.191	0.113	549	1064	-2.89
Ba	56	2.8.18.18.8.2	714	1640	0.198	0.135	502	965	-2.90
Ra	88	2.8.18.32.18.8.2	700	1140		0.148	508	975	-2.92

Some general remarks:

i) Because of the small size, beryllium differs considerably from the other members of the group and its compounds when anhydrous show a considerable degree of covalent character.

ii) The first ionization energies of this group of metals are considerably higher than those of the corresponding alkali metals; and the energy necessary to form M^{2+} ions is therefore very much greater than that necessary to form uni-positive alkali metal ions. Because of this, their compounds are less ionic than those of alkali metals.

iii) The melting and boiling points of alkaline earth metals are higher than those of the alkali metals because they contribute 2 electrons to the stronger metallic bond.

iv) The melting and boiling points of beryllium are higher than expected, because beryllium has small atomic size and shows non-metallic behavior, it forms Be-Be covalent bonds that are quite stable. These covalent bonds persist even in the liquid state and since melting and boiling involve breaking these bonds, the melting and boiling points are high.

v) Calcium, strontium and barium give characteristic flame coloration which can be used to identify them: calcium (yellowish-red), strontium (crimson) and barium (apple-green).

vi) The tendency to form complexes by M^{2+} ions decrease with increasing size of M^{2+} ions. Thus Be^{2+} on account of its small size and high charge density form stable complexes, e.g., $[BeF_4]^{2-}$ while Ba^{2+} form very few.

Uses of group II metals

Beryllium is used for making containers for uranium-238 as it does not absorb neutrons and therefore does not become radioactive.

Magnesium

1. Is alloyed with aluminium to make Duralumin® used in construction of airplanes and small boats.
2. Used in extraction of titanium
3. As sacrificial anode to prevent iron from rusting
4. The intense white light of burning magnesium is used in flares and distress signals.

Trial 4.16

1. (a) Give one reason in each case to explain why the elements of group II of the periodic table
   - *(i) are divalent*
   - *(ii) form mainly ionic compounds*
   - *(iii) are strongly reducing agents*
   - *(iv) do not form many complex compounds (8marks)*

(b) Although lithium is in group 1 of the periodic table, it resembles elements of group II in some of its properties
   - *(i) State two properties in which Lithium resembles group II elements (2marks)*
   - *(ii) Give one reason for the anomalous of lithium. (2marks)*

(c) Explain why lithium has a smaller electrical conductivity than caesium ion when ions are in aqueous solution (2marks)

(d) Compare the reactivity of the elements of group I and group II with water and indicate the trend in the reactivity within each. (5marks)

(e) Explain why beryllium carbonate is thermally less stable than barium carbonate (2marks)

2. Beryllium chloride is more soluble in ethanol than in water whereas magnesium chloride is more soluble in water and in ethanol.

3. (a) Group II metals form few complexes. However, the tendency to form complexes decrease down the group. Explain
   - *i. What is meant by the term complex (3marks)*
   - *ii. Why the tendency to form complexes decrease down the group (4marks)*

(b) $[V(CN)_6]^{3-}$ and $[CoCl_4]^{2-}$ are complexes formed by vanadium and cobalt respectively State
   - *i. The oxidation state of Vanadium …………………*

Cobalt ……………………….

ii. *The coordination number of*
Vanadium …………………

Cobalt ……………………….

4. *The elements beryllium, magnesium and barium belong to group II in the periodic Table*
   (i) State three chemical properties shown by the elements. For each property, write an equation to illustrate your answer (6marks)

   (ii) Explain the trend in the solubilities of the hydroxides of the element in water (3marks)

5. *(a) Write electronic configuration of strontium(1mark)*
   (b) What is the principal oxidation state of strontium? (1mark)
   (c) State whether the first ionization energy of strontium is greater or less than of barium. Explain your answers (2marks)
6. *The elements Be, Mg, Ca, Sr, Ba are in group II of the Periodic Table.*
   (a) Explain how the following factors vary with the group
      (i) atomic radius
      (ii) ionization energy
      (iii) electropositivity

7. *(a) (i) State how the melting point and boiling points of group II metals differ from those of group (I) (1mark)*
   (ii) Give reason for your answer in (a)(i)

   (b) State why beryllium differ from the rest of the elements in group (II) of the periodic Table (4marks)

Physical properties of group 2A metals

The metals are very much harder than those of alkali metals; they are good conductors of heat and electricity. When pure are silvery colored but quickly tarnish on exposure to air due to formation of oxides.

Chemical reactions of group elements

The reactivity of alkali earth metals increases with increasing atomic number.

Reaction of group 2 elements with water

Beryllium has no reaction with water; magnesium reacts with steam form oxide and hydrogen gas, while calcium, strontium and barium react with water with increasing vigor to give hydrogen and hydroxides. This is due to increase in electropositivity of metals down the group.

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

$$Ca\ (s) + 2H_2O\ (l) \rightarrow Ca(OH)_2\ (s) + H_2\ (g)$$

Trial 4.17

1. *State what would be observed and write an equation for the reaction which takes place when:*

   (i) magnesium is reacted with steam (3½ marks)

   (ii) barium is reacted with water (03marks)

2. *Describe and explain the trend in the reactivity of elements of group II* ***with cold water*** *down the group (8marks)*

Reaction group 2 elements with nonmetals

At suitable conditions alkaline earth metals react with a variety of nonmetals to give oxides, sulphides, halides and nitrides; calcium, strontium and barium also combine with hydrogen to give hydrides.

$$2Mg\ (s) + O_2\ (g) \rightarrow 2MgO\ (s)$$

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

$$Ba\ (s) + O_2 \rightarrow BaO_2\ (s)$$

$$Ca\ (s) + S\ (s) \rightarrow CaS\ (s)$$

$$Sr\ (s) + Cl_2(g) \rightarrow SrCl_2\ (s)$$

$$3Ba\ (s) + N_2(g) \rightarrow Ba_3N_2\ (s)$$

$$Ba\ (s) + H_2\ (g) \rightarrow BaH_2\ (s)$$

Reaction of group 2 elements with acids

With dilute hydrochloric and dilute sulphuric acids, they give the corresponding salts and hydrogen.

$$Be\ (s) + H_2SO_4\ (aq) \rightarrow BeSO_4\ (aq) + H_2\ (g)$$

$$Ca\ (s) + 2HCl\ (aq) \rightarrow CaCl_2\ (aq) + H_2\ (g)$$

Calcium and barium do react slowly with cold dilute dilute sulphuric acid but the reaction immediately stops because their sulphates are insoluble. Calcium dissolves in hot concentrated sulphuric acid to form calcium hydrogen sulphate

$$Ca\ (s) + 2H_2SO_4\ (aq) \rightarrow Ca(HSO_4)_2\ (aq) + H_2\ (g)$$

With hot concentrated acid group II metals react to form sulphates, sulphur dioxide and water but the reaction of barium is stopped by the formation of insoluble sulphate.

$Be(s) + 2H_2SO_4(aq) \longrightarrow BeSO_4(aq) + SO_2(g) + 2H_2O(l)$

Beryllium does not react with concentrated nitric acid because it forms a layer of insoluble BeO

Trial 4.18

Compare the reaction of beryllium and barium with sulphuric acid under various conditions (7 ½ marks)

Beryllium evolves H_2 when reacted with sodium hydroxide solution.

$$Be\ (s) + 2^-OH\ (aq) + 2H_2O\ (l) \longrightarrow [Be(OH)_4]^{2-}\ (aq) + H_2\ (g)$$

Oxides of the Group 2A metals

Alkaline earth metals form oxides of the form $M^{2+}O^{2-}$ when heated with oxygen. Strontium and barium form peroxide $M^{2+}\ O_2^{\ 2-}$ on prolonged heating, particularly if pressure is used.

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

$$Ba(s) + O_2\ (g) \xrightarrow{\text{Heat, pressure}} BaO_2\ (s)$$

The normal oxides can also be produced by thermal decomposition of corresponding carbonates.

$$CaCO_3(s) \xrightarrow{\text{Heat,}} CaO(s) + CO_2\ (g)$$

Reaction of oxides of group elements with water

i) Beryllium oxide (BeO) does not react with water.

ii) Oxides react with water to give hydroxides.

$$CaO\ (s) + H_2O\ (l) \longrightarrow Ca(OH)_2\ (aq)$$

iii) The peroxide form hydroxides and hydrogen peroxide with water.

$$BaO_2\ (s) + 2H_2O\ (l) \longrightarrow H_2O_2\ (l) + Ba(OH)_2\ (aq)$$

Uses of group 2 oxides

1. Magnesium oxide (m.pt. 2800^{0}C) is used for manufacturing lining for open-hearth steel furnaces.
2. CaO is used in many metallurgical operations to produce a slag with impurities in metal ores and as a drying agent.

Hydroxides of alkaline earth metal

They can be prepared by reaction of corresponding oxides with water.

$$CaO\ (s) + H_2O\ (l) \longrightarrow Ca(OH)_2\ (aq)$$

Reaction of group 2 hydroxides

i) Hydroxides neutralize acids to form salts.

$Ca(OH)_2$ (s) + 2HCl (aq) → $CaCl_2$ (aq) + $2H_2O$ (l)

or ^{-}OH (aq) + H^+ (aq) → H_2O (l)

ii) They displace ammonia from ammonium salts.

$Ca(OH)_2$ (s) + $2NH_4Cl$ (aq) → $CaCl_2$ (aq) + $2H_2O$ (l) +NH_3 (g)

or ^{-}OH (aq) + NH_4^+ (aq) → H_2O (l) + NH_3 (g)

Solubility of group 2 metal hydoxide

The solubility of group 2 metal hydroxides increases down the group from $Be(OH)_2$ to $Ba(OH)_2$ due to (i) an increase in ionic character of the hydroxides and (ii) the decrease in lattice energy as the size of the cations increases. The solubility of group II hydroxide is less than that of group I hydroxide because (i) they are less ionic and (ii) they have higher lattice energy due to their high charge density. The low solubility of group 2 metal hydroxides in water make them weaker alkalis.

Trial 4.19

1. *Explain how the solubility of hydroxides of elements of group II elements of the Periodic Table vary down the group (6marks)*
2. *Compare the solubility and bacisity of hydroxides of group II elements with the hydroxides of group I elements (3½ marks)*
3. *Be, Mg, Ca, Sr and Ba are elements in group II of the periodic table*

(a) Describe and explain the trend in the reactivity of elememts with cold water down the group (08marks)

(b) Compare the solubility and basicity of hydroxide of group II elements with the hydroxides of group I elements (3½marks)

(c) (i) Explain why beryllium and aluminium show diagonal relationship. (2marks)

(ii) Write equations to show how beryllium and aluminium each react with concentrated sodium hydroxide solution (3marks)

(d) A chloride of beryllium Z, contains 11.25% beryllium and 88.75% of chlorine

(i) Calculate the empirical formula of Z. (1½ marks)

(ii) Determine the molecular formula of Z (the vapor density of Z =80) (01mark)

(iii) Write the structural formula of Z (1mark)

Uses of group2 metal hydroxides

Calcium hydroxide has many uses, e.g.

a. Calcium hydroxide displaces ammonia from ammonium salts when heated within the mixture. It is thus used in recovering ammonia form ammonium chloride in Solvay process. It is also employed in laboratory preparation of this gas.
b. It can be used as a cheap alkali for neutralizing acidic soils.
c. It is used to soften temporary hard water by addition of calculated amounts.
d. It is used in preparation of calcium hydrogen sulphite solution that removes the lignin from wood, leaving cellulose used in paper manufacturing.
e. Bleaching powder is manufactured by passing chlorine over moist calcium hydroxide. Its composition is complex but it is known to contain Ca^{2+}, OCl^-, Cl^- and OH^- ions and water. However, it is usually represented by simplified formula $CaOCl_2$.

$$Ca(OH)_2\ (s) + Cl_2\ (g) \rightarrow CaOCl_2\ (s) + H_2O\ (l)$$

f. A dilute solution of $Ca(OH)_2$ is used for testing carbon dioxide.

$$Ca(OH)_2\ (aq) + CO_2\ (g) \rightarrow \underset{\text{White ppt.}}{CaCO_3\ (s)} + H_2O\ (l)$$

The carbonates of alkaline earth metals

Magnesium carbonate occurs naturally as magnetite ($MgCO_3$) and in association with calcium carbonate as dolomite, $MgCO_3.CaCO_3$. Calcium carbonate occurs naturally as Calcite, Marble, Lime stone and Chalk.

Preparation of carbonates of group 2 elements

By reaction of corresponding soluble salts with alkali carbonates, e.g.,

$$Mg(NO_3)_2\ (aq) + Na_2CO_3\ (aq) \rightarrow MgCO_3\ (s) + 2NaNO_3\ (aq)$$

Properties of carbonates of group 2 elements

i) Carbonates of alkaline earth metals are sparingly soluble in water.

ii) Carbonates decompose on heating to form oxides and carbon dioxide.

$$MgCO_3\ (s) \rightarrow MgO\ (s) + CO_2\ (g)$$

iii) They react with acids to give off carbon dioxide.

$$CaCO_3\ (s) + 2HCl\ (aq) \rightarrow CaCl_2\ (aq) + H_2O\ (l) + CO_2\ (g)$$

Uses of carbonates of group 2 elements

1. $MgCO_3$ is used to provide MgO for the lining for open-hearth steel furnaces.

2. $CaCO_3$ is a basic raw material in Solvay process and glass industry.

3. It is used to remove impurities form iron ore during extraction of iron.

4. Calcium carbonate used is used for chalk

4. Calcium carbonate is used to make antacids (drugs for ulcers).

Fig. 4.13 Tums (calcium carbonate antacid) made by GlaxoSkithKline limited from calcium Carbonate for treatment of peptic ulcers

Portland cement

Portland cement, so called because it resembles Portland stone when set.

Fig 4.14 Bags of one brand of Tororo cement

Composition

The average composition of different oxides is

Lime (CaO)	50-60%
Silica (SiO_2)	20-25%
Alumina (Al_2O_3)	5-10%
Magnesia (MgO)	2-3%
Ferric oxide (Fe_2O_3)	1-2%
Sulphur trioxide (SO_3)	1%
Sodium oxide (Na_2O)	1%
Potassium oxide (K_2O)	1%

The essential constituents of cement are lime (obtained from lime stone) silica and alumina (present in clay). For cement of good quality, the oxides should be kept in following ratio

$$\frac{\%SiO_2}{\%Al_2O_3} = 25-40, \quad \frac{\%CaO}{\%SiO_2 + \%Al_2O_3 + Fe_2O_3} = 1.9-2.3$$

Cement containing no iron is white but hard to burn. If less lime is present than given in the ration above, the cement is low in strength and sets very soon. If more lime is there, cement cracks. Excess of silica provides slow hardening cement, while excess of alumina gives a quick-setting product.

Raw materials of cement

1. Lime stone $CaCO_3$ provides CaO
2. Clay ($Al_2O_3.SiO_3. Fe_2O_3.H_2O$) supplies silica ($SiO_2$) and alumina ($Al_2O_3$) some clay don't contain Fe_2O_3 and the cement obtained in this case is white and hard to burn
3. Gypsum, $CaSO_4.2H_2O$; these decreases the setting time of cement.

Manufacture of cement

1. Preparation of raw material or slurry. Lime stone and clay are mixed in proper ratio by any of the following process

(a) Dry process: This is employed when the raw material (clay and sand) are hard. In this process the lime stone is first broken into small pieces. It is then mixed with clay in proper ratio and finally pulverized to such a fine powder that 90-95 pass through a 100 mesh size. The mixture is made homogeneous to produce a raw meal.

(b) Wet process. This process is used when raw material Vis lime stone and clay are soft, the fuel is fairly cheap. In this process, lime stone is crushed to suitable size and the clay is washed with water in a wash mill to remove foreign materials like flint. The powdered lime stone is mixed with the clay paste in proper proportion (limestone 75% and clay 25%) and the mixture is finely ground and made homogeneous by means of a compressed air mixing arrangement. The result is a paste or slurry contains about 40% water.

2. Burning (calcinations) of raw meal or slurry in rotary Kiln.

The raw meal or slurry is heated in rotary Kiln up to 1500^{0}C between 2-3 hours to cover the journey in the Kiln. Reaction taking place in the rotary Kiln can be divided into the following

(i) Reaction taking place in moderate temperature zone (up to 800^{0}C)
In this zone, free moisture is removed and clay ($Al_2O_3.2SiO_2Fe_2O_3.2H_2O$ is broken into Al_2O_3, SiO_2 and Fe_2O_3

$Al_2O_3.2SiO_2.Fe_2O_3.2H_2O \longrightarrow Al_2O_3 + 2SiO_2 + Fe_2O_3 + H_2O$

(ii) Reaction taking place in average temperature zone (800^0-1000^{0}C)

Lime stone ($CaCO_3$) decomposes into lime (CaO) and CO_2.

$CaCO_3 \longrightarrow CaO\ (s) + CO_2\ (g)$

(iii) Reaction in maximum temperature zone (1000^0-1500^{0}C)

The oxides CaO, Al_2O_3, SiO_2, Fe_2O_3 combine to together to form silicates $2CaO.SiO_2$, 3 $CaO.SiO_2$, calcium aluminates, $2CaO.Al_2O_3$ and tetra calcium aluminoferrite $4CaO.Al_2O_3.Fe_2O_3$

$2CaO + SiO_2 \longrightarrow 2CaO.SiO_2$

$3CaO + SiO_2 \longrightarrow 3CaO.SiO_2$

$2CaO + Al_2O_3 \longrightarrow 2CaO.Al_2O_3$

$3CaO + Al_2O_3 \longrightarrow 3CaO.Al_2O_3$

$4CaO + Al_2O_3 + Fe_2O_3 \longrightarrow 4CaO.Al_2O_3\ .Fe_2O_3$

The mixture of the above silicates, aluminates and tetracaliumaluminaferrite is called cement clinker.

3. Mixing the cement clinker with gypsum.

The cooled clinker is mixed with 2-3% of its weight of gypsum ($CaSO_4.2H_2O$). $3CaO.Al_2O_3$, which is a fast setting constituent of the clinker, reacts with gypsum to form the crystal calcium suphoaluminate, $3CaO.Al_2O_3.3CaSO_4.2H_2O$.

$3CaO.Al_2O_3 + 3(CaSO_4.2H_2O).2H_2O \rightarrow 3CaO.Al_2O_3.3CaSO_4.2H_2O + 6H_2O$

Thus, gypsum removes the fast setting $3CaO.Al_2O_3$ and the process of setting cement gets retarded and these results in better strength of the mass which results. The product is crushed into powder and packed.

Setting of cement

The use of cement in construction of building is based on its property of setting to a hard mass, when its paste with water is allowed to stand for some time. When cement is mixed with water, it absorbs water and the mass become hard and resistant to pressure. This is known as setting of cement. The reactions involved are

(i) Reactions taking place during first 24hours/short time after the cement is mixed with water, the following reactions take place

(a) $3CaO.Al_2O_3$ absorbs water and forms a hydrated colloidal gel of tricalciumaluminate

$$3CaO.Al_2O_3 + 6H_2O \xrightarrow{\text{Hydration}} 3CaO.Al_2O_3.6H_2O$$

The gel of $3CaO.Al_2O_3.6H_2O$, so formed start crystalling slowly

(b) $3CaO.Al_2O_3$, which is a fast setting material, reacts with gypsum ($CaSO_4.2H_2O$) to form crystals of calcium sulphoalminates, $3CaO.Al_2O_3.3CaSO_4.2H_2O$

$$3CaO.Al_2O_3 + 3(CaSO_4.2H_2O).2H_2O \rightarrow 3CaO.Al_2O_3.3CaSO_4.2H_2O + 6H_2O$$

Thus, gypsum removes the fast setting $3CaO.Al_2O_3$ and the process of setting cement gets retarded and these results in better strength of the mass which results.

(ii) Reactions taking place between 1-7days

$3CaO.SiO_2$ and $3CaO.Al_2O_2$ get hydrolyzed according to equation

$$3CaO.SiO_2 + H_2O \xrightarrow{\text{hydrolyzed}} Ca(OH)_2\ (s) + 2CaO.SiO_2\ \text{(colloidal gel)}$$

$Ca(OH)_2$ formed starts changing into needle shaped crystals which get studded in the colloidal gel

$2CaO.SiO_2$ formed above, and thus impart strength to it.

$$3CaO.Al_2O_3 + 6H_2O \xrightarrow{\text{hydrolyzed}} 3Ca(OH)_2\ (s) + 3Al(OH)_3\ (s)$$

$Al(OH)_3$ formed fills the interstices in hardening mass.

(iii) Reactions taking place between 7-28days

$2CaO.SiO_2$ begin to hydrate (very slowly) and form the hydrate colloidal gel of dicalcium silicate of composition $2CaO.SiO_2.xH_2O$

$$2CaO.SiO_2 + xH_2O \xrightarrow{\text{hydration}} 2CaO.SiO_2.xH_2O$$

Needles of $Ca(OH)_2$ form get studded in hydrate colloidal gel of $2CaO.SiO_2.xH_2O$ and thus impart strength to it.

$4CaO.Al_2O_3.Fe_2O_3$ also gets hydrated to form a colloidal gel of composition $3CaO.Al_2O_3.\ 6H_2O$.

$$4CaO.Al_2O_3.Fe_2O_3 + 6H_2O \xrightarrow{\text{hydration}} 3CaO.Al_2O_3.6H_2O + CaOFe_2O_3$$

The gel formed above starts losing water partially by evaporation and partly by forming hydrate constituent. Thus cement sets to a hard mass.

Concrete

Concrete is a mixture of cement with sand and gravel. Concrete is used to make floors, brick and pillars that build strong building and roads, figure 9.2

(a) (b)

Fig. 4.15 (a) concrete bricks and pillars for strong wall and (b) concrete pavers

Trial 4.20

Cement is a mixture of inorganic compounds and is widely used in construction industry.

(a) Namethe two processes by which cement can be made (1 ½ marks)

(b) Describe briefly how cement is manufactured and explain the main physical and chemical changes involved. (13marks)

(c) Describe what happens when water is added to a mixture of cement and sand (4½ marks)

Hydrogen carbonates of alkaline earth metals

The solid hydrogen carbonates are unknown at room temperature, but they cause temporary hardness in water. Rain water attacks any rocks containing calcium and magnesium carbonates and a dilute solution of hydrogen carbonate is formed.

$$CaCO_3\ (s) + H_2O\ (l) + CO_2\ (aq) \longrightarrow Ca(HCO_3)_2\ (aq)$$

When water containing the dissolved hydrogen carbonates is boiled deposits of the carbonates are formed.

$$Ca(HCO_3)_2\ (aq) \xrightarrow{\text{heat}} CaCO_3\ (s) + H_2O\ (l) + CO_2\ (g)$$

The halide of alkaline earth metals

a. The **solubility** of group 2A halides increases in order $MF_2 \rangle MCl_2 \rangle MBr_2 \rangle MI_2$ and for any given halide, X, the solubility increases in order $BeX_2 \langle MgX_2 \langle CaX_2 \langle SrX_2 \langle BaX_2$.

Reasons

1. Because the lattice (bond) energy decreases in order $MF_2 \rangle MCl_2 \rangle MBr_2 \rangle MI_2$.

2. The ionic character increases in order $BeX_2 \langle MgX_2 \langle CaX_2 \langle SrX_2 \langle BaX_2$.

The bromides and iodides are appreciably soluble in organic solvents such as alcohols, carbonyl compounds and ether, due to formation of complexes.

The nitrates of alkaline earth metals

They are obtained by reacting metal, oxides, hydroxides and carbonates with dilute nitric acid and crystallization of the resultant solutions.

$$MgO\ (s) + 2HNO_3\ (aq) \rightarrow Mg(NO_3)_2\ (aq) + H_2O\ (l)$$

$$Ca\ (s) + 2HNO_3\ (aq) \rightarrow Ca(NO_3)_2\ (aq) + H_2\ (g)$$

$$Ba(OH)_2\ (s) + 2HNO_3\ (aq) \rightarrow Ba(NO_3)_2\ (aq) + 2H_2O\ (l)$$

$$CaCO_3\ (s) + 2HNO_3\ (aq) \rightarrow Ca(NO_3)_2\ (aq) + H_2O\ (l) + CO_2\ (g)$$

Reactions of nitrates of group 2 elements

1. The nitrates decompose on heating to give the oxide, nitrogen dioxide and oxygen.

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

Uses: Calcium nitrate is used as nitrogenous fertilizer e.g, CAN, NPK.

The sulphate of alkaline earth metals

Occurrence

Magnesium sulphate occurs as Epsom salts, $MgSO_4.7H_2O$. Calcium sulphate occurs as anhydrite, $CaSO_4$, and as gypsum $CaSO_4.2H_2O$.

Preparation

By reaction of corresponding metal oxides, hydroxide or carbonates with dilute sulphuric acid.

$$Mg\ (s) + H_2SO_4\ (aq) \rightarrow MgSO_4\ (aq) + H_2\ (g)$$

$$CaO\ (s) + H_2SO_4\ (aq) \rightarrow CaSO_4\ (aq) + H_2O\ (l)$$

$$Ba(OH)_2\ (aq) + H_2SO_4\ (aq) \rightarrow BaSO_4\ (s) + 2H_2O\ (l)$$

$$CaCO_3\ (s) + H_2SO_4\ (aq) \rightarrow CaSO_4\ (aq) + H_2O\ (l) + CO_2\ (g)$$

Solubility of group 2 metal sulphates

Like carbonates the solubility of the sulphates decreases down the group; thus, Ba^{2+} ions are used for testing for sulphate ions.

$$Ba^{2+} (aq) + SO_4^{2-} (aq) \rightarrow BaSO_4 (s) \text{ (white ppt)}$$

Reactions

1. Sulphates of calcium and magnesium are reduced to sulphide when heated with carbon.

$$MgSO_4 (s) + 2C (s) \rightarrow MgS (s) + 2CO_2 (g)$$

Uses

When heated to a temperature a little in excess of 100^0C, gypsum loses three-quarter of its water of crystallization and become **plaster of Paris**, $CaSO_4.\frac{1}{4}H_2O$.
Plaster of Paris sets solid, with slight expansion, on mixing with water and standing. It is widely used in making plaster casts, as a surface for walls, molds and in surgery to keep an injured limb rigid. Plaster board, which is used in building, is also made from plaster of Paris.

Ammonium sulphate, a nitrogenous fertilizer, is manufactured by reacting powdered calcium sulphate with an ammonia solution containing carbon dioxide $(NH_4)_2CO_3$). This reaction takes place because calcium carbonate is less soluble in water that is calcium sulphate.

$$2NH_3(aq) + CO_3(g) + H_2O(l) \rightarrow (NH_4)_2CO_3(aq)$$
$$CaSO_4 (s) + (NH_4)_2CO_3 (aq) \rightarrow CaCO_3 (s) + (NH_4)_2SO_4 (aq)$$

The carbides and nitride of group 2A metals

i) Calcium and magnesium carbides $M^{2+}(C{\equiv}C)^{2-}$ can be made by reaction of the oxides with carbon at high temperatures.

$$CaO (s) + 3C (s) \rightarrow Ca^{2+} (C{\equiv}C)^{2-} (s) + CO (g)$$

Addition of water to calcium carbide gives ethyne, which as a vital raw material in Organic chemistry.

$$Ca (C{\equiv}C) (s) + 2H_2O (l) \rightarrow Ca(OH)_2 (s) + HC{\equiv}CH (g)$$

(Ethyne)

ii) The nitrides, $(M^{2+})_3 (N^{3-})_{2,}$ are formed on heating metals with nitrogen.

$$3Mg (s) + N_2 (g) \rightarrow Mg_3N_2 (s)$$

Addition of water causes hydrolysis and ammonia is liberated.

$$Mg_3N_2 (s) + 6H_2O (l) \rightarrow 3Mg(OH)_2 (s) + 2NH_3 (g)$$

Hydrolysis of group 2A metal salts

Soluble salts of Beryllium with strong acids undergo some acidic hydrolysis. Although the majority of beryllium compound are essentially covalent, the hydrated ion $[Be(H_2O)_4]^{2+}$ exist in solution because, beryllium ions are small and attract water molecules more strongly. The hydrolysis beryllium salts can be represented, thus:

$$[Be(H_2O)_4]^{2+} (aq) + H_2O (l) \longrightarrow [Be(H_2O)_3(OH)]^{+} (aq) + H_3O^{+} (aq)$$

However, salts of weak acid with the most electropositive alkaline earth metals undergo an alkaline hydrolysis in water:

$$H^{+} + H_2O (l) \longrightarrow OH^{-} (aq) + H_2 (g)$$

$$(C{\equiv}C)^{2-} + 2H_2O (l) \longrightarrow 2OH^{-} (aq) + HC{\equiv}CH (g)$$

$$N^{3-} + H_2O (l) \longrightarrow 3OH^{-} (aq) + NH_3 (g)$$

Magnesium chloride hydrolyzes to give a basic chloride

$$MgCl_2(aq) + H_2O(l) \quad MgOHCl \rightleftharpoons (s) + HCl(aq)$$

Qualitative analysis of Mg^{2+}, Ca^{2+}, and Ba^{2+}

Compound containing these ions are white and when dissolved in water to give colorless solution. Unlike other white compound (Pb^{2+}, Zn^{2+}, and Al^{3+}) which form white precipitates with sodium hydroxide that **are soluble in excess**, these form curdy /powder like precipitate with sodium hydroxide solutions which are **insoluble in excess**. Because alkaline earth metal behave very similarly to each other in aqueous solutions, it is very difficult to distinguish them and especially to separate them. The different reactions used to distinguish between them are shown in the table 9.2 below:

Table 4.16. Qualitative analysis of group 2 elements

Test	Mg^{2+}	Ca^{2+}	Ba^{2+}
To 1 ml of solution add dilute sodium hydroxide drop wise until in excess	White curdy ppt. Insoluble in excess	White curdy ppt. Insoluble in excess	White curdy ppt. Insoluble in excess
To 1 ml of solution add Freshly prepared ammonia solution drop wise until in excess, followed with ammonium chloride solution	White gelatinous ppt. Soluble in ammonium chloride solution	No Observable change	* No Observable change * With an aged solution a precipitate may occur due to formation of Barium carbonate
Ammonium carbonate followed by a few drops of ammonium chloride	White ppt. Soluble in ammonium chloride	White ppt. insoluble in ammonium chloride	White ppt. insoluble in ammonium chloride
Potassium chromate followed with dilute acetic acid (30%)	No ppt. formed	No ppt. formed	A yellow ppt insoluble in acetic acid
Dilute sulphuric acid followed by conc. H_2SO_4 and boil	No ppt. formed	White ppt. Soluble in hot concentrate sulphuric acid	White ppt. Insoluble
Saturated solution of calcium sulphate	No ppt. formed	No ppt. formed	*White ppt
Ammonium oxalate solution followed by drop acetic acid	No precipitate formed	White ppt insoluble in acetic acid	White ppt soluble in acetic acid
To 1 ml add a spatula end-full of ammonium chloride followed by 1 ml disodium hydrogen phosphate	No Precipitate formed	White ppt formed	White ppt formed

Trial 4.21

(a) *Name a reagent that can be used to distinguish between the following pairs of ions. In each case, state what is observed if each ion is separately treated with the reagent.*

(i) Ba^{2+} and Ca^{2+} *(03 marks)*

Thank you
Dr. Bbosa Science