Chapter 13

S and P-Block Elements

Learning Periodic Table

Group IA :

Hayee	Laila	Nay	Ki	Rb	Csey	Fryad
H	Li	Na	K	Rb	Cs	Fr

Group II A :

Bhai	Meray	Camray	Say	Bahar	Raho
Be	Mg	Ca	Sr	Ba	Ra

Group III A

Baba	Ali	Gayee	India se	Til lenay
B	Al	Ga	In	Ti

Group IV A

Chemistry	Sir	Gives	Silly	Problems
C	Si	Ge	Sn	Pb

Group VA

Nice	Pal	As	Sibi	Bhai
N	P	As	Sb	Bi

Group VIA

Oye	Sun	Seni	Tera	Pol (Kolhu)
O	S	Se	Te	Po

Group VII A

Faltu Class mein Boring Instructor Ata hein

F Cl Br I At

Group VIII A

He Never Argue, Kal Xero Run pe out

He Ne Ar Kr Xe Rn

D-Block Period 4

Some Times Very Critical Measures For Completely Nasty Cures 2 taken.

Sc Ti V Cr Mn Fe Co Ni Cu Zn

IN A PERIOD

1. Atomic Radius → Decrease
2. Ionization Energy → Increase
3. Electronegativity → Increase
4. Electron affinity → Increase
5. Metallic character → Decrease
6. Melting and boiling point → I-A to IV-A increase
   V to VIII decrease
7. Electrical Conductance → Decrease
8. Hydration Energy → Increase

IN A GROUP

1. Atomic Radius → Increase
2. Ionization Energy → Decrease
   however II A > III A
   VA > VI A
3. Electronegativity → Decrease
4. Electronaffinity → Decrease
   however IA > II A
   IV A > VA
   VII A > VIII A
5. Metallic character → Increase
6. Electrical Conductance → Decrease
7. Hydration energy → Decrease

ATOMIC RADIUS

Atomic Radius depends on:

1. Effective nuclear charge
2. Number of shells

IONIZATION ENERGY

1. Size of atom $I.E \propto 1/d$
2. Nuclear charge $I.E \propto Z$
3. $I.E \propto$ 1/Screening Effect of Electrons
4. Electronic Configuration
5. For Same Shell

$s > p > d > f$

ELECTRON AFFINITY

1. $E.A \propto$ Nuclear Charge
2. $E.A \propto$ 1/Atomic Size
3. Electronic Configuration: Stable the configuration of an atom, lesser will be its tendency to accept an electron and hence lower value of electron affinity

* Chlorine has highest electron affinity in periodic table.

ELECTRONEGATIVITY

1. $E.N \propto$ Nuclear charge
2. $E.N \propto$ 1/Atomic Size

METALLIC CHARACTER

Tendency to lose electron and form cations and basic oxides.

TRENDS IN PERIOD 3

1. Atomic Radius:

   Na > Mg > Al > S > P > S > Cl < Ar

2. FIRST IONIZATION ENERGY:

   Na < Mg > Al < Si < P > S < Cl < Ar

3. Electronegativity:

   Na < Mg < Al < Si < P < S < Cl

4. Melting Point:

   Na < Mg < Al < Si > P_4 < S_8 > Cl_2 > Ar

REACTION OF PERIOD 3 ELEMENTS WITH WATER, OXYGEN AND CHLORINE

1. SODIUM (Na)

1. $2Na + 2H_2O \rightarrow 2NaOH + H_2$
2. $4Na + O_2 \rightarrow 2Na_2O$ (white solid)
3. $2Na + O_2 \rightarrow Na_2O_2$ (white or white-yellowish solid)
4. $2Na + Cl_2 \rightarrow 2NaCl$

2. MAGNESIUM (Mg)

1. $Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$
2. $Mg + H_2O \rightarrow MgO + H_2$ (white solid)
3. $2Mg + O_2 \rightarrow 2MgO$ (white solid)
4. $Mg + Cl_2 \rightarrow MgCl_2$ (white solid)

3. ALUMINIUM (Al)

1. $2Al + H_2O \rightarrow Al_2O_3 + 3H_2$
2. $4Al + 3O_2 \rightarrow 2Al_2O_3$
3. $2Al + 3Cl_2 \rightarrow 2AlCl_3$ (pale yellow)

4. SILICON (Si)

1. $Si + 2H_2O \rightarrow SiO_2 + 2H_2$
2. $Si + O_2 \rightarrow SiO_2$
3. $Si + 2Cl_2 \rightarrow SiCl_4$ (colorless liquid)

5. PHOSPHORUS (P)

1. $P_4 + 3O_2 \rightarrow P_4O_6$
2. $P_4 + 5O_2 \rightarrow P_4O_{10}$

White phosphorus catches fire spontaneously in air, burning with flame and producing clouds of white smoke.

3. $P_4 + 6Cl_2 \rightarrow 4PCl_3$ (colorless fuming liquid)
4. $P_4 + 10Cl_2 \rightarrow 4PCl_5$ (straw color liquid)

→ Phosphorous doesnot react with water.

6. SULPHUR (S)

1. $S + O_2 \rightarrow SO_2$ (colorless gas)

Sulphur burns in oxygen or air with pale blue flame.

2. $2S + Cl_2 \rightarrow S_2Cl_2$

Sulphur produces orange, foul smelling disulphide S_2Cl_2 with chlorine.

7. CHLORINE (Cl)

1. $Cl_2 + H_2O \rightleftharpoons HCl + HOCl$

$2HOCl \rightarrow 2HCl + O_2$

$2Cl_2 + 2H_2O \rightarrow 4HCl + O_2$

In presence of sunlight, chloric (I) acid slowly decomposes to produce more hydrochloric acid releasing oxygen gas.

→ Despite having several oxides, chlorine would not react directly with oxygen.

MELTING POINT OF OXIDES

- → MgO has highest melting point due to smaller size of Mg^{+2} and O^{-2} ions.
- → Both ions have same magnitude of charge.
- → MgO has greater ionic character

$Na_2O < MgO > Al_2O_3 > SiO_2 > P_4O_{10} > P_4O_6 > SO_3 > SO_2 > Cl_2O_7 > Cl_2O$

ELECTRICAL CONDUCTIVITY OF OXIDES

None of them will conduct electricity when they are solid. The ionic oxides can, however, undergo electrolysis in molten state.

Good Conductors : Na_2O , MgO , Al_2O_3

Very Poor Conductors : SiO_2

Insulators : P_4O_{10}, P_4O_6 , SO_3 , SO_2 , Cl_2O_7, Cl_2O

ACID BASE BEHAVIOUR OF OXIDES

- → Oxides of metals are basic
- → Oxides of non-metals are acidic.

Basic Oxides :

Na_2O (alkaline) , MgO (weakly amphoteric)

Amphoteric Oxide :

Al_2O_3

Acidic Oxides :

P_4O_{10}, P_4O_6, SO_3 , SO_2 , Cl_2O_7 , Cl_2O

STATES OF OXIDES OF PERIOD 3

1. Solid : Na_2O , MgO , Al_2O_3 , SiO_2 , P_4O_{10} , P_4O_6
2. Liquid : SO_3 , Cl_2O_7
3. Gas : SO_2 , Cl_2O

* Amphoteric oxides are usually water insoluble.

CHLORIDES OF PERIOD 3

1. State

   Solid : $NaCl$, $MgCl_2$, Al_2Cl_6 ,

   Liquid : $SiCl_4$, PCl_3 , S_2Cl_2

2. Electrical Conductivity :

   Good Conductors : $NaCl$, $MgCl_2$

   Very Poor : Al_2Cl_6

   Insulators : $SiCl_4$, PCl_3 , S_2Cl_2

3. Effect of Adding chloride to water:

   Dissolves readily : $NaCl$, $MgCl_2$, Al_2Cl_6

   Reacts with water producing fumes of HCl : $SiCl_4$, PCl_3 , S_2Cl_2

HYDROXIDES OF PERIOD 3

The word "hydroxide" refers to anything which contains either a hydroxide ion (OH^-) or an $-OH$ group covalently bonded to the element.

ACIDIC AND BASIC NATURE

1. Basic : $NaOH$, $Ca(OH)_2$, $Mg(OH)_2$
2. Amphoteric : $Al(OH)_3$
3. Acidic : $Si(OH)_4$, H_3PO_4, H_2SO_4, $HClO_4$

SODIUM AND MAGNESIUM HYDROXIDES

- → white solids having soapy touch
- → very hygroscopic (tending to absorb moisture)

SODIUM HYDROXIDE:

- → slightly soluble in alcohols
- → highly soluble in water
- → form a number of hydrates e.g $NaOH.2H_2O$
- → used in soap industry, petroleum refining and reclaiming of rubber.

MAGNESIUM HYDROXIDES $Mg(OH)_2$

Preparation :

$$MgCl_2 + 2KOH \longrightarrow Mg(OH)_2 + 2KCl$$

→ Solubility of $Mg(OH)_2$ is enhanced tremendously by addition of NH_4Cl

* ALUMINIUM HYDROXIDE

Preparation :

1. $Al_2(SO_4)_3 + 6NH_4OH \rightarrow 2Al(OH)_3 + 3(NH_4)_2SO_4$
2. $AlCl_3 + 3H_2O \rightarrow Al(OH)_3 + 3HCl$

Reaction With Base:

$2Al(OH)_3 + 2NaOH \rightarrow 2Na[Al(OH)_4]$

Aluminate

Reaction with Acid:

$Al(OH)_3 + HCl \rightarrow AlCl_3 + H_2O$

SILICON HYDROXIDE

Preparation :

$SiO_2 + 2H_2O \underset{}{\overset{800^\circ C}{\rightleftharpoons}} Si(OH)_4$

→ $Al(OH)_3$ has the capacity of absorbing various dyes forming coloring matter known as "lakes"

GROUP IA ELEMENTS

→ They are known as alkali metals since they form oxides and hydroxides which combine with water to form alkaline solution.

→ Only the first three of these are safe to keep in the school lab. The rest are violently reactive.

→ Lithium is lightest metal known

→ They have low densities. Li, Na and K will float on water.

→ They react quickly with water producing hydroxides and hydrogen gas.

TRENDS

1. Atomic Radius

$Li < Na < K < Rb < Cs$

2. First Ionization Energy

$Li > Na > K > Rb > Cs$

Due to their low I.E these metals have a greater tendency to lose valence electron to change into M^+ ions.

3. Electronegativity:

$Li > Na > K > Rb > Cs$

Since these metals are highly electropositive, their electronegativity values are very low.

4. Melting and Boiling Point:

Li > Na > K > Rb > Cs

The forces binding the atoms in crystal lattices of these metals are relatively weak. Consequently these metals are soft with low melting and boiling points.

Metal becomes softer from top to bottom.

5. Density

Li < Na > K < Rb < Cs

Potassium is lighter than Sodium due to vacant d-orbital.

REACTION OF GROUP IA ELEMENTS

1. WITH WATER

→ Group IA elements are very good reducing agents

→ They all react vigorously with water reducing it to hydrogen gas

→ excluding Lithium which reacts slower than all other elements of IA

$$M + H_2O \rightarrow MOH + \frac{1}{2}H_2\uparrow$$

→ Potassium reacts even more vigorously, it cracks and pops as hydrogen explodes

→ Cs and Rb explodes violently in water.

REACTION OF IA WITH CHLORINE

$2X + Cl_2 \rightarrow 2XCl$

$2Na + Cl_2 \rightarrow 2NaCl$

$2K + Cl_2 \rightarrow 2KCl$

* Sodium burns with an intense orange flame in exactly the same way as it does in pure oxygen

REACTION OF IA WITH OXYGEN

→ Very reactive metals and reactivity increase down the group.

→ Rb and Cs catch fire in air and produce 'superoxid sush as RbO_2 and CsO_2.

Normal Oxide : O^{-2}

Peroxide : O^{-1}

Super oxide : O_2^-

$4Li + O_2 \rightleftharpoons 2Li_2O$

$6Li + N_2 \rightarrow 2Li_3N$

$4Na + O_2 \rightarrow 2Na_2O$ (sodium oxide)

$2Na + O_2 \rightarrow Na_2O_2$ (sodium peroxide)

$2K + O_2 \rightarrow K_2O_2$ (Potassium peroxide)

$K + O_2 \rightarrow KO_2$ (Potassium superoxide)

$Rb + O_2 \rightarrow RbO_2$ (Rubidium superoxide)

$Cs + O_2 \rightarrow CsO_2$ (Cesium superoxide)

REACTION OF OXIDES WITH WATER

$X_2O + H_2O \rightarrow 2XOH$

(normal oxide) (Metal Hydroxide)

$X_2O_2 + 2H_2O \rightarrow 2XOH + H_2O_2$

(Peroxide) (Hydrogen Peroxide)

$2XO_2 + 2H_2O \rightarrow 2XOH + H_2O_2 + O_2$

(Super oxide)

REACTION OF OXIDES WITH DILUTE ACIDS

$X_2O + 2HCl \rightarrow 2XCl + H_2O$

(Normal Oxide) salt

$X_2O_2 + 2HCl \rightarrow 2XCl + H_2O_2$

(Peroxide)

$2XO_2 + 2HCl \rightarrow 2XCl + H_2O_2 + O_2$

(Super oxide)

FLAME TEST

Li → Red
Na → Yellow
K → Lilac
Rb → Red
Cs → Blue / Violet
Be → No colour
Mg → No colour
Ca → Orange-red
Sr → Red
Ba → Pale Green

EFFECT OF HEAT ON NITRATES

* Lithium Nitrate :

$$4LiNO_3 \rightarrow 2Li_2O + 4NO_2 + O_2$$

* Other Alkali metals :

$$2XNO_3 \rightarrow 2XNO_2 + O_2$$

NO_3 : Nitrate
NO_2 : Nitrite

EFFECT OF HEAT ON CARBONATES

$$Li_2CO_3 \rightarrow Li_2O + CO_2$$

The rest of Group I carbonates donot decompose even at higher temperature.

EFFECT OF HEAT ON HYDROGEN CARBONATES

Bicarbonates are less stable than carbonates. They decompose on heating forming carbonates.

$$2NaHCO_3 \longrightarrow Na_2CO_3 + CO_2 + H_2O$$

Thermal stability of hydrogen carbonates of Group I and II increase down the group due to increasing polarizing power. The polarizing power of a cation increases with increasing charge on the ion and decreasing radius of ion.

Cations of greater polarizing power distorts the HCO_3^- ion more and facilitates its decomposition than a cation of larger size and lesser polarizing power.

Due to the same reason, bicarbonates of Group-I are more stable than those of Group II

POLARIZING POWER:

The ability of a cation to distort an anion is known as its polarizing power and the tendency of the anion to be polarized by the cation is known as its polarizibility.

FACTORS AFFECTING POLARIZING POWER:

1. Small Cation

Mg^{+2} is smaller cation than Na^{+1} and hence have high polarizing power.

2. Large Anion
3. Large charges

(Mg^{+2} has larger charge than Na^{+1} and hence more polarizing power)

PECULIAR BEHAVIOUR OF LITHIUM

1. Lithium is much harder and lighter than other I-A elements
2. Gives normal oxides on burning
3. Compounds are more covalent and soluble in organic solvent.
4. LiOH on heating gives Li_2O and less basic.
5. Gives nitride on reaction with N_2
6. Gives carbide directly
7. Compounds are either insoluble or sparingly soluble in water.
8. Li^+ ions and its compounds are heavily hydrated.

OTHER ALKALI METALS

1. Other elements are soft.
2. Give peroxide and super oxide.
3. Compounds are ionic and insoluble in organic solvent
4. Hydroxides are stable and more basic.

5. No reaction with nitrogen.
6. No reaction with carbon directly
7. Compounds are highly soluble in water.
8. Ions and compounds are less hydrated.

GROUP IIA ELEMENTS

→ Alkaline Earth metals donot exist free in nature.

→ Mg and Ca are very abundant in rocks of earth's crust.

→ Chief Sources of Magnesium:

Sea water, Underground brines, Mineral dolomite, Magnesite ($MgCO_3$)

→ Chief Sources of Calcium:

$CaCO_3$ (sea shell), Gypsum ($CaSO_4 \cdot 2H_2O$)

→ All alkaline earth metals except Be are white in colour.

Be → Grey

TRENDS

1. Atomic Radius

IA > IIA

Be < Mg < Ca < Ba < Ra

2. IONIZATION ENERGY:

Be > Mg > Ca > Sr > Ba < Ra

I.E of Ra is higher than Ba

REACTION OF GROUP IIA WITH WATER

→ Group II elements become more reactive towards water as we go down the group.

→ These metals react slowly with water liberating hydrogen and forming hydroxides.

→ Beryllium doesnot react with water or steam at red heat.

→ Magnesium react with steam to give MgO or $Mg(OH)_2$ with excess steam.

$$Mg + H_2O_{(g)} \rightarrow MgO + H_2$$

or $$Mg + 2H_2O_{(g)} \rightarrow Mg(OH)_2 + H_2$$

REACTION OF GROUP IIA WITH OXYGEN

→ Be is reluctant to burn unless it is in the form of dust or powder. Be has very strong layer of BeO on its surface which prevents any more oxygen reaching the underlying beryllium to react with it.

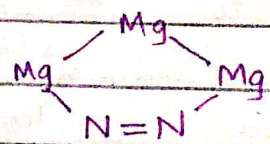
→ Ca and Mg form normal oxides, MgO and CaO

→ Sr, Ba and Ra form peroxides SrO_2 and BaO_2

REACTION OF GROUP IIA WITH NITROGEN

The nitrides of alkaline earth metals are ionic in nature except that of Be which is covalent and unpredictable.

Nitride contain the nitrite ion (N^{-3})



$$Mg + N_2 \xrightarrow{\Delta} Mg_3N_2$$

$$Be + N_2 \xrightarrow{\Delta} Be_3N_2$$

SOLUBILITY OF HYDROXIDES OF IIA

Solubility of hydroxides in water increase from $Be(OH)_2$ to $Ba(OH)_2$.

SOLUBILITY OF ~~HYDRO~~ SULPHATES OF IIA

→ Decrease down the group
→ $BeSO_4$ and $MgSO_4$ are fairly soluble in water
→ $CaSO_4$ is sufficiently soluble
→ $SrSO_4$ and $BaSO_4$ are almost insoluble.

SOLUBILITY OF CARBONATES OF IIA

→ Stability of carbonates increase down the group
→ All carbonates insoluble in neutral medium
→ All dissolve in acids
→ All decompose at red heat

* Both carbonates and nitrates become more thermally stable as we go down the group. And hence more heat is required to decompose down the group.

PECULIAR BEHAVIOUR OF BERYLLIUM

1. Harder as iron
2. Too small that it form covalent compounds
3. High melting and boiling point
4. ~~Form insoluble~~ Beryllium doesnot react with water or steam even at red heat
5. Have resistance to complete oxidation (not easily affected by dry air
6. Salts of Be donot impart any color to the flame.

OTHER MEMBERS

1. Softer than Be but harder than IA group
2. Makes ionic compounds usually
3. Less Melting and boiling point than Be
4. Form soluble oxide with water
5. oxidize completely
6. Salts impart color to flame

GROUP IV ELEMENTS

ns^2, np^2

NATURE

Non-metal : Carbon, ~~Silicon (Si)~~
Metalloid : Silicon (Si), Germanium (Ge)
Metal : Tin (Sn), Lead (Pb)

→ All of them form hydrides (XH_4)
→ They form tetrachlorides (XCl_4)
→ They also form dioxides (XO_2)

OXIDATION STATE

→ Carbon and silicon show +4 oxidation state in carbonates and silicates e.g CCl_4, $SiCl_4$, SnO_2
→ The remaining members can show +2 as well as +4 oxidation state.
→ Carbon forms C^{-4} and C_2^{-2} ions in certain compounds.
Be_2C ($C = -4$, $Be = +2$), Na_2C_2 ($C_2 = -2$, $Na = +1$)

FAJAN'S RULE

The smaller the cation, the greater is the amount of covalent character in the compounds. e.g
Compounds of Sn^{+4} are covalent, while those of Sn^{+2} are ionic.

CHLORIDES OF CARBONS, SILICON AND LEAD

→ All these elements give tetrachlorides (MCl_4) which are covalent and tetrahedral due to sp^3 hybrid orbitals.

→ MCl_4 are fuming liquids at room temperature. The stability of MCl_4 decreases from CCl_4 to $PbCl_4$. $PbCl_4$ decomposes to give $PbCl_2$ and Cl_2 gas

$$PbCl_4 \rightarrow PbCl_2 + Cl_2$$

REACTION OF CHLORIDES WITH WATER

→ CCl_4 doesnot react with water. This is due to bulky nature of chlorine atoms around small carbon atom. As a result oxygen of water cannot penetrate to reach carbon atom.

→ $SiCl_4$ to $PbCl_4$ react violently with water to produce their respective oxides (SiO_2, PbO_2) and fumes of HCl.

→ $PbCl_2$ is ionic in nature and sparingly soluble in cold water, but more soluble in hot water

OXIDES OF GROUP IV

The elements of Group IV form two types of oxides i-e monoxide and dioxide in which these exist in +2 and +4 oxidation state.

SILICON DIOXIDE (SiO_2)

→ Every silicon atom is tetrahedrally attached with four oxygen atoms.

→ Every oxygen atom is attached to two silicon atom

→ Ratio of atoms Si and O is 1:2

→ Both Si and O atoms are sp^3 hybridized

NATURE OF OXIDES

1. CO : covalent, neutral

2. CO_2 : Covalent, Acidic

$CO_2 + H_2O \rightarrow H_2CO_3$ (Acidic)

3. SiO_2 : covalent, acidic

$SiO_2 + NaOH \rightarrow 2Na_2SiO_3 + H_2O$ (Acidic)

4. SnO_2 : Ionic, Amphoteric

$SnO + NaOH + H_2O \rightarrow Na_2Sn(OH)_4$ (Acidic)

$SnO + HCl \rightarrow SnCl_2 + H_2O$ (Basic)

5. PbO : Ionic, Amphoteric

$PbO + 2HCl \rightarrow PbCl_2 + H_2O$ (Basic)

$PbO + NaOH + H_2O \rightarrow Na_2Pb(OH)_4$ (Acidic)

GROUP VII A ELEMENTS

- → Halogens due to salt forming tendency.
- → Electronic Configuration : $ns^2 p^5$
- → very reactive with non metals.
- → have highest E.N among the elements of their period
- → Fluorine is the most electronegative element of the periodic table.
- → Held together by weak Van der Waal's forces which explain volatile nature of them
- → Astatine is a radioactive element.

PHYSICAL APPEARANCE:

Flourine → Pale yellow gas
Chlorine → Greenish yellow gas
Bromine → Red brown liquid
Iodine → shiny greyish black solid

OXIDATION STATES

Fluorine → -1
Chlorine → -1, +1, +3, +5, +7
Bromine → -1, +1, +3, +5, +7
Iodine → -1, +1, +3, +5, +7

ATOMIC RADIUS

I > Br > Cl > F

ELECTRONEGATIVITY

F > Cl > Br > I

ELECTRON AFFINITY

Cl > Br > F > I

Fluorine has extraordinary small size and dense electronic cloud which decrease its electron affinity value.

MELTING AND BOILING POINT

F > Cl > Br > I

BOND ENTHALPY OF HALOGENS

→ Bond enthalpy is the amount of heat needed to break one mole of covalent bonds to form indiudual atoms

→ It begins with original substance in a gaseous state and ending with gaseour atoms

→ Bond breaking is endothermic so sign will be +ve.

$Cl_2 > Br_2 > F_2 > I_2$

The abnormal behaviour of F-F bond is due to its smaller size and repulsion between nuclei of fluorine atoms.

BOND ENTHALPY OF HX

HF > HCl > HBr > HI

HF, HCl → very stable to heat and donot decompose

HBr → may or may not decompose depending on temperatu..

HI → easily decomposed

ACIDITY OF HALOGEN ACIDS

HI > HBr > HCl > HF

HF is a weak acid due to intermolecular forces while HI is strong acid bcz it can give proton easily

STRENGTH OF HALOGEN AS OXIDIZING AGENT

$F_2 > Cl_2 > Br_2 > I_2$

→ Fluorine is most reactive halogen and most powerful oxidizing agent.

→ Fluorine and chlorine can oxidize various colored dyes to colorless substances for bleeching e.g litmus and universal indicator can be decolorized when exposed to fluorine or chlorine.

→ When chlorine water is added to KI solution, the solution becomes brown. This is due to formation of iodine i-e

$$2I^- \rightarrow I_2 + 2e^-$$

$$Cl_2 + 2KI \rightarrow I_2 + 2KCl$$

→ Chlorine can oxidize I^- to I_2 and Br^- to Br_2

REDUCING POWER OF HALOGEN ACID

HI > HBr > HCl > HF

* In a redox reaction, the hydrogen halides are oxidized to elemental halogens e.g

$$H_2SO_4 + 8H^+ + 8I^- \rightarrow 4I_2 + H_2S + 4H_2O$$

→ The larger the size of halide ion, the stronger is the reducing agent.

Bromide reduces sulphuric acid to SO_2. There is a decreas of oxidation state of sulphur from '+6' in H_2SO_4 to '+4' in SO_2

$$H_2SO_4 + 2H^+ + 2Br^- \longrightarrow Br_2 + SO_2 + 2H_2O$$

IMPORTANT POINTS

1. Milk of Magnesia $Mg(OH)_2$: used for treatment of acidity in stomach

2. ZnO is amphoteric

3. Size of cation (+ve) is always less than parent neutral atom

4. Size of anion (-ve) is always greater than parent neutral atom.

5. Metal oxides are ionic oxides, when they are in their lower oxidation state

6. Diagonal Relationship:

A diagonal relationship is said to exist b.w certain pairs of diagonally adjacent elements in the second and third periods of the periodic table. These pairs exhibit similar properties.

Only 3 pairs of diagonal relationship:

1. Lithium diagonal to Mg
2. Berryllium diagonal to Aluminium
3. Boron diagonal to Silicon

Boron is not a metal as its properties are similar to silicon. Some consider it metalloid.

7. Alkali metals always form ionic bond

8. Group I halides are more ionic (than Group II) and their ionic character increase down the group

9. Sodium carbonate when fused with sand forms sodium silicate which is commonly known as water glass

10. Fluorine can never be in positive oxidation state

11. Silica → SiO_2

12. The metals in their lower oxidation states are reducing agents (can lose e^-) while in their higher oxidation states are oxidizing agents (can gain e^-)

13. Water is said to be permanently hard when it contains sulphates of Na^+ and Mg^{+2} ions.

PERIODIC TABLE

* Acid Strength

    Along a Period → Increase
    Along a Group → Decrease

* Basic Strength

    Along a Period → Decrease
    Along a Group → Increase

* Melting Points

    - Metals generally possess a high melting point
    - Most non-metals posess low melting points

→ The non-metal carbon possess the highest boiling point of all elements

→ The semi-metal boron also posess a high melting point

* Metallic Character

Metallic character relates to the ability to lose electrons, and non-metallic character relates to the ability to gain electrons.

Along A Period: Metallic character decrease from left to right across a period. This is caused by the decrease in radius of the atom that allows the outer electrons to ionize less readily.

Along A Group: Metallic character increase down the group. Electron shielding causes the atomic radius to increase thus the outer electrons ionize more readily than electrons in smaller atoms.

* Reactivity Of Metals

Period → Decrease

Group → Increase

* Reactivity of Non-metals

Period → Increase

Group → Decrease

* OXIDES

Acidic Strength

Along Period : Increase

Along Group : Decrease

Basic Strength :

Along Period : Decrease

Along Group : Increase

* Basicity Order of Oxides of Third Period

$Na_2O > MgO > Al_2O_3 > P_4O_{10}$

* Acidity Order of Oxides of Third Period

$Cl_2O_7 > SO_3 > SO_2 > P_4O_{10} > SiO_2$