S and p block

~Sketch of Periodic table

Periodic table :-

table in which elements are arranged by increasing order of their atomic number. / groups, periods.

Groups :

vertical columns

no. of valence electron

Period :

horizontal rows

Shows number of shell / value of n

Blocks :-

divided into 4 :-

s. block :- electronic configuration end at 's' subshell

p- block :- electronic configuration end at 'p' subshell.
↳ IIIA – VIIIA or 13th – 18th

d- block :- electronic configuration end at 'd' subshell
↳ IB - VIIIB or 3rd - 10th

f. block :- electronic configuration end at 'f' subshell
↳ Lathanide or Actainide.

periodic and group properties

. atomic radius

Average distance between the nucleus and outermost shell, election is called atomic radius.

trend :-

down the group (Top to bottom) :-

- Increases
- As the atomic number and mass increases, the shell number also increase. This lead to significant increase in distance between nucleus and outermost electronic shell. [$H = 1s^1$, $Li = 1s^2 2s^1$, $Na = 1s^2 2s^2 2p^6 3s^1$]

Along the period (Left to right) :-

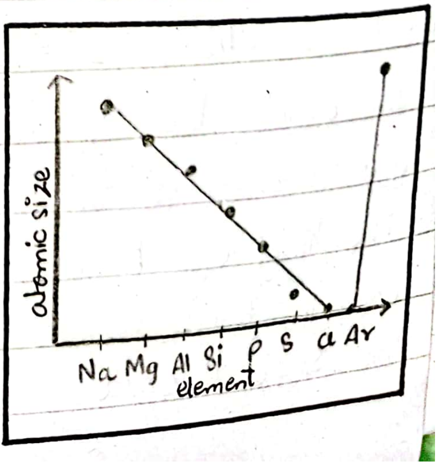
- decreases
- Nuclear charge increases. leading to greater attractive force between nucleus and outermost shell hence shrinking of size.

[$Na = 1s^2 2s^2 2p^6 3s^1$, $Mg = 1s^2 2s^2 2p^6 3s^2$, $Al = 1s^2 2s^2 2p^6 3s^2 3p^1$]

+1 +2 +3

except :-

-. Argon having larger atomic size than chlorine?



- Noble gases including Argon tend to having exceptional atomic size because of their sophisticated property of inter-electronic repulsion and shielding effect.
- When octet gets completed, the e's

start to produce repulsive forces and have better shielding effect leading increase size.

•-ionic radius

Average distance between the nucleus and outermost e^- in an ion.

anion (-)

Cation (+).

radius of cation

- Radius of cation is always less than its parent atom.
- Na = 1.54A°, Na^{+1} = 0.95A°
- In respective cations, the number of total electrons are less than protons. This indeed decreases the shielding effect but also in result to that increases effective nuclear charge. That causes more attraction, shrinking the size all overly.

radius of anion

- Radius of anion is always greater than its parent atom.
- Cl = 0.99A°, Cl^- = 1.81A°
- In anion, the number of total electrons are greater than proton. This increases significant shielding effect but also decreases nuclear charge that causes less attraction, expanding the size overall.

. ionization energy

Ionization energy, potential is the minimum amount of energy require to remove outermost electron from gaseous atom.

$X_{(g)} \longrightarrow X^{+}_{(g)} + 1e^{-}$

$M_{(g)} \longrightarrow M^{+}_{(g)} + 1e^{-}$

$Na_{(g)} \longrightarrow Na^{+}_{(g)} + 1e^{-}$ I.E = + 496kJ/mol

note:

- To remove first e^{-} is easy.
- But grodually removing e^{-} one by one, we need high I.E.

$I.E_3 > I.E_2 > I.E_1$

$Mg_{(g)} \longrightarrow Mg^{+1} + 1e^{-}$ I.E = 738 kJ/mol

$Mg^{(+)}_{(g)} \rightarrow Mg^{+2} + 1e^{-}$ I.E = 1451 kJ/mol

$Mg^{+2}_{(g)} \rightarrow Mg^{+3}_{(g)} + 1e^{-}$ I.E = 7730 kJ/mol

Types:-

Low I.E:- more easy to remove e^{-}

High I.E:- difficult to remove e^{-}

Trend:-

Dawn The Group:-

→ I.E decreases.

→ The Atomic Size increase leading increase distance b/w outermost electron and nucleus. Furthermore, force of attraction decrease. Hence less energy will be required to remove e^{-}.

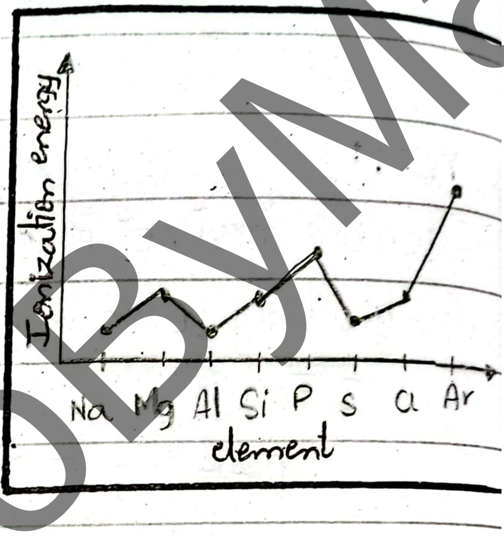
Along the period:

→ Increases

→ The atomic size decrease meaning that more force of attraction will be present leading to a difficult way by providing more energy.

exception:-

- Al has less I.E than Mg?
- S has less I.E than P?



Al and Mg:

- The full-filled orbitals are more stable than half filled and then partially filled.
- Mg = $1s^2 2s^2 2p^6 3s^2$
- Al = $1s^2 2s^2 2p^6 3s^2 3p^1$

A/o the given rule, the Aluminim is partially filled leading to unstability that's why it need less energy than magnesium because it stable.

S and P:-

- S = $1s^2 2s^2 2p^6 3s^2 3p^4$
- P = $1s^2 2s^2 2p^6 3s^2 3p^3$

A/o to given rule, the Phospurus is quite half filled than Sulphur that's partially fully. That's why phosphorus requires more energy than Sulphur.

. electronic affinity (KJ/mol)

The amount of energy released or absorbed when electron is added to gaseous atom to form negative ion.

$F(g) + 1e^- \longrightarrow F^-$ E.A = -ve

(uninegative ion)

↓
negative = Exothermic.

1st and Successive E.A :-

- The first E.A is negative because element/atom wants to be stable by releasing energy.
- The 2nd E.A is positive because atom wants to reduce inter-electron repulsion inorder to gain another e^- that's why its positive.

$O(g) + 1e^- \longrightarrow O^-(g)$ $E.A_1 = -141$ KJ/mol

$O^-(g) + 1e^- \longrightarrow O^{-2}(g)$ $E.A_2 = +780$ KJ/mol

trends:-

down the group:-

- Decreases
- The atomic size increases, the distance between outermost electron and nucleus decreases, the addition of electron.

along the group:

- Increases
- The atomic size decreases, the force of attraction increases thus addition of electron

exception or anomalous trend:-

Why electron affinity of Cl is greater than F?
The flourine has small atomic size. Thus the inner electron causes electronic repulsion due which incoming electron faces difficulty to be add and hence it has low electronic affinity than Cl.

Why group IIA, VA, VIIIA has low E.A along the period?

- The group IIA elements are stable at s^2. Thus stability causes incoming e^- to have repulsion and leading to low electron affinity value.
- The group VA is half filled at ns^2np^3. Thus difficult to add.
- The group VIIIA elements are stable at ns^2np^6. They are highest stable. Thus low electron affinity.

factors affecting

- atomic size $\propto \frac{1}{E.A}$
- shielding effect $\propto \frac{1}{E.A}$
- Nuclear charge $\propto$ E.A

• electronegativity no unit

The power of an atom to attract shared paired of e towards itself in bonding.

- measured by pauling scale.

most and least E.N:-

F = 4.0 , Cs = 0.7
O = 3.5 Fr = 0.7
N = 3.0

trend:-

down the group:-

- electronegativity decrease
- Atomic size increase leading to increase distance b/w nucleus and outer-most electron. Thus force of attraction decrease.

along the period:

- electronegativity increase
- Atomic size decrease, leading to small distance b/w nucleus and outer-most electron. Thus force of attraction increase.

exception:

Why Argon donot have E.N value rather Cl have the highest?

E.N is not assigned to noble gases because it's highly Stable due to its complete octet Thus it doesnot participate in any type of bonding.

Ar = $1s^2 2s^2 2p^6 3s^2 3p^6$

factors affecting :-

- Atomic size $\propto \frac{1}{E.N}$
- shielding effect $\propto \frac{1}{E.N}$.
- Nuclear charge $\propto$ E.N.

Bonds and E.N

ΔE.N = type of bond
0 - 0.4 = non-polar covalent
0.4 - 1.8 = Polar bond
1.87 = bond ionic

physical properties

. electrical conductivity

The ability of substance to conduct electric current due to presence of free electrons | ions.

factors :-

electrical conductivity $\propto$ no. of free electrons
electrical conductivity $\propto$ no. of free ion
electrical conductivity $\propto$ electropositivity

metals and electrolyte :-

→ metals can conduct electricity due to no. of free electrons.

→ electrolyte can conduct electricity due to no. of free ions.

trend:

along period: 3rd

- Na, Mg, Al are good conductors. (+1, +2, +3)
- electrical conductivity increase from Na to Al due to increase no. of free electrons.
- Silicon is a semi-conductor. (metaloid)
- P, S, Cl, Ar are non-conductor of electricity because they don't have free electrons.

down the group: IA, IIA, IVA

- increase down the group.
- The Atomic size increase, thus the respective distance b/w nucleus and valence electron So, the force of attraction b/w nucleus and valence electron will be decreased thus it will easily loose e^- hence conductivity will be carried out in easier manner.

Coinish / Coinage metal (mcq)

↳ Group IB (copper, silver, gold)

- These have extra-ordinary electrical conductivity.

• melting and boiling point

transition temp

↳ The Temp at which solid is converted into liquid but a equilibrium is maintained.

The temperature at which liquid is converted into gases but equilibrim is maintained.

factors :-

Melting /Boiling point $\propto$ No.of Binding e's

Melting/ Boiling point $\propto$ Intermolecular forces

Comporision :-

Group IIA melting point/ boiling > Group IA melting/ boiling point

Trend :-

down the group :-

- decreases except halogen and noble gases
- Group IIA has no trend.
- Atomic size increase, the distance also increases thus the force of attraction decrease, hence bonding e^- becomes loosely held.

• halogen and noble cases

- The atomic size increase, Thus they are non-polar molecules.
- As size increases, distance increases, the london dispersion force increases. So M.P & B.P also increases.

along period :-

- Increase upto group IVA and then decrease from VA to VIIIA.

Why : Sulphur , has greater m.p and b.p compared to phosphorus

The phosphorus exists in P_4 form having small

size creating small london dispersion force and less m.p and b.p.
But Sulphur is non-polar having exceptional increase size that causes more london dispersion force having high m.p and b.p. x———x

- No Of binding e^- increases from IA to IVA group element or IVA have giant covalent structure.
- M.P & B.P decreases from VA to VIIIA due to simple covalent structure size. In addition the size is small hence LDF is less.

chemical rxn of 3rd peri

third period

Na Mg Al Si P S Cl Ar

a- water

Sodium:-

- Sodium has a highly exothermic reaction with cold water.

$$2Na + \underset{(cold)}{2H_2O} \longrightarrow 2NaOH + H_2 + \text{highly exothermic}$$

Magnesium:-

- Magnesium has slight reaction with cold water.

$$Mg + 2H_2O \longrightarrow Mg(OH)_2 + H_2.$$

$$Mg + \underset{(steam)}{H_2O} \longrightarrow MgO + H_2$$

Aluminium:-

- Aluminium <u>metal</u> neither reacts with cold water nor steam
- Aluminium <u>powder</u> surely reacts but slow.

$$\underset{(powder)}{2Al} + 3H_2O \longrightarrow Al_2O_3 + 3H_2.$$

- During process, it actually forms Al_2O_3 non porous unreactive layer thus some unreactive Al is left behind.

Silicon, sulphur, phosphorus :-

- Silicon forms a protective oxide layer (SiO_2) on its surface when exposed to air which prevents rxn with water. In addition, its strong covalent network makes it stable.
- Sulphur exists in several allotropes but S_8 is stable making it unreactive also its non-polar.
- Phosphorus exists in several allotrope that doesnot reacts with water also it makes protective layer.

→ It doesn't react.

Chlorine :

- Cl_2 dissolves in water to some extent and forms green solution.
- Reaction is reversible.

+1

$$2Cl_2 + 2H_2O \rightleftharpoons 2HOCl + 2HCl$$

$$\underline{2HOCl \xrightarrow{\text{sunlight}} 2HCl + O_2}$$

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

Argon :

The Argon is unreactive with water.

- full valence shell
- Inert nature
- No chemical affinity
- Non-polar nature

b) oxygen

Sodium :-

→ burns in Oxygen with yellow flame to produce solid mixture of Na_2O & Na_2O_2

$$4Na + O_2 \xrightarrow{\text{limited}} 2Na_2O$$

$$2Na + O_2 \xrightarrow{\text{excess}} Na_2O_2$$

Magnesium :-

→ Magnesium burns in Oxygen with an intense white flame

$$2Mg + O_2 \longrightarrow 2MgO$$

Aluminium :-

→ Al reacts with O_2 to form thin protective oxide layer to prevent metal from further corrosion.

$$4Al + 3O_2 \longrightarrow \underbrace{2Al_2O_3}_{\text{amphoteric oxide}}.$$

→ burns with O_2 at high temp of 800°c

Silicon :-

→ Burns when strongly heated producing a Giant covalent structure

$$Si + O_2 \longrightarrow SiO_2$$

Phosphorus :-

→ White phosphorus (P_4) catches fire spontaneously in air burning with white flame and producing a smoke of (P_4O_6 & P_4O_{10}).

→ Products depend upon supply of O_2.

$$P_4 + 3O_2 \xrightarrow{\text{limited}} P_4O_6$$ (Phosphorus III oxide)
(diameric form of P_4O_3).

$$P_4 + 5O_2 \xrightarrow{\text{excess}} P_4O_{10}$$ (Phosphorus V oxide)
(diameric form of P_4O_5).
→ dehydrating agent.

Sulphur :-

Sulphur burns in air or O_2 with blue pale flame

$$S + O_2 \rightarrow SO_2$$
→ colorless gas.

Argon :- / chlorine :-

- Chlorine and Oxygen donot directly react under normal conditions because both are diatomic molecule with strong covalent bonds making them relatively stable.
- Thus producing conditions like catalysts, high temp will cause them to react producing Cl_2O & Cl_2O_7.

Cl_2O → diachloro monoxide

Cl_2O_7 → dichloro heptaoxide

- Argon doesn't react because.
  - full valence shell
  - Inert nature
  - No chemical affinity
  - Non-polar nature

c) chlorine

a) Sodium:-

burns with a bright orange flame.

$$2Na + Cl_2 \rightarrow \underbrace{2NaCl}_{\text{white solid}}$$

Magnesium:-

burns with its usual intense white flame

$$Mg + Cl_2 \rightarrow \underbrace{MgCl_2}_{\text{white solid}}$$

Aluminium:-

→ reacts when Cl_2 gas passes over aluminium foil heated in long tube to produce pale-yellow product.

$$2Al + 3Cl_2 \rightarrow 2AlCl_3$$

Silicon:-

→ reacts with Silicon powder in tube.

$$Si + 2Cl_2 \rightarrow \underbrace{SiCl_4}_{\text{colorless liquid}}$$

Phosphorus:-

→ White phosphorus req (P_4) reacts with diff amount of chlorine to produce diff products

$P_4 + 6Cl_2 \xrightarrow{\text{limited}} \underbrace{4PCl_3}_{\text{colurless fuming liq}}$ → phosphorus trichloride

$P_4 + 10Cl_2 \xrightarrow{\text{excess}} \underbrace{4PCl_5}_{\text{off white} \rightarrow \text{yellow solid}}$ → phosphorus Pentachloride

Sulphur:-

- Stream of Cl_2 passes over heated S, to produce orange evil-smelling liq.

$2S + Cl_2 \rightarrow S_2Cl_2$ → disulphur dichloride
Sulphur monochloride

Argon:-

- Argon don't react due to all above mentioned reasons.

Physical Properties

oxide

a) Oxides:-

Na_2O, MgO, Al_2O_3, SiO_2, P_4O_{10}/P_4O_6, SO_3/SO_2, Cl_2O_7/Cl_2O.

→ These all are known as Higher Oxides.
→ Their outer shell is wholly involved in bond formation.

b) Structure :-

The structure varies :-

Na_2O, MgO, Al_2O_3 = metallic structure (ionic giant)

SiO_2, = giant Covalent structure,

P_4O_x, SO_x, Cl_2O_x = Simple molecular structure

c) melting and Boiling point :-

i) Na_2O, MgO, Al_2O_3, SiO_3

They have high melting and boiling point because they giant ionic Structure. Alot of energy is required to break the bonds.

ii) P_4O_x, SO_x, Cl_2O_x

They have comparatively low melting and boiling point because they are covalent structure having london dispersion & Dipole & dipole forces.

d) electrical Conductivity :-

Solid metallic Oxide

They are non-conductor.

Molten metallic Oxide

They are good conductor.

- **Na_2O, MgO, Al_2O_3**

They are non-conductor in solid State but they conduct in liq/molten or aqueous form.

- **SiO_2, P_4O_x, SO_x, Cl_2O_x,**

They are non-conductor, can't be ionizable.

acid-base behavior

Na_2O, MgO, Al_2O_3, SiO_2, P_4O_{10}/P_4O_6, SO_3/SO_2 ; Cl_2O_7/Cl_2O

strongly basic ; basic, amphoteric, weakly basic, Acidic, strongly acidic, very strongly acid

a) Sodium oxide

Sodium oxide Na_2O is a strong bas. because it has O^{-2} being a strong base.

i- reaction with water :-

$$Na_2O + H_2O \rightarrow 2NaOH$$

- pH=14
- reacts exothermically with cold water

ii- reaction with acid :-

$$Na_2O + 2HCl \rightarrow 2NaCl + H_2O$$

- strong base always reacts with acids to produce salt showing a valid justification

b) Magnesium oxide

It's basic due to presence of O^{-2}. But as comparative to Na_2O it's not that much strong because O^{-2} are not truly free. As, there's strong attraction between O^{-2} & Mg^{+2} increasing its lattice energy.

i- reaction with water :-

- white MgO powder reacts with H_2O
- $Mg(OH)_2$ is produced that's insoluble
- Most of hydroxide doesn't get into sol.

$MgO + H_2O \rightarrow Mg(OH)_2$

- pH = 9.

ii) Reaction with acids

- reacts with warm dilute HCl

$$MgO + 2HCl \rightarrow MgCl_2 + H_2O$$

O) Aluminium oxide

It's amphoteric :-

Why?

→ The presence of ionic oxide allow it accept (H^+) making it basic.

→ The ion having high charge density & partially covalent structure allow it to donate (H^+) making it acidic.

i- Reaction with acid

$$Al_2O_3 + 6HCl \xrightarrow[(hot)]{} 2AlCl_3 + 3H_2O.$$

ii- Reaction with base

→ reacts with hot, conc NaOH to form colorless sol.

$$Al_2O_3 + 2NaOH + \overset{3}{\cancel{H}}H_2O \rightarrow 2Na[Al(OH)_4]$$

Sodium tetra hydroxo Aluminate III

iii- Reaction with water

doesn't react due to strong oxide.

d) Silicon dioxide

i- **Reaction with water :-**

It's insoluble in water due to it large & giant covalent structure.

ii- **Reaction with base :-**

- reacts with hot & conc. NaOH

$$SiO_2 + 2NaOH \rightarrow Na_2SiO_3 + H_2O.$$

iii- **Reacts with further**

reacts with Na_2O, K_2O, CaO, PbO, ZnO or mixture to form silicates & glasses.

e) phosphorus oxide

i- **Reaction with water :-**

- Phosphorus V oxide reacts violently with water.
- Giving a sol containing mixture of acids.

$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$ (Phosphoric acid)

$P_4O_6 + 6H_2O \rightarrow 4H_3PO_3$ (phosphorus acid)

ii. **Reaction with base :-**

- $P_4O_{10} + NaOH \rightarrow Na_3PO_4$

↳ sodim hydroxide sol.

or

$H_3PO_4 + NaOH \rightarrow NaH_2PO_4 + H_2O$

$NaH_2PO_4 + NaOH \rightarrow Na_2HPO_4 + H_2O$

$Na_2HPO_4 + NaOH \rightarrow Na_3PO_4 + H_2O$

↳ acidic.

f- Sulphur oxide

i- SOx reaction with water:-

SO2 is fairly sol in water.

$$SO_2 + H_2O \rightarrow H_2SO_3$$

SO3 reacts violently with H2O to produce fog of Conc. H2SO4 droplets

$SO_3 + H_2O \rightarrow H_2SO_4$ → strong acidic

ii- SOx reaction with Base:-

SO2 :-

$SO_2 + 2NaOH \rightarrow Na_2SO_3 + H_2O$ sodim sulphite.

$$Na_2SO_3 + H_2O + SO_2 \rightarrow 2NaHSO_3$$

$$H_2SO_3 + 2NaOH \rightarrow Na_2SO_3 + 2H_2O$$

SO3 :-

$$SO_3 + 2NaOH \rightarrow Na_2SO_4 + H_2O$$

$$Na_2SO_4 + H_2O + SO_3 \rightarrow 2NaHSO_4$$

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

CaO :-

$CaO + SO_2 \rightarrow CaSO_3$ (calcium sulphite)

$CaO + SO_3 \rightarrow CaSO_4$

g) Chlorine oxides

i- Cl2Ox reaction with water:-

chlorine VII oxide :-

$Cl_2O_7 + H_2O \rightarrow HClO_4$ (Perchloric acid).

$Cl_2O + H_2O \rightarrow 2HOCl$ (Hypchlorous acid).

~~Base+~~

ii- Reaction with base :-

Cl_2O_7 :-

$2NaOH + Cl_2O_7 \rightarrow 2NaClO_4 + H_2O$

$NaOH + HClO_4 \rightarrow NaClO_4 + H_2O$

Cl_2O :-

$2NaOH + Cl_2O \rightarrow 2NaOCl + H_2O$

$NaOH + HOCl \rightarrow NaOCl + H_2O$

order of acidity

$HClO_4 > HClO_3 > HClO_2 > HClO > HCl$.

Chloride of Peroid 3

Nacl, $MgCl_2$, $AlCl_3$, $SiCl_4$, PCl_3/PCl_5, S_2Cl_2

• Sodium Chloride :

- ionic compound
- High melting & Boiling point
- have strong electrostatic force of attraction
- It's non-conductor in solid but it is a good conductor in molten, aqueous form
- forms neutral solution in water.

$$NaCl \rightleftharpoons Na^+ + Cl^-$$

•Magnesium chloride :-

- Ionic Compound
- High melting & Boiling point
- have strong electrostatic force of attraction
- It's non-conductor in solid form but in water it forms faintly acidic sol.

$$MgCl_2 + 6H_2O \rightarrow [Mg(H_2O)_6]^{+2} + 2Cl^-$$

$\downarrow$ pH = 6

hexaaquamagnesim II ion

• Aluminium Chloride :-

- ionic compound but alot of covalent character
- Sublimation occurs at 180°c
- Non-conductor.
- Converted into Al_2Cl_6.
- forms acidic sol with wate

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

$AlCl_3 + H_2O \rightarrow [Al(6H_2O)_6]^{+3} + 3Cl^-$

order :-

$NaOH > Mg(OH)_2 > Al(OH)_3$.

- leftover from book.

Group-IA

Chemical Properties

Li, Na, K, Rb, Cs, Fr

reaction of water

- Reactivity of IA element increases down the group because metallic character increases
- Reaction will be fast & highly exothermic producing metal hydroxide & hydrogen gas
- for their special characteristic, they are stored in Kerosene/ paraffin oil.

$2Na + 2H_2O \longrightarrow 2NaOH + H_2 +$ exothermic

- The K, Rb, Cs are so reactive that they even react with ice at -100°C.

b) rxn with oxygen

- Alkali metals react with O_2 or air rapidly & thus get tarnished by the formation of their oxide.

$$4Li + O_2 \longrightarrow 2Li_2O$$

- Sodium burns **yellow/golden flame** in O_2 to form peroxide.

$$2Na + O_2 \longrightarrow Na_2O_2$$

- All the metals after Na forms superoxide.

$$K + O_2 \longrightarrow KO_2$$

Q. How to form normal oxides of Sodium?

Normal oxide is not produced by direct method rather via indirect procedures :-

Peroxide :

$$Na_2O_2 + 2Na \longrightarrow 2Na_2O +$$

Sodium nitrate :

$$2NaNO_3 + Na \longrightarrow Na_2O + N_2$$

Sodium nitrite :

$$2NaNO_2 + 6Na \longrightarrow 4Na_2O + N_2$$

oxidation of oxide :-

-2

oxidation of peroxide :-

-1

oxidation of superoxide :-

$-1/2$

properties

Q. Write down the reactions of different types oxide with water.

Following are the reactions.

normal oxide:-

$$\underbrace{O^{-2} + \underbrace{H_2O \rightarrow OH^-}_{\text{loss } H^+}}_{\text{Gain } H^+} + OH^-$$

Peroxides :-

$$O_2^{-2} + 2H_2O \rightarrow 2H_2O_2$$

Superoxides :

$$O_2^{-1} + 2H_2O \rightarrow 2H_2O_2 + O_2.$$

rxn with halogen

- All alkali metals reacts with halogen (Cl₂) to produce halide
- Li, Na, [illegible] reacts slowly with halogen.
- All other IA elements have vigorous rxn with halogen.

$$2Na + Cl_2 \rightarrow 2NaCl$$

Group - II

Chemical properties

Be, Mg, Ca, Sr, Ba, Ra

Rxn with water

- Reactivity of IIA element increases down the group.
- Be doesn't react with water even at high temp.
- Mg burns in steam to produce

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

- It reacts slightly with cold water that lasts for few min. Some bubbles of H_2 seen on surface. After that reaction stops because $Mg(OH)_2$ is produced that's insoluble & forms protective layer.

$$Mg + H_2O \longrightarrow Mg(OH)_2 + H_2$$

- Remaining reacts vigorously with cold water.

$$M + 2H_2O \longrightarrow M(OH)_{2\,(aq,s)} + H_2.$$

rxn with oxygen

- All alkaline earth metals reacts with oxygen to form normal Oxide.

$$2M + O_2 \longrightarrow 2MO.$$

$$2Ca + O_2 \longrightarrow 2CaO.$$

- Barium reacts with O_2 to form Peroxide

$$2Ba + O_2 \xrightarrow{excess} Ba_2O_2$$

- Barium oxide is amphoteric.

Proof:

$BeO + H_2SO_4 \rightarrow BeSO_4 + H_2O$.

$BeO + NaOH \rightarrow Na_2BeO_2 + H_2O$.

Normal oxide :-

$MCO_3 \xrightarrow{\Delta} MO + CO_2$

$BaCO_3 \xrightarrow{\Delta} BaO + CO_2$

- BaO_2 is basic oxide

$BaO_2 + H_2SO_4 \rightarrow BaSO_4 + H_2O_2$

rxn with nitrogen

- All IIA element react with nitrogen & form nitride.

$3Mg + 2N \longrightarrow Mg_3N_2$

- Nitride of group IIA when dissolve in water produce NH_3 gas & hydroxide.

$Mg_3N_2 + H_2O \rightarrow Mg(OH)_2 + NH_3$

- Be_3N_2 is volatile & other all are non-volatile.

trends of Solubility of Hydroxides, Sulphates, Carbonate

Solubility depend on 2 factors :-

- Lattice energy ΔH_L
- Hydration energy ΔH_h.

amount of energy 1 mole of ionic ion in gaseous state forms 1 mole of ionic compound. (can be endo/exo).

amount of energy released when 1 mole of gaseous ion is dissolved in water to forms infinite dil solo

→ $Na^+ + Cl^- \longrightarrow NaCl$ $\Delta H_L = -787\ kJ/mol$

$NaCl \longrightarrow Na^+ + Cl^-$ $\Delta H_b = +787\ kJ/mol$

→ $H^+ + H_2O \longrightarrow H_3O^+$ $\Delta H_n = -ve$.

Imp:-

- If hydration energy > Lattice energy of Compound
  - → compound is soluble.

- If hydration energy < Lattice energy of Compound

- Hydration energy ∝ Solubility, charge density
- Charge density ∝ $\frac{charge}{size}$

Bulky anion:

CO_3^{-2}, SO_4^{-2}, PO_4^{-2}, NO_3^{-1}.

Small anion:

O^{-2}, OH^-, Cl^-

- lattice energy ∝ $\frac{1}{solubility}$
- lattice energy ∝ $\frac{1}{size}$
- Solubility ∝ difference of size of anion & cation

: Hydroxides

- Solubility of IIA elements increases down the group.

  ↳ Lattice energy $\propto \frac{1}{\text{size/solubility}}$

- Size is increasing, lattice energy is decreasing, solubility inc.

$Be(OH)_2$, $Mg(OH)_2$, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

insoluble, slightly soluble, most soluble

$\xrightarrow{\hspace{4cm}}$

increase in solubility

- Ra is radioactive, its hydroxide are not discussed
- Smaller ion (OH^-) attached with IIA, lattice energy overcomes hydration energy.

: Carbonates & Sulphates

- The solubility decreases down the group.

$BeCO_3$, $MgCO_3$, $CaCO_3$, $\underbrace{SrCO_3, BaCO_3}_{\text{insoluble}}$

$\xrightarrow{\hspace{4cm}}$

decreases.

$BeSO_4$, $MgSO_4$, $CaSO_4$, $\underbrace{SrSO_4, BaSO_4}_{\text{insoluble}}$

$\xrightarrow{\hspace{4cm}}$

decrease

Bulky anions CO_3^{-2}, SO_4^{-2} is attached to, ~~lattice energy~~ So hydration energy overcomes lattice energy.

Hydration energy $\propto$ Solubility
Hydration energy $\propto$ charge density
charge density $\propto \frac{\text{charge}}{\text{size}}$

As we move down group, the size increases, leading to decrease in charge density thus hydration energy decrease so solubility does so.

Group IA CO_3^{-2} vs Group IIA CO_3^{-2}

Group IA are large enough in size having low lattice energy so they are more soluble. On the other hand, the size is small leading to high enough lattice energy hence decreasing solubility.

Group IV

physical properties

1. melting & boiling point

Melting & boiling point decreases down the group but not in regular manner.

→ Carbon & Silicon have high melting point due to giant structure.

→ Tin (Sn) has low melting point due to distorted 12 coordinate structure than Pb (lead).

trends from non-metal to metal in group IV element

- The metallic character increase down the group. in group IV elements.
  - ↳ The atomic size increases
  - ↳ The internuclear distance increase
  - ↳ The effective nuclear charge decrease
  - ↳ Thus it become good metal.

- $\underbrace{C, Si}_{\text{non-metal}}, \underbrace{Ge}_{\text{metalloid}}, \underbrace{Sn, Pb}_{\text{metals}}$

$\xrightarrow{\hspace{4em}}$ increase in metallic character.

Oxidation State

Apparent charge that is positive or negative on atom.

$\underbrace{C, Si}_{\substack{+4 \\ \text{compound} \\ \text{stable}}}, \underbrace{Ge, Sn, Pb}_{\substack{+2, +4 \\ \text{compound} \\ \text{stable}}}$

- Relative stability of +2 oxidation state increase down in Group IVA due to increase in inert pair effect.
- Relative stability of +4 oxidation state increase down in Group IVA.
- **inert pair effect:-**

The ns^2 pair of electron remains inert & does not take part in bond formation. This pair of ns^2 electrons are called inert pair of e^- & the effect is called inert pair effect.

trend:-

- The inert pair effect increase down the group.
  ↳ (post-transition elements).
    ↳ IIIA, IVA, VA
- This effect is shown by those element whose np^2 e^-s take part in bonding & ns^2 doesn't thus showing +2 oxidation state.

Order:-

+2:-

$Ge^{+2} < Sn^{+2} < Pb^{+2}$

→ highest inert pair of $\bar{e}$ ~~pair~~

+4:-

$Ge^{+4} > Sn^{+4} > Pb^{+4}$

↳ when all ns^2p^2 $\bar{e}$s are lost.

Group VII

Strength of Halogen as oxidizing agent

$F_2 > Cl_2 > Br_2 > I_2$.

★ Strongest Oxidizing agent among halogen is due:-

→ Flourine has small atomic radius, which allow it to attract $\bar{e}$s more effectively.

→ florine has most electronegativity of nearly 3.98. This allow it to strongly attract e^- from other atoms.

→ Therefore, it reduces itself but oxidizes other.

factors affecting oxidizing power of halogen:-

→ electron affinity

halogen have high e.a, meaning that they readily gain e^-s to form neg ions. A high affinity means a stronger tendency to accept e^-s, enhancing oxidizing power.

→ Hydration energy

high hydration energy stabilizes anion after halogen gains e^-. This stabilization makes reduction process more favorable.

→ Bond dissociation energy

Lower bond dissociation energy means that the halogen atom can more easily dissociate into individual atoms which can act as oxidizing agent.

→ Heat of vaporization

lower heat of vaporization means that the halogen can more easily transition to gaseous state, where it can act as oxidizing agent.

acidity:-

$HI > HBr > HCl > HF$

→ The HI is considered as 'strongest acid'

→ The HF is not considered as "strong" rather counted as 'weak acid'.

reason :-

- → The size of atom down the group increase.
- → So far, the distance also increases.
- → This case a significant drop in force of attracting creating a ease to break bond. & release H^+.

reducing agent

$I^- > Br^- > Cl^- > F^-$

→ I^- is counted as strong reducing agent. (It oxidize itself & reduce other).

reason :-

- large size
- Increase distance
- less force of attraction
- easy to lose e^-s.

redox rxn with conc. H_2SO_4 :-

Chloride :

$$2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$$

It's not good reducing agent.

Bromide :-

$$2NaBr + H_2SO_4 \rightarrow Na_2SO_4 + 2HBr.$$

$$H\overset{-1}{Br} + H_2\overset{+6}{S}O_4 \rightarrow \overset{0}{Br_2} + \overset{+4}{S}O_2 + H_2O.$$

oxidation

reduction

Bromide is a good reducing agent.

Iodide :-

$$NaI + H_2SO_4 \rightarrow Na_2SO_4 + HI$$

$$\overset{-1}{HI} + H_2\overset{+6}{S}O_4 \rightarrow \overset{0}{I_2} + H_2\overset{-2}{S} + H_2O$$

oxidation

reduction

- It is excellent reducing agent.

SLO

Q. Why phosphorus trichloride is more acidic than the silicon dioxide or aluminium trichloride.

- The acidity of a compound usually depends on the ability to donate proton (H^+) or accept $\bar{e}s$.
- PCl_3 is a lewis acid, which can mean accept pair of $\bar{e}s$.
- The phosphorus atom has an empty d-orbital which can accept $\bar{e}$.
- further its reaction with water produces 2 types of acids:-

$$PCl_3 + 3H_2O \rightarrow H_3PO_3 + 3HCl$$

Q. Which of the following gives acidic solution with water $MgCl_2$, $AlCl_3$, $SiCl_4$, $NaCl$.

By evaluating these compound, we can see that $AlCl_3$ & $SiCl_4$ gives acidic solution only.

- $2NaCl + 2H_2O \rightarrow 2Na(OH) + \cancel{H_2O}\, 2HCl$
- $MgCl_2 + H_2O \rightarrow [Mg[H_2O)_6]^{+2} + 2Cl^-$
- $AlCl_3 + H_2O \rightarrow [Al(H_2O)_6]^{+3} + 3Cl^-$
- $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4HCl$

They give acidic solution.

Q Why NaCl have high melting & boiling point.

- NaCl have high melting & boiling point due to its ionic bond.
- In an Ionic bond, the electrostatic forces b/w oppositely charged ions (Na^+ & Cl^-) are very strong, requiring a significant amount of energy to break these bonds.

Q. Why PCl_3 is liquid but PCl_5 is solid?

- The PCl_3 is a covalent molecule & exist as liquid at room temp. Its intermolecular forces, like Van der Waals forces & dipole - dipole interaction are relatively weak compared to ionic thus remain liquid.
- The PCl_5 exists as solid at room temperature. This is because PCl_5 forms crystal lattice

structue in solid state.

Q. Why aluminium chloride is non-conductor in solid as well as liquid but NaCl & $MgCl_2$ are conductor in liq state & non-conductor in solid state.

- In solid state, $AlCl_3$ forms a network of covalent bonds. It doesn't have free e^-s or ions which are required for electrical conductivity. Thus non-conductor.
- When melted $AlCl_3$ exists as discrete molecules rather than ions. In this state, It doesn't have free moving charged particles to conduct electricity. making it non-conductor.
- But in case of NaCl & $MgCl_2$, they form crystal Lattice & in aqueous they form free ions.

Q. Why lithim gives red color in flame tests.

- Lithium ions emit a characteristic red color when heated in flame.
- They have small size leading to large ionization energy this means less energy is required for excitation thus more λ visible light is shown that's red.

Q. Why alkali metals shows an Oxidation state of +1 only?

Alkali metals exhibits an oxidation state of +1 due to their tendency to lose only $1e^-$.

Q. Why alkali metals are good reducing agents.

The Alkali metals are good reducing agents due to their following :-

- Alkali metals have single e^- in their outermost shell.

They have large size so weak nuclear forces thus low I.E.

They are good reducing agent therefore.

Q. Why sodim is stored under Kerosene oil & why alkali metals forms +1 rather than +2?

- Sodim is extremely reactive metal. When exposed to air, it reacts with Oxygen & moisture forming Sodim Oxide & Sodim Hydroxide.
- Stored in Kerosene allow it prevent from coming in contact with air & moisture.

⟷

The alkali metals have only $1e^-$ in outermost shell. & it becomes stable. But removing 2 electrons makes it unstable & requires high energy.

Q. Why $MgSO_4$ is more soluble than $BaSO_4$ ...

$MgSO_4$ is more soluble in water than $BaSO_4$ due to difference in lattice & hydration energy.

- The $MgSO_4$ being smaller results in higher charge density allow them to strongly interact with water molecules.
- The $BaSO_4$ being larger results in lower charge density allow them to less strongly attract with water molecules.

Q. Why do hydroxide of group 2 metals are weaker basis than those of group I metals.

The reasons are following:

- Lattice Energy & Ionization Energy

  Group 2 are less soluble in H_2O than Group 1 due to higher lattice energy.
- Charge density of metal ions.

  Group 2 metals form ions, which have high charge density than ion of 1.
- Solubility trends

  Group I are more likely to be soluble than II.

Q. Why potassium hydroxide is stronger base than the barium hydroxide.

- The KOH is highly soluble in water. When it dissolve, it dissociates completely into K^+ & OH^-. This higher solubility making it stronger base.
- The $Ba(OH)_2$ is less soluble in water compared to KOH. It completely doesn't dissociate into ions.

Q. Which of the following metals are semi-conductor.

a) Pb b) C c) Si d) Ge

The Si & Ge are go-to used in Semi-Conductor.

Q. The HI is stronger acid than HF?

The HI is stronger acid than HF due to:-

- larger size of atom
- less polarity
- Stable conjugate base.

Q. Electron affinity of flourine is less than chlorine.

The florine has small size thus when e^- is added to 2p orbital it becomes quite crouded. This experience significant repulsion from existing e^-. In chlorine, its opposite

Q. Chlorine is stronger oxidizing agent than the iodine.

Yes, the chlorine is strong oxidizing agent.

- more electronegative
- effectively gain e⁻
- less bond dissociation energy

Q. Why florine is better oxidizing agent than the iodine.

This is because :-

- flourine is most electronegative
- standard Reduction potential
- ~~Low~~ Bond dissociation energy is low
- Atomic Size is smaller.

Q. Why electronegativity of halogen decrease in order of F > Cl > Br > I.

- As we move down, the atomic size increase.
- As we down, the shielding effect also increases.