CHAPTER 13

s AND p - BLOCK ELEMENTS

After completing this lesson, you will be able to:

- Recognize the demarcation of Periodic Table into s block, p block d block, and f block.
- Describe how physical properties like atomic radius, ionization energy, electro negativity, electrical conductivity and melting and boiling points of element change within a group and within a period and the Periodic Table.
- Describe reactions of period 3 elements with water, oxygen and chlorine.
- Describe reaction of oxides and chlorides of period 3 elements with water.
- Describe reaction of group I elements with water, oxygen and chlorine.
- Discuss the trend in solubility of the hydroxides, sulphates and carbonates of group II elements.
- Discuss the trends in thermal stability of the nitrates and carbonates of group II elements.
- Explain the trends in physical properties and oxidation states in group I, II IV and V of the periodic table.
- Explain effects of heat on nitrates, carbonates and hydrogen carbonated of Group I elements.
- Differentiate beryllium from other members of its group.

Period 3 (Na to Ar)

Atomic And Physical Properties Of The Period 3 Elements

(This period contains Na, Mg, Al, Si, P, S, Cl and Ar)

This topic describes and explains the trends in atomic and physical properties of the Period 3 elements from sodium to argon. It covers ionisation energy, atomic radius, electronegativity, electrical conductivity, melting point and boiling point.

a) Atomic Properties

13.1.1.1 Electronic structures

In Period 3 of the Periodic Table, the 3s and 3p orbitals are filling with electrons. Just as a reminder, the shortened versions of the electronic structures for the eight elements are:

Na [Ne] $3s^1$

Mg [Ne] $3s^2$

Al [Ne] $3s^2\ 3p_x^1$

Si [Ne] $3s^2\ 3p_x^1\ 3p_y^1$

P [Ne] $3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1$

S [Ne] $3s^2\ 3p_x^2\ 3p_y^1\ 3p_z^1$

Cl [Ne] $3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^1$

Ar [Ne] $3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^2$

In each case, [Ne] represents the complete electronic structure of a neon atom.

Trends in Atomic radius

We know that the number of shells in all the elements of a given periods remains the same but the value of effective nuclear charge, increases from left to right. The increased effective nuclear charge pulls the electron cloud of the atom nearer to the nucleus and thus the size of the atoms and ions goes on decreasing from left to right. Thus in going from left to right in a period of s-and p-block elements atomic and ionic radii decrease with the increase of atomic number. This fact can be illustrated by considering the atomic (covalent) and ionic radii of the elements as shown below.

Table 13.1 *Ionic radii (in Å) of representative elements (s- and p-block elements). In parentheses are given the oxidation states of the elements.*

Group / Period	s-block elements		p-block elements				
	I A	II A	III A	IV A	V A	VI A	VII A
1	H 2.08 (-1) 0.29 (+1)						
2	Li 0.60 (+1)	Be 0.31 (+2)	B 0.20 (+3)	C 2.60 (-4) 0.15 (+4)	N 1.71 (-3) 0.11 (+5)	O 1.40 (-2) 0.09 (+6)	F 1.36 (-1) 0.07 (+7)
3	Na 0.95 (+1)	Mg 0.65 (+2)	Al 0.50 (+3)	Si 2.71 (-1) 0.41 (+4)	P 2.12 (-3) 0.34 (+5)	S 1.84 (-2) 0.29 (+6)	Cl 1.81 (-1) 0.26 (+7)
4	K 1.33 (+1)	Ca 0.99 (+2)	Ga 1.13 (+1) 0.62 (+3)	Ge 0.93 (+2) 0.53 (+4)	As 2.22 (-3) 0.47 (+5)	Se 1.98 (-2) 0.42 (+6)	Br 1.95 (-1) 0.39 (+7)
5	Rb 1.48 (+1)	Sr 1.13 (+2)	In 1.32 (+1) 0.81 (+3)	Sn 1.12 (+2) 0.71 (+4)	Sb 0.45 (-3) 0.62 (+5)	Te 2.21 (-2) 0.56 (+6)	I 2.16 (-1) 0.50 (+7)
6	Cs 1.69 (+1)	Ba 1.35 (+2)	Tl 1.40 (+1) 0.95 (+3)	Pb 1.20 (+2) 0.84 (+4)	Bi 1.20 (+3) 0.74 (+5)	Po - -	At - -
7	Fr 1.76 (+1)	Ra 1.40 (+2)					

Q1. Give Trends in Electronegativity in modern periodic table.

Answer

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons.

The Pauling scale is the most commonly used. Fluorine (the most electronegative element) is assigned a value of 4.0, and values range down to caesium and francium which are the least electronegative at 0.7.

The trend

The trend across Period 3 looks like this:

Fig. 13.2 Electronegativities

In going from left to right in a period of s- and p-block elements, the electronegativity values increase. This increase can be explained on the basis of any of the following facts.

i) On moving from left to right in a period, there is a decrease in the size of the atoms. Smaller atoms have greater tendency to attract the electrons towards themselves i.e. smaller atoms have higher electronegativity values.

ii) On moving from left to right in a period there is an increase of ionization energy and electron affinity of the elements. The atoms of the elements which have higher value of ionization energies and electron affinities also have higher electronegativities.

The variation of electronegativity in a period and a group of representative elements (s- and p-block elements) is show in Fig. 13.3

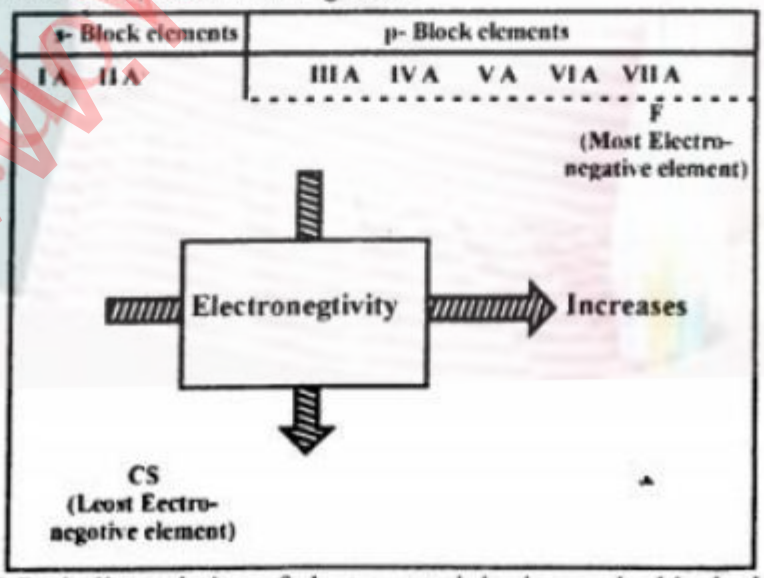


Fig 13.3 Periodic variation of electronegativity in s and p block elements.

Notice that argon isn't included. Electronegativity is about the tendency of an atom to attract a *bonding* pair of electrons. Since argon doesn't form covalent bonds, you obviously can't assign it electronegativity.

Explaining the trend

The trend is explained in exactly the same way as the trend in atomic radii.

As you go across the period, the bonding electrons are always in the same level - the 3-level. They are always being screened by the same inner electrons.

All that differs is the number of protons in the nucleus. As you go from sodium to chlorine, the number of protons steadily increases and so attracts the bonding pair more closely.

B) Physical Properties

This section is going to look at the electrical conductivity and the melting and boiling points of the elements. To understand these, you first have to understand the structure of each of the elements.

Structures of the elements

The structures of the elements change as you go across the period. The first three (i.e. Na, Mg, Al) are metallic, silicon is giant covalent, and the rest (i.e. P, S, Cl, Ar) are simple molecules.

Three metallic structures

Sodium, magnesium and aluminium all have metallic structures.

In sodium, only one electron per atom is involved in the metallic bond - the single 3s electron. In magnesium, both of its outer electrons are involved, and in aluminium all three.

The other difference you need to be aware of is the way the atoms are packed in the metal crystal.

Sodium is 8-co-ordinated - each sodium atom is touched by only 8 other atoms.

Both magnesium and aluminium are 12-co-ordinated (although in slightly different ways). This is a more efficient way to pack atoms, leading to less wasted space in the metal structures and to stronger bonding in the metal.

Q2. Draw and discuss structure giant molecules.

Answer

Silicon has a giant covalent structure just like diamond. A tiny part of the structure looks like this:

Fig 13.4

The structure is held together by strong covalent bonds in all three dimensions.

Four simple molecular structures

The structures of phosphorus (i.e. white etc) and sulphur (i.e. rhomic or nomoclinic etc) vary depending on the type of phosphorus or sulphur you are talking about.

a P_4 molecule an S_8 molecule a Cl_2 molecule an Ar molecule

The atoms in each of these molecules are held together by covalent bonds (apart, of course, from argon).

In the liquid or solid state, the molecules are held close to each other by van der Waals dispersion forces.

Electrical conductivity

- Sodium, magnesium and aluminium are all good conductors of electricity. Conductivity increases as you go from sodium to magnesium to aluminium as they have free electrons
- Silicon is a semiconductor.
- None of the rest conduct electricity.

The three metals, of course, conduct electricity because the delocalised electrons (the "sea of electrons") are free to move throughout the solid or the liquid metal.

In the silicon case, explaining how semiconductors conduct electricity is beyond the scope of this level. With a diamond structure, you mightn't expect it to conduct electricity, but it does!

The rest don't conduct electricity because they are simple molecular substances. There are no electrons free to move around.

Q3. Draw and discuss Trends in Melting and Boiling Points of si, p, Cl and Ar.

Answer

The chart shows how the melting and boiling points of the elements change as you go

across the period. The figures are plotted in Kelvin rather than °C to avoid having negative values.

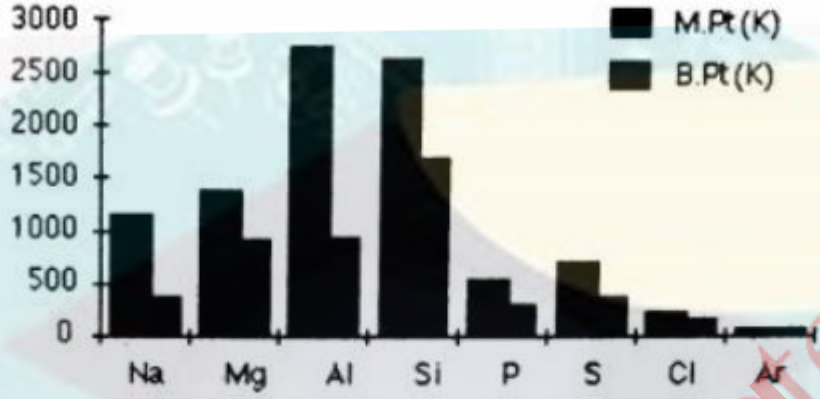


Fig 13.5 Melting and Boiling Point

Silicon

Silicon has high melting and boiling points because it is a giant covalent structure. You have to break strong covalent bonds before it will melt or boil.

Because you are talking about a different type of bond, it isn't profitable to try to directly compare silicon's melting and boiling points with aluminium's.

The four molecular elements

Phosphorus, sulphur, chlorine and argon are simple molecular substances with only van der Waals attractions between the molecules. Their melting or boiling points will be lower than those of the first four members of the period which have giant structures.

The sizes of the melting and boiling points are governed entirely by the sizes of the molecule: . Remember the structures of the molecules:

Phosphorus

Phosphorus contains P_4 molecules. To melt phosphorus you don't have to break any covalent bonds - just the much weaker van der Waals forces between the molecules.

Sulphur

Sulphur consists of S_8 rings of atoms. The molecules are bigger than phosphorus molecules, and so the van der Waals attractions will be stronger, leading to a higher melting and boiling point.

Chlorine

Chlorine, Cl_2, is a much smaller molecule with comparatively weak van der Waals attraction , and so chlorine will have a lower melting and boiling point than sulphur or phosphor. s.

Argon

Argon molecules are just single argon atoms, Ar. The scope for van der Waals attractions between these is very limited and so the melting and boiling points of argon are lower again.

Q4. Give Chemical Reactions of the Period 3 Elements.

Answer

This section describes the reactions of the Period 3 elements from sodium to argon with water, oxygen and chlorine.

I) Reactions with water

Sodium

Sodium has a very exothermic reaction with cold water producing hydrogen and a colourless solution of sodium hydroxide.

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

Magnesium

Magnesium has a very slight reaction with cold water, but burns in steam.

A very clean coil of magnesium dropped into cold water eventually gets covered in small bubbles of hydrogen which float it to the surface. Magnesium hydroxide is formed as a very thin layer on the magnesium and this tends to stop the reaction.

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

Magnesium burns in steam with its typical white flame to produce white magnesium oxide and hydrogen.

$$Mg + H_2O \longrightarrow MgO + H_2$$

Aluminium

Aluminium powder heated in steam produces hydrogen and aluminium oxide. The reaction is relatively slow because of the strong aluminium oxide layer on the metal, and the build-up of even more oxide during the reaction.

$$2Al + 3H_2O \longrightarrow Al_2O_3 + 3H_2$$

Phosphorus and sulphur

These have no reaction with water.

Chlorine

Chlorine dissolves in water to some extent to give a green solution. A reversible reaction takes place to produce a mixture of hydrochloric acid and chloric(I) acid (hypochlorous acid).

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

In the presence of sunlight, the chloric(I) acid slowly decomposes to produce more hydrochloric acid, releasing oxygen gas, and you may come across an equation showing the overall change:

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

Argon

There is no reaction between argon and water.

II) Reactions with oxygen

Sodium

Sodium burns in oxygen with an orange flame to produce a white solid mixture of sodium oxide and sodium peroxide.

For the simple oxide:

$$4Na + O_2 \longrightarrow 2Na_2O$$

For the peroxide:

$$2Na + O_2 \longrightarrow Na_2O_2$$

Magnesium

Magnesium burns in oxygen with an intense white flame to give white solid magnesium oxide.

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

Silicon

Silicon will burn in oxygen if heated strongly. Silicon dioxide is produced.

$$Si + O_2 \longrightarrow SiO_2$$

Phosphorus

White phosphorus catches fire spontaneously in air, burning with a white flame and producing clouds of white smoke - a mixture of phosphorus (III) oxide and phosphorus(V) oxide.

The proportions of these depend on the amount of oxygen available. In an excess of oxygen, the product will be almost entirely phosphorus(V) oxide.

For the phosphorus(III) oxide:

$$P_4 + 3O_2 \longrightarrow P_4O_6$$

For the phosphorus(V) oxide:

$$P_4 + 5O_2 \longrightarrow P_4O_{10}$$

Sulphur

Sulphur burns in air or oxygen on gentle heating with a pale blue flame. It produces colourless sulphur dioxide gas.

$$S + O_2 \longrightarrow SO_2$$

Chlorine and argon

Despite having several oxides, chlorine won't react directly with oxygen. Argon doesn't react either.

Properties of the oxides of elements in period 3

Formula of Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
					(P_4O_6)	(SO_2)	(Cl_2O)
State of Oxide	solid	solid	solid	solid	solid	liquid	liquid
Conduction of Electricity by Molten or liquid Oxide	good	good	good	v.poor	Nil.	Nil.	Nil.
Structure of oxide	Giant Structures				Simple molecular structure		
Enthalpy change of Formation of oxide at 298K/kJ mol-1	-416	-602	-1676	-910	-2984	-395	80
Enthalpy change of Formation of oxide at 298K/kJ mol-1 O/kJ	-416	-602	-559	-455	-298	-132	80
Effect of adding oxide to water	reacts to form NaOH (aq) alkaline solution	reacts to form Mg(OH)$_2$	does not react with water but it is amphoteric	does not react with water but it is acidic	P_4O_{10} reacts to form H_3PO_4 acid solution	SO_3 reacts to form H_2SO_4 acid solution	Cl_2O_7 reacts to form $HClO_4$ acid solution
Nature of Oxide	Basic (alkaline)	Basic (weakly alkaline)	Amphoteric	Acidic	Acidic	Acidic	Acidic

iii) Reactions with chlorine

Sodium

Sodium burns in chlorine with a bright orange flame. White solid sodium chloride is produced.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

Magnesium

Magnesium burns with its usual intense white flame to give white magnesium chloride.

$$Mg + Cl_2 \longrightarrow MgCl_2$$

Aluminium

Aluminium is often reacted with chlorine by passing dry chlorine over aluminium foil heated in a long tube. The aluminium burns in the stream of chlorine to produce very

pale yellow aluminium chloride. This sublimes (turns straight from solid to vapour and back again) and collects further down the tube where it is cooler.

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

Silicon

When chlorine is passed over silicon powder heated in a tube, it reacts to produce silicon tetrachloride. This is a colourless liquid which vaporises and can be condensed further along the apparatus.

$$Si + 2Cl_2 \longrightarrow SiCl_4$$

Phosphorus

White phosphorus burns in chlorine to produce a mixture of two chlorides, phosphorus(III) chloride and phosphorus(V) chloride (phosphorus trichloride and phosphorus pentachloride).

Phosphorus (III) chloride is a colourless fuming liquid.

$$P_4 + 6Cl_2 \longrightarrow 4PCl_3$$

Phosphorus(V) chloride is an off-white (going towards yellow) solid.

$$P_4 + 10Cl_2 \longrightarrow 4PCl_5$$

Sulphur

When a stream of chlorine is passed over some heated sulphur, it reacts to form an orange, evil-smelling liquid, disulphur dichloride, S_2Cl_2.

$$2S + Cl_2 \longrightarrow S_2Cl_2$$

Chlorine and argon

It obviou_ly doesn't make sense to talk about chlorine reacting with itself, and argon doesn't react with chlorine.

Q5. Give Physical Properties of the Oxides of 3rd period elements.

Answer

This section explains the relationship between the physical properties of the oxides of Period 3 elements (sodium to chlorine) and their structures. Argon is obviously omitted because it doesn't form an oxide.

A quick summary of the trends

The oxides

The oxides we'll be looking at are:

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

Those oxides in the top row are known as the ***highest oxides*** of the various elements. These are the oxides where the Period 3 elements are in their highest oxidation states. In these oxides, all the outer electrons in the Period 3 element are being involved in the bonding - from just the one with sodium, to all seven of chlorine's outer electrons.

i) structures

The trend in structure is from the metallic oxides containing giant structures of ions on the left of the period via a giant covalent oxide (silicon dioxide) in the middle to molecular oxides on the right.

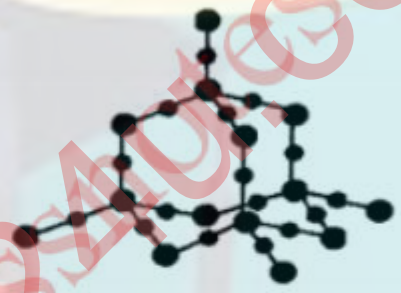
ii) The metallic oxides (e.g Sodium, Magnesium, aluminium etc.)

Sodium, magnesium and aluminium oxides structure/diagram is same as sodium chloride which you have read in previous classes.

ii) Giant Covalent Oxides (e.g Silicon dioxide (silicon(IV) oxide))

i. structure

Crystalline silicon has the same structure as diamond. To turn it into silicon dioxide, all you need to do is to modify the silicon structure by including some oxygen atoms.



Notice that each silicon atom is bridged to its neighbours by an oxygen atom.

iii) The molecular oxides (e.g. Phosphorus, sulphur and chlorine oxides)

A) The phosphorus oxides

Phosphorus has two common oxides, phosphorus(III) oxide, P_4O_6, and phosphorus(V) oxide, P_4O_{10}.

i) Phosphorus(III) oxide

Phosphorus(III) oxide is a white solid, melting at 24°C and boiling at 173°C.

The structure of its molecule is best worked out starting from a P_4 molecule which is a little tetrahedron.

a P_4 molecule

Fig. 13.6

Pull this apart so that you can see the bonds . . .

a P_4 molecule

Fig. 13.7

and then replace the bonds by new bonds linking the phosphorus atoms via oxygen

atoms. These will be in a V-shape (rather like in water).

a P_4O_6 molecule

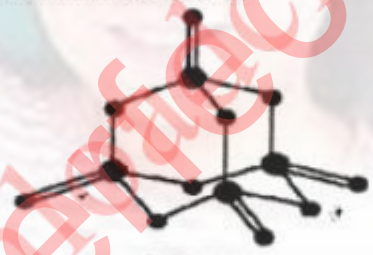
The phosphorus is using only three of its outer electrons (the 3 unpaired p electrons) to form bonds with the oxygens.

ii) Phosphorus(V) oxide

Phosphorus(V) oxide is also a white solid, subliming (turning straight from solid to vapour) at 300°C. In this case, the phosphorus uses all five of its outer electrons in the bonding.

Solid phosphorus(V) oxide exists in several different forms - some of them polymeric. We are going to concentrate on a simple molecular form, and this is also present in the vapour.

This is most easily drawn starting from P_4O_6. The other four oxygens are attached to the four phosphorus atoms via double bonds.



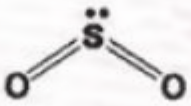
a P_4O_{10} molecule

B) The sulphur oxides

Sulphur has two common oxides, sulphur dioxide (sulphur(IV) oxide), SO_2, and sulphur trioxide (sulphur(VI) oxide), SO_3.

i) Sulphur dioxide

Sulphur dioxide is a colourless gas at room temperature with an easily recognised choking smell. It consists of simple SO_2 molecules.



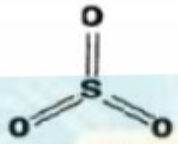
The sulphur uses 4 of its outer electrons to form the double bonds with the oxygen, leaving the other two as a lone pair on the sulphur. The bent shape of SO_2 is due to this lone pair.

ii) Sulphur trioxide

Pure sulphur trioxide is a white solid with a low melting and boiling point.

Gaseous sulphur trioxide consists of simple SO_3 molecules in which all six of the

sulphur's outer electrons are involved in the bonding.



There are various forms of solid sulphur trioxide. The simplest one is a trimer, S_3O_9, where three SO_3 molecules are joined up and arranged in a ring.

Fig. 13.8

There are also other polymeric forms in which the SO_3 molecules join together in long chains. For example:

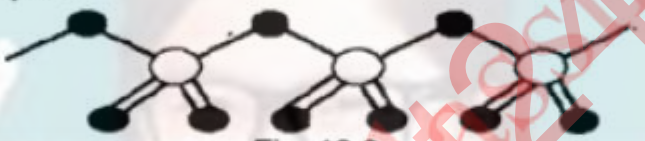


Fig. 13.9

The fact that the simple molecules join up in this way to make bigger structures is what makes the sulphur trioxide a solid rather than a gas.

C) The chlorine oxides

Chlorine forms several oxides. Here we discuss only two which are - chlorine(I) oxide, Cl_2O, and chlorine(VII) oxide, Cl_2O_7.

Chlorine(I) oxide

Chlorine(I) oxide is a yellowish-red gas at room temperature. It consists of simple small molecules.

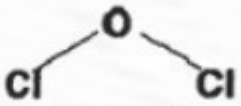


Fig. 13.10

In this structure chlorine uses its one outer electron and bonds with oxygen.

Chlorine(VII) oxide

Chlorine(VII) oxide is a colourless oily liquid at room temperature.

In chlorine(VII) oxide, the chlorine uses all of its seven outer electrons and bonds with oxygen. This produces a much bigger molecule.

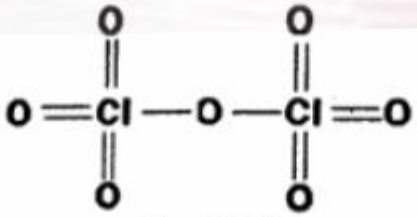


Fig. 13.11

ii) Melting and boiling points

The giant structures (the metal oxides and silicon dioxide) will have high melting and boiling points because a lot of energy is needed to break the strong bonds (ionic or covalent) operating in three dimensions.

The oxides of phosphorus, sulphur and chlorine consist of individual molecules - some small and simple; others polymeric.

The attractive forces between these molecules will be van der Waals dispersion and dipole-dipole interactions. These vary in size depending on the size, shape and polarity of the various molecules - but will always be much weaker than the ionic or covalent bonds you need to break in a giant structure.

These oxides tend to be gases, liquids or low melting point solids.

Q6. Give Acid-Base Behaviour of the Oxides of 3rd period.

Answer

This topic looks at the reactions of the oxides of Period 3 elements (sodium to chlorine) with water and with acids or bases where relevant.(Take quick review from table 13.2)

iii) Electrical conductivity

None of these oxides has any free or mobile electrons. That means that none of them will conduct electricity when they are solid.

The ionic oxides can, however, undergo ***electrolysis*** when they are molten. They can conduct electricity because of the movement of the ions towards the electrodes and the discharge of the ions when they get there.

i. Trend in acid-base behaviour

The trend in acid-base behaviour is shown in various reactions, but as a simple summary:

The trend is from strongly basic oxides on the left-hand side to strongly acidic ones on the right, via an amphoteric oxide (aluminium oxide) in the middle. An amphoteric oxide is one which shows both acidic and basic properties.

ii) Reactions of Oxides with water, Acids and Bases

Chemistry of the individual oxides

Sodium oxide (Na_2O)

Sodium oxide is a simple strongly basic oxide. It is basic because it contains the oxide ion, O^{2-}, which is a very strong base with a high tendency to combine with hydrogen ions.

Reaction with water

Sodium oxide reacts exothermically with cold water to produce sodium hydroxide solution. Depending on its concentration, this will have a pH around 14.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

Reaction with acids

As a strong base, sodium oxide also reacts with acids. For example, it would react with dilute hydrochloric acid to produce sodium chloride solution.

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

Magnesium oxide (MgO)

Magnesium oxide is again a simple basic oxide, because it also contains oxide ions. However, it isn't as strongly basic as sodium oxide because the oxide ions aren't so free.

In the sodium oxide case, the solid is held together by attractions between 1+ and 2- ions. In the magnesium oxide case, the attractions are between 2+ and 2-. It takes more energy to break these.

Reaction with water

If you shake some white magnesium oxide powder with water, nothing seems to happen (it doesn't look as if it reacts). However, if you test the pH of the liquid, you find that it is somewhere around pH 9 (showing that it is slightly alkaline).

There must have been some slight reaction with the water to produce hydroxide ions in solution. Some magnesium hydroxide is formed in the reaction, but this is almost insoluble - and so not many hydroxide ions actually get into solution.

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

Reaction with acids

Magnesium oxide reacts with acids as you would expect any simple metal oxide to react. For example, it reacts with warm dilute hydrochloric acid to give magnesium chloride solution.

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

Aluminium oxide (Al_2O_3)

As it is amphoteric oxide, It has reactions as both a base and an acid.

Reaction with water

Aluminium oxide doesn't react in a simple way with water and doesn't dissolve in it. Although it still contains oxide ions, they are held too strongly in the solid lattice to react with the water.

Reaction with acids

Aluminium oxide will react with hot dilute hydrochloric acid to give aluminium chloride solution.

$$Al_2O_3 + 6HCl \longrightarrow 2AlCl_3 + 3H_2O$$

Reaction with bases

Aluminium oxide has also got an acidic side to its nature, and it shows this by reacting with bases such as sodium hydroxide solution.

Various aluminates are formed, compounds where the aluminium is found in the negative ion. This is possible because aluminium has the ability to form covalent bonds with oxygen.

With hot, concentrated sodium hydroxide solution, aluminium oxide reacts to give a colourless solution of sodium tetrahydroxoaluminate.

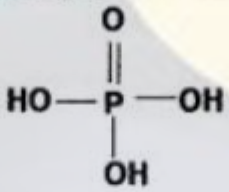
$$Al_2O_3 + 2NaOH + 3H_2O \longrightarrow 2NaAl(OH)_4$$

Phosphorus(V) oxide (P_4O_{10})

Reaction with Water

Phosphorus(V) oxide reacts violently with water to give a solution containing a mixture of acids, the nature of which depends on the conditions. We usually just consider one of these, phosphoric(V) acid, H_3PO_4 (also known just as phosphoric acid or as orthophosphoric acid).

$$P_4O_{10} + 6H_2O \longrightarrow 4H_3PO_4$$



Reaction with Base:

As it is acidic so it reacts with NaOH as follows: .

$$NaOH + H_3PO_4 \longrightarrow NaH_2PO_4 + H_2O$$
$$2NaOH + H_3PO_4 \longrightarrow Na_2HPO_4 + 2H_2O$$
$$3NaOH + H_3PO_4 \longrightarrow Na_3PO_4 + 3H_2O$$

Again, if you were to react phosphorus(V) oxide directly with sodium hydroxide solution rather than making the acid first, you would end up with the same possible salts.

This is getting ridiculous, and so I will only give one example out of the possible equations:

$$12NaOH + P_4O_{10} \longrightarrow 4Na_3PO_4 + 6H_2O$$

The sulphur oxides (SO_x)

We are going to be looking at sulphur dioxide, SO_2, and sulphur trioxide, SO_3.

Sulphur dioxide

Reaction with Water:

Sulphur dioxide is fairly soluble in water, reacting with it to give a solution of sulphurous acid, H_2SO_3.

$$SO_2 + H_2O \longrightarrow H_2SO_3$$

Reaction with Base:

As it is acidic so it reacts with NaOH and CaO as follows:

$$SO_2 + 2NaOH \longrightarrow Na_2SO_3 + H_2O$$
$$Na_2SO_3 + H_2O + SO_2 \longrightarrow 2NaHSO_3$$
$$CaO + SO_2 \longrightarrow CaSO_3$$

Sulphur trioxide

Sulphur trioxide reacts violently with water to produce a fog of concentrated sulphuric acid droplets.

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

Reaction with Base:

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

In principle, you can also get sodium hydrogensulphate solution by using half as much sodium hydroxide and just reacting with one of the two acidic hydrogens in the acid. In practice, I personally have never ever done it.

Sulphur trioxide itself will also react directly with bases to form sulphates. For example, it will react with calcium oxide to form calcium sulphate. This is just like the reaction with sulphur dioxide described above.

$$CaO + SO_3 \longrightarrow CaSO_4$$

The chlorine oxides (Cl_2O_n)

Chlorine forms several oxides, but the only two are chlorine(VII) oxide, Cl_2O_7, and chlorine(I)oxide, Cl_2O. Chlorine(VII) oxide is also known as dichlorine heptoxide, and chlorine(I) oxide as dichlorine monoxide.

Chlorine(VII) oxide

Chlorine(VII) oxide is the highest oxide of chlorine - the chlorine is in its maximum oxidation state of +7. It continues the trend of the highest oxides of the Period 3 elements towards being stronger acids.

Reaction with Water:

Chlorine(VII) oxide reacts with water to give the very strong acid, chloric(VII) acid - also known as perchloric acid. The pH of typical solutions will, like sulphuric acid, be around 0.

$$Cl_2O_7 + H_2O \longrightarrow 2HClO_4$$

Reaction with Base:

Chloric(VII) acid reacts with sodium hydroxide solution to form a solution of sodium chlorate(VII).

$$NaOH + HClO_4 \longrightarrow NaClO_4 + H_2O$$

Chlorine(VII) oxide itself also reacts with sodium hydroxide solution to give the same product.

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

Chlorine(I) oxide

Reaction with Base:

Chlorine(I) oxide is far less acidic than chlorine(VII) oxide. It reacts with water to some extent to give chloric(I) acid, HOCl - also known as hypochlorous acid.

$$Cl_2O + H_2O \rightleftharpoons 2HOCl$$

Reaction with Base:

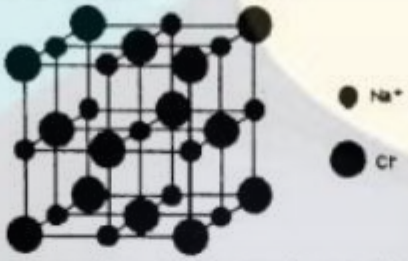
Chloric(I) acid reacts with sodium hydroxide solution to give a solution of sodium chlorate(I) (sodium hypochlorite).

$$NaOH + HOCl \longrightarrow NaOCl + H_2O$$

Chlorine(I) oxide also reacts directly with sodium hydroxide to give the same product.

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

This is normally drawn in an exploded form as:



The strong attractions between the positive and negative ions need a lot of heat energy to break, and so sodium chloride has high melting and boiling points.

It doesn't conduct electricity in the solid state because it hasn't any mobile electrons and the ions aren't free to move. However, when it melts it undergoes electrolysis.

Sodium chloride simply dissolves in water to give a neutral solution.

Magnesium chloride, $MgCl_2$

Magnesium chloride is also ionic, but with a more complicated arrangement of the ions to allow for having twice as many chloride ions as magnesium ions.

Again, lots of heat energy is needed to overcome the attractions between the ions, and so the melting and boiling points are again high.

Solid magnesium chloride is a non-conductor of electricity because the ions aren't free to move. However, it undergoes electrolysis when the ions become free on melting.

Magnesium chloride dissolves in water to give a faintly acidic solution (pH = approximately 6).

When magnesium ions are broken off the solid lattice and go into solution, there is enough attraction between the 2+ ions and the water molecules to get co-ordinate (dative covalent) bonds formed between the magnesium ions and lone pairs on surrounding water molecules.

Hexaaquamagnesium ions are formed, $[Mg(H_2O)_6]^{2+}$.

$$MgCl_{2(s)} + 6H_2O_{(l)} \longrightarrow [Mg(H_2O)_6]^{2+}{}_{(aq)} + 2Cl^-_{(aq)}$$

Ions of this sort are acidic.

Aluminium chloride, $AlCl_3$

$$Al_2Cl_6 \rightleftharpoons 2AlCl_3$$

Solid aluminium chloride doesn't conduct electricity at room temperature because the ions aren't free to move. Molten aluminium chloride (only possible at increased pressures) doesn't conduct electricity because there aren't any ions any more.

The reaction of aluminium chloride with water is surprising. If you drop water onto solid aluminium chloride, you get a violent reaction producing clouds of steamy fumes of hydrogen chloride gas.

The aluminium chloride reacts with the water rather than just dissolving in it. In the

first instance, hexaaquaaluminium ions are formed together with chloride ions.

$$AlCl_{3(s)} + 6H_2O_{(l)} \longrightarrow [Al(H_2O)_6]^{3+}_{(aq)} + 3Cl^-_{(aq)}$$

You will see that this is very similar to the magnesium chloride equation given above - the only real difference is the charge on the ion.

Silicon tetrachloride, $SiCl_4$

Silicon tetrachloride is a simple covalent chloride. There isn't enough electronegativity difference between the silicon and the chlorine for the two to form ionic bonds.

> **Do you know?**
> Silicon tetrachloride is a colourless liquid at room temperature which fumes in moist air. The only attractions between the molecules are van der Waals dispersion forces.
> It doesn't conduct electricity because of the lack of ions or mobile electrons.

It fumes in moist air because it reacts with water in the air to produce hydrogen chloride. If you add water to silicon tetrachloride, there is a violent reaction to produce silicon dioxide and fumes of hydrogen chloride. In a large excess of water, the hydrogen chloride will, of course, dissolve to give a strongly acidic solution containing hydrochloric acid.

$$SiCl_4 + 2H_2O \longrightarrow SiO_2 + 4HCl$$

The phosphorus chlorides

There are two phosphorus chlorides - phosphorus(III) chloride, PCl_3, and phosphorus(V) chloride, PCl_5.

Phosphorus(III) chloride (phosphorus trichloride), PCl_3

This is another simple covalent chloride - again a fuming liquid at room temperature.

> **Do you know:**
> Phosphorus trichloride is a liquid because there are only van der Waals dispersion forces and dipole-dipole attractions between the molecules. It doesn't conduct electricity because of the lack of ions or mobile electrons.

Phosphorus(III) chloride reacts violently with water. You get phosphorous acid, H_3PO_3, and fumes of hydrogen chloride (or a solution containing hydrochloric acid if lots of water is used).

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

Phosphorus(V) chloride (phosphorus pentachloride), PCl_5

Unfortunately, phosphorus(V) chloride is structurally more complicated.

> Phosphorus(V) chloride is a white solid which sublimes at 163°C. The higher the temperature goes above that, the more the phosphorus(V) chloride dissociates (splits up reversibly) to give phosphorus(III) chloride and chlorine.
>
> $$PCl_5 \rightleftharpoons PCl_3 + Cl_2$$
>
> Solid phosphorus(V) chloride contains ions - which is why it is a solid at room temperature. The formation of the ions involves two molecules of PCl_5.

Phosphorus(V) chloride has a violent reaction with water producing fumes of hydrogen chloride. As with the other covalent chlorides, if there is enough water present, these

will dissolve to give a solution containing hydrochloric acid.

The reaction happens in two stages. In the first, with cold water, phosphorus oxychloride, $POCl_3$, is produced along with HCl.

$$PCl_5 + H_2O \longrightarrow POCl_3 + 2HCl$$

If the water is boiling, the phosphorus(V) chloride reacts further to give phosphoric(V) acid and more HCl. Phosphoric(V) acid is also known just as phosphoric acid or as orthophosphoric acid.

$$POCl_3 + 3H_2O \longrightarrow H_3PO_4 + 3HCl$$

The overall equation in boiling water is just a combination of these:

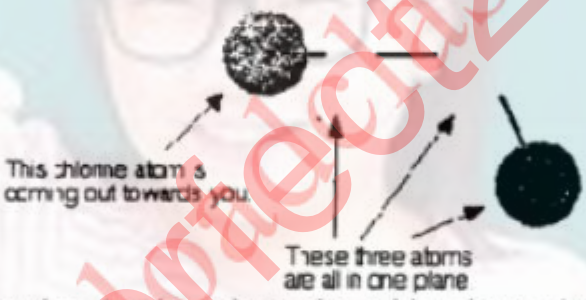
$$PCl_5 + 4H_2O \longrightarrow H_3PO_4 + 5HCl$$

Disulphur dichloride, S_2Cl_2

Disulphur dichloride is formed when chlorine reacts with hot sulphur.

Disulphur dichloride is a simple covalent liquid (orange and smelly).

The shape is difficult to draw convincingly. The atoms are all joined up in a line but twisted:



The reason for drawing the shape is to give a hint about what sort of intermolecular attractions are possible. There is no plane of symmetry in the molecule and that means that it will have an overall permanent dipole.

Do you know?
Disulphur Dichloride has van der Waals dispersion forces and dipole-dipole attractions.
There are no ions in disulphur dichloride and no mobile electrons - so it never conducts electricity.

Disulphur dichloride reacts slowly with water to produce a complex mixture of things including hydrochloric acid, sulphur, hydrogen sulphide and various sulphurcontaining acids and anions (negative ions).

Q8. Discuss Hydroxides Of The Period 3 Elements.

Answer

This topic looks briefly at how the chemistry of the "hydroxides" of the Period 3 elements from sodium to chlorine varies as you cross the period.

A quick summary of the trends
Sodium and magnesium hydroxides
These contain hydroxide ions, and are simple basic hydroxides.
Aluminium hydroxide

Aluminium hydroxide, like aluminium oxide, is amphoteric - it has both basic and acidic properties.

much the negative charge can be spread around the rest of the ion.

If the negative charge stays entirely on the oxygen atom left behind from the -OH group, it will be very attractive to hydrogen ions. The lost hydrogen ion will be easily recaptured and the acid will be weak.

On the other hand, if the charge can be spread out (delocalised) over the whole of the ion, it will be so "dilute" that it won't attract the hydrogen back very easily. The acid will then be strong.

Q9. Discuss Group 1 Elements.

Answer

1) Atomic and Physical Properties of the Group 1 Elements

This section explores the trends in some atomic and physical properties of the Group 1 elements - lithium, sodium, potassium, rubidium and caesium. You will find separate sections below covering the trends in atomic radius, first ionisation energy, electronegativity, melting and boiling points, and density.

2) Trends in Atomic Radius

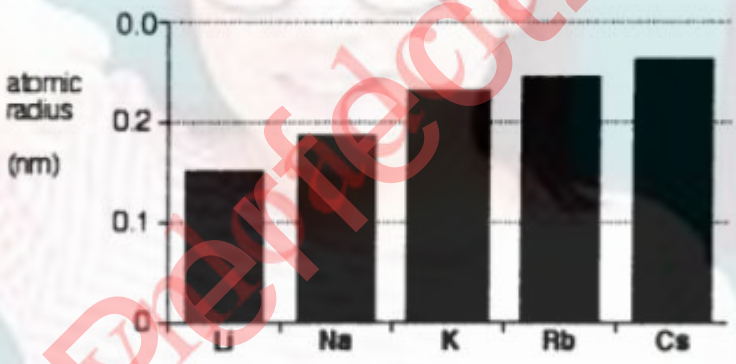


Fig. 13.12 Atomic Radii of the Group 1 elements

As we move from lithium to caesium, an extra shell of electrons is added to each element. The addition of an extra shell increases the atomic volume. We find therefore, that there in an in increase of atomic and ionic radii (of M^+ ions) as we move from lithium to caesium.

Some Physical Properties of alkali metals					
Property	Li	Na	K	Rb	Cs
Atomic weight	6.94	22.99	39.1	25.47	132.91
Atomic volume	12.97	23.63	45.36	55.8	69.95
Atomic (i.e. metallic radius for coordination number 12	1.55	1.9	2.35	2.46	2.6
covalent radius	1.23	1.54	2.03	2.16	2.36
Ionic radius of M+ ions	0.6	0.95	1.33	1.48	1.69
Melting point	180.5	97.8	63.7	38.9	28.7
Boiling point	1330	892	760	688	670
Ionisation energies (kJ/mol) (I_1)	520.3	495.8	418.9	403.0	375.7
I_2	7298.1	4562.4	3051.4	2633.0	2230.0
Standard oxidation potential	3.04	2.71	2.99	2.99	2.99
Sublimation energy (eV/ion)	1.7472	1.2432	1.032	0.984	0.9024
Hydration energy (eV/ion)	5.904	3.792	3.6955	3.36	0.624
Electronegativity	1	0.9	0.8	0.8	0.7
Colour of the flame	Crimson red	Golden yellow	Violet	Violet	Violet
Heat of atomisation at 25C (eV/atm)	1.7472	1.2432	1.032	0.984	0.9024
Ionic conduction of M+ ion	33.5	43.5	64.6	67.5	63

2) Trends in First Ionization Energy

First ionisation energy is the energy needed to remove the most loosely held electron from each of one mole of gaseous atoms to make one mole of singly charged gaseous ions - in other words, for 1 mole of this process:

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

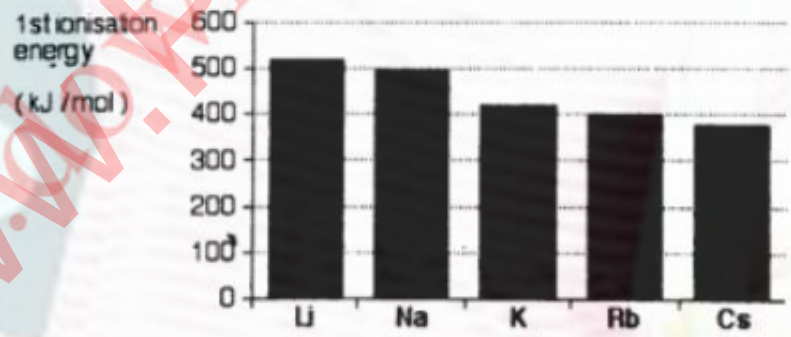


Fig. 13.13 First Ionization Energy of the Group 1 elements

Notice that first ionisation energy falls as you go down the group.

We know that alkali metal have only one electron in their outermost shell (ns1 electron). This ns1 electron is so weakly held with the nucleus that it can be removed very easily. Alkali metals therefore have low ionisation energies.

As the distance of ns1 electron from the nucleus increases on moving from Li to Cs its removal becomes more and more easy as we proceed from Li to Cs i.e. the amount of

energy (ionisation energy) used in the removal of ns1 electron is maximum in case of energies of alkali metals go on decreasing from Li to Cs as shown in table 13.3.

The second ionisation energies are fairly high, since the loss of the second electron from M+ cation which has a noble gas configuration is quite difficult.

3) Trends in Electronegativity

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. It is usually measured on the Pauling scale, on which the most electronegative element (fluorine) is given an electronegativity of 4.0.

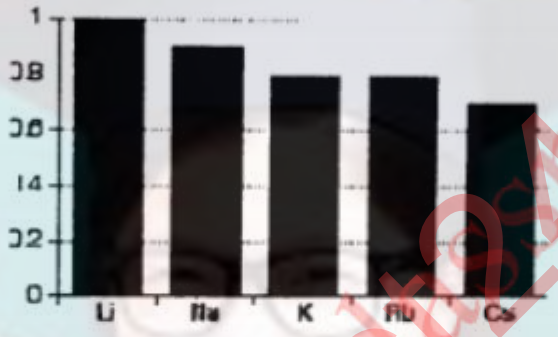


Fig 13.14 Electronegativties of the Group 1 elements

We have seen that the outer electron (i.e. ns1 electron) of the atom of alkali metals is loosely held with the nucleus and hence it can be easily excited to the higher energy levels even by a small amount of heat energy (e.g. by heating the metals or their salts into Bunsen burner). During the excitation process the electron absorbs some energy and when this excited electron comes back to its original position, it gives out absorbed energy in the form of light in visible region of the electromagnetic. Since the amount of energy absorbed during the excitation process is different in different atoms, different colours are imparted by the atoms to the flame. The property of alkali metals to give colouration in the burner flame has been used to detect their presence in salts by a test, known as flame test.

4) Trends in Melting and Boiling Points

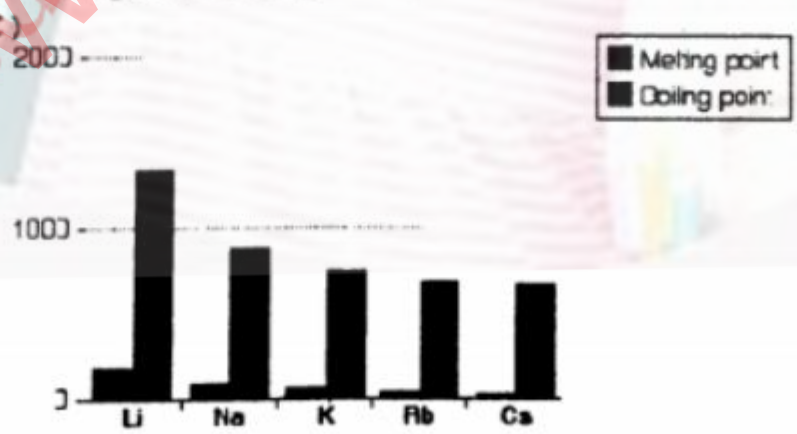


Fig 13.15 Melting and Boiling Points of the Group 1 elements

You will see that both the melting points and boiling points fall as you go down the Group.

The melting and boiling points are very low because of the presence of weak inner atomic bonds in the solid state of the alkali metals. These bonds are due to their atomic radii and mainly due to the their electronic configuration having a single valence electrons as compared to large number of available vacant orbital. As the size of the metal atoms increases, the repulsion of the non-bonding electrons also increases. This increase in the repulsion of non-bonding electron decreases the melting and boiling points of alkali metals when we move from Li to Cs (as shown in table 13.3).

Trends in Density

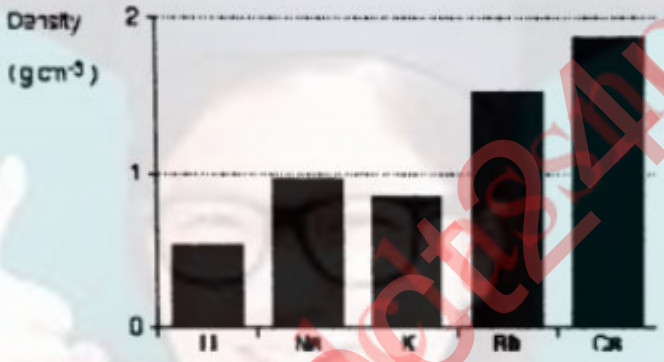


Fig. 13.16 Densities of the Group 1 elements

The densities of alkali metals are quite low due to the large atomic volumes. Li, Na and K are lighter than water. The densities increase with the increase in atomic from Li to Cs indicating that greater atomic with more than compensates for the bigger size of the atoms. K is however, lighter than Na which is due to an unusual increase in atomic size of K.

Elements	Li	Na	K	Rb	Cs
Densities at 0^0C (g/c.c)	0.534	0.972	0.859	1.525	1.903

6) Trends in Reactivity with Water

In this topic, we discuss the reactions of the Group 1 elements - lithium, sodium, potassium, rubidium and caesium with water. It uses these reactions to explore the trend in reactivity in Group 1.

With the exception of Li, the alkali metals are extremely soft and readily fused. They are highly malleable (i.e. can be pressed out into sheets) and ductile (i.e. can be drawn into wires). When freshly cut, they have a bright lustre which quickly tarnished as soon as metal comes in contact with atmosphere.

Summary of the trend in reactivity
The Group 1 metals become more reactive towards water as you go down the Group.

Explaining the trend in reactivity
Looking at the enthalpy changes for the reactions

The overall enthalpy changes

As you go down the Group, the amount of heat given off increases as you go from lithium to caesium. Not so!

The table gives estimates of the enthalpy change for each of the elements undergoing the reaction:

$$X_{(s)} + H_2O_{(l)} \longrightarrow XOH_{(aq)} + {}^1/_2 H_{2(g)}$$

	enthalpy change (kJ / mol)
Li	-222
Na	-184
K	-196
Rb	-195
Cs	-203

You will see that ***there is no pattern at all in these values***. They are all fairly similar and, surprisingly, lithium is the metal which releases the most heat during the reaction!

When these reactions happen, the differences between them lie entirely in what is happening to the metal atoms present. In each case, you start with metal atoms in a solid and end up with metal ions in solution.

Overall, what happens to the metal is this:

$$X_{(s)} \longrightarrow X^+_{(aq)} + e^-$$

Reactions with Oxygen

This topic mainly looks at the reactions of the Group 1 elements (lithium, sodium, potassium, rubidium and caesium) with oxygen - including the simple reactions of the various kinds of oxides formed.

A) The Reactions with Air or Oxygen

Alkali metals react with O_2 or air rapidly and thus get tarnished due to the formation of their oxide on the surface of the metals. It is for this reason that alkali metals are stored in kerosene or paraffin oil.

Li when burnt in O_2 gives mainly lithium monoxide, (normal oxide) Li_2O.

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

Na when burnt in O_2 forms sodium peroxide, Na_2O_2

$$2Na + O_2 \rightarrow Na_2O_2$$

Other alkali metals react with O_2 to form super oxide of MO_2 type.

$$M + O_2 \rightarrow MO_2$$

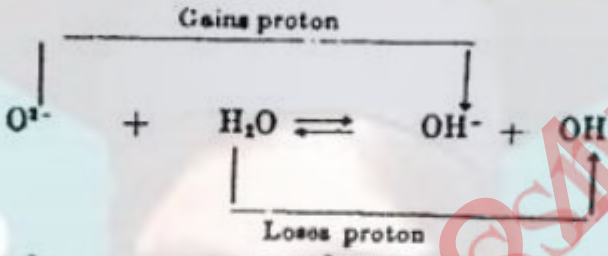
Since the normal oxides of alkali metals other than that of Li (Li_2O) are not formed by the direct reaction between the metals and O_2 they are formed by indirect methods, e.g. by reducing peroxides, nitrite and nitrates with the metals itself.

$$2Na_2O_2 + 4Na \rightarrow 4Na_2O$$

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

$$2NaNO_2 + 10Na \rightarrow 6Na_2O + N_2$$

Properties. Normal oxides (O^{2-}) react with H_2O to form hydroxides by proton exchange.



The peroxides (O_2^{2-}) and superoxides (O^{2-}) are strong oxidising agents and react with H_2O to give H_2O_2 and O_2.

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

$$2[O\text{-}O] + 2H_2O \rightarrow H_2O_2 + O_2$$

Normal oxides have anti-fluorite structure and are ionic in nature since they contain monoxide in, O^{2-} Peroxides contain peroxide ion, O_2^{2-} or $[\text{-O-O}]^{2-}$.

The super oxide ion has a three electron bond as shown below.

The presence of one unpaired electron in it makes this in paramagnetic and coloured.

B) Reactions of the Oxides with Water and Dilute Acids

i) The simple oxides, X_2O

Reaction with water

These are simple basic oxides, reacting with water to give the metal hydroxide.

For example, lithium oxide reacts with water to give a colourless solution of lithium hydroxide.

$$X_2O + H_2O \longrightarrow 2XOH$$

Reaction with dilute acids

These simple oxides all react with an acid to give a salt and water. For example, sodium oxide will react with dilute hydrochloric acid to give colourless sodium chloride solution and water.

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

ii) The peroxides, X_2O_2

Reaction with water

If the reaction is done ice cold (and the temperature controlled so that it doesn't rise even though these reactions are strongly exothermic), a solution of the metal hydroxide and hydrogen peroxide is formed.

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

If the temperature increases (as it inevitably will unless the peroxide is added to water very slowly), the hydrogen peroxide produced decomposes into water and oxygen. The reaction can be very violent overall.

Reaction with dilute acids

These reactions are even more exothermic than the ones with water. A solution containing a salt and hydrogen peroxide is formed. The hydrogen peroxide will decompose to give water and oxygen if the temperature rises again, it is almost impossible to avoid this. Another potentially violent reaction!

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

iii) The superoxides, XO_2

Reaction with water

This time, a solution of the metal hydroxide and hydrogen peroxide is formed, but oxygen gas is given off as well. Once again, these are strongly exothermic reactions and the heat produced will inevitably decompose the hydrogen peroxide to water and more oxygen. Again violent!

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

Reaction with dilute acids

Again, these reactions are even more exothermic than the ones with water. A solution containing a salt and hydrogen peroxide is formed together with oxygen gas. The hydrogen peroxide will again decompose to give water and oxygen as the temperature rises. Violent!

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

C) Reactions of the elements with Chlorine

Sodium burns with an intense orange flame in chlorine is exactly the same way that it does in pure oxygen. The rest also behave the same in both gases.

In each case, there is a white solid residue which is the simple chloride, XCl. There is nothing in any way complicated about these reactions!

$$2X + Cl_2 \longrightarrow 2XCl$$

D) Effect of heat on Nitrates, Carbonates and Hydrogen-Carbonates

i) The facts

Group 1 compounds are more stable to heat than the corresponding compounds in Group 2. You will often find that the lithium compounds behave similarly to Group 2 compounds, but the rest of Group 1 are in some way different.

Nature of carbonates, bicarbonates and nitrates

The carbonates (M_2CO_3) and bicarbonates ($MHCO_3$) are highly stable to heat. With increase of electropositive character from Li to Cs, the stability of these salts increases.

Their nitrates decompose on strong heating to the corresponding nitrite and O_2, (Exception is $LiNO_3$).

$$2NaNO_3 \rightarrow 2NaNO_2 + O_2$$

ii) Explaining the trend in terms of the polarising ability of the positive ion

When alkali metal cations approach near an anion attracts the outer most electrons of the anion and repels the nucleus. Thus the distortion or polarisation of the anion takes place. This distortion results in the sharing of electrons between two oppositely charged ions, i.e. the bond between the cation and anion becomes partly covalent in character. In general the smaller cations polarise the anions more effectively than bigger one. Therefore, the lithium salts are slightly covalent while other alkali metal salts are ionic.

E) Flame Tests

This topic describes how to do a flame test for a range of metal ions, and briefly describes how the flame colour arises.

Flame tests are used to identify the presence of a relatively small number of metal ions in a compound. Not all metal ions give flame colours.

For Group 1 compounds, flame tests are usually by far the easiest way of identifying which metal you have got. For other metals, there are usually other easy methods which are more reliable - but the flame test can give a useful hint as to where to look.

Q10. Give Flame Tests of Group I-A elements.

Answer

Practical details

Clean a platinum or nichrome (a nickel-chromium alloy) wire by dipping it into concentrated hydrochloric acid and then holding it in a hot Bunsen flame. Repeat this until the wire doesn't produce any colour in the flame.

When the wire is clean, moisten it again with some of the acid and then dip it into a small amount of the solid you are testing so that some sticks to the wire. Place the wire back in the flame again.

If the flame colour is weak, it is often worthwhile to dip the wire back in the acid again and put it back into the flame as if you were cleaning it. You often get a very short but intense flash of colour by doing that.

The colours

Different colours shown by different elements are given below.

	flame colour
Li	red
Na	golden yellow

K	lilac (pink)
Rb	red (reddish-violet)
Cs	Blue
Ca	orange-red
Sr	red
Ba	pale green
Cu	blue-green (often with white flashes)
Pb	greyish-white

Note:
What do you do if you have a red flame colour for an unknown compound and don't know which of the various reds it is?
Get samples of known lithium, strontium (etc) compounds and repeat the flame test, comparing the colours produced by one of the known compounds and the unknown compound side by side until you have a good match.

The origin of flame colours

We have seen that the outer electron (i.e. ns^1 electron) of atom of alkali metals is loosely held with the nucleus and hence it can be easily excited to the higher energy levels even by a small amount of heat energy (e.g. by heating the metals or their slats into Bunsen burner). During the excitation process the electron absorbs some energy and when this excited electron comes back to its original position, it gives out absorbed energy in the form of light in visible region of the electromagnetic spectrum and hence the colour is imparted by the atoms to the flame. Sine the amount of energy absorbed during the excitation process is different in different atoms, different colour are imparted by the atoms to the flame. The property of alkali metals to give coloration in the Bunsen flame has been used to detect their presence in salts by a test known as flame test.

Group 2 Elements

Q11. Give Trends of Properties of II-A elements Atomic And Physical Properties.

Answer

This section explores the trends in some atomic and physical properties of the Group 2 elements - beryllium, magnesium, calcium, strontium and barium. You will find

separate sections below covering the trends in atomic radius, first ionization energy, electronegativity and physical properties.

1) Trends in Atomic Radius

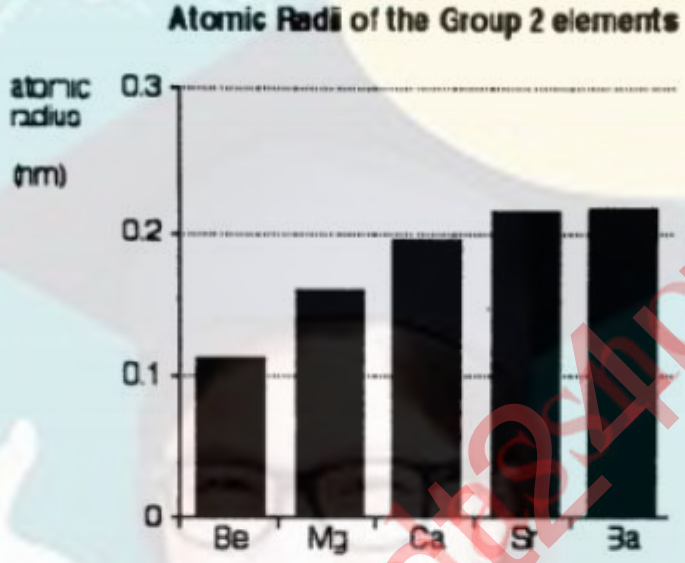


Fig. 13.17 Atomic Radii of the Group 2 elements

You can see that the atomic radius increases as you go down the Group. Notice that beryllium has a particularly small atom compared with the rest of the Group.

Explaining the increase in atomic radius

Atomic volume, atomic and ionic radii

Because of the addition of an extra shell of electrons to each elements from Be to Ra, the atomic volume increases from Be to Ra. With the increases of atomic volume the atomic and ionic radii (of M^{2+} ions) also increase from Be to Ra. The atomic radii of these elements are however, smaller than those of alkali metals in the same period. This is due to the fact that the alkaline earth metals in the same period. This is due to the fact that the alkaline earth metals have higher nuclear charge which tends to draw the orbit electrons towards the nucleus. The smaller values of atomic radii result in that the alkaline earth metals are harder, have higher densities and higher melting points than alkali metals.

Some physical properties of alkaline earth metals

Property	*Be*	*Mg*	*Ca*	*Sr*	*Ba*	*Ra*
Atomic weight	9.01	24.31	40.08	87.62	137.34	226
Abundance (% of earth's crust)	6.4 x 10.4	2.0	3.45	0.915	0.040	1.3x10.10
Density (gm/c.c.)	1.84	1.74	1.55	2.54	3.75	6.00
Melting point (°C)	1277	650	838	763	714	700
Boiling point (°C)	2770	1107	1440	1380	1610	–
Atomic volume (c.c.)	4.90	13.97	25.9	34.54	36.7	38.0
Atomic (i.e., metallic) radius for co-ordination number 12 (A°)	1.12	1.60	1.97	2.15	2.22	–
Covalent radiue (A°)	0.90	1.36	1.74	1.91	1.98	–
Tonic (crystal radius of M^2 + ion for co-ordination number + (A°)	0.31	0.65	0.99	1.13	1.35	140
Ionisation energies (KJ/mole)						
I_1	899.5	737.7	829.8	547.5	502.9	509.4
I_2	1757.1	1450.7	1145.4	1064.3	965.2	979.06
$I_1 + I_2$	2656.6	2188.4	1735.2	1613.8	1468.1	1488.46
Oxidation stato	+2	+2	+2	+2	+2	+2
Electronegativity	1.5	1.2	1.0	0.9	0.9	0.9
Flame colouration	None	None	Brick red	Crim-son	Apple green	Red
Oxidation potentials (volts) for M (s) $\longrightarrow$ M^2 + (aq) + 2e	1.70 1.70	1.37	2.87	2.89	2.90	2.92
Heat of nutomisation at 25°C and 1 atm pressure (KJ/mole)	327.26	146.89	181.21	163.21	175.77	–
Heat of hydration (KJ/mole)	2385.45	1925.1	1653.07	1458.67	1276.42	–
Ionic potential of M2 + ion (i.e., charge/ radius ratio).	6.66	3.08	2.12	1.82	1.55	1.33

2) Trends in First Ionisation Energy

First ionisation energy is the energy needed to remove the most loosely held electron from each of one mole of gaseous atoms to make one mole of singly charged gaseous ions - in other words, for 1 mole of this process:

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

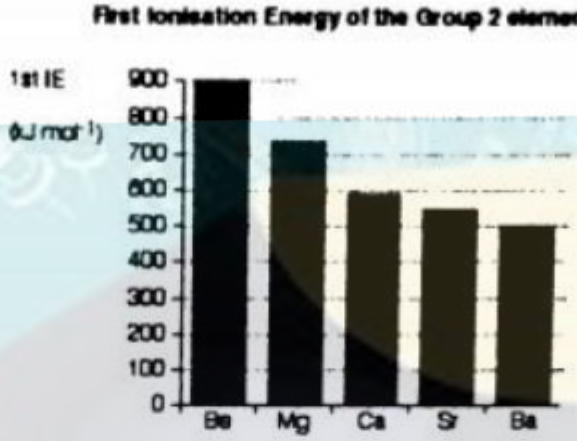


Fig. 13.18 First Ionisation energy of the Group 2 elements

Notice that first ionisation energy falls as you go down the group.

Explaining the decrease in first ionisation energy

The first and second ionisation energies of these elements decrease with the increase of atomic radii from Be to Ba. However, both these values for Ra are slightly higher than those of Ba (for values see table no 13.4)

3) Trends in Electronegativity

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. It is usually measured on the Pauling scale, on which the most electronegative element (fluorine) is given an electronegativity of 4.0.

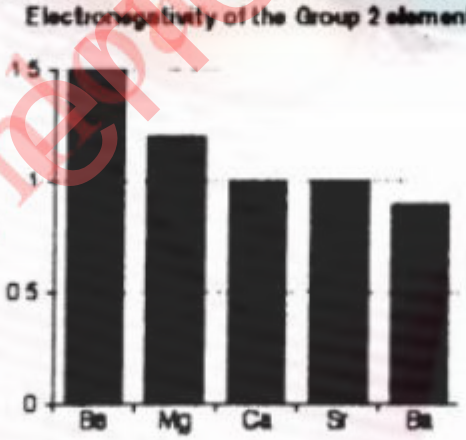


Fig. 13.19 Electronegativity of the Group 2 elements

Notice that electronegativity values fall as you go down the group (for values see table)

Summarising the trend down the Group

As the metal atoms get bigger, any bonding pair gets further and further away from the metal nucleus, and so is less strongly attracted towards it. In other words, as you go down the Group, the elements become less electronegative.

As you go down the Group, the bonds formed between these elements and other things such as chlorine become more and more ionic. The bonding pair is increasingly attracted away from the Group 2 element towards the chlorine (or whatever).

4) Trends in Melting Point and Boiling Point

A) Melting points

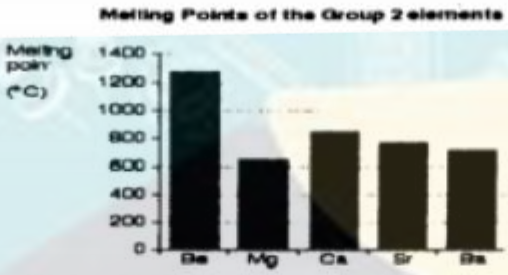


Fig. 13.20 You will see that (apart from where the smooth trend is broken by magnesium) the melting point falls as you go down the Group.

B) Boiling points

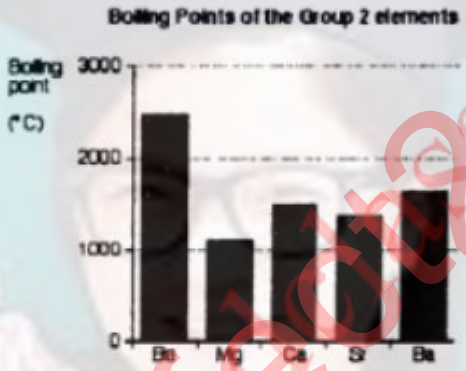


Fig. 13.21 You will see that there is no obvious pattern in boiling points. It would be quite wrong to suggest that there is any trend here whatsoever.

If we look at figure, the ionisation energies of carbon at the top of the group that there is no possibility of it forming simple positive ions, while Sn and Pb have low energies so form positive ion easily.

5) Trends in Reactivity with Water

This section looks at the reactions of the Group 2 elements - beryllium, magnesium, calcium, strontium and barium - with water (or steam). It uses these reactions to explore the trend in reactivity in Group 2.

The Facts

Beryllium

Beryllium has no reaction with water or steam even at red heat.

Magnesium

Magnesium burns in steam to produce white magnesium oxide and hydrogen gas.

$$Mg_{(s)} + H_2O_{(g)} \longrightarrow MgO_{(s)} + H_{2(g)}$$

Very clean magnesium ribbon has a very slight reaction with cold water. After several minutes, some bubbles of hydrogen form on its surface, and the coil of magnesium ribbon usually floats to the surface. However, the reaction soon stops because the magnesium hydroxide formed is almost insoluble in water and forms a barrier on the magnesium preventing further reaction.

$$CaO_{(s)} + SiO_{2(s)} \longrightarrow CaSiO_{3(l)}$$

Gernamium, tin and lead oxides

The monoxides

All of these oxides are amphoteric (they show both basic and acidic properties).

The basic nature of the oxides

These oxides all react with acids to form salts.

For example, they all react with concentrated hydrochloric acid. This can be summarised as:

$$XO_{(s)} + 2HCl_{(aq)} \longrightarrow XCl_{2(aq)} + H_2O_{(l)}$$

Where X can be Ge and Sn, but unfortunately needs modifying a bit for lead.

Lead(II) chloride is fairly insoluble in water and, instead of getting a solution, it would form an insoluble layer over the lead(II) oxide if you were to use *dilute* hydrochloric acid - stopping the reaction from going on.

$$PbO_{(s)} + 2HCl_{(aq)} \longrightarrow PbCl_{2(s)} + H_2O_{(l)}$$

However, in this example we are talking about using *concentrated* hydrochloric acid.

The large excess of chloride ions in the concentrated acid react with the lead(II) chloride to produce soluble complexes such as $PbCl_4^{2-}$. These ionic complexes are soluble in water and so the problem disappears.

$$PbCl_{2(s)} + 2Cl^-_{(aq)} \longrightarrow PbCl_4^{2-}{}_{(aq)}$$

The acidic nature of the oxides

All of these oxides also react with bases like sodium hydroxide solution.

$$XO_{(s)} + 2OH^-_{(aq)} \longrightarrow XO_2^{2-}{}_{(aq)} + H_2O_{(l)}$$

Lead(II) oxide, for example, would react to give PbO_2^{2-} - plumbate(II) ions.

The dioxides

These dioxides are again amphoteric - showing both basic and acidic properties.

The basic nature of the dioxides

The dioxides react with concentrated hydrochloric acid first to give compounds of the type XCl_4:

$$XO_2 + 4HCl \longrightarrow XCl_4 + 2H_2O$$

These will react with excess chloride ions in the hydrochloric acid to give complexes such as XCl_6^{2-}.

$$XCl_4 + 2Cl^- \longrightarrow XCl_6^{2-}$$

In the case of lead(IV) oxide, the reaction has to be done with ice-cold hydrochloric acid. If the reaction is done any warmer, the lead(IV) chloride decomposes to give lead(II) chloride and chlorine gas. This is an effect of the preferred oxidation state of lead being +2 rather than +4.

Q22. Write note on the acidic nature of the dioxides of Group IV-A elements.

Answer

The dioxides will react with hot concentrated sodium hydroxide solution to give soluble complexes of the form $[X(OH)_6]^{2-}$.

$$XO_{2(s)} + 2OH^-_{(aq)} + 2H_2O_{(l)} \longrightarrow [X(OH)_6]^{2-}_{(aq)}$$

Some sources suggest that the lead(IV) oxide needs molten sodium hydroxide. In that case, the equation is different.

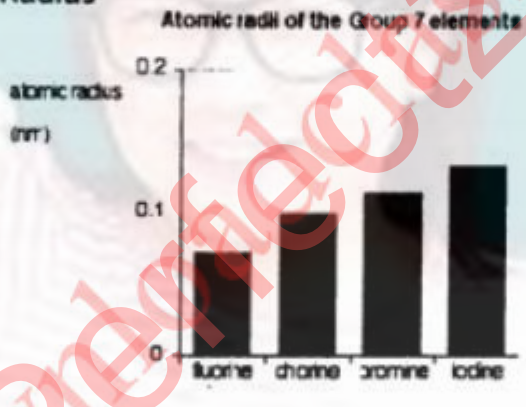
$$PbO_{2(s)} + 2NaOH_{(l)} \longrightarrow Na_2PbO_{3(s)} + H_2O_{(g)}$$

Group 7-Elements: (Halogens)

Atomic And Physical Properties

This article explores the trends in some atomic and physical properties of the Group 7 elements (the halogens) - fluorine, chlorine, bromine and iodine. You will find separate sections below covering the trends in atomic radius, electronegativity, electron affinity, melting and boiling points, and solubility. There is also a section on the bond enthalpies (strengths) of halogen-halogen bonds (for example, Cl-Cl) and of hydrogen-halogen bonds (e.g. H-Cl)

1) Trends in Atomic Radius



Trend

Atomic radius increases as we go down the group due to increase number of shell greater shielding effect and less nuclear change.

2) Trends in Electronegativity

Halogens have large values of electronegativity. These values decrease as we proceed from F to I in the group. Large electronegativities values of halogen atoms indicate that X atoms have a strong tendency to form X-ions.

3) Trends in First Electron Affinity

Electron affinity values decrease from Cl to I. why the electron affinity value of F is less than that of Cl has already been explained.

4) Trends in Melting and Boiling Points

The melting and boiling points of the halogens regularly increase form F to I. This indicates that the attractive forces between molecules become progressively more prominent as the molecules increase in halogens. F and Cl are gases at ordinary temperature Br is a heavy liquid while I is a solid.

==========

5) Bond enthalpies (bond energies or bond strengths)

Bond enthalpy is the heat needed to break one mole of a covalent bond to produce individual atoms, starting from the original substance in the gas state, and ending with gaseous atoms.

So for chlorine, $Cl_{2(g)}$, it is the heat energy needed to carry out this change per mole of bond:

$$Cl-Cl_{(g)} \longrightarrow 2Cl_{(g)}$$

For bromine, the reaction is still from *gaseous* bromine molecules to separate gaseous atoms.

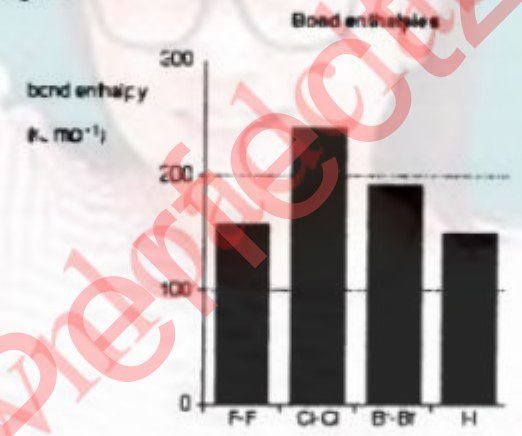
$$Br-Br_{(g)} \longrightarrow 2Br_{(g)}$$

Note GAS not liquid!

Q23. Discuss Bond enthalpy in the halogens, $X_{2(g)}$.

Answer

Look at following figure



The bond enthalpies of the Cl-Cl, Br-Br and I-I bonds fall just as you would expect, but the F-F bond is deviated from the sequence.

This is because of:

Due to very small F-F bond length very large as compared to other X-X bond lengths. This makes the F atoms in F_2 molecule repel each other and helps the dissociation of F_2 molecule into F atoms. (ii) X-X bond in Cl_2, Br_2 and I_2 molecules is stronger than F-F bond in F_2 molecule. This is due to the possibility of the existence of multiple bonds in X-X bond involving d- orbitals.

Q24. Discuss Bond enthalpies in the hydrogen halides, $HX_{(g)}$

Answer

Where the halogen atom is attached to a hydrogen atom, this effect doesn't happen. There are no lone pairs on a hydrogen atom.

==========

The acidity of the hydrogen halides

Hydrogen chloride as an acid

All the halogen acids in the gaseous states are essentially covalent but in the aqueous solution they ionise to give solvated proton (H_3O^+) and hence acts as acids.

$$H_2O + HX \rightarrow H_3O + X^-$$

HF ionises only slightly while HCl, HBr, and HI ionise completely. Hence HF is the weakest acid and strength of these acids increases from HF to HI, i.e. HF (weakest acid)<HCl<HBr<HI (strongest acid). The weakest acidic nature of HF is due to the fact that the dissociation energy of H-F bond is HF molecule is the highest and hence this molecule has least tendency to splits up into H^+ and F^- ions in aqueous solution. Another explanation of the above order of the acidic strength of HX acids can be given by finding out the relative order of the basicity of the conjugate bases viz F^-, Cl^-, Br^- and I^- of these acids. The hydrides show no acidic character when perfectly dry.

Q27. Discuss Halide Ions as Reducing Agents and Trends in Reducing Strength Ability of Halide Ions.

Answer

THE REDOX REACTIONS BETWEEN HALIDE IONS AND CONCENTRATED SULPHURIC ACID

This section describes and explains the redox reactions involving halide ions and concentrated sulphuric acid. It uses these reactions to discuss the trend in reducing ability of the ions as you go from fluoride to chloride to bromide to iodide.

Fluorides and Chlorides do not reduce concentrated sulphuric acid.

1) With bromide ions

The bromide ions are strong enough reducing agents to reduce the concentrated sulphuric acid. In the process the bromide ions are oxidised to bromine.

$$2Br^- \longrightarrow Br_2 + 2e^-$$

The bromide ions reduce the sulphuric acid to sulphur dioxide gas. This is a decrease of oxidation state of the sulphur from +6 in the sulphuric acid to +4 in the sulphur dioxide.

$$H_2SO_4 + 2H^+ + 2e^- \longrightarrow SO_2 + 2H_2O$$

You can combine these two half-equations to give the overall ionic equation for the reaction:

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

2) With iodide ions

Iodide ions are stronger reducing agents than bromide ions are. They are oxidised to iodine by the concentrated sulphuric acid.

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

The reduction of the sulphuric acid is more complicated than before. The iodide ions are powerful enough reducing agents to reduce it

3) first to sulphur dioxide (sulphur oxidation state = +4)

4) then to sulphur itself (oxidation state = 0)

5) and all the way to hydrogen sulphide (sulphur oxidation state = -2).
The most important of this mixture of reduction products is probably the hydrogen sulphide. The half-equation for its formation is:

$$H_2SO_4 + 8H^+ + 8e^- \longrightarrow H_2S + 4H_2O$$

Combining these last two half-equations gives:

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

Summary of the trend in reducing ability
1. Fluoride and chloride ions won't reduce concentrated sulphuric acid.
2. Bromide ions reduce the sulphuric acid to sulphur dioxide. In the process, the bromide ions are oxidised to bromine.
3. Iodide ions reduce the sulphuric acid to a mixture of products including hydrogen sulphide. The iodide ions are oxidised to iodine.
4. Reducing ability of the halide ions increases as you go down the Group.

Explaining the trend

1) When a halide ion acts as a reducing agent, it gives electrons to something else. That means that the halide ion itself has to lose electrons.
2) The bigger the halide ion, the further the outer electrons are from the nucleus, and the more they are screened from it by inner electrons. It therefore gets easier for the halide ions to lose electrons as you go down the Group because there is less attraction between the outer electrons and the nucleus.

Q28. Write a note on: Food and Beverage Canning.

Answer

As early as 1940, can manufacturers began to explore adapting cans to package carbonated soft drinks. The can had to be strengthened to accommodate higher internal can pressures created by carbonation (especially during warm summer months), which meant increasing the thickness of the metal used in the can ends. Otherwise, distortion of the end would strain the seal, creating potential leaks or making cans unstackable for storage and transit.

Another concern for the new beverage can was its shelf life. Even small amounts of dissolved tin or iron from the can could impair the drinking quality of both beer and soft drinks. Fortunately, beer, which is only mildly acidic, is relatively noncorrosive. In addition, beer ages naturally, so it has a limited shelf life of about three months in any package. In contrast, the food acids, including carbonic, citric and phosphoric, in soft drinks present a risk for rapid corrosion of exposed tin and iron in the can. The consequences of off-flavors, color changes and leakage through the metal needed to be addressed. At this point, the can was upgraded by improving the organic coatings used to line the inside, making cans heavier and more encasing.

The use of cans for carbonated beverages was delayed because of wartime material limitations mandated by the U.S. government. When the restriction ended in 1953 after the Korean War, the improved can was introduced and marketed nationwide. The can manufacturers then embarked on a program of material and cost savings by reducing both the amount of steel and the amount of coatings used in can making. These efforts were in part inspired by a new competitor — aluminum.

Q29. Write a note on Mining and Extraction of metals.

Answer

Different elements/metals are not obtained such rather these are obtained after passing through different steps.

These steps are discussed as follows:

1. Mining and enrichment
2. Reduction
3. Refining and Casting

In fact some special methods are used to obtain each metal from its ores and to develop it into useful articles, yet few steps are common in the metallurgy of every metal. These are follows.

1. Mining

i) Crushing

Obtaining ores by digging the rocks and hills is called mining. This work is done by engineers and laborers with the help of machines. But prior to this work it is confirmed by survey and analysis that obtaining metals from this is economical or not.

ii) **Grinding**

Breaking of rocks and larger stones into smaller size stones is called crushing. This is done by jaw crushers.

iii) **Hand Picking, Jugging and Shaking**

In Pakistan and other under developed countries where labour is cheap, metallic stones are picked and separated by hands. Heavy metals are separated from useless material i.e. gangue, by shaking with "chaage". In some countries this process is done by pressurized water.

iii) **Magnetic Separation**

The ground ore is passed over a magnetic belt which separates the magnetic metal from gangue. This process is sued for metals which have magnetic properties like iron.

2) Reduction

For the complete separation of a metal from gangue, ores are heated at high temperature. At its melting pint, molten metal is separated from solid gangue. It must be remembered that different metals are mixed with different compounds according to the type of impurities present in the metal ore and then they are passed through the process of reduction. The process of reduction is carried out in the blast furnace.

Blast Furnace

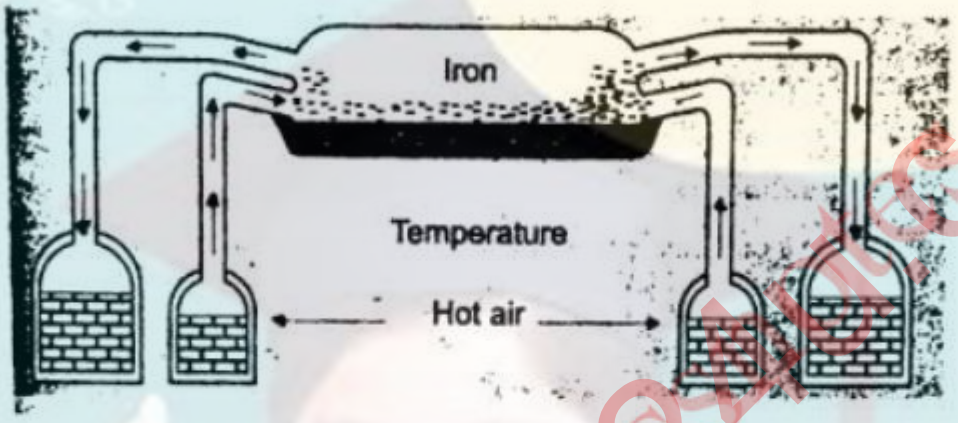
It is lined inside with fire bricks. Its height and capacity are kept according to the requirement. Hot gases enter from lower side and ores are charged from the tope of the furnace. Temperature is maintained at 1500°-3000°C. This furnace is usually used for iron and copper metallurgy.

3) Refining of Metals

Metals extracted in the above process are further refined by the following process.

Open-Hearth Process:

A fire furnace is used to remove the impurities of metal. It is lined inside with fire bricks and is just like a room. Burning gases are entered from one side and exhaust gases are removed from the opposite end. The process is operated form opposite ends after an interval. Metals melt in a shorter time by this two way heating.



Q30. Give applications of bleaching powder.

Answer

Bleaching powder is actually a mixture of calcium hypochlorite ($Ca(OCl)_2$) and the basic chloride $CaCl_2$, H_2O with some slaked lime, $Ca(OH)_2$.

Bleaching powders take time to dissolve in water and longer to work but have a longer shelf life in comparison to liquid bleaches and can be used on items like upholstery, carpet and some delicate fabrics. However, bleaching powder should never be combined with ammonia or used on colored fabrics as it will cause fading.

1) Bleaching powder is highly effective for cleaning inside the home and outdoors. It can be used for removing mildew from fabric, cleaning countertops and for removing mold from grout between tiles, bathmats and shower curtains. Outside, the agent can be used on plastic furniture, unpainted cement, paving and painted surfaces to eliminate mildew and other stubborn stains.

2) Bleaching powder can be used to safely disinfect and sterilize many things around the home including secondhand goods, trash cans, pet accessories and baby toys and furniture.
Bleaching powder is a highly effective means of returning the luster to white porcelain and glassware. Glassware can regain its sparkle by adding a small amount of powdered bleach to dishwater when washing glasses. Gardening

3) To kill any annoying weeds growing from cracks and crevices in the garden a strong mixture of bleaching powder and water is applied. Moss and algae on garden walkways can be easily eliminated by scrubbing with bleaching powder diluted in water. Powdered bleach is also useful for sanitizing garden tools to avoid diseases spreading between plants. Adding powdered bleach to the water of cut flowers will help to preserve their freshness by preventing the growth of bacteria in the vase.

4) Bleaching powder is used for the disinfection of drinking water or swimming pool water. It is used as a sanitizer in outdoor swimming pools in combination with a

cyanuric acid stabilizer, which reduces the loss of chlorine due to ultraviolet radiation. The calcium content hardens the water and tends to clog up some filters; hence, some products containing calcium hypochlorite also contain anti-scaling agents.
Bleaching powder is used for bleaching cotton and linen. It is also used in bathroom cleaners, household disinfectant sprays, moss and algae removers, and weedkillers.

In addition, bleaching powder may be used to manufacture chloroform.

Bleaching powder is used also in sugar industry for bleaching sugar cane juice before its crystallization.

Q31. Give Commercial Uses of Halogens.

Answer

1. Chlorine is used as a cheap industrial oxidant in the manufacture of bromine
2. Iodine is dissolved in alcohol, commonly known as tincture of iodine is used as a mild antiseptic for cuts and scratches. Iodine is also mixed with the detergents used in cleaning diary equipment.
3. Small quantities of fluorine are used in rocket propulsion. Much larger quantities are used make uranium (VI) fluoride for the separation of ^{238}U and ^{235}U:

$$UF_4\,(s) + F_2\,(g) \rightarrow UF_6\,(s)$$

4. Fluorine is also used to make a wide range of fluorocarbon compounds for use as refrigerants, aerosol propellants, anaesthetics and fire-extinguisher fluid. One of the most important fluorocarbons is poly (tetrafluoroethene), PTFE, frequently sold under the trade name Fluon or Teflon.

Q32. Write a note on: Iodine Deficiency and Goiter.

Answer

Iodine Deficiency

Iodine is an element that is needed for the production of thyroid hormone. The human body cannot synthesize iodine, so it is an essential element.

1. The deficiency of iodine leads to enlargement of thyroid a condition called goiter. Hypothyroidism and mental retardation in children and infants is observed if their mothers suffered from iodine deficiency during pregnancy.

   Before 1920, iodine deficiency was common in Appalachian, north-western US regions, and in most of Canada. Approximately, 40% of the world's population remains at the risk of iodine deficiency.

Goiter

- The term goiter refers to the abnormal enlargement of thyroid gland due to deficiency of iodine in diet. It results in swelling in neck. It is important to know that the presence of goiter does not necessarily mean that the thyroid gland is malfunctioning (hypothyroidism). A goiter can also occur in a gland that is producing too much thyroid hormone (hyperthyroidism) or even the correct amount of hormone (euthyroidism). A goiter indicates there is a condition present which is causing the thyroid to grow abnormally.

==========

Q33. What is Fluoride Deficiency and Toxicity?

Answer

Fluoride Toxicity

Fluoride toxicity or fluoride poisoning is a condition in which more fluoride is taken than the amount required for normal growth, development and metabolism. Fluoride toxicity is characterised by a variety of signs and symptoms. Poisoning most commonly occurs following ingestion of conspicuous amount of fluoride entraining products. Symptoms or set usually occurs within minutes of exposure. Fluoride is found in many common household products e.g. toothpaste, dietary supplementary, insecticides, roderticides etc. fluoride toxicity results,

i) Arthritis
ii) Stiff painful joints with or wistful swelling
iii) Asthma, especially after showering
iv) Painful bony lumps where tendons and ligaments attach to bones

Fluoride Deficiency

Fluoride deficiency results when the amount of its up take is less than required.

Fluoride deficiency results in

Brittle bones or demineralization of bones

Brittle bones or demineralization of bones

Cavities

Weakened tooth enamel

Fluoride deficiency can lead to a higher likelihood of developing bone fractures and possibly even steoporosis.

Halogens and their compounds are used or bleaching, refrigeration and as aerosols, etc.

EXERCISE

Q1: Select the right answer from the choices given with each question.

1) **Oxides and hydroxides of Group1 elements are:**
   a) Acidic b) Alhaline
   c) Neutral d) Amphoteric

2) **The flame colour of sodium metal or its compounds is:**
   a) A, Bright crimson b) Voilet
   c) Golden yellow d) Bright blue

3) **When sodium burn in air , it forms sodium:**
   a) Monixide b) Peroxide
   c) Oxide d) Superoxid

4) **The carbonates of alkali metal are not affected by heat except:**
   a) Li_2CO_3 b) Na_2CO_3

==========

c) K_2CO_3 d) Rb_2CO_3

5) **Green is characteristic flame colour of:**
a) Calcium b) Barium
c) Strontium d) Sodium

6) **All the carbonates, sulphates and phosphates of alkaline earth metals are _____ in water.**
a) Sprangly b) Soluble
c) Insoluble d) Less soluble

7) **The first ioniziatio energy is higher for the:**
a) Alkaline earth metals b) Alkali metals
c) Halogens d) Nobel gases

8) **Which one of the element has the maximum electron affinity?**
a) F b) Cl c) Br d) I

9) **Which pair has both members from same period of periodic table?**
a) Na – Ca b) Na – Cl c) Ca – Cl d) Cl – Br

10) **Melting points and boiling points of alkali metals:**
a) Decreases from top to bottom b) Increases from top to bottom
c) First increases then decreases d) Remains uncharged

11) **Which of the following oxides is amphoteric in nature?**
a) Rubidium oxide b) Barium oxide
c) Antimony d) Sulphur oxide

12) **Oxidizing power of halogen depends upon:**
a) Energy of dissociation b) Electron affinity
c) Heat of vaporization d) All of above

13) **Which of following oxides amphoteric is nature?**
a) Mgo b) BeO c) CO_2 d) SnO_2

14) **Select the correct increasing order of atomic radius?**
a) Ne > O > S > Al b) Ne < O > S > Al
c) Ne < O > S > Al d) Ne > O < S > Al

15) **Due to inert pair effect _____ oxidation state is more stable than _____ For Sn and Pb.**
a) 2+, 4+ b) 1+, 4+ c) 4+, 2+ d) 2+, 3+

16) **Highest electron affinity is shown by?**
a) F_2 b) I_2 c) Br_2 d) Cl_2

17) **Which is the strongest reducing agent?**
a) HF b) HCl c) Br_2 d) HBr

18) **Substance boiling at higher temperature among following is?**
a) HI b) HF c) HCl d) HBr

19) **Group VII-A elements are generally called:**

a) Halogens b) Nobel gases
c) Inert gases d) Metallods

20) **The radioactive element in halogen group is:**

a) Radon b) Radium c) Astatine d) Bromine

Answers

1)	b)	2)	c)	3)	b)	4)	a)	5)	b)
6)	c)	7)	d)	8)	b)	9)	b)	10)	a)
11)	c)	12)	b)	13)	b)	14)	c)	15)	a)
16)	d)	17)	d)	18)	b)	19)	a)	20)	c)

Q2: Give brief answers for the following questions.

Q1. Although Na and P are present in the same period their oxides are different in nature Na_2O is basic while P_2O_5 is acidic why?

Answer

Na and P are present in same period But Na_2O is basic and P_2O_5 is acidic:

Na is an alkali metal present in IA group and has one electron in its outermost shell, It gives a strong base (NaOH) on dissolving in water.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

Phosphorus is member of group – VA. It is non-metal and has five electrons on its outermost shell. P_2O_5 on dissolving in cold water produce meta phosphoric acid and on dissolving in hot water give orthophosphoric acid.

$$P_2O_5 + H_2O \xrightarrow[\text{(Cold)}]{} 2HPO_3 \quad \text{(metaphosphoric acid)}$$

$$P_2O_5 + 3H_2O \xrightarrow[\text{(Hot)}]{} 2H_3PO_4 \quad \text{(orthophosphoric acid)}$$

Q2. How acidic basic and amphoteric behaviour of oxides is explained?

Answer

Basic, Acidic and amphoteric behaviour of oxides:

Odides of alkali metals are basic in nature.

$$\underset{\text{(Base)}}{Na_2O} + \underset{\text{(Acid)}}{H_2SO_4} + H_2O \longrightarrow Na_2SO_4$$

Oxides of Be Saluminium are amphoteric in nature e.g $SO_3 + P_2O_5$

$$\underset{\text{(acid)}}{SO_3} + \underset{\text{(Base)}}{Na_2O} \longrightarrow \underset{\text{(Salt)}}{Na_2SO_4}$$

Oxides of Be saluminium are amphoteric

$$\underset{\text{(Acid)}}{BeO} + 2NaOH \longrightarrow \underset{\text{(Sodium beryllate)}}{Na_2BeO_2} + H_2O$$

$$\underset{\text{(Base)}}{BeO} + H_2SO_4 \longrightarrow BeSO_4 + H_2O$$

Q3. Why the elements of group 1 are called alkali metals?

Answer

Elements of group IA are metals and their oxides and hydroxides are alkaline in nature.

e.g $Na_2O + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$

$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$

Q4. Why all group of metals have low ionization energies?

Answer

All IA – group elements have one electron on their outermost shell. Except lithium all alkali metal have strong shielding or screening effect. As a result pull of nucleus on outermost electron decreases thats why outermost electron of alkali methods may be removed easily using low energy.

Q5. Why do the group 1 metals show strong electropositive character?

Answer

The metals which lose outermost electrons easily and give a stable cation are called (electropositive. And group IA metals lose their outermost single electron easily to complete their outermost stable octet Therefore they are electropositive metals.

Q6. Why do group 1 metals show strong reducing properties?

Answer

All alkali metals have one electron in their outermost shell and on losing one outermost electron their octet is completed which is a stable state. Therefore they try to lose their valence electron and all as good reducing agent.

Q7. Why different colours are imported by the atoms of the group 1 metals to the flame?

Answer

The outermost electron of group IA is ns^1. This electron is loosely hold with the nucleus.

When an element is brought to a flame of a burner the electron absorbs some energy and goes to higher energy orbital and is called excited. This excited electron absorbs energy during process of excitation Atoms of different elements absorb different energies therefore on dexitation they impart different colours to flame.

Q8. Why the elements of group 2 are called alkaline earth metals?

Answer

II – A group members are metals, their salts are found in earth and their oxides & hydroxides alkaline in nature. Therefore they are called alkaline earth metals.

e.g $CaO + H_2SO_4 \longrightarrow CaSO_4 + H_2O$

$Ca(OH)_2 + H_2SO_4 \longrightarrow CaSO_4 + 2H_2O$

Q9. Why do the group 2 earth metals have high melting and boiling points than alkali metals?

Answer

The alkaline earth metals have two electrons in their outermost shell whereas group II – A have one electron in their outermost shell. Therefore there are strong metal–metal

bonds among group IIA. Therefore II – A members are harder and stranger than group I–A members. Therefore the melting and boiling point of group II–A members are higher than group I – A members.

Q10. How do group 1 metals resemble with group 2 metals?

Answer

Both group I–A and group II–A members

1) are metals.
2) are iso-electronic up to subshell.

e.g Li = 3 $Is^2, 2S^1$

Be = 4 $Is^2, 2S^2$

3) have alkaline oxides and hydroxides

e.g Na2O, MgO

NaOH, $CaCOH_2$

4) are good conductors for heat and electricity can react with halogens to form salts. e.g NaCl, $CaCl_2$

Q11. How do group 1 metals differs from group 2 metals?

Answer

Group IA Metals	Group IIA Metals
1) Their have one electron in their outermost shell. 2) They are monovalent i.e M^{+1}.	Their atoms have two electons in their outermost shell. They are divalent i.e M^{+2}
3) They are soft metals.	They are hard metals
4) Their melting and boiling points are lone.	Their melting and boiling points are high.
5) Their carbonates are solube in water.	Their carbonates are insoluble in water.
6) Their carbonates are stable towards heat $Na_2CO_3 \xrightarrow{\Delta} X$	Their carbonates decompose on heating. $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$
7) Their hydroxides are stable towards heat $NaOH \xrightarrow{\Delta} X$	Their hydroxides decompose on heating. $Ca(OH)_2 \xrightarrow{\Delta} CaO + H_2O$

Q12. Discuss the metallic and non-metallic character of group 4 elements.

Answer

Metallic and Non Metallic charcter of Group IV A Elements.

Metallic character of Group iv – A elements increase from top to bottom.

iv – A.

C, Si — Non – metals

Ge – semi metals

Sn, Pb – Metals

Q13. Discuss the general trends of group 7 elements.

Answer

General Group Trends of Group VII – A Elements

i) Trends in physical states:

As attractive forces among their molecules increase from F_2 to I_2, therefore.

F_2, Cl_2 — gases

Br_2 – Liq

I_2 – solid

Trends in Electronegativity

Electronegativity decrease from top to bottom in halogen group. i.e. Decreases from F to I.

iii) Trends in Electron Affinity

Due to high electron density, the E.A value of E is low. And E.A value of remaining halogens decrease from Cl to I.

iv) Bond Enthalpy in halogens

Due to very small size and high electron density F atoms repell each other in F_2. As a result the bond dissociation. Energy of F_2 is lower than the remaining halogens B D E of remaining halogens decrease from Cl_2 to I_2

v) Trends as oxidising agents

Oxidising ability of halogens decrease from F2 to I_2.

$$F_2 > Cl_2 > Br_2 > I_2$$

vi) Trends of Activity of Halogen halides

Due to strong H – bonds, the HF ionizes slightly, therefore is a weak acid. Other halo acids ionize completely and are strong acids.

Their acidic strength increases from HF to HI

$$HF < HCl < HBr < HI$$

(weakest acid) (Strong acid)

Q14. Why the term halogen is used for group 7 elements.

Answer

Halogen means salt forming. The elements of Group VII–A on reaction with alkali and

(As Base)

Q8. Explain the trends in oxidation states with suitable examples.

Answer

Trends In Oxidation State

The ability of an tom to loose or gain the number of electrons is called its oxidation state.

Trends within a Period

When we move from left to right oxidation state increases from upto mid of the periodic table

Q9. Discuss the inert pair effect in the

i) Formation of ionic bonds **ii) Formation of covalent bonds**

Answer

Inert pair Effect and Ionic Bond

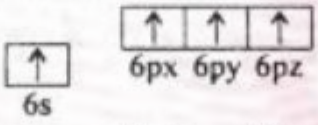
In ground state the electronic configuration of group IV-A elements is n^2p^2. When we move down the group atomic size increases due to addition of new shells. They loos p-electrons leaving the s^2 pair unused this is called inert pair. For example to form Pb^{+2} ion lead will loos two 6p– electrons but the 6s – electrons will be left unused as an inert pair. Also the ionization energy decreases to form 2 + ions therefore the chances of formation of ionic bonds increases specially in case of Sn to form Sn^{+2} and Pb to form Pb^{+2} that is whey Sn and Pb like to form ions bonds.

ii) Inner pair Effect and Covalent Bond

According to theory of relativity the behaviour elements like Sn, Pb, Pt, Si difficult to remove the electros due to relativistic contact on. This effects "S" electrons more than "P" electrons. And one electron moves from S-orbital to empty p-orbital

Electrons configuration of Pb in ground state.

Electronic configuration of Pb in exited state =



The energy is released when atoms of halogen like chlorine come to attach with tin or lead. Therefore no extra energy is needed to promote an electron from 6s orbital to 6p-orbital. Now these four unpaired electrons like to form covalent bonds.

Q10. Discuss in detail acid base trend of group IV–A oxides?

Answer

Acid-base trends in grup IV–A oxides.

Group IV–A oxides are as follows:

Co – Acidic

SiO_2, GeO_2 – Neutral

SnO, PbO_2 – Basic

Top most oxide of group IV–A are acidic in nature. CO_2 on dissolving in water produce carbonic acid

$$CO_2 + H_2O \longrightarrow H_2CO_3 \text{ (carbonic acid)}$$

Lowermost oxide of group IV–A may gave O^- ion which on dissolving in water produce OH ion and behave as Lu- flood bases.

$$PbO_2 \longrightarrow Pb^{4+} + 2O^{2-}$$

$$O^{-2} + H_2O \longrightarrow 2OH^{-1}$$

Q11. Explain in detail the trends in group VII– of following physical properties.

i) Electronegativity **ii) Electroneffinity**

Answer

Trends of Electronegativityin group VIII–A

Due to increase in atomic size and increase in shielding effect the electronegativityin group VII–A decreases from top to bottom

Element	E.N
F	4.0
Cl	3.2
Br	3.0
I	2.5
At	2.2

Trends of Electron effinity in group VII–A

Electron effinity of F has thick value due to high charge density (thick electronic cloud) whereas electron effinity in group VII–A decreases from Cl to At.

Element	Electron effinity
F	–322
Cl	–349
Br	–325
I	–295
At	–270

Q12.a) Why is the bond enthalpy of F–F less as compared to Cl – Cl and Br– Br?

==========

b) Explain the order F>Cl> Br>I with respect to oxidizing agent or power.

Answer

a) Why is the bond enthalpy of F–F less as compared to Cl – Cl and Br– Br?

Due to very high electronegeitivity of F attract the electrons towards their nuclei with greater force which hardly allows the electrons to be shared to form a covalent bond. Small amount of supply of energy increases the Kinetic energy of atom and break the bond between atoms.

In case of Cl_2 and Br_2 the electronegativityof Cl and Br atoms are moderate and easily allow atoms to share the electrons to form stable covalent bond therefore Cl_2 and Br_2 exist stable gaseous molecules.

The order F>Cl> Br>I with respect to oxidizing agent

The electronegativityorder of halogen is:

i) Element F> Cl> Br> I

Electronegetivity 4.0>, 3.2>, 3.0>, 2.5

Due to highest value of electronegativityF may attract and lake the electron from a substance and oxidize with more force as the order of electronegativitydecreases towards I therefore oxidizing power towards I also decreases.

Q13. a) Why is flourine much stronger oxidizing agent than chlorine?
b) HCl is strong acid as compared to HF. Why?

Answer

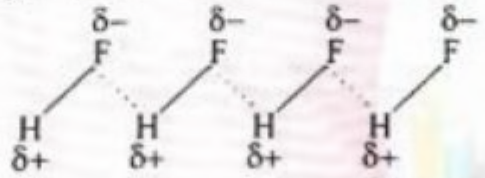
a) Why is flourine much stronger oxidizing agent than chlorine?

The electronegativityof F is 4.0 which is maximum in periodic table due to its high electonegitivity it may take an electrons from a sub stance to complete its octate whereas Cl has somewhat lower electronegativityand has low oxidizing power.

b) HCl is strong acid as compared to HF. Why?

Due to very high electronegativity F, the molecule HF is highly polar.

There are strong hydrogen bonds between HF molecules.



As a result the power of HF to release proton is decreases. As a result HF behaves as weaker acid than HCl which is completely ionized on dissolving in water.

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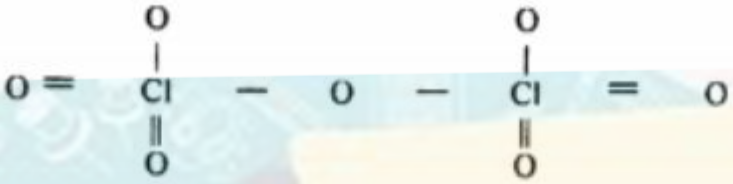


Q14: Why the melting and boiling points of metal oxide and silicon dioxide is high?

Answer

The metal oxide s and silicon dioxide (the gaint structures) will have high melting and boiling point because a lot of energy is needed to break the strong bonds operating in 3 dimensions.

The attractive forces between the oxides of phosphorus, sulphur and chorine will be van der waals dispersion and dipole dipole interaction. These way in size depending on the size shape and polarity of the various molecules but will always be much weaker than the ionic on covalent bonds need to break in a giant structure.

Q15: Write down the reaction of sodium oxide (Na_2O) with acids and H_2O.

Answer

Sodium oxide (Na_2O) is a simple strongly basic oxide. It is basic because it contains the oxide ion, O^{2-} which is very strong base with a high tendency to combine with hydrogen ions.

Rex with H_2O

Na_2O reacts exothermically will cold H_2O produce sodium hydroxide solution.

$$Na_2O + H_2O \longrightarrow 2Na_2OH$$

Rex with acids

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

Q16: Write down the reaction of Magnesium Oxide (MgO) with H_2O and acids.

Answer

MgO is simple basic oxide because it also it also contain oxide ions. In Magnesium oxide case the attraction are between 2 + and 2 – it takes more energy to break these.

Rex with H_2O

If your shake some write magnesium oxide with H_2O it doesn't look as if it reacts. If you test the PH of the liquid, you find that it is somewhere around PH 9. The must be some slight rex with H_2O to produce hydroxide ions in solution. Some $Mg(OH)_2$ is formed in the reaction, but this is almost insoluble and so not many hydroxide ion actually get into solution.

Rex with acid

MgO react with acid to form $MgCl_2$ solution.

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

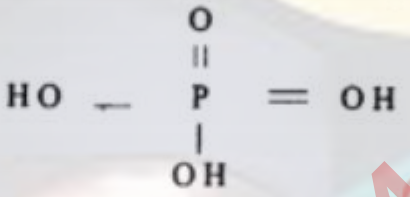
==========

Q17: Write down the reaction of phosphorus (V) oxide (P_4O_{10}) with H_2O and Base?

Answer

Phosphorus (V) oxide reacts violently with H_2O to give a solution containing a mixture of acids, the nature of which depends on the conditions, use usually ust consider one of these, phosphoric acid H_3PO_4 (also known just as phosphoric acid or as orthophospheric acid).

$$P_4O_{10} + 6H_2O\ 4H_3PO_4 \longrightarrow 4H_3PO_4$$



Reactions with Base

$$NaOH + H_3PO_4 \longrightarrow NaH_2PO_4 + H_2O$$
$$2NaOH + H_3PO_4 \longrightarrow Na_2HPO_4 + 2H_2O$$
$$3NaOH + H_3PO_4 \longrightarrow Na_3HPO_4 + 3H_2O$$

Again, if you were to react phosphorus (V) oxide directly with sodium hydroxide solution rather than making the acid first you would end up with the same possible salts.

This is getting ridiculers, and so will only give one example out of the possible equations.

$$12NaOH + P_4O_{10} \longrightarrow 4Na_3PO_4 + 6H_2O$$

Q18: Write down the reaction of sulphur dioxides (SO_2) with H_2O and base?

Answer

Reaction with H_2O

Sulphur dioxide is fairly soluble in H_2O reacting which it to give a solution of sulphurs acid H_2SO_3.

$$SO_2 + H_2O \longrightarrow H_2SO_3$$

Reaction with Base

As it is acidic so it reacts with NaOH and CAO as follows.

$$SO_2 + 2NaOH \longrightarrow Na_2SO_3 + H_2O$$
$$Ca + SO_2 \longrightarrow CaSO_3$$

Q19: Write down the reaction of sulphur trioxide with Base?

Answer

Sulphur trioxide reacts violently with H_2O to produce a fog of concentrated sulphuric acid dioplets.

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

==========

Rex with Base

$$2NaOH + H_2SO_4 \longrightarrow Na2SO_4 + 2H_2O$$

Sulphur trioxide itself will also reacts directly will bases to form sulphates. e.g. it. will react with CaO to form calcium sulphates.

$$CaO + SO_3 \longrightarrow CaSO_4$$

Q20: Write down the reaction of chlorine oxides (Cl_2O) with H_2O and Base.

Answer

Chlorine forms several oxides but the only two are chlorine (VII) oxide Cl_2O_7 and chlorine (1) oxide, Cl_2O, and chlorine (1) oxide, Cl_2O. Chlorine (VII) oxide is also known is dichlorine heptoxide and chlorine (1) oxide as dichlorine monoxide.

Chlorine (VII) oxide

Chlorine (VII) oxide is the highest oxide of chlorine the chlorine is in its maximum oxidation state of +7. It continues the trend of the highest oxides of the eriod 3 elements towards being stronger acids.

Reaction with H_2O

Chlorine (VII) oxide reacts with H2O to give the very strong acid chloric (VII) acid also known as perchloric acid. The PH of typical solution will sulphuric acid. The PH of typical solution will sulpohuric acidbe around O.

$$Cl_2O_7 + H_2O \longrightarrow 2HClO_4$$

Reaction with Base

Chloric (VII) acid reacts with NaOH solution to form a solution of sodium chlorate (VII).

$$NaOH + HClO_4 \longrightarrow NaClO_4 + H_2O$$

Chlorine (VII) oxide itself also react with sodium hydroxide solution to give the same product.

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

Q21: Write down the reaction of chlorine (1) oxide with Base?

Answer

Chlorine (1) oxide is for less acidic than chlorine (VII) oxide. It reacts with water to some extent to give chloric (1) acid, HOCl also known as hypochloeous acid.

$$Cl_2O + H_2O \rightleftharpoons 2HOCl$$

Reaction with Base

Chloric (1) acid reacts with sodium hydroxide solution to give a selection of sodium chlorate (1) (sodium hypochlrite).

$$NaOH + HOCl \longrightarrow NaOCl + H_2O$$

Chlorine (1) oxide also reacts directly with sodium hydroxide to give the same product.

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

==========

Q22: Give a detail note on sodium chloride NaCl.

Answer

Sodium Chloride is a simple ionic compound consisting of a gaint arrays of sodium and chloride ions.

The strong attractions between the positive and negative ion need a lot of heat energy to break, and so sodium chloride has high melting and boiling points.

If doesn't conduce electricity in the solid state because it hasn't any mobile electrons and the ions aren't free to move. However, when it melts it undergoes electrolysis.

Sodium chloride simply dissolves in H_2O to give a neutral solution.

Q23: Give a electric note on Magnesium chloride $MgCl_2$.

Answer

Magnesium chloride is also ionic, but with a more complicated arrangement of the ions to allow for having twice as many chloride ions as magnesium ions.

Again, lots of heat energy is needed to overcome the attractions between the ions, and so the melting and boiling points are again high.

Solid magnesium chloride is a non conductor of electricity because the ions are'nt free to move. However it underogoes electrolysis when the ions become free on melting.

$MgCl_2$ dissolves in H_2O to give a faintly acidic solution (PH = approximately 6). When magnesium ions are broken off the solid lattice and go into solution, there is enough action between the 2+ ions and the water molecules to get co-ordinate bonds formed between the magnesium ions and lone pairs on surrounding H_2O molecules.

Hexaquamangnesium ions are formed $\left[Mg(H_2O)_6\right]^{2+}$

$$MgCl_2 + 6H_2O \longrightarrow \left[Mg(H_2O)_6\right]^{2+} + 2Cl^- \text{ ions are acidic.}$$

Q24: Give a detail note on Aluminium chloride $AlCl_3$.

Answer

At room temperature, solid aluminium chloride has an ionic lattice with a lot of covalent character.

At temperature around 180 – 190°C (depending on the pressure aluminium chloride connects to a molecular form Al_2Cl_6. This causes it to melt or vaporise because there are now only comparatively weak inter molecular attractions. As the temperature increases a bit more, it increasingly breaks up into simple $AlCl_3$ molecules.

Solid aluminium chloride doesn't conduct electricity at room temperature because the ion are not free to move.

Molten aluminium chloride doesn't conduct electricity because there are not any ions any more.

The aluminium chloride react with H_2O ratio than just dissolving in it. In the first existence, hexaaquaaluminium ions are formed together with chloride ions.

$$AlCl_3\,6H_2O \longrightarrow \left[Al(H_2O)_6\right]^{3+} + 3Cl^-$$

==========

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Q25: Write down the details note on silicon tetrachloride $SiCl_4$.

Answer

$SiCl_4$ is a simple covalent chloride between the silicon and the chlorine for the two to form ionic bonds.

Silicon tetrachloride is a colourless liquid at room temperature which fumes in moist air. The only attractions between the molecules are van der waals dispersion forces.

It doesn't conduct electricity because of the lake of ions mobile electrons.

It fumes in moist air because it least with H_2O in the air to produce hydrogen chloride. If you add H2O to silicon tetrachloride there is a violent reaction to produce silicon dioxide and fumes of hydrogen chloride.

In a large excess of H_2O the hydrogen chloride will of cause, dissolve to give a strongly acidic solution containing hydchloride acid.

$$SiCl_4 + 2H_2O \longrightarrow SiO_2 + 4HCl$$

Q26: Write down the reaction of phosphorus chlorides with H_2O.

Answer

There are two phosphorus chlorides phosphorus (III) chloride, Pcl_3 and phosphorus (V) chloride Pcl_5. Phosphorus (III) chloride (phosphorus trichlorde) Pcl_3.

This is simple covalent chloride-again a fuming liquid at room temperature. Phosphorus (III) chloride react violently with H_2O.

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

Phosphorus (V) chloride (Phosphorus pentachloride) PCl_5.

Phosphorus (V) chloride is a white solid which sublimes at 163°C. The higher the temperatue goes above that, the more the phosphorus (V) chloride dissociate to give phosphorus (III) chloride and chlrine.

$$PCl_5 \rightleftharpoons PCl_3 + Cl_2$$

Solid phosphorus (V) chloride contain ions that is a solid at room temperature. The formation of ions molecules two molecules of PCl_5.

Phosphorus Chloride has a violent reaction will H_2O producing fumes of hydrogen chloride.

As with the other covalent chlorides, if there is enough H_2O present, these will dissolve to give a solution containing hydrochloric acid.

$$PCl_5 + H_2O \longrightarrow POCl_3 + 2HCl$$

If H_2O is boiling, the phosphorus (V) chloride reacts further to give phosphorus (V) acid and more HCl. Phosphoric (V) acid is also known just as phosphoric acid or as ortho phosphoric acid.

$$POCl_3 + 3H_2O \longrightarrow H_3PO_4 + 3HCl$$

The overall equation in boiling H_2O is just a combination of these.

$$PCl_5 + 4H_2O \longrightarrow H_3PO_4 + 5HCl$$

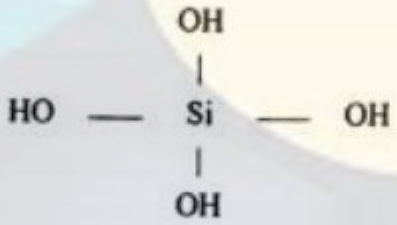
==========

Q27: Write down the formula of following:

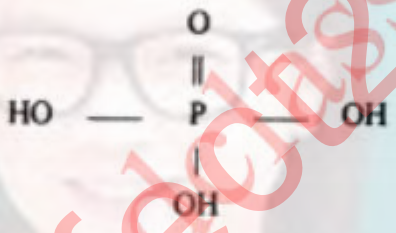
1) **Ortho silicic acid** 2) **Phospheric acid**
3) **Sulphuric acid** 4) **Chloric acid**

Answer

Orthosilicic acid.

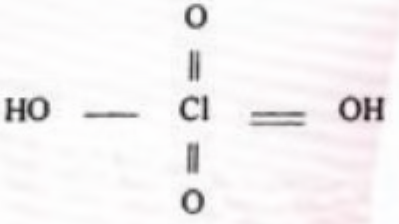


Phospheric acid



Sulphuric acid

Chloric acid



Q28: Aluminium hydronuclie is amphoteric. Justify it.

Answer

Aluminium hydroxide is amphoteric like Na and Magnesium hydroxides, it will least will acids. This is showing the basic side of its nature.

With dilute hydrochloric acid, a colourless solution of aluminium chloride is formed.

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

But aluminium hydroxide also has an acidic side to its nature. It will react with sodium

hydroxide solution to give a colourless solution of sodium tetrahydro aluminate.

$$Al(OH)_3 + NaOH \longrightarrow NaAl + (OH)_2$$

Q29: Why atomic Radii increases as we move from Li to caesium?

Answer

As we move from lithium to caesium an extra shell of electrons of an extra shell increases there in an in increases of atomic and ionic radii (of M^+ cons) as we move from lithium to caesium.

Q30: Define ionization energy and why Alkali metals have low ioinizaton energy?

Answer

First ionization energy is the energy needed to remove the most loosely held electron from each of one mole of gaseous atoms to make one mole of singly charged gaseous ions, in other words for 1 mole of this process.

$$^{x}(g) \longrightarrow X^{+} e$$

As you go down the graph ionization energy falls. Alkali metal have only one electron in their outermost shell (ns1 electron). The ns1 electron is so weekly held with the nucleus that it can be removed very easily. Alkali metals therefore have low ionisation energies.

As the distance ns1 electron from the nucleus increases on moving from Li to Cs its removal becomes more and more easy as are proceed from Li to Cs i.e the amount of energy used in the removal of ns1 electron is maximum in case of energies of alkali metals go on decreasing from Li to Cs.

The second ionisation energies are fairly high, since the loss of the second electron from M^+ cation which has a noble gas configuration is quite difficult.

Q31: Define electronegativity. And write down its trend in alkali metals.

Answer

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. In alkali metals to the outer electron of the atom of alkali metals is loosely held with the nucleus and hence it can be easily excited to the higher energy levels even by a small amount of heat energy. During the excitation process the electron absorbs some energy and when this excited electron comes back to its original position, it gives out absorbed energy in the form of light in visible region of the electromagnetic. Since the amount of energy absorbed during the excitatoin process is different in different atoms, different colours are impacted by the atom to the flame. The property of alkali metal to give colouration in the burn flame has been used to detect their presence in salts by a test, known as flame test.

Q32: Why the melting and boiling point of group 1 elements go down the group?

Answer

The melting and boiling points are very low because of the presence of weak inner atomic bonds in the solid state of the alkali metals. These bonds are due to their atomic radii and mainly due to their electronic configuration having a single valance electrons are compared to large number of available vacant orbital.

As the size of the metal atoms increases the repulsion of non bonding electrons

decreases the melting and boiling points of alkali metals when we move from Li to Cs.

Q33: The densities of alkali metal are low, why?

Answer

The densities of alkali metals are low due to large atomic volumes Li, Na, and K are lighter than H_2O. The densities increase with the increases in atomic from Li to Cs indicating the greater atomic with more than compensates for the bigger size of the atoms. K is however lighter than Na which is due to an _____ increase in atomic size of K.

Q34: Write down the reaction of first group alkali metal will oxygen.

Answer

Alkali metal lead with O_2 OI air rapidly and thus trainshed due to the formation of their oxide on the surface of the metals. It is for this reason that alkali metals are stored in kerosene or paraffin oil.

$$4Li + O_2 \longrightarrow Li_2O$$

Na when burnt in O_2 from sodium peroxide Na_2O_2.

$$2Na + O_2 \longrightarrow Na_2O_2$$

Other alkali metal react with O2 to from space oxide of MO_2 type.

$$M + O_2 \longrightarrow MO_2$$

Normal oxides of alkali metals other than that of Li (Li_2O) are not formed by the direct reaction b/w the metals and O2 they formed by indirect methods by producing peroxides, nitrates, nitrate with metals itself.

$$2Na_2O_2 + 4Na \longrightarrow 4Na_2O$$

$$2Na_2O_2 + 6Na \longrightarrow 4Na_2O + N_2$$

$$2Na_2O_2 + 10Na \longrightarrow 6Na_2O + N_2$$

Normal oxide react with H_2O to form hydroxides by proton exchange.

$$O_2^- + H_2O_2 \rightleftharpoons OH^- + OH^-$$

The peroxides (O_2^{2-}) and superoxides (O_2^{2-}) are strong oxidising agent and react with H_2O to give H_2O_2 and O_2.

$$O_2^{2-} + 2H_2O_2 + H_2O_2 + 2H_2O_2$$

$$2[O-O] + 2H_2O \longrightarrow H_2O_2 + O_2$$

Q35: Write down effect of Nitrates, carbonates and hydrogen carbonates.

Answer

Group 1 compounds are non stable to heat than the corresponding compounds in group 2. Li compound behave similarly to group 2 compounds but the rest of group 1 are in same way difference.

==========

Nature of carbonates, bicarbonates and Nitrates

The carbonates (M_2CO_3) and bicarbonates (MHCO) an highly stable to heat with increase of electropositive character from Li to Cs, the ability of these salts increases.

Their nitrates decomposes a strong heating to the corresponding nitrate and O_2.

$$2NaNO_3 \longrightarrow 2NaNO_2 + O_2$$

Q36: Explaining the trend in terms of the polarising ability of the positive ion. Justify.

Answer

When alkali metal cations approach near an anion attracts the outer most electron of the anion and repels the molecules. Thus the distortion of polarisation of the anion talks place. This distortion usual showing of electron L/w two oppositely charged ions the bond between the cations and anion becomes partly covalent in character. In general the smaller cations polarise the anions more effectively than bigger one. Therefore, the lithium salts are slightly covalent which other alkali metal salts are ionic.

Q37: Define flame test? And the property of alkali metals to give also reaction in the Bunsen flame has been used to detect the presence in salts by a test known as flame test? Justify.

Answer

Flame test are used to identify the presence of a relatively small number of metal ions in a compound.

We have seen that outer electron of atom of alkali metals is loosely hold will nucleus and hence it can be easily excited to the higher energy links even by a small amount of heat energy. During the excitation process the electron absorbs some energy and under this excited electron comes back to its original position, it guess out absorbed energy in the form of light in visible region of the electromagnetic spectrum and hence the colour is imparted by the atoms to the flame. Since the amount of energy absorbed to the flame. The property of alkali metals to give coloration in the Bunsen flame has been used to detect their presence in salts by a test known as flame test.

Q38: Alkaline earth metals have higher nuclear charge which tends to draw the orbit electrons towards the nucleus?

Answer

Because of the addition of an extra shell of electrons to each elements from Be to Ra, the atomic volume increases from Be to Ra. With the increase of atomic volume the atomic and ionic radii (of M^{2+} ions) also increases from Be to Ra. The atomic radii of these elements are however smaller than those of alkali metals have higher nuclear charge which tends to draw the electrons towards the nucleus. The smaller values of atomic radii result in that the alkaline earth metals are harder, have higher densities and higher melting points than alkali metals.

Q39: $Mg(OH)_2$ is weakly basic while $Be(OH)_2$ is the amphoteric why?

Answer

$Be(OH)_2$ is not all basic infant it is amphoteric since it react with acids to from sals and with alkalies to give beryllates.

==========

$$Be(OH)_2 + 2HCl \longrightarrow BeCl_2 + 2H_2O$$

$$Be(OH)_2 + 2NaOH \longrightarrow Na_2BeO_2 + 2H_2O$$

The hydroxides or other metals are basic character. Their basic character increase on moving down the group. Thus Mg (OH) is weekly basic while Be $(OH)_2$ Be $(OH)_2$ is the amphoteric. The increase in basic character of the hydroxides on moving down the group is due to the fact that will the increase in size of M^{2+} cation both the polarity of M– OH bond and the internuclear distance between oxygen of OH ion and the metals atom increase. As a result of this there is greater ionisation of M $(OH)_2$ and hence basic character increases.

Due to high polarising power of small Be^{2+} in Be $(OH)_2$ are almost insoluble in H_2O while the hydroxides of other metals are slightly soluble. Their solubility increases on moving down the group as shown by the increasing value of the solubility products of these hydroxides.

Q40: Discuss the effects of heat on the group 2 nitrates.

Answer

All the nitrates in this group undergo thermal decomposition to give the metal oxide nitrogen dioxide and oxygen.

The mixtures are while solids, and the oxides produced are also white solids. Brown nitrogen dioxide gas is given off together with oxygen, Magnesium and calcium nitrates normally have water of crystallization and the solid may dissolve in its over H2O of crystallization to make a colourless Skelton before its starts to decompose.

$$2x\,(NO_3)_2 \longrightarrow 2 \times O + 4NO_2 + O_2$$

As go down the group mixtures also have to be heated more strongly before they will decompose.

Q41: How Beryllium differ from Members of its family?

Answer

Be the first element of the group differs from rest of alkaline earth metals due to its small atomic size and comparatively high electronegativity.

Hardness and Melting and boiling point

Be is hardest by all the element of its group and their melting and boiling point are also high.

Formation of Covalent Bond

Be has a tendency to form covalent compounds. Thus when its react with other elements the electronegativity difference is _____ layer and the bond is therefore covalent.

Reaction with H_2O

Be does not react with H_2O even at high temperature other alkaline earth metal decompose water librating H_2 gas.

$$Mg + H_2O \longrightarrow MgO + H_2$$

Rex with alkalis

Be react with alkalis to form hydrogen

$$Be + 2NaOH \longrightarrow Na_2 BeO_2 + H_2$$

Behaviour of oxides and hydroxides

The oxides and hydroxides of Be are amphoteric dissolve in both acid and alkalis to form salts.

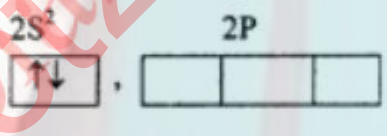
$$Be O + H_2SO_4 \longrightarrow BesO_4 + H_2O$$

Q42: Write down the number of molecules of H_2O of crystallization of Be^{2+}.

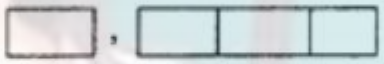
Answer

The salts of Be^{2+} ion cannot have more than four molecules of H_2O of crystallisation while other alkaline earth melts have more than four molecules of H_2O of crystallisation. This is explained as follow. In case of Be^{2+} ion there are only four orbitals (namely one orbitals can accept lone pairs of electrons slenobed by 0 – atoms on each of the H_2O molecules are as.

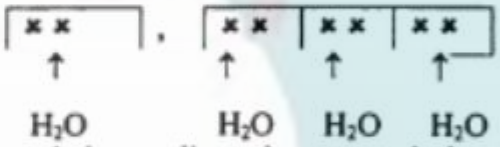
Valance shell electronic configuration of Be atom = $2s^2, 2P^o$



Valance shell configuration of Be^{2+} $(2S^o, 2P^o)$



Attachment of $4H_2O$ molecules with Be^{2+} ion



Other alkaline earth metals like Mg can ____ their coordinate ion no to six by using one their outer most shells.

Q43: The melting and boiling pointing C to Pb are gradually decreases down the group. Why?

Answer

As we move down the group from c to Pb the melting point as well as boiling points generally decreases, although the decrease is not in a regular order. This decrease in melting points as well as in boiling point indicates that inter-atomic forces also decreases in the same direction. The melting and boiling point of c and Si are notably high because of the tendencies of these elements to form giant molecules. The low value for tin's melting point compared with lead is presumably due to forming a distorled 12 co – ordinate structure rather than a pore one.

Q44: Carbon and silicon doesn't conduct electricity.

Answer

Carbon as diamond doesn't conduct electricity. In diamond the electrons are all tightly

bond and not free to move.

Unlike diamond (which doesn't conduct electricity) silicon, germanican and grey tin are semiconductor.

White tin and lead are normal conductors of electricity. There is clear trend from the typically non metallic conductivity behaviour of carbon as diamond, and the typically metallic behaviour of white tin and lead.

Q45: Electronegativity or ionization energy decreases of moving down type group from c to Pb. Why?

Answer

Carbon is the most electronegative element of this sub group and the electronegativities decreases with like of atomic no but not a regular manner. This is probably due to the filling of the d orbitals in case of Ge and S_n and f orbitals in case of Pb.

The ionisation energy values decrease on moving down the group from c to Pb although the decreases does not occur in a regular order. The irregulating in the decreases of those values is due to the filling of interviewing d – orbitals in case of Ge and S_n and f – orbitals in case of Pb which are not able to seen the valance electrons effectively in elements.

Q46: Carbon and silicon show +4 oxidation state while occurrence of +2 and +4 oxidations states in case of Ge, Sn and PB. Explain.

Answer

Carbon and silicon show + 4 oxidation state while occurance of + 2 and +4 oxidation states in case of Ge, Sn and Pb are when only two np electorns from the ns^2P^2 configuration are lost. The element in +2 oxidation states remain infect and hence are not lost in the formation of M^{2+} cations. This pair of ns^2 electrons is called inter pair of electrons. Since the group the stability of +2 oxidation state also increases from Ge^{2+} to Pb^{2+} i.e. Ge2+ < Sn^{2+} <Pb^{2+}.

When all four ns^2P^2 electrons are lost we get the elements in +4 oxidation state i.e. M^{4+} cations are formed. On decreasing the group stability of +4 oxidations state decreases i.e. the stability of M^{4+} cations decreases from Ge^{4+} to Pb^{4+}. i.e. $Ge^{4+} > Sn^{4+} > Pb^{4+}$.

Q47: The dipole moment of CO_2 and explain it?

Answer

The dipole moment of CO_2 is zero. Therefore it is a linear molecule.

$$\longleftarrow \quad \longrightarrow$$
$$O = C = O$$

Q48: Draw and explain the structure of SiO_2.

Answer

It is a macro molecular compounds in which silicon and oxygen atoms are linked together covalently in tetrahedral basic unit.

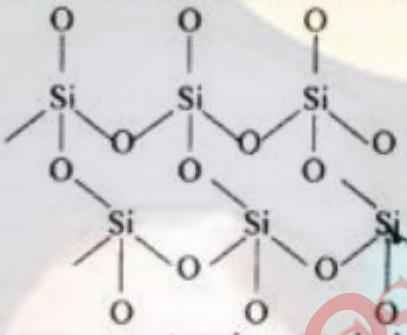
In crystablite, these unit are joined as in diamond, circle in qwartz and tridymite they are arranged spirally around an axis. Because of its structure of SiO_2 is non volatile and hard unlike CO. Triatomic molecules of SiO_2 and CO_2, carbon and silicon are similar

in having.

1) 4 valence electron
2) 4 covalent bond formation

But they show a lot of difference in their physical properties. It is due to the fact that.

1) Silicon atoms are much layer is size than c atoms and thus tend to be surrounded by more oxygen atoms.
2) Silicon from only single bond with oxygen atom while carbon forms double bonds.



Si atom is bound to four oxygen atom in a tetrahedral structure which result in the formation of SiO_2 crystal. The simplest formula for silica is SiO_2.

Q49: Write down the trends of electronegativity of Halogens and also its first electron of finally.

Answer

Halogens have high values of electronegativity in the group. These value decreases as the proceed from F to I in the group. Large electnetivities values of halogen atom indicates that x atoms have a strongly tendency to form x ions. Electronaffinity values decreases from Cl to I.

Q50: Define bond enthalpies and also explain the bond enthalpies in halogens?

Answer

Bond entholpyh is the heat needed to break one mole of a covalent Bond to produce individuals atoms, stating from the original substance in the gas state, and ending with gaseous atoms.

The bond enthalpies of the Cl–Cl, Br – Br and I–I bonds fall just as you would expect, but the F – F bond is deviated from the sequence. This is because duet to very small F–F bond length very large as compared to other x– x bond lengths. This makes the F atoms in F_2 molecules repel each other and helps the dissociation of F_2 molecules into factors.

X – X bonds in Cl_2, Br_2 and I_2 molecules is stronger than F–F bond in F_2 molecule. This is due to the possibility of the existence of multiple bond in x – x bonds involving d – orbitals.

Q51: HF is the weakest acid. Why?

Answer

HF ionises only slightly while HCl, HBr, and HI ionise completely. Hence HF is the weakest acid and strength of these acids increases from HF to HI

i.e. HF (weakest acid) <HCl<HBr <HI.

The weakest acidic nature of HF is due to the fact that the dissociation energy of H–F bonds is HF molecule is the highest and hence this molecule has least tendency to splits up into H^+ and F^- ion in aqueous selection.

Q52: Define open Hearth process.

Answer

A fie furnace is used to remove the impurities of metals. It is lined inside with fire bucks and is just like a room. Burning gases are entered from one side and exhaust gases are removed from the opposite end. The process is operated from opposite ends after an interval. Metals melts in a shorter times by this two way heating.

Q53 :Write down the application of bleaching powder.

Answer

Bleaching powder is actually a mixture of calcium hypochlocite ($Ca(OCl)_2$) and the basic chloride $CaCl_2$, H_2O will some stalked lime $Ca(OH)_2$.

Bleaching powder is used for the disinfection of drinking H_2O or swimming pools H_2O. It is used as a sanitizer in outdoor swimming pool in combination with a cyanuric acid stabilizer which reduces the loos of chlorine due to ultraviolet radiation.

The calcium content hardness the H_2O and tends to clog up some filters hence some produce containing calcium antiscaling agents.

Bleaching powder is used for bleaching cloth and limen. It is also used in bathroom, cleaners, household disinfectant sprays, mors and algae removers and cured/____.

In addition bleaching powder may be used to manufacture chlocoform. Bleaching powder is used also in sugar industry for bleaching sugar cane juice before its crystallization.

Q54: Define Goiter?

Answer

The term goiter refers to the abnormal enlargement of thyroid gland due to deficiency of iodine in diet. It results in swelling in neck. It is important to know that the presence of goiter does not necessarily mean that the thyroid gland is malfunctioning hypothrodism. A goiter can also occur in a gland that is producing too much thyroid hormone (hyperthyroidism) or even the current amount of hormone (euthyroidism) goiter indicates there is a condition present which is carrying the thyroid to grow abnormally.

Q55: Define fluoride deficiency and toxicity?

Answer

Fluoride Toxicity

Fluoride toxicity or fluoride poisoning is a condition in which more fluoride is taken than the amount required for normal growth development and metabolism. Fluoride toxicity is characterized by a variety of signs and symptoms poisoning most commonly occurs following ingestion of conspicuous amount of fluoride entraining products. Symptoms or set usually occurs within minutes of exposure. Fluoride is formed in

many common household products toothpaste, insecticides etc..

Fluoride deficiency

Fluoride deficiency result when the amount of its up taken is less than required. Fluoride deficiency results in Brittle bones or demineralization of bones. Fluoride deficiency can lead to a higher likehood of developing bone fractures and possibly even steoporesis.

We can combine these two half equations to give the overall reaction as follows:

$$H_2SO_4 + 2H^+ + 2BR^{-1} \longrightarrow BR_2 + SO_2 + 2H_2O$$

Q56. What happen when HI is added to conc. H_2SO_4?

Answer

Iodide ions are stronger reducing agents they are oxidized to I_2 to I2 by the conc. H_2SO_4. ($2I^- \longrightarrow I_2 + 2e^-$).

The reduction of the sulphuric acid is more complicated than before. The iodide ions are powerful enough reducing agents to reduce it.

This reduction takes place in following steps:

i) First to sulphur dioxide (oxidalation state of S = +4).

ii) Then to sulphur itself (oxidation state of S = O).

iii) Then formation of H2S (oxidation state of S = – 2).

The most important of this mixture of reduction products is probably the H_2S. The half-reaction for its formation is:

$$H_2SO_4 + 8H^+ + 8e^- \longrightarrow H_2S + 4H_2O$$

On combing these text two half equations

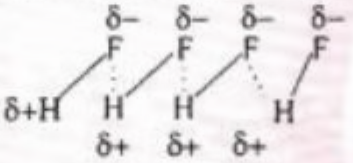
$$H_2SO_4 + 8H^+ + 8I^- \longrightarrow 4I_2 + H_2S + 4H_2O$$

Quick Quiz

Q57. HI is stronger acid then HF why?

Answer

Due to very high electronegativity of F (EN = 4). HF is highly polar molecule. Due to this reason strong hydrogen bonds are formed ammg H–F molecules.

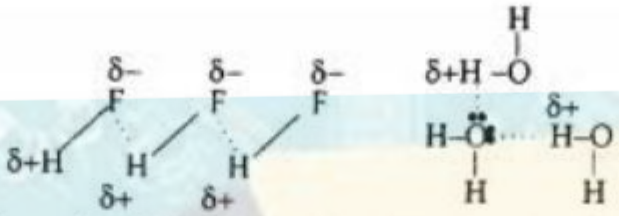


On the other hand $I^{\ominus}$ is a large sized atom which can exist more freely. So HI releases more hydrogen ions. That is why HI is stronger acid then HF.

Q58. Although hydrogen bonding in HF is stranger than in H_2O but H_2O has much higher boiling point then that of H_2O, Why?

Answer

HF molecule may form one hydrogen bond whereas one H2O molecule may form two H-bonds.



That is why boiling point of H_2O (100°C) is higher than that of HF (98°C).

Q59. The acidic character of hydrides of group VII–A elements increases on descending the group,Why?

Answer

Atomic sizes of halogeins increase from F to I in group VII–A as a result stability of halogen ions increases. That's why the ability of halo acid to release H^+ ion increases, from HF to HI.

Q60. Illustrate the oxidizing properties of halogens by giving examples of two typical reactions.

Answer

A substance which as high electronegativity and ability to accept one or more electrons is said to show oxidizing property. Also the electron affinity values of halogens are high. But the electron affinity decrease from F to I. So the oxidizing power of halogens, also decrease from F to I.

Upper halogen in group VII–A may displace lower halogen that is F_2 may displace lower halogen that is F_2 may displace Cl^- ions then Cl_2 may displace Br^{-1} ions from their solution.

$$F_2 + 2Cl^{-1} \longrightarrow 2F^{-1} + Cl_2$$

$$Cl_2 + Br^{-1} \longrightarrow 2Cl - Br_2$$

Q61. Give reason

i) F is better oxidizing agent then Cl

ii) Electronegativity of halogens decreases in the order F>Cl>Br>I

Answer (i)

F has the highest electronegativity (EN = 4) whereas Cl has elecronegativity (EN = 3.2). Therefore F may attract the electrons with more force and may order any substance more than that of Cl.

Answer (ii)

Atomic sizes of halogens increase from F to I due to increase in number of shells from F to I As a result pull of nucleus to attract the shared pair of electrons decreases from F to I.

	F	>	Cl	>	Br	>	I
	F	>	Cl	>	Br	>	I
E.N = Value =	4.0		3.2		3.0		2.5

Q62. Give different steps to obtain iron from its ores.

Answer

Iron has the following ores:

Manetite	Fe_3O_4
Haematite	Fe_2O_3
Limonite	$Fe_2O_3.3H_2O$

Iron is mostly obtained from magnetite and hematite.

Steps to extract metal from their ores:

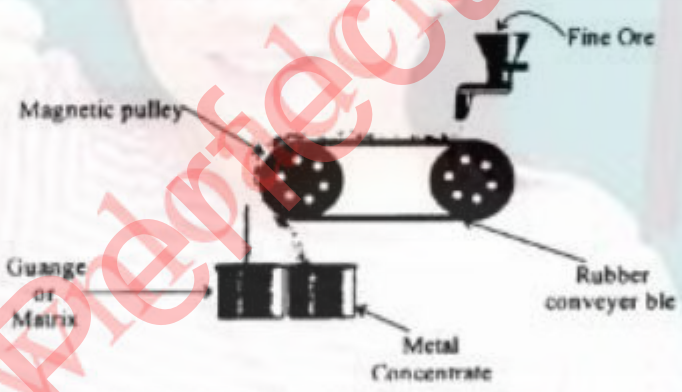
i) **Removal of gangue or matrix:**
In this step useless material from ores is removed manually or mechanically.

ii) **Crushing**
In this step large pieces are broken down in to small pieces with the crushers.

iii) **Grinding**
Teema Ball mill is oftenly used to grind the ores without contamination.

iv) **Concentration (Magnetic concentration of Iron)**
Different methods are used to concentrate the ore before sending to furnace. In case of iron magnetic separation method is used.

In this method powdered iron ore is passed over a magnetic pulley using rubber conve belts.



VI) **Refining of metals (Open hearth Process)**
In this process metal is heated is a fire bricks linening furnace. Hot air is through the molten metal which oxidized paned and remove the impurities on the top of molten metal or skimed off.

Q63. What is bleaching powder? Give its common applications.

Answer

Bleaching powder is a mixture of calcium hypo chloride and calcium chloride $Ca(OCl)_2$ $CaCl_2$, with some slaked lime ($Ca(OH)_2$).

Application of bleaching powder

i) **Removal of stains**
Bleaching powder is used to remove stain from clothe and different materials to climinate mildew and other stubborn stains.

ii) To Disinfect and sterilize the baby toys and other home applications.

iii) Removal of weeds, moss and algae from gardens and walk ways.

iv) Bleaching powder is used to keep freshness of cut flowers by adding it to the solutions in which flowers are placed as it kills bacteria.

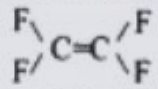
v) Disinfection of drinking water and swimming pool water.

vi) Bleaching of sugar in sugar industry.

Q64. Give different uses of halogens.

Answer

i) Manufacturing Teflon: (Polytetra flowroethlene) i.e.



Teflon is used to store HF and other acids which corrode metals.

Flourine isused in extraction of uranium and their isotope. Separation UF_6 is manufactured to separate ^{238}U and ^{235}U.

$$UF_{4(s)} + F_{2(g)} \longrightarrow UF_{6(s)}$$

Q65. Write notes an toxicity of fluorides and deficiency of fluorides.

Answer

Flouride Toxicity

Flouride toxicity or fluoride poisoning is a condition is which more fluoride is takes than the amount required for normal growth, development an metabolism. Flouride toxicity is characterized by a variety of signs and symptoms. Poisoning most commonly occurs following ingestion of conspicuous amount of flouride entraining products symptoms or set usually occurs within mixtures of exposure. Flouride is found in many common household products e.g. toothpaste, dietary supplementary insecticides rodenticide etc. fluoride toxicity results.

1) Arthrit
2) Stiff painful joints with or wistful swelling.
3) Asthma, especially after showering.
4) Painful bony lumps where tendons and ligaments attach to bones.

Q65. What problems are caused in human body by deficiency of Iodine?

Answer

As human body cannot synthesize iodine so it is an essential element. The deficiency of iodine in human body leads to a disease goiter.

Goiter

The term goiter refers to abnormal enlargement of thyroid glands due to deficiency of iodine in diet. It results in swelling in neck. The presence of goiter does not means that the thyroid gland is malfunctioning (hypothyroidism): A goiter can also occur in a gland that is producing too much thyroid ramones (hyperthyroidism).

I–A	II–A	III–A	IV–A
+1	+2	+3, 5	2+,4+,6

Trends within a Group

The oxidation state of group I–A and II–A elements is fixed. But in P–block when we move from top to bottom the atomic size increases as a result the ability of atoms to loose electrons also increases.

For example in group IV–A

C and Si show +4 oxidatin state. In case of Ge, Sn and Tb only two nP electrons from the ns^2P^2 configuration are lost. We get the elements in +2 oxidatin sates remain inert and hence are not lost in the formation of M^{2+} cations. This pair of ns^2 electrons is called inert pair of electrons. Since the group the stability of +2 oxidation state also increases from Ge^{2+} to Pb^{2+} i.e $Ge^{2+} < Sn^{2+} < Pb^{2+}$.

When all the foru ns^2P^2 electrons are lost we get the element in +4 oxidation state, i.e. M^{4+} cations are formed. On descending the group stability of +4 oxidation state decreases i.e the stability of M4+ cations decreases from Ge^{4+} to Pb^{4+} i.e. $Ge^{4+} > Sn^{4+} > Pb^{4+}$.

Compounds of Ge^{2+} are less stable than those of Ge^{4+} and hence the compounds of Ge^{2+} are readily charged oxidised into those of Ge^{4+} act as strong reducing agents while those of Ge^{4+} act as oxidising agents.

Ge^{2+} Compounds ⟶ Ge^{4+} compounds

(Less stable) (More stable)

Reducing agent oxidising agent

On similar grounds it can be shown that the compounds of Sn^{2+} are less stable than those of Sn^{2+} into those of Sn^{4+}. In other words, compounds of Sn^{2+} act as strong reducing agents while those of Sn^{4+} act as oxidising agents.

Sn^{2+} Compounds ⟶ Sn^{4+} compounds

(Less stable) (More stable)

Reducing agent oxidising agent

When we compare the stability of the compounds of Pb^{2+} and Pb^{4+} ions, we find that Pb^{2+} than those of Pb^{4+} ($PbCl_4$) and hence the compounds of Pb^{4+} are readily charged (reduced) into those of Pb^{2+}. In other words compounds of Pb^{4+} act as strong oxidising agents while those of Pb^{2+} act as reducing agents.

Pb^{2+} Compounds ⟶ Pb^{4+} compounds

(Less stable) (More stable)

Reducing agent oxidising agent