NEET Revision Notes
Chemistry
The p-block Elements

p-block elements

- p-block elements include elements of groups 13, 14, 15, 16, 17, and 18.
- These elements are called p-block elements because their valence shell is p-orbital which means their last electron is added to a p-orbital.

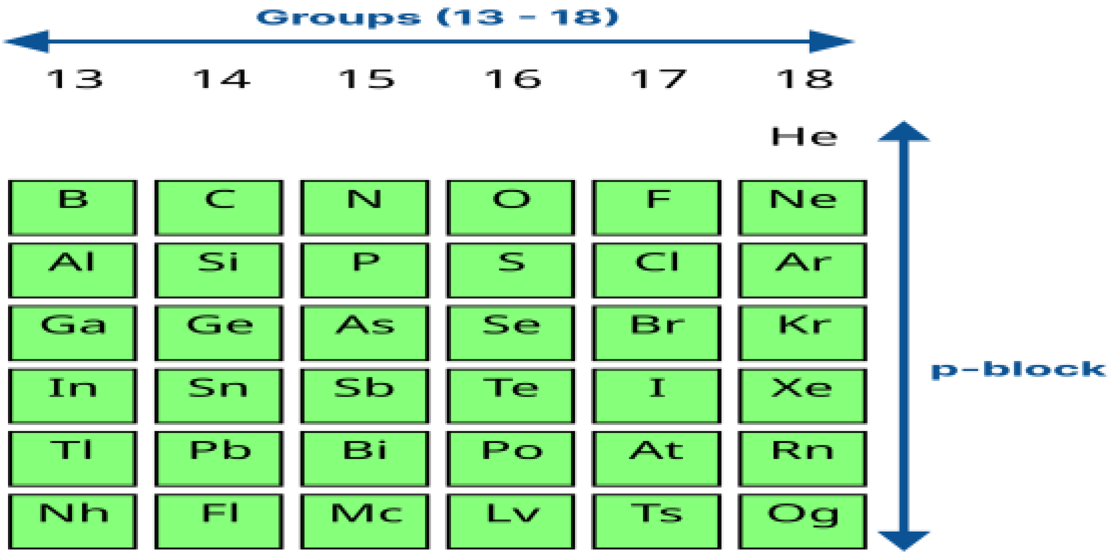


Image: p-block elements

Group 13

- This group is known as the Boron family.
- The members of group 13 are Boron (B), Aluminum (Al), Gallium (Ga), Indium (In), and Thallium (Tl).
- **General configuration:** $ns^2\ np^1$
- **Order of atomic radii:** B < Al> Ga < In < Tl.
  In general, size is increasing on moving down the group due to the addition of an extra shell but Gallium is an exception. Gallium is smaller than Aluminum due to the poor shielding effects of d electrons in Gallium.
- **General oxidation state:** +3
- **Trend of ionization energy:** From B to Al, it decreases due to an increase in size but then from Al to Ga, it increases slightly due to poor shielding of d-electrons. From Ga to In, it decreases slightly and then again increases for Tl due to poor shielding of f-electrons.

- **Trend of electronegativity:** Decreases from B to Al and then increases from Al to Tl. This trend is again due to the poor shielding of d and f electrons.
- **Metallic character:** Boron has some character of metalloids. Aluminum is most metallic and then metallic character is also present Ga to Tl.
- **Reactivity of Boron family with oxygen:** These elements react with oxygen at higher temperatures to form trioxides, M_2O_3.

  $4Al + 3O_2 \rightarrow 2Al_2O_3$

  $4B + 3O_2 \rightarrow 2B_2O_3$
- **Reactivity with water:** At high temperature only, boron reacts with steam to form boron oxide.

  $2B + 3H_2O \rightarrow B_2O_3 + H_2$
- Aluminum reacts with cold water and liberates hydrogen gas. Ga and In do not react with water.

  Tl form TlOH in moist air- $4Tl + 2H_2O + O_2 \rightarrow 4TlOH$
- **Reactivity towards metals:** Only Boron reacts with the metal to form borides (M_3B_2).

  Example: $3Mg + 2B \rightarrow Mg_3B_2$
- **Reaction with acid and alkalis**: Boron reacts with hot and concentrated nitric acid to form boric acid and nitrogen dioxide.

  $B(s) + 3HNO_3(aq) \rightarrow H_3BO_3(aq) + 3NO_2(g)$
- Other elements of this group react with acids to produce hydrogen gas.
- At temperatures above 773 K, Boron reacts with alkalis to form borates.

  $2B(s) + 6KOH(aq) \rightarrow 2K_3BO_3(aq) + 3H_2(g)$

Important compounds of Boron:

Orthoboric acid (H_3BO_3):

- It is a weak and monobasic acid of boron.
- Structure: Trigonal structure

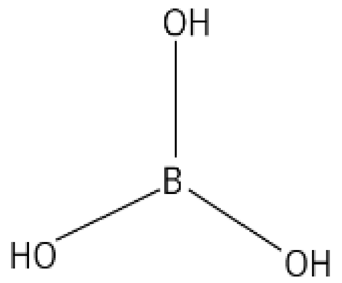


Image: Orthoboric acid

Preparation of Orthoboric acid:

- From borax: By heating a concentrated solution of borax with sulphuric acid or hydrochloric acid.

  $Na_2B_4O_7.10H_2O + 2HCl \rightarrow 4H_3BO_3 + 2NaCl + 5H_2O$
- By hydrolysis of diborane: $B_2H_6 + 6H_2O \rightarrow 2B(OH)_3 + 6H_2$
- By hydrolysis of borane trihalides: $BX_3 + 3H_2O \rightarrow B(OH)_3 + 3HX$

Properties of Orthoboric acid

- Action of Heat:

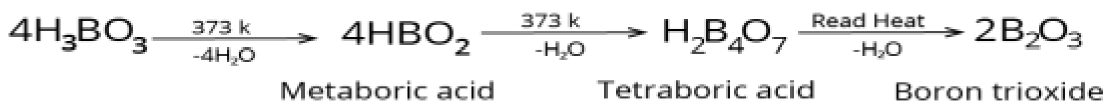


Image: Action of heat

- Reaction with Metal oxide:

Image: Reaction with metal oxide

- Reaction with Ammonium borofluoride:

Image: Reaction with ammonium bifloride

Borax (sodium tetraborate)

- Formula $Na_2B_4O_7.10H_2O$
- Preparation from Boric Acid:

  $4H_3BO_3 + Na_2CO_3 \rightarrow Na_2B_4O_7 + 6H_2O + CO_2$

Properties of Borax

- Basic Nature: The aqueous solution of borax is alkaline in nature.

  $Na_2B_4O_7 + 3H_2O \rightarrow NaBO_2 + 3H_3BO_3NaBO_2 + 2H_2O \rightarrow NaOH + H_3BO_3$
- Action of heat:

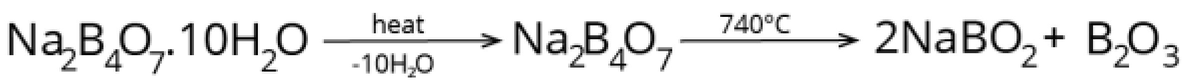


Image: Action of heat

Diborane

- Chemical formula: B_2H_6
- Structure:

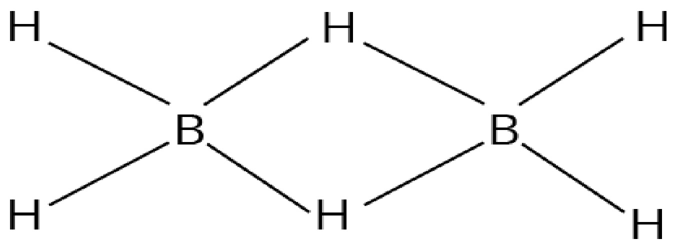


Image: B-H-B bond

Each B-H-B bond has only 2 electrons.

Preparation of Diborane:

- Reduction of Boron Trifluoride: $BF_3 + 3LiAlH_4 \rightarrow 2B_2H_6 + 3LiAlF_4$
- From $NaBH_4$:

$$2NaBH_4 + H_2SO_4 \rightarrow B_2H_6 + 2H_2 + Na_2SO_4$$

$$2NaBH_4 + H_3PO_4 \rightarrow B_2H_6 + 2H_2 + NaH_2PO_4$$

Properties of Diborane:

- Reaction with water: Boric acid is produced.

$$B_2H_6 + H_2O \rightarrow 2H_3BO_3 + 6H_2$$

- Combustion: Boric oxide is produced.

$$B_2H_6 + 3O_2 \rightarrow B_2O_3 + 3H_2O \Delta H = -2615\,kJ/mol$$

Important compounds of Aluminium:

Aluminium Oxide or Alumina

- Chemical formula :Al_2O_3
- Preparation: Aluminium oxide is produced by heating aluminium hydroxide or aluminium sulphate.

$$2Al(OH)_3 + \text{ Heat } \rightarrow Al_2O_3 + 3H_2O$$

- This oxide is amphoteric in nature.

Aluminium Chloride

1. Chemical formula :$AlCl_3$
2. Structure of Aluminium Chloride:

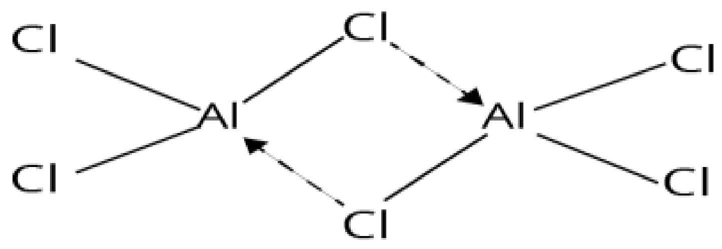


Image: Aluminium chloride

3. Properties of Aluminium Chloride
    - It exists in dimer form.
    - White, hygroscopic solid. Thus, it absorbs moisture from air.
    - There are weaker intermolecular forces due to which it sublimes at 183 ^{0}C.
    - Hydrolysis: $AlCl_3 + 3H_2O \rightarrow Al(OH)_3 + 3HCl + 3H_2O$
    - Action of Heat: $2AlCl_3.6H_2O \rightarrow 2Al(OH)_3 + Al_2O_3 + 6HCl + 3H_2O$

Group 14

- This group is known as Carbon family.
- The members of group 14 are Carbon (C), Silicon (Si), Germanium (Ge), Tin (Sn), and Lead (Pb).
- **General configuration:** $ns^2\ np^2$
- Valency of this group elements is 4.
- General oxidation states are +4 and +2.
- Carbon shows different properties as compared to the rest of the elements because of its small size and its catenation property.
- Size of the atoms increases on going down the group but after Si, there is only a slight increase in the size due to the poor shielding effect of d-electrons.
- Trend in ionization energy: C>Si>Ge>Pb> Sn.
- The size of Pb and Sn are comparable, so due to higher charge density, it is more difficult to ionize lead.
- Down the group, metallic character increases. Carbon and Silicon are non-metals, Germanium is a metalloid, and Tin and Lead are metals.
- Catenation is the property of carbon atoms to bond chains with other carbon atoms.

Important compounds of Group 14

Carbon monoxide:

- Chemical formula: CO
- Structure: $:C=O:$
- CO is a very poisonous gas due to the formation of carboxyhemoglobin in the blood.

Preparation of CO:

- On heating the mixture of powdered zinc and Calcium carbonate (Lab method): $Zn + CaCO_3 \rightarrow ZnO + CaO + CO$
- By dehydrating methanoic acid in presence of conc. H_2SO_4: $HCOOH \xrightarrow{H_2SO_4} CO + H_2O$
- In industries, carbon monoxide is generated by passing air over red-hot coke. $2C + O_2 \rightarrow 2CO + heat$
- In industries, it is also produced by the reaction of steam and carbon. $2C + H_2O + heat \rightarrow 2CO + H_2$

Chemical properties of CO:

- CO behaves as string reducing agent and reduces metal oxides.
  $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
  $CuO + CO \rightarrow Cu + CO_2$
- CO reacts with nickel and forms tetracarbonyl nickel.
  $Ni + 4CO \rightarrow Ni(CO)_4$
- CO reacts with water vapors at high temperatures to produce carbon dioxide.
  $CO + H_2O(g) \rightarrow CO_2 + H_2$

Carbon dioxide

- Chemical formula -CO_2
- Structure:

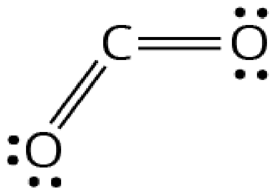


Image: Carbon dioxide

- It is a greenhouse gas, and thus causes global warming.

Preparation of Carbon dioxide

- By action of acids on carbonates: $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$
- By combustion of carbon: $C + O_2 \rightarrow CO_2$

Properties of Carbon dioxide:

- When carbon dioxide is passed to lime water, it turns milky due to the formation of calcium carbonate. $Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$
- The above milkiness disappears in presence of excess CO_2. $CaCO_3 + H_2O + CO_2 \rightarrow Ca(HCO_3)_2$
- It behaves as acid and reacts with base to form carbonates and bicarbonates.
  $CO_2 + NaOH \rightarrow NaHCO_3$
  $NaHCO_3 + NaOH \rightarrow Na_2CO_3 + H_2O$

Compounds of Silicon:
Sodium silicate (Na_2SiO_3):

- Prepared by melting soda ash in pure sand at high temperature:
  $Na_2CO_3 + SiO_3 \rightarrow Na_2SiO_3 + CO_2$
- **Silicon**: Hydrolysis of alkyl or aryl substituted chlorosilane forms silicon polymer with Si-O-Si bonds.
- **Silicate:** The structure of the silicate is SiO_4^{4-}, in which four oxygen atoms are bonded to one silicon atom. Assembling the silicate unit forms the ring, chain, and 3D structure. Glass and cement are two important silicates made by humans**.**

Group 15

- This group is known as Nitrogen family.
- The members of group 15 are nitrogen (N), phosphorus (P), arsenic (As), antimony(Sb) and bismuth (Bi).
- **General configuration:** $ns^2\ np^3$
- General oxidation states are -3 to +5. But +3 is most stable.
- Size of the atoms increases on going down the group.
- Trend in ionization energy: N>P>As>Sb> Bi.
- Down the group, metallic character increases.
- N is a diatomic gas while others are solids in nature.
- Allotropy: Except Bismuth, all other group 15 elements show allotropy.
- Phosphorus is present in many allotropic structures. Two of these important allotrope structures are red phosphorus and white phosphorus. Arsenic exists in three important allotrope structures: black, grey, and

yellow. Antimony also has three allotropes: yellow, metallic, and explosive.

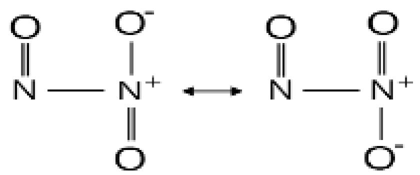
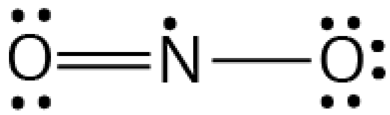
Nitrogen

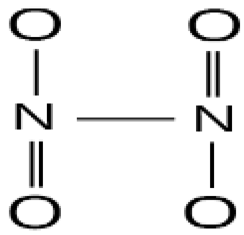
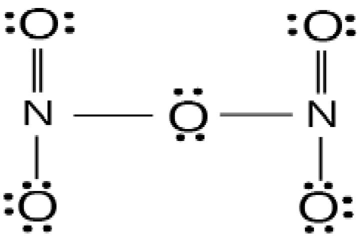
Preparation:

- Laboratory Method: In the laboratory, N_2 is produced by heating an aqueous solution of ammonium chloride and sodium nitrite.
  $NH_4Cl\ (aq) + NaNO_2\ (aq) \rightarrow NaCl\ (aq) + 2H_2O\ (l) + N_2\ (g)$
- By heating the red crystals of ammonium dichromate:
  $(NH_4)_2Cr_2O_7 \rightarrow N_2 + 4H_2O + Cr_2O_3$
- Oxidation of ammonia: Nitrogen is produced when ammonia is oxidized by red hot copper oxide or chlorine.
  $2NH_3 + 3CuO \rightarrow N_2 + 3H_2O + 3Cu$
  $8NH_3 + 3Cl_2 \rightarrow N_2 + 6NH_4Cl$
- Very pure nitrogen can be obtained by heating sodium azide.
  $NaN_3 \rightarrow 2Na + 3N_2$

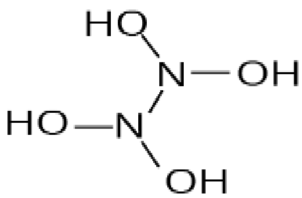
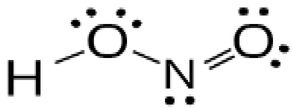
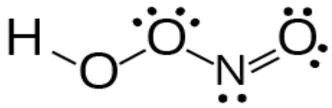
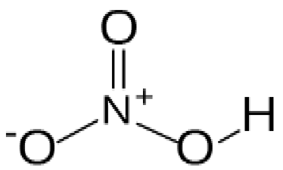
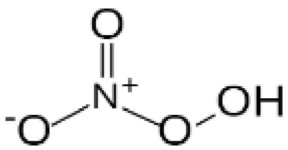
Chemical properties:

- N_2 binds to some strong electropositive metals at high temperatures to form their nitrides.
  $6Li + N_2 \rightarrow 2Li_3N_2$
  $3Mg + N_2 \rightarrow Mg_3N_2$
  $2Al + N_2 \rightarrow 2AlN$
- N_2 combines with O_2 at temperature above 3273 K to form nitric oxide.
  $N_2 + O_2 \rightarrow 2NO$
- Oxides of nitrogen:

Nitric oxide	NO	
Dinitrogen trioxide	N_2O_3	
Nitrogen dioxide	NO_2	

Nitrogen tetraoxide	N_2O_4	
Nitrogen pentaoxide	N_2O_5	

Oxyacids of nitrogen

Hyponitrous acid	$H_2N_2O_2$	HO—N=N—OH
Hydronitrous acid	HN_2O_2	
Nitrous acid	HNO_2	
Pernitrous acid	HOONO	
Nitric acid	HNO_3	
Pernitric acid	HNO_4	

Ammonia:

- Chemical formula- NH_3
- Structure:

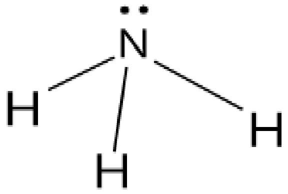


Image: Ammonia

Preparation:

- Ammonium salt is heated with a strong alkali.
  $NH_4Cl + NaOH \rightarrow NH_3 + NaCl + H_2O$
- By hydrolysis of magnesium nitride: $Mg_3N_2 + 6H_2O \rightarrow 3Mg(OH)_2 + 2NH_3$
- Haber's process: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Chemical properties:

Basic nature: Ammonia is basic in nature.

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

Reaction with halogens :

$8NH_3 + 3Cl_2 \rightarrow 6NH_4Cl + N_2$

$NH_3 + 3Cl_2(\text{in excess}) \rightarrow NCl_3 + 3HCl$

$8NH_3 + 3Br_2 \rightarrow 6NH_4Br + N_2$

$NH_3 + 3Br_2(\text{in excess}) \rightarrow NBr_3 + 3HBr$

$2NH_3 + 3I_2 \rightarrow NH_3.NI_3 + 3HI$

$8NH_3.NI_3 \rightarrow 6NH_4I + 9I_2 + 6N_2$

Complex formation:

$Ag^+ + NH_3 \rightarrow [Ag(NH_3)]^{2+}$

$Cu^{2+} + 4NH_3 \rightarrow [Cu(NH_3)4]^{2+}$

$Cd^{2+} + 4NH3 \rightarrow [Cd(NH_3)4]^{2+}$

Precipitation of heavy metal ions from the aq. solution of their salts:

$FeCl_3 + 3NH_4OH \rightarrow Fe(OH)_3 + 3NH_4Cl$
Brown ppt.

$AlCl_3 + 3NH_4OH \rightarrow Al(OH)_3 + 3NH_4Cl$
White ppt.

$CrCl_3 + 3NH_4OH \rightarrow Cr(OH)_3 + 3NH_4Cl$
Green ppt.

Group 16

- Group 16 is known as Oxygen Family
- Group 16 elements: Oxygen (O), Sulfur (S), Selenium (Se), Tellurium (Te), and Polonium (Po)
- General electronic configuration: ns^2np^4
- Together, all these elements are also called chalcogens.

General characteristics of the oxygen family

- Atoms and ionic radii: As the number of shells increases, the atom and ionic radii increase from top to bottom within the group.
- Ionization enthalpy: Due to the expansion of atoms, the ionization enthalp y decreases within the group.
- Electron gain enthalpy: Due to the compact nature of oxygen, the electron gain enthalpy is lower than that of sulfur. After sulfur, the electron gain enthalpy is reduced within the group.
- Electronegativity: Electronegativity decreases within the group. This means that metallic properties increase in the oxygen-to-polonium group.
- Oxidized state: Elements in group 16 show oxidized state -2, +2, +4, +6. -2.

Reactivity to Hydrogen: All 16 elements in the group form hydrides H_2E (E = S, Se, Te, Po).

Thermal Stability: The thermal stability of the hydrides of Group 16 elements is- $H_2O > H_2S > H_2Se > H_2Te > H_2Po$

Acidity: The acidic character of hydrides of group 16 elements increases down the group. $H_2O < H_2S < H_2Se < H_2Te$.

Reducing character: The reducing character also decreases down the group due to the decreasing bond dissociation enthalpy. $H_2O < H_2S < H_2Se < H_2Te < H_2Po$

GROUP 17

Introduction:

Halogens are highly reactive nonmetals. These elements are very similar in their properties. Group 17 elements are collectively called halogens (in Greek: halo means salt and gen means to produce, so together they produce salt) and consist of fluorine (F), chlorine (Cl), bromine (Br), iodine (I) and astatine (As).

- Similarity to this extent is not found in other periodic table groups. They have a regular gradation in physical and chemical properties.

- Astatine is the only radioactive element in the group. They have seven electrons in their outermost shell (ns^2np^5) and are missing one electron from the nearest noble gas configuration.
- An element's chemical properties and reactivity are determined by the oxidation state they exhibit.

Oxidation state:

- All halogen group elements show an oxidation state of -1. However, elements such as chlorine, bromine, and iodine also exhibit +1, +3, +5, and +7 states.
- This higher oxidation state of chlorine, bromine, and iodine is realized when these halogens are combined with small and highly electronegative atoms of fluorine and oxygen.
- The oxides and oxoacids of chlorine and bromine have the +4 and +6 states. There are no valence d orbitals in the fluorine atom, so it cannot expand its octet.
- Fluorine, the most electronegative element, only has an oxidation state of -1.

Trends in periodic table:

1) Ionic and atomic radii
The nuclear and atomic radii of these elements increase steadily as we move down the group. This is done due to the addition of another level of energy. They have minimal atomic radii compared to other elements in related periods. This can be attributed to the fact that their atomic charge is relatively strong.
2) Ionization enthalpy
These elements have a higher ionization enthalpy. This value continues to decrease as we move down the group. This happens because of the increase in kernel size. However, it is interesting to note that fluorine, due to its tiny size, has the highest ionization enthalpy of any other halogen!
3) Enthalpy of electron gain
The electron gain enthalpy of these elements becomes less negative as you move down the group. Fluorine has a lower enthalpy than chlorine. We can attribute this to the small size and smaller 2p subshell of the fluorine atom.
4) Electro-negativity

Halogens show high values of electronegativity. However, it slowly decreases as you move down the group from fluorine to iodine. This can be attributed to the increase in nuclear radii as one moves down the group.

Physical properties

- Physical state: Group 17 elements are found in various physical states. For example, fluorine and chlorine are gases. On the other hand, bromine is a liquid, and iodine is a solid.
- Color: These elements have different colours. For example, while fluorine has a pale yellow colour, iodine has a deep purple colour.
- Solubility: Fluorine and chlorine are soluble in water. On the other hand, bromine and iodine are much less soluble in water.
- Melting and boiling points: These elements' melting and boiling points increase as we move up the group from fluorine to iodine. Therefore, fluorine has the lowest boiling and melting points.

Chemical properties

1) Oxidizing power
All halogens are excellent oxidizing agents. Of the list, fluorine is the most effective oxidizing agent. It is able to oxidize all halide particles to halogen. The oxidizing power decreases as we move down the group. Halide particles also act as reducing agents. However, their reduction capacity also decreases in the group.
2) Reaction with hydrogen
All halogens react with hydrogen to produce acidic hydrogen halides. The acidity of these hydrogen halides increases from HF to HI. Fluorine reacts violently and chlorine requires sunlight. On the other hand, bromine reacts when heated and iodine needs a catalyst.
3) Reaction with oxygen
Halogens react with oxygen to form oxides. However, it was found that the oxides are not permanent. In addition to oxides, halogens also form a number of halogen oxoacids and oxoanions.
4) Reaction with metals

- Because halogens are very reactive, they react with most metals immediately to form the resulting metal halides. For example, sodium reacts with chlorine gas to form sodium chloride.

- This process is exothermic and emits a bright yellow light and large amounts of heat energy.
- Metal halides are ionic in nature. This is due to the highly electronegative nature of halogens and the high electropositivity of metals.
- This ionic character of the halides is reduced from fluorine to iodine.

GROUP 18:

Introduction:

Group 18 elements include helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). They are referred to as rare gases or inert gases. This means that these elements are chemically inert and do not participate in any reaction.

The general valence shell configuration is ns^2np^6. The exception is helium. It has a $1s^2$ configuration. Because they already have the octet configuration in their valence shells, they are completely chemically inert. They all have a valency of zero.

Trends across Periodic table:

- Atomic Radius: The radius of nuclei increases moving down a group with an increasing number of protons. This is a consequence of the expansion of the additional level at each progressive element as it moves down the group.
- Electron Gain Enthalpy: Group 18 elements exhibit very stable electron configurations. They do not tend to accept electrons.
- Ionization Potential: They have a high ionization potential due to their closed electronic configurations. This value decreases as you move down the group due to the expansion of core size.

Physical properties

- Due to their stable nature, we find these elements as monatomic gases in the free state.
- They are colorless, tasteless and odorless gases. Particles of these elements have weak Van der Waals forces. This power increases as you move down the group. This is due to the expansion of the polarization capacity of the molecules.
- They exhibit low melting and boiling points. We can attribute this to weak Van der Waals forces. Melting and boiling points increase as we move down the group.

- We can condense these elements at extremely low temperatures. As the size of atoms in a group increases, the ease of liquefaction also increases.

Chemical properties

- These elements are chemically inert due to their stable electronic configuration.
- In 1962, Neil Bartlett hypothesized that xenon should react with platinum hexafluoride. He was the first to create a compound of xenon, called xenon hexafluoroplatinate (V). Later, many xenon compounds were integrated, including fluorides, oxyfluorides, and oxides.
- The ionization enthalpies of helium, argon, and neon are too high to form compounds.
- Krypton only forms krypton difluoride because its ionization enthalpy is slightly higher than that of xenon.
- Although radon has a lower ionization enthalpy than xenon, it forms only a few compounds, such as radon difluoride, and a few complexes because radon has no stable isotopes. In any case, xenon forms a particularly significant number of compounds.