

Chapter

5

GROUP VA AND VIA ELEMENTS

HALOGENS

The halogens, VIIA group of the periodic table consists of fluorine, chlorine, bromine, iodine and astatine. All these elements are **non-metals** quite similar in chemical properties. First four elements are more common halogens and Astatine is radioactive and its most stable isotope has half life of only 8.3 hours. Chemistry of astatine is not discussed here.

Table – Electronic configuration of halogens

Name	Symbol	At. Number	Electronic configuration
Fluorine	F	9	$1s^2, 2s^2, 2p^5$
Chlorine	Cl	17	$1s^2, 2s^2, 2p^6, 3s^2, 3p^5$
Bromine	Br	35	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^5$
Iodine	I	53	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^5$
Astatine	At	85	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 5p^5$

NAME OF HALOGENS

Origin of Halogens:

With the exception of fluorine, names of halogens are derived from Greek (Gk) words.

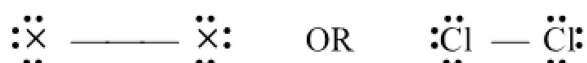
Name	Origin	Meaning
Halogen	Hals (Gk)	Salt
	Gen (Gk)	Producing
Flourine	Fluere (Land F)	Flow or flux (until AD 1500, the mineral fluorspar, CaF_2 was used as flux in metallurgy)
Chlorine	Chloros (Gk)	Greenish yellow
Bromine	Bromos (Gk)	Stench (Offensive odour, ill smell)
Iodine	Iodes (Gk)	Violet
Astatine	Astatos (Gk)	Unstable

L = Latin, F = French, Gk = Greek

PHYSICAL PROPERTIES OF HALOGENS

Some common properties of halogens are given below:

- (1) The electronic configuration of halogens is ns^2, np^5 one electron short from the octet.
- (2) In the free elemental state, all halogens exist in diatomic molecules in which they complete octets by sharing their single unpaired electrons, e.g., F_2 , Cl_2 , Br_2 , I_2 .



- (3) At room temperature and pressure, fluorine is pale yellow gas, chlorine is greenish yellow gas, bromine is red-brown coloured liquid and iodine is a metallic-appearing shiny grayish black solid. Bromine is the only non-metallic liquid.
- (4) The halogens have irritating odours, and they attack the skin. Bromine is particularly nasty, it causes burns that heal slowly.
- (5) Bromine and iodine have very high vapour pressure and even at room temperature, dark-red vapours are always present above the liquid bromine and violet vapours above the solid iodine.
- (6) Halogens have high value of ionization potential due to their smaller size. However ionization potential decreases downward in a group, due to increase in the size of atom.
- (7) Electron affinity values of halogens are large and negative, outer most shell of halogen can absorb one electron very easily.
- (8) They have high value of standard reduction potentials and can easily be reduced.

- (9) Halogens have high electronegativities.

$$F = 4.0, Cl = 3.0, Br = 2.8 \text{ and } I = 2.5$$

- (10) The intermolecular attraction is greater in the larger molecules due to greater masses and greater polarizability. Due to their large size, the Van Der Waal's forces in iodine molecules are stronger than the smaller molecules and it exists in solid state.
- (11) None of the halogens is found in free form. Most of the halogens exist in the form of negative ions due to their greater electronegativities.

Table – Physical properties of group VIIA

Properties	Flourine	Chlorine	Bromine	Iodine
Atomic number	9	17	35	53
Ionic radii (pm)	136	181	196	216
Covalent radii (pm)	72	99	114	133
Melting point (°C)	-220	-101	-7	+114
Boiling point (°C)	-188	-35	+59	+184
First ionization energy (kJ/mol)	1681	1251	1140	1008
Electron affinity, (kJ/mol)	-322	-349	-325	-295
Electronegativity	4.0	3.0	2.8	2.5
Bond energy (x-x) (kJ/mol)	154.80	242.67	192.46	150.6
Oxidation states	-I	-I, I, III, V, VII	-I, I, III, V, VII	-I, I, III, V, VII
Physical form at	Pale	Greenish	Red-brown	Shiny
Room temperature	Yellow gas	Yellow gas	Liquid	Greyish black solid

- (12) Bond dissociation energy (to break X-X bond) decreases from Cl_2 to I_2 due increase in size of atoms. The bond dissociation energy of F_2 is less than Cl_2 due to large repulsive forces between the electrons in the small, highly electronegative fluorine atom.

OCCURRENCE

The halogens exist in nature primarily in the form of compounds. Their most common state is the halide ions F^- , Cl^- , Br^- and I^- . These halides are soluble in water and are found in sea, in salt lakes and underground beds of salts.

Element	Mineral	Formula
Flourine	Fluorspar Cryolite Apatite	CaF_2 Na_3AlF_6 $[\text{CaF}_2.3\text{Ca}_3(\text{PO}_4)_2]$
Chlorine	Halite (Salt beds, brine wells, sea water) Carnallite	NaCl $\text{KCl.MgCl}_2.6\text{H}_2\text{O}$
Bromine	Brine wells, sea water	NaBr , KBr , MgBr_2
Iodine	Deposits in chile brine wells	NaIO_3 , NaIO_4

PECULIAR BEHAVIOUR OF FLUORINE

Fluorine has some difference in properties due to the following reasons:

- (i) Small size of F atom and F^- ion.
- (ii) High first ionization energy and electronegativity.
- (iii) Low dissociation energy of F_2 as compared to Cl_2 and Br_2 .
- (iv) Restriction of valence shell to an octet.

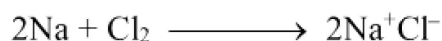
Flourine has following peculiar properties:

- (1) Bond of fluorine with other elements is shorter and stronger due to better overlap as compared to other elements like Cl, Br.
- (2) Ionic fluoride like NaF, LiF have high lattice energy as compared to other halides.
- (3) Fluorides of Ca, Mg, Ba, Sr are insoluble in water due to high lattice energy while other halides of these elements are water soluble.
- (4) Due to low dissociation value of F_2 as compared to Cl_2 or Br_2 it has greater reactivity. Other halogens have low reactivity.
- (5) Fluorine has only -1 oxidation state while other halogens have oxidation states -1 , $+1$, $+3$, $+5$ and $+7$.
- (6) Due to restriction of the valence shell, to an octet, many fluoro compounds are inert e.g., CF_4 , SF_6 , SiF_4 .
- (7) Fluorine is the only element which combines directly with noble gases, like xenon, krypton and radon forming their fluorides.

OXIDIZING PROPERTIES OF HALOGENS

All the free halogens (F_2 , Cl_2 , Br_2 , I_2) act as oxidizing agents when they react with metals or non-metals. The reactant element acquires positive oxidation state in compound form.

On forming ionic compounds with metals, the halogen gain electrons and are converted into negative halide ions.



The oxidizing power of halogens decreases with increase in atomic number. Fluorine has the highest oxidizing power and iodine the least. The order of decreasing power as an oxidizing agent is:



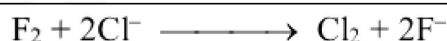
The oxidizing power of halogens depends upon the following factors:

- (i) Energy of dissociation
- (ii) Electron affinities of atoms
- (iii) Hydration energy
- (iv) Heats of vaporization (for Br_2 and I_2)

If a halogen has low energy of dissociation, a high electron affinity and a higher hydration energy of its ions, it will have a high oxidizing power.

Oxidizing power of F_2 is higher because it has low energy of dissociation and higher hydration energy of its ions. Due to the relative strength as oxidizing agents, it is possible for each free halogen to oxidize the ions of other halogens next to it, in the family.

Fluorine can oxidize all the halide ions to molecular halogens.



In the similar way chlorine will oxidize both bromide and iodide ions, while bromine can oxidize only iodide, ion.



Fluorine and chlorine can **oxidize various coloured dyes** to colourless substances. Colour of litmus or universal indicator can be decolourized when exposed to fluorine or chlorine. When chlorine acts as bleaching agent, it behaves as an oxidising agent.

COMPOUNDS OF HALOGENS

Hydrogen Halides, HX (Hydrides of Halogens):

Hydrogen halides are HF, HCl, HBr and HI.

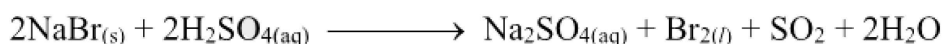
Preparation:

- (1) All halogens react with hydrogen to form halogen halides. e.g.,



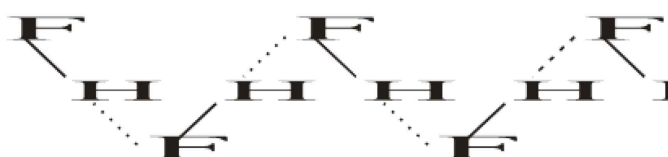
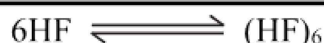
The reaction of molecular hydrogen and fluorine is very fast and explosive. With chlorine, molecular hydrogen reacts in the presence of sunlight. Bromine and iodine react with molecular hydrogen at a higher temperature. The reaction with iodine is very slow and reversible. Direct combination is used as a preparative method only for HCl and HBr.

- (2) Hydrogen fluoride and hydrogen chloride can also be obtained by the action of concentrated sulphuric acid on fluorides and chlorides, but analogous reactions with bromides and iodides result in partial oxidation of the hydrogen halide to the free halogen.



PROPERTIES OF HYDROGEN HALIDES

- (1) HF is a colourless volatile liquid whereas other hydrogen halides (HCl, HBr, HI) are colourless gases at room temperature. They give fumes in moist air. They are strong irritants.
- (2) **Hydrogen fluoride attacks glass** and has found applications as a non-aqueous solvent. It can be handled in teflon (polytetrafluorethylene) containers or if absolutely dry, in copper or stainless-steel vessels kept under vacuum.
- (3) Pure liquid HF is strongly hydrogen bonded and is **viscous liquid**. Its viscosity is less than that of water due to the absence of a three-dimensional network of H-bonds which occur in H₂O. Hydrogen bonding is also responsible for the association of HF molecules in the vapour phase. Various test results indicate that gaseous HF consists of an equilibrium mixture of monomers and cyclic hexamers.



Hydrogen bonding in HF

Chain polymers may also exist under certain conditions. Chains and rings of HF are of various sizes, some of these persist in the vapour phase as well. Some of the physical properties of hydrogen halides are given in table.

Table – Some physical properties of hydrogen halides

Property	HF	HCl	HBr	HI
M.P (°C)	–83.8	–114.2	–86.9	–50.8
B.P (°C)	19.5	–85.0	–66.7	–35.3
Heat of fusion at M.P. (kJ/mol)	4.58	1.99	2.41	2.87
Heat of vaporization at B.P. (kJ/mol)	30.3	16.2	17.6	19.7
Heat of formation / kJ mol ^{–1} (ΔH)	–27.0	–92.0	–36.0	+26.0
Bond energy kJ/mol ^{–1}	566	431	366	299
H-X Bond length (pm)	92	128	141	160
Dissociation into elements at 1000°C (%)	0	0.014	0.5	33
Dipole moment (Debye)	1.8	1.1	0.8	0.4

- (4) **Melting points, boiling points, heats of fusion and heats of vaporization** increase regularly from HCl to HI. The HF has much higher values for these properties due to hydrogen bonding. A very high boiling point of hydrogen fluoride is a major evidence of the presence of hydrogen bonding among its molecules. The relative volatility of HCl, HBr and HI reflects the strengthening of the van der Waal's forces due to increasing size of halogens. Since the dipole moment of molecules decreases from HCl to HI, probably dipole-induced dipole forces play an important role in the intermolecular binding of the heavier HX molecules.
- (5) **The strength of the hydrogen halogen bond** is very high in HF. It decreases with increasing size of the halogen atom. The bond strength is reflected in the ease of dissociation of hydrogen halides at elevated temperatures.
- (6) HF, HCl, HBr and HI act as reducing agents in the following order.



Hydrogen iodide is a stronger reducing agent. In redox reactions the hydrogen halides are oxidized to elemental halogens. e.g.,



- (7) In water, hydrogen halides give hydrofluoric, hydrochloric, hydrobromic and hydroiodic acids. Hydrofluoric acid is a weak acid due to limited ionization. The other three acids are very strong acids.



The strength of acids increases in the order of:



OXIDES OF HALOGENS

“The compounds which contain oxygen and halogens chemically combined together are called oxides of halogens.”

Halogens do not react directly with oxygen. These halogen oxides are formed indirectly. Some oxides of halogens with their names are given below in the table.

Table – Oxides of halogen

Fluorine	Chlorine	Bromine	Iodine
Oxygen difluoride OF_2	Dichlorine monoxide Cl_2O	Bromine monoxide Br_2O	Iodine tetraoxide I_2O_4
Dioxygen difluoride O_2F_2	Chlorine dioxide ClO_2	Bromine dioxide BrO_2	Iodine iodate I_4O_9
Trioxxygen difluoride O_3F_2	Chlorine heptaoxide Cl_2O_7	Bromine trioxide BrO_3	Iodine pentoxide I_2O_5

TRIOXYGEN DIFULORIDE (O_3F_2)

Preparation:

This oxide can be prepared when a mixture of fluorine and oxygen is subjected to the electric discharge.



Properties:

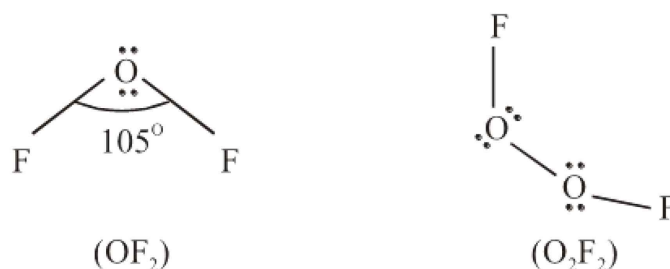
At 363°C , it is a dark red viscous liquid but turns to reddish brown solid at 350°C . on decomposition, it gives oxygen and other oxide of fluorine.



O_3F_2 reacts with F_2 in the presence of electric discharge to produce O_2F_2 .



Structure of OF_2 and O_2F_2 are shown below.



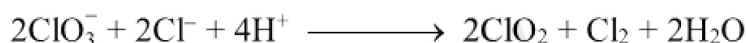
OXIDES OF CHLORINE

The oxides of chlorine are generally unstable. It is not possible to synthesize them by direct combination of the elements Cl_2 and O_2 . They have extensive industrial use as commercial bleaching agents for wood, paper-pulp and for water treatment.

(1) Chlorine Dioxide ClO_2 :

Preparation:

- (a) It is prepared by reducing NaClO_3 with NaCl or SO_2 or CH_3OH in strongly acidic solution.



- (b) ClO_2 can also be prepared by the action of concentrated H_2SO_4 on KClO_3 . This reaction is violent. To control the reaction, oxalic acid should be added.

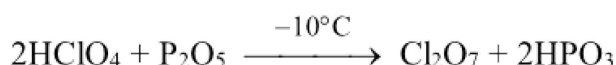


Properties:

- (1) It is pale yellow gas. ClO_2 explodes into Cl_2 and O_2 on warming.
- (2) It is soluble in water and is stable in dark.
- (3) It decomposes slowly in H_2O to HCl and HClO_3 .
- (4) It is a paramagnetic substance.
- (5) It is used as an antiseptic, for purification of water and to bleach cellulose material.

(2) Chlorine Heptaoxide (Cl_2O_7):

Cl_2O_7 is an anhydride of perchloric acid (HClO_4). It can be obtained at -10°C by dehydration of HClO_4 with P_2O_5 .



- (3) It is explosive in nature.
- (4) It dissolves in H_2O .

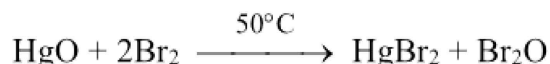


OXIDES OF BROMINE

Oxides of bromine are dark volatile liquids with low thermal stability.

(1) Bromine Monoxide Br₂O:

It can be prepared by the reaction of bromine vapours with mercuric oxide.



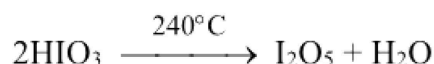
Br₂O can also be prepared by treating the suspension of mercuric oxide in CCl₄ with bromine. It is stable in dark in CCl₄ at -20°C. It has oxidizing properties.

OXIDES OF IODINE

Out of all the oxides of iodine, only iodine, pentoxide I₂O₅ is important. The other compounds, I₂O₄ and I₄O₉ are salt like and are considered as iodine-iodates.

(1) Iodine Pentoxide (I₂O₅):

It can be prepared by heating iodic acid at 240°C.



It is a white crystalline solid, stable up to 300°C. It has a polymeric structure. It is insoluble in organic solvents. It forms iodic acid with water.



It reacts with H₂S, HCl and CO and acts as an oxidizing agent.

It is used for the quantitative analysis of CO.



Reaction of chlorine with cold and hot NaOH (Disproportionation reaction):

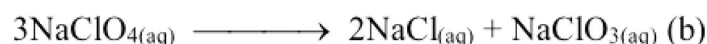
“A reaction in which a species (molecule, atom or ion) is simultaneously oxidized and reduced, is called a “disproportionation reaction”.

In cold (15°C) state chlorine will react with NaOH_(aq) to form hypochlorite and chloride.

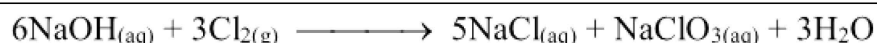
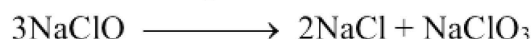
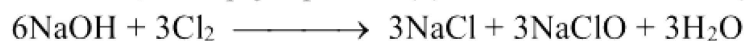


The reaction is a disproportionation reaction, because the zero oxidation state of chlorine atom in Cl₂ is converted to -1 in chloride and +1 in hypochlorite.

Sodium hypochlorite which is produced in cold state in the above reaction decomposes forming sodium chloride and sodium chlorate at 70°C.



The reaction (b) involves the disproportionation of hypochlorite ion. To balance overall reaction in hot state, multiply equation (a) with 3 and then add (a) and (b).



The above two reactions (a) and (b) show that chlorine atoms are both reduced and oxidized.

OXYACIDS OF HALOGENS

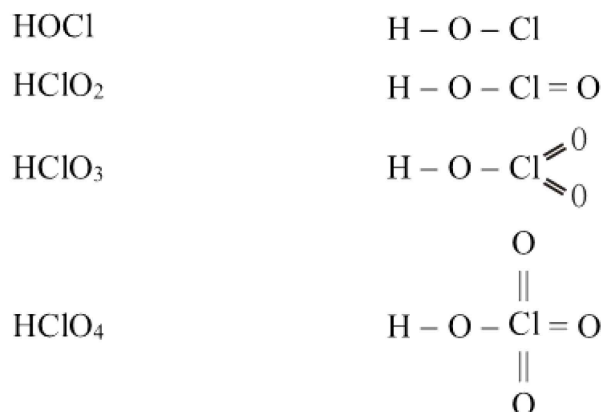
Halogens form four kinds of oxyacids and their anions. No stable oxyacid of fluorine (HOF) is known. Other halogens form oxyacids which cannot be isolated in pure form and are stable only in aqueous solutions in the form of their salts.

Names of Oxyacids and their Salts:

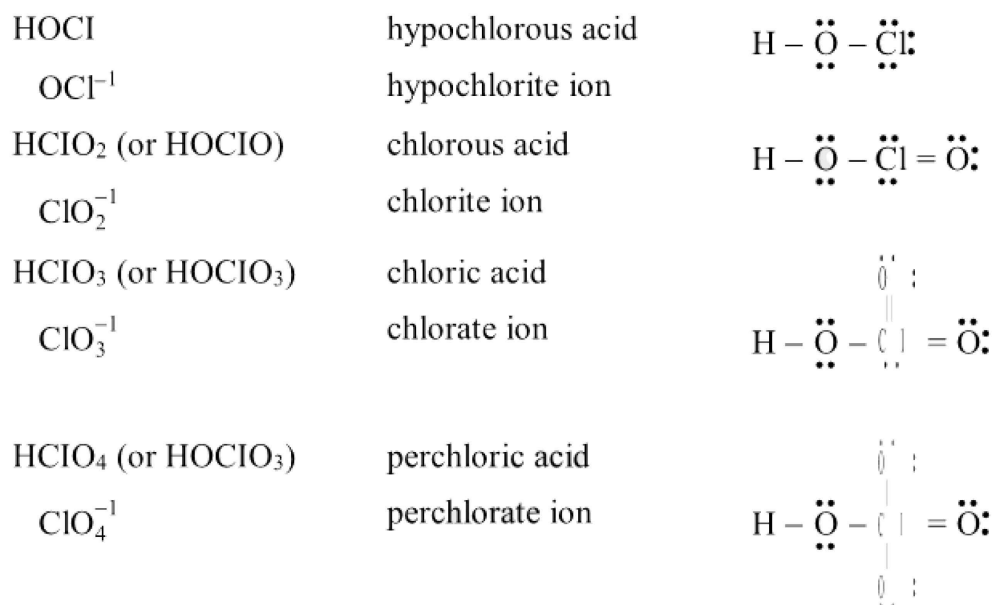
- (1) Oxyacids in which central atom halogen has oxidation state +1 ends at **hypo-ous** acid and salt ends as **hypo-ite**. e.g., HClO. (Hypochlorous acid) NaClO sodium hypochlorite.
- (2) Oxyacid in which central atom halogen has oxidation state +3 ends at **-ous** acid and salt ends at **-ite**. e.g., HClO₂ (chlorous acid) NaClO₂ (sodium chlorite).
- (3) Oxyacid in which central atom halogen has oxidation state +5 ends at **-ic** acid and salt ends at **-ate**. e.g., HClO₃ (chloric acid) NaClO₃ (sodium chlorate).
- (4) Oxyacids in which central atom halogen has oxidation state +7 ends at **per-ic** acid and salt ends at **per-ate**. e.g., HClO₄ (per chloric acid), NaClO₄ sodium perchlorate.

Table – Oxyacids of Cl, Br and I and anions of their salts

Oxidation state	Formula	Acids name	Anion	Name
+1	HClO	Hypochlorous acid	ClO ⁻	Hypochlorite
+3	HClO ₂	Chlorous acid	ClO ₂ ⁻	Chlorite
+5	HClO ₃	Chloric acid	ClO ₃ ⁻	Chlorate
+7	HClO ₄	Perchloric acid	ClO ₄ ⁻	Perchlorate
+1	HBrO	Hypobromous acid	BrO ⁻	Hypobromite
+3	HBrO ₂	Bromous acid	BrO ₂ ⁻	Bromite
+5	HBrO ₃	Bromic acid	BrO ₃ ⁻	Bromate
+7	HBrO ₄	Perbromic acid	BrO ₄ ⁻	Perbromate
+1	HIO	Hypoiodous acid	IO ⁻	Hypoiodite
+3	HIO ₂	Iodous acid	IO ₂ ⁻	Iodite
+5	HIO ₃	Iodic acid	IO ₃ ⁻	Iodate
+7	HiO ₄	Periodic acid	IO ₄ ⁻	Periodate



Structure of oxyacids of chlorine is given below:



PROPERTIES OF OXYACIDS

(i) Acidic Strength of Oxyacids:

The acid strength increases with increase in the number of oxygen atom. As the oxidation state of the halogen increases, the bonding electrons are shifted away from the H-atom and the tendency for the molecule to lose a proton increases. This accounts for the change of strength of oxyacids. The oxyacids of halogens show their strength in the order given below:



(ii) Thermal Stability:

Thermal stability of oxyacid molecules increase by increasing the number of oxygen atoms. Perchloric acid HClO_4 is more stable than HClO hypochlorous acid.

(iii) Oxidizing Power of Oxyacids:

Oxidizing strength of oxyacid decreases by increasing oxidation number. HClO and HClO_2 are good oxidizing agent. Oxidizing strength is given in the following order.

**(iv) Polarity:**

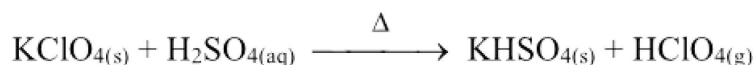
In oxyacids the halogens serves as a central atom to which one or more oxygen atoms are covalently bonded. These bonds are polar in characters due to the electronegativity difference between halogen and oxygen.

(v) Relative Strength of Oxyacids of Different Halogens:

The oxyacids of chlorine are stronger than the corresponding oxyacids of bromine which are, in turn, stronger than the corresponding oxyacids of iodine. It is due to decrease in the electronegativity and increase in the size of the halogen. We will discuss perchloric acid (HClO_4) only at the stage.

PERCHLORIC ACID HClO_4 **Preparation:**

Perchloric acid (HClO_4) is commonly obtained in aqueous solution. Pure anhydrous compound can be prepared by distilling a mixture of perchlorate (KClO_4) and conc. H_2SO_4 under reduced pressure.

**Properties:**

- (1) Perchloric acid is a colourless, hygroscopic liquid.
- (2) At normal pressure it freezes at -112°C and boils with decomposition at 90°C .
- (3) In the cold and dilute state perchloric acid is a very weak oxidizing agent but when hot and concentrated its oxidizing power is enhanced. Dissolving power of perchloric acid is enhanced due to its oxidizing strength. Perchloric acid is strongest of all the acid in an aqueous medium.



- (4) Pure perchloric acid decomposes explosively when heated. That is why it is stored and used as 67% solution in water. Perchloric acid reacts with organic substances violently.
- (5) Due to its oxidizing effect, acidic strength and solubility of its salts, it is considered as a valuable analytical reagent.

BLEACHING POWDER $\text{Ca}(\text{OCl})\text{Cl}$

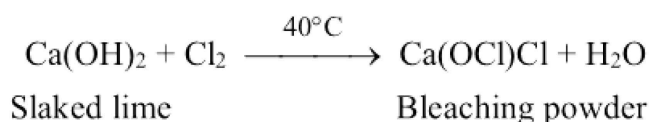
The chemical name of bleaching powder is calcium hypochlorite chloride. Formula of bleaching powder was given by **oddling**. It is mixed salt.

Preparation:

Bleaching powder can be manufactured by the action of chlorine on dry slaked lime using any one of the following methods:

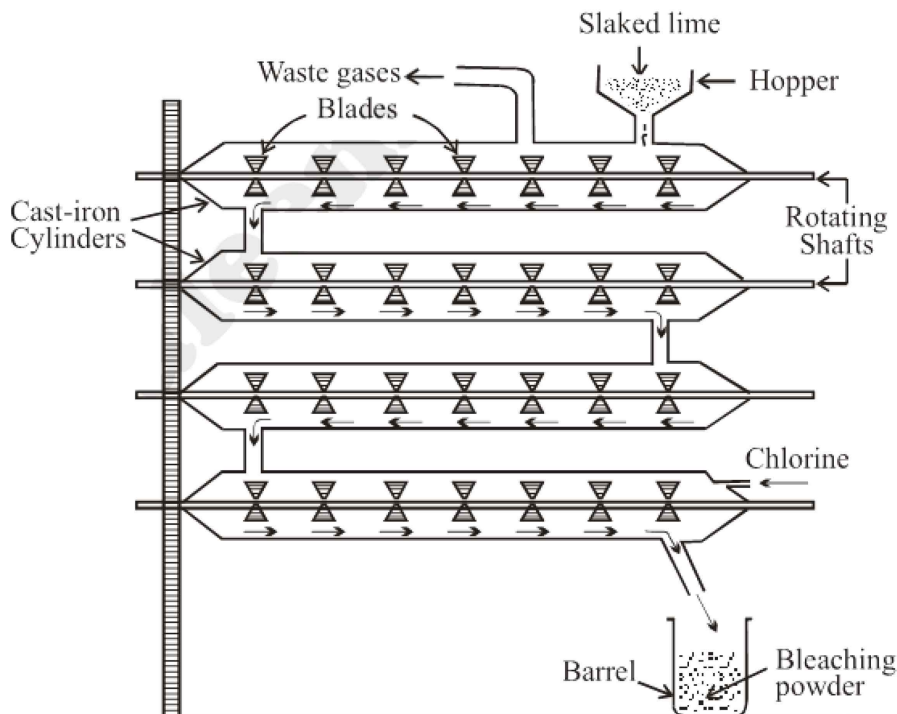
- (a) Hasenclever's method (old method)
- (b) Beckmann's method (modern method)

The reaction in both the cases will be:



(1) Hasenclever's Method:

The apparatus used in this method consists of 4 to 8 iron cylinders placed one above the other horizontally. They are interconnected and provided with stirrers. The slaked lime is added in through a hopper in the upper cylinder and is transported from one cylinder to the other with rotating stirrers. Chlorine introduced into the lowest cylinder rises up and reacts with slaked lime to form bleaching powder which is collected through the outlet in the lowest cylinder.

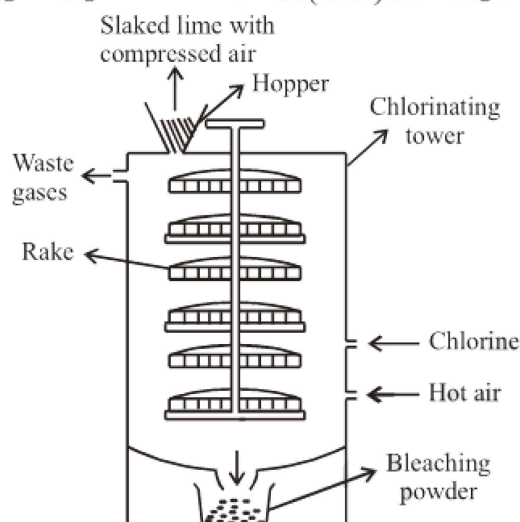
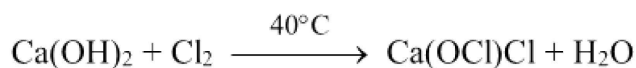


Hasenclever's method

(2) Beckmann's Method:

In this method a cast iron tower with eight horizontal shelves is used. In each shelf there is a rotating rake.

Powdered slaked lime is introduced through hopper at the top with compressed air. A mixture of hot air and chlorine are introduced from the base of the tower. The slaked lime is pushed down by the rotating rakes while chlorine rises up. Reaction between slaked lime and chlorine produces bleaching powder which is collected at the bottom of the tower.



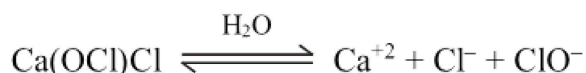
The apparatus works on **counter-current principle**. In counter-current principle reacting substances move in opposite directions and intermix together.

Thus maximum reaction of slaked lime and chlorine is brought about with very little loss of chlorine. Bleaching powder should always be packed in air tight containers to avoid the loss of chlorine.

Bleaching powder is a **yellowish white powder** with strong smell of chlorine.

CHEMICAL PROPERTIES OF BLEACHING POWDER

- (1) It is an oxidizing agent. This property is due to the generation of hypochlorite ion (OCl^-) in water.



- (2) With dilute acid it gives hypochlorous acid.



- (3) If excess of an acid (weak or strong) is added to bleaching powder, chlorine is given out.



(b) Uses of Chlorine and its Compounds:

- (1) Chlorine is used in the manufacture of **bleaching powder**.
- (2) It is used as a disinfectant in swimming pools and water treatment plants.
- (3) A number of antiseptics, insecticides, weed killers and herbicides are manufactured from chlorine.
- (4) It is also used in the manufacture of hydrochloric acid, which is the cheapest industrial acid.
- (5) Chlorine is also used in the manufacture of **polyvinyl chloride** (PVC) plastics.
- (6) **Chloroform** and carbon tetrachloride are prepared from chlorine which are used as solvents.
- (7) A salt of chlorine, sodium chloride is used in food and important ingredient of our diet.

(c) Uses of bromine and its Compounds:

- (1) Ethylene dibromide ($C_2H_4Br_2$) is added to leaded gasoline to save the engine from lead oxide and lead sulphate deposits.
- (2) Bromine is also used as **fungicide**.
- (3) Silver bromide $AgBr$ is used in **photography**. It is light sensitive compound.

(d) Uses of Iodine and its Compounds:

- (1) The major applications of iodine are in pharmaceutical industry. It is used as disinfectant and germicide. Tincture of iodine and **iodex** are popular preparations of iodine. (I_2 in C_2H_5OH and H_2O containing KI is called tincture of iodine).
- (2) Diet with insufficient iodide ion leads to an enlargement of the **thyroid** (Goiter). To ensure the presence of iodide ion the diet, sodium or potassium iodide is added to the common salt which is known as iodized salt.
- (3) Deficiency of iodine in the body is cause of dwarfism.

NOBLE GASES**Introduction:**

Noble gases are present in group VIIIA or the zero group of periodic table. Noble gases consist of helium, neon, argon, krypton, xenon and radon. All the elements of zero group are colourless and odourless gases. They exist in mono-atomic form. At low temperature and high pressure, they can be liquefied or solidified. These gases are also called rare gases. The noble gases are present in atmosphere and their percentage in atmosphere is less than 1%.

The amount of chlorine thus set free is called “**available chlorine**”. The activity of bleaching powder is measured in terms of available chlorine. The average percentage of available chlorine in bleaching powder is 35–40 percent. The bleaching action of bleaching powder is due to its oxidative character.

- (4) It oxidizes HCl, HBr and HI giving the corresponding halogens.



- (5) It oxidizes ammonia to nitrogen.



- (6) It reacts with CO_2 to form CaCO_3 and Cl_2 .



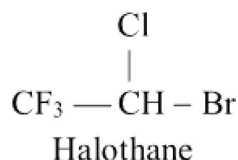
Uses:

- (1) Bleaching powder is used for the preparation of Cl_2 and O_2 .
- (2) It is also used for the preparation of chloroform CHCl_3 .
- (3) It is used as disinfectant and for the sterilization of water.
- (4) It is used for making unshrinkable wool.
- (5) It is used to bleach cotton, linen and paper pulp. Delicate fabrics like wool, silk are not bleached with bleaching powder because they can be damaged by chlorine produced from bleaching powder.
- (6) For the removal of poisonous gases from the atmosphere.

COMMERCIAL USES OF HALOGENS AND THEIR COMPOUNDS

(a) Uses of Fluorine and its Compounds:

- (1) Fluorine is used for the preparation of Feons. **Freon** is the commercial name of low molecular mass fluorochlorocarbons, CCl_2F_2 , CClF_3 . These are being used as refrigerants and aerosol propellants.
- (2) Fluorine is used to prepare **Teflon** $(-\text{CF}_2 - \text{CF}_2)_n$. It is a polymerized tetrafluoro ethylene compound. It is a valuable plastic which resists the action of oxidants, acids and alkalies. Corrosion-proof parts of machinery are made of it. It is used for coating the electrical wiring. Teflon is also used as a non-stick coating for cooking pans.
- (3) **Halothane** is used as an anesthetic.



- (4) Fluorides in toothpastes build a protective coating on teeth.
- (5) HF is used for etching glass.
- (6) Sodium fluoro-acetate is rat poison.

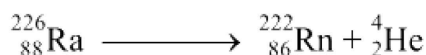
Table – Electronic configuration of noble gases

Name	Symbol	Atomic number	Electronic configuration
Helium	He	2	$1s^2$
Neon	Ne	10	$1s^2, 2s^2, 2p^6$
Argon	Ar	18	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6$
Krypton	Kr	36	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6$
Xenon	Xe	54	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6$
Radon	Rn	86	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 6p^6$

OCCURRENCE OR EXTRACTION OF NOBLE GASES

- (1) The noble gases are isolated from air, either by fractional distillation or by some chemical method. The principal commercial source of Ne, Ar, Kr and Xe is air.
- (2) **Helium** is present on earth as a result of radioactive decay. After hydrogen, it is the second most abundant element in the universe. α -particles are doubly ionized helium atoms He^{2+} . It is simple and economical to isolate the helium gas from certain natural gases by liquefaction method. In 1868, He was detected from the chromosphere of sun. The name helium is from the Greek word “Heilos” which means sun.
- (3) **Neon** is $1/65000^{\text{th}}$ part of the atmosphere and it is also isolated during liquefaction of air in a discharge tube, neon glows reddish (of all the noble gases, the discharge of neon is the most intense at ordinary voltage and current). Liquid neon has over 40 times more refrigeration capacity than liquid helium.
- (4) **Argon** was discovered by “Ramsay”. It is a colourless and odourless gas. It is very inert and not known to form any true chemical compound. It is obtained as a by product during the liquefaction of air.
- (5) Traces of **krypton** are present in air. It is a colourless, odourless and fairly expensive gas. It is characterized by its brilliant green and orange spectral lines. Its compound krypton difluoride (KrF_2) can be prepared by various methods.
- (6) **Xenon** is present in the atmosphere to a very small extent (0.08 ppm). It is obtained as a by product during the fractional liquefaction of air. Xenon is available commercially in cylinders at high pressure. It reacts with fluorine but not with water. However it is slightly soluble in water to the extent of about 110 ml/litre at 20°C .

- (7) **Radon** is the α -decay product of the radium. Radon is present to a very small extent in the atmosphere and it could be obtained as a by product from the liquefaction of air. However the small quantities of this gas which are usually needed can be collected from the radioactive decay of radium isotopes.



PHYSICAL PROPERTIES OF NOBLE GASES

- (1) The noble gases have valence shells which are closed octets (except He). Due to closed shells their ionization energy values are very high.
- (2) They have low boiling points. The boiling point of helium is the lowest of any known substance. Their boiling points increase with increasing atomic number down the group.
- (3) The very low values of melting and boiling points and low heats of vaporization show that noble gases have weak forces of attraction between their atoms. As there are no ordinary electron pair interaction, these weak forces must be of the van der Waal's type.
- (4) The solubility of the noble gases in water increases with increasing atomic number. This is because the bigger atoms are more readily polarized by water molecules.

Table – Some physical properties of noble gases

Property	He	Ne	Ar	Kr	Xe	Rn
At. mass (a.m.u)	4	20.18	39.95	83.8	131.3	222
M.P (°C)	-272	-249	-189	-157	-112	-71
B.P (°C)	-269	-246	-186	-153	-108	-61
Ionization energy (1 st) kJ mol ⁻¹	2372	2081	1521	1351	1170	1037
Water solubility (ml/lit) at 20°C	13.8	14.7	37.9	73.00	110.9	—
Heat of vaporization kJ mol ⁻¹	0.08	1.77	6.5	9.7	13.7	18.0

COMPOUNDS OF XENON

Xenon reacts directly with fluorine only. The known oxidation state of Xe in its compounds range from +2 to +8. The compounds are stable and can be obtained in large quantities. Some important compounds of Xe are given in the table. The first compound of Xe (Xenon Hexafluoroplatinate) was prepared by N-Bartlet in 1962 in Chicago University.

Table – Compounds of Xenon

Oxidation state of xenon	Compound	Physical form	Milting point (°C)
+2	XeF ₂	Colourless crystals	140
+4	XeF ₄	Colourless crystals	114
	XeOF ₂	Colourless crystals	90
+6	XeF ₆	Colourless crystals	48
	XeOF ₄	Colourless liquid	–28
	XeO ₃	Colourless crystals	25 (Explodes)
+8	XeO ₄	Colourless gas	–39.9 (Explodes on warming)

(a) FLOURIDES OF XENON

Three known fluorides of Xenon are XeF₂, XeF₄ and XeF₆.

XeF₂ can be prepared by direct interaction of the elements. The compound formed should be removed immediately from the reaction zone, otherwise further reaction with F₂ will give XeF₄. The reaction is completed in about **8 hours**. XeF₂ is a crystalline solid. It is stored in nickel vessels. XeF₂ is a mild fluorinating agent.



XeF₄ can be prepared by heating a mixture of **Xe and F₂ in 1 : 5 ratio** in a nickel container under **6 atmospheric** pressure for a **few hours**. It can be stored in nickel vessels. Its properties are similar to that of **XeF₂** but it is a strong fluorinating agent.

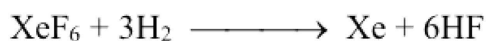
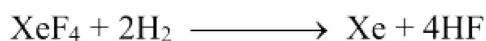


XeF₆ require more severe conditions. **Xe and F₂ are taken in 3 : 20 ratio** in a stainless steel vessel and heated to **300°C at 50 atmospheric** pressure. More than 95% conversion to XeF₆ takes place. XeF₆ is a crystalline solid. It is colourless in the solid state but yellow in liquid and gaseous forms.



CHEMICAL REACTIONS OF FLUORIDES OF XENON

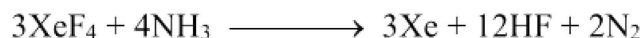
- (1) Fluorides of xenon can be reduced with hydrogen at 400°C giving xenon and hydrofluoric acid.



- (2) Xenon tetra-fluoride is a good fluorinating agent and can be used to prepare metal fluorides as follows.



- (3) Reaction occurs with explosion when XeF_4 is brought in contact with liquid ammonia.



- (4) Hydrolysis of XeF_6 with small amount of water gives XeOF_4 .



XENON OXYFLUORIDES

Xenon oxytetrafluoride, XeOF_4 is also formed by a rapid reaction of XeF_6 with silica (quartz).



XeOF_4 is colourless volatile liquid. It can be kept in nickel vessel. It reacts with water to give XeO_3 .

Xenon oxydifluoride XeOF_2 , is obtained when xenon reacts with oxygen difluoride in an electric discharge.



OXIDES OF XENON

There are two oxides are xenon:

- (1) Xenon trioxide
- (2) Xenon tetraoxide

(1) Xenon Trioxide XeO_3 :

Xenon trioxide can be obtained when XeF_6 is hydrolysed slowly.



It is a crystalline solid. It explodes at very low temperature. It is weakly acidic and its aqueous solution is almost non-conductor.

(2) Xenon Tetraoxide XeO_4 :

It is obtained by the addition of barium or sodium perxenate to conc. H_2SO_4 .

