13

HALOGENS

Student Learning Outcomes [C-11-B-19 to C-11-B-41]

After studying this chapter, students will be able to:

- Describe the colours and the trends in volatility of chlorine, bromine, and iodine. **(Understanding)**
- Describe the trend in bond strength of halogen molecules. **(Understanding)**
- Interpret the volatility of the elements in terms of instantaneous dipole-induced dipole forces. **(Understanding)**
- Describe the relative reactivity of halogen elements as oxidizing agents. **(Understanding)**
- Describe the reactions of elements with hydrogen and explain their relative reactivity in these reactions. **(Understanding)**
- Describe the relative thermal stabilities of hydrogen halides and explain these in terms of bond strength. **(Understanding)**
- Describe the relative reactivity of halide ions as reducing agents. **(Understanding)**
- Explain the reactions of halide ions with aqueous silver nitrate and concentrated sulphuric acid. **(Understanding)**
- Describe the reactions of halides with aqueous silver ions followed by aqueous ammonia. **(Understanding)**
- Interpret the reaction of Chlorine with cold and hot aqueous sodium hydroxide as disproportion reactions. **(Understanding)**
- Explain the use of chlorine in water purification, including the production of the active species $HOCl$ and ClO^- which kill bacteria. **(Understanding)**

Elements present in Group 17 of the periodic table are termed as halogens. It includes fluorine (F), chlorine (Cl), bromine (Br), iodine (I), astatine (At) and tennessine (Ts). The halogen elements form a group of very reactive non-metals and are quite similar to each other in their chemical properties. First four elements are the common elements of the halogen family but last two astatine (At) and tennessine (Ts) are very rare and radioactive elements.

Halogens exist as diatomic molecules in all phases (gas, liquid or solid). Fluorine (F_2) and chlorine (Cl_2) are gases of pale yellow and greenish yellow colours respectively at room temperature and pressure. Bromine (Br_2) is a volatile liquid of reddish-brown colour at room temperature. It has corrosive and toxic fumes. Iodine (I_2) is shiny greyish black solid at room temperature. It sublimes directly from solid to a violet vapor. The colors of halogens (X_2) darken progressively from chlorine to iodine. The trends of colour changes from chlorine to iodine is due to changes in the absorption of light as a result of electron transitions within the molecules of respective halogens.

Figure 13.1 Greenish yellow chlorine (left), orange bromine (middle) and purple iodine (right)

Did You Know!

The name halogen comes from the Greek words "halos", meaning "salt", and "gen", meaning "to make." The first halogen to be isolated and recognized as an element was chlorine. Despite the fact that chlorine is poisonous, small amount is essential to human health and life in the form of chloride.

Table 13.1 Atomic and Physical properties of the common halogens.

Element	Fluorine	Chlorine	Bromine	Iodine
Proton number	9	17	35	53
Electron shell structure	2, 7	2, 8, 7	2, 8, 18, 7	2, 8, 18, 18, 7
Outer shell electron configuration	$2s^2 2p^5$	$3s^2 3p^5$	$4s^2 4p^5$	$5s^2 5p^5$
Relative atomic mass (μ)	19.0	35.5	79.9	126.9
Physical state at of 20 ºC	gas	gas	liquid	solid
Colour	pale yellow	pale green	red–brown	dark gray
Melting point/ºC	–220	–101	–7	113
Boiling point/ºC	–188	–35	59	183
Enthalpy change of vaporisation /kJ mol^{-1}	+3.3	+10.2	+15	+30
Solubility/g per 100 g of water at 20ºC	reacts readily with water	0.59 (reacts slightly)	3.6	0.018

13.1 VOLATILITY OF CHLORINE, BROMINE AND IODINE

Chlorine being a liquid is very volatile at room temperature and it disperses quickly in the air. Bromine being a liquid is less volatile than chlorine but more volatile than iodine. It evaporates readily releasing toxic fumes at room temperature. Iodine is the least volatile among the three. At room temperature, its solid states show lower volatility as compared to chlorine and bromine. Generally, volatility decreases from chlorine to iodine. This trend is due to the increasing molecular mass, increase in size of outer shell and stronger intermolecular forces (London dispersion forces) as we move down the group in the periodic table.

Interesting Information

Bromine liquid evaporates easily at room temperatures emitting an orange vapor. Bromine has a very strong and bad odor. It gets its name from the Greek word "bromos" which means "stench."

13.2 TREND IN VOLATILITY OF THE HALOGENS

Halogens are non-polar and instantaneous dipole-induced forces play a significant role in determining the volatility of halogens. The forces depend on factors like molecular size, shape and polarizability, with stronger forces leading to lower volatility and higher boiling points. Substances with weak id-id forces, smaller and less polarizable molecule have lower boiling points and higher volatility. Larger, more polarizable molecules with stronger London dispersion forces have higher boiling points and lower volatility.

Volatility is inversely related to the boiling point of a substance. A more volatile substance will have a lower boiling point. Stronger intermolecular forces require more energy to separate the molecules from the liquid phase to the gaseous phase. A summary about polarizability and volatility is given in **Table 13.2.**

Table 13.2 Effect of London forces on physical properties

Size	Polarizability	Instantaneous dipole-induced dipole forces	Boiling Point	Volatility
Small	Small	Weak	Lower	Higher
Large	High	Strong	Higher	Lower

The first two halogens, i.e, fluorine (colourless or very light green) and chlorine (greenish yellow) are gases due to weaker id-id forces. Bromine is a liquid as its size is bigger and it possesses stronger intermolecular forces than fluorine and chlorine. Iodine has the strongest forces among the group, so it is in solid state at room temperature.

Quick Check 13.1

a) Which halogens elements are radioactive?
b) What is the reason behind the different colours of halogens?
c) Why chlorine is more volatile than bromine and iodine?

13.3 THE BOND STRENGTH OF HALOGEN MOLECULES

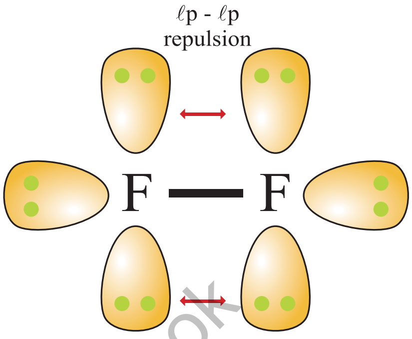

Figure. 13.2 F_2 molecule with lone pairs of electrons

Bond strength in halogens decreases as we move down the group from chlorine to iodine. This is due to the increase in the atomic size down the group which results in longer bond lengths and weaker bonds. Moving from top to bottom in group 17, the bond energies of halogens decrease gradually from chlorine to iodine. However, fluorine is an exception in this group. In the case of fluorine, the bond strength is relatively weak because fluorine atoms are very small. Due to the small size, there is significant electron–electron repulsion between the lone pairs on the fluorine atoms, which weakens the bond despite the high electronegativity of fluorine. The bond energy of halogens is listed in **Table 13.3.**

Table 13.3 Bond energies of halogen molecules

Halogen Molecule	Bond Energy ($kJmol^{-1}$)
F_2	156
Cl_2	243
Br_2	193
I_2	151

13.4 RELATIVE REACTIVITIES OF THE HALOGENS AS OXIDIZING AGENTS

All the free halogens act as oxidizing agents when they react with metals and most of non-metals. On forming ionic compounds with metals, the halogens gain electrons and are converted to negative halide ions.

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

The oxidizing power of halogens decreases down the group. Fluorine has the highest oxidizing power and iodine the least. The order of decreasing power as an oxidizing agent is:

$$F_2 > Cl_2 > Br_2 > I_2$$

The reactivity of halogens is directly related to their ability to acquire an electron and form halide ions (F^-, Cl^-, Br^- and I^-) when they react with other elements. Fluorine has the highest tendency to acquire an electron and form fluoride. Fluorine molecule can oxidize and displace all the halide ions from their solutions (Cl^-, Br^- and I^-) to free halogens.

$$F_{2(g)} + Cl^{1-}_{(aq)} \longrightarrow F^{1-}_{(aq)} + Cl_{2(g)}$$
$$F_{2(g)} + Br^{1-}_{(aq)} \longrightarrow F^{1-}_{(aq)} + Br_{2(g)}$$
$$F_{2(g)} + I^{1-}_{(aq)} \longrightarrow F_{(aq)}^{1-} + I_{2(g)}$$

Similarly, chlorine can oxidize and displace Br^- and I^-. Bromine can oxidize and displace I^-. Iodine cannot oxidize any halide ion.

The oxidizing power of halogens can be related to standard electrode potential (E°) values. Fluorine is the most reactive halogen and the most powerful oxidizing agent. The standard electrode potential $E^\circ(X_2/X^-)$ for halogens shown in **Table 13.4** become less positive from fluorine to iodine. This reflects the decreasing oxidizing power. The oxidizing power of halogens depends upon various factors, i.e., energy of dissociation, electron affinity of atoms, hydration energies of ions, and heats of vaporization (for Br_2 and I_2). A halogen having low energy of dissociation, high electron affinity and higher hydration energy of its ions, will have a high oxidizing power.

Table 13.4 Standard electrode potential $E^\circ(X_2/X^-)$

Halogen Molecule (X_2)	Standard Reduction Potential, E°(V)
F_2	+ 2.87
Cl_2	+1.36
Br_2	+1.07
I_2	+0.54

Quick Check 13.2

a) The F-F bond is weaker than Cl-Cl bond although fluorine is the most electronegative element, Explain.

b) Is the reaction between $NaCl_{(aq)}$ and F_2 gas possible?

i) Give reason whether yes or no.

ii) If yes, write the equation for this reaction.

c) What is the relationship between the oxidizing power of halogens and their standard reduction potential values?

13.5 REACTIONS OF THE HALOGENS WITH HYDROGEN

When halogen elements react with hydrogen, they produce hydrogen halides.

$$H_2 + X_2 \longrightarrow 2HX \qquad \text{Where } X_2 = F_2, Cl_2, Br_2 \text{ and } I_2$$

Hydrogen halides are colourless gases that dissolve in water to form hydrohalic acid. When we move down the group from fluorine to iodine, the reactivity of halogens with hydrogen decreases.

$$F_2 > Cl_2 > Br_2 > I_2$$

At low temperature and in the dark, fluorine reacts explosively with hydrogen.

$$H_{2(g)} + F_{2(g)} \longrightarrow 2HF_{(g)}$$

When this gas is dissolved in water, it forms hydrofluoric acid.

In the presence of UV light or a spark, chlorine reacts readily with hydrogen and produces colourless HCl gas which forms hydrochloric acid in water.

$$H_{2(g)} + Cl_{2(g)} \xrightarrow{hv} 2HCl_{(g)}$$

Hydrogen bromide (HBr) gas is produced when bromine reacts with the hydrogen upon heating. HBr gas is less reactive than HCl and HF. It forms a strong hydrobromic acid in water. This is an exothermic reaction.

$$H_{2(g)} + Br_{2(g)} \xrightarrow{\Delta} 2HBr_{(g)}$$

At high temperature and in presence of a catalyst, iodine reacts with hydrogen to form hydrogen iodide (HI) gas, which forms hydroiodic acid in water.

$$H_{2(g)} + I_{2(g)} \overset{\text{catalyst}}{\rightleftharpoons} 2HI_{(g)}$$

This is a reversible reaction and occurs very slowly.

13.6 RELATIVE THERMAL STABILITIES OF HYDROGEN HALIDES IN TERMS OF THEIR BOND STRENGTH

As we move down the halogen group from fluorine to iodine, the thermal stabilities of hydrogen halides (H-X) decrease due to decrease in bond dissociation energies as given in **Table 13.5**. The order of thermal stability of HX is as follows:

$$HF > HCl > HBr > HI$$

Bond strength between the hydrogen and the halogen atom in the H-X molecule explain this trend. Hydrogen fluoride (HF) is the most thermally stable hydrogen halide. Due to the high electronegativity of fluorine and its small atomic radius, a strong overlap of orbitals produces a very strong H-F bond. The bond dissociation energy of H-F (569 kJ/mol) is the highest among the hydrogen halides. Hydrogen chloride (HCl) is less thermally stable than hydrogen fluoride but more stable than other hydrogen halides. Due to the larger atomic radius of chlorine than fluorine H-Cl bond is weaker than H-F bond. Its bond dissociation energy is less than that of H-F. Its bond dissociation energy is 431 KJ/mol. Hydrogen bromide (H-Br) bond is weaker than H-F and H-Cl due to the larger atomic radius of bromine. There is a reduced overlapping of orbitals. Its bond dissociation energy is 366 kJ/mol. Among the common hydrogen halides, hydrogen iodide (HI) is the least thermally stable. Atomic radius of iodine is very large leading to poor orbital overlap due to which hydrogen iodide bond is the weakest among H-F, HCl and HBr. The bond dissociation energy of hydrogen iodide is 299 kJ/mol.

Table 13.5 Bond dissociation energy of H-X bonds

Hydrogen halide (HX)	Bond dissociation energy ($kJmol^{-1}$)
H-F	569
H-Cl	431
H-Br	366
H-I	299

Quick Check 13.3

a) The reaction between H_2 and F_2 is explosive but that between H_2 and I_2 is slow and reversible. Explain why.

b) Refer to Table 13.4 and 13.5 to predict which of the following reactions would be more exothermic. Explain your answer.

$H_{2\,(g)} + Cl_{2\,(g)} \longrightarrow 2HCl_{(g)}$ and $H_{2\,(g)} + Br_{2\,(g)} \longrightarrow 2HBr_{(g)}$

c) How thermal stability of hydrogen halides is related to their bond dissociation energies?

d) HF is the most thermally stable hydrogen halide. Give reasons.

13.7 RELATIVE REACTIVITY OF HALIDE IONS AS REDUCING AGENTS

The reducing ability of halide ions increases as we move down the group from fluorine to iodine. This trend is mostly due to the decreasing electronegativity and increasing atomic radius down the group. This result in lower charge density and greater ease of electron donation. Fluoride ion is the weakest reducing agent while iodide is the strongest reducing agent among the halides.

The order of decreasing power as a reducing agent is **$I^- > Br^- > Cl^- > F^-$**. The following is an explanation of this trend in terms of their properties.

The high electronegativity of fluorine and the strong bond between the fluorine atom and its extra electron make it difficult for fluoride ion (F^-) to donate electrons. The small size of fluoride ion results in a high charge density, further stabilizing the fluoride ion and reducing its tendency to lose an electron.

Chloride ion (Cl^-) is a stronger reducing agent than fluoride but still weaker than bromide and iodide. This is due to its larger size and lower electronegativity.

Bromide (Br^-) is a stronger reducing agent than chloride. Bromine is less electronegative than chlorine. Bromide ion (Br^-) has a larger ionic radius and lower charge density due to which it easily loses an electron and act as a reducing agent.

Iodide (I^-) is much stronger reducing agent than both chloride and bromide. Electronegativity of iodine is much lower than chlorine and bromine. Iodide ion (I^-) is large in size. This results in a lower charge density and a tendency to donate an electron by the iodide ion, making it strong reducing agent.

13.8 REACTIONS OF HALIDES WITH AQUEOUS SILVER ION FOLLOWED BY AQUEOUS AMMONIA

13.8.1 Reactions of halides with aqueous silver ion

Insoluble silver halides (AgX) are formed when halide ions react with aqueous silver nitrate ($AgNO_3$). These reactions are used in qualitative analysis to identify halides ions. The general reaction is given below:

$$X^{1-}_{(aq)} + Ag^{1+}_{(aq)} \longrightarrow 2AgX_{(s)}$$

No reaction is visible when silver ions and fluoride ions are mixed in the aqueous medium. As silver fluoride is soluble in water, fluoride (F^-) does not form precipitate as silver fluoride.

When silver ions and chloride ions are mixed in the aqueous medium a white precipitate of silver chloride forms, which is soluble in dilute ammonia.

$$Ag^{1+}{}_{(aq)} + Cl_{(aq)}{}^{1-} \longrightarrow AgCl_{(s)}$$

On mixing of aqueous silver ions with aqueous Bromide ions, a cream-coloured precipitate of silver bromide forms. This precipitate of silver bromide is sparingly soluble in concentrated ammonia.

$$Ag_{(aq)}{}^{1+} + Br_{(aq)}{}^{1-} \longrightarrow AgBr_{(s)}$$

A yellow precipitate of silver iodide forms on mixing aqueous silver ion with aqueous halide solution. The product AgI is insoluble in ammonia.

Figure 13.3 Colours of the silver halide precipitates: silver chloride (left), silver bromide (middle) and silver iodide (right).

$$Ag^{1+}_{(aq)} + I^{1-}_{(aq)} \longrightarrow AgI_{(s)}$$

This type of reactions is used to identify halide ions and is called **silver nitrate test**.

13.8.2 Reaction of silver halides (AgX) with aqueous ammonia

Addition of aqueous ammonia tests the solubility of initially formed silver halide precipitates. AgF is soluble in water so it does not form precipitate, so ammonia has no effect. Silver chloride dissolves in dilute ammonia to form diamminesilver(1) complex. The white precipitate of AgCl dissolves in dilute ammonia forming a colourless solution.

$$AgCl_{(s)} + 2NH_{3(aq)} \longrightarrow [Ag(NH_3)_2]^+{}_{(aq)} + Cl_{(aq)}{}^{1-}$$

The cream-coloured precipitate of AgBr dissolves in concentrated ammonia, forming a colourless solution.

$$AgBr_{(s)} + 2NH_{3(aq)} \longrightarrow [Ag(NH_3)_2]^+_{(aq)} + Br^{1-}_{(aq)}$$

The yellow precipitate of silver iodide (AgI) does not dissolve in both dilute and concentrated ammonia, so there is no change in the presence of ammonia.

The below sequence of reactions in **Table 13.6** provides a systematic way to differentiate between halide ions using their solubility with aqueous silver ions and ammonia.

Quick Check 13.4

a) F^- is a weaker reducing agent than Cl^-. Explain why.
b) What is the cause of the different solubilities of silver halides in aqueous ammonia?
c) Write down the equation for the reaction of KI with Ag^+ followed by aq. NH_3. What would you observe at the completion of this reaction?

Table 13.6 Action of Ag^+ followed by ammonia on halides

Halide ion	Action of aq. Ag^+ ion	Action of Aqueous Ammonia
Fluoride ion (F^-)	No precipitate	No reaction with aq. NH_3
Chloride ion (Cl^-)	White precipitate (AgCl)	Soluble in dil. aq. NH_3
Bromide ion (Br^-)	Cream colour precipitate (AgBr)	Soluble in conc. aq. NH_3
Iodide ion (I^-)	Pale yellow precipitate(Agl)	Insoluble in aq. NH_3

13.9 REACTIONS OF HALIDES (X^-) WITH CONCENTRATED SULFURIC ACID

The reactions of different halide ions with concentrated sulfuric acid are different from one another. The nature of the product and nature of reaction changes down the group from fluoride to iodide.

Fumes of hydrogen fluoride gas are produced when concentrated sulfuric acid reacts with NaF. HF is a weak reducing agent, therefore, it does not react with H_2SO_4 further.

$$NaF_{(s)} + H_2SO_{4(\ell)} \longrightarrow NaHSO_{4(\ell)} + HF_{(g)}$$

When concentrated sulfuric acid reacts with sodium chloride, fumes of hydrogen chloride gas are produced.

$$NaCl_{(s)} + H_2SO_{4(\ell)} \longrightarrow NaHSO_{4(\ell)} + HCl_{(g)}$$

In case of Br^{1-} ion, steamy fumes of hydrogen bromide (HBr) gas and brown fumes of bromine (Br_2) are produced along with the smell of sulfur dioxide (SO_2). It is a redox reaction. Hydrogen bromide acts as a reducing agent by reducing sulfuric acid to sulfur dioxide and is self-oxidized to bromine.

$$NaBr + H_2SO_{4(\ell)} \longrightarrow NaHSO_{4(l)} + HBr_{(g)}$$

$$2HBr_{(s)} + H_2SO_{4(\ell)} \longrightarrow Br_{2(g)} + 2SO_{2(g)} + 2H_2O_{(\ell)}$$

When concentrated sulfuric acid reacts with sodium iodide, fumes of hydrogen iodide gas, purple fumes of solid iodine I_2 and smell of H_2S gas produced. HI acts as a strong reducing agent. It reduces H_2SO_4 to H_2S and is itself oxidized to I_2.

$$NaI_{(aq)} + H_2SO_{4(\ell)} \longrightarrow NaHSO_{4(\ell)} + HI_{(g)}$$

$$8HI_{(s)} + H_2SO_{4(\ell)} \longrightarrow 4I_{2(s)} + H_2S_{(g)} + 4H_2O_{(\ell)}$$

The trend of increasing reducing power of the halide ions (X^-) from fluoride to iodide, leads to more complex reactions with concentrated sulfuric acid.

13.10 REACTIONS OF CHLORINE WITH COLD AND HOT AQUEOUS SODIUM HYDROXIDE

Those reactions in which a single element undergoes both oxidation and reduction simultaneously are known as disproportionation reactions. When chlorine reacts with cold and hot aqueous sodium hydroxide (NaOH), it undergoes disproportionation and forms different products according to the temperature of the reaction.

13.10.1 Reaction with cold aqueous sodium hydroxide

Chlorine undergoes disproportionation when it reacts with cold aqueous sodium hydroxide (NaOH) forming sodium chloride (NaCl) and sodium chlorate (I) (NaClO).

$$Cl_{2(g)} + 2NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + NaClO_{(aq)} + H_2O_{(\ell)}$$

Oxidation states of chlorine in the above reaction are given below:

Chlorine in Cl_2 = 0
Chlorine in NaCl = -1
Chlorine in NaClO = +1

Keep in Mind!

Sodium chlorate (I) is previously called sodium hypochlorite.

Chlorine in Cl_2 is oxidized from 0 to +1 in NaOCl. Chlorine in Cl_2 is reduced from 0 to 1 in NaCl.

The above reaction shows the simultaneous oxidation and reduction of chlorine and is example of disproportionation reaction.

13.10.2 Reaction with hot aqueous sodium hydroxide

When chlorine reacts with hot aqueous sodium hydroxide (NaOH), it forms sodium chloride (NaCl) and sodium chlorate (V) ($NaClO_3$). It is another example of **disproportionation** reaction.

$$3Cl_{2(g)} + 6NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + NaClO_{3(aq)} + 3H_2O_{(\ell)}$$

Chlorine in Cl_2 = 0

Chlorine in NaCl = -1

Chlorine in $NaClO_3$ = +5

Chlorine is oxidized from 0 to +5 in $NaClO_3$. Chlorine is reduced from 0 to -1 in NaCl

This disproportionation reaction shows how the temperature of the reaction influences the products formed, demonstrating the versatility of chlorine in undergoing redox reactions.

Quick Check 13.5

a) How would KI react with conc. H_2SO_4? What does this reaction indicate about the reducing power of iodide?

b) Show that the reaction of Cl_2 with cold and hot aqueous KOH is a disproportionation reaction.

c) HI acts as strong reducing agent. Explain it with chemical reactions.

13.11 USE OF CHLORINE IN WATER PURIFICATION

Chlorine gas is very poisonous. However, in small quantities, they are harmless to humans but poisonous to the bacteria which cause diseases. Due to its strong disinfectant properties, chlorine is widely used at the treatment plants for water purification. The process involves adding chlorine to water, where it forms active species that kill bacteria and other pathogens. Water in the swimming pools is also chlorinated with slightly higher concentrations of chorine because there is likely to be a higher concentration of bacteria in the water. The primary active species are chloric (I) acid or hypochlorous acid (HOCI) and the chlorate (I) or hypochlorite ions (OCl^-). Chlorination is a relatively inexpensive method of water disinfections.

13.11.1 Chlorine Addition to Water

When chlorine gas (Cl_2) is added to water, it undergoes hydrolysis to form a mixture of hydrochloric acid (HCl) and chloric (I) acid (HClO).

$$Cl_{2(g)} + H_2O_{(\ell)} \longrightarrow HCl_{(aq)} + HClO_{(aq)}$$

Chloric (I) acid (HOCI) is a weak acid and partially dissociate in water to form hydrogen ions (H^+) and chlorate (I) ion (OCl^-).

13.11.2 Disinfection Activity

HOCl and OCI^- are effective disinfectants, but HOCl is more effective due to its neutral charge. The neutral charge of HOCl allows to penetrate the cell walls of micro-organisms easily. Essential cellular components such as proteins and lipids are oxidized by HOCl and OCI^-, which disrupt the cell function leading to cell death. HOCI and OCI can oxidize and inactivate enzymes that are crucial for survival and replication of bacteria.

Interesting Information

Nucleic acids (DNA and RNA) can be oxidized by HOCI and OCl^- and thus preventing bacteria from replicating and vital cellular functions.

13.11.3 Factors Affecting Disinfection

pH

i) At pH around 6-7.5, HOCI predominates and makes the disinfection process more effective.

ii) At higher pH (above 7.5), OCl^- predominates. It is less effective but still provides disinfection.

Chorine dose

The higher the amount of chlorine, the more effective the disinfection. Sufficient chlorine must be added to get enough HOCI and OCI^- to kill bacteria.

Contact time

Contact time of water with chlorine must be long enough, to allow the disinfectants to penetrate and kill bacteria, viruses and protozoa.

Quick Check 13.6

a) Why HOCl is more effective disinfectant than OCl^- to kill bacteria in water?

b) What are the factors that affect disinfection of bacteria in water?

c) What are the primary active species in the chlorination of water? Give equation that shows their formation.

EXERCISE

MULTIPLE CHOICE QUESTIONS

Q.1 Four choices are given for each question. Select the correct choice.

I. Which halogen molecule has the strongest bond?

a) F_2 b) Br_2
c) I_2 d) Cl_2

II. The volatility of the halogens (Group 17) generally __________ as you move down the group (from Fluorine to Iodine).

a) Increases b) Decreases
c) Remains the same d) Fluctuates unpredictably

III. Which one of the following halogen molecules has strongest oxidizing power?

a) Br_2 b) F_2
c) I_2 d) Cl_2

IV. The decreasing thermal stability of the halogens down the group is primarily due to the:

a) Increasing electronegativity of the atoms.
b) Decreasing bond length between the halogen atoms.
c) Increasing atomic radius, leading to a weaker covalent bond.
d) Increasing strength of van der Waals forces

V. Which one of the following halides has strongest reducing power?

a) F^- b) Cl^-

b) Br^- d) I^-

VI. Which statement about the reaction between halogens and hydrogen is correct?

a) Iodine reacts most vigorously with hydrogen.
b) Chlorine and hydrogen explode in darkness.
c) Fluorine combines explosively with hydrogen even in cold and dark conditions.
d) Bromine and hydrogen do not react at all.

VII. How does the acidic strength of hydrogen halides change as you move down the group?

a) It remains constant. b) It decreases from HF to HI.

c) It increases from HF to HI. d) It fluctuates erratically.

VIII. Why is fluorine the most reactive halogen?

a) Bond length in the halogen molecule

b) Bond strength in the halogen molecule

c) Electronegativity of the halogen

d) Number of electrons in the halogen molecule

IX. When aqueous silver nitrate is added to a solution containing bromide ions, a cream precipitate forms. What is the solubility of this precipitate in aqueous ammonia solution?

a) Soluble in dilute ammonia solution

b) Partially soluble in dilute ammonia solution

c) Insoluble in dilute ammonia solution

d) Soluble only upon heating with ammonia

X. Concentrated sulfuric acid is added to solid sodium chloride. What is the initial observation?

a) Reddish-brown fumes are evolved.

b) A purple vapor is evolved.

c) Steamy white fumes of hydrogen chloride are evolved.

d) A black solid is formed.

SHORT ANSWER QUESTIONS

Q.2 Attempt the following short-answer questions:

a. Which halogen is the least reactive, which is the most reactive? Give reason.

b. The ionic equation for a reaction is:

$$Br_2 + 2I^- \rightarrow Br^- + I_2$$

Explain which species is oxidized in this reaction. Why?

c. What is role of London dispersion forces in the trend of volatility of halogens?

d. How does the reactivity of halogens with hydrogen vary?

e. Which halogen is used as an antiseptic? How does it work?

f. What is the colour change when chlorine displaces bromine?

g. How the halogen acids are ionized in water?

h. Why HF is weaker acid than HCl?

i. Describe a simple chemical test that could be used to distinguish between aqueous solutions of potassium bromide and potassium iodide. Include the reagents and expected observations.

j. Explain the chemical principles behind the use of chlorine as a disinfectant in water purification. Include relevant chemical equations in your explanation.

k. Describe one significant disadvantage associated with the use of chlorine in water purification.

l. What is disproportionation reaction? Give an example.

m. Chlorine gas reacts differently with sodium hydroxide solution depending on the temperature and concentration.

n. Write balanced chemical equations for the reaction of chlorine (Cl_2) with:

i) Cold, dilute sodium hydroxide (NaOH).

ii) Hot, concentrated sodium hydroxide (NaOH).

o. For each reaction in question l), identify the oxidation states of chlorine in the reactant (Cl_2) and in each of the chlorine-containing products. Use these oxidation states to explain why both reactions are classified as disproportionation reactions.

DESCRIPTIVE QUESTIONS

Q.3 Describe and explain the relative thermal stabilities of the halogen hydrides in terms of bonds strength.

Q 4. Discuss the relative reactivity of the halogen elements as oxidizing agents. Arrange F_2, Cl_2, Br_2, I_2 in increasing order of the oxidizing power.

Q 5. Discuss the reducing power of halide ions with relevant reactions. Also explain the factors affecting it.