

# 10. THE HALOGENS AND THEIR COMPOUNDS

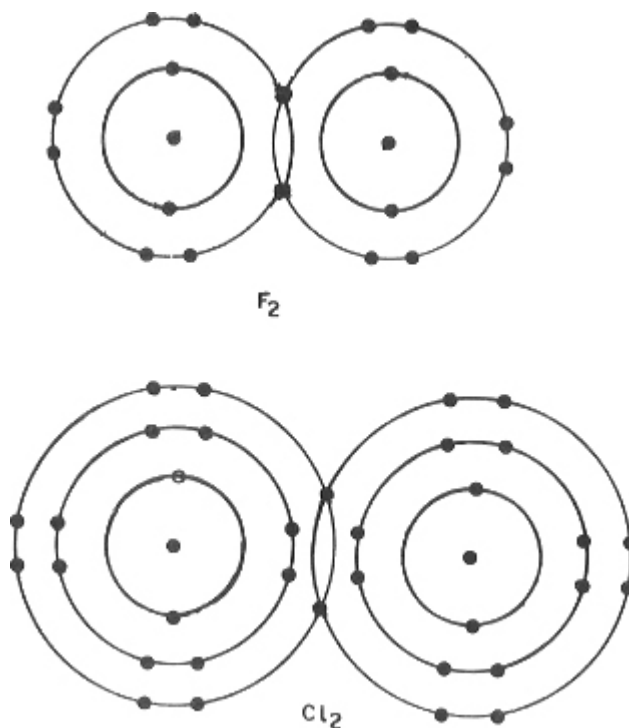
## 10.1 Introduction to the Halogens

In Chapter 1 we discussed the similarities in properties among the halogens as members of the same group of the Periodic Table. We saw then that their similarity in chemical properties was due to their similar electronic configurations. Some physical properties of the halogens are listed in Table 10.1

**TABLE 10.1 PHYSICAL PROPERTIES OF HALOGENS**

Halogen	Atomic number	Electronic configuration	M.P. (°C)	B.P. (°C)	Appearance	Solubility
Fluorine	9	2,7	-223	-187	Yellow gas	Soluble in water.
Chlorine	17	2,8,7	-102	-34.6	Greenish-yellow gas	Soluble in water.
Bromine	35	2,8,18,7	-7.3	58.8	Reddish-brown liquid	Fairly soluble in water
Iodine	53	2,8,18,18,7	114	183	Grey solid	Sparingly soluble in water.

The halogens exist as covalently bonded diatomic molecules (Figure 10.1),

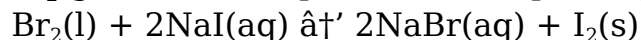
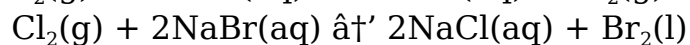
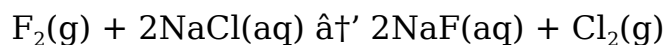


**Figure 10.1 Molecules of fluorine and chlorine.**

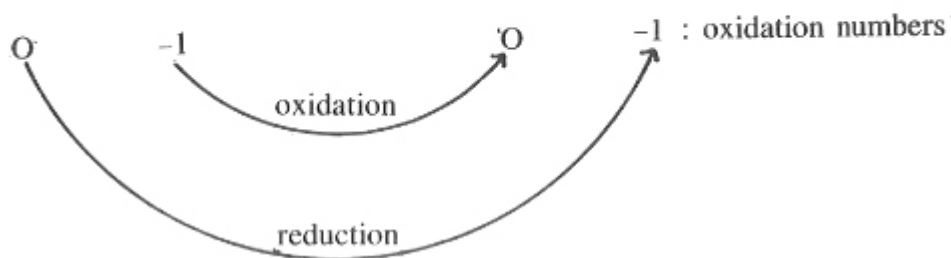
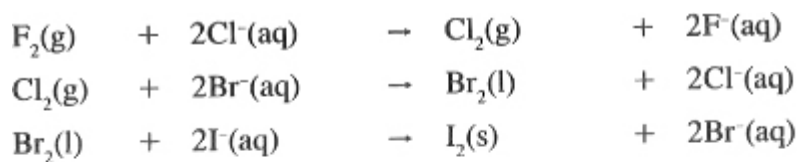
The deepening colours from fluorine, through chlorine, bromine to iodine, and the change in physical state, from gaseous fluorine and chlorine, to liquid bromine and then solid iodine, are attributed to the increasing strength of the intermolecular Van der Waals<sup>™</sup> forces. The more the number of electrons in the molecules, the stronger is this force.

Among the halogens, reactivity decreases with increasing atomic number. Fluorine is thus the most reactive. Their characteristic reaction is oxidation, because they readily form negative ions by gaining electrons from metals. They therefore combine directly with metals.

A halogen displaces the one below it in the group from a solution of its salt. The displacement reaction is an indication of their relative reactivities.



Since the oxidation state of the halogens in their elemental state (the diatomic molecules) is zero, and -1 in the metallic halides, these displacement reactions are redox reactions in which the free halogen is reduced to the corresponding halide ion.



## 10.2 Chlorine

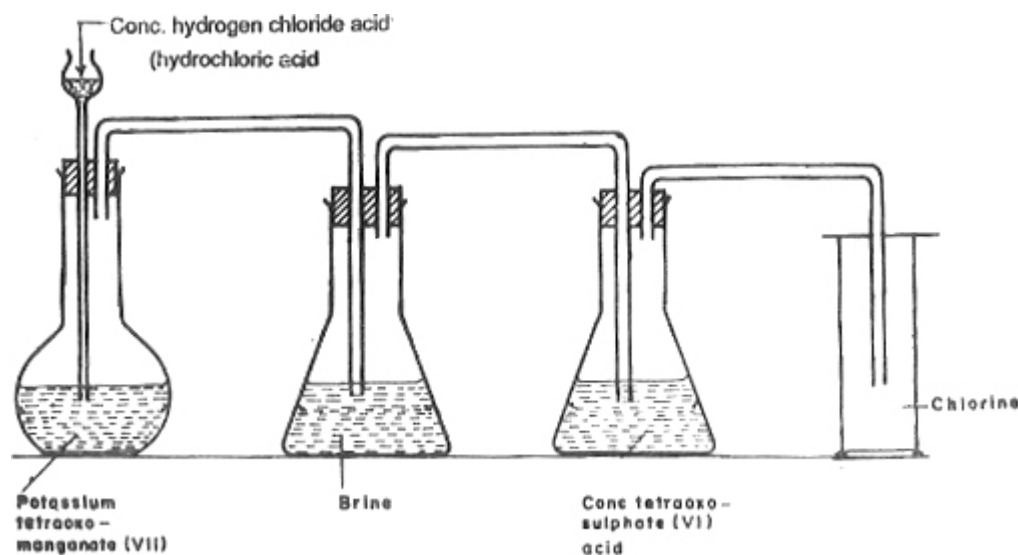
Chlorine occurs naturally in combination with metals, e.g., as sodium chloride, magnesium chloride, etc.

*Experiment 10.1: Preparation of chlorine from hydrogen chloride acid (hydrochloric acid)*

Put some crystals of potassium tetraoxomanganate(VII) into a flat-bottomed flask. Fit a cork carrying a delivery tube and a thistle funnel to the flask. Pour concentrated hydrogen chloride acid (hydrochloric acid) through the thistle funnel into the flask. Effervescence occurs and a misty gas is liberated. Pass the gas liberated through saturated sodium chloride solution to remove hydrogen chloride vapours which escape along with the chlorine, then through concentrated tetraoxosulphate(VI) acid to dry the gas (Figure 10.2). Collect the gas by downward delivery (upward displacement of air).

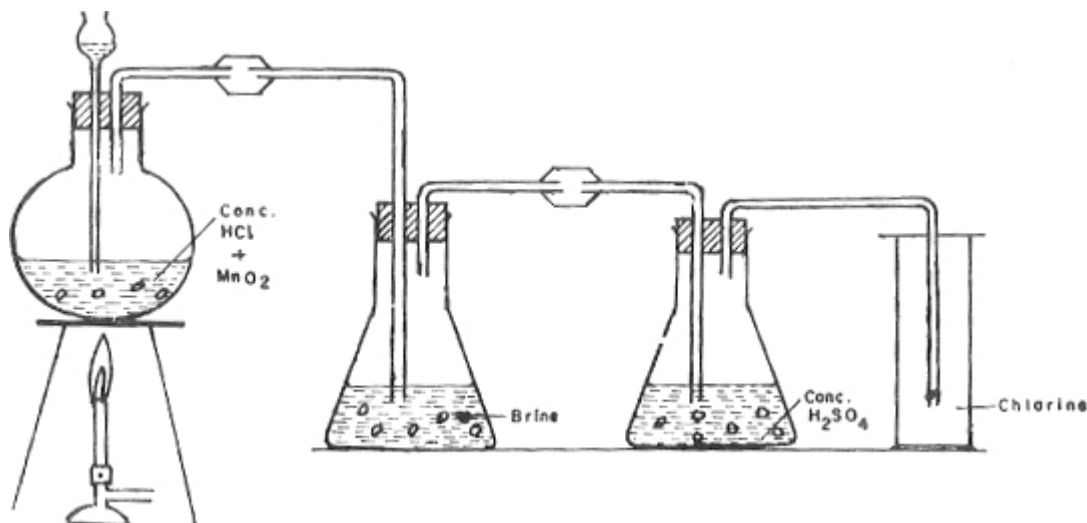


Collect several gas jars for use to study the properties of chlorine later.



**Figure 10.2 Laboratory preparation of chlorine from hydrogen chloride acid (hydrochloric acid) and potassium tetraoxomanganate(VII)**

Alternatively, the concentrated hydrogen chloride acid (hydrochloric acid) could be heated with manganese(IV) oxide in a round-bottomed flask as in Figure 10.3, to liberate chlorine gas, which is passed through brine to absorb hydrogen chloride, and through concentrated tetraoxosulphate(VI) acid, to dry it.



**Figure 10.3 Laboratory preparation of chlorine from hydrogen chloride acid (hydrochloric acid) and manganese(IV) oxide**

## 10.3 Properties of Chlorine

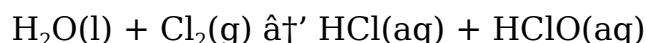
*Experiment 10.2: Investigating the Properties of Chlorine.*

1. Note the colour of chlorine. It is a greenish yellow gas.

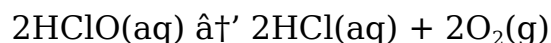
2. It is collected by upward displacement of air because it is denser than air. It is about twice as dense as air.
3. Hold a wet blue litmus paper near the mouth of an open gas jar of chlorine. The litmus first turns red, then becomes bleached. This shows the bleaching action of chlorine.
4. Shake one of the gas jars of chlorine with water, then test the resulting solution with blue litmus. The litmus turns red and is then bleached, showing that chlorine dissolves in water to form an acidic solution which has a bleaching action.

Pour the solution into a beaker. Fill a small test-tube with it and invert the test-tube inside the beaker (Figure 10.4). Then expose the test-tube to sunlight. A gas soon collects in the test-tube. Test this gas with a glowing splint. The splint is rekindled, showing the gas to be oxygen.

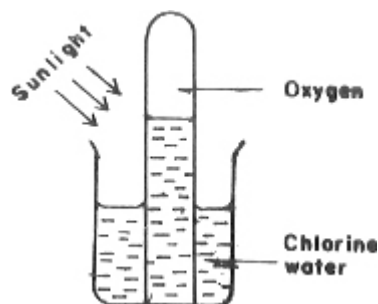
Chlorine forms two acids in water, hydrogen chloride acid (hydrochloric acid) and oxochlorate(I) acid.



The oxochlorate(I) acid readily decomposes into hydrogen chloride acid (hydrochloric acid) and oxygen.

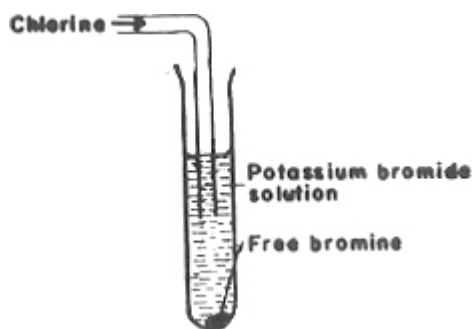
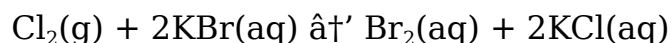


The oxygen formed collects in the test-tube.



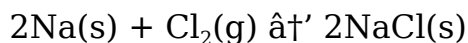
**Figure 10.4 Chlorine water exposed to sunlight**

5. Pass chlorine into potassium bromide solution contained in a test-tube. The colourless potassium bromide solution turns brown. The solution now contains free bromine displaced by chlorine.



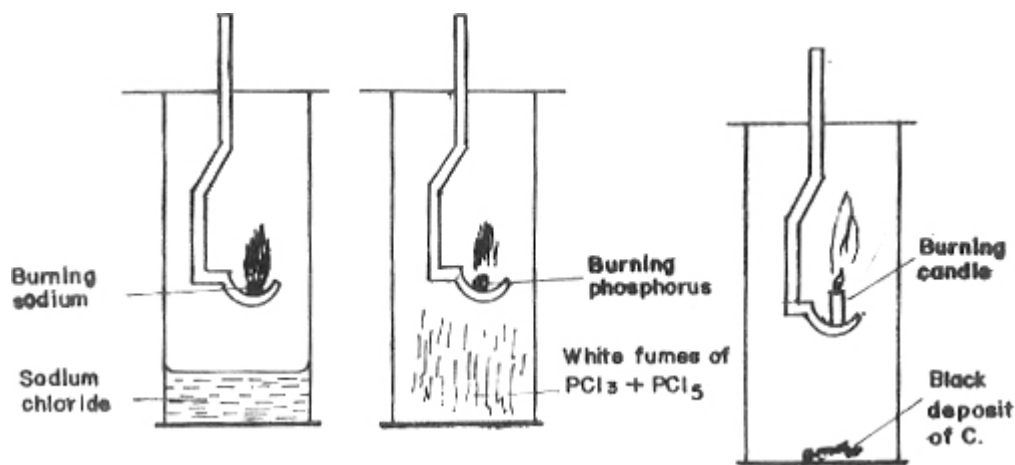
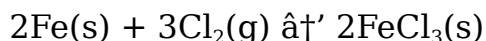
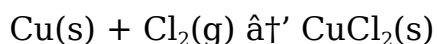
### Figure 10.5 Displacement of bromine by chlorine

6. Lower heated sodium metal in a deflagrating spoon into a gas jar of chlorine. The sodium burns spontaneously, producing sodium chloride.



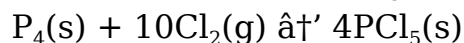
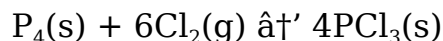
Burning magnesium continues to burn in a gas jar of chlorine, forming magnesium chloride. Most other metals combine directly with chlorine.

If a metal exhibits two oxidation states, the chloride of the higher one is usually formed. Dutch metal, an alloy of copper and a little amount of zinc, burns with a green flame in chlorine, forming copper(II) chloride. Iron forms iron(III) chloride when chlorine is passed over heated iron.

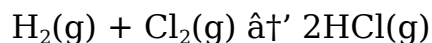


### Figure 10.6 Burning sodium, phosphorus and candle in chlorine

7. Put a small piece of yellow phosphorus into a deflagrating spoon, then lower it into a gas jar of chlorine. It burns to produce dense white fumes of phosphorus(III) chloride and phosphorus(V) chloride.



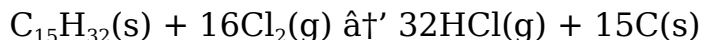
Many other non-metals also combine directly with chlorine on heating. Burning hydrogen continues to burn in a jar of chlorine in the dark. In the presence of light, the reaction is explosive.



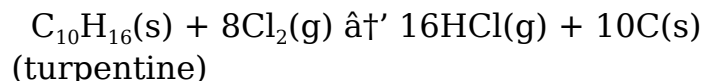
Carbon, nitrogen and oxygen, however, do not combine directly with chlorine.

8. Stand a candle in a deflagrating spoon and light it. Lower the burning candle into a gas jar of chlorine. A red flash of light is observed. White fumes of hydrogen chloride are formed, and a

black deposit of carbon is left in the gas jar after the burning.



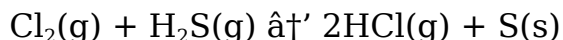
Filter paper soaked in warm turpentine will also burn in a gas jar of chlorine with a red flash of light.



*Experiment 10.3: Investigating the oxidizing property of chlorine.*

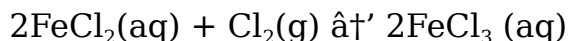
Although in most of its reactions chlorine acts as an oxidizing agent, only some of these reactions are used to demonstrate this characteristics because of the prominent visible changes which occur during such reactions. For example:

1. Invert a gas jar of hydrogen sulphide over a gas jar of chlorine and remove the gas jar covers for both gases to mix. A yellow deposit of sulphur and white fumes of hydrogen chloride are formed.



Chlorine oxidises hydrogen sulphide to sulphur. At the same time hydrogen sulphide reduces chlorine to hydrogen chloride. The yellow deposit of sulphur is evidence of this redox reaction (Figure 10.7).

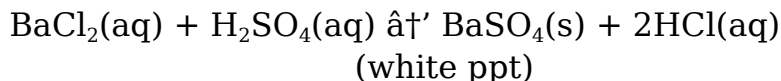
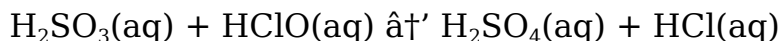
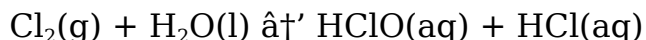
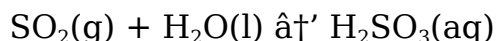
2. Bubble chlorine into iron(II) chloride solution. The light-green solution turns reddish-brown as iron(II) is oxidised to iron(III).



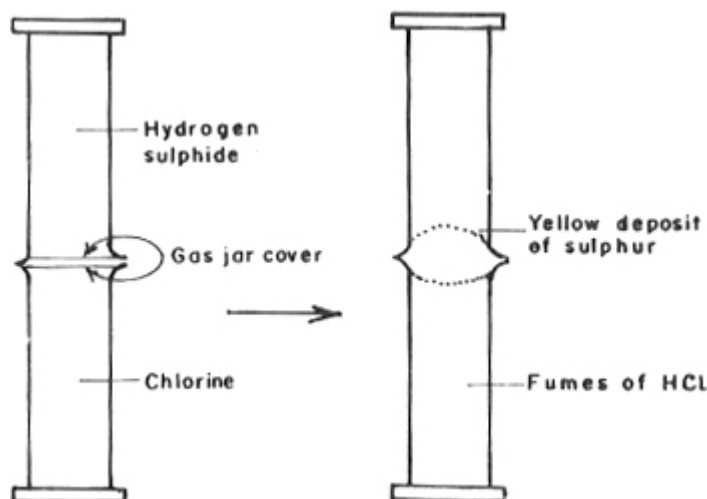
The colour change is evidence of the oxidation.

3. Bubble sulphur(IV) oxide into a beaker of water till it is saturated. Then bubble chlorine into the same water. Put a little of the resulting solution into a test-tube and add some hydrogen chloride acid (hydrochloric acid), followed by barium chloride solution. A white precipitate of barium tetraoxosulphate(VI) is formed.

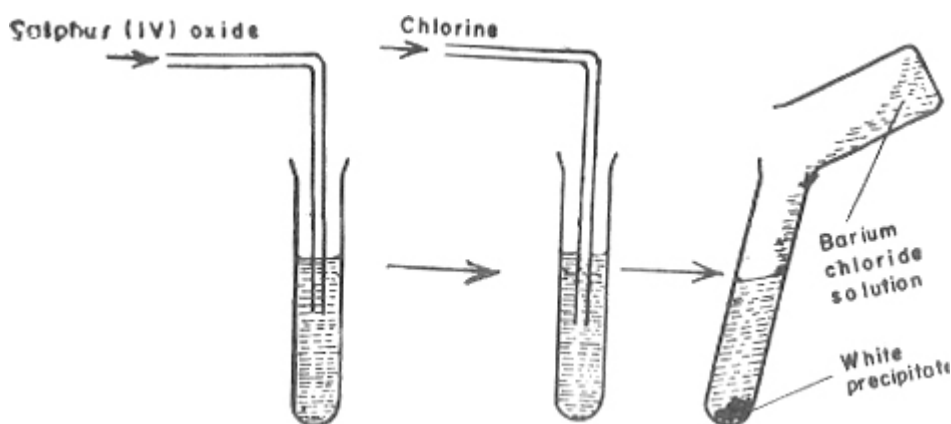
Sulphur(IV) oxide dissolves in water to form trioxosulphate(IV) acid. Chlorine oxidises this acid to tetraoxosulphate(VI) acid. With a barium chloride solution acidified with hydrogen chloride acid (hydrochloric acid) (test for the tetraoxosulphate(VI) ion), white precipitate of barium tetraoxosulphate(VI) is formed.



Without the oxidation caused by chlorine, barium trioxosulphate(IV) dissolves in hydrogen chloride acid (hydrochloric acid), and there is no precipitate formed.



**Figure 10.7 Action of chlorine on hydrogen sulphide**



**Figure 10.8 Oxidation of trioxosulphate(IV) acid by chlorine**

## **Oxochlorates(I) and trichlorates(V)**

Chlorine is used industrially as an oxidizing agent. The chlorine gas for this purpose is obtained as required from oxochlorates(I) of sodium or calcium. Dilute acids, including carbonate(IV) acid formed when atmospheric carbon(IV) oxide dissolves in water vapour present in air, can release chlorine from these chlorates(I).

*Experiment 10.4: Preparation and investigation of the bleaching action of sodium chlorate (I).*

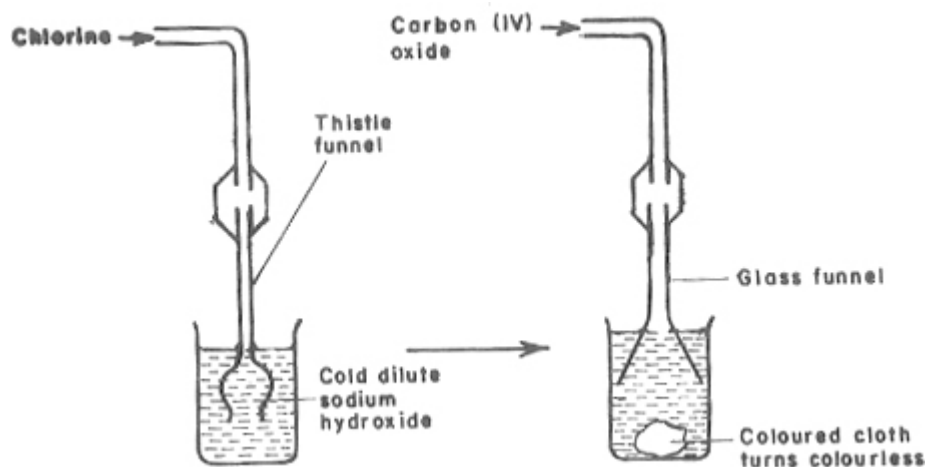
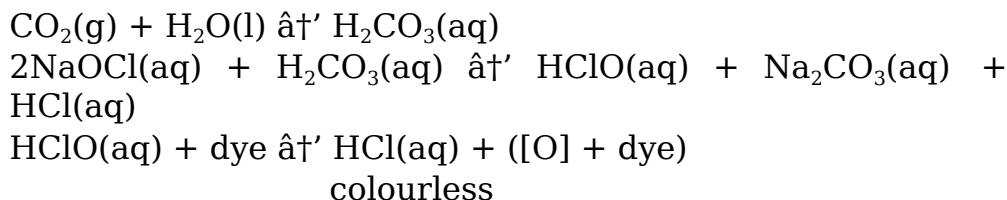
Bubble chlorine into cold dilute sodium hydroxide solution, using the apparatus in Figure 10.9. A pale-yellow solution with a characteristic disinfectant odour is formed. This is a solution of sodium oxochlorate(I) and sodium chloride.



Dip a piece of coloured cloth into the solution, then bubble carbon(IV)



oxide into it. The cloth is bleached. The carbon(IV) oxide forms trioxocarbonate(IV) acid which provides hydrogen for displacing the sodium in sodium chlorate(I). The oxochlorate(I) acid so formed decomposes readily, donating oxygen to the dye of the cloth which then becomes colourless.



**Figure 10.9: Preparation and bleaching action of sodium oxochlorate(I)**

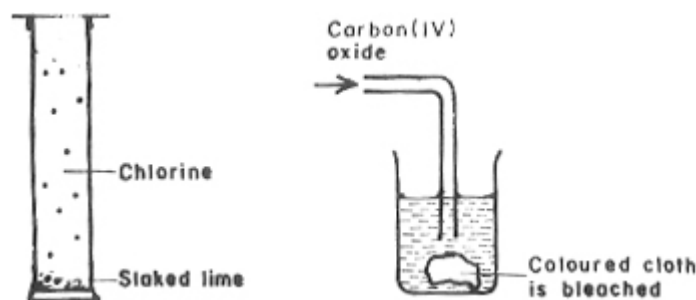
If slaked lime is used instead of sodium hydroxide solution, chlorine forms calcium oxochlorate(I), known as **bleaching powder**. The bleaching action of this compound is similar to that of sodium oxochlorate(I).

*Experiment 10.5: Investigating the Preparation and Bleaching Action of Bleaching Powder.*

Collect a gas jar of chlorine, then put a few drops of slaked lime into it. Shake the gas jar vigorously. The slaked lime absorbs chlorine, forming bleaching powder.



Shake the bleaching powder in water. Put a piece of coloured cloth into the water, then bubble carbon(IV) oxide into the water. The cloth is bleached.



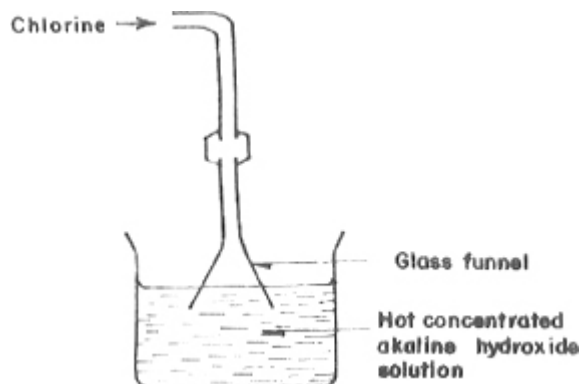
**Figure 10.10 Preparation and bleaching action of bleaching powder**

## Potassium trioxochlorate(V)

Hot concentrated alkaline hydroxide solution forms trioxochlorates(V) salts with chlorine gas.

*Experiment 10.6: Preparation of potassium trioxochlorate(V),  $KClO_3$*

Prepare a concentrated solution of potassium hydroxide. Heat the solution nearly to boiling, then bubble chlorine gas into it using the apparatus in Figure 10.11. White crystals separate out as chlorine is bubbled into the solution. Cool the solution and decant out the clear solution. Redissolve the crystals in a little hot water and cool again. Potassium trioxochlorate(V) separates out, leaving potassium chloride in solution.



**Figure 10.11 Preparation of potassium trioxochlorate(V)**

Sodium hydroxide, if hot and concentrated, also forms sodium trioxochlorate(V) with chlorine. The trioxochlorates(V) are dangerous chemicals which explode readily. On heating, they readily give off oxygen.



## Summary of the properties of chlorine

1. It is a greenish-yellow gas with a choking odour. It is moderately soluble in water and denser than air.
2. With water, chlorine forms two acids: oxochlorate(I) acid, and hydrogen chloride acid (hydrochloric acid). The oxochlorate(I) acid readily gives up its oxygen atom to a coloured material, thus decolourising it by oxidation. Chlorine water is decomposed by sunlight, liberating oxygen.
3. Chlorine combines very readily with some metals (Cu, Zn, Pb), and others (Na, Mg, Fe) on heating, forming their chlorides. The chlorides of the higher oxidation states are formed if the metals exist in two oxidation states.
4. Burning sulphur, phosphorus and hydrogen continue to burn in a gas jar of chlorine.
5. The following observations demonstrate the oxidizing action of chlorine:
  - i) with ammonia, dense white fumes of ammonium chloride are seen.
  - ii) with hydrogen sulphide, a yellow deposit of sulphur is formed.
  - iii) with hydrocarbons, a red flash is first observed, followed by a cloud of soot.
  - iv) with iron(II) chloride solution, the green colour becomes reddish brown, indicating the formation of iron(III) chloride.
6. Chlorine displaces bromine and iodine from solutions of their ions.
7. With cold dilute alkalis, oxochlorates(I) are formed while with hot concentrated alkalis, trioxochlorates(V) are formed. With slaked lime, bleaching powder, calcium oxochlorate(I) is formed.

## 10.4 Uses of Chlorine

1. The most important use of chlorine is as an oxidizing agent for oxidizing wood pulp in the paper industry, and in the textiles industry.
2. It is also used as a disinfectant to sterilise water.
3. It is used for making some organic chemicals such as trichloromethane (chloroform), tetrachloromethane, chloroethene (vinyl chloride), trichloroethanal (chloral), etc.
4. It is used in making inorganic industrial chemicals such as hydrochloric acid, sodium oxochlorate(I) (for bleaching wood pulp and treating sewage), sodium trioxochlorate(V) (a weed killer), etc.

## 10.5 Industrial Preparation of Chlorine

Chlorine is obtained on a large scale from the electrolysis of brine (solution of sodium chloride), using a mercury cathode. Sodium is liberated at the mercury cathode and forms sodium hydroxide on dissolution in water. Chlorine is liberated at the anode. Because of the possibility of reaction between the two products, the anode and cathode must be separated.

### EXERCISE 10A

Write equations to show the reactions occurring when

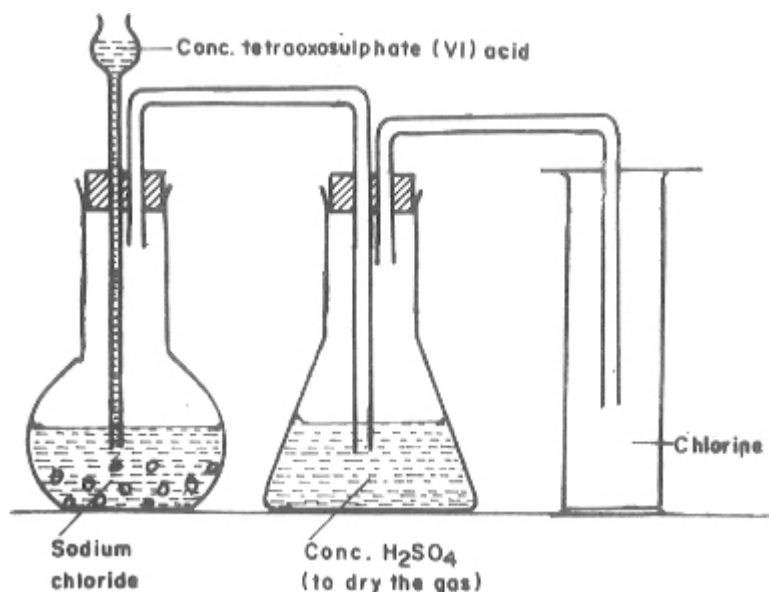
- chlorine is bubbled into cold dilute sodium hydroxide solution.
- chlorine is bubbled into hot, concentrated sodium hydroxide solution.
- slaked lime is shaken in a gas jar of chlorine.
- a gas jar of hydrogen sulphide is inverted over a gas jar of chlorine.
- chlorine is bubbled into iron(II) chloride solution.

What observations (if any) can be made in each case?

## 10.6 Hydrogen Chloride

*Experiment 10.7: Laboratory preparation of hydrogen chloride.*

Set up the apparatus shown in Figure 10.12. Put some sodium chloride into the flat-bottomed flask. Pour concentrated tetraoxosulphate(VI) acid into the flask through a thistle funnel. Pass the gas liberated through concentrated tetraoxosulphate(VI) acid to dry it, then collect by upward displacement of air (downward delivery). Collect five gas jars for use in studying the properties of hydrogen chloride.



**Figure 10.12: Preparation of hydrogen chloride.**

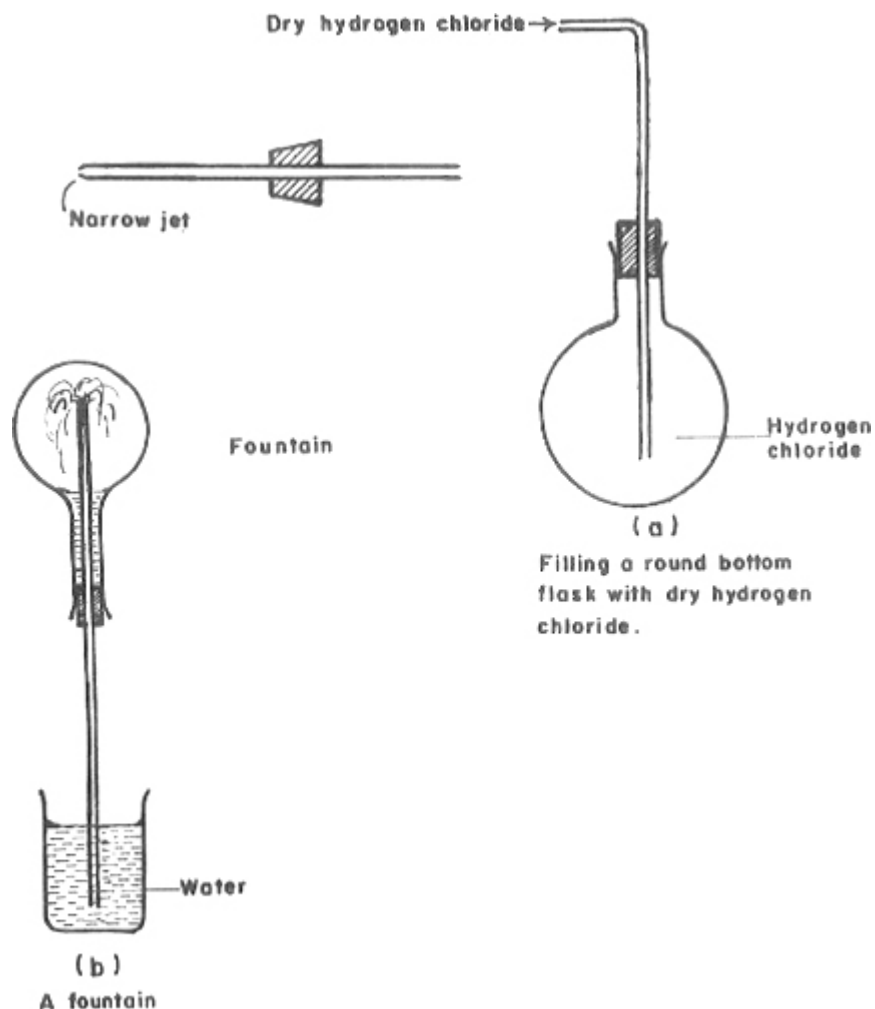
## 10.7 Properties of Hydrogen Chloride

### *Experiment 10.8: Investigating the Properties of Hydrogen Chloride*

1. Note the colour of the gas. It is colourless.
2. Smell (do not inhale) the gas. It has a choking odour.
3. The gas is denser than air. That is why you can collect it by upward displacement of air.
4. Collect some of the dry gas in a round-bottomed flask. Fit the flask with a cork carrying a delivery tube drawn to a narrow jet at the end of the tube inside the flask. Invert the flask over water in a beaker as shown in Figure 10.13 (b). A fountain of water is seen at the narrow jet end of the delivery tube inside the flask.

This experiment, known as the **fountain experiment**, demonstrates the high solubility of hydrogen chloride in water.

Water rushes up the delivery tube into the flask to occupy the space created by the dissolution of the gas which was in the flask.

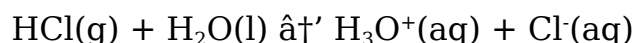


### Figure 10.13 Fountain experiment with hydrogen chloride

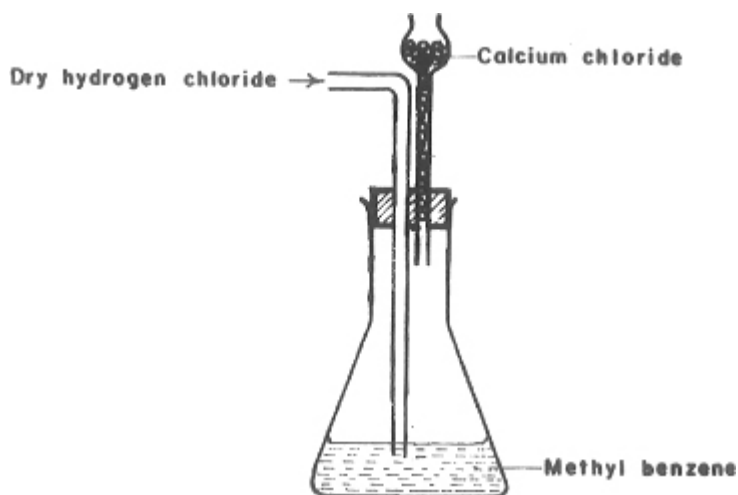
Test the solution in the flask with blue litmus paper. The litmus paper turns red, showing that an acid is formed when hydrogen chloride dissolves in water.

5. Make a solution of dry hydrogen chloride in dry methylbenzene (toluene) by passing the gas into methylbenzene, using the apparatus shown in Figure 10.14. Test the resulting solution with blue litmus paper. There is no change in its colour.

This shows that hydrogen chloride is acidic only in water. Water is a polar solvent. The hydrogen chloride molecule gives up a proton to the oxygen atom of a water molecule, forming a hydroxonium ion which is responsible for the acidity of the solution.

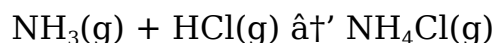


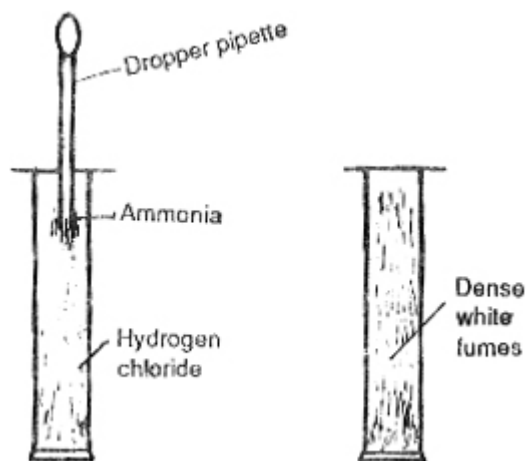
Methylbenzene does not cause the ionization of hydrogen chloride because it is non-polar.



### Figure 10.14 Dissolving hydrogen chloride in toluene

6. Put a drop of concentrated ammonia solution into a gas jar of hydrogen chloride. White fumes of ammonium chloride are formed. Ammonia, an alkaline gas, reacts with hydrogen chloride, an acidic gas, to form ammonium chloride.

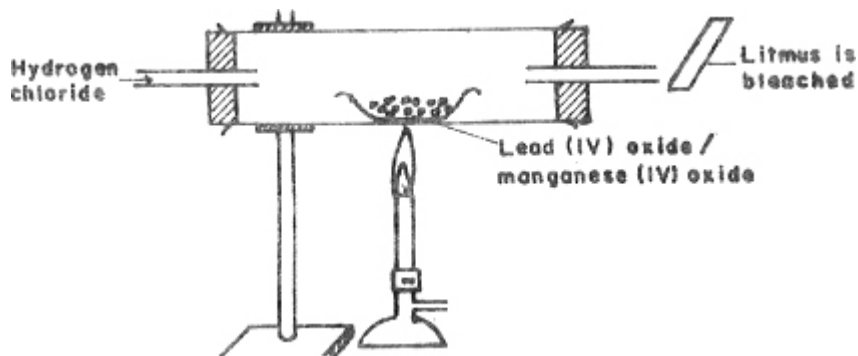




**Figure 10.15 Action of ammonia on hydrogen chloride**

7. Pass the gas over heated lead(IV) oxide, or manganese(IV) oxide in a porcelain boat placed inside a combustion tube. Heat the combustion tube. Test the gas at the opposite end of the tube with wet blue litmus paper. The litmus paper turns red and is then bleached, showing that the gas coming over is chlorine.

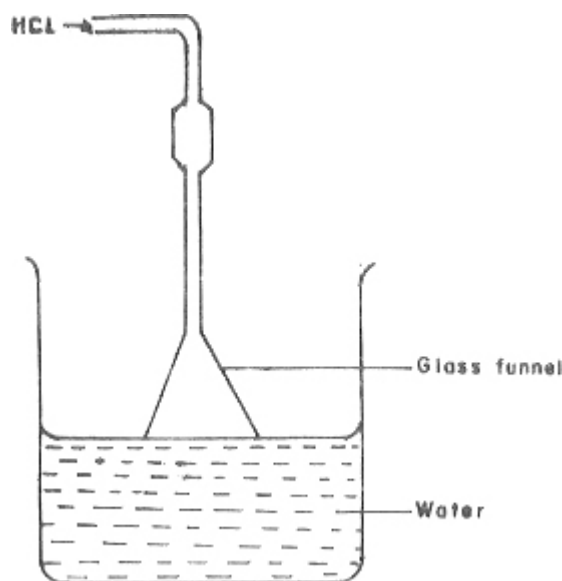
Oxidizing agents such as lead(IV) oxide, oxidise hydrogen chloride to chlorine.



**Figure 10.16 Action of oxidizing agents on hydrogen chloride**

## 10.8 Hydrogen Chloride Acid (Hydrochloric Acid)

Hydrogen chloride acid (hydrochloric acid) is a solution of hydrogen chloride gas in water. If hydrogen chloride is bubbled into water through a delivery tube, there is sucking back because of the high solubility of the gas in water. The device shown in Figure 10.17 is used for the dissolution to avoid such sucking back.



**Figure 10.17 Dissolving hydrogen chloride gas in water**

Hydrogen chloride acid (hydrochloric acid) shows all the properties of a strong acid. In addition, oxidizing agents such as potassium tetraoxomanganate(VII) and manganese(IV) oxide, liberate chlorine from it (see preparation of chlorine).

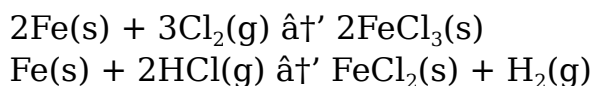
## 10.9 Chlorides

Salts of hydrogen chloride acid (hydrochloric acid) are called chlorides.

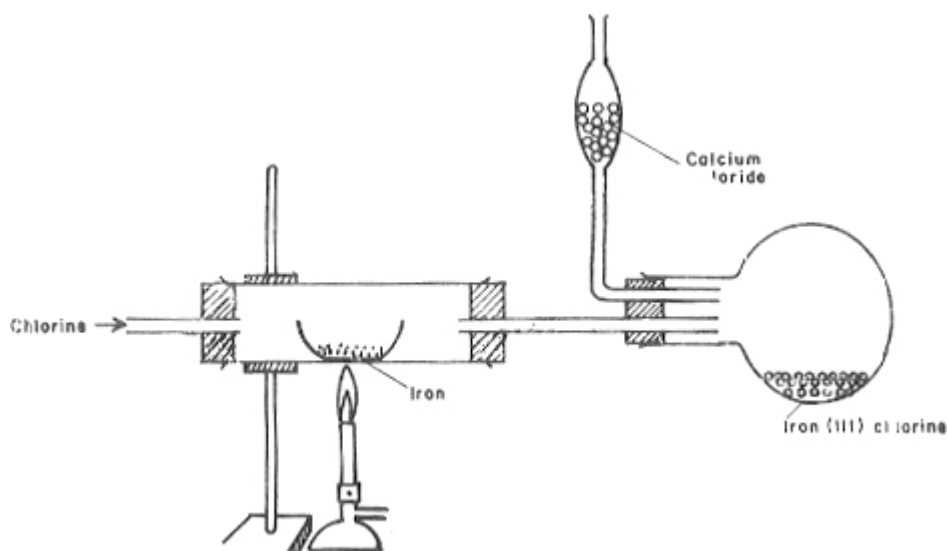
*Experiment 10.9: Preparation of anhydrous iron (II) chloride and iron(III) chloride.*

THIS EXPERIMENT SHOULD BE PERFORMED IN THE FUME CUPBOARD.

Pass chlorine gas over heated iron in a combustion tube. The vapour of the iron(III) chloride formed is carried over to the flask fitted at the opposite end of the combustion tube (Figure 10.18), where it solidifies. If hydrogen chloride is used in place of chlorine in this experiment, iron(II) chloride is formed.



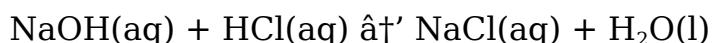




**Figure 10.18 Preparation of anhydrous iron(III) chloride**

Other metallic chlorides are prepared by

1. **neutralisation:**



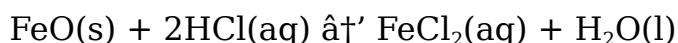
An indicator is used to determine the exact quantity of acid solution to add to the alkaline solution. The dilute salt solution produced is concentrated by evaporation, then cooled for crystals to form.

2. action of hydrogen chloride acid (hydrochloric acid) on trioxocarbonate(IV):

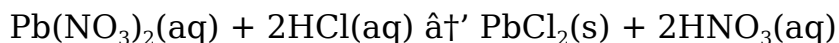


The solution of calcium chloride is either evaporated to dryness, or concentrated and crystallised.

3. **action of hydrogen chloride acid (hydrochloric acid) on metallic oxides:**



4. **double decomposition (for insoluble chlorides):**



The solid lead chloride is filtered off.

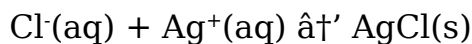
Chlorides of most metals are soluble in water. The only common insoluble ones are lead(II) chloride (soluble in hot water), silver chloride, and mercury(I) chloride.

## Test for the chloride ion

The insolubility of silver chloride is employed in testing for the chloride ion.

*Experiment 10.10: Test for chloride ion.*

To solutions of hydrogen chloride acid (hydrochloric acid), trioxonitrate(V) acid, tetraoxosulphate(VI) acid, sodium chloride, sodium tetraoxosulphate(VI), calcium trioxonitrate(V) and calcium chloride in separate test tubes, add a few drops of silver trioxonitrate(V) solution, then trioxonitrate(V) acid. Those solutions which contain the chloride ion, hydrogen chloride acid (hydrochloric acid), calcium chloride and sodium chloride) form white precipitates of silver chloride, insoluble in the trioxonitrate(V) acid.



If a precipitate is formed in any of the other solutions, it dissolves on the addition of trioxonitrate(V) acid, leaving a clear solution.

## Use of concentrated hydrogen chloride acid (hydrochloric acid) in the flame test

Chlorides are the most volatile salts. Concentrated hydrogen chloride acid (hydrochloric acid) is therefore used in the flame test to form chlorides of the metals. These easily vapourise in the bunsen flame to give the characteristic colour of the vapour of each metal.

*Experiment 10.11: Flame test.*

Dip a platinum wire into concentrated hydrogen chloride acid (hydrochloric acid). With the wet wire pick a speck of each of the following salts in turns. Heat each salt in the bunsen flame. The salts to use and the colour of the flame observed are shown in Table 10.2.

**TABLE 10.2 FLAME COLOURATION OF CATIONS**

<b>Salt</b>	<b>Metal ion Colour of flame</b>	
Sodium chloride	Sodium	Golden yellow
Potassium chloride	Potassium	Lilac blue
Copper tetraoxosulphate(VI)	Copper	Green
Calcium chloride	Calcium	Brick red

### EXERCISE 10B

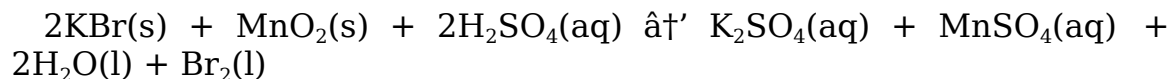
1. With suitable examples, list four methods of preparing metallic chlorides.
2. Explain what happens when copper(II) tetraoxosulphate(VI) crystal is picked on a platinum wire which has been dipped into concentrated hydrogen chloride acid (hydrochloric acid), and the crystal heated in a bunsen flame.

## 10.10 Bromine

The reason for the similarity between chlorine and bromine has already been stressed. The two halogens are prepared in the laboratory by similar reactions.

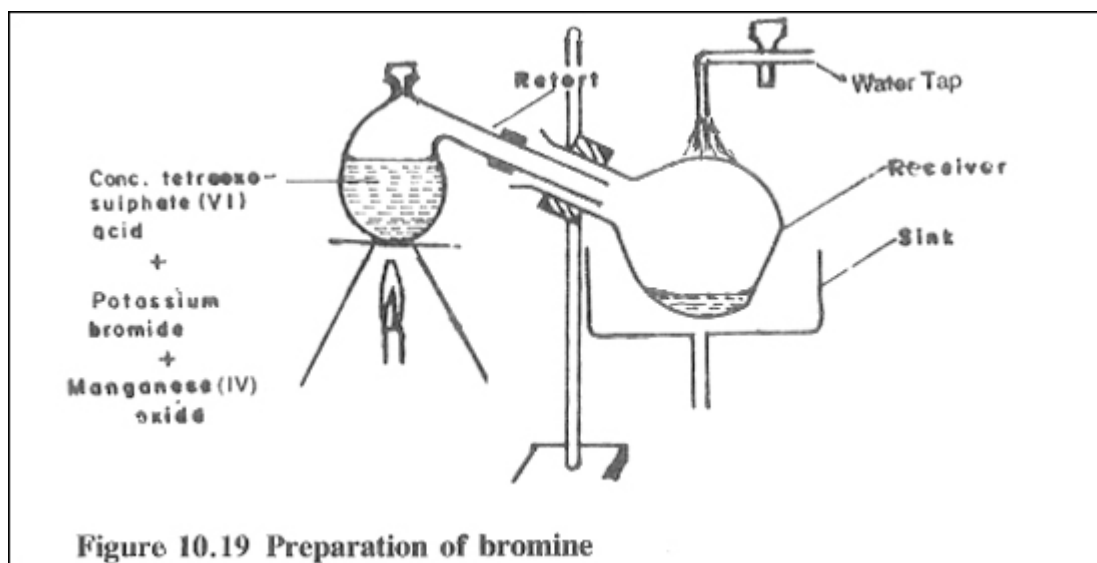
*Experiment 10.12: Laboratory Preparation of Bromine.*

Put a mixture of potassium bromide and manganese(IV) oxide into a glass retort. Add concentrated tetraoxosulphate(VI) acid to it. Connect a receiver under tap water as shown in Figure 10.19. Heat the retort. Reddish brown fumes of bromine are liberated. The fumes condense to liquid bromine in the receiver which is cooled by the running tap water.



The reaction can be broken down into two stages thus:

1.  $\text{H}_2\text{SO}_4\text{(aq)} + \text{KBr(s)} \rightarrow \text{HBr(g)} + \text{KHSO}_4\text{(aq)}$
2.  $4\text{HBr(g)} + \text{MnO}_2\text{(s)} \rightarrow \text{MnBr}_2\text{(aq)} + 2\text{H}_2\text{O(l)} + \text{Br}_2\text{(g)}$



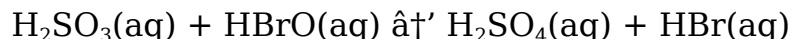
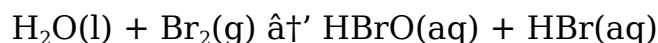
## 10.11 Properties of Bromine

*Experiment 10.13: Investigating the Properties of Bromine.*

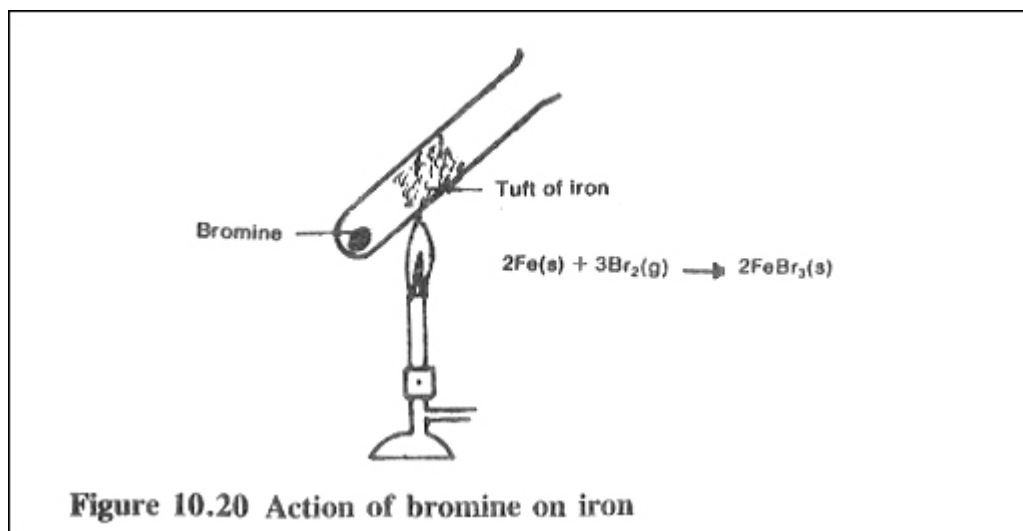
1. Place a piece of wet blue litmus paper in contact with vapours of bromine. The blue litmus is slowly bleached. The bleaching action of bromine is very slow compared to that of chlorine.
2. Bubble sulphur(IV) oxide into water to get trioxosulphate(IV) acid. Add a few drops of bromine to the solution and shake. Warm if the bromine colour is not discharged quickly.

Add barium chloride solution, then hydrogen chloride acid (hydrochloric acid) to the resulting solution. A white precipitate of barium tetraoxosulphate(VI) shows that the trioxosulphate(IV) acid

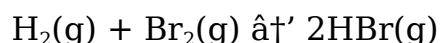
is oxidised to tetraoxosulphate(VI) acid by bromine.



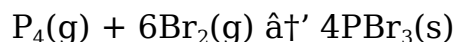
3. Put two drops of bromine into a test-tube, then hang a tuft of iron in the test-tube above the bromine. Heat the iron strongly, occasionally flipping the bunsen flame to the bromine to vapourise it. The bromine vapour produced does not go beyond the tuft of iron, showing that it reacts with the iron.



Many metals and non-metals combine directly with bromine, as with chlorine. Sodium, zinc, iron, tin and copper form their bromides by direct combination. Hydrogen and bromine combine directly, when heated, to form hydrogen bromide.

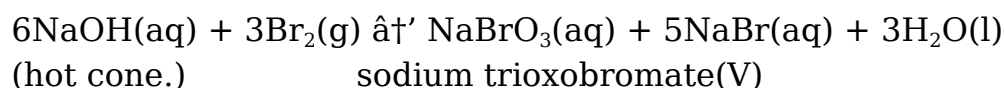
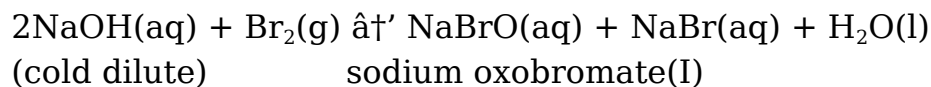


Yellow phosphorus and bromine explode violently, forming phosphorus(III) bromide.



To avoid the explosion, phosphorus(III) bromide is prepared by adding bromine slowly to red phosphorus in tetrachloromethane.

4. Bubble bromine into cold dilute sodium hydroxide solution, and again into hot concentrated sodium hydroxide solution. Compounds similar to those with chlorine are formed.



Bromine resembles chlorine closely in all these reactions, but the

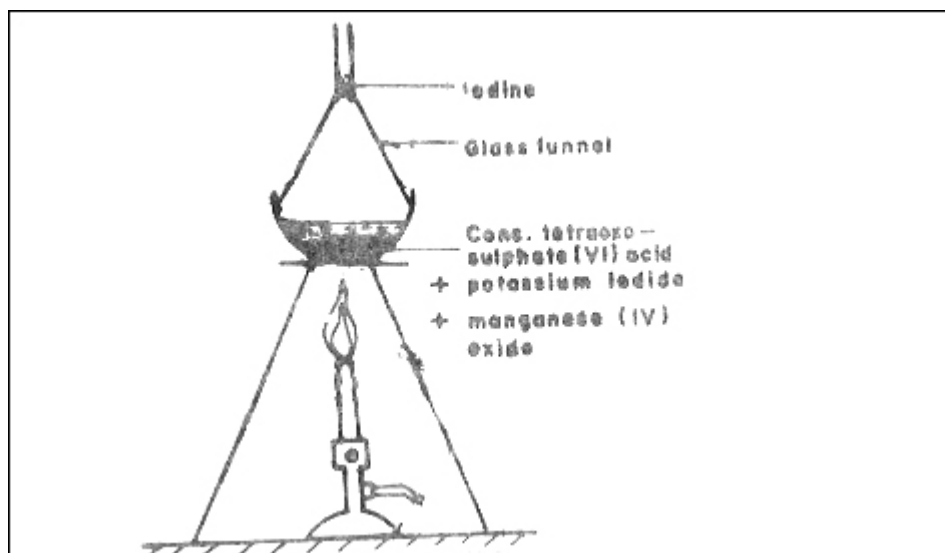
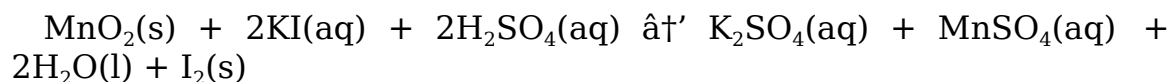
reactions are less vigorous with bromine than with chlorine.

## 10.12 Iodine

Though iodine is a shiny solid, it still resembles the other members of the group VII elements in its chemical reactions.

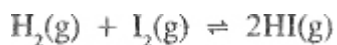
### *Experiment 10.14 Laboratory Preparation of Iodine.*

Add concentrated tetraoxosulphate(VI) acid to a mixture of potassium iodide and manganese(IV) oxide in an evaporating dish with an inverted glass funnel over it, (Figure 10.21). Heat the mixture. Grey shiny crystals of iodine are deposited at the neck of the funnel.

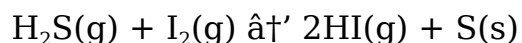


**Figure 10.21 Laboratory preparation of iodine**

The reactivity of iodine is highly reduced. Thus., iodine does not bleach. It reacts with hydrogen only at elevated temperatures, and the reaction is reversible.



However, iodine still combines directly with many metals to form their iodides. It shows limited oxidizing property. For example, it oxidizes hydrogen sulphide to sulphur.



## Uses of iodine

1. The antiseptic action of iodine which is due to its oxidation property is evident in the use of **tincture of iodine** (a solution of

iodine in ethanol) for treating cuts.

2. Iodine is also used in small amounts for treating goitre.

## 10.13 Uses of Halogen Compounds

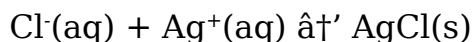
1. **Hydrogen chloride acid (Hydrochloric acid):**
  - (a) For the manufacture of chlorine.
  - (b) For cleaning metal surfaces before galvanization.
2. **Sodium chloride:**
  - (a) For the manufacture of sodium hydroxide and sodium trioxocarbonate(IV).
  - (b) As an important food additive.
  - (c) For preserving food.
3. **Ammonium chloride:** For the manufacture of dry batteries.
4. **Potassium chloride:** Added to common salt to avoid deliquescence.
5. **Aluminium chloride:** As a catalyst in petrochemical industry.
6. **Zinc chloride:** As a flux in soldering.
7. **Tin chloride:** As a mordant in dyeing.
8. **Sodium oxochlorate(I):** For bleaching wood pulp and textiles.
9. **Sodium trioxochlorate(V):** As a weed killer.
10. **Silver chloride and silver bromide:** For making photographic films and plates.

## Chapter Summary

1. All the halogens, (F, Cl, Br, I, At), contain seven electrons in their outermost shells.
2. Chlorides, bromides and iodides of sodium and potassium occur in sea water.
3. The laboratory preparation of chlorine is by the oxidation of a chloride, hydrogen chloride acid (hydrochloric acid). If potassium tetraoxomanganate(VII) is used as the oxidising agent, heating is not necessary. If, however, manganese(IV) oxide is the oxidising agent, heat is required.
4. (a) Chlorine reacts as
  - (i) an acid gas,
  - (ii) a bleaching agent, and
  - (iii) an oxidising agent.(b) (i) It also displaces bromine and iodine from bromides and iodides respectively.  
(ii) It forms oxochlorates(I) with cold dilute alkalis, and

trioxochlorates(V) with hot, concentrated alkalis.

5. Hydrogen chloride is prepared by the action of concentrated tetraoxosulphate(VI) acid on sodium chloride.
6. The gas is acidic, forms dense white fumes with ammonia, and a white precipitate with silver trioxonitrate(V) solution.
7. Test for the chloride ion: It forms a white curdy precipitate with silver trioxonitrate(V) solution acidified with trioxonitrate(V) acid.



8. Silver halides are used in photography because they are sensitive to light.

## Assessment

1. Describe three reactions to show that
  - (a) chlorine and bromine belong to the same group of the periodic table,
  - (b) chlorine is more reactive than bromine.
2. List the uses of chlorine. Explain the oxidation action of chlorine.
3. Describe what is observed and explain what happens when
  - (a) chlorine is passed into iron(II) chloride solution;
  - (b) burning candle is lowered into a gas jar of chlorine;
  - (c) a gas jar of hydrogen sulphide is inverted over a gas jar of chlorine;
  - (d) bromine is bubbled into hot concentrated solution of potassium hydroxide.
  - (e) a mixture of potassium iodide, manganese(IV) oxide and concentrated tetraoxosulphate(VI) is heated in an evaporating dish with a glass funnel inverted over it.
4. Explain the following observations:
  - (a) A piece of red litmus paper held inside a test-tube in which ammonium chloride is heated first turns blue, then red, and finally a white solid forms near the mouth of the test-tube.
  - (b) Solid sodium chloride disappears when dropped into water in a test-tube, and shaken. Addition of silver trioxonitrate(V) solution to the resulting clear liquid causes the appearance of a white solid.
  - (c) A wet litmus paper held near the mouth of a test-tube in which calcium chloride and manganese(IV) oxide are heated together in the presence of a little concentrated tetraoxosulphate(VI) acid first turns red, then becomes colourless.
5. Which of the following reactions is not used to prepare chlorine?
  - A.  $\text{MnO}_2(\text{s}) + 4\text{HCl}(\text{aq}) \rightarrow \text{MnCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) + \text{Cl}_2(\text{g})$

- B.  $2\text{KMnO}_4(\text{s}) + 16\text{HCl}(\text{aq}) \rightarrow 2\text{KCl}(\text{aq}) + 2\text{MnCl}_2(\text{aq}) + 8\text{H}_2\text{O}(\text{l}) + 5\text{Cl}_2(\text{g})$
- C.  $\text{CaOCl}_2(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{Cl}_2(\text{g})$
- D.  $\text{Pb}_3\text{O}_4(\text{s}) + 8\text{HCl}(\text{aq}) \rightarrow 3\text{PbCl}_2(\text{s}) + 4\text{H}_2\text{O}(\text{l}) + \text{Cl}_2(\text{g})$
- E.  $2\text{NaCl}(\text{aq}) + \text{I}_2(\text{aq}) \rightarrow 2\text{NaI}(\text{aq}) + \text{Cl}_2(\text{g})$
6. (a) With the aid of a labelled diagram, describe how you would prepare and collect a few gas jars of dry chlorine gas.
- (b) What happens when chlorine gas is passed into
- a cold dilute solution of sodium hydroxide,
  - a hot concentrated solution of sodium hydroxide?
- (WAEC)
7. Chlorine gas is bubbled through an aqueous solution of sodium bromide.
- Name the products formed
  - Which of the reactants provides the reducing agent?
  - Write the two half-reaction equations for the process
- (WAEC.)
8. Describe what is observed and explain what happens when
- a piece of wet blue litmus paper is introduced into a gas jar of chlorine;
  - a jar of chlorine is placed above a jar of hydrogen sulphide and the lids removed;
  - burning phosphorus in a deflagrating spoon is lowered into a jar of chlorine;
  - a stick soaked in turpentine ( $\text{C}_{10}\text{H}_{16}$ ) is lighted and the burning tapper plunged into a jar of chlorine;
  - hot dutch metal (Cu/Zn alloy) is lowered into a jar of chlorine.
9. Name two other elements which with chlorine, are members of the group VII of the Periodic Table. By what symbols are they known? Describe what happens, and explain the reaction taking place when chlorine is bubbled into (i) potassium iodide solution; (ii) iron(II) chloride solution.