

## INTERCHAPTER Q

# The Halogens



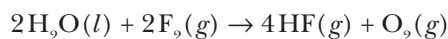
Chlorine, bromine, and iodine.

The elements fluorine, chlorine, bromine, iodine, and astatine are collectively called the **halogens** (Frontispiece). At 25°C fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid. There are no stable (nonradioactive) isotopes of astatine. These nonmetals, which occur in Group 17 of the periodic table, are all highly reactive diatomic molecules that are not found free in nature. They occur primarily as **halide** ( $F^-$ ,  $Cl^-$ ,  $Br^-$ ,  $I^-$ ) salts. The free elements are strong oxidizing agents and combine directly with almost all the other elements. The halogens have pungent, irritating odors and are poisonous.

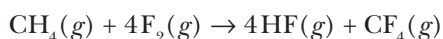
Table Q.1 shows that most of the properties of the halogens vary smoothly from the lightest,  $F_2(g)$ , to the heaviest,  $I_2(s)$ . The relative sizes of the diatomic halogen molecules and the halide ions are shown in Figure Q.1. Note that the size increases with increasing atomic number. Table Q.2 summarizes the principal sources and commercial uses of the halogens.

### Q-1. Fluorine Is the Most Reactive Element

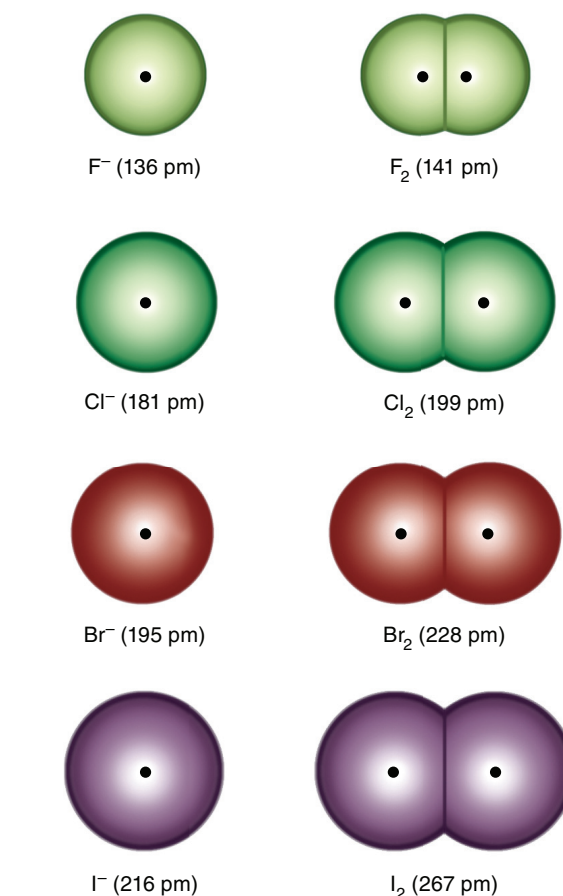
Fluorine is a pale yellow, corrosive gas that reacts directly, and in most cases vigorously, at room temperature with almost every element. It is the most reactive and the strongest oxidizing agent of all the elements. The extreme reactivity of fluorine is evidenced by its reactions with glass, ceramics, and most alloys. Water burns vigorously in fluorine according to



and the light hydrocarbons react spontaneously; for example,



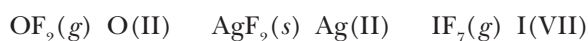
Even xenon and krypton, which were once thought to be completely inert, react with fluorine to form fluo-



**Figure Q.1** Relative sizes of the halide ions (*circles*) and diatomic halogen molecules (*attached pairs*). Distances are in picometers.

rides such as  $KrF_2(s)$ ,  $XeF_2(s)$ ,  $XeF_4(s)$ , and  $XeF_6(s)$  (see Interchapter K).

Because of its high electronegativity, fluorine is capable of stabilizing unusually high ionic charges of other elements. Some examples are



**TABLE Q.1** Physical properties of the diatomic halogen molecules

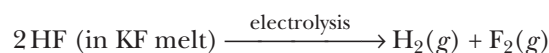
Element	Molecular mass of diatomic species	Melting point/°C	Boiling point/°C	Density at 20°C/g·cm <sup>-3</sup>	Bond length/pm	Bond energy/kJ·mol <sup>-1</sup>
fluorine, $F_2$	37.9968064	-219.67	-188.12	$1.553 \times 10^{-3}$	141.19	158.78
chlorine, $Cl_2$	70.906	-101.5	-34.04	$2.898 \times 10^{-3}$	198.78	242.580
bromine, $Br_2$	159.808	-7.2	58.8	3.1028	228.11	192.807
iodine, $I_2$	253.80894	113.7	184.4	4.933	266.63	151.088

TABLE Q.2 Sources and uses of the halogens

Element	Principal sources	Uses
fluorine	<i>fluorspar</i> , $\text{CaF}_2(s)$	production of $\text{UF}_6(g)$ for the nuclear industry; production of fluorocarbons (Teflon, Freon, etc.)
chlorine	<i>halite</i> , $\text{NaCl}(s)$ ; <i>sylvite</i> , $\text{KCl}(s)$ ; seawater	production of organic compounds; water purification; bleaches; flame-retardant compounds; dyes; textiles; insecticides; plastics
bromine	natural brines; salt lakes and salt beds	production of pesticides; fire-retardant materials; photochemicals; dyestuffs
iodine	brines associated with certain oil well drillings; Chilean deposits of saltpeter; seaweeds	production of organic compounds; iodized salt and tincture of iodine

In effect, fluorine stabilizes high ionic charges in other elements because its electronegativity is so great that even strong oxidizing agents, such as  $\text{Ag(II)}$ , cannot remove electrons from it. Because its electronegativity is greater than that of any other element, fluorine does not occur with a positive ionic charge in any compound. All fluorine compounds are fluorides, where the ionic charge of fluorine is  $-1$ . Fluorides are widely distributed in nature. The main natural sources of fluorine are the minerals fluorapatite,  $\text{Ca}_{10}\text{F}_2(\text{PO}_4)_6(s)$ ; fluorspar,  $\text{CaF}_2(s)$ ; and cryolite,  $\text{Na}_3\text{AlF}_6(s)$  (Figure Q.2). Because of its extreme reactivity, elemental fluorine wasn't isolated until 1886. The French chemist Henri Moissan, who first isolated fluorine, received the 1906 Nobel Prize in Chemistry for his work. Elemental fluorine is obtained by the

electrolysis of hydrogen fluoride dissolved in molten potassium fluoride according to



The modern method of producing  $\text{F}_2(g)$  is essentially a variation of the method first used by Moissan. Prior to World War II, there was no commercial production of fluorine. The atomic bomb project required huge quantities of fluorine for the production of uranium hexafluoride,  $\text{UF}_6(g)$ , a gaseous compound that is used in the separation of uranium-235 from uranium-238. It is uranium-235 that is used in nuclear devices. The production of uranium hexafluoride for the preparation of fuel for nuclear power plants is today a major commercial use of fluorine.

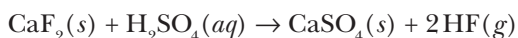


Figure Q.2 Samples of the minerals (left to right) fluorapatite,  $\text{Ca}_{10}\text{F}_2(\text{PO}_4)_6(s)$ , fluorspar,  $\text{CaF}_2(s)$ , and cryolite,  $\text{Na}_3\text{AlF}_6(s)$ .

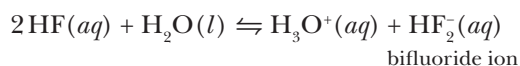


The compound  $^{235}\text{UF}_6(g)$  is separated from  $^{238}\text{UF}_6(g)$  by gaseous effusion. The lighter  $^{235}\text{UF}_6(g)$  effuses more rapidly than does the heavier  $^{238}\text{UF}_6(g)$  (see Section 13-11).

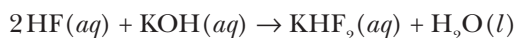
All the halogens form binary hydrogen compounds with the general formula  $\text{HX}(g)$ , where X is one of the halogens. Hydrogen fluoride,  $\text{HF}(g)$ , is a highly irritating, corrosive, colorless gas. It can be prepared by the reaction described by the equation



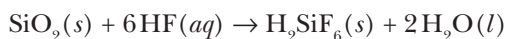
Hydrogen fluoride is used in petroleum refining and in the production of fluorocarbons such as methoxyflurane (an anesthetic) and fluorocarbon polymers such as Teflon (see Interchapter S). Unlike the other **hydrohalic acids**, hydrofluoric acid,  $\text{HF}(aq)$ , is a weak acid in aqueous solution. Furthermore, owing to the formation of strong hydrogen bonds, in aqueous solution it forms bifluoride ions,  $\text{HF}_2^-(aq)$ , according to



The bifluoride ion is important in the chemistry of  $\text{HF}(aq)$ . For example, the reaction of hydrofluoric acid with potassium hydroxide produces potassium bifluoride,  $\text{KHF}_2(aq)$ , according to



The other hydrohalic acids produce only normal salts such as  $\text{KCl}(s)$ . Hydrofluoric acid reacts with glass and so must be stored in plastic bottles. The reaction of hydrofluoric acid with glass is described by the equation



This reaction can be used to etch, or “frost,” glass for lightbulbs and decorative glassware (Figure Q.3). Various fluorides, such as tin(II) fluoride,  $\text{SnF}_2(s)$ , and sodium monofluorophosphate,  $\text{Na}_2\text{PO}_3\text{F}(s)$ , are used as toothpaste additives, and sodium fluoride is added to some municipal water supplies

Tin(II) fluoride is also known as stannous fluoride.



Figure Q.3 Glass can be etched by letting it react with hydrofluoric acid,  $\text{HF}(aq)$ .

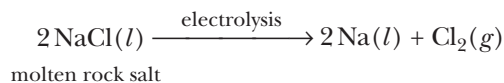
to aid in the prevention of tooth decay. The fluoride converts tooth enamel from hydroxyapatite,  $\text{Ca}_{10}(\text{OH})_2(\text{PO}_4)_6(s)$ , to fluorapatite,  $\text{Ca}_{10}\text{F}_2(\text{PO}_4)_6(s)$ , which is harder and more resistant to acids and decay than hydroxyapatite.

Oxygen is very soluble in numerous liquid fluorocarbons, and this unusual property of fluorocarbons has led to their study for potential use as artificial blood fluids, as a replacement for oxygen mixtures for deep-sea diving, and other applications. Figure Q.4 shows a mouse totally submerged in dichlorofluoromethane saturated with oxygen. The mouse is able to breathe by absorbing oxygen from the oxygen-containing fluorocarbon that fills its lungs.

## Q-2. Chlorine Is Obtained by Electrolysis; Bromine and Iodine by Oxidation

Chlorine is the most abundant of the halogens. The most abundant source of chlorine in nature is the chloride ion found in the oceans and salt lakes. The major mineral sources of chloride are rock salt,  $\text{NaCl}(s)$ , sylvite,  $\text{KCl}(s)$ , and carnallite,  $\text{KMgCl}_3 \cdot 6\text{H}_2\text{O}(s)$ .

Chlorine is prepared commercially by the electrolysis of either brines or molten rock salt:



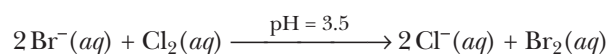
About 12 million metric tons of chlorine are produced annually in the United States, making it one of the top ten industrial chemicals. Most chlorine produced in the United States is used as a bleaching agent in



**Figure Q.4** This submerged mouse is breathing oxygen dissolved in a liquid fluorocarbon. The solubility of oxygen in this liquid is so great that the mouse is able to breathe by absorbing oxygen from the oxygen-containing fluorocarbon that fills its lungs. When the mouse is removed from the liquid, the fluorocarbon vaporizes from its lungs and normal breathing resumes. Liquids such as these may one day enable divers to descend to great depths without pressure suits, may help reduce lung inflammation in premature babies whose lung tissue is not yet fully developed, and may allow astronauts to survive greater forces upon acceleration into space.

the pulp and paper industry. It is also used extensively as a germicide in water purification; in the manufacture of plastics; in solvents used for dry cleaning; in insecticides; in herbicides; and in pharmaceuticals.

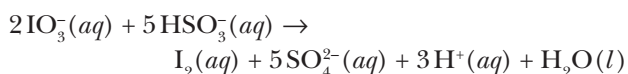
Bromine is a dense, red-brown liquid. Bromine vapor and solutions of bromine in nonpolar solvents are red (Figure Q.5). The major source of bromine in the United States is from brines that contain bromide ions. The pH of the brine is adjusted to 3.5, and chlorine is added; chlorine oxidizes the bromide ion to bromine, which is swept out of the brine with a current of air:



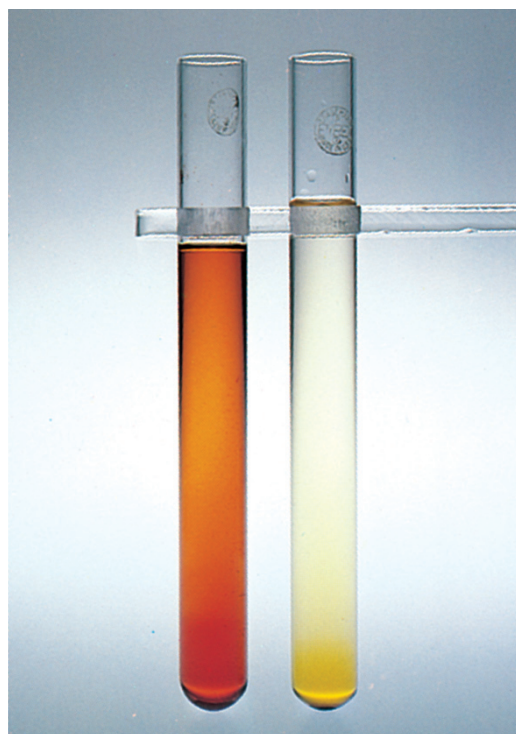
About one kilogram of bromine can be obtained from 15 000 liters of seawater.

Bromine is used to prepare a wide variety of metal bromide and organobromide compounds. Bromine is used as a fumigant and in the synthesis of fire retardants, photochemicals, dyes, and pharmaceuticals, especially sedatives.

Iodide ion is present in seawater and is assimilated and concentrated by many marine animals and by seaweed. Certain seaweeds are an especially rich source of iodine. The iodide ion in seaweed is converted to iodine by oxidation with chlorine. Iodine is also obtained from Chilean mineral deposits of sodium iodate,  $\text{NaIO}_3(\text{s})$ , and sodium periodate,  $\text{NaIO}_4(\text{s})$ . The free element is obtained by the reaction of iodine in  $\text{IO}_3^{-}(\text{aq})$  or  $\text{IO}_4^{-}(\text{aq})$  with sodium hydrogen sulfite,  $\text{NaHSO}_3(\text{aq})$ , for example,



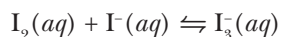
Solid iodine is dark gray with a slight metallic luster. Iodine vapor and solutions of iodine in nonpolar solvents are purple. Solutions of iodine in water and alcohols are brown as a result of the specific polar



**Figure Q.5** Bromine,  $\text{Br}_2(\text{l})$ , dissolved in carbon tetrachloride,  $\text{CCl}_4(\text{l})$  (left) and in water (right).



interactions between molecular iodine and the oxygen-hydrogen bond. Iodine is only slightly soluble in pure water, but when the iodide ion is present, the iodine forms a linear, colorless triiodide complex according to



and the solubility of  $\text{I}_2(s)$  is greatly enhanced. The presence of very low concentrations of aqueous triiodide can be detected by adding starch to the solution. The triiodide ion combines with starch to form a brilliant deep-blue species (Figure Q.6).

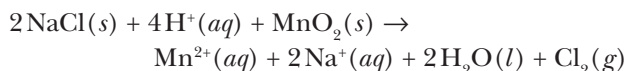
Iodide ion is essential for the proper functioning of the thyroid gland, which in humans is located in the base of the throat. Iodide deficiency disease is manifested as a goiter, which is a swelling in the neck caused by an enlargement of the thyroid gland. Potassium iodide is added to ordinary table salt that is marketed as iodized salt. Iodized salt acts to prevent iodide deficiency. Alcohol solutions of iodine, known as tincture of iodine, were once used extensively as an antiseptic.

Table Q.3 lists other commercially important compounds of the halogens.

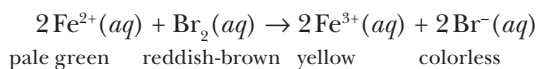
### Q-3. Chlorine, Bromine, and Iodine Have Very Similar Chemical Properties

Chlorine, bromine, and iodine have such similar chemical properties that it is convenient to consider them together. All three halogens can be prepared by

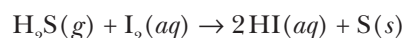
the oxidation of a halide salt by manganese dioxide in concentrated sulfuric acid; for example,



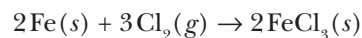
The halogens act as oxidizing agents. For example, if bromine water is added to an aqueous solution of  $\text{FeSO}_4(aq)$  in dilute sulfuric acid, the color of bromine disappears and the yellow color of a solution of  $\text{Fe}_2(\text{SO}_4)_3(aq)$  appears:



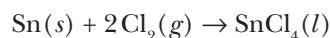
If  $\text{H}_2\text{S}(g)$  is passed into an aqueous solution of chlorine, bromine, or iodine, then the halogen color disappears and a yellow precipitate of sulfur appears; for example,



Chlorine reacts with most metals at moderate temperatures. For example,



and



**Figure Q.6** Various solutions of iodine. From left to right,  $\text{I}_2(s)$  dissolved in  $\text{CCl}_4(l)$ ,  $\text{I}_2(s)$  dissolved in  $\text{KI}(aq)$ ,  $\text{I}_2(s)$  dissolved in water, and  $\text{I}_2(s)$  dissolved in  $\text{KI}(aq)$  with starch added to the solution.

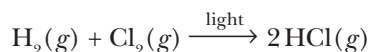
TABLE Q.3 Some important compounds of the halogens

Compound	Uses
hydrogen fluoride, $\text{HF}(g)$	catalyst in the petroleum industry; refining of uranium
sulfur hexafluoride, $\text{SF}_6(g)$	gaseous electrical insulator
sodium hexafluoro-aluminate(III) ( <i>cryolite</i> ), $\text{Na}_3\text{AlF}_6(s)$	production of aluminum
hydrochloric acid, $\text{HCl}(g)$	ore refining; metallurgy; boiler-scale removal; food processing
sodium chloride, $\text{NaCl}(s)$	table salt; preservative; soap and dye manufacture; deicer for roads
calcium chloride, $\text{CaCl}_2(s)$	drying agent; dust control on roads; paper industry; refrigeration; fireproofing
potassium chlorate, $\text{KClO}_3(s)$	oxidizing agent; explosives; matches
calcium hypochlorite, $\text{Ca}(\text{ClO})_2(s)$	algaecide; bactericide; deodorant; bleaching agent for paper and textiles
sodium bromide, $\text{NaBr}(s)$	photochemicals; sedatives
potassium bromide, $\text{KBr}(s)$	gelatin bromide; photographic papers
potassium bromate, $\text{KBrO}_3(s)$	oxidizing agent; food additive; permanent wave compound for hair
silver bromide, $\text{AgBr}(s)$	photographic films and plates
sodium iodide, $\text{NaI}(s)$	photochemicals; feed additive
silver iodide, $\text{AgI}(s)$	photochemicals; cloud seeding

Bromine and iodine react in a similar way but usually require higher temperatures.

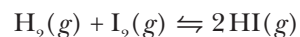
The halogens also react with many nonmetallic elements. For example, phosphorus forms a trichloride,  $\text{PCl}_3(l)$ , and a pentachloride,  $\text{PCl}_5(s)$ , with chlorine (see Interchapter N). Phosphorus tribromide,  $\text{PBr}_3(l)$ , results when carbon disulfide,  $\text{CS}_2(l)$ , solutions of white phosphorus and bromine are mixed, and phosphorus triiodide,  $\text{PI}_3(s)$ , can be made in a similar manner. Sulfur reacts with both chlorine and bromine to form a compound of the formula  $\text{S}_2\text{X}_2(l)$ , but iodine does not. Except for the reaction between fluorine and oxygen, none of the halogens reacts directly with oxygen or nitrogen.

The most important binary compounds of the halogens and a nonmetal are the hydrogen halides. Chlorine reacts explosively with hydrogen in direct sunlight to produce hydrogen chloride according to



whereas bromine and hydrogen combine slowly in sunlight. Iodine combines with hydrogen even less

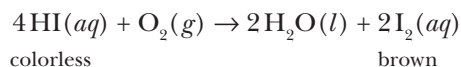
readily and requires a catalyst for any appreciable reaction rate. The reaction is reversible as indicated by the double arrow in the equation



Hydrogen chloride,  $\text{HCl}(g)$ , is a colorless, corrosive gas with a pungent, irritating odor. It is a major industrial chemical and is produced on a huge scale as a by-product in the chlorination of hydrocarbons, although the direct combination of hydrogen and chlorine is used if a high-purity product is desired. The annual world production of  $\text{HCl}(g)$  is well in excess of five million metric tons. Hydrochloric acid, which is also called **muriatic acid**, is an aqueous solution of hydrogen chloride. Its largest single use is in the treatment of steel and other metals to remove an adhering oxide coating, although it has many other industrial uses.

Like hydrochloric acid, hydrobromic acid and hydroiodic acid are strong acids. Hydrobromic acid,  $\text{HBr}(aq)$ , and hydroiodic acid,  $\text{HI}(aq)$ , are both reducing agents. Of the two,  $\text{HI}(aq)$  is the stronger and more

useful reducing agent. Solutions of hydroiodic acid turn brown upon standing owing to air oxidation:

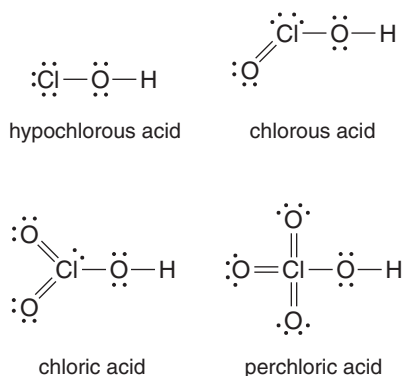


#### Q-4. The Halogens Form Numerous Oxygen-Halogen Compounds

The best-known and most important oxygen-halogen compounds are the halogen oxyacids. For example, the oxyacids of chlorine are

$\text{HClO}(aq)$	hypochlorous acid
$\text{HClO}_2(aq)$	chlorous acid
$\text{HClO}_3(aq)$	chloric acid
$\text{HClO}_4(aq)$	perchloric acid

The Lewis formulas for these acids are

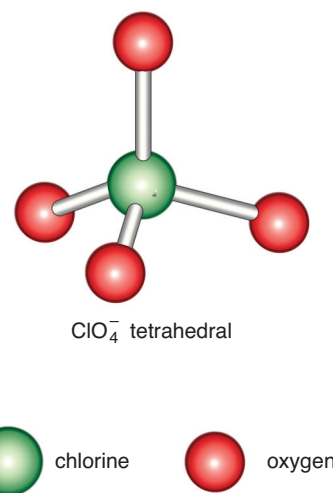
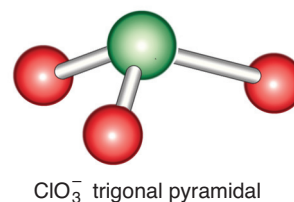
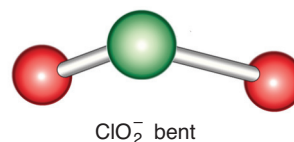


Notice that in each case the hydrogen is attached to an oxygen atom. The anions of the chlorine oxyacids are

$\text{ClO}^-$ hypochlorite	$\text{ClO}_3^-$ chlorate
$\text{ClO}_2^-$ chlorite	$\text{ClO}_4^-$ perchlorate

The shapes of these ions are predicted correctly by VSEPR theory (Chapter 8) and are shown in Figure Q.7. Table Q.4 gives the known halogen oxyacids and their anions. Note that there are no oxyacids of fluorine.

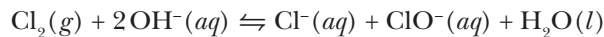
The strength of an oxyacid depends strongly upon the number of oxygen atoms without attached hydrogen atoms that are attached to the central atom. Of the four oxyacids of chlorine for which the Lewis formulas are given above, perchloric acid is the strongest, followed by chloric acid, chlorous acid, and finally hypochlorous acid, which is the weakest.



**Figure Q.7** The shapes of the oxyacid anions of chlorine.

We shall learn more about acids and acid strength in Chapter 20.

The various oxygen-halogen compounds, often called simply **oxyhalogens**, are prepared by the reactions of the halogens with water under various conditions. When  $\text{Cl}_2(g)$ , for example, is dissolved in aqueous alkaline solution, the reaction described by the following equation occurs:



A solution of  $\text{NaClO}(aq)$  is a bleaching agent, and many household bleaches are a 5.25% aqueous solution of sodium hypochlorite. Commercially, solutions of  $\text{NaClO}(aq)$  are manufactured by the electrolysis of cold aqueous solutions of sodium chloride (see

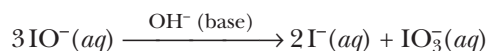


TABLE Q.4 The halogen oxyacids

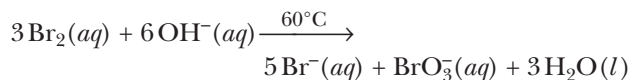
Chlorine	Bromine	Iodine	Acid	Salt
HClO	HBrO	HIO	hypohalous	hypohalite
HClO <sub>2</sub>	—	—	halous	halite
HClO <sub>3</sub>	HBrO <sub>3</sub>	HIO <sub>3</sub>	halic	halate
HClO <sub>4</sub>	HBrO <sub>4</sub>	HIO <sub>4</sub>	perhalic	perhalate

Chapter 25). Sodium hypochlorite is also employed as a disinfectant and deodorant in dairies, creameries, water supplies, and sewage disposals.

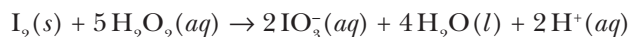
Hypohalite ions decompose in basic solution via reactions of the type



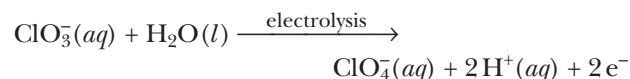
The analogous reaction with  $\text{ClO}^-(aq)$  is slow, which makes it possible to use hypochlorite as a bleach in basic solutions. The rate of decomposition of  $\text{ClO}^-(aq)$ ,  $\text{BrO}^-(aq)$ , and  $\text{IO}^-(aq)$  in hot alkaline aqueous solution is sufficiently fast that when  $\text{Cl}_2(g)$ ,  $\text{Br}_2(l)$ , or  $\text{I}_2(s)$  is dissolved in basic solution, and the resulting solution is heated to  $60^\circ\text{C}$ , reactions such as that described by the following equation go essentially to completion:



Chlorates, bromates, and iodates also can be prepared by the reaction of the appropriate halogen with concentrated nitric acid,  $\text{HNO}_3(aq)$ , or hydrogen peroxide,  $\text{H}_2\text{O}_2(aq)$ , or (commercially) by electrolysis of the halide. For example, the reaction of  $\text{I}_2(s)$  with  $\text{H}_2\text{O}_2(aq)$  is given by the equation



Perchlorate and periodate are prepared by the electrolysis of chlorate and iodate, respectively (see Chapter 25). For example,



The perchlorate is obtained from the electrolyzed cell solution by adding potassium chloride in order to pre-

cipitate potassium perchlorate, which is only moderately soluble in water. Perchloric acid,  $\text{HClO}_4(aq)$ , also can be obtained from the electrolyzed solution containing perchlorate by adding sulfuric acid and then distilling. Concentrated perchloric acid should not be allowed to come into contact with reducing agents, such as organic matter, because of the extreme danger of a violent explosion. Solutions containing perchlorates should not be evaporated because of their treacherously explosive nature. Perchlorates are used in explosives, fireworks, matches, and as an oxidizer in solid rocket fuels. Recent data from the Mars Phoenix Lander suggest that perchlorates may be one of the natural soil constituents of Mars.

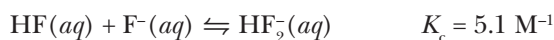
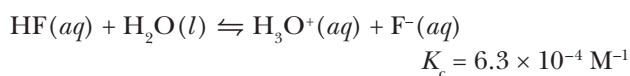
## TERMS YOU SHOULD KNOW

halogen Q1  
halide Q1  
hydrohalic acid Q3  
muriatic acid Q6  
oxyhalogen Q7

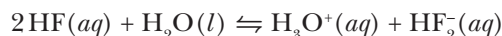
## QUESTIONS

Q-1. Why do the boiling points of the halogens increase going from fluorine to iodine?

Q-2. Given



calculate the value of  $K_c$  for



**Q-3.** Explain briefly why fluorine is capable of stabilizing unusually high oxidation states in many elements.

**Q-4.** Describe how each of the halogens is prepared on a commercial scale.

**Q-5.** Give the chemical formulas and names of the oxyacids of chlorine.

**Q-6.** Describe how the glass used in frosted light bulbs is etched.

**Q-7.** Describe the role of fluoride in the prevention of tooth decay.

**Q-8.** Write a chemical equation for the laboratory preparation of bromine.

**Q-9.** What is household bleach?

**Q-10.** Describe, using balanced chemical equations, how you would prepare  $\text{KIO}_3(s)$  starting with  $\text{I}_2(s)$ .

**Q-11.** Name the following oxyacids:

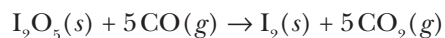
- |                         |                        |
|-------------------------|------------------------|
| (a) $\text{HBrO}_2(aq)$ | (b) $\text{HIO}(aq)$   |
| (c) $\text{HBrO}_4(aq)$ | (d) $\text{HIO}_3(aq)$ |

**Q-12.** Use VSEPR theory to predict the shapes of the following species:

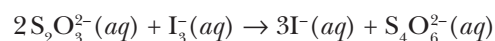
- |                      |                      |
|----------------------|----------------------|
| (a) $\text{I}_3^-$   | (b) $\text{IF}_3$    |
| (c) $\text{BrF}_5$   | (d) $\text{ICl}_4^-$ |
| (e) $\text{BrF}_6^+$ |                      |

**Q-13.** Bromine is present in seawater as  $8.3 \times 10^{-4} \text{ M Br}^-(aq)$ . Calculate how many kilograms of elemental bromine can be obtained from 15000 liters of seawater.

**Q-14.** Iodine pentoxide,  $\text{I}_2\text{O}_5(s)$ , is a reagent for the quantitative determination of carbon monoxide. The reaction equation is



The iodine produced is dissolved in  $\text{KI}(aq)$  to form  $\text{I}_3^-(aq)$ , which is then determined by titration with  $\text{Na}_2\text{S}_2\text{O}_3(aq)$  according to



Suppose that 10.0 liters of air at  $0^\circ\text{C}$  and one bar is bubbled through a bed of  $\text{I}_2\text{O}_5(s)$  and that it required 4.76 mL of  $0.0100 \text{ M Na}_2\text{S}_2\text{O}_3(aq)$  to titrate the  $\text{I}_2(s)$  produced. Calculate the mass in milligrams of  $\text{CO}(g)$  in the 10.0-liter air sample.