REVIEW OF ADVANCED INORGANIC OXIDIZERS

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I. Introduction

In this paper the properties and chemistry of advanced inorganic oxidizers will be reviewed with emphasis on recent developments. The classes of compounds will be limited to the oxygen fluorides, the nitrogen fluorides, the nitrogen fluoride oxides, the halogen fluorides and the halogen fluoride oxides. In general, all of the compounds described here are highly reactive, toxic materials, and very volatile except for ionic derivatives. The growth of chemistry involving O-F, N-F and Cl-F compounds has been remarkable: from about nil in 1946 to over 1000 today.

II. Oxygen Fluorides and Derivatives

The reported oxygen fluorides, their derivatives and some of their properties are summarized in Table I. Oxygen difluoride is the only binary compound of this type which is stable at room temperature, but O_2F_2 is stable below about -80°C and is well characterized. A reasonably good understanding has been gained recently of the apparent equilibrium system $O_4F_2 \rightleftharpoons 2O_2F$.

The existence of O_3F_2 as a separate entity is in doubt, and the natures of the reported "O_5F_2" and "O_6F_2" are quite uncertain. Oxygen difluoride is commercially available, while the others are prepared by electrical discharge at low temperatures from appropriate O_2-F_2 mixtures.

The structures of OF_2 and O_2F_2 have been determined and are compared in Figure 1 with those of related compounds. The abnormally long O-F distance and the greatly shortened O-O bond in O_2F_2 have led Linett to formulate its structure as having an essentially four-electron O-O bond (as in O_2 itself, 1.21 Å bond length) and a one-electron O-F bond (resonating between the two positions). Dioxygen difluoride can, therefore, readily dissociate at the O-F bond to give F atoms and the stable O_2F radical:

$$O_2F_2 \longrightarrow O_2F\cdot + F\cdot$$

and is an extremely energetic fluorinating agent. At low temperatures it is more reactive than fluorine itself and far more reactive than OF_2. The latter compound begins to decompose near 200°C and many of its reactions require activation by heat or light. The $O_2F\cdot$ radical can be generated by ultraviolet radiation of liquid OF_2, of OF_2-O_2, F_2-O_2 or F_2-N_2O mixtures in a matrix, at 4°K, by the decomposition of O_4F_2 or FSO_2OOF, and in the reaction of CF_4 with O_2. The OF radical has never been detected in the gas or liquid phase despite an intensive search.

The reactions of the oxygen fluorides are usually simple fluorinations, but oxygen addition by OF_2 can occur especially in aqueous solution or in reactions

TABLE I

OXYGEN FLUORIDES AND DERIVATIVES

Compound	Properties: MP, °C	BP, °C	Decomp. Temp.	^{19}F NMR[a]	Infrared[b]
OF·	Observed in N_2 or Ar matrix at 4°R			--	1028
OF_2	-223.8	-144.8	~ 200	-249 liq. -248 g	1740, 929, 909, 880, 826, 461
O_2F_2	-163.5	-57 dec.	-78	-865 -825	1024, 628, 463
$O_4F_2 \rightleftharpoons 2O_2F\cdot$	-191	-80 ext	-183	--	1519, 588 (O_4F_2) 1494, 484 ($O_2F\cdot$)
"O_3F_2"	-189 to 190	-60 dec.	<-158	-857	--
"O_5F_2"	<-196	--	--	--	--
"O_6F_2"	<-213	--	--	--	--
CF_3OF	<-215	-95	--	72, -147	1282, 1259, 1223, 1217, 945, 880, 679
$CF_2(OF)_2$	--	-62.5	--	-159 tr, +84.2 tr	1300, 1270, 1150, 950 qt
$CF_3C(O)OF$	--	-21 -21.5	--	--	--
SF_5OF	-86	-35	--	-189	935, 888, 614, 585
NO_2OF	-175	-45.9	--	--	1760, 1300, 920, 810, 708
ClO_3OF	-167.3	-15.9	--	-226	1298, 1049, 666
FSO_2OF	-158.5	-31.3	--	-244.9, -36.2	--
CF_3OOF	--	-50 (est)	--	+69.2 db, -291.5 qt	1290, 1175, 955, 755, 725
C_2F_5OOF	--	-15 (est)	--	+84.1 db of tr (CF_3) +97.4 db of qt (CF_2) -291.6 tr of qt	1380, 1285, 1235, 1175, 740
FSO_2OOF	--	0	--	--	--

a/ ^{19}F chemical shift (ppm vs. $CFCl_3$=0) for fluorine of -OF group.
b/ Positions (cm^{-1}) for O-F stretching fundamental.

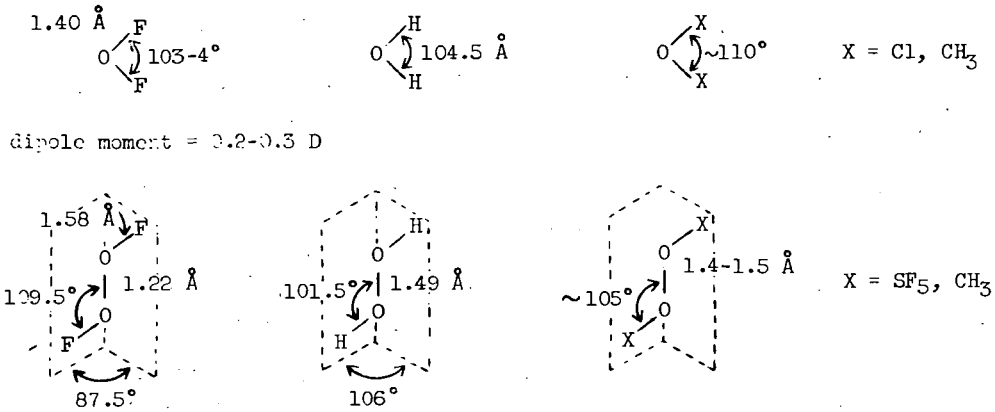


Figure 1 - Comparison of Structures of OF_2, O_2F_2 and Related Compounds

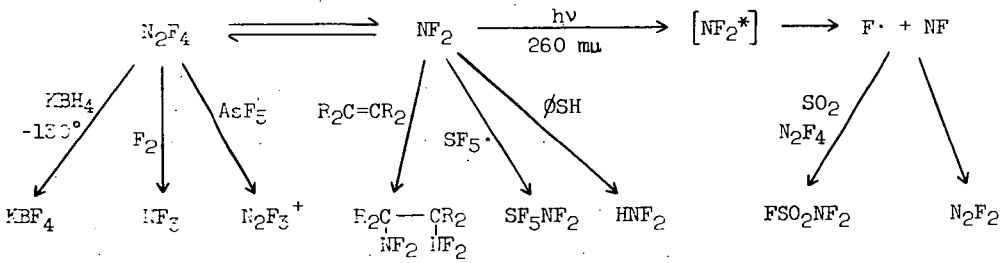


Figure 2 - Examples of Reaction Types for N_2F_4

with some nitrogen, chlorine, or sulfur-containing materials. Dioxygen difluoride (or the O_2F radical) undergoes the fluoride abstraction reaction with strong Lewis acids at low temperatures to give stable dioxygenyl salts. Oxygen difluoride undergoes a similar reaction at elevated temperatures.

$$O_2F_2 + BF_3 \xrightarrow{-140^\circ} O_2BF_4 + 1/2F_2$$

$$2OF_2 + SbF_5 \xrightarrow{200^\circ} O_2SbF_6 + F_2$$

Although considerable doubt now exists about the existence of trioxygen trifluoride, the trifluoromethyl derivative CF_3OOOCF_3 is known as well as the peroxide CF_3OOCF_3. The authors are not aware of reports of the simple ether $(CF_3)_2O$, although the sulfide $(CF_3)_2S$ is stable at 400°C.

III. Nitrogen Fluorides and Derivatives

The reported nitrogen fluorides, their derivatives and some of their properties are summarized in Table II.

The NF molecule is known only as a reactive intermediate which dimerizes readily to N_2F_2. It has been observed by flash photolysis of the NF_2 radical in an argon matrix at 4°K and by pyrolysis of N_2F_4 at high dilution in argon at 2550°K.

The NF_2 radical is well established to exist in equilibrium with N_2F_4 in analogy to the N_2O_4-NO_2 equilibrium

$$N_2F_4 \rightleftharpoons 2NF_2$$

The structure of NF_2 (N-F = 1.36 Å; ∠FNF = 103°) is closely related to that of NF_3. The NF_2 radical has fundamental stretching frequencies at 1074 and 935 cm^{-1} in the infrared. In the ultraviolet, NF_2 absorbs at 260 mμ to give an excited radical, NF_2*, which decomposes to NF and F·, but may undergo unique reactions. The reactions of NF_2 are discussed in the section on N_2F_4.

Nitrogen trifluoride is commercially available and its properties are well known. It has an ammonia-like structure with N-F = 1.37 Å and ∠FNF = 103°. Only 57 kcal/mol are required to dissociate a F· atom from NF_3, but an average of 71 kcal/mol are required for each of the remaining two N-F bonds of the stable NF_2 radical. NF_3 is relatively unreactive at mild temperatures, but functions as a fluorinating agent at elevated temperature. The most interesting reaction of NF_3 to be reported recently is that with fluorine and Lewis acids to give stable salts of the NF_4^+ ion.

$$NF_3 + F_2 + AsF_5 \xrightarrow[-80^\circ]{\text{elect. disc.}} NF_4^+AsF_6^-$$

$$NF_3 + F_2 + SbF_6 \xrightarrow[3000\ \text{psi}]{300^\circ} NF_4^+SbF_6^-$$

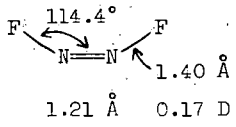
Difluorodiazine is now well established to exist in the <u>cis</u> and <u>trans</u> isomeric forms and the properties and structures of each are known.

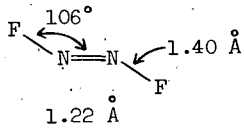
TABLE II

NITROGEN FLUORIDES AND DERIVATIVES

Compound	MP, °C	BP, °C	^{19}F NMR	IR
NF_2	--	--	--	1070, 931
NF_3	-208.5	-129	-142	1031, 1010, 907, 642
N_2F_2 (cis)	-195	-105.7	-129	1524db, 952tr, 896db, 737tr
(trans)	-172	-111.4	-88	989 (1996m, 1581m, 1434m)
N_2F_4	-164.5	-73	-58	1010, 998, 959, 946
N_3F	-154	-82	--	--
NFO	-132.5	-59.9	-479	1844, 766, 521
NFO_2	-166	-72.5	-393	1793, 1312, 822, 742, 570
NF_3O	-160	-85	-363	1687, 887, 743, 528
NF_2NO	--	dec.	-65	--
HNF_2	-116 ± 3	<23.6	6	3193w, 1428db, 1280db, 978db
$ClNF_2$	-183	-67	-141	920tr, 853db, 694tr
Cl_2NF	-10 to 0 (est.)		--	--
N_2F^+			-103	1050
$N_2F_3^+$			-144 (db.) -192	1521, 1304, 1128, 924, 518, 497
NF_4^+			-214.7 (tr.)	1163, 611
NF_2O^+			-331	1857, 1162, 905

Other derivatives include compounds of the types: MNF_2 (M = SF_5-, SF_5O-, CF_3SF_4-, FSO_2-, FSO_2O-, CF_3S-); $R(NF_2)_x$ (R = R_f, alkyl, anthracene, etc.; X = 1-4); R-C(=NF)-R'; and R-N(O)=NF.





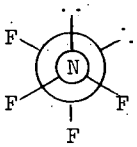
The N-F bond is lengthened over that in NF_3 and the N=N bond is shortened compared to that (1.24 Å) in $N_2(CH_3)_2$. The N-F and N=N bond energies of the trans form is 68 and 108 kcal/mol, respectively. The trans N_2F_2 is conveniently prepared by reaction of N_2F_4 at low pressure with $AlCl_3$ at -80°. The trans form is converted to the cis form in over 90% yield at 75° in a well-passivated stainless steel cylinder. The N_2F_2 is obtained nearly quantitatively by the decomposition at 20° of the complex $KF \cdot HNF_2$ formed at -80°C. Difluorodiazine has also been offered commercially. The cis N_2F_2 is more reactive chemically than trans N_2F_2, but both can act as fluorinating agents. The so-called N=N double bond is unusually inert and does not undergo addition reactions. Detonations, especially with the cis isomer, have occurred upon application of high pressure. The fluorination of N_2F_2 to N_2F_4 has apparently not been reported. cis-N_2F_2 undergoes the fluoride abstraction with Lewis acids to give a stable salt of the N_2F^+ ion

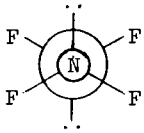
$$N_2F_2 + AsF_5 \longrightarrow N_2F^+AsF_6^-$$

The trans isomer also gives this reaction with SbF_5.

Tetrafluorohydrazine has the structure in the gas phase on the left below (d and l isomers) but the symmetrical isomer on the right has also been observed in the liquid at low temperature.

N-F 1.39 Å
N-N 1.53 Å
∠ FNF 104°
∠ NNF 102°
dihedral angle 69°
^{19}F NMR -65.2, -53.2, -44.0, -31.9 ∅
(in NF_3 at -155°)





-60.4 ∅

The abnormally long N-N bond (compared to 1.45 Å in N_2H_4) has a low dissociation energy (20 kcal/mol) and N_2F_4 exists in equilibrium with stable NF_2 radicals. Tetrafluorohydrazine has been available commercially since 1960 and its reactions have been studied intensively. It enters into at least five types of reactions: (1) normal NF_2 radical reactions including addition to olefins or a second radical, and abstraction reactions; (2) excited NF_2 reactions resulting from photolysis at about 260 mμ; (3) fluorination (oxidation) reactions; (4) reduction reactions; (5) fluoride abstraction reaction. Examples of each are illustrated in Figure 2.

Trifluoramine oxide, NF_3O, was reported in 1965. It can be prepared by electric discharge at -196° in mixtures of NF_3 and O_2 or OF_2, by the flame fluorination of NO with fast quench, or more conveniently at 25° by the fluorination of NO with F_2 photochemically or with certain metal hexafluorides such as IrF_6. The NF_3O has a tetrahedral structure in which the N-F and N → O bonds are slightly

weaker than those in NFO_2. The NF_3O has good thermal and hydrolytic stability, but undergoes the fluoride abstraction reaction with strong Lewis acids, to give salts such as $NF_2O^+AsF_6^-$, and adds to perfluoroolefins (under BF_3 catalysis) to give stable R_fONF_2 compounds. It apparently reacts slowly with NO to give NOF.

Difluoramine has an ammonia-like structure with the following parameters: N-H, 1.03-1.08; N-F, 1.38; FNF, 103°, HNF, 102°; dipole moment, 1.93 D. It is best prepared from difluorourea (which is obtained by aqueous fluorination of urea) by treatment with H_2SO_4 at 90° or by the reaction of N_2F_4 with C_6H_5SH at 50°. The HNF_2 is stable and can be stored, but the usual procedure is to generate it as needed and pass it directly into a reaction vessel, since it has a tendency to explode when frozen. The reactions of HNF_2 are usually complex, but it undergoes three general types of reactions as follows: oxidation, for example with aqueous Fe^{+3} solution to give N_2F_4 (perhaps the NF_2^- ion is involved); reduction, as in the reaction with aqueous HI to give NH_4F and HF; complex formation with ethers, Lewis acids, and metal fluorides.

Chlorodifluoramine $ClNF_2$ is well known, Cl_2NF and $BrNF_2$ are known as unstable compounds, and the other halogen fluoramines appear to be very unstable. The $ClNF_2$ (or $BrNF_2$) can be prepared by the reaction of aqueous NaOCl (or NaOBr) with N,N-difluoroureas or N,N-difluorosulfuryl amide. The Cl_2NF is prepared by the reaction of ClF with ClN_3 at 25°C or with NaN_3 at 0°. The $ClNF_2$ is stable but dissociates readily to give Cl atoms and NF_2 radicals (which defines the reaction chemistry) while Cl_2NF is explosively unstable in the liquid state.

A few remaining N-F compounds such as NF_2NO and N_3F are of limited interest, but a number of inorganic compounds and a host of organic compounds have been prepared in recent years in which NF_2 groups may be regarded as substituents, e.g., SF_5NF_2, $C(NF_2)_4$, $C(NF)(NF_2)_2$, CF_3ONF_2 and $CF_3N(O)=NF$.

A summary of the interconversions of the nitrogen fluorides is found in Figure 3.

Chlorine Fluorides and Related Compounds

The halogen fluorides of prime interest as propellant oxidizers have been ClF_3, ClF_5, ClO_3F and BrF_5, but a number of other halogen fluorides have been studied. Properties of halogen fluorides are summarized in Table III.

Chlorine monofluoride appears to have considerable ionic character as reflected in the ^{19}F NMR chemical shift of ϕ = +441 ppm (vs. $CFCl_3$). The Cl-F bond distance is 1.63 Å, the dipole moment, 0.88 D, the bond dissociation energy, 60.4 kcal/mol, and the heat of formation, -13.5 kcal/mol. The ClF is an energetic fluorinating agent. It reacts with fluorides such as CsF or NOF to give $Cs^+ClF_2^-$ and $NO^+ClF_2^-$, respectively, and has been reported to react with the Lewis acid AsF_5 to give $Cl^+AsF_6^-$ but substantiating evidence is lacking. The high volatility of ClF (b.p. -100°C) suggests that little or no association or self-ionization (to Cl^+ and ClF_2^- ions or to Cl_2F^+ and ClF_2^- ions) exists, but the electrical conductivity is higher than that of ClF_3. Chlorine monofluoride is prepared by reaction of ClF_3 and Cl_2.

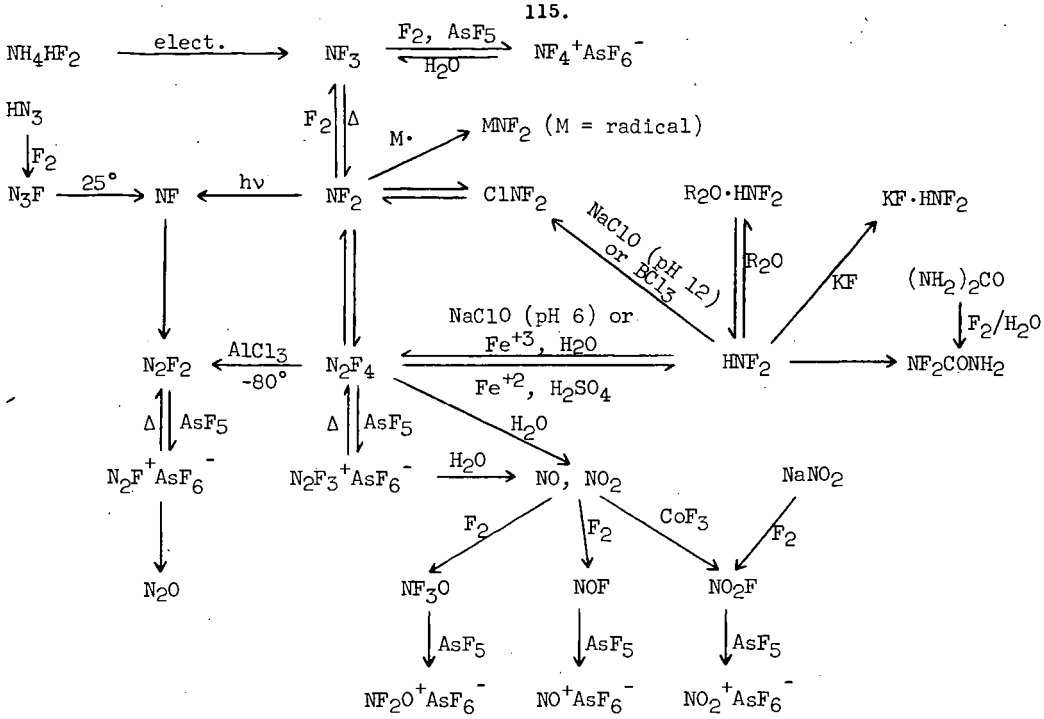


Figure 3 - Interconversion of Nitrogen Fluorides

TABLE III

PROPERTIES OF HALOGEN FLUORIDES AND DERIVATIVES

Compound	MP, °C	BP, °C	NMR	IR
ClF	-154 ± 5	-100.8	440	793, 785
ClF_3	-76.3	11.75	-14(a), -123(e)	761, 746, 713, 694
ClF_5	-103 ± 4	-14	-412(a), -247(e)	786, 732, 541, 486
BrF_3	8.8	127	39	674, 613, 384, 300
BrF_5	-60.5	40.8	-132(e)	683, 644, 587, 369
IF_5	8.5	97.0	-53(a), -4(e)	710, 640, 318
IF_7	4.8[a]	4 (1 atm.)	-177 (-168)	670, 547, 426, 368, 250
ClO_2F	above -115	-6	-329 (330)	1270, 1259, 1104 (pqr) 632, 548
ClO_3F	-146	-46.8	-287	1315-1300, 725, 717, 714, 590
IF_5O	4.5	--	--	--

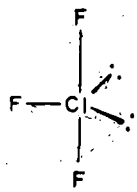
a Triple point 6.4.

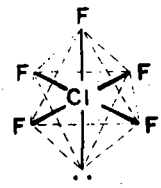
Chlorine trifluoride is usually described as having a T-shape, but is better regarded as a trigonal bipyramid derivative. The two apical Cl-F bonds are elongated (1.70 Å) compared to the equatorial Cl-F bond (1.60 Å). The bond angles are 87.5° and 175°. The ^{19}F NMR of gaseous ClF_3 shows the apical fluorines to be further downfield (-145.6 ϕ) than the equatorial fluorine -19.7 ϕ. In liquid ClF_3 (with HF carefully removed) the resonances are -159.0 and -21.8 ϕ. The high dipole moment determined with the liquid (1.00-1.03 D) compared to that in the gas (~ 0.6 D) suggests considerable interaction and the relatively low volatility (b.p. 12°) confirms this, but the electrical conductivity shows that this interaction is essentially not in the form of self-ionization to ClF_2^+ and ClF_4^- ions. Evidence for some dimerization in the gas has been suggested. However stable salts of both of these ions are known, e.g., $ClF_2^+SbF_6^-$ and $Cs^+ClF_4^-$. Chlorine trifluoride is commercially available in quantity, being prepared by fluorination of chlorine. It is an extremely vigorous fluorinating agent, but can in turn be fluorinated to ClF_5 under pressure. Moisture converts ClF_3 to ClO_2F and ClO_2.

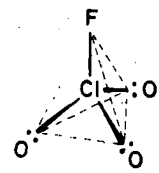
Chlorine pentafluoride is the newest member of the halogen fluoride family and has very recently become available commercially in laboratory quantities. It is prepared by the high pressure fluorination of ClF_3 or a salt such as $CsClF_4$ at 150° or above. It has also been prepared by the fluorination of ClF_3 under ultraviolet radiation or electric discharge at low temperatures and by platinum hexafluoride. The structure of ClF_5, a square pyramid, is an octahedral derivative with the apical Cl-F bond length of 1.62 Å and the equatorial bonds 1.72 Å. The ^{19}F chemical shifts are -412 ϕ (apical fluorine) and -247 ϕ (basal fluorines). The former is one of the most unshielded fluorine atoms known, being exceeded only by those in F_2, O_2F_2, and NOF. The dipole moment is probably 0.2-0.4 D. The boiling point indicates very little interaction exists in the liquid and self-ionization is not expected. The ClF_4^+ ion is an intense peak in the mass spectrum and the possibility of obtaining

stable salts such as $ClF_4^+SbF_6^-$ appears reasonably good. The addition of fluoride ion to ClF_5 to give ClF_6^- does not appear likely (although BrF_6^- salts have been reported) since the chlorine atom would be pseudo-heptacoordinate and this type of coordination is without precedence in elements of the second row of the Periodic Table.

The structures of ClF_3, ClF_5, and ClO_3F have been determined and are shown below.







The commercially available higher halogen fluorides are BrF_3, BrF_5 and IF_5, with IF_7 available on special order. Bromine trifluoride is structurally and chemically very similar to ClF_3 except that self-ionization is extensive in liquid BrF_3:

$$2BrF_3 \rightleftharpoons BrF_2^+ + BrF_4^-$$

Bromine trifluoride has therefore an exceptionally high boiling point (126°) and electrical conductivity. It is an excellent solvent for many ionic materials. Stable salts of both the BrF_2^+ and BrF_4^- ions are known, but in some instances with BrF_2^+ the complexes may be fluorine bridged rather than truly ionic. Bromine pentafluoride and iodine pentafluoride are also structurally similar to ClF_5. Self-ionization occurs to an appreciable extent in IF_5, but far less in BrF_5.

$$2IF_5 \rightleftharpoons IF_4^+ + IF_6^-$$

Stable salts have been reported for BrF_4^+ and BrF_6^- as well as for IF_4^+ and IF_6^-, but structural studies are incomplete. Unlike BrF_5 or ClF_5, IF_5 undergoes fluorine exchange with HF. The BrF_3 and BrF_5 are quite energetic fluorinating agents. While IF_5 is a considerably milder fluorinating agent, it inflames or forms explosive mixtures with many organics. It can be dissolved in excess of such solvents as sulfolane, but forms crystalline complexes with certain oxygenated solvents such as dioxane.

Iodine heptafluoride is formed by the reaction of IF_5 with F_2 at about 100°. The dissociation of IF_7 to IF_5 and F_2 sets in at higher temperatures, ~300°, and ΔH_{dissoc} is 28.5 kcal/mol. The structure of IF_7 in the gas phase is the pentagonal bipyramid, but the structure in the solid is still uncertain. The ^{19}F NMR of the gas is a singlet (ϕ = -336 ppm) suggestiong intramolecular exchange. (The spectrum of the liquid is a doublet at lower temperatures, a singlet at higher temperatures because of quadrapole effects.) The symmetrical structure of IF_7 gives it an extremely short (~2°) liquid range (sublimes, 4.77°; triple point, 6.45°). The reaction of IF_7 as a fluorinating agent are much more vigorous than those of IF_5. With strong fluoride acceptors IF_7 forms salts such as $IF_6^+AsF_6^-$. No reaction to

form such salts as $Cs^+IF_8^-$ have been reported, but the fluorine exchange which occurs between HF and IF_7 in the gas phase may involve a HIF_8 or a IF_6HF_2 intermediate.

The halogen fluoride oxides of major interest are ClO_2F and ClO_3F, but BrO_2F, IO_2F and IF_5O have been characterized. Only perchloryl fluoride, ClO_3F, is commercially available. Chloryl fluoride, ClO_2F, is readily prepared by allowing ClF_3 or ClF_5 to stand with excess of a chlorate or chlorite salt in a Monel cylinder:

$$2ClF_3 + 3NaClO_3 \xrightarrow{25^\circ} 3ClO_2F + 3NaF + Cl_2 + 1.5O_2$$

$$ClF_5 + 3NaClO_2 \xrightarrow{25^\circ} 2ClO_2F + 3NaF + Cl_2 + O_2$$

(The O_2 and Cl_2 are removed at -80°.) The ClO_2F is also formed by the reaction of ClF_3 or ClF_5 with traces of moisture, but is in turn readily hydrolyzed with additional water. The ClO_2F is isoelectronic with PF_3 and SF_2O and should have approximately the same structure, i.e., a tetrahedral arrangement with an electron pair at one apex. The ^{19}F NMR shows the fluorine atom far downfield, ϕ = -332 ppm. This structure makes ClO_2F an extremely reactive oxidizer. It is in turn fluorinated only with difficulty to ClF_5. The ClO_2F undergoes the fluoride abstraction reaction to give stable salts such as $ClO_2^+AsF_6^-$, but efforts to form such salts as $Cs^+ClF_2O_2^-$ have not been successful. Pyrolysis of ClO_2F at 300° and 0.5 atm. pressure gives ClF and O_2 apparently via an intermediate ClFO species.

Perchloryl fluoride has a closed tetrahedral structure with the approximate parameters: Cl-F, 1.55-1.60 Å; Cl→O, 1.40 Å, dipole moment 0.023 D. The ^{19}F NMR shows the fluorine at ϕ = -287 ppm. The structure of ClO_3F gives it considerable kinetically derived stability and it is relatively unreactive at mild temperatures, e.g., it does not hydrolyze, or react with sodium. Above 150° ClO_3F is a potent oxidizer for most organics and attacks many metals in the presence of moisture. Attempts to react ClO_3F with fluoride acceptors such as AsF_5 have not led to stable salts, but the $AlCl_3$-catalyzed reaction of ClO_3F with aromatics (to give $ArClO_3$ compounds) likely involves an intermediate ClO_3^+ ion. No reaction between CsF and ClO_3F has been effected and ClO_3F is quite resistant to fluorination. With NH_3, ClO_3F is converted to $NH_4^+NHClO_3^-$, from which metal salts such as $KNHClO_3$ and K_2NClO_3 can be prepared. All of these salts are shock sensitive. Perchloryl fluoride is generally quite soluble without reaction with other oxidizers and has a small (0.5-5 g/liter) solubility in organic solvents and water.

Bromyl fluoride BrO_2F (formed by the reaction $KBrO_3$ or BrO_2 and BrF_5 at -50°) is thermodynamically unstable and has been studied very little. Of the reported iodine fluoride oxides, only IF_5O appears to be unequivocally established as a molecular entity. It is readily formed by reaction of IF_7 with traces of moisture or with glass and has a relatively low reactivity because of the nearly octahedral structure. The remaining iodine fluoride oxides, IF_3O, IO_2F and IO_3F are all white crystalline solids and may not be molecular entities. The "IF_3O" and "IO_2F" can be prepared by reacting I_2O_5 respectively with IF_5 and F_2. The pyrolysis of IF_3O gives IO_2F and IF_5 and the observed reversibility of this reaction suggested the ionic structure $IO_2^+IF_6^-$ for IF_3O. The IO_2F may exist in a bridged form or possibly in the ionic form, $IO_2^+IF_2O_2^-$, since stable salts of both of these ions are formed. The IO_3F (formed by fluorination of periodic acid in HF) is reported to undergo hydrolysis slowly and thus may in fact be structurally analogous to ClO_3F.

Suggestions for Additional Reading

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