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SOME REACTIONS OF PEROXYDISULFURYL DIFLUORIDE

by

Jeanine M. Shreeve and George H. Cady

Department of Chemistry
University of Washington
Seattle 5, Washington

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SOME REACTIONS OF PEROXYDISULFURYL DIFLUORIDE

The reactions of peroxydisulfuryl difluoride with several substances have been examined. Three new fluorosulfonates, octafluorobis(fluorosulfonato)cyclopentane ($C_5F_8(SO_3F)_2$), tetrafluorobis(fluorosulfonato)ethane ($C_2F_4(SO_3F)_2$), and tetrafluorobis(fluorosulfonato)sulfur (VI) ($SF_4(SO_3F)_2$) were produced by direct combination of peroxydisulfuryl difluoride with perfluorocyclopentene (C_5F_8) and tetrafluoroethylene (C_2F_4) at room temperature and with sulfur tetrafluoride (SF_4) at -128° respectively. These new compounds have been characterized. Structures have been confirmed by mass, nuclear magnetic resonance and infrared spectra. A white solid, dioxobis(fluorosulfonato)molybdenum (VI) ($MoO_2(SO_3F)_2$), was obtained with either molybdenum hexacarbonyl or molybdenum metal.

It has been shown that peroxydisulfuryl difluoride reacts with some substances to give fluorosulfonates.^{1,2,3} The reactions of

- 1 John E. Roberts and George H. Cady, J. Am. Chem. Soc., 81, 4166 (1959).
- 2 Ibid., 82, 353 (1960).
- 3 Ibid., 82, 352 (1960).

$S_2O_5F_2$ with several additional substances have now been studied and it has been found that the peroxide acts in three ways:

(1) As an oxygenating agent through the addition of oxygen to the central atom of the reacting species, e.g., with CO, PF_3 , SOF_2 , $COCl_2$, CCl_4 or S; (2) as a fluorosulfonating agent, e.g., with HgO, KCl, ICl, C_5F_8 , C_2F_4 or SF_4 ; (3) as an oxygenating and fluorosulfonating agent by adding oxygen and fluorosulfonate groups to the central atom, e.g., $SOClF$, Mo, or $Mo(CO)_6$. Pyrosulfuryl fluoride ($S_2O_5F_2$) is obtained as a product of reactions of types (1) and (3).

EXPERIMENTAL

Preparation of peroxydisulfuryl difluoride.

Peroxydisulfuryl difluoride can be prepared by reaction of fluorine with an excess of sulfur trioxide at about 250°, by the combination of fluorine fluorosulfonate with sulfur trioxide at 300°⁴ or by the electrolysis of fluorosulfonic acid.⁵ The catalytic

 4 F. B. Dudley and G. H. Cady, J. Am. Chem. Soc., 79, 513 (1957).

5 F. B. Dudley, Thesis, University of New England, Australia (1960)

fluorination of sulfur trioxide vapors by fluorine in the presence of a heated catalyst of copper ribbon coated with silver difluoride provided an easy and simple method to prepare the easily purified product.

Fluorination of sulfur trioxide carried by a stream of dry nitrogen with a slight excess of fluorine in the presence of a silver difluoride catalyst in a "catalytic reactor"⁶ at 155° produced the

 6 K. B. Kellogg and G. H. Cady, J. Am. Chem. Soc., 70, 3086 (1948).

$S_2O_6F_2$, which was purified by prolonged pumping at -78° to remove the contaminants silicon tetrafluoride, sulfuryl fluoride and fluorine fluorosulfonate. In a fifteen hour run using a nitrogen flow of eight liters per hour and a catalyst contact time of about 15 minutes, a 96 gram sample of product was obtained. It contained 97% peroxydisulfuryl difluoride. In another run, 26.1 grams of sulfur trioxide yielded 31.6 grams of $S_2O_6F_2$ (yield, 97% theoretical).

Storage of the compound in a glass vessel at -78° was satisfactory. Although the glass container apparently was not attacked

at room temperature, a nonvolatile oily material was slowly formed in the $S_2O_6F_2$. Apparently this did not introduce contaminants into the $S_2O_6F_2$ distilled from the vessel. The yellow oil was not identified.

Materials

The Swarts reaction⁷ provided a general method for the conversion

7 H. S. Booth and F. C. Mericola, J. Am. Chem. Soc., 62, 640 (1940).

of chlorides to the corresponding fluorides by reaction of the former with antimony (III) fluoride in the presence of antimony (V) chloride as a catalyst. Thionyl fluoride and thionyl chlorofluoride were obtained using thionyl chloride while phosphorus (III) fluoride was prepared from phosphorus (III) chloride. Tetrafluoroethylene was produced by the pyrolysis of Teflon (polytetrafluoroethylene) in an iron vessel at 550°. Sulfur tetrafluoride was used directly from a cylinder supplied by the E. I. duPont de Nemours Co., Inc. All other materials were of reagent grade.

General Methods

Infrared spectra were studied using a Perkin-Elmer Model 21 Infrared Spectrometer with a sodium chloride prism. The gaseous samples were contained in a 10 cm glass cell with silver chloride windows.

Nuclear magnetic resonance spectra were obtained through the use of a Varian Model 4311 B high resolution spectrometer with a forty or sixty megacycle oscillator. For some of the samples areas of the absorption peaks were used as a quantitative measure for fluorine.⁸

8 Wayne P. Van Meter and George H. Cady, J. Am. Chem. Soc., 82, 6005 (1960).

Mass spectra were recorded using a Consolidated Engineering Corporation type 21-103 mass spectrometer.

Vapor densities were determined using Regnault's method. Single capillary pycnometers were used for liquid densities. Melting points were obtained by warming the solid at a rate of about 0.2° per minute in a n-propanol bath which had been cooled by adding solid carbon dioxide. Because of the tendency for the compounds to supercool, freezing points were not found by cooling.

Vapor pressures were determined using a method previously described.⁶

Elemental analyses were obtained for $C_5F_8(SO_3F)_2$ and $C_2F_4(SO_3F)_2$ after heating the material with potassium at 450° for thirty hours. Basic hydrolysis of $SP_4(SO_3F)_2$, KSO_3F , $I(SO_3F)_3$ or $MoO_2(SO_3F)_2$ was used to obtain aqueous solutions for analysis. Sulfur was determined iodometrically as sulfide when the potassium fusion technique was employed. The technique involving basic hydrolysis allowed sulfate to be precipitated as barium sulfate. Fluorine, after distillation,⁹ was precipitated as lead chlorofluoride, dissolved,

9 R. H. Kimball and L. E. Tufts, Ind. Eng. Chem., Anal. Ed., 19, 150 (1947).

and chloride was determined by the Volhard method. After hydrolysis of $MoO_2(SO_3F)_2$, molybdate was removed¹⁰ before sulfate precipitation

10 D. A. Lambie and W. R. Schoeller, Analyst, 65, 281 (1940).

as barium sulfate. Molybdenum was determined gravimetrically as the 8-hydroxyquinolate.

Since $S_2O_6F_2$ and its compounds attack glass only slowly if at all, all reactions studied were carried out in Pyrex glass vessels, unless otherwise specified. Because of the ease of hydrolysis of

Procedure 1 for the reaction of $S_2O_6F_2$ with an equimolar amount of the reacting compound was as follows: (a) The reacting compound was added to a 500 ml Pyrex glass bulb to give a pressure of about 125 mm. (b) The $S_2O_6F_2$ was slowly introduced at room temperature. (When reaction occurred at this temperature it was accompanied by the production of a fleeting brown color in the area of mixing and was exothermic.) (c) After cooling to room temperature, separation of reaction products was effected successfully using fractional codistillation.¹¹

11 G. H. Cady and D. P. Siegwirth, Anal. Chem., 31, 618 (1959).

Procedure II, usually used with materials which did not react at room temperature, was as follows: (a) An equimolar mixture of $S_2O_6F_2$ and the reacting compound was placed in a nickel reactor previously described⁸. (b) As the temperature was slowly increased from 26 to 200°, temperature and pressure were recorded frequently. (c) A plot of this temperature-pressure relationship, if different from that expected for an ideal gas, was regarded as an indication that reaction had occurred. (d) The mixture was cooled, removed from the reactor by condensing in a tube at -183° and later separated by fractional codistillation.

Identification of previously known reaction products, after separation, was made using vapor density and infrared spectra determinations.

Reactions Involving Oxygenation by $S_2O_6F_2$.

Reaction with carbon monoxide.

In an attempt to prepare carbonyl bis(fluorosulfonate) the reaction between $S_2O_6F_2$ and CO was examined using procedures 1 and 2. Two compounds were found. The more volatile one was CO₂ (M.W. 46.2,

I.R. 4.35 μ) and the less volatile was $S_2O_5F_2$ (M.W. 183, I.R. 6.65, 8.05, 11.45, 12.10 and 13.60 μ) which indicated that the reaction $CO + S_2O_6F_2 = CO_2 + S_2O_5F_2$ had occurred at room temperature.

Reaction with carbonyl chloride.

A second attempt to establish the existence of carbonyl bis-(fluorosulfonate) and also chlorine fluorosulfonate led to the study of the reaction between carbonyl chloride and $S_2O_6F_2$. No reaction occurred at room temperature. However, a green gas was produced by u.v. irradiation. Three products were identified: CO_2 (M.W. 46.3, I.R. 4.35 μ); Cl_2 (M.W. 70.5, green) and $S_2O_5F_2$ (M.W. 182.9, I.R. as above). This indicated that the reaction $COCl_2 + S_2O_6F_2 = CO_2 + S_2O_5F_2 + Cl_2$ had occurred.

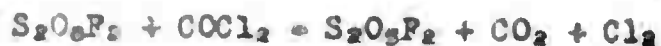
Reaction with molybdenum hexacarbonyl.

When an excess of peroxydisulfuryl difluoride was distilled onto 0.100 g of sublimed molybdenum carbonyl in a 500 ml bulb and the mixture allowed to warm to room temperature, an exothermic, rapid reaction occurred producing 0.108 g. of a gas identified as carbon dioxide (theoretical weight of CO_2 obtainable, 0.100 g.) Pyrosulfuryl fluoride ($S_2O_5F_2$) was also a gaseous product. In addition to the gases a white solid ($MoO_2(SO_2F)_2$) and a nonvolatile liquid, to be discussed later, were produced.

Reaction with carbon tetrachloride.

Unsuccessful attempts were made to prepare chlorine fluorosulfonate by the reaction of $S_2O_6F_2$ with various chlorides including anhydrous carbon tetrachloride. No immediate reaction was noted but on standing at room temperature, a green gas was formed. Four products were identified as follows: CO_2 (I.R. 4.35 μ); Cl_2 (green); $COCl_2$ (I.R. 5.48, 11.75 μ) and $S_2O_5F_2$ (I.R. as above). These indicated the following reactions had occurred:

Excess $S_2 O_8 F_2$



Excess CCl_4



Reaction with sulfur.

An excess of peroxydisulfuryl difluoride was distilled on to finely divided, vacuum dried sulfur, and the mixture was irradiated with ultraviolet light for an extended period. A color change from yellow to orange to white to medium blue to green was observed.

Volatile products were identified as follows: SO_2 (M.W. 60); $S_2 O_5 F_2$ (I.R. as above) and $S_3 O_8 F_2$ (M.W. 262, I.R. 6.70, 8.00, 8.18, 11.50, 12.65 μ). Two equations may be written: $S + 2S_2 O_8 F_2 = SO_2 + 2S_2 O_5 F_2$ and $SO_2 + S_2 O_5 F_2 = S_3 O_8 F_2$. The second of these reactions had previously been recognized.¹ No explanation has been found for the color change which appeared to be a surface effect.

Reaction with phosphorus (III) fluoride.

The reaction of $S_2 O_8 F_2$ and PF_3 was found to be extremely vigorous and exothermic. Products found were: POF_3 (I.R. 6.15, 7.08, 7.42, 10.11, 10.58, 11.45 μ) and $S_2 O_5 F_2$ (I.R. as above). The same products were obtained by mixing at 0° followed by ultraviolet irradiation.



Reaction with thionyl fluoride.

A temperature vs pressure plot obtained from heating a mixture of SOF_2 and $S_2 O_8 F_2$ to 200° was identical with that expected for increasing the temperature of an ideal gas at constant volume. Examination of the products showed small amounts of $SO_2 F_2$ (I.R. 6.65, 7.85, 11.50 μ)

and $S_2O_3F_2$ (I.R. as above). Most of the reactants remained unchanged. A slow oxidation of thionyl fluoride to sulfuryl fluoride had occurred with no pressure change according to $SOF_2 + S_2O_3F_2 = SO_2F_2 + S_2O_3F_2$.

Reactions Involving Fluorosulfonation by $S_2O_3F_2$

The second type of reaction displayed by peroxydisulfuryl difluoride was fluorosulfonation where the $S_2O_3F_2$ molecule was split into two fluorosulfonate groups which then either: (1) replaced volatile anions from the reacting compound; (2) added across carbon-carbon double bonds; or (3) added to the central atom of the reacting compound increasing its oxidation number.

Reaction with mercury (II) oxide.

A slight excess of peroxydisulfuryl difluoride was distilled on to 0.2536 g. (1.170 millimoles) of yellow mercury (II) oxide which had been dried at 300° . No reaction occurred at room temperature, but on heating at 150° for 24 hours, a white solid was obtained according to the reaction, $HgO + 2S_2O_3F_2 = Hg(SO_3F)_2 + O_2 + S_2O_3F_2$. After volatile materials had been removed, the remaining mercury bis(fluorosulfonate) weighed 0.3490 g. (1.164 millimoles). On addition of water, the yellow solid $Hg_3F_4(OH)_2 \cdot 3H_2O$, noted by Roberts and Cady² on hydrolyzing $Hg(SO_3F)_2$, was produced. Because $Hg(SO_3F)_2$ had been previously synthesized and characterized, no further study was undertaken.

Reaction with potassium chloride.

An excess of peroxydisulfuryl difluoride was distilled on to 0.0620 g. (0.832 millimole) of powdered potassium chloride which had been dried at 130° . When the bulb was allowed to warm to room temperature, bubbles could be seen leaving the surface of the solid. Subsequently, a green gas formed by the reaction $2KCl + S_2O_3F_2 = 2KSO_3F + Cl_2$.

was observed. After the reaction had become complete, the excess unchanged $S_2O_8F_2$ and chlorine were distilled away leaving 0.1158 g. (0.333 millimole) of a white solid, KSO_3F . When the potassium fluorosulfonate was subjected to basic hydrolysis at 100° for 48 hours, 1.96 g. equivalents of base were consumed per mole of compound. (Theoretical, 2.00 g. equivalents) Sulfur found: 22.9%; calculated: 22.2% for KSO_3F .

Reaction with iodine monochloride

Addition of peroxydisulfuryl difluoride to ICl resulted in a vigorous reaction producing swirling smoke in the reaction flask and green gaseous chlorine. A red-orange liquid was formed which upon the addition of more $S_2O_8F_2$ lightened in color. When the excess $S_2O_8F_2$ was removed, the liquid crystallized to a yellow solid, $I(SO_3F)_3$.³ The compound was analyzed for sulfur. Sulfur found: 23.42, 22.68%. (Theoretical 22.63%) The equation for the reaction is $2ICl + 3S_2O_8F_2 \rightarrow 2I(SO_3F)_3 + Cl_2$.

Reaction with perfluorocyclopentene to give octafluorobis(fluorosulfonato)cyclopentane.

Peroxydisulfuryl difluoride, in excess, reacted vigorously with perfluorocyclopentene according to the equation $C_5F_8 + S_2O_8F_2 \rightarrow C_5F_8(SO_3F)_2$. On occasion, when the $S_2O_8F_2$ was added too rapidly, flashes of light were observed. After the reaction was complete, the remaining $S_2O_8F_2$ was removed by first pumping while the products were maintained at -20° , and second by a stream of dry helium passing through the remaining material at -20° for four hours. Nearly pure octafluorobis(fluorosulfonato)cyclopentane was obtained.

Properties of octafluorobis(fluorosulfonato)cyclopentane.

(a) Molecular weight. Experimental measurements gave an average molecular weight of 410 ± 10 , compared to a calculated 410.2 for $C_5F_8(SO_3F)_2$. The error was rather large because of the low vapor

pressure (3 mm) of the compound at room temperature. (b) Density. The density of $C_3F_8(SO_2F)_2$ was determined at five different temperatures. The interpolated value at 25° was 1.886 g/cc. The volume

Table I

Densities of $C_3F_8(SO_2F)_2$

Temp., $^\circ C$	10.0	20.0	29.4	40.1	47.7
Density, g/cc	1.917	1.895	1.874	1.850	1.831

coefficient of expansion, calculated from the data, was 1.17×10^{-5} at 25° . (c) Melting point. The melting point was not found since all attempts to crystallize the compound resulted in a glass.

(d) Vapor pressure and boiling point. The vapor pressure values of the compound over the temperature range 61.2 to 149.4° which are given in Table II indicated a boiling point of 148.0° , a molar heat of vaporization of 10.2 kcal. and a Trouton constant of 24.3.

Table II

Vapor pressures of $C_3F_8(SO_2F)_2$

Pmm	T, $^\circ K$	Pmm	T, $^\circ K$
24.9	334.4	280.9	388.6
45.0	343.4	321.9	392.3
65.1	351.8	349.8	395.1
83.4	357.2	429.4	401.7
107.1	362.8	459.9	403.7
129.4	367.5	519.6	407.6
151.8	371.6	571.6	411.2
186.5	376.8	645.2	415.1
226.0	383.0	755	420.0
		(760)	(421.2)
		798	422.6

(e) Infrared spectrum. The infrared spectrum of $C_5F_8(SO_3F)_2$ was obtained in the range 2 to 15 μ at 3 mm pressure. Some of the bands of the product were identified by comparison with those of the reactants.^{12,13} Strong bands at 1496 cm^{-1} (6.68 μ) and 1253 cm^{-1}

12 R. N. Haszeldine, J. Chem. Soc., p. 4423 (1952).

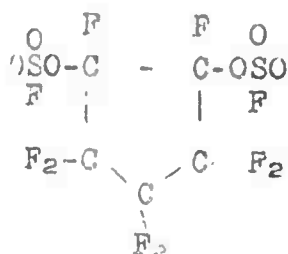
13 P. Torkington and H. W. Thompson, Trans. Faraday Soc., 41, 236 (1945).

(7.97 μ) were attributed to S-O stretch and at 853 cm^{-1} (1173 μ) to S-F stretch, and at 1217 cm^{-1} (8.22 μ) to C-F stretch. Weak bands at 1410 cm^{-1} (7.05 μ) and at 1152 cm^{-1} (8.68 μ) were also attributed to C-F stretch. Unidentified bands occurred at 365 cm^{-1} (7.32 μ), 1069 cm^{-1} (9.35 μ), 990 cm^{-1} (10.11 μ) and 801 cm^{-1} (12.47 μ).

(f) Mass spectrum. The principal peaks in the spectrum for mass numbers above 55 corresponded to the following ions: SO_2^+ , SOF^+ , CF_3^+ , SO_2F^+ , $C_3F_3^+$, $C_2F_3O^+$, $C_2F_4^+$, $C_3F_3O^+$, $C_2F_5^+$, $C_3F_7^+$, $C_4F_5O^+$, $C_3F_7^+$, $C_4F_7^+$, $C_4F_7O^+$, and $C_5F_8^+$. A parent peak ($C_5F_8(SO_3F)_2^+$), although very weak, was noted at mass number 410.

(g) Nuclear magnetic resonance spectrum. A comparison of areas under absorption peaks for $C_5F_8(SO_3F)_2$ and for $S_2O_6F_2$ gave 2.3 for the number of S-F bonds per molecule of $C_5F_8(SO_3F)_2$. The ratio of S-F/C-F was 1/3.9 compared to theoretical 1/4.

(h) Elementary analysis. Sulfur found: 14.9%; calculated: 15.6%. The sample was treated by potassium fusion in preparation for analysis. It apparently did not react with water.



Reaction with tetrafluoroethylene to form tetrafluorobis(fluorosulfonato)ethane.

Since peroxides act as initiators for the polymerization of C_2F_4 to polytetrafluoroethylene, only low pressures of C_2F_4 were used and during the reaction $S_2O_8F_2$ was always present in excess. The reaction was extremely vigorous and was carried out by slowly adding tetrafluoroethylene to $S_2O_8F_2$ containing 10% nitrogen by volume to reduce the reaction rate. Although $S_2O_8F_2$ was present in excess at all times to insure minimum polymerization of C_2F_4 , the entrance to the reaction vessel and other parts of the vacuum line were soon coated with a white solid. When the addition of C_2F_4 was very slow, a white smoke was observed which gradually condensed into fine droplets which very slowly enlarged and collected at the bottom of the vessel. A slightly more rapid addition resulted in a large temperature rise and quantities of brown color (probably due to SO_3F radicals). Separation and identification of products indicated the following reactions: $C_2F_4 + S_2O_8F_2 = C_2F_4(SO_3F)_2$ and $C_2F_4 + 2S_2O_8F_2 = 2COF_2 + 2S_2O_5F_2$. Polymeric products were not identified. The products, carbonyl fluoride and pyrosulfuryl fluoride, of the second reaction constituted only about 5% of the total product. After the sample had been purified by fractional codistillation, a portion was subjected to gas chromatography using a tricresyl phosphate-fire brick column at 72°. Since a single peak was obtained, it was concluded that the distilled material was over 99% pure.

Properties of tetrafluorobis(fluorosulfonato)ethane.

(a) Molecular weight. Molecular weight determinations on the sample gave an average of 298.2 compared to a calculated value of 298.15 for $C_2F_4(SO_3F)_2$.

(b) Density. The density of $C_2F_4(SO_3F)_2$ was determined at the four temperatures given in Table III. The density at 25° by interpolation, was 1.778 g/cc.

Table III

Densities of $C_2F_4(SO_3F)_2$

Temp., °C	1.0	12.5	24.3	35.4
Density, g/cc	1.839	1.811	1.780	1.750

The volume coefficient of expansion, calculated from the above data, was 1.49×10^{-3} at 25°.

(c) Melting point. $C_2F_4(SO_3F)_2$ melted at -27.8°.

(d) Vapor pressure and boiling point. The vapor pressures given in Table IV indicate a normal boiling point of 102.7°. From the Clausius-Clapeyron equation, the molar heat of vaporization was 9.1 kcal and the Trouton constant 24.1.

Table IV

Vapor pressures of $C_2F_4(SO_3F)_2$

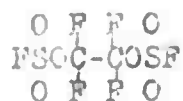
Pmm	T, °K	Pmm	T, °K
21.8	295.8	371.3	355.1
46.0	306.9	424.1	359.0
82.7	318.0	496.8	363.7
124.1	327.4	561.1	367.0
177.8	336.0	641.4	370.9
247.6	343.5	693.8	373.2
283.6	348.6	743	375.7
		(760)	(375.9)
		787	377.2

(e) Infrared spectrum. Infrared spectra were obtained at 1 and 10 mm pressure. By comparing the frequencies of the tetrafluorobis(fluorosulfonato)ethane with those of the reactants,^{4,13} the molecular motions for some of the bands were assigned. Strong bands at 1496 cm^{-1} ($6.68\text{ }\mu$) and 1263 cm^{-1} ($7.92\text{ }\mu$) were attributed to S-O stretch, while those at 905 cm^{-1} ($11.04\text{ }\mu$) and 846 cm^{-1} ($11.81\text{ }\mu$) were due to S-F motions. A C-F stretch was responsible for the band at 1163 cm^{-1} ($8.60\text{ }\mu$). Other strong bands occur at 1230 cm^{-1} ($8.13\text{ }\mu$), 1077 cm^{-1} ($9.28\text{ }\mu$) and 758 cm^{-1} ($13.20\text{ }\mu$).

(f) Nuclear magnetic resonance spectrum. Using perfluorocyclopentane (C_5F_{10}) as a standard, the number of C-F bonds found per molecule of $\text{C}_2\text{F}_6(\text{SO}_3\text{F})_2$ was 3.4 (theor. 4.0). The experimental number of S-F bonds was 1.8 (theor. 2.0) when $\text{S}_2\text{O}_8\text{F}_2$ was used as the standard.

(g) Elementary analysis. Sulfur content found was 20.6% compared to 21.5% calculated. The sample was prepared for analysis by heating with potassium.

General. Tetrafluorobis(fluorosulfonato)ethane was observed to be a colorless liquid which dissolved Kel F stopcock grease but was inert toward mercury and glass and, for at least short periods of contact, toward water. The properties were consistent with the structure:



Reaction with sulfur tetrafluoride to form tetrafluorobis(fluorosulfonato)sulfur (VI).

Peroxydisulfuryldifluoride did not react with sulfur tetrafluoride at room temperature. To study the effect of increasing

temperature, equivalent amounts of SF_4 and $\text{S}_2\text{O}_6\text{F}_2$ were treated as described under Procedure 2 above. At 85° , the pressure began to fall gradually. At 106° , the decrease was very rapid. In subsequent runs the reactor was maintained at 128° while equal pressures of the reactants were added and allowed to remain until the pressure stopped decreasing. An 86% yield of tetrafluorobis(fluorosulfonato)-sulfur (VI) was obtained based on the reaction $\text{SF}_4 + \text{S}_2\text{O}_6\text{F}_2 \xrightarrow{128^\circ} \text{SF}_4(\text{SO}_3\text{F})_2$. This indicated that the SF_4 as received was at least 86% pure. SOF_2 was the main contaminant in the SF_4 as shown by vapor density and infrared spectrum determinations. The more volatile SOF_2 was removed by distillation leaving pure $\text{SF}_4(\text{SO}_3\text{F})_2$ which required no further purification.

Properties of tetrafluorobis(fluorosulfonato)sulfur (VI).

- (a) Molecular weight. Molecular weight determinations on the sample gave an average of 305.8 compared to 306.12 for $\text{SF}_4(\text{SO}_3\text{F})_2$.
- (b) Density. Densities were measured at the three temperatures given in Table V. From these data the extrapolated density at 25° was found to be 1.984 g/cc.

Table V

Densities of $\text{SF}_4(\text{SO}_3\text{F})_2$

Temp. $^\circ\text{C}$	0.0	10.6	22.9
Density, g/cc	2.049	2.022	1.989

The volume coefficient of expansion was calculated from the above data to be 1.34×10^{-3} at 25° .

- (c) Melting point. The melting point found for $\text{SF}_4(\text{SO}_3\text{F})_2$ was -62.9° .
- (d) Vapor pressure and boiling point. A normal boiling point of 116.6° was obtained from the vapor pressures given in Table VI. The molar heat of vaporization from vapor pressure data was 9.3 kcal. and the Trouton constant was 23.9.

Table VI

Vapor Pressures of $\text{SF}_4(\text{SO}_3\text{F})_2$

Pmm	T, °K	Pmm	T, °K
22.9	305.8	272.4	359.0
30.7	310.5	330.8	363.7
45.2	317.7	415.6	370.8
60.0	322.3	455.4	373.9
89.2	331.2	532.1	378.1
99.7	333.6	634.8	384.2
149.8	343.5	690.1	386.9
189.1	349.6	724.2	388.6
232.2	354.4	(760)	(389.8)
		769	390.4
		800	391.5

(e) Infrared spectrum. Infrared spectra were obtained at 1 and 10 mm pressure. A comparison of the spectrum of $\text{SF}_4(\text{SO}_3\text{F})_2$ with those of the reactants^{4,14} enabled identification of the molecular

14 R. E. Dodd, L. A. Woodward and H. L. Roberts, Trans. Faraday Soc., 52, 1052 (1956).

motions responsible for some of the bands. S-O stretch appeared at 1491 cm^{-1} ($6.71\text{ }\mu$) and 1251 cm^{-1} ($7.99\text{ }\mu$) while S-F stretch at 887 cm^{-1} ($11.28\text{ }\mu$) and 839 cm^{-1} ($12.06\text{ }\mu$). Other bands occurred at 943 cm^{-1} ($10.60\text{ }\mu$) and 704 cm^{-1} ($14.20\text{ }\mu$).

(f) Mass spectrum. The mass spectrum of $\text{SF}_4(\text{SO}_3\text{F})_2$ above 56 included the following ions: SO_2^+ , SOF^+ , SF_2^+ , SO_3^+ , SO_2F^+ , SF_3^+ , SF_3O^+ , and SF_4^+ . No peaks for masses greater than that of SF_4^+ were noted.

(g) Nuclear magnetic resonance spectrum. The NMR spectrum was run at 23° and at a frequency of 40 megacycles. The spectrum contained fluorine resonance peaks for the SF_4 group and for the SO_3F groups.

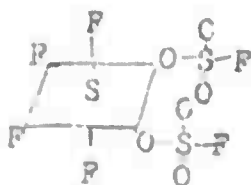
These corresponded to a large "chemical shift", the peaks for SO_3F coming at a higher magnetic field than those for the SF_4 group. The SF_4 group caused two symmetrical clusters of peaks. The two clusters of peaks suggested that two pairs of structurally different fluorine atoms were present in the $>\text{SF}_4$ group. Since the clusters were not triplets, it appeared that the spin-spin coupling constant, J , and the chemical shift, δ , were of the same order of magnitude. Professor Kenneth B. Wiberg, a colleague of the authors, has kindly applied his computer program to the problem and has calculated the theoretical spectrum for the $>\text{SF}_4$ group based upon a value of δ of 501 cycles per second. Various values of J were tried, the best agreement with the experimental spectrum being obtained for $J = 156$ cycles per second. Agreement between theory and experiment was good enough to validate the assumed structure of the $>\text{SF}_4$ group. It was therefore concluded that the fluorosulfonate groups occupied cis positions in an octahedral structure about the sulfur atom. If the fluorosulfonate groups had been trans, the fluorine atoms of the $>\text{SF}_4$ group would probably have been identical. Spin-spin splitting due to the fluorine atoms of the fluorosulfonate groups caused the "lines" of the SF_4 group to be triplets.

(h) Elementary analysis. When the compound was hydrolyzed in sodium hydroxide solution at 100° it was found that 11.6 g. equivalents of base (theoretical, 12.0) were consumed per mole as shown in the equation:



The sulfur and fluorine content of the compound (as found in the hydrolyzed product) were: S, 30.4%; F, 37.6% (Theor. S, 31.4%; F, 37.2%).

(1) General. The compound was found to be a clear, colorless liquid which was inert toward mercury but apparently attacked glass slowly and was readily dissolved in Kel-F stopcock grease. The properties were consistent with the structure:



Reactions Involving Both Fluorosulfonation and Oxygenation

With some compounds peroxodisulfuryl difluoride acted both as an oxygenating and a fluorosulfonating agent through the addition of one or more oxygen atoms and one or more fluorosulfonate groups to the central atom of the reacting species.

Reaction with thionyl chlorofluoride.

Thionyl chlorofluoride is the only known thionyl compound which contains two different groups attached to the central atom. Although it seemed that $S_2O_2F_2$ might replace the chlorine with a fluorosulfonate group and give rise to a second such compound, this was not found to be the case.

The reaction was quite exothermic and the gas in the bulb acquired a slight greenish tinge. Products found were chlorine (green, M.W. 72)

and $S_2O_6F_2$ (I.R. as above) indicating the reaction: $2SOClF + 3S_2O_6F_2 \rightarrow 4S_2O_6F_2 + Cl_2$.

Reaction with molybdenum hexacarbonyl or molybdenum metal to form dioxobis(fluorosulfonato) molybdenum (VI).

As was mentioned earlier, the reaction of $S_2O_6F_2$ with $Mo(CO)_6$ proceeded rapidly and was accompanied by the quantitative displacement of the carbonyl groups to form carbon dioxide, and by evolution of heat. As the reaction progressed, the white carbonyl, which had been sublimed into the reaction vessel, gradually became emerald green. Finally a bright orange liquid and a white solid were produced. On occasion, if the gaseous products were not almost continuously removed, the bulb containing the reaction mixture was completely destroyed due to the sudden and rapid release of carbon monoxide and carbon dioxide. This removal required frequent cycles of holding the vessel at -78° while the gas was pumped away and then warming the bulb to room temperature to allow the reaction to continue.

Because of the inconvenience associated with this reaction, the reaction of excess $S_2O_6F_2$ with molybdenum metal was tried. It was found that at room temperature the reaction occurred slowly but that at $60-85^\circ$, the metal was rather rapidly converted to a white solid and a solution of a bright orange substance in liquid $S_2O_6F_2$. The reaction was allowed to occur in a fifty or one hundred ml Pyrex glass bulb equipped with a 2 mm Kern stopcock attached to a 10/30 inner standard taper joint. Two break seals were also attached to the bulb. Finely divided molybdenum metal was weighed into the bulb and the stopcock and break seals were then attached. The bulb was evacuated and, while pumping was continued, was heated strongly to remove any trace of moisture. Peroxylisulfuryl difluoride was distilled on to the metal until present in at least a three-

fold excess. The stopcock was pulled off and the bulb was maintained at 75° until all visible reaction had ceased (no metal left). Through use of one break seal, the excess $S_2O_5F_2$ and other volatile materials (largely $S_2O_5F_2$) were removed. The white solid was still moist with an orange liquid. The bulb was then wrapped with uniformly spaced nichrome resistance wire and covered with asbestos tape. When a small electric current was passed through the wire to warm the bulb, the more volatile orange liquid was distilled under good vacuum into the side arm to which the second break seal was attached. Heating was continued until the solid became white or cream colored, and the side arm was then pulled off. The bulb was opened in a dry box and portions of the solid were placed in a weighed bulb and in a melting point tube. A sample of the solid (1.3419 grams) was hydrolyzed and analyzed for molybdenum, sulfur and fluorine. Oxygen was determined by difference. The ratio, in terms of atoms, found for Mo:S:F:O was 1.0:2.0:2.1:8.0 compared to 1.0:2.0:2.0:8.0 theoretical for $MoO_2(SO_3F)_2$. The experimentally obtained percentage composition for dioxobis (fluorosulfonate) molybdenum (VI) was Mo, 29.13%; S, 19.76%; F, 12.22%; and O, 38.89%. Calculated: Mo, 29.42%; S, 19.66%; F, 11.65% and O, 39.24%. In an attempt to determine the melting point the tube was placed in a furnace and the temperature was allowed to increase slowly. At about 230° decomposition to a blue solid occurred and the Pyrex tube was badly etched.

The very viscous orange liquid (collected in the side arm as stated above) was transferred under anhydrous conditions and 0.2067 gram was subjected to basic hydrolysis. The experimental ratio of Mo: SO_3F :O was 1:2.3:0.7. This suggests the formula $MoO(SO_3F)_2$. It is very unlikely,

however, that this compound would: (1) exist under such highly oxidizing conditions; (2) show a definite but somewhat broadened NMR spectrum (the two unpaired electrons in a Mo(IV) compound would prevent one from obtaining an NMR spectrum);;and (3) be a liquid which melts sharply at approximately -65° when the analogous oxyfluoride is a high melting solid. Further work is indicated.

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