Organic Disulfides and Related Substances. XXVII. Reactions and Synthetic Utility of Cyclic Disulfides, Dioxides, and Tetroxides[1a–c]

LAMAR FIELD AND ROBERT B. BARBEE[1d]

Department of Chemistry, Vanderbilt University, Nashville, Tennessee 37203

Received September 20, 1968

Hydrolysis, polarographic reduction, and other reactions were studied of the unsubstituted five-, six-, and seven-membered disulfides, the 1,1-dioxides, and the 1,1,2,2-tetroxides. 1,2-Dithiane 1,1,2,2-tetroxide reacted less readily with thiophenol than the 1,1-dioxide but oxidized a thiolate quantitatively to the disulfide by a mild method of possible general use; its reactivity with nucleophiles resembled that of a disulfide, except for greater susceptibility to alkali (all three tetroxides were readily cleaved at pH 8); its pyrolysis gave tetrahydrothiophene dioxide but in low yield. Generalizations are difficult but seem usually to be for easier cleavage of the five-membered systems than of the six-membered ones (with the seven-membered systems variable) and for greater resistance to self-polymerization or to attack of a thiol by the more oxidized forms but for lesser resistance to hydrolysis and electrochemical reduction. The dioxides and tetroxides are quite promising intermediates for synthesis. 1,2-Dithiane 1,1-dioxide underwent "oxodisulfide cleavage" by thiolate ion to give disulfides **1**, **2**, and **3** containing a sulfinate moiety, which in turn were converted into alkyl sulfones (**4** and **5**), an aryl sulfone (**6**), or a sulfonate (**8**). A typical disulfide product (**4**) disproportionated to the symmetrical disulfides comparably with resistant classes, affording a synthesis of a disulfide sulfone. 1,2-Dithiane tetroxide underwent "oxodisulfide cleavage" with alkali to give a sulfonate salt containing a sulfinate moiety (**11**), which was converted into a disulfide dioxide (**12**). Product **3** was active as an antiradiation drug.

The previous paper in this series compared syntheses for the five-, six-, and seven-numbered cyclic disulfides (1,2-dithiolane, 1,2-dithiane, and 1,2-dithiepane), their 1,1-dioxides, and their 1,1,2,2-tetroxides.[1c] The present paper first considers relative stabilities and reactivities of these substances under various circumstances. It then demonstrates utility of such substances for synthesis of a variety of structures otherwise difficultly obtainable. Whether the future will justify the hope that model reactions of the types described will prove general for cyclic disulfides and their oxidation products remains to be seen, but we are optimistic that these prototypes will suggest useful approaches for many ring-size and ring-substitution situations.

Instability of 1,2-dithiolane, apparently at least partly a consequence of repulsion between nonbonding electrons on the two sulfur atoms,[2] originally led us to wonder whether the 1,1,2,2-tetroxide would be likewise destabilized (made more reactive) by interaction of the four negatively charged oxygen atoms and how the 1,1-dioxide would compare with the disulfide and tetroxide (since it lacks these nonbonding electrons of the disulfide on one of its sulfur atoms and the two oxygen atoms of the tetroxide on the other). Some evidence on these points was adduced previously: 1,2-dithiolane monoxide is stable as a liquid for at least 2 hr and the dioxide for at least a month,[1c] in contrast to 1,2-dithiolane which polymerizes rapidly if solvent is removed.[2,3] That 1,2-dithiolane 1,1,2,2-tetroxide is quite stable (melting point unchanged after 9 months) also is noteworthy;[1c] the fact that it is a solid probably is important to its stability, but no polymerization was evident when either the dioxide or tetroxide was recrystallized.[1c] Other questions also are intriguing, such as reactivity of the oxides with nucleophiles. Even though it may be some time before rigorously satisfying answers to such questions as the foregoing can be had, information bearing on them seems likely to have fairly immediate practical use.

The desire for convenient, preparatively significant assessments of the influence of ring size and oxidation level on properties of the disulfides, the 1,1-dioxides, and the 1,1,2,2-tetroxides led first to a study of hydrolysis in refluxing aqueous dioxane. Although the hydrolyses undoubtedly are complicated by side reactions (*e.g.*, of sulfenic and sulfinic acids), the speed with which acidic groups were formed (measured by alkali consumption) was taken as a reasonable measure of reactivity in hydrolysis for the dioxides and tetroxides. Results are shown in Figure 1.

The three disulfides were followed by using uv spectra instead of by titration. Since all apparently were characterized by nearly complete resistance to the hydrolytic conditions, results with them are shown for comparison as the flat base line in Figure 1. Thus, after 50 hr, 100% of 1,2-dithiolane survived and change was insignificant even after 70 hr (in confirmation of identity at the latter point, a sample underwent 90% decrease in its optical density at 330 mμ upon exposure to sunlight, a property useful for distinguishing the dithiolane).[1c] 1,2-Dithiane and 1,2-dithiepane showed no change in absorption after 48 and 52 hr, and only little after 120 hr.

Figure 1 suggests that the dioxides were hydrolyzed more readily than the disulfides but still relatively slowly. Qualitative checks on this conclusion, obtained by allowing the dioxides to react with excess thiol and titrating the sulfinic acid formed in the manner of Barnard and Cole,[4] revealed that 84% of the dithiane dioxide remained after 75 hr, 72% of the dithiolane dioxide after 51 hr, and 62% of the dithiepane dioxide after 46 hr; both these results and the data shown in Figure 1 suggest that 1,2-dithiane dioxide is the least reactive of the three dioxides, but that all are fairly resistant and roughly comparable.

Figure 1 suggests that the tetroxides are hydrolyzed much more readily than the dioxides, and that reactivity increases in order six < seven < five. For these three,

---

[1] (a) This investigation was supported by the U. S. Army Medical Research and Development Command, Department of the Army, under Research Contract No. DA-49-193-MD-2030. (b) Taken mainly from the Ph.D. dissertation of R. B. B., Vanderbilt University, Aug 1968; the dissertation may be consulted for greater detail. (c) Paper XXVI: L. Field and R. B. Barbee *J. Org. Chem.*, **34**, 36 (1969). (d) Eastman Kodak Fellow, 1965–1966.

[2] J. A. Barltrop, P. M. Hayes, and M. Calvin, *J. Amer. Chem. Soc.*, **76**, 4348 (1954).

[3] A. Schöberl and H. Gräfje, *Ann.*, **614**, 66 (1958).

[4] D. Barnard and E. R. Cole, *Anal. Chim. Acta*, **20**, 540 (1959).

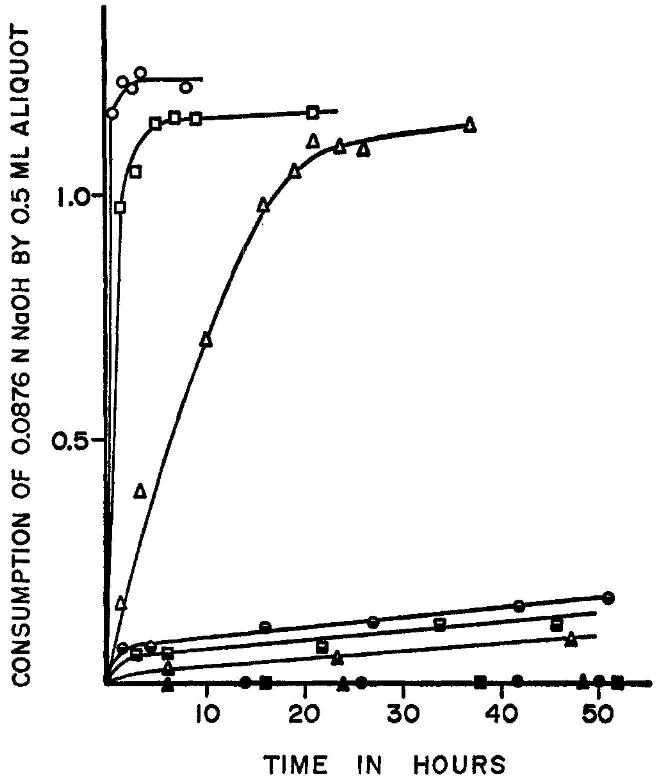


Figure 1.—Hydrolysis of 0.11 *M* solutions in 2:1 dioxane–water of cyclic sulfur compounds at 100°: ○, five-membered ring; △, six-membered ring; □, seven-membered ring; filled figure, disulfide; half-filled figure, disulfide dioxide; open figure, disulfide tetroxide. Hydrolysis of the cyclic disulfides was followed spectroscopically.

incidentally, alkali consumption (after a nearly constant value was observed) amounted to 87, 90, and 96%, respectively, of expectation of hydrolysis, assuming that eq 1 applies. Recovery of only 19% of 1,2-dithiane

$$(\overline{CH_2)_nSO_2S}O_2 + H_2O \longrightarrow HO_3S(CH_2)_nSO_2H \qquad (1)$$

1,1,2,2-tetroxide after 21 hr qualitatively confirmed its relatively rapid hydrolysis, and recovery of 85% of 1,2-dithiane 1,1-dioxide after 21 hr confirmed that a dioxide is less reactive than a tetroxide.

Reactions of the dioxides and tetroxides with nucleophiles were considered next. Like arenethiolsulfonates,[5] 1,2-dithiane 1,1-dioxide reacted with thiophenol in ether according to eq 2 but, again similarly, not to very

$$(\overline{CH_2)_4SO_2S} + C_6H_5SH \longrightarrow HO_2S(CH_2)_4SSC_6H_5 \qquad (2)$$

great extent. After 1 hr with 1 molar proportion of thiophenol, the reaction of the dioxide was about 16% complete, based on an average from unreacted thiophenol, titration of sulfinic acid formed, and isolation of the unchanged dioxide; phenyl benzenethiolsulfonate reacted under rather similar conditions to the roughly comparable extent of 27–31%.[5] The extent of reaction of the dithiane dioxide increased to 65% in ethanol and to 86% in ethanol with 2 molar proportions of thiophenol, points noteworthy for their possible general applicability to reactions of disulfide dioxides.

In contrast to the reaction extent of 86% in ethanol shown by the dioxide (1 hr, *ca.* 25°), only 15% of 1,2-dithiane tetroxide reacted in 4 days; since the tetroxide was only partly soluble in ethanol, however, the reaction also was done homogeneously for 1 hr in acetonitrile, but no more than 7% reacted. In a 20-hr experiment (different concentrations), 9% reacted. In an identical 20-hr experiment, but in ethanol, 1,2-dithiepane tetroxide also was unreactive with thiophenol, 10% reacting. Dithiolane tetroxide was more reactive than the other two tetroxides, however, reacting to the extent of 50% (20 hr, ethanol); unlike the other two tetroxides, it also gave a positive acidic reaction with thiophenol in the Barnard–Cole test for disulfide dioxides.[4]

Although the 1,2-dithiane tetroxide resisted thiophenol, it was readily cleaved by 2 equiv of sodium thiophenoxide, with quantitative formation of phenyl disulfide and disodium 1,4-butanedisulfinate (eq 3).

$$(\overline{CH_2)_4SO_2S}O_2 \xrightarrow{C_6H_5SNa} [NaO_2S(CH_2)_4SO_2SC_6H_5] \xrightarrow{C_6H_5SNa}$$
$$NaO_2S(CH_2)_4SO_2Na + (C_6H_5S)_2 \qquad (3)$$

This mild means of quantitatively oxidizing a thiol to its disulfide has attractive general possibilities worthy of further exploration, particularly for biochemical systems.

To gain guidelines to its relative reactivity with typical nucleophiles, 1,2-dithiane tetroxide was treated with several potassium salts in aqueous acetonitrile. Starting material unconverted to a sulfinate salt (removed with water) then was recovered (*cf.* eq 3). To the extent such recovery signifies the order of reactivity, relative S nucleophilicities were in the following order: $C_6H_5S^-$, $OH^- > CN^- > SCN^-$, I^-.

Except that OH^- is much higher in this series, the order is that for attack upon an aryl disulfide;[6] formation of a sulfene intermediate with the tetroxide may explain the greater reactivity of OH^- and, indeed, cannot be discounted for other ions in the series. Since potassium hydroxide dissolved only partly, and perhaps because of unrecognized complications as well, this order is best regarded merely as an approximation of S nucleophilicity, pointing only to high reactivity of cyanide, hydroxide, and thiophenoxide ions and low reactivity of iodide and thiocyanate ions.

In initial attempts to carry this aspect further by comparing stabilities of the three tetroxides toward hydroxyl ion, 1,2-dithiane tetroxide reacted very rapidly at 60° (conversely, the pH of an aqueous alcoholic solution of the dioxide at pH 10 did not change during 2 hr at 60°, reflecting again the lower reactivity of the dioxide noted in the discussion of Figure 1).

Relative reactivities of the cyclic disulfide tetroxides with alkali at room temperature therefore were examined by using a pH-Stat to follow the amount of alkali required to maintain pH 8. The times at which uptake of alkali ceased (in parentheses) suggest the following order of increasing reactivity for the three tetroxides: six (30 min) < five (20 min) $\lesssim$ seven (15 min); since the curves for the five- and seven-membered compounds were much alike however,[1b] as they were in Figure 1, their difference probably is less significant than the times imply. In this reaction, too, sulfene intermediates

---

[5] T. F. Parsons, J. D. Buckman, D. E. Pearson, and L. Field, *J. Org. Chem.*, **30**, 1923 (1965).

[6] A. J. Parker and N. Kharasch, *J. Amer. Chem. Soc.*, **82**, 3071 (1960).

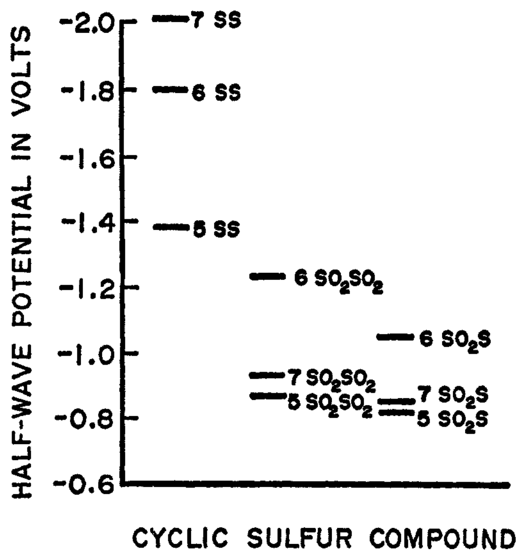


Figure 2.—Correlations of oxidation state and ring size of cyclic sulfur compounds with polarographic meaurements in diglyme.

can by no means be excluded at present; they are well established in not dissimilar reactions,[7] and indeed we found 1,2-dithiane tetroxide to be destroyed by triethylamine in refluxing acetonitrile (1 hr), with development of a striking red color seen also during formation of a sulfene.[8]

The polarographic behavior of the disulfides, dioxides, and tetroxides next was examined, because of its possible bearing on the relative reactivity of the various ring sizes and oxidation levels. Figure 2 depicts relationships which emerged from the polarographic reduction of the nine compounds in diglyme.[9] Figure 2 implies that increasing ease of reduction falls in the order $SS < SO_2SO_2 < SO_2S$; Lindberg and Bergson likewise found a dithiolane to be more resistant toward reduction than its dioxide at pH 7.2 (although less so at pH 1.2).[10] Figure 2 also implies that, although the potentials for reduction of the six- and seven-membered rings of a given oxidation level varied relative to each other, these rings always were more resistant to reduction than the five-membered ring. The polarographic waves of all nine compounds showed that each compound was irreversibly reduced, and thus that the reduction processes are slow in diglyme.

With the disulfides, the apparent increasing ease of electrochemical reduction of the rings (seven < six < five) agrees with data of Nygård and Schotte for the α,α'-dicarboxylic acids in buffer.[11] Although Nygård and Schotte stated that a greater ring stability was apparent for the six- and seven-membered cyclic disulfides compared with the five-membered one,[11a] Schotte pointed out that apparent involvement of intermediate mercury complexes made a direct comparison inadequate for establishment of possible differences in ring stability.[11b]

Although the half-wave potentials for the unsubstituted disulfides (seven, −2.01 V; six, −1.80 V; five, −1.36 V) are higher than values of Nygård and Schotte for their substituted disulfides (seven, −0.92 V; six, −0.75 V; five, −0.33 V),[11a] the order is the same; ethyl disulfide in diglyme had a half-wave potential of −1.99 V.

With the dioxides, reduction waves were characterized by a second small wave near −1.8 V of undetermined origin. The half-wave potentials of the primary waves suggests the following order of increasing ease of reduction (half-wave potentials in volts in parentheses): six (−1.05) < seven (−0.85) < five (−0.83).

With the tetroxides, the polarographic waves were characterized by a small wave prior to the appearance of the normal wave, suggesting adsorption. Apparent ease of reduction increased in the order six (−1.23) < seven (−0.93) < five (−0.88).

For rigorous comparison of half-wave potentials, the mechanism of electrode processes must be the same. This requirement seems satisfied in comparing potentials of those cyclic sulfur systems which contained the same electroactive group. Thus the implication of the data seems valid that in any one oxidation level the six- or seven-membered ring is less easily reduced than the five-membered one. Unfortunately for comparison of relative reactivity of the linkages $-SS-$, $-SO_2S-$, and $-SO_2SO_2-$ in a given ring size, differences in the polarographic curves suggest that different mechanisms of reduction may be operative. Hence the conclusion from Figure 2 that the ease of electrochemical reduction necessarily increases in the order $SS < SO_2SO_2 < SO_2S$ must be regarded with considerable caution, since "Any comparison of the reactivity based on a comparison of half-wave potentials corresponding to different groups following different reduction paths is only a rough approximation." [12]

An effort to trace patterns in the effects of ring size and oxidation level may be ventured, but with a strong reminder of possible pitfalls and exceptions noted both earlier[1c] and above. With respect to *ring size*, the tendency usually seems to be for the six-membered systems to be the least reactive of the three, and for the five-membered systems to be the most reactive, with the seven-membered systems taking a variable position. Among the disulfides, the order of Schöberl and Gräfje for increasing reactivity (six < seven < five)[3] was seen earlier in that the six-membered disulfide could be formed in situations where the five- and seven-membered disulfides formed in low yield or not at all, in oxidation of the six-membered-disulfide to the dioxide in higher yield, and in that only the six-membered disulfide could be oxidized directly to the tetroxide;[1c] presumably, the five- and seven-membered systems are more reactive, and hence more prone to undergo undesired reactions. Among the dioxides, the trend toward increasing ease of ring cleavage in the order six < seven < five seemed to be reflected for the more nearly coplanar five-membered one in its longer wavelength uv absorption[1c] and its easier electrochemical reduction, and for the six-membered one in better conversion of the dioxide into tetroxide[1c] and in its

---

(7) T. J. Wallace, *Quart. Rev.* (London), 67 (1966).

(8) W. E. Truce and J. R. Norell, *J. Amer. Chem. Soc.*, **85**, 3231 (1963).

(9) We are indebted for these measurements to Dr. L. C. Hall and Dr. G. J. Clark of Vanderbilt University. Details may be found in the forthcoming Ph.D. dissertation by G. J. Clark. Thanks also are due to Dr. Hall and Dr. Clark for many helpful discussions on interpretation of the results.

(10) B. Lindberg and G. Bergson, *Ark. Kemi*, **23**, 319 (1965).

(11) (a) B. Nygård and L. Schotte, *Acta Chem. Scand.*, **10**, 469 (1956); (b) L. Schotte, *Ark. Kemi*, **9**, 441 (1956).

(12) P. Zuman, *Chem. Eng. News*, **46** (No. 13), 94 (1968).

somewhat greater resistance to simple hydrolysis. Among the tetroxides, the six-membered system followed the order of increasing reactivity six < seven < five in simple hydrolysis and in reaction with alkali, and the five-membered system did so in reaction with a thiol. Since the oxides of the disulfides thus seem to follow the reactivity sequence of the disulfides, it may be reasonable to conclude that destabilizing effects in the disulfides have counterparts in the oxides. Respecting the effect of *oxidation level*, the presence of oxygen atoms seemed to increase resistance toward thiols and toward self-polymerization (making the five-membered cyclic monoxide, dioxide, and tetroxide tractable, in contrast to the unstable disulfide,[1c] and probably improving resistance of the six-membered systems as well). On the other hand, presence of oxygen atoms seemed to decrease resistance toward hydrolysis and toward electrochemical reduction in diglyme.

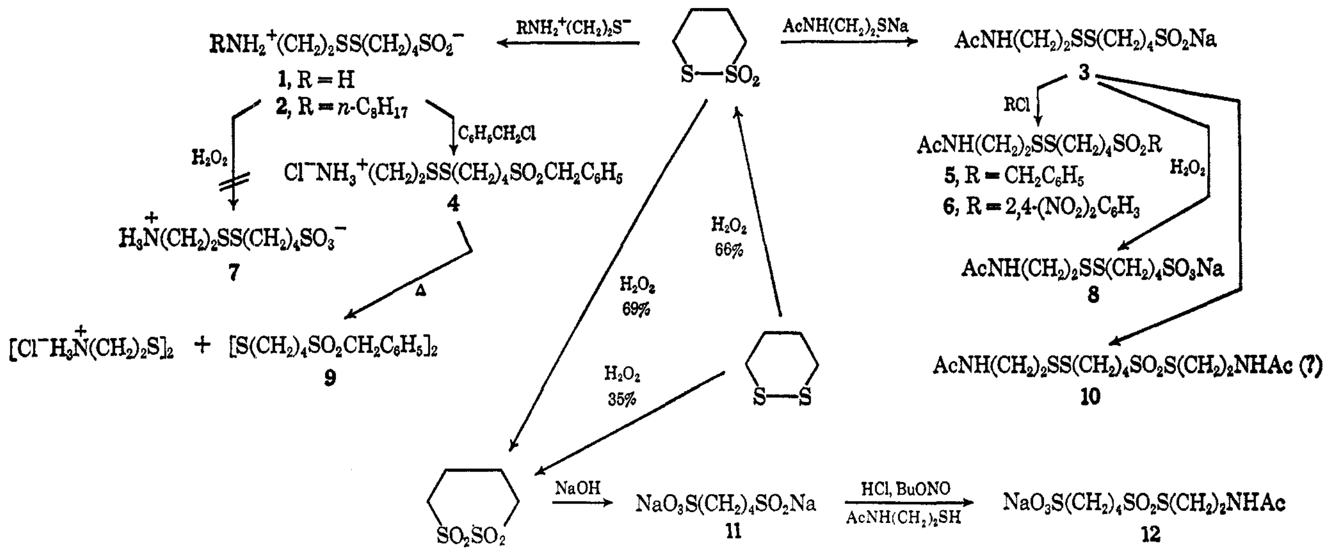
Elimination of sulfur dioxide by thermal decomposition has been reported for a variety of organic sulfur compounds.[13] Such a reaction occurs when 1,2-dithiane tetroxide is heated briefly at 280° (eq 4). The yield of

$$(CH_2)_4SO_2SO_2 \longrightarrow (CH_2)_4SO_2 + SO_2 \qquad (4)$$

tetrahydrothiophene 1,1-dioxide was only 12%, however, and could not be improved. Both this tetroxide and 1,2-dithiane 1,1-dioxide actually are rather stable thermally, being recovered in 92–95% yield after 7 days in chlorobenzene at 132°. 1,2-Dithiepane 1,1,2,2-tetroxide apparently gave no pentamethylene sulfide 1,1-dioxide when heated at 280° for 10 min; the dithiolane tetroxide was not tried because of insufficient material.

Several applications of 1,2-dithane dioxide and tetroxide to synthesis of otherwise difficultly obtainable classes of disulfides are shown in Scheme I. Thus the dioxide led to disulfides containing sulfininate, sulfonate, and aryl or alkyl sulfone moieties, and the tetroxide led to a disulfide dioxide containing the sulfonate moiety. (Yields of the dioxide and tetroxide shown are previous results.)[1c] Presumably, preparation of similar products containing a varied number of substituted or unsubstituted methylene groups greater than two separating the –SS– moiety and a functional group is possible, since cyclic disulfides containing from three to thirteen methylene groups are known.[3] Use of other known reactions of disulfide dioxides and tetroxides hopefully should permit extension of the general approach, which for convenience is called "oxodisulfide cleavage," to other nucleophiles than thiols to afford an even more broadly useful new synthetic tool.

SCHEME I



Acetamido- and aminothiols were used as models in these studies, as shown in Scheme I, in the hope that the products also might be of interest as antiradiation drugs (*cf.* earlier papers in this series). 2-Aminoethanethiol and 2-(n-octylamino)ethanethiol were sufficiently nucleophilic to convert 1,2-dithiane 1,1-dioxide into the disulfide sulfinates **1** and **2** in 86 and 97% yields, respectively; **1** and **2** are formulated in Scheme I as dipolar ions (but are named as aminosulfinic acids in the Experimental Section for convenience). 2-Acetamidoethanethiol was converted into the thiolate with a sodium alkoxide for an otherwise similar synthesis of the disulfide sulfinate **3** (94%). Infrared spectra of **1**, **2**, and **3**, as well as of other products described below were consistent with the structures assigned to Scheme I. The hydrochlorides of the aminothiols also could be used to prepare **1** and **2** in the presence of 1 equiv of sodium alkoxide.

"Oxodisulfide cleavage" can lead to products which are themselves useful intermediates. Formation of the disulfide sulfone **4** (11%) from the sulfinate **1** and benzyl chloride is an illustration. Sulfones such as **4** were of interest, since activity of prototypes as antiradiation drugs could lead to use of this flexible route for preparing a wide variety of ring-substituted counterparts. The reaction of disulfide sulfinate **3** with benzyl chloride similarly gave the disulfide sulfone **5** (57% yield). Homogeneity of **4** and **5** was confirmed by thin

---

[13] J. L. Kice in "The Chemistry of Organic Sulfur Compounds," Vol. 2, N. Kharasch and C. Y. Meyers, Ed., Pergamon Press, New York, N. Y., 1966, p 115.

layer chromatography (tlc). Conversion of disulfide sulfinate **3** into an aryl sulfone (**6**) was effected by its reaction with 2,4-dinitrochlorobenzene, but sulfone **6** proved extremely difficult to purify.

A further extension is the oxidation of disulfide sulfinates to form a third class of compounds, the disulfide sulfonates (Scheme I). Oxidation of disulfide sulfinate **3** with aqueous hydrogen peroxide produced the desired disulfide sulfonate **8** (26% yield). The major product appeared to be a disproportionation product, $[NaO_3S(CH_2)_4S]_2$, however, since it had strong ir absorption only at 1200 and 1060 cm^{-1} (characteristic of sulfonic acid salts) and contained no nitrogen. A similar result occurred upon oxidation of disulfide sulfinate **1** with hydrogen peroxide, although here only the presumed disproportionation product was isolated; the ir absorption spectra of the bissulfonate disproportionation product from oxidation of both **1** and **3** were identical. Sodium metaperiodate proved better for oxidizing the disulfide sulfinate **3** to the disulfide sulfonate **8**; the yield was 57%, with no observable disproportionation. With disulfide sulfinate **1** and the metaperiodate, however, disproportionation still predominated and little if any of the disulfide sulfonate **7** seemed to result.

Another use of the disulfide sulfinate **3** involved an attempt to convert it into a disulfide disulfide-dioxide (**10**) by a method introduced by Kresze and Kort[14] and subsequently developed for 2-acetamidoethanethiol by Field and Lacefield.[15] The method involves simultaneous oxidation of a thiol and a sulfinic acid by an alkyl nitrite. In it, disulfide sulfinate **3** was converted *in situ* into the sulfinic acid (ordinarily unstable as such) using hydrochloric acid, which in turn seemingly was converted into the disulfide dioxide hemihydrate **10** (37% yield). The identity as **10** was suggested by elemental analysis, the ir spectrum, and a molecular weight of 383 (calcd 398) by the method of Barnard and Cole,[4] and its homogeneity was indicated by tlc; however, a completely satisfactory elemental analysis could not be obtained despite several attempts and some reaction products showed two close tlc spots.

The susceptibility of 1,2-dithiane tetroxide to nucleophilic attack suggested the synthetic application outlined in Scheme I. Cleavage of the tetroxide by alkali provided the sulfinate sulfonate **11**, from which the sulfinic acid was generated *in situ* with mineral acid. Application of the Kresze-Kort reaction to this acid and acetamidoethanethiol gave the disulfide dioxide sulfonate **12** (79% yield), which gave a positive disulfide dioxide test;[4] some difficulty was encountered in separation of the sodium salt **12** from sodium chloride. Sulfinate sulfonates like **11** should have numerous other uses as intermediates.

Compound **1** was rated no better than slightly active as an antiradiation drug (dose, 150 mg/kg; LD_{50}, 500 mg/kg), and **8** was rated only "fair" (dose, 800 mg/kg; LD_{50}, 1200 mg/kg).[16] However, the disulfide sulfinate **3** gave particularly promising results ("good," 87–100% survival, at a dose level of 47 mg/kg; LD_{50}, 694 mg/kg);[16] the other compounds are still to be tested.

Earlier papers of this series have considered factors which influence disproportionation of unsymmetrical disulfides to two symmetrical ones. To rank compounds of the present type among those studied earlier, the disulfide sulfone **4** was chosen as a representative, because one of the products, cystamine dihydrochloride, should be easily separable from the other (**9** of Scheme I) by washing with water. Thermal disproportionation of **4** was effected by heating it in water at 100° for 72 hr (Scheme I). The extent of disproportionation, determined by isolation of the water-insoluble **9**, was 79%. The value of 79% for **4** is comparable with that of 80% under much the same conditions obtained with 1,4-bis(2-aminoethyldithio)butane dihydrochloride.[19] This result suggests that compounds like **4** should be among the most stable of the disulfides studied.[19] The result also both confirms the structure of **4** and demonstrates a useful synthesis of disulfide sulfones typified by **9**.

Experimental Section[20]

Stability in Refluxing Aqueous Dioxane.—Each sulfur compound (1,2-dithiolane, 1,2-dithiane, and 1,2-dithiepane, their dioxides, and tetroxides) was dissolved in enough 66% dioxane–water to give an 0.11 *M* solution, which then was heated under reflux (100°). Samples of the dioxides and tetroxides (0.50 ml), withdrawn at suitable intervals, were titrated with 0.0876 *N* NaOH, and base consumption in milliliters for the 0.5-ml aliquot was plotted *vs.* time in Figure 1. With cyclic disulfide dioxides, the end point was determined using a Beckman pH meter with a calomel–glass electrode pair. The end point of each aliquot containing tetroxide was determined using bromophenol blue indicator; the Bernard–Cole titrations[4] for the dioxides were done after similar preliminary neutralization. With the cyclic disulfides, lack of hydrolysis was concluded by following the characteristic uv maxima at 330 mμ for 1,2-dithiolane, at 288 mμ for 1,2-dithiane, and at 258 mμ for 1,2-dithiepane.[1c] Isolations by evaporating a sample, washing the solid with water, and drying gave 1,2-dithiane 1,1-dioxide in 75% yield and the 1,1,2,2-tetroxide in 18% yield after 21 hr; reliability of this isolation in turn was checked by isolating the dioxide from an unheated control solution in 88% yield and the tetroxide in 93% yield [so that the corrected value for survival of the dioxide is (75) (100)/88 = 85%, and for the tetroxide is (18) (100)/93 = 19%].

Reaction of 1,2-Dithiane 1,1-Dioxide with Thiophenol.—In a typical experiment, thiophenol (0.110 g, 1.00 mmol) and 1,2-dithiane 1,1-dioxide (0.152 g, 1.00 mmol) were stirred in 5 ml of dry ether at *ca.* 25° for 1 hr. Ether and unreacted thiol then were distilled at 1 mm into a Dry Ice chilled trap. Iodine titration of the trap contents showed presence of 85% of the original thiophenol, representing a reaction not more than 15% complete. The undistilled residue was titrated with 0.5 ml of 0.0876 *N* NaOH to a bromophenol blue end point, which represents a reaction extent of not less than 4%. The solution then was extracted with ether and the extract was washed with water, dried, and evaporated to give 0.109 g (72%) of 1,2-dithiane 1,1-dioxide, mp 47–48° (lit.[21] mp 56–57°), indicating a reaction extent of not

---

[14] G. Kresze and W. Kort, *Chem. Ber.*, **94**, 2624 (1961).

[15] L. Field and W. B. Lacefield, *J. Org. Chem.*, **31**, 599 (1966).

[16] These results were kindly provided by Dr. D. P. Jacobus, Dr. T. R. Sweeney, and Dr. E. A. Steck of the Walter Reed Army Institute of Research. General procedures, the meaning of activity ratings, etc., are referred to in previous papers.[17,18]

[17] L. Field, H. K. Kim, and M. Bellas, *J. Med. Chem.*, **10**, 1166 (1967).

[18] L. Field, A. Ferretti, R. R. Crenshaw, and T. C. Owen, *ibid.*, **7**, 39 (1964).

[19] L. Field, A. Ferretti, and T. C. Owen, *J. Org. Chem.*, **29**, 2378 (1964).

[20] Details were as given in footnote 24 of ref 1c. Polarographic measurements were made of the five-, six-, and seven-membered disulfides, 1,1-dioxides, and tetroxides in diglyme with a Sargent polarograph Model XXI at 25° by Dr. Larry C. Hall and Dr. Gale J. Clark.[9] Tetrabutylammonium perchlorate was used as supporting electrolyte with a modified calomel electrode as reference. Glpc analyses were performed on an F & M Model 720 gas chromatograph with thermal conductivity detector using a 2-ft column packed with 5% silicone–gum rubber on Chromosorb-P. Thin layer chromatography was conducted on silica gel (Eastman Chromagram Type K 301R), with acetone for development and with location of spots by exposure to iodine vapor. Preparations were as described previously for 1,2-dithiolane, 1,2-dithiane, 1,2-dithiepane, their 1,1-dioxides, and their 1,1,2,2-tetroxides.[1c]

[21] N. Isenberg, Ph.D. Thesis, Rensselaer Polytechnic Institute, 1963.

more than 28%. From these percentages, the average value of 16% was estimated.

Reactions of Cyclic Disulfide Tetroxides with Thiophenol.—The procedure used with 1,2-dithiepane 1,1,2,2-tetroxide also illustrates that used with the tetroxide of 1,2-dithiolane and (as a complete solution in acetonitrile) with that of 1,2-dithiane.

Thiophenol (49 mg, 0.44 mmol) and 1,2-dithiepane tetroxide (31 mg, 0.16 mmol) were stirred in ethanol (10 ml) at *ca.* 25° for 20 hr. Solvent and excess thiophenol were evaporated, and residue was washed with ether to remove phenyl disulfide and with water to remove any sulfinic acid; 28 mg (90%) of 1,2-dithiepane tetroxide remained, mp 157–159° (lit.[1c] mp 159–160°).

Reaction of 1,2-Dithiane 1,1,2,2-Tetroxide with Sodium Thiophenoxide.—1,2-Dithiane tetroxide (0.500 g, 2.71 mmol) was suspended in 25 ml of 95% ethanol, and sodium thiophenoxide (5.44 mmol) in 25 ml of methanol was added dropwise (15 min) at 25°. The solution then was heated to reflux for 15 min and let cool (reaction occurs rapidly and completely without reflux however). After a stirring period of 5 hr more, the solvent was evaporated, and the remaining solid was washed with benzene to remove phenyl disulfide, leaving a white residue, yield 0.67 g (107%); in both melting point behavior (charring, 335°) and ir absorption (990 cm^{-1}, s), the properties were those of authentic disodium 1,4-butanedisulfinate.[15] The benzene wash contained 0.610 g (103%) of phenyl disulfide, identical in melting point, mixture melting point (56–58°), and ir absorption with an authentic sample.

Reactions of Nucleophiles with 1,2-Dithiane 1,1,2,2-Tetroxide.—1,2-Dithiane tetroxide (0.543 mmol) in 10 ml of 4:1 acetonitrile–water was stirred for 30 min at *ca.* 25° with 1.09 mmol of nucleophile (KCN, KOH, KI, KSCN, and KSC_6H_5); complete solution resulted in all instances, except with KOH. Removal of solvent, washing of the residue with 10 ml of water, and drying gave recovered tetroxide. After reaction of the tetroxide with potassium thiophenoxide, the residue first was extracted with ether to remove phenyl disulfide before being washed with water. The amount of tetroxide (identified by melting point and ir spectrum) isolated from each reaction, given in order of the nucleophiles above, was 0.014 g (14%), 0.004 g (4%), 0.090 g (90%), 0.086 g (86%), and 0.001 g (1%).

Stability of Cyclic Disulfide Tetroxides at pH 8.—Each cyclic disulfide tetroxide (0.036 mmol) was dissolved in 80% acetonitrile–water (7.2 ml). By means of a pH-Stat (Fisher automatic titrimeter),[22] 0.0876 *N* NaOH was added as rapidly as possible to a pH of 8 at *ca.* 25°. More alkali then was added automatically to maintain pH 8. The approximate milliequivalents of alkali required, before uptake ceased, with times for that point in parentheses follow: for 1,2-dithiane tetroxide, 0.057 mequiv (30 min, alkali uptake still not quite complete); for 1,2-dithiolane tetroxide, 0.066 mequiv (20 min); for 1,2-dithiepane tetroxide, 0.064 mequiv (15 min).[1b]

Pyrolysis of 1,2-Dithiane 1,1,2,2-Tetroxide.—1,2-Dithiane tetroxide (0.300 g, 1.63 mmol) was heated under N_2 at 275–280° for 10 min. Decomposition occurred immediately. Chloroform extracted from the resulting dark solids 0.045 g of yellow paste. Extraction of the paste with water left 0.014 g (5% recovery) of 1,2-dithiane tetroxide, mp 239° dec, and removed 0.028 g (14%) of crude tetrahydrothiophene 1,1-dioxide which contained all the ir bands of authentic material and had a melting point of *ca.* 5° (lit.[23] mp 20–21°). Glpc analysis[20] of this material indicated 85% purity for the main component (retention time, 50 sec), which was shown to be tetrahydrothiophene 1,1-dioxide by peak enhancement using authentic material; the yield therefore was 14%(0.85) = 12%.

A decomposition time of 40 min gave essentially the same result. Pyrolysis in nitrobenzene at 190–210° for varied lengths of time gave no improvement in yield, nor did addition of $AlCl_3$. Pyrolysis, with sand or chlorobenzene, also gave no improvement (132°, 7 days; 95% recovery of tetroxide, mp 238° dec); when 1,2-dithiane 1,1-dioxide was similarly heated in chlorobenzene, the recovery was 92%, mp 54–56°.

4-(2-Aminoethyldithio)butanesulfinic Acid (1).—2-Mercaptoethylamine (0.25 g, 3.3 mmol) and 1,2-dithiane 1,1-dioxide (0.50 g, 3.3 mmol) were stirred together for 30 min in 15 ml of 95% ethanol. The solid formed was removed to give 0.64 g (86%) of **1**, mp 157–159°. Recrystallizations from methanol–ether gave **1** with a constant melting point of 163–164° and ir absorption at 1640, 1560, and 1000 (s) cm^{-1}.

Anal. Calcd for $C_6H_{15}NO_2S_3$: C, 31.42; H, 6.59; N, 6.11; S, 41.93. Found: C, 31.13; H, 6.56; N, 5.99; S, 42.13.

Much the same result was obtained by use of 2-mercaptoethylamine as its hydrochloride (using also 1 equiv of an alkoxide), but with much greater difficulty in purification owing to presence of sodium chloride.

4-[2-(*n*-Octylamino)ethyldithio]butanesulfinic Acid (2).—A solution of 1,2-dithiane 1,1-dioxide (1.00 g, 6.60 mmol) and 2-(*n*-octylamino)ethanethiol[24] (1.24 g, 6.55 mmol) in 20 ml of absolute ethanol was stirred for 2 hr and then was cooled in an ice bath. The product, **2**, which precipitated then was removed by filtration and quickly returned to a flask before it could melt. Residual solvent then was removed, leaving 2.17 g (97%) of **2** having mp 25–30°. Several recrystallizations (absolute ethanol) gave rather waxy **2** with a melting point of 35–39°, which was not improved by further recrystallization, and ir absorption at 1620, 1000 (s), and 960 (s) cm^{-1}.

Anal. Calcd for $C_{14}H_{31}NO_2S_3$: C, 49.22; H, 9.15; N, 4.10; S, 28.16. Found: C, 49.10; H, 9.06; N, 4.25; S, 28.25.

After 2 weeks, the melting point had changed from *ca.* 35 to 73–78°, either because of decomposition (became pale green) or, more likely, because of progressive crystallization with some (but slight) decomposition, since the ir spectrum did not change significantly.

Sodium 4-(2-Acetamidoethyldithio)butanesulfinate (3).—Methanolic sodium methoxide (1.3 *N*) was added dropwise to 1,2-dithiane 1,1-dioxide (14.3 g, 93.9 mmol) and 2-acetamidoethanethiol (11.2 g, 93.9 mmol) in 250 ml of ethanol until the solution became basic to moist pH test paper (67.7 ml, 88 mmol). Removal of solvent left 26.0 g (94%) of crude **3**, mp 205–210°, with softening at 90°. For recrystallization, this **3** was dissolved in a minimum of methanol, to which acetone was added to precipitate 24.3 g (88%) of **3**, mp 210–220° dec, with softening at 110°. Purification for analysis was effected by adding ether to a solution of the latter in a minimum of methanol at 25° until a little precipitate formed; precipitate was removed, and more ether then was added to the filtrate to precipitate a second fraction. Continuation of this process produced several fractions, the largest of which was again fractionally precipitated similarly. Three such fractionations gave **3** with mp 215–220° dec and ir absorption at 3300 (s), 1650 (s), 1550, and 1000 (s) cm^{-1}.

Anal. Calcd for $C_8H_{16}NNaO_3S_3$: C, 32.75; H, 5.50; N, 4.77; S, 32.78. Found: C, 32.63; H, 5.53; N, 4.54; S, 32.57.

4-(2-Aminoethyldithio)butyl Benzyl Sulfone Hydrochloride (4).—A mixture of benzyl chloride (1.85 g, 14.6 mmol) and **1** (3.00 g, 13.1 mmol) in 350 ml of methanol was heated under reflux for 9 hr and let stand for 3 days. Solvent then was evaporated, and the residue was taken up in water (50 ml). Extraction with ether removed unreacted benzyl chloride. Benzene (50 ml) was added, followed by 3.5 ml of 3.4 *N* KOH solution (11.9 mmol).[25] The organic layer was separated, washed with water (10 ml), and filtered into 7 ml of ice-cold 3.4 *N* HCl, with which it was shaken well. A white precipitate formed immediately. The alkali layer was extracted twice more with benzene and the extract also was shaken with the acid. The total precipitate then amounted to 0.507 g (11%), mp 162–163°. Three recrystallizations from methanol–ether gave **4** having a melting point of 164–164.5° and ir absorption at 1600, 1500, 1310 (s), 1280 (s), and 1140 (s) cm^{-1}; tlc showed only one spot, R_f 0.36.

Anal. Calcd for $C_{13}H_{22}ClNO_2S_3$: C, 43.86; H, 6.23; N, 3.94; S, 27.02. Found: C, 44.11; H, 6.22; N, 4.04; S, 27.04.

4-(2-Acetamidoethyldithio)butyl Benzyl Sulfone (5).—A mixture of benzyl chloride (1.52 g, 12.0 mmol) and **3** (3.00 g, 10.2 mmol) in 225 ml of ethanol was heated under reflux for 3 hr and let stand at 25° for 24 hr. Sodium chloride then was removed by filtration, and the solvent was evaporated. The residue was rubbed well with hot acetone which, when evaporated, gave 2.10 g (57%) of **5**, mp 80–85°. Four recrystallizations from 95% ethanol gave **5** with mp 82–83° and ir absorption at 3300, 3220, 1640 (s), 1560, 1320, 1290 (s), 1130 (s), 780, and 700 cm^{-1}; tlc showed only one spot, R_f 0.76.

---

[22] We thank Dr. M. M. Jones of Vanderbilt University for use of this equipment.

[23] H. J. Backer and C. C. Bolt, *Rec. Trav. Chim. Pays-Bas,* **54,** 538 (1935).

[24] We thank the Walter Reed Army Institute of Research for supplying this material.

[25] It is probably quite important to minimize the time during which **4** is present as a free base, since our experience has been that free bases of aminoalkyl disulfides rapidly disproportionate to the symmetrical disulfides.

Anal. Calcd for $C_{15}H_{23}NO_3S_3$: C, 49.83; H, 6.41; N, 3.87; S, 26.60. Found: C, 50.34; H, 6.45; N, 3.88; S, 26.30.

4-(2-Acetamidoethyldithio)butyl 2,4-Dinitrophenyl Sulfone (6).—2,4-Dinitrochlorobenzene (3.46 g, 17.1 mmol) and **3** (5.0 g, 17.0 mmol) were heated under reflux for 30 min in 150 ml of ethanol. After 7 days, solvent was evaporated, and the residue was taken up into acetone and filtered to remove sodium chloride. Evaporation of acetone left 7.9 g (106%) of crude **6**, mp 60–67°. The **6** was a poorly crystalline material which proved very difficult to purify, even partly. Thus numerous recrystallizations using isopropyl alcohol, *t*-butyl alcohol, and chloroform–carbon tetrachloride failed to give **6** with a constant melting point, the last recrystallization yielding still somewhat impure yellow solid with mp 85–88° and strong ir bands at 3300, 3110, 1650, 1550, 1360, 1330, 1310, and 1160 cm^{-1}. Chromatography on both alumina and silica gel also did not improve the melting point.

Anal. Calcd for $C_{14}H_{19}N_3O_7S_3$: C, 38.43; H, 4.38; S, 21.99. Found: C, 39.42; H, 4.66; S, 22.36.

Sodium 4-(2-Acetamidoethyldithio)butanesulfonate (8). A. *Via* Hydrogen Peroxide.—Hydrogen peroxide (1.6 ml of "30%" solution, 8.8 *M*, 14.1 mmol) in 25 ml of water was added dropwise (30 min) to **3** (4.0 g, 13.6 mmol) in 100 ml of water at *ca.* 25°. After 2 days, the water was evaporated and the residue was dissolved in a minimum of methanol. Acetone then was added until a small amount of precipitate formed. This precipitate was removed and more acetone was added to precipitate a second fraction. Repetition of this process produced several fractions, which were combined into two different groups according to similar infrared spectra. The first fractions (2.0 g, mp 280° dec) showed only two strong bands at 1200 and 1060 cm^{-1}, but not at 3300, 1650, or 1550 cm^{-1} as is characteristic of **8**, indicating they were the disproportionation product, $[NaO_3S(CH_2)_4S]_2$; a negative test for nitrogen was obtained after sodium fusion. The last fractions were the desired product **8**, yield 1.1 g (26%). Purification of similarly prepared **8** was achieved by precipitating fractions from methanol at 25° using ether, as described for the sulfinate salt **3**. Four such fractionations gave **8**, which showed ir absorption at 3300, 1650 (s), 1550, 1200 (s), and 1060 (s) cm^{-1}; **8** had no well-defined decomposition point. Oxidation of $-SO_2Na$ to $-SO_3Na$, rather than of –SS– to –S(O)S–, is shown by disappearance of the 1000-cm^{-1} ir band of **3** and appearance of the 1060- and 1200-cm^{-1} bands in **8**.

Anal. Calcd for $C_8H_{16}NNaO_4S_3$: C, 31.06; H, 5.21; N, 4.53; S, 31.09. Found: C, 31.01; H, 5.22; N, 4.30; S, 30.82.

B. *Via* Sodium Metaperiodate.—Sodium metaperiodate (0.48 g, 2.2 mmol) in 10 ml of water was added to **3** (0.45 g, 1.5 mmol) in 10 ml of water at *ca.* 25°. After 15 hr, the water was evaporated, and the residue was extracted with 15 ml of hot methanol. Acetone was added to the methanolic solution until a small amount of precipitate formed. Since the ir spectrum of this first fraction was identical with that of **8** described in A, apparently little disproportionation had occurred. Addition of more acetone precipitated the remaining **8**, which was combined with the first fraction to give 0.27 g (57%) of disulfide sulfinate **8**.

Acetamidoethyl 4-(2-Acetamidoethylsulfenylsulfonyl)butyl Disulfide (10) Hemihydrate.—The disulfide sulfinate **3** (5.00 g, 17.0 mmol) was dissolved in 60 ml of ethanol and concentrated HCl (17.0 mmol) was added. A solution of 2-acetamidoethanethiol (2.03 g, 17.0 mmol) and *n*-butyl nitrite (2.65 g, 25.7 mmol) in 50 ml of anhydrous ethanol was heated at reflux for 15 min. An intense red color developed. *n*-Butyl nitrite (2.65 g, 25.7 mmol) then was added in one portion, followed by dropwise addition (30 min) of the above acidic solution. The reaction mixture was heated under reflux until the red color disappeared and evolution of nitric oxide ceased (1 hr). The mixture was cooled, the solvent was evaporated, and the residue was washed with 25 ml of ether, then with three 40-ml portions of water, and dried to yield crude hemihydrate **10**: 2.5 g (37%); mp 58–60°; ir absorption at 3300, 1650, 1550, 1320, 1300, and 1125 cm^{-1}. Equivalent weight determination[4] gave 383 (calcd 398 for **10** hemihydrate). Recrystallization from nitromethane and isopropyl alcohol gave **10** hemihydrate with mp 77–78°; tlc gave a single spot, R_f 0.40.

Anal. Calcd for $C_{12}H_{24}N_2O_4S_4 \cdot 0.5H_2O$: C, 36.25; H, 6.34; N, 7.04; S, 32.26. Found: C, 36.11; H, 6.26; N, 7.07; S, 33.02.

When this work was repeated on a larger scale and the crude product was recrystallized several times from isopropyl alcohol–water, a solid with mp 55–58° was obtained which contained two components by tlc indicating possible disproportionation. Recrystallization of the crude product several times from water alone gave solid with mp 90–91.5°, which probably was mainly disproportionation product, $[CH_3C(O)NH(CH_2)_2SSO_2(CH_2)_4S]_2$.

Disodium 1,4-Butanesulfinatesulfonate (11).—An aqueous solution of sodium hydroxide (21.5 ml, 0.97 *M*, 20.8 mmol) was added dropwise (15 min) to 1,2-dithiane 1,1,2,2-tetroxide (2.00 g, 10.9 mmol) suspended in 75 ml of 66% dioxane–water mixture, at which point pH test paper showed a pH of *ca.* 7 for at least 10 min. Evaporation of solvent left **11** as white solid, yield 2.83 g (106%). An approximate molecular weight of 280 (calcd 246) was obtained by titrating part of this **11** (after acidification with sulfuric acid) using sodium nitrite to a positive starch–iodide test.[26] Purification was effected by dissolving crude **11** in a minimum of methanol and adding acetone until *ca.* 70% of the solid had been precipitated. This solid was removed, redissolved in methanol, and reprecipitated with acetone. Several repetitions gave analytically pure **11**: strong absorption at 1195, 1050, and 980 cm^{-1}; no well-defined decomposition point.

Anal. Calcd for $C_4H_8Na_2O_5S_2$: C, 19.51; H, 3.27; S, 26.05. Found: C, 19.59; H, 3.30; S, 26.25.

Sodium 4-(2-Acetamidoethylsulfenylsulfonyl)butanesulfonate (12).—The disodium salt **11** (5.32 g, 21.6 mmol) was dissolved in 600 ml of methanol and 1.80 ml (21.6 mmol) of 12 *N* HCl was added. The mixture was stirred for 20 min. A solution of 2-acetamidoethanethiol (2.58 g, 21.6 mmol) and *n*-butyl nitrite (3.32 g, 32.2 mmol) in 50 ml of anhydrous ethanol was heated at reflux for 15 min, at which time the red color had become intense. *n*-Butyl nitrite (3.32 g, 32.2 mmol) then was added in one portion, followed by dropwise addition (40 min) of the above sulfinic acid solution. The color disappeared after the addition of the sulfinic acid, but reflux was continued for 30 min, after which the mixture was allowed to cool and stand overnight. Evaporation of solvent then gave white solid, which was washed with two 50-ml portions of acetone and extracted with 400 ml of hot ethanol. Removal of the ethanol gave 5.83 g (79%) of **12**, mp 126–128°. Recrystallization from ethanol of similarly prepared material gave **12** with a constant melting point of 128–129.5° and ir absorption at 3310, 1650 (s), 1550, 1325, 1210, 1165 (s), 1125, and 1070 cm^{-1}.

Anal. Calcd for $C_8H_{16}NNaO_6S_3$: C, 28.14; H, 4.72; N, 4.10; S, 28.17. Found: C, 27.94; H, 5.14; N, 3.97; S, 27.93.

Disproportionation of 4-(2-Aminoethyldithio)butyl Benzyl Sulfone Hydrochloride (4).—The benzyl sulfone **4** (0.287 g, 0.806 mmol) in 10 ml of water was heated at 100° for 72 hr in a sealed vial wrapped with aluminum foil, and the vial then was chilled in ice. Insoluble material was extracted into benzene, which was dried and evaporated to give 0.156 g (79%) of 1,14-diphenyl-2,7,8,13-tetrathiatetradecane-2,2,13,13-tetroxide (**9**), mp 118–118.5°. The water containing the remaining products was evaporated leaving 0.129 g of residue, mp 194–199°, the ir spectrum of which was consistent with that expected of a mixture of **4** (mp 164–164.5°) and cystamine dihydrochloride, mp 218–220°. The disulfide **9** gave a single tlc spot (R_f 0.68); elemental analysis showed absence of nitrogen but presence of sulfur. The water-soluble material on tlc produced two spots, one consistent with cystamine dihydrochloride (R_f 0.00) and the other with **4** (R_f 0.36) (authentic samples were done simultaneously).

The per cent of disproportionation (79%) was calculated as (100 × 2 × moles of disulfide **9** isolated)/(0.806 mmol of disulfide **4**).

Similarly prepared disulfide **9**, recrystallized from methanol, gave **9** with a constant melting point of 120–120.5°.

Anal. Calcd for $C_{22}H_{30}O_4S_4$: C, 54.29; H, 6.21. Found: C, 54.41; H, 5.98.

Registry No.—**1,** 19293-54-0; **2,** 19293-55-1; **3,** 19293-56-2; **4,** 19293-57-3; **5,** 19293-91-5; **6,** 19293-92-6; **8,** 19293-93-7; **9,** 19293-94-8; **10,** 19293-95-9; **11,** 19293-96-0; **12,** 19293-97-1; 1,2-dithiane 1,1,2,2-tetroxide, 18321-18-1; thiophenol, 108-98-5; sodium thiophenoxide, 930-69-8.

---

(26) J. L. Kice and K. W. Bowers, *J. Amer. Chem. Soc.*, **84**, 605 (1962).