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Authors

Barltrop, J.A
Hayes, P.M.
Calvin, M.

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AS A MODEL FOR THE PRIMARY QUANTUM CONVERSION ACT IN PHOTOSYNTHESIS

J. A. Barltrop, P. M. Hayes and M. Calvin

March, 1954

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ABSTRACT

Some chemical and photochemical observations of 1,2-dithiolane and its

derivatives are reported with particular reference to the possible mode of

function of the naturally occurring system, 6-thioctic acid. Experimental

evidence is presented to demonstrate that the strain energy in this 5-membered

ring is not less than 6.5 Kcals. and probably larger. Reagents which both

oxidize and reduce this ring are described together with the conditions required

for its reformation from the corresponding dithiol. Evidence is adduced to in-

dicate that the primary product of photolysis of this ring in acetic media is

very likely a thiol and sulfenic acid or derivative thereof.

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Although plants which are allowed to photosynthesize in $C^{14}O_2$ rapidly

assimilate labelled carbon into a series of compounds, very little of the C^{14}

finds its way into the intermediates of the Krebs tricarboxylic acid cycle

during illumination. If, after a period of photosynthesis, the light is turned

off, the compounds of the Krebs cycle rapidly become labelled.[1] Thus there is

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(1)   A. A. Benson and M. Calvin, J. Exptl. Bot., 1, 63 (1950).

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a reaction path linking the photosynthesis and Krebs cycles which becomes

blocked during illumination. With the discovery[2,3,4,5,6] that 6-thioctic acid

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(2)   L. J. Reed, I. C. Gunsalus, et al., J. Am. Chem. Soc., 73, 5920 (1951).

(3)   E. L. Patterson, et al., J. Am. Chem. Soc., 73, 5919 (1951).

(4)   I. C. Gunsalus, L. Struglia and D. I. O'Kane, J. Biol. Chem., 194, 859
   (1952).

(5)   L. J. Reed, B. G. DeBusk, J. Am. Chem. Soc., 74, 3457 (1952).

(6)   M. W. Bullock, et al., J. Am. Chem. Soc., 74, 3455 (1952).

is a coenzyme for the oxidative decarboxylation of pyruvate to active acetyl

groups, which through CoA feed carbon into the Krebs cycle,[7,8] Calvin and Massini[9]

(7)   S. Ochoa, J. R. Stern, M. C. Schneider, J. Biol. Chem., 193, 691 (1951).

(8)   S. Korkes, A. DelCamilo, I. C. Gunsalus and S. Ochoa, J. Biol. Chem.,
       193, 721 (1951).

(9)   M. Calvin and P. Massini, Exper., 8, 445 (1952).

suggested that the process could be formulated thus:
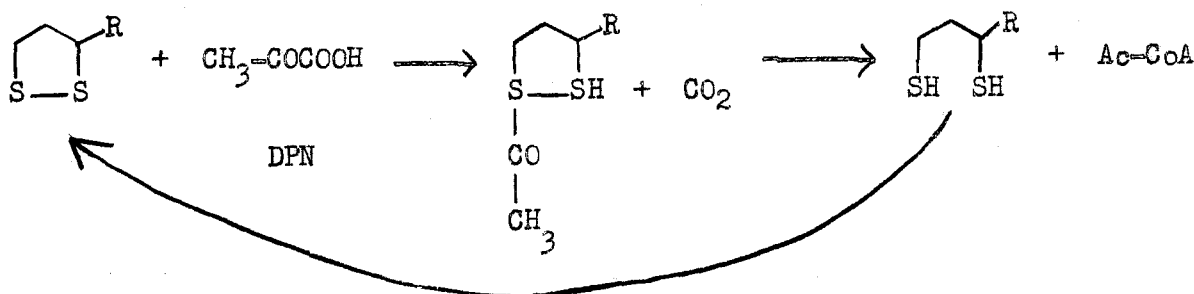
Moreover, since the coenzyme must be present in its oxidized form in

order that the oxidation of pyruvic acid may proceed, these authors suggested

that the reducing power formed in the presence of light shifted the steady-state

conditions of the coenzyme toward its reduced (dithiol) form and thus reduced

the rate at which intermediates of the photosynthetic cycle entered the TCA

system. At this point, specimens of the isomeric 4-, 5-, and 6-thioctic acids

became available to us through the courtesy of Dr. T. H. Jukes of Lederle

Laboratories. The u.v. absorption spectra of these compounds (Figure 1) showed

a displacement of the absorption peak to progressively longer wave lengths as

the size of the disulfide ring diminished a phenomenon which might well be due

to ring-strain. If one supposes that the absorption band is due to a transition:

$$\binom{(CH_2)_n}{S - S} \xrightarrow{h\nu} \binom{(CH_2)_n}{S \quad S}$$
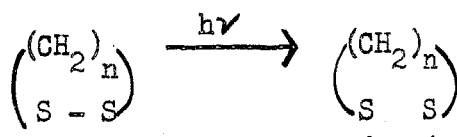
then assuming to a first approximation that the excited states of all the

disulfides have the same energy curves (Figure 2) and that the transitions

all reach the same point on the upper curve, the differences in λ_{max} would

correspond to differences in the energies of the ground states - i.e., to ring

strain. On this basis the shift from $\lambda_{max} = 2500$ Å for straight chain ali-

phatic disulfides to $\lambda_{max} = \underline{ca}.$ 3300 Å for 6-thioctic acid implies that the

latter experiences a strain of $\underline{ca}.$ 25-30 Kcals. ($\Delta = \Delta_6 + \Delta_U$ in Figure 2),

and that hence the dissociation energy of the -S-S-bond (E_D ub Figure 2) will be

reduced by a like amount. Any difference (Δ_U in Figure 2) in the energy of the

excited state reached in the ring compound from that reached in an open chain

compound will have to be subtracted from this value. The published values

for D (R-S-S-R) vary from 50^{10} - 70^{11} Kcals; thus the ring strain might reduce

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(10)   A. H. Sehon, J. Am. Chem. Soc., 74, 4723 (1952); M. L. Huggins, J. Am
        Chem. Soc., 75, 4125 (1953).

(11)   J. L. Franklin and H. E. Lumpkin, J. Am. Chem. Soc., 74, 1024 (1952).

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the dissociation energy of the ring S-S- bond to $\underline{ca}.$ 30-40 Kcals. There is

another quite independent method of estimating the strain in this ring based

upon purely structural considerations, together with some measured single bond

rotational barriers. Of prime consideration is the apparent[12,13] requirement

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(12)   D. W. Scott, H. L. Finke, M. E. Gross, G. B. Guthrie and H. M. Huffman,
        J. Am. Chem. Soc., 72, 2424 (1950).

(13)   D. W. Scott, H. L. Finke, J. P. McCullough, M. E. Gross, R. E. Pennington
        and Guy Waddington, J. Am. Chem. Soc., 74, 2478 (1952).

that the stable configuration about the S-S bond be the one having the dihedral

angle equal to $90°$ and a rotational barrier of at least 10 Kcal. This barrier

is presumably in the $180°$ or trans position. The $0°$ or cis position which would

be the one demanded in the trimethylene disulfide ring might be expected to have

a somewhat higher barrier. A similar qualitative suggestion was made by Affleck

and Dougherty[14] based upon observations of the relative stability of a series

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(14)   J. G. Affleck and G. Dougherty, J. Org. Chem., 15, 865 (1950).

---

of polymethylene disulfides. We do not propose here to elaborate further on

the origin of this barrier but simply to suggest that it is at least partly due

to the repulsion between the two non-bonding pairs of primarily p electrons on

the sulfur atoms.[15,16,17] Presumably it is the changing hybridization of these

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(15)   L. Pauling, Proc. Nat. Acad. Sci., 35, 495 (1949).

(16)   E. N. Lassettre and L. B. Dean, J. Chem. Phys., 17, 317 (1949).

(17)   K. S. Pitzer, Farad. Soc. Disc., 1951, "Hydrocarbons."

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and the non-bonding s electrons as well, which is induced by the geometrical

requirements of the ring which is responsible for the shift in ultraviolet ab-

sorption. An examination of the constitution of this ring shows that the

largest dihedral angle that it is possible to achieve without major changes in

the C-C bond is approximately $45°$. Under this circumstance, the three pairs

of methylene hydrogens lie opposed to each other in a position of maximum

potential energy as judged from considerations of the rotational barriers in

ethane and propane.[17] Pitzer has given the potential barrier for the rotation

of a methyl group in propane as 3.4 Kcals. If we take a minimum of 6 Kcals. for

the CH (six) repulsions and use a cosine law with a maximum of 10 Kcals. to es-

timate the potential energy remaining in the S-S bond at a dihedral angle which

approaches 45° (5 Kcals.) we arrive at a minimum strain energy of approximately

11 Kcals. In view of the great uncertainty in the normal S-S bond energy it does

not seem fruitful to pursue these estimates any further.

The evidence is that light absorbed by the chlorophyll and associated

pigments in plants is passed on as quanta of less than 40 Kcals. of energy, and

this led directly to the notion that 6-thioctic acid was itself the substance

involved in the conversion of the electromagnetic energy of the quantum into

chemical bond energy and that the process of quantum conversion in photosynthe-

sis was the fission of the disulfide bond of 6-thioctic acid to give the di-

thiyl radical, i.e.

$$Chl^* + \underset{S-S}{\overset{R}{\bigwedge}} \longrightarrow Chl_{(ground)} + \underset{S\quad S}{\overset{R}{\bigwedge}}$$
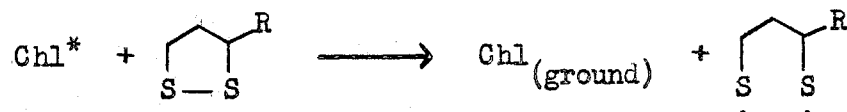
Suggestions as to the subsequent transformations of the diradical have

already been outlined in a preliminary communication.[18] The purpose of the

(18)   M. Calvin and J. A. Barltrop, J. Am. Chem. Soc., 74, 6153 (1952).

present paper is to describe in detail some of the experiments mentioned there

as well as new ones, and to amend and amplify the suggestions.

In order to assess the probability of the preceding ideas, it became

necessary to learn something of the fundamental properties of dithiolane ring

systems. Since 6-thioctic acid is difficult to prepare, trimethylene disulfide

(dithiolane) (I, R = H) was selected as a model substance. It had been

prepared by Affleck and Dougherty[14] from the bis-Bunte salt derived from

trimethylene dibromide, but we were not aware of this at the time. The only
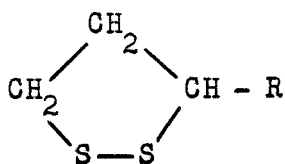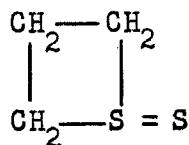
previous work on this substance was recorded by Hagelberg[19] and Autenrieth

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(19)   L. Hagelberg, Ber., 23, 1086 (1890).

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and Wolff[20] who describe it as a colorless crystalline substance m.p. ca.

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(20)   W. Autenrieth and K. Wolff, Ber., 32, 1368 (1899).

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71-75°, and which the latter authors believed to be a dimer. When trimethylene

dibromide was condensed with sodium disulfide in aqueous ethanol and the ethanol

evaporated, the distillate had a bright yellow color like 6-thioctic and a

u.v. absorption spectrum which showed a peak at λ_{max} 3300 Å and had a general

form almost identical with 6-thioctic acid. This observation led to the method

of isolation and purification described in the experimental section. A sub-

stance, almost certainly thiacyclobutane was formed as a by-product. The tri-

methylene disulfide was obtained in only 9% yield; the main reaction product

appears to be the "dimer" of Autenrieth and Wolff[20] since it can be crystallized

to a colorless solid m.p. ca. 75°. It can be slowly depolymerized to tri-
methylene disulfide by steam distilling with Na_2S_2 and NaOH.[21] This, rather

(21)   F. O. Davis and E. M. Fetters, J. Am. Chem. Soc., 70, 2611 (1948).

than monomeric trimethylene disulfide, appears to have been the product ob-
tained by Yur'ev and Levi[22] by pyrolysis of thiacyclobutane.

(22)   Y. K. Yur'ev and I. S. Levi, Dokl. Akad. Nauk, SSSR, 73, 953 (1950).

Tetramethylene disulfide, similarly prepared and described by Affleck
and Dougherty[14] as an oil, was found to be a crystalline solid m.p. 32-33° C.
Trimethylene disulfide could only be obtained in solution. All attempts to
isolate it in a pure state were terminated by polymerization. It was recog-
nized by the close similarity between its u.v. spectrum and that of 6-
thioctic acid, the constitution of which has been established, and also by
an independent molecular weight estimation. The fact that trimethylene di-
sulfide is yellow seems to have escaped previous notice, although this seems
to be a characteristic property of the dithiolane system since 6-thioctic acid,
4:4-dimethyldithiolane[23] and 4:4-spirocyclohexyldithiolane[24] are also yellow.

(23)   H. J. Backer and N. Evenhuis, Rec. trav. chim., 56, 129 (1937).

(24)   H. J. Backer and A. F. Tamsma, Rec. trav. chim., 57, 1183 (1938).

Trimethlyene disulfide is stable to reduction by sodium borohydride
but is very rapidly reduced by zinc and dilute acid to trimethylene dithiol,
from which trimethylene disulfide can be quantitatively recovered by oxidation

with iodine. This observation also provides proof of structure since it excludes the alternative formulation II (cf. Backer and Evenhuis[23]). Trimethylene disulfide was also obtained in excellent yield from trimethylene dithiol by oxidizing with air in the presence of alkali or of hydroidic acid, the latter reaction presumably depending on the formation of iodine as an intermediate.

The suggestion as formulated earlier as to the possible function of 6-thioctic acid in photosynthesis included the photochemical fission of the -S-S- bond into thiyl radicals. The polymerization of trimethylene disulfide in light is consistent with the suggestion -- but in order to obtain more definitive evidence, trimethylene disulfide was photolyzed in an ether-pentane-alcohol glass[25] at -196° C. Under these conditions, one could expect that

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(25)   G. N. Lewis, D. Lipkin and T. T. Magel, J. Am. Chem. Soc., 63, 3005 (1941).

---

the thiyl radicals produced by photolysis would be preserved from further reaction (a) because the rigid solvent would prevent their approaching each other and thus recombining and (b) because at the low temperature employed the insufficiency of thermal activation energy would prevent their attacking the solvent. Hence it should be possible to build up a considerable concentration of thiyl radicals. When trimethylene disulfide was photolyzed under these conditions, a clear, pale salmon-colored glass resulted, which appeared to be stable in liquid N_2, but which on warming a few degrees to the point at which the glass lost its rigidity, rapidly became turbid with polymer. The clarified solution now contained more than 50% of the original trimethylene disulfide (Figure 6). These results are consistent with the interpretation:
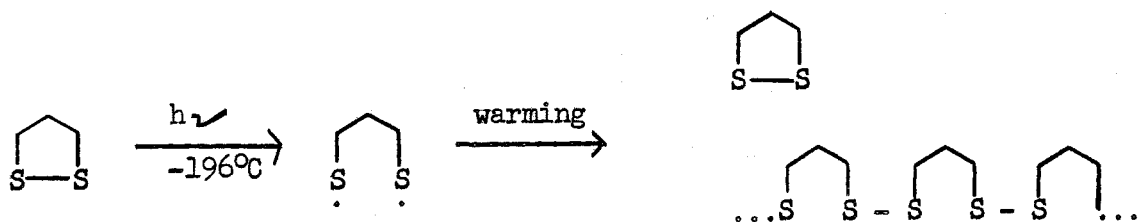
Confirmation of the production of free radicals[26] by photolyzing trimethylene

(26)   Since the initial preparation of this manuscript there appeared some
        confirmatory evidence for the photo dissociation of aromatic disulfides
        into free radicals. M. S. Kharasch, W. Nudenberg and T. Meltzer.,
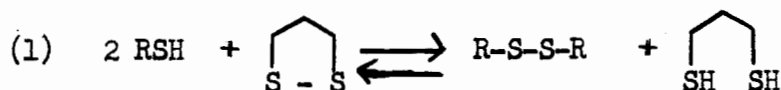   J. Org. Chem., 18, 1233 (1953).

disulfide was sought by using diphenylpicrylhydrazyl which is itself a free

radical, stable to air, and which has been used before in the detection of

radicals.[27] Diphenylpicrylhydrazyl, with and without trimethylene disulfide

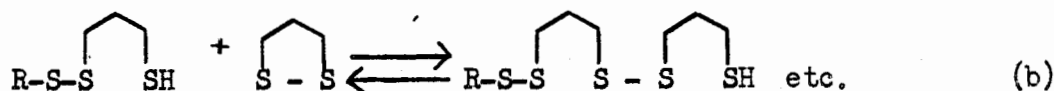(27)   C. E. H. Bawn and S. F. Mellish, Trans. Farad. Soc., 47, 1217 (1951).

solution, was both illuminated and kept in the dark. The color of the hydrazyl

faded only when mixed with trimethylene disulfide and exposed to light (see Figure

4). This result was also consistent with the notion that trimethylene disulfide

on photolysis gave radicals, these then reacting with the hydrazyl. However,

we realize now that this experiment no longer affords proof of the formation of

radicals for subsequent work has shown that the picrylhydrazyl color fades on

adding thiols and as is shown below, thiols are generated when the disulfide is

illuminated in ethanol solution.

The theory of quantum conversion as it was outlined previously also

required that the -S-S-bond be in a condition to be ruptured by a quantum of

less than 40 Kcals. The estimate of ring strain given earlier derived from

spectral data is based on several assumptions of uncertain validity. That based o

on purely structural considerations is somewhat better. In order to arrive at

a more accurate balue for the difference in the -S-S- dissociation energy be-

tween an aliphatic disulfide and trimethylene disulfide (i.e., ring strain) we

exploited the observation that aliphatic thiols reacted in a reversible manner

with trimethylene disulfide. After adding a thiol, the trimethylene disulfide

was partly destroyed, the amount of destruction being correlated with the amount

of added thiol, and dilution of the solution causing a regeneration of tri-

methylene disulfide. The equilibrium being observed could be either:

(1)   2 RSH + ⌒ ⇌ R-S-S-R + ⌒
      S - S   SH SH

or

(2)   RSH + ⌒ ⇌ R-S-S ⌒ SH   (a)
      S - S

      R-S-S ⌒ SH + ⌒ ⇌ R-S-S ⌒ S - S ⌒ SH etc.   (b)
              S - S
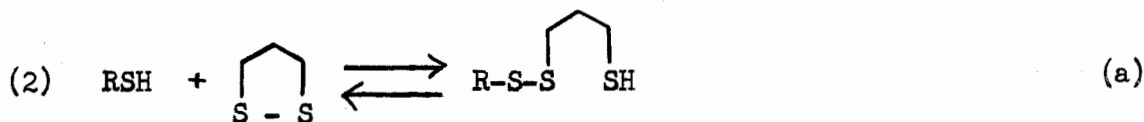
The equilibrium constant calculated for reaction (1) does not fit the

data obtained. The complex of reactions (2) can be given a simple mathematical

formulation by noting that for each added thiol group reacting with trimethylene

disulfide a new thiol group is generated. If it is assumed that all the thiols

in reaction (a) have the same equilibrium constant for the reaction with tri-

methylene disulfide, then since in the experiments herein reported there is

always a large excess ($<$10-fold) of added thiol, a good approximation to the

overall equilibrium constant can be represented by

$$K = \frac{[\text{aliphatic disulfide}]}{[\text{cyclic disulfide}] \times [\text{thiol}]}$$

Values of K calculated from the data by this equation are reasonably constant for the reactions of trimethylene disulfide with both benzyl mercaptan and β-mercaptoethanol.

For the reaction with benzyl mercaptan conducted at 24.1 and 35.85° C, the free energies of reaction are $\Delta F_{24.1} = -1.67$ Kcals. and $\Delta F_{35.1} = -1.74$ Kcals. respectively. Hence $\Delta H = -6.3$ Kcals. ΔH, which is a measure of the differences in bond energy between the cyclic disulfide and an aliphatic disulfide is thus a measure also of the ring strain in trimethylene disulfide. It should be noted that although there is ca. 6 Kcals of heat energy available to drive the reaction to the right, this is opposed by an entropy term $T\Delta S$ of ca. 4.5-5 Kcals. so that only ca. 1.5-2 Kcals. of free energy is left.

As an alternative method of deriving information concerning the stability of the S-S bond in trimethylene disulfide, the kinetics of the polymerization of the substance in n-octane were investigated. The activation energy of such a reaction would be the same as the dissociation energy of the S-S- bond, if the reaction were unimolecular in the disulfide and if the rate-determining step required the homolytic fission of the bond. However, it was found that the reaction was abnormal in that the activation energy was extremely low (ca. 8.7 Kcals.) and the apparent PZ factor fantastically small. Whatever may be the interpretation of these results it is clear that the reaction is not a simple one, and the values obtained have little relevance to the problem under investigation.

As an approach to an inquiry into the intermediate oxidation levels of
the disulfide the oxidation of trimethylene disulfide by air, by hydrogen
peroxide and by ammonium persulfate was investigated. Preliminary tests showed
that trimethylene disulfide was not readily oxidized by air, but was rapidly
photooxidized in the presence of a suitable photosensitizer. Figure 7 shows
that while trimethylene disulfide and zinc tetraphenylporphin individually
were almost stable to oxygen both in visible light and in the dark, a mixture
of the two rapidly absorbed oxygen upon illumination. When the reaction was
followed quantitatively (Figure 8) it emerged that the oxygen was absorbed
in the ratio of one atom of oxygen per mole of trimethylene disulfide suggest-
ing that the overall reaction was:
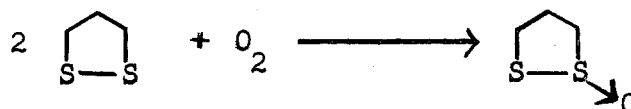
This formulation of the oxidation product as a disulfide monoxide was supported
by the close similarity between its u.v. absorption spectrum and the spectrum
of the monoxide of 6-thioctic[28] acid (Figure 10). At this point we should,

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28   We are indebted to Dr. T. H. Jukes of Lederle Laboratories for a
      specimen of this substance.

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perhaps, mention that further experiments have shown that photochemical polymer-
ization of the disulfide sensitized by zinc tetraphenylporphin which was men-
tioned in the earlier communication[18] was probably caused by traces of oxygen.
The ammonium persulfate oxidation of trimethylene disulfide gave comparable

results. Figure 11 shows the rapid course of the reaction and the isobestic

point indicates that the oxidation followed one course only. The diluted

reaction mixture had a u.v. spectrum virtually identical (allowing for the

end absorption of the ammonium bisulfate formed) with those of 6-thioctic acid

monoxide and the photooxidation product of trimethylene disulfide described

previously. Furthermore, only one mole of ammonium persulfate can probably

have been used per mole of trimethylene disulfide since the same product was

obtained even with an excess of trimethylene disulfide. Also 6-thioctic acid,

oxidized with one mole of ammonium persulfate gave a product the spectrum of

which resembled closely the spectrum of authentic 6-thioctic acid monoxide,

and the extinction coefficient of the 2440 Å peak of the oxidation product

assuming 100% yield (which is justified by the isobestic point) is within the

limits of experimental error identical with that of 6-thioctic monoxide.

This accumulation of data proves that the course of the reaction is

$$\text{(disulfide-R)} + (NH_4)_2S_2O_8 + H_2O \longrightarrow \text{(disulfide monoxide)} + 2\ NH_4HSO_4$$

In this context it is interesting to note that di-n-propyl disulfide is stable

to oxidation under these conditions. Trimethylene disulfide monoxide was

found to be stable to acid in the cold but it was instantaneously destroyed

by alkali. Disulfide monoxides are a little-investigated group of compounds

of which the best known are allicin[29] and 6-thioctic acid monoxide.[30] They

(29)   C. J. Cavallito, J. S. Buck and C. M. Suter, J. Am. Chem. Soc., 66, 1950 (1944).

(30)   M. W. Bullock, S. A. Brockman, E. L. Patterson, J. V. Pierce and E. L. R. Stokstad, J. Am. Chem. Soc., 74, 3455 (1952).

have been previously prepared by oxidizing disulfides with _t_-butyl hydro-peroxide[30] and with per-acids.[31] The course of the reaction between tri-

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(31)   C. J. Cavallito, et al.,   J. Am. Chem. Soc., 69, 1710 (1947).

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fide and hydrogen peroxide was too complicated to be elucidated spectroscopically, but little, of any monoxide appeared to be formed directly.

The fact that thioctic acid seems to be the coenzyme directly responsible for the oxidative decarboxylation of α-keto acids led us to investigate the non-enzymic reactions between α-keto acids and trimethylene disulfide. When pyruvic acid or α-ketoglutaric acid was heated in alcoholic solution with trimethylene disulfide very little change occurred. Next, α-ketoglutaric acid and trimethylene disulfide were exposed to u.v. irradiation in a nitrogen atmosphere. It was found that whereas control tubes containing trimethylene disulfide only, rapidly became turbid due to polymerization, the tubes containing trimethylene disulfide and α-ketoglutaric acid remained clear although spectroscopic examination showed the complete destruction of the trimethylene disulfide. A series of similar experiments were performed with pyruvic, oxalic and acetic acids. The following facts emerged. (1) With pyruvic and oxalic, as with α-ketoglutaric acid, the solution remained clear upon photolysis but with acetic acid polymerization of the trimethylene disulfide was only partly inhibited. (2) No detectable quantity of CO_2 was formed during the pyruvic and oxalic acid experiments. (3) The H^+ concentration was unchanged by photolysis. (4) Amperometric titration for thiols with Ag^+ showed that rather less than one thiol group was formed per molecule of trimethylene disulfide. (5) Iodine titration revealed the presence of reducing substances in the solution in addition to the thiols.

Since the added acids were apparently unchanged by the photolysis and since

acetic acid which is a weaker acid than pyruvic and oxalic acids was less

effective as a polymerization inhibitor than the others, it appeared likely

that these substances were inhibiting polymerization merely by virtue of their

acidity and that mineral acids would exercise a similar effect. This was

found to be the case. Photolysis of trimethylene disulfide in the presence

of hydrochloric acid, again caused no change in the acid titre. Also ca.

1 thiol group was liberated per molecule of trimethylene disulfide and excess

reducing power was revealed by iodine oxidation as with the organic acids.

A small quantity of hydrogen sulfide was detected by a lead acetate paper.

The difference between the iodine and silver titrations posed the

question as to which should be accepted as the index of thiol concentration.

It was found in model titrations with other thiols that with freshly prepared

solutions the I_2 and Ag^+ titres agreed, but that with aged solutions, partially

oxidized by air, the iodine titre was much higher than the silver titre. This

suggested that while both silver and iodine were equally effective for es-

timating pure thiols, and therefore the silver was probably the more specific

reagent. We attempted to confirm this value (from Ag^+ titration) for the

-SH concentration by developing a new analytical method. It is known that

both alcohols and thiols react with carbon disulfide under alkaline conditions

to form xanthates (III, X = O) and trithiocarbonates (III, X = S) respectively:

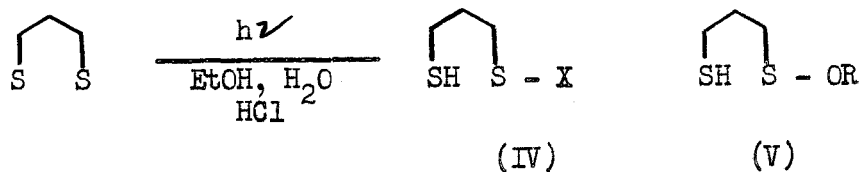$$R - S - C \underset{SNa}{\overset{S}{\lessgtr}}$$
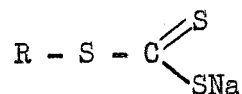
III

Model experiments revealed that thiols react instantaneously while alcohols form xanthates only slowly. Furthermore, the spectra of trithiocarbonates are readily distinguishable from those of xanthates (see Figure 13). By adding approximately the calculated amount of carbon disulfide and sodium hydroxide to the solution, and observing the instantaneous formation of the characteristic peaks at ca. 3000 and 3350 Å the presence of microscopic amounts of thiols in alcoholic solution may readily be detected. Unfortunately, the formation of trithiocarbonate appears to be reversible and the uncertainty as to the extent of reaction means that the method, while useful for qualitative detection, is of little value for quantitative estimation. Nevertheless, comparison between the heights of the 3000 Å and 3350 Å peaks obtained from the reaction between the photolysis solution and carbon disulfide on the one hand and those of the trithiocarbonates derived from benzyl mercaptan and β-mercaptoethanol on the other indicated the formation of about one and less than two thiol groups from one molecule of trimethylene disulfide.

The finding that approximately one thiol group was formed from each molecule of trimethylene disulfide on photolysis raised the question of whether one group was attached to each 3-carbon fragment derived from the trimethylene disulfide or whether the photolysis produced a mixture containing some trimethylene dithiol and some 3-carbon units without thiol groups, the average being one thiol for each C_3 unit.
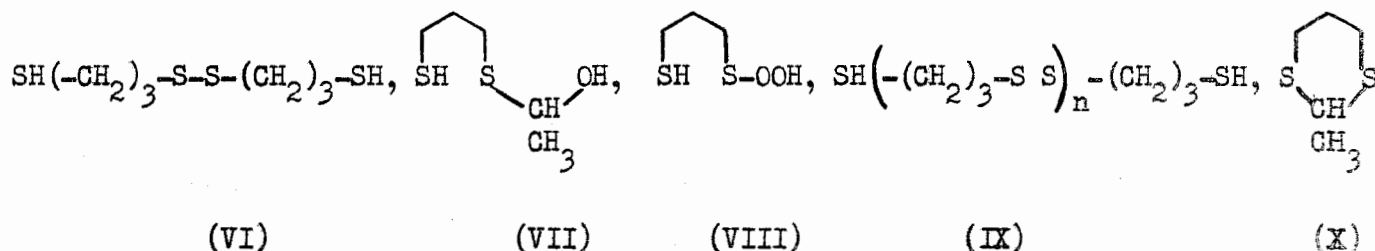
Separate experiments had shown that trimethylene dithiol could readily be oxidized and in excellent yield to trimethylene disulfide by iodine under acid conditions or by hydroxylamine and alkali. In spite of repeated experiments no trimethylene disulfide could be found in any of the photolysis

solutions after these treatments, thus demonstrating that the course of the photolysis could be represented schematically as:

$$\underset{S \qquad S}{\bigcap} \quad \xrightarrow[\substack{\text{EtOH, H}_2\text{O} \\ \text{HCl}}]{h\nu} \quad \underset{\substack{SH \quad S - X \\ (IV)}}{\bigcap} \quad \underset{\substack{SH \quad S - OR \\ (V)}}{\bigcap}$$
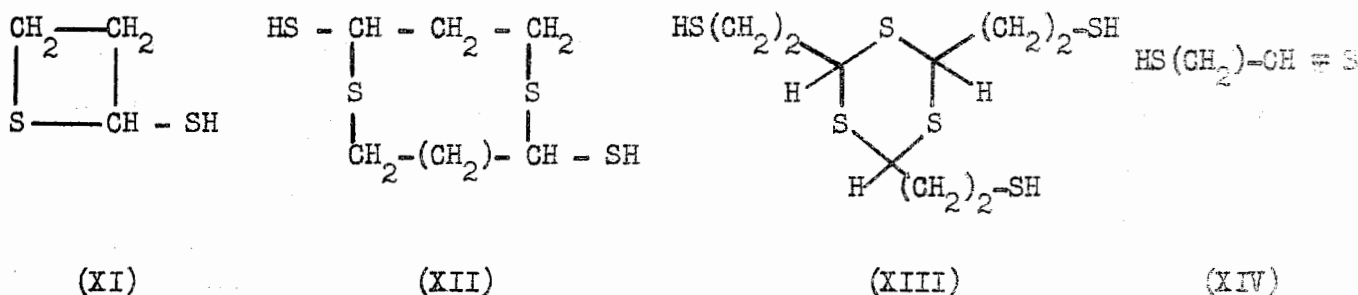
Having now established that one of the sulfur atoms of trimethylene disulfide was transformed on photolysis into a thiol, it became pertinent to inquire into the nature of the other, i.e., what is "-S-X" in IV?   Since one of the original sulfur atoms has been reduced, it follows that one or more of the components of the initial solution must have been oxidized. This could be accomplished in two ways. The components could either be directly dehydrogenated to some oxidation product which then became attached to the other sulfur atom or direct addition of one of the components to the S-S bond might take place.

The only substances present in the solution before photolysis are trimethylene disulfide, hydrochloric acid, water and ethanol. The simple oxidation of any of these to e.g. chlorine, acetaldehyde, acetic acid, hydrogen peroxide, etc. is excluded since for reasons of stoichiometry, trimethylene dithiol would also have to be formed. Their formation can only be reconciled with the absence of trimethylene dithiol if they subsequently react with one of the thiol groups of the dithiol giving (VI), (VII), (VIII).

$SH(-CH_2)_3-S-S-(CH_2)_3-SH$, SH S $\overset{\displaystyle CH}{\underset{CH_3}{|}}$ OH, SH $S-OOH$, $SH\left(-(CH_2)_3-S\ S\right)_n-(CH_2)_3-SH$, $S\overset{\displaystyle}{\underset{CH_3}{\underset{|}{CH}}}S$

     (VI)   (VII)   (VIII)   (IX)   (X)

All these are improbable. It is difficult to see why interaction between chlorine (or hydrogen peroxide) and trimethylene dithiol should stop at (VI) -- one would expect polymers of the type (IX) to be simultaneously formed. Also aliphatic disulfides have a peak $\lambda_{max} \cong 2500$, $\mathcal{E}_{max} \cong 500$ (see Figure 1) no evidence for which can be found in the u.v. spectrum of the photolysis solution (Figure 12). The structure (VII), a thio semi-acetal, would certainly dehydrate into the cyclic thio acetal (X) and thus eliminate all thiol groups. Having considered direct oxidation of one of the components, the only other possibility for the primary photochemical act is addition of one of the solution components to the S-S bond of trimethylene disulfide. This leads to structures (V, R = H), (V, R = C_2H_5) (IV, X = Cl) and to (XI)

$\underset{\displaystyle \begin{array}{c} | \\ S \end{array}}{\overset{CH_2---CH_2}{\underset{S---CH - SH}{}}}$ $HS - \underset{\displaystyle CH_2-(CH_2)-}{\overset{\displaystyle |}{\underset{|}{\overset{S}{|}}}}CH - CH_2 - \underset{\displaystyle CH - SH}{\overset{\displaystyle CH_2}{\underset{|}{\overset{S}{|}}}}$ $HS(CH_2)_2 \cdots S \cdots (CH_2)_2-SH$ structure $HS(CH_2)-CH = S$

     (XI)   (XII)   (XIII)   (XIV)

The structure (V, R = H) would be a sulfenic acid. With the exception of a sulfenic acid derived from anthroquinone[32] none are known to exist. All at-

(32)   K. Fries, Ber., 45, 2965 (1912).

tempts to prepare them have resulted in mixtures of dismutation products such

as sulfinic acids, disulfides, thiolsulfonic esters and thio ketones.[33]

(33)   N. Kharasch, S. J. Potempa and H. L. Wehrmeister, Chem. Rev., 39, 263
        (1946).

The substance (V, R = C_2H_5 is a sulfenic ester. Little is known of

these substances.[33] The sulfenyl chloride (IV, R = Cl) would certainly not

preserve its identity in the photolysis solution but might be expected to yield

mixtures of substances such as sulfenic anhydride[34] and the usual dismutation
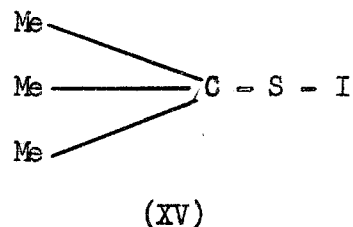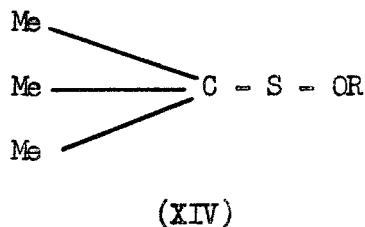
(34)   Th. Zincke, Ann., 406, 106 (1914) and Ann., 391, 57 (1912).

products of sulfenic acids. So far as we have been able to ascertain, no semi-

thioacetals are known. This does not exclude the possible formation of (XI)

as the primary photochemical product but on structural grounds we would expect

it to be transformed into (XII) or more likely (XIII) through equilibration

with the isomeric thiol thioaldehyde (XIV). Thus general considerations sug-

gest the conclusion that the primary photochemical transformation occurring

when trimethylene disulfide is photolyzed should lead to (XI), or to (IV,

X = OH, OEt, or Cl) and that the actual solution obtained at the end of pho-

tolysis should consist either of one or more of these entities or of some

transformation product or products derived therefrom.

At this point it became necessary to investigate the photolysis solution

chemically in order to attempt to discriminate among these possibilities, but

first in order to orientate the investigations we decided to experiment with

model sulfenic esters. The only aliphatic sulfenic esters described in the

literature are ethyl and methyl t-butyl sulfenate (XIV) prepared by Rheinholdt and Motzkus.[35]

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(35)   H. Rheinholdt and E. Motzkus, Ber., 74, 657 (1939) and Chem. Abs., 32, 484 (1938).

$$\begin{array}{ccc}
\text{Me} & & \\
\text{Me} & \!\!\!\!\!\!\!\!>\!\! C - S - OR & \\
\text{Me} & & \\
\end{array}
\qquad
\begin{array}{ccc}
\text{Me} & & \\
\text{Me} & \!\!\!\!\!\!\!\!>\!\! C - S - I & \\
\text{Me} & & \\
\end{array}$$

(XIV)   (XV)

The u.v. absorption spectrum of the ethyl ester was found to have a characteristic peak at 2655 Å $\mathcal{E}_{max}$ = 70.5 (see Figure 16). Spectroscopic investigation then showed that the substance was stable to ca. M/40 sodium hydroxide in aqueous alcohol for many hours. Hydroxylamine also had no detectable effect. Dilute acid on the other hand rapidly destroyed the ester in the cold to give a solution with a u.v. peak at ca. 2880 Å. Thus the sulfenic ester seems chemically to be more closely related to ethers than to esters. Because of the possible simultaneous occurrence in the photolysis solution of thiols and sulfenic esters, the reaction between (XIV, R = C_2H_5) and thiols was investigated. Under neutral conditions, thiols were without effect upon t-butyl sulfenic ester in the cold. On making alkaline, however, a rapid reaction occurred. (Figure 17.) With trimethylene dithiol, ethyl t-butyl sulfenate rapidly gave a moderate yield of trimethylene disulfide identified by its characteristic u.v. absorption peak at 3300 Å. Thus the ester oxidizes thiols to disulfides under alkaline conditions. It can also function as an oxidizing agent under acid conditions as shown by the rapid liberation of iodine from

dilute hydriodic acid solutions.

The known tendency of sulfenic derivatives to dismute[36] into disulfides

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(36)   There appears to be no verb in the English language dictionary corres-
         ponding to the noun dismutation. In this paper, we propose to rectify
         this regrettable omission by introducing a new verb "to dismute."

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and other products suggested that the 2880 Å peak mentioned earlier which

develops on treating ethyl t-butyl sulfenate with acids might be due to

di-t-butyl disulfide. In order to elucidate this point some of the disulfide

was synthesized by oxidizing t-butyl mercaptan with iodine in aqueous alcohol.

It was observed during the preparation that although the iodine was rapidly

destroyed, no definite end point could be discerned; the solution became

gradually yellow and then brown and the unforgettable and unmistakable smell

of sulfenyl iodides developed suggesting that the disulfide was fissioned by

the slight excess of iodine into t-butyl sulfenyl iodide (XV). The formation

of sulfenyl iodides from aliphatic disulfides under these conditions is unique

in our experience although Fuson, et al.[37] have shown that it is possible to

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(37)   R. C. Fuson, C. C. Price, R. A. Bauman, O. H. Bullitt, Jr., W. R.
         Hatchard and E. W. Maynerd, J. Org. Chem., 11, 469 (1946).

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fission di-β-chloroethyl disulfide with chlorine in CCl_4 solution into the

sulfenyl chloride. This anomaly associated with the t-butyl disulfide was

reflected also in its u.v. spectrum (Figure 16). The substance is yellow, in

contrast with all other known aliphatic disulfides, and instead of showing the

characteristic 2500 Å peak of aliphatic disulfides, only end absorption was

manifested. Thus the 2880 Å peak mentioned above cannot be due to di-t-butyl

disulfide. At this point it may be pertinent to point out that the spectrum of

sodium t-butyl trithiocarbonate (Figure 12) is also markedly different from

those of the other trithiocarbonates.

These interesting observations on t-butyl disulfide suggested a more

extended investigation in which it was discovered that the substance was an

oxidizing agent capable of both liberating iodine from dilute hydriodic acid

and oxidizing trimethylene dithiol to trimethylene disulfide though much less

effectively than ethyl t-butylsulfenate.

In view of the spectral anomolies associated with t-butyl compounds,

we decided to make another sulfenic ester for comparison purposes. Trichloro-

methyl sulfenyl chloride was commercially available. It was reacted with

sodium ethoxide under conditions similar to those described for the synthesis

of the t-butyl analogue by Rheinholdt and Motzkus[35] to give ethyl trichloro-

methylsulfenate (XVI). We experienced difficulty in obtaining good

$$Cl_3C - S - OEt$$
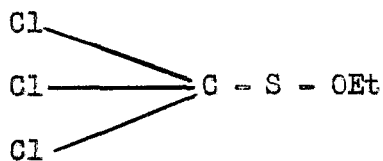
(XVI)

analytical data for this substance — a difficulty mentioned by Rheinholdt and

Motzkus[35] in connection with the t-butyl analogue. The substance had an ab-

sorption maximum in the same region as the t-butyl sulfenate (Figure 19) but

inverse chemical properties. Thus in contrast to ethyl t-butylsulfenate, the

trichloromethyl sulfenic ester was stable to acid and sensitive to alkali. We

are thus placed in the embarrassing position of having to decide which, if

either, of these two model sulfenic esters is in fact a model for the supposed

substance (V, R = C_2H_5). In view of the profound change in chemical properties

caused by introducing a trichloromethyl group in a compound (cf. chloral and acetaldehyde), we inclined to the belief that the t-butylsulfenic ester was the better chemical analogue although it might not be good as a spectroscopic model.

A large number of experiments were performed on the solution obtained by photolyzing trimethylene disulfide in the presence of acid in the attempt to determine its precise constitution. The facts which emerged are these: On photolysis the 3300 $\overset{o}{A}$ band of trimethylene disulfide gradually disappeared and a more intense absorption built up at wavelength $\langle 2500$ $\overset{o}{A}$, but even when all the trimethylene disulfide had been destroyed, there was a residual absorption in the region just short of 3300 $\overset{o}{A}$ (Figure 12). No absorption maximum could be detected in the spectrum of the solution photolyzed at room temperature. However, thiols which have already been shown to be present in the solution show a similar absorption (see Figure 12b) and it seemed possible that the thiol absorption was masking the bands of the groups present in the photolyzed solution. In order to neutralize as far as possible this masking thiol absorption, a solution of β-mercaptoethanol of concentration equal to the concentration of the initial trimethylene disulfide (silver titrations described earlier demonstrated the formation of ca. one thiol group per initial trimethylene disulfide molecule present at the commencement of photolysis) was introduced into the compensating cell of the spectrophotometer. In this way, the instrument will automatically subtract the absorption due to the thiol and the resultant spectrum should then represent the absorption of other groups present in the photolyzed solution. When this was done (Figure 12b) a maximum appeared at 2320 $\overset{o}{A}$. That this band is probably not an artefact is indicated by the fact that it remains in essentially the same position

$(\lambda_{max}$ = 2290 $\overset{o}{A}$) when only 1/2 of the previous concentration of thiol is intro-duced into the compensating cell. That the photolysis solution contains entities sensitive to acid is demonstrated by the fact that while no change occurs in the spectrum if the solution is kept at -20° for 24 hours, keeping it at room temperature causes the 2320 $\overset{o}{A}$ peak to move to 2420 $\overset{o}{A}$ (Figure 12b).

The photolysis solution is very sensitive to alkali which causes instantaneous elimination of the 2320 $\overset{o}{A}$ peak and the development of maxima at 2629 $\overset{o}{A}$ and more slowly at 3340 $\overset{o}{A}$ and 4100 $\overset{o}{A}$ and still more slowly at 2310 $\overset{o}{A}$. On reacidifying, the bands at 3340 and 4100 $\overset{o}{A}$ are immediately destroyed and replaced by bands at 2820 $\overset{o}{A}$ and 2350 $\overset{o}{A}$. Making alkaline again causes reappearance of the 3340 and 4100 $\overset{o}{A}$ maxima. With variations in degree, reference to Figures 14 and 21 shows that similar phenomena were encountered both with solutions of trimethylene disulfide photolyzed at low temperatures and in solutions photolyzed in the presence of sodium hydroxide. Expecially notable is what appears to be a definite maximum of low intensity at -3100 $\overset{o}{A}$ in the low temperature photolysate.

In the time at our disposal, it has proved impossible to unravel completely the complexities in these transformations occurring in photolyzed solutions of trimethylene disulfide. But certain conclusions are inescapable.

Firstly, the photolyzed solution is markedly unstable, particularly to alkali. Secondly, the solution possessed oxidizing power. This is demonstrated by the oxidation of trimethylene dithiol to trimethylene disulfide. Also the bands which appear at 3340 $\overset{o}{A}$ and -4200 $\overset{o}{A}$ when the solution is made alkaline and their disappearance when it is reacidified are strongly reminiscent of dithiocarboxylic acids. Hantsch and Bucerius[38] give for dithioacetic acid, λ_{max} =

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(38)   A. Hantsch and W. Bucerius, Ber.. 59, 793 (1926).

2950 Å and for the ammonium salt, λ_{max} = 3330 Å. The formation of dithio acids from a substance which before photolysis was at the oxidation level of a disulfide must involve oxidation. Whether the tentative assignment of those bands to dithio acids is correct or not, the transformation of compounds of the photolysis solution into substances absorbing at such long wavelengths can only mean the development of multiple carbon or sulfur bonds (i.e., oxidation) for these bonds can be developed in the absence of halogens. Furthermore, the formation of unequivocally acid substances such as those giving rise to these bands requires the presence of oxidizing agents in the original photolysis solution. Thirdly, the solution possesses reducing power in addition to that furnished by the liberated thiol groups, as shown by the value for the iodine titre.

The arguments advanced earlier in this paper limited the structure of the primary photolytic product to (V, R = OH, OEt, Cl) or to (XI). The fact that similar products appear to be formed when trimethylene disulfide is photolyzed under both alkaline and acid conditions as well as with a variety of acid anions eliminates structure (V, R = Cl). We can find no rationale by which to explain the diversity of products derivable from the photolyzed solution in terms of the structure (XI) and we are left only with (V, R = OH or OEt).

We can thus postulate with reasonable confidence the idea that when trimethylene disulfide is photolyzed in ethanol solution the sulfenic ester (V, R = OEt) is the primary chemical product
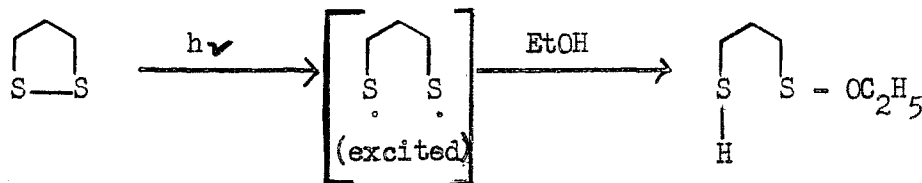
and that the solution at the end of the photolysis contains its transformation

products. The well-established instability of sulfenic acid derivatives, their

extreme tendency to dismute (shown also by sulfur monoxide, the inorganic ana-

logue), and their oxidizing power are completely in accord with the data obtained

from photolyzed solutions of trimethylene disulfide. Furthermore, hydrogen

sulfide which was detected in the photolyzed solutions is an expected de-

composition product of certain sulfenic acids (cf. Schoberl[39]). The various

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(39)   S. Schoberl, Ber., 70B, 1186 (1937).

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decomposition products could be derived by such reactions as:
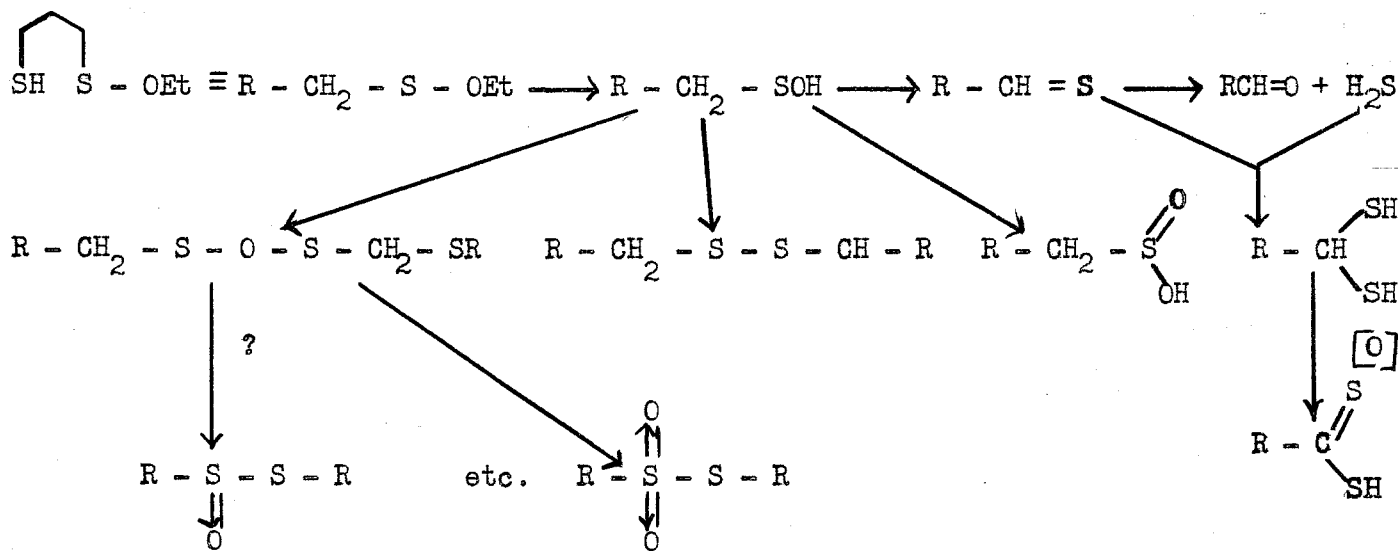
and what has been discovered of the properties of the photolyzed solution is

entirely consistent with this formulation.

If the suggestion that photolysis of trimethylene disulfide in ethanol

solution leads to a sulfenic ester be correct, the question immediately arises,

as to what is its relevance to the process of photosynthesis. In an earlier

communication, Calvin and Barltrop[18] advanced arguments in support of the idea
that 6-thioctic acid was directly involved in the quantum conversion step of
photosynthesis but the steps by which water was oxidized to molecular oxygen
were not specified. The above work with trimethylene disulfide leads us
directly to the further suggestion that water itself or some very closely
related hydroxylic substance is the primary hydrogen donor, i.e., that the pri-
mary photochemical act of photosynthesis involves the splitting of the O-H bond
thus:

(XVII)
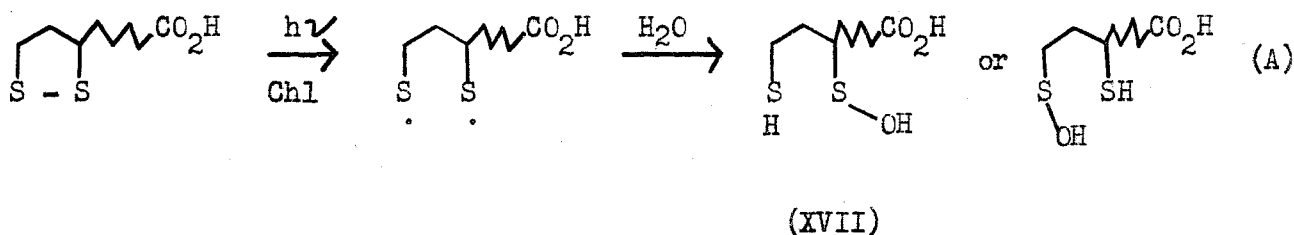
It is difficult to assess the feasibility of this process from consider-
ations of the energy changes involved since most of the critical data is lacking.
However, assigning values of 120 and 85 Kcals. respectively to the bond energies
of the O-H (E_{O-H}) and S-H (E_{S-H}) bonds, the change in energy of the system
would be

$$\Delta E_{(Kcals)} = 35 + (E_{S-S} - E_{S-O})$$

No thermodynamic measurements involving the S-O link have been reported
in the literature and there seems to be a considerable diversity of opinion on
the S-S bond. However, we may estimate the difference $(E_{S-S} - E_{S-O})$ by analogy
with F_2 (35), Cl_2 (57), and FCl (60)[40] to be zero or somewhat negative. Since

(40)   K. S. Pitzer, "Quantum Chemistry," Prentice Hall, New York, 1953.

E_{S-S} for the dithiolane ring seems to be at least about 6 Kcals. less than that

for an open chain disulfide the net energy requirement for the reaction would be

about 30 Kcals. If the size of the quantum passed on by the chlorophyll is

something less than 40 Kcals. then this process would not only be possible but

would represent a relatively efficient utilization of the quantum.

It should be noted that sulfenic acids are the sulfur analogues of alkyl

hydroperoxides. Thus in one step, the oxygen of water (or its hydroxylic

relative) has been brought to the oxidation level of peroxide while one of the

hydrogen atoms has been raised to a reduction potential approaching that of

molecular hydrogen.[41] In order to liberate oxygen from water the energy of

(41)   M. Calvin, "Mercaptans and Disulfides," presented at the Conference on
        Glutathione, Ridgefield, Conn., Nov. 20, 21, 22, 1953. Academic Press,
        1954.

several quanta is required, and how these may be accumulated has been the subject

of prolonged discussion.[42] Assuming only a single type of quantum conversion

(42)   E. I. Rabinowitch, "Photosynthesis I", 1951; "Photosynthesis II, pt. 1,"
        1945, Interscience Publishers, Inc., New York.

process always producing the same chemical entity, we must have a dismutation

between two such systems whereby one drops to a low energy state and the other

is raised to a higher energy level; repetition of these dismutations, if necessary,

would lead ultimately to a system with sufficient energy to effect the fission of

water. Now, what emerges more clearly than anything else from a survey of the

literature on sulfenic acids is their tendency to dismute.[34] The only known

sulfenic acids are those of Fries[32] — all other attempts to prepare them have

resulted merely in the isolation of their dismutation products. Similarly, all

reactions in which sulfenic acids have been postulated as intermediates lead to
the decomposition of dismutation products of sulfenic acids. We thus feel
justified in suggesting that the thioctic-sulfenic acid (XVII), once formed, would
dismute. In view of our limited understanding of the chemistry of sulfenic acids
the nature of this dismutation is difficult to predict, but in order to crystal-
lize the situation, we tentatively propose a dismutation into 6:8-dithioloctanoic
acid (XVIII) and a molecule at the energy level of hydrogen peroxide (possible
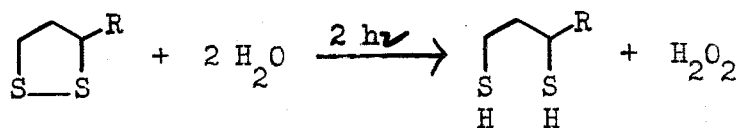(XIX), (XX), or (XXI) which then decomposes into hydrogen peroxide and 6-
thioctic acid:
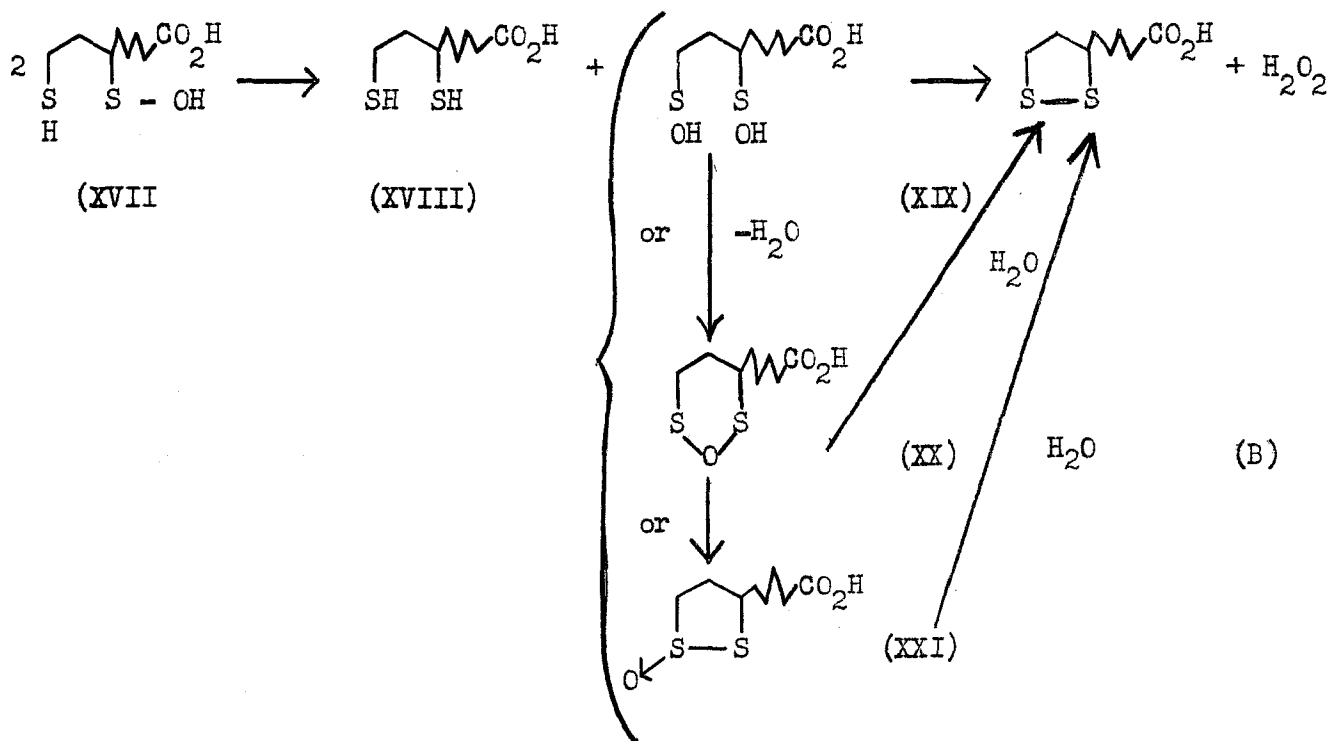
The dithiol (XVII) can then provide the reducing power required to con-
vert carbon dioxide into sugars — possibly through the reduction of TPN or DPN.[43]
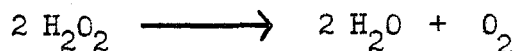
(43)   I. C. Gunsalus, McCollum-Pratt Inst. Symp. on "Mechanics of Enzyme Action,"
        1953. Johns Hopkins Univ. Press., 1954.

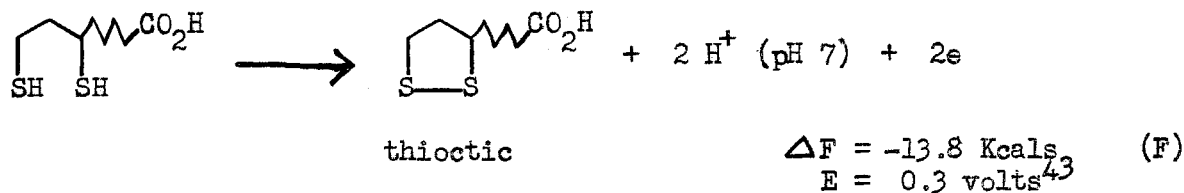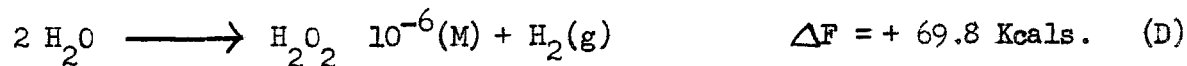The hydrogen peroxide may lead through catalase to molecular oxygen.

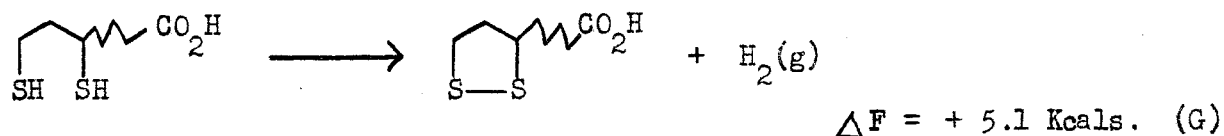Adding equations (A) and (B) gives the overall reaction:

$$\text{(structure)} + 2\ H_2O \xrightarrow{2\ h\nu} \text{(structure)} + H_2O_2 \qquad \text{(C)}$$

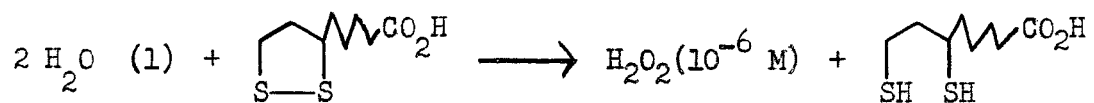$$2\ H_2O_2 \longrightarrow 2\ H_2O + O_2$$

and this leads, on purely stoichiometric grounds, to the requirement that four quanta be utilized for each molecule of oxygen evolved, not considering losses or any other concommitent energy requirements that might exist. The question still remains to be answered -- is quantum conversion by the mechanism of equation (C) thermodynamically possible? The literature contains the following data:

$$2\ H_2O \longrightarrow H_2O_2\ 10^{-6}(M) + H_2(g) \qquad \Delta F = +\ 69.8\ \text{Kcals.} \quad \text{(D)}$$

$$H_2(g) \longrightarrow 2\ H^+(pH\ 7) + 2e^- \qquad \Delta F = -18.9\ \text{Kcals.} \quad \text{(E)}$$

$$\text{(structure)} \longrightarrow \text{(structure thioctic)} + 2\ H^+\ (pH\ 7) + 2e$$

$$\Delta F = -13.8\ \text{Kcals.} \quad \text{(F)}$$
$$E = 0.3\ \text{volts}^{43}$$

Subtracting (E) from (F)

$$\text{(structure)} \longrightarrow \text{(structure)} + H_2(g)$$

$$\Delta F = +\ 5.1\ \text{Kcals.} \quad \text{(G)}$$

From (D) and (G) by subtraction we get

$$2\ H_2O\ (1)\ +\ \text{[structure: S—S ring with }CO_2H\text{]}\ \longrightarrow\ H_2O_2(10^{-6}\ M)\ +\ \text{[structure: SH SH with }CO_2H\text{]}$$
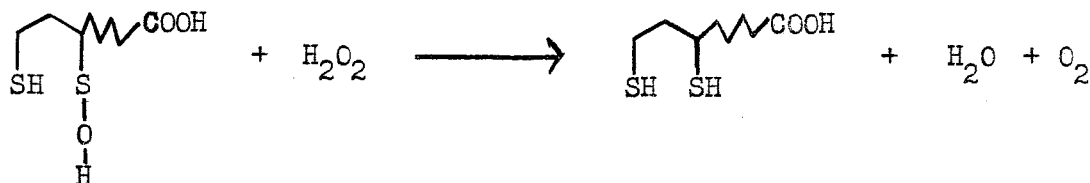
$$\Delta F = +\ 64.7\ \text{Kcals} \quad (C)$$

Hence the increase in free energy involved in the proposed photosynthetic equation (C) is $\cong$ 65 Kcals. Since this has to be provided by two quanta, the energy of the excitation passed on by chlorophyll should be not less than $\doteq$ 32 Kcals., a figure which is well within the bounds of possibility. The energy required for a photosynthesis process following the above mechanism might well be less than ca. the 65 Kcals per atom of oxygen liberated calculated above, for two reasons. Firstly, it is entirely possible that the plant might be able to evolve oxygen directly from the dismutation product (B) without necessarily using the energy-rich hydrogen peroxide as an intermediate.[44] Secondly, should

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(44)   A. H. Mehler, Arch. Biochem. and Biophys., 33, 65 (1951); ibid. 34, 339 (1951).

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hydrogen peroxide by an intermediate, then it is, in principle, possible usefully to recover the energy stored in the molecule by using it as a reducing agent, e.g.

$$\text{[structure: SH S—O—H with }COOH\text{]}\ +\ H_2O_2\ \longrightarrow\ \text{[structure: SH SH with }COOH\text{]}\ +\ H_2O\ +\ O_2$$

thus providing a three quantum route to one molecule of oxygen.

These considerations of the energy relationships lends support to the proposed

scheme for the quantum conversion process in photosynthesis.

EXPERIMENTAL PART

1. <u>Synthesis of Trimethylene Disulfide:</u> - Pulverized sodium sulfide nonahydrate
(120 g.) was dissolved in boiling 95% ethanol (500 cc.) and heated under reflux with
sulfur (17 g.). When the sulfur had dissolved, the brown two-phase solution was
dispersed by shaking and added during 20 minutes to a boiling solution of tri-
methylene dibromide (49 cc., 0.48 mols.) in 95% ethanol (250 cc.). The mixture
was then distilled under reduced pressure: 750 cc. of a bright yellow distillate
(A) was collected. More 95% ethanol was added to the residue which was again
distilled under reduced pressure to give 600 cc. of a pale yellow distillate (B).
The optical densities (at 3300 $\overset{o}{A}$) of solutions A and B were 6.7 and 1.98, respec-
tively. Since the molar extinction coefficient of the 3300 $\overset{o}{A}$ band of trimethylene
disulfide is 147 (see below and Figure 1), the yield of product is approximately
9%. The solution of trimethylene disulfide prepared this way sometimes contains
a significant amount of thiacyclobutane. This may be removed by fractionating
the solution through a 15-plate column until no more distillate with a u.v. peak
λ_{max} = 2720 $\overset{o}{A}$ distills.[45] The bulk of the reaction product is a solid which may

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(45)   Dr. G. Waddington, (priv. comm.) gives λ_{max} (in iso octane) <u>ca.</u> 2760 $\overset{o}{A}$)

---

be obtained from benzene-chloroform as colorless crystals m.p. <u>ca.</u> 73° C. This is
presumably the "dimer" of Antenreith and Wolff.[20]   It may be slowly depolymerized
to trimethylene disulfide by steam distilling with strong sodium hydroxide and sodium
disulfide.

    (a)   <u>Molecular Weight of Trimethylene Disulfide.</u>[46] - Ten cc. of crude

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(46)   We are indebted to Dr. D. F. Bradley, of our laboratory, for the freezing
        point determinations.

---

trimethylene disulfide in petroleum ether (ca. 1 M) was evaporated in a vacuum desiccator leaving 1.36 g. of crude trimethylene disulfide, a few mg. of polymer, and possibly remaining traces of petroleum ether. This was taken up in 10 ml. benzene. The freezing point of 43.75 g. benzene to which 1 ml. aliquots of this solution were added successively was determined with a Beckman thermometer. The maximum depression (ΔT_f) was 0.703° C. (for 47.57 g. of benzene and 0.684 g. of trimethylene disulfide) with a maximum deviation from linearity (of the intermediate points) in the weight of trimethylene disulfide/weight of benzene vs. ΔT_f plot of $< 0.005^\circ$ C. Using 5.49° C. as the freezing point and 30.4 cal./g. as the heat of fusion of benzene, the found molecular weight is 104 for trimethylene disulfide (calc. molecular weight 106).

2. **Synthesis of Tetramethylene Disulfide.** - This preparation is similar to that of trimethylene disulfide except that tetramethylene dibromide is used. The reaction mixture was extracted with light petroleum instead of being distilled. The extract was concentrated and distilled under reduced pressure. Tetramethylene disulfide was collected at 82° C., 150 mm. pressure and was purified by sublimation. It formed colorless needles (25% yield), melting point $32-33^\circ$.

Anal: Calcd. for $C_4H_8S_2$: S, 53.34; Found: S, 53.30%.

3. **Properties of Trimethylene Disulfide.** - The substance can be obtained in the form of an unstable, unpleasant smelling oil by cautiously evaporating a pentane solution under strongly reduced pressure in a stream of nitrogen. Attempts to distil it even under reduced pressure caused it to be transformed into a colorless rubbery polymer on the walls of the condenser. In solution it is destroyed by oxygen and is polymerized by light. Even in the dark in an atmosphere of nitrogen at sub-zero temperatures, solutions will sometimes be-

come turbid if kept in certain flasks. However, pure solutions of the disulfide
can readily be obtained from the partially polymerized solutions by distilling
under reduced pressure. Iodoacetamide does not react with it in the cold and
it is stable to reduction by sodium borohydride, as is the related 6,8-thioctic
acid. It is, however, much more rapidly reduced by zinc dust in the presence
of dilute hydrochloric acid than are other aliphatic disulfides. Also, in
contrast to other aliphatic disulfides, trimethylene disulfide appears to be
stable to the action of potassium cyanide solution.

4. <u>Extinction Coefficient of Trimethylene Disulfide.</u> - A freshly distilled
aqueous ethanol solution of trimethylene disulfide was found spectroscopically
to be free from dimer and polymers (optical density of 3300 $\overset{o}{A}$ band 0.801). One
hundred cc. of this solution was reduced in a stoppered measuring cylinder by
shaking with zinc dust (200 mg.) and concentrated hydrochloric acid (1.5 cc.).
Spectroscopic examination showed reduction to be almost complete in 5 minutes
and quantitative in 10 minutes. The small amount of unreacted zinc was spun
down and the amount of trimethylene dithiol present in the solution estimated
by iodine titration which had already been shown to be quantitative by model
titrations on β-mercaptoethanol. Five cc. of dithiol solution was treated
under nitrogen with 10 cc. of 1.022×10^{-2} $\underline{N}$ iodine (I_2) and immediately back
titrated under nitrogen with 1.00×10^{-2} $\underline{N}$ sodium thiosulfate solution;
4.83 cc. of thiosulfate was required as the average of several titrations.
Hence, normality of trimethylene dithiol $= \dfrac{10.22 - 4.83}{5} \times 10^{-2}$, i.e.,
molarity $= 1/2 \times$ normality $= 0.539 \times 10^{-2}$ $\underline{M}$. Therefore, molar extinction
coefficient of trimethylene disulfide $= \dfrac{0.801}{0.539 \times 1.015} \times 10^{2}$ (where the factor
1.015 corrects for dilution of disulfide by added hydrochloric acid) $= \underline{147}$.

5. Quantitative Re-oxidation of Trimethylene Dithiol to Trimethylene Disulfide by Iodine. - Three cc. of the solution of trimethylene dithiol obtained in the preceding experiment was treated with 2.8 cc. of 1.022 x 10^{-2} $\underline{N}$ aqueous iodine solution. (The data of the preceding experiment show that this corresponds to 88.5% oxidation of the dithiol.) The partially oxidized solution examined spectroscopically showed trimethylene disulfide (optical density of the 3300 Å band is 0.37). The solution was quite clear, indicating absence of polymers. The original disulfide solution of the preceding experiment (optical density 0.801) was diluted during reduction by a factor of 1.015 and during the iodine reoxidation diluted further by a factor of $\frac{5.8}{3}$. Thus, if the original disulfide had been quantitatively reduced and reoxidized without dilution, the calculated optical density of the 3300 Å band would be:

$$0.37 \; x \; \frac{5.8}{3} \; x \; \frac{100}{88.4} \; x \; 1.015 \; = \; 0.82.$$

Since the actual optical density of the initial disulfide is 0.801 the yield of trimethylene disulfide obtained by oxidizing trimethylene dithiol with iodine equals 102%.

6. Catalyzed Aerial Oxidation of Trimethylene Dithiol. -

    (a) Twenty cc. of a freshly distilled aqueous ethanol solution of trimethylene disulfide (3300 Å band optical density 0.89) was reduced with zinc dust (50 mg.) and concentrated hydrochloric acid (0.2 cc.). To 2.5 cc. of this solution was added 0.5 cc. of an approximately 10^{-2} $\underline{N}$ iodine in aqueous potassium iodide. Some trimethylene disulfide was formed (optical density of 3300 Å band 0.23). After various intervals of time the spectrophotometer cell containing the reaction mixture was opened to the atmosphere, closed, shaken

vigorously to aerate the mixture and then re-examined spectroscopically. By
repetition of these operations the optical density of the 3300 $\overset{o}{A}$ band increased
gradually to a maximum value of 0.74. Had no dilution taken place during the
addition of the iodine solution, the height of the band would have been
0.74 x $\frac{3}{2.5}$ = 0.89. Since the optical density of the initial trimethylene di-
sulfide is 0.89, this corresponds to a 100% yield in the reduction and aerial
reoxidation of the disulfide.

(b)   A solution of trimethylene dithiol in 95% ethanol (approximately
10^{-2} M) was shaken with air. No trimethylene disulfide was found spectro-
scopically. Ten cc. of the solution was agitated for 10 minutes with zinc
dust (25 mg.) and concentrated hydrochloride (0.1 cc.) and centrifuged from
unreacted zinc dust. The supernatant solution was shaken with air and examined
spectroscopically. No detectable amount of trimethylene disulfide was formed.
Potassium iodide (23 mg.) dissolved in water (0.2 cc.) was added to 3 cc. of
the supernatant solution and the mixture then shaken with air. Trimethylene
disulfide was formed rapidly. Trimethylene dithiol was also rapidly oxidized
to trimethylene disulfide by shaking with air in the presence of sodium
hydroxide.

7.   <u>Photolysis of Diphenylpicralhydrazyl and Trimethylene Disulfide.</u> - The
following solutions were prepared: diphenylpicrylhydrazyl (8.3 mg.) in ethanol
(100 cc.) and freshly distilled trimethylene disulfide in aqueous ethanol
(1.19 x 10^{-2} M). The apparatus shown in Figure 3 was charged with the hydrazyl
solution (10 cc.) and the disulfide solution (1 cc.). After deoxygenating the
mixture in a stream of pure nitrogen saturated with ethanol for half an hour,
the fading of the diphenylpicrylhydrazyl, both in the light and in the dark,
was followed by density measurements in a Klett colorimeter using a red filter.

Illumination was effected between opposed photospot lamps behind infra-red absorbing filters, the apparatus being kept cool in an air blast. A control experiment was performed by introducing into the apparatus a mixture of the diphenylpicrylhydrazyl (10 cc.) and ethanol (1 cc.) and repeating the above sequence of operations. The data obtained are shown in Figure 4. Previous experiments had shown that the colorimeter density measurements were pro- portional to the concentration of diphenylpicrylhydrazyl.

8.   Pyrolysis of Trimethylene Disulfide in n-Octane. - A solution of tri- methylene disulfide in light petroleum was prepared by diluting a freshly dis- tilled alcoholic solution of the substance with water and extracting with several small portions of light petroleum and washing and drying the extract. This extract was diluted with approximately twenty times its volume of n-octane to give a solution of approximately 6.8×10^{-2} $\underline{M}$. Five cc. portions of this mixture were pipetted into thick-walled tubes, deoxygenated, sealed and heated in the dark at 141° and 174°. After appropriate time intervals, the tubes were opened and the spectra of the contents recorded with a Cary recording spectrophotometer. The height of the 3300 $\overset{o}{A}$ band was used to estimate the amount of the disulfide which had been destroyed. The graph (Figure 5) shows the data obtained. From the linear plot of unreacted disulfide against time, the pyrolysis would seem to be first order with respect to the disulfide. Because the products of the reaction appeared to have an adsorption band at short wavelengths, the tail of which extended as far as 3300 $\overset{o}{A}$, an attempt was made to correct for this by taking the height of the main adsorption band of trimethylene disulfide at wavelengths larger than 3300 $\overset{o}{A}$. The fact that the slope of the graphs for various wavelengths, changes little indicates that this effect is probably small. The most accurate kinetic data are probably derived

from the results at 3300 $\overset{o}{A}$.

$\underline{174^\circ\ C}$. Velocity constant $k_2 = \dfrac{2.3}{t} \log \dfrac{X_o}{X}$ where X_o and X concentrations

of trimethylene disulfide at $t = 0$ and $t = t$ sec.

Substitute, $X_o = 0.65$; $X = 0.44$, $t = 14,400$ secs.

$$k_2 = 2.70 \times 10^{-5} \text{ secs}^{-1}$$

$\underline{141^\circ C}$. $X_\bullet = 0.74$, $X = 0.60$, $t = 16,800$ secs.

$$k_\bullet = 1.24 \times 10^{-5} \text{ secs}^{-1}$$

Substituting these values for k_1 and k_2 into the relation:

$$\Delta H_2 = \frac{RT_1 T_2}{T_2 - T_1} \times 2.3 \log \frac{k_2}{k_1} \quad \text{gives} \quad \Delta H = 8.7 \text{ Kcals.}$$

9. <u>Photolysis of Trimethylene Disulfide at Low Temperatures in Ether-Pentane-Alcohol/Glass</u>. - (Figure 6)   A freshly distilled solution of trimethylene disulfide in aqueous ethanol was diluted with water and extracted into n-pentane. The extract was washed with water, dried over calcium sulfate and then diluted with ether and alcohol in the ratio: pentane extract 5 parts, ether 5 parts, alcohol 2 parts, and the spectrum recorded. The solution was introduced into a one centimeter Corex spectrophotometer cell, deoxygenated in nitrogen and sealed. The cell was mounted in a quartz Dewar attached to a Beckman spectrophotometer and rapidly cooled in liquid nitrogen. When the solution had set to a transparent glass, the spectrum was taken after first pumping off the liquid nitrogen to a level below that of the light path. The cell was reimmersed in liquid nitrogen and ultraviolet radiation from a high pressure mercury arc filtered by a 2 mm. Corning filter No. 7740 was focussed upon the disulfide with a quartz lens. The spectrum of the solution was taken again after 20 and 30 minutes, by which time very little, if any, of the disulfide was left. At this point the solution was a delicate pale salmon pink color. Sudden recrystalliza-

tion of the glass prevented our examining the spectrum at wavelengths longer than 4200 Å. The bulk of the liquid nitrogen was now distilled off, but the cells were kept in the Dewar flask so that they might warm up very slowly in the dark. When they had attained room temperature, the cell originally containing the disulfide was found to be turbid. The solution was clarified by ultracentrifuging and its spectrum taken. The reappearance of the characteristic 3300 Å band of the disulfide with an optical density of 0.95 indicates a $>50\%$ recovery of the disulfide.

10. **Equilibrium between Trimethylene Disulfide and β-Mercaptoethanol. -**

(a) A freshly distilled aqueous ethanol solution of trimethylene disulfide examined spectroscopically in a 1 cm. cell had an optical density of 1.42 at 3300 Å. β-Mercaptoethanol (23.5 mg. = ca. 20 λ) was weighed into a spectrophotometer cell, trimethylene disulfide solution (3.0 cc.) added, and the optical density of the 3300 Å band recorded on a Cary spectrophotometer as a function of time. The optical density dropped rapidly at first and then leveled off to a constant value of 1.06.

(b) The sequence of operations described in (a) was repeated using β-mercaptoethanol (110 mg. $\cong$ 100 λ) and trimethylene disulfide solution of initial optical density of 1.68. Final optical density = 0.65.

(c) The equilibrated solution of (b) was diluted with two volumes of ethanol. This caused the optical density to _increase_ to a final value of 0.365.

Assume equilibrium being measured is:

$$\text{(trimethylene disulfide)} + RSH \longrightarrow R-S-S-CH_2-CH_2-CH_2-SH$$
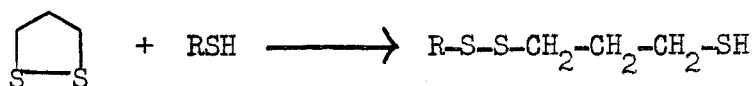
if a, b and 0 are the initial concentrations of these three entities respectively, then at equilibrium after $\underline{x}$-moles of reaction, their concentrations are (a-x), (b-x),[47] and x and the equilibrium constant will be:

_____

(47)   Note that the total thiol group concentration remained constant and equal to b for any amount of reaction x.

_____

$$K = \frac{\text{aliphatic disulfide}}{\text{cyclic disulfide x thiol}^{47}} = \frac{x}{b(a-x)}$$

$$\text{For (a) above, a} = \frac{1.42}{147} \times \frac{3.0}{3.02} = 0.960 \times 10^{-2} \text{ moles/liter}$$

where 147 = extinction coefficient of trimethylene disulfide

and the factor $\frac{3.0}{3.02}$ corrects for dilution of disulfide

by the added β-mercaptoethanol

$$a - x = \frac{1.06}{147} = 0.721 \times 10^{-2} \text{ moles/liter; } x = 0.239 \times 10^{-2} \text{ moles/liter}$$

$$b = \frac{23.5}{3 \times 78} = 10.05 \times 10^{-2} \text{ moles/liter}$$

Hence K = 3.30

for (b) and (c), K = 3.31 and 3.20, respectively.

11.   <u>Equilibrium between Trimethylene Disulfide and Benzyl Mercaptan.</u> -

(A)   (1)   The experiment described above for β-mercaptoethanol was repeated using benzyl mercaptan (20 λ) which was added to trimethylene disulfide, initial and final optical densities = 1.75, 0.97, respectively. Since density of benzyl mercaptan = 1.058, K = 13.9.

(2)   More benzyl mercaptan (40 λ) was added to the equilibrated solution; new final optical density = 0.545. Hence, K = 12.8.

(3) The equilibrated solution of (2) was diluted with two volumes of ethanol; new equilibrium optical density = 0.33. Hence K = 12.7.

(B)   (1)   Benzyl mercaptan (5 λ) added to trimethylene disulfide, initial and final optical densities of 3300 $\overset{o}{A}$ band were 1.44 and 1.20, respectively. Hence, K = 13.8.

(2)   More benzyl mercaptan (15 λ) added; new equilibrium optical density = 0.81. Hence K = 13.4.

(3)   Equilibrated solution of (2) was diluted with 15 volumes of ethanol.

Equilibrium optical density measured in 5 cm. cell = 0.43. Hence, K = 11.1.

- - - - -

Strain Energy of Trimethylene Disulfide

Equilibrium determinations similar to those described above were performed in 1 cm. photometer cells surrounded by a jacket through which water at constant temperature circulated. At equilibrium the actual temperature in the reaction vessel was recorded.

(a)   24.1° C: Freshly distilled trimethylene disulfide solution (optical density = 1.55) (3.0 cc.) was added to benzyl mercaptan (26.0 mg.). At equilibrium optical density = 0.71. Hence $K_{24.1}$ = 16.9.

(b)   35.8°C: Trimethylene disulfide solution (optical density = 1.51) (3.0 cc.) added to benzyl mercaptan (28.6 mg.). At equilibrium optical density = 0.81. Hence, $K_{35.8}$ = 11.23.

Substitution of these data into the isochore:

$$\log \frac{(K)_2}{(K)_1} = \frac{-\Delta H}{4.576} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

gives $\Delta H = -6.34$ Kcals.

12. <u>Catalyzed Photo-oxidation of Trimethylene Disulfide.</u> -

(a) The photo-oxidation was followed in a Warburg apparatus at $18°$ C. using solutions of zinc tetraphenylporphin in benzene and trimethylene disulfide in ligroin. The following mixtures were introduced into Warburg vessels;

(A) Prophin solution 0.5 cc., benzene 2.5 cc. (B) Porphin solution 0.5 cc., ligroin 2.5 cc. (C) Porphin solution 0.5 cc., ligroin 2.3 cc., trimethylene disulfide 0.2 cc. (D) Trimethylene disulfide 0.2 cc., ligroin 2.8 cc.

(E) Trimethylene disulfide 0.2 cc., benzene 2.8 cc. Oxygen consumption by these mixtures was followed both in the dark and in the light. Figure 7 exhibits the data obtained.

(b) The solution of zinc tetraphenylporphin in ethanol (2.4×10^{-4} $\underline{M}$) and a freshly distilled aqueous ethanol solution of trimethylene disulfide (0.972×10^{-2} $\underline{M}$) were made up. The following mixtures of these solutions were introduced into Warburg vessels. (A) into two (No. 1 and 2) vessels trimethylene disulfide 1 cc., porphin 0.2 cc., ethanol 1.5 cc., (B) into two vessels (No. 3 and 4) trimethylene disulfide 2 cc., prophin 0.3 cc., ethanol 0.5 cc. After equilibrating at $11.2°$ C. in the dark, the vessels were illuminated and the oxygen consumption obtained as a function of time. The graph (Figure 8) records the results obtained and shows that slightly more than one atom of oxygen was consumed per mole of trimethylene disulfide. In a similar experiment conducted at $24°$ C. rather more oxygen was consumed (Figure 9). When the oxidations were

complete the solutions were combined, diluted with a little water, and fractionated under reduced pressure. The spectrum of the ethanol distillate which apparently contains the sulfoxide derived from trimethylene disulfide is shown in Figure 10 and compared with the sulfoxide of 6,8-thioctic acid.

13. **Chemical Oxidation of Disulfide.** - Since H_2O_2 did not give a simple oxidation of the disulfide other oxidizing agents were tried. Ammonium persulfate was found to be suitable. To 9.4 cc. of trimethylene disulfide in 90% ethanol (O.D. 1.67 = 1.17 x 10^{-2} $\underline{M}$) was added 110 λ of 1 $\underline{M}$ $(NH_4)_2S_2O_4$ (aq. sol'n). This was allowed to stand at room temperature for 16 hours. The disulfide band rapidly disappears and is replaced by the monoxide band at 2450 $\overset{o}{A}$. Assuming the extinction coefficient of the sulfoxide equals 1025 the calculated height of the 2450 $\overset{o}{A}$ band (diluted 10 times) = $\dfrac{1.67 \times 1025}{147 \times 10}$ = 1.16. Found = 1.15 (Figure 11). A corresponding result was obtained with 6-thioctic acid.

<center>PHOTOLYSIS OF TRIMETHYLENE DISULFIDE IN</center>

<center>PRESENCE OF ACIDS</center>

A. **With α-ketoglutaric acid.** - Into each of two pyrex test tubes was introduced 6 cc. of trimethylene disulfide (ca. M/100) in 80% EtOH + 1 cc. of M/2 solution α-ketoglutaric acid in 80% EtOH. Two control tubes each containing 6 cc. of the trimethylene disulfide solution alone were also made up. The tubes were deoxygenated in a stream of N_2, sealed and exposed in pairs to the following conditions: (a) one pair was kept in the dark for 24 hours; (b) the other pair was illuminated for 3 hours under an 8 W. u.v. lamp. Spectroscopic investigation of the tube contents demonstrated that no change occurred with (a). With (b) the tube containing the α-ketoglutaric acid remained clear, while the control rapidly became turbid then milky. Further, in the clear solution, all the trimethylene disulfide had been destroyed but no new characteristic absorption

peak appeared.

B.   <u>With acetic, oxalic and pyruvic acids - Quantitative Data</u>. - In these

experiments, the concentration of trimethylene disulfide in a freshly distilled

ethanolic solution was estimated spectroscopically. The trimethylene disul-

fide mixed with the required amount of the organic acid was diluted with water

to $\underline{ca}$. 80% EtOH, deoxygenated in nitrogen, sealed in Pyrex glass tubes and

photolyzed under a bank of six 4-W tubular u.v. lamps for 1-3/4 hours. With

acetic acid, a turbidity slowly developed but with the other acids, the solu-

tions were clear after photolysis. Controls without added acid rapidly became

turbid. The cell contents were examined spectroscopically in order to confirm

the complete destruction of the trimethylene disulfide and were then titrated

with microburettes (a) potentiometrically against NaOH, (b) against KI_3, (c)

against Ag^+ amperometrically to estimate thiol groups (following the procedure

of Rosenberg, <u>et al</u>.[48] The results of these titrations are exhibited in Table 1.

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(48)   S. Rosenberg, J. C. Perrone and P. L. Kirk, Anal. Chem., <u>22</u>, 1186 (1950).

<u>Estimation of thiols by titration against I_2 and Ag^+</u>. - In view of the results

of the preceding experiments it became desirable to learn something of the

validity of these methods for estimating thiols. N/100 solutions of β-mercapto-

ethanol and trimethylene dithiol in aqueous EtOH ($\underline{ca}$. 80%) which had been made

up several months previously were titrated on the micro scale against $\underline{N}$/10 I_2

(using starch) and amperometrically against $\underline{N}$/10 $AgNO_3$ using the procedure of

Rosenberg, <u>et al</u>.[48] with the results shown in Table 2.

Table 1

ANALYSIS OF SOLUTIONS OBTAINED BY

PHOTOLYZING TRIMETHYLENE DISULFIDE IN THE

PRESENCE OF ORGANIC ACIDS

Added Acid	Acidity of solution before photolysis	Acidity of solution after photolysis	Conc. of -SH after photolysis (Ag$^+$ titration)	Conc. of -SH after photolysis (I_2 titration)	Conc. of trimethylene disulfide before photolysis
Acetic	2.43×10^{-2} N	2.46×10^{-2} N	—	2.1×10^{-2} N	2.0×10^{-2} M
Pyruvic	2.3×10^{-2} N	2.22×10^{-2} N (a)	1.2×10^{-2} N	2.3×10^{-2} N	2.0×10^{-2} M
Oxalic	4.38×10^{-2} N	4.44×10^{-2} N (a)(b)	1.5×10^{-2} N	2.4×10^{-2} N	2.0×10^{-2} M

Notes: - (a) Solutions tested for presence of CO_2 by bubbling pure N_2 through portions of the photolyzed mixtures into NaOH then adding $BaCl_2$ - no CO_2 was found. (b) Titre virtually unchanged after heating to 200°C. for ca. 10 minutes. Also unchanged oxalic was found to be present by treating with alkaline $CaCl_2$, collecting the precipitate, washing, precipitating the Ca++ with 2 N H_2SO_4, and testing the supernatant solution for oxalic acid by reduction to glycolic acid and treating with 2:7-dihydroxynaphthalene and conc. H_2SO_4.[49] Also the presence of oxalic acid was demonstrated by paper chromatography.[50]

(49) F. Geigl, "Qualitative Analysis by Spot Tests," 1946, Elsevier, New York.

(50) A. A. Benson, Private communication.

Table 2

	-SH estimated by I_2	-SH estimated amperometrically Ag^+
β-mercaptoethanol	5.1×10^{-3} N	3.55×10^{-3} N
Trimethylene dithiol	9.2×10^{-3} N	5.6×10^{-3} N

However, new solutions made up from freshly distilled thiols and from pure cysteine hydrochloride gave the results shown in Table 3.

Table 3

	-SH estimated by I_2 titration	-SH estimated by Ag^+ titration
β-mercaptoethanol	1.1×10^{-1} N	1.08×10^{-1} N
Cysteine hydrochloride	1.2×10^{-1} N	1.15×10^{-1} N
Trimethylene dithiol	1.13×10^{-1} N	1.08×10^{-1} N (a)

Note: - (a) Precipitate formed during titration.

It thus appears that in the absence of certain unidentified interfering substances I_2 and Ag^+ titrations give, within the limits of experimental error, the same value for thiol concentration and that the presence of ethanol does not noticeably interfere.

C. Photolysis of trimethylene disulfide in presence of hydrochloric acid - Quantitative Data. - A freshly distilled aqueous ethanol solution of trimethylene disulfide had optical density for 3300 Å peak = 2.05 corresponding to 1.39×10^{-2} M. 25.0 cc. of this solution and 400 λ of 0.95×10^{-1} N aqueous HCl (calc. for 1 equiv. = 366 λ) was introduced into a modified Pyrex "lollipop" (a vessel with flat sides furnished with an exit tube and a capillary reaching almost to the

bottom of the vessel, both tubes being sealed to stopcocks). The solution was deoxygenated for 20 minutes in a stream of pure N_2 saturated in alcohol vapour introduced through the capillary and then isolated from the atmosphere by turning the stopcocks. Photolysis was conducted for 1-3/4 hours at a distance of ca. 2" from a bank of eight 4-W tubular u.v. lamps, the apparatus being kept cool with an air blast. At the end of the photolysis, the solution was quite clear and spectroscopic examination (Figure 12) revealed the complete disappearance of the trimethylene disulfide 3300 $\overset{o}{A}$ band. Lead acetate papers showed the presence of a trace of H_2S in the photolysis solution. The solution was then titrated.

(a) **Potentiometric**

Initial solution before photolysis; 300 $\lambda = 46.6 \lambda$ of 0.1 $\underline{N}$ NaOH,

$\therefore [H^+] = 1.56 \times 10^{-2} \underline{M}$.
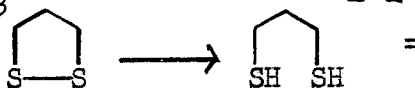
Solution after photolysis, 300 $\lambda = 47.95 \lambda$ of 0.1 $\underline{N}$ NaOH, $\therefore [H^+]$
$= 1.60 \times 10^{-2} \underline{N}$.

(b) **Iodometric**

Photolysis solution 250 $\lambda = 58.8, 58.2 \lambda$ of $1.06 \times 10^{-1} \underline{N}$ I_2,

$\therefore [-SH] = 2.48 \times 10^{-2} \underline{M}$. The end point, observed with the aid of starch, was not good — the blue color rapidly disappeared.

(c) **Amperometric**

300 λ of photolysis solution = 35.4, 36.2, 35.6, 35.8, 34.2, 33.6 λ
of $1.02 \times 10^{-1} \underline{N}$ AgNO$_3$. 35.1 λ average $\therefore [SH] = 1.2 \times 10^{-2} \underline{N}$.

Calc. value of -SH assuming reaction =

$2 \times 1.39 \times 25/25.6 = 2.7 \times 10^{-2} \underline{N}$.

<u>Xanthates and Trithiocarbonates.</u> - β-Mercaptoethanol, benzyl mercaptan and <u>t</u>-butyl mercaptan dissolved in ethanol were mixed in a Cary spectrophotometer cell with slightly more than the calculated amount of carbon disulfide, then treated with 1 <u>N</u> sodium hydroxide. The yellow color of trithiocarbonates developed instantaneously. Figure 13 shows the spectra obtained. A similar experiment, conducted with a specimen of the solution obtained by photolyzing trimethylene disulfide in the presence of hydrochloric acid, gave the same characteristic peaks. When alkali was similarly added to a solution of carbon disulfide, alone, in ethanol, the paler color of the xanthate (see Figure 11) developed much more slowly.

<u>Properties of Photolyzed Solution</u> .

(a)   Spectrum of photolyzed solution is virtually unchanged by heating 5 cc. to 70° C. for 20 minutes with 50 λ of conc. HCl.

(b)   On making just alkaline with NaOH, the photolysis solution slowly becomes yellow as a peak gradually develops at <u>ca</u>. 3400 $\overset{o}{A}$ (Figure 14) and later turbid. After standing overnight and centrifuging, the solution possesses a good peak at 3380 $\overset{o}{A}$. This is not due to trimethylene disulfide since, on acidifying, the solution becomes turbid, and the peak largely disappears and distillation of the solution under reduced pressure gives a distillate with no absorption at 3300 $\overset{o}{A}$.

(c)   Making the solution strongly alkaline and distilling under reduced pressure gives no trimethylene disulfide. Apart from end absorption below <u>ca</u>. 2300 $\overset{o}{A}$, the distillate shows only a minor peak at <u>ca</u>. 2700 $\overset{o}{A}$ due to traces of thiacyclobutane in the original trimethylene disulfide.

(d)   Reduced pressure distillation of the photolysis solution gives a spectrum similar to that in (c) except that the 2700 $\overset{o}{A}$ peak is now an inflexion on side of greater end absorption.

(e)   Oxidizing the thiol groups in photolysis solution with I_2 and then distilling under reduced pressure gave a distillate with a peak at 3380 Å, which for some time was believed to be due to impure trimethylene disulfide (λ_{max} = 3300 Å) derived by oxidation of trimethylene dithiol present in the photolysis solution. However, this possibility was eliminated because reduced pressure distillation of the acidic photolysis solution gave no dithiol since oxidation of the distillate with I_2 gave a turbid solution which on redistillation gave no trimethylene disulfide. A control experiment in which a small amount of synthetic dithiol was first added to the photolysis solution before distillation, I_2 oxidation and redistillation gave the 3300 Å peak of trimethylene disulfide, thus demonstrating that the photolysis solution did not destroy trimethylene dithiol and that hence there was no trimethylene dithiol present in the original photolysis solution. Another control experiment (Figure 15) in which trimethylene disulfide was added to the photolysis solution which was then distilled showed trimethylene disulfide in the distillate, thus proving that trimethylene disulfide also was not destroyed by the photolysis solution.

(f)   Further proof of the absence of trimethylene dithiol in the photolysis solution was obtained from the observation that trimethylene dithiol could be oxidized virtually quantitatively to trimethylene disulfide by hydroxylamine. A few λ of trimethylene dithiol in 3 cc. of $\underline{ca}$. 80% aqueous ethanol was warmed to 50° C. in a Cary spectrophotometer cell with 3.5 $\underline{N}$ NaOH (600 λ) and 4 $\underline{N}$ NH_2OH, HCl solution (300 λ). The rapid formation of trimethylene disulfide could be observed from the development of the characteristic 3300 Å peak. In a similar experiment in which the photolysis solution was treated with alkali and hydroxylamine, no development of the 3300 Å peak could be observed.

(g)   No H_2O_2 was present in the photolysis solution since (i) adjusting the

pH to 7, diluting with water and adding catalase caused no liberation of O_2

although catalase added to alcoholic H_2O_2 of equivalent ethanol concentration

did give off oxygen; (ii) bleaching of PbS-impregnated paper failed to occur.

Also a photolysis experiment conducted with a piece of PbS paper suspended in

the solution, failed to cause bleaching of the pigment.

Ethyl t-butylsulfenate

This was prepared according to the directions of Rheinholdt and Motzkus.[35]

It was a colorless mobile, pleasant smelling liquid, b.p. 65°/90mm. Kept in

sealed glass ampoules at 5° C. it is quite stable but exposed to the air it

rapidly turns yellow and appears to change at least partly into di-t-butyl

disulfide. The spectrum in ethanol showed $\lambda_{max} = 2655$ Å, $\varepsilon_{max} = 70.5$

(Figure 16).

Effect of acid and alkali on ethyl t-butylsulfenate

(a)   The sulfenic ester (5λ) was dissolved in ethanol (3 cc. of 80%) and the

spectrum taken. 70 λ of 1 N NaOH was added and the spectrum retaken. A slight

increase in absorption throughout the spectrum was observed but no specific

change. No further change occurred even after 12 hours.

(b)   The sulfenic ester (5λ) in 3 cc. of 80% aqueous ethanol was treated in

a spectrophotometer cell with 70 λ of 1 N HCl and spectra recorded at appropriate

intervals of time. A rapid change occurred in the spectrum leading to increased

absorption throughout the spectrum. After standing overnight, all the sulfenic

esters had been destroyed and a new as yet unidentified peak had appeared at

ca. 2900 Å. (Figure 16)

Di-t-butyl disulfide

This was prepared by oxidizing an alcoholic solution of t-butyl mercaptan with aqueous KI_3 solution. No definite end point could be observed. As the iodine solution was added, the solution became browner and the characteristic smell of sulfenyl iodides gradually developed. The excess I_2 was destroyed with $Na_2S_2O_3$ and the disulfide isolated with pentane and distilled. The disulfide was obtained as a yellow oil, b.p. $85^{\circ}/25$ mm. Anal: Found C, 54.25; H, 9.86; S, 36.1, $C_8H_{18}S_2$ require C, 53.4; H, 10.1; S, 36.0%. Repeated fractionation failed to remove the yellow color. The u.v. spectrum (in EtOH) in contrast to other aliphatic disulfides showed no characteristic band at 2500 $\overset{o}{A}$; instead it revealed only an end absorption extending into the visible (Figure 16).

Oxidations by Ethyl-t-butylsulfenate and by di-t-butyl disulfide.

A Cary spectrophotometer cell was fitted with an attachment provided with two side arms to hold reagents, with an inlet capillary and an exit tube, both tubes being furnished with stopcocks.

(a) Into the Cary cell was introduced 90% ethanol (3 cc.) and 1 N aqueous KI (50λ). Into the side arms of the attachment was introduced the sulfenic ester (5λ) in ethanol (50λ) and 1 N, HCl (50λ) in ethanol (50λ). The attachment was assembled and the system deoxygenated in a stream of purified nitrogen for 10 minutes beyond the time the effluent gases showed zero oxygen content in a Pauling oxygen analyzer. The taps were then closed. The HCl was then dumped into the Cary cell and the very slow development of the I_2 color induced by traces of O_2 was plotted as a function of time. The sulfenic ester was then tipped into the HI solution. I_2 was immediately liberated and continued to be formed at a considerable speed ca. 500 times greater than that due to traces of O_2.

(b) A similar experiment was performed using di-t-butyl disulfide (5λ) in EtOH

(50 λ) in one side arm, HCl (50 λ of 1 $\underline{N}$) in EtOH (50 λ) in the other and with

1 $\underline{N}$ KI (50 λ) in aqueous EtOH (3 cc.) in the spectrophotometer cell. Again on

adding the disulfide a considerable increase in the rate of liberation of iodine

was observed.

(c)   To investigate the oxidation on thiols, the sulfenic ester (10 λ) in EtOH

(50 λ) and trimethylene dithiol (5 λ) in EtOH (50 λ) were introduced into the

side arms. The cell contained 1 $\underline{N}$ NaOH (50 λ) in EtOH (3 cc. of 90%). After

removing oxygen, the dithiol was added to the alkali; a small amount of tri-

methylene disulfide was formed slowly. Addition of the sulfenic ester caused

an instantaneous production of trimethylene disulfide (O.D. at 3300 $\overset{o}{A}$ = 0.75).

(d)   A similar experiment demonstrated that the sulfenic ester appears to be

relatively stable to thiols under neutral conditions (Figure 17).

(e)   A similar experiment conducted with di-$\underline{t}$-butyl disulfide (4 λ), and tri-

methylene dithiol (4 λ) also showed a production of trimethylene disulfide but in

this case the rate of formation was slow and the ultimate yield small.

Ethyl Trichlormethylsulfenate

This was prepared from trichlormethylsulfenylchloride by reaction, with

sodium ethoxide under conditions similar to those used for the preparation of

$\underline{t}$-butyl analogue. Anal: Found, C, 21.4; H, 3.1; OEt, 22.4. $C_3H_5SOCl_3$ re-

quires C, 18.4; H, 2.6; OEt, 23.0%. The absorption spectra of the chloride and

ester are shown in Figure 18 and Figure 19.

Oxidation of trimethylene dithiol by the photolysis solution.

In the apparatus in the diagram (Figure 20), the Cary cell (A) was

charged with ethanol (3 cc. of 90%) and 90 λ of 1 $\underline{N}$ NaOH, the side arm with

trimethylene dithiol (5 λ) + ethanol (150 λ) and the photolysis cell (B) with

trimethylene disulfide solution (3 cc., 1.2×10^{-2} $\underline{M}$) + 40 λ of 0.95×10^{-1}

$\underline{N}$ HCl ($\triangleq$ 1 mole prop.). Pure N_2 was bubbled through the capillaries for 10

minutes after the effluent gas had zero O_2 content according to a Pauling oxygen

analyzer and the taps were then turned. The trimethylene dithiol was dumped

into the Cary cell and a spectrum taken. A small amount of trimethylene

disulfide was found. The Cary cell was then wrapped in foil, the disulfide in

the photolysis cell exposed to a bank of eight 4-W tubular u.v. lamps at a dis-

tance of $\underline{ca}$. 1-1/2" for 2 hours, the apparatus being kept cool by an air blast.

The spectrum of the dithiol and alkali in the spectrophotometric cell was taken

again but no increase in the concentration of trimethylene disulfide had taken

place. The contents of the Cary cell were then tipped into the photolysis

cell, mixed thoroughly and poured back into the Cary cell. The spectrum showed

an increased absorption at 3300 $\overset{o}{A}$ which developed with time. The ultimate in-

crease in the height of the band corresponded to $\underline{ca}$. 15% yield in the oxidation

of trimethylene dithiol to trimethylene disulfide. Reduced pressure distilla-

tion of the acidified mixture gave a distillate containing trimethylene disulfide.

<u>Photolysis of Trimethylene Disulfide in the Presence of Sodium Hydroxide.</u>

A solution of trimethylene disulfide in $\underline{ca}$. 95% ethanol ($\underline{M}$/100) was

treated with 70λ of 1 $\underline{N}$ NaOH, freed from oxygen and photolyzed in a Cary spectro-

photometer cell before a bank of 6-W tubular u.v. lamps. Figure 21 shows the

progress of the photolysis and the effect of acidifying the photolysate.

Figure Captions

Figure 1 - Absorption spectra of some disulfides in alcohol.

Figure 2 - Schematic representation of the potential curves and transitions

for the cyclic and open-chain disulfides.

Figure 3 - Apparatus for the measurements on photosensitized fading of picryl-

hydrazyl.

Figure 4[51]- Photolysis of diphenylpicralhydrazyl in the presence of trimethylene

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(51)   Readings taken against a dark red filter No. 66.

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disulfide.

Figure 5 - Kinetics of pyrolysis of ⟨S—S ring⟩ in n-octane.

Figure 6 - Low temperature photolysis of trimethylene disulfide and its reversal.

Figure 7 - Photo-oxidation of ⟨S—S ring⟩ catalyzed by zinc tetraphenylporphin.

Figure 8 - Stoichiometry of photo-oxidation of ⟨S—S ring⟩ catalyzed by zinc

tetraphenylporphin at 11° C.

Figure 9 - Stoichiometry of photo-oxidation of ⟨S—S ring⟩ catalyzed by zinc

tetraphenylporphin at 24° C.

Figure 10 - Absorption spectra of disulfide monoxides in alcohol.
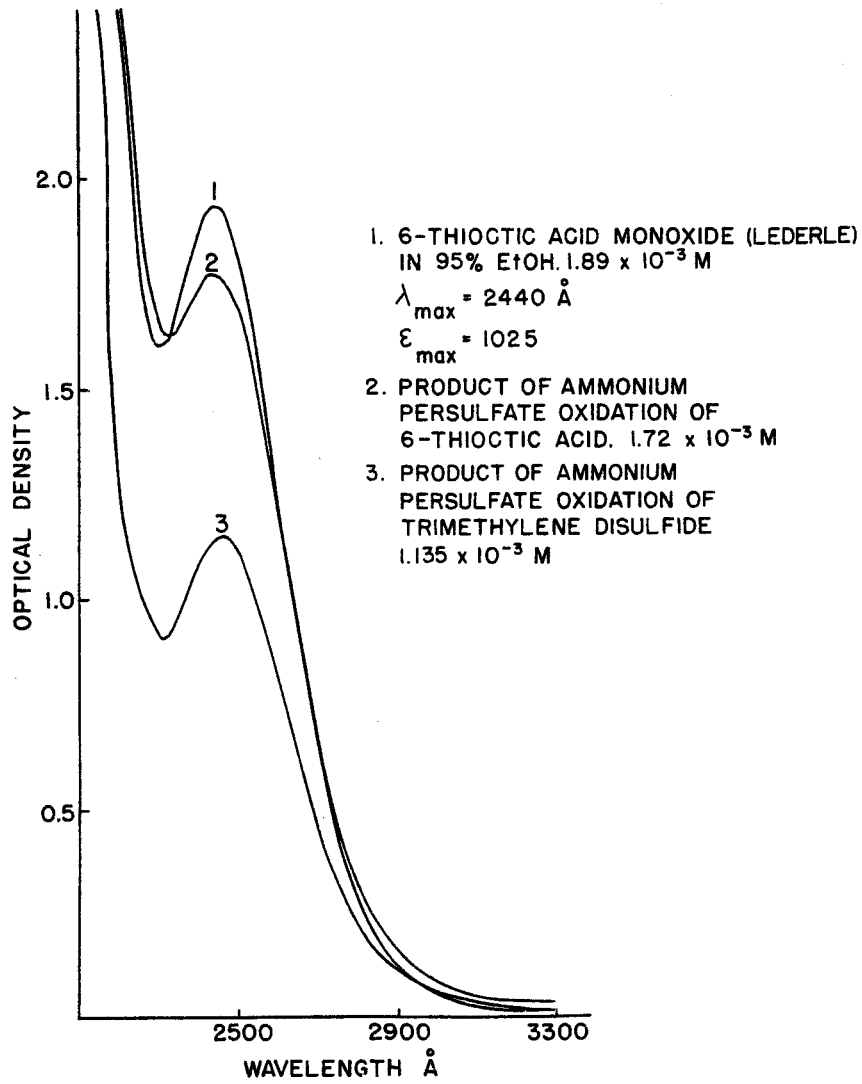
Figure 11 - Oxidation of trimethylene disulfide by ammonium persulfate.

Figure 12a - Photolytic behavior of acidic aqueous alcoholic solution of trimethylene

disulfide at room temperature.

Figure 12b - Trimethylene disulfide acid photolysis solution after standing and

with thiol compensation.

Figure 13 - Trithiocarbonate formed by photolysis solution.

Figure 1

MOLAR EXTINCTION COEFFICIENT

N-PROPYL DISULFIDE

4,8-THIOCTIC ACID

5,8-THIOCTIC ACID

TETRAMETHYLENE DISULFIDE

TRIMETHYLENE DISULFIDE

6,8-THIOCTIC ACID

WAVELENGTH-Å

Figure 2

MU-6999

Figure 3

PHOTOLYSIS OF DIPHENYL PICRYL HYDRAZYL IN THE PRESENCE OF TRIMETHYLENE DISULFIDE

1. IN DARK AFTER ILLUMINATION
2. PICRYL HYDRAZYL + ⬡ IN LIGHT
3. PICRYL HYDRAZYL IN DARK
4. PICRYL HYDRAZYL IN LIGHT
5. PICRYL HYDRAZYL + ⬡ IN DARK
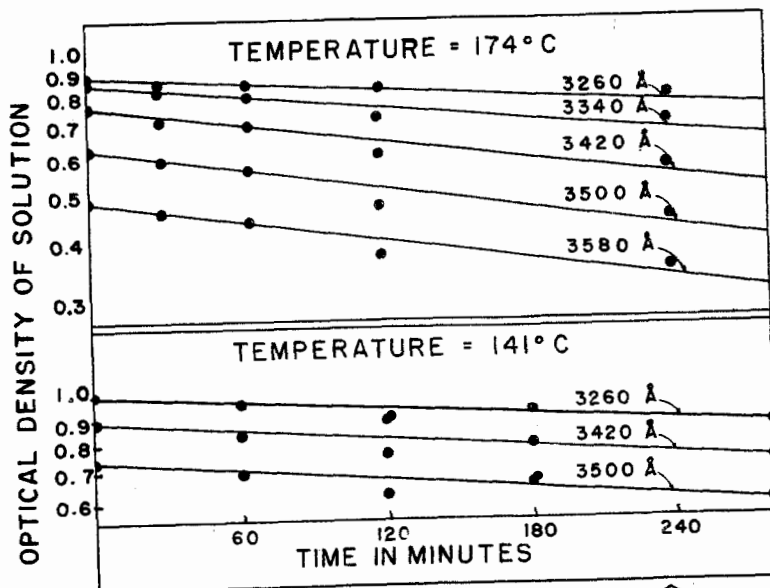
Figure 4

KINETICS OF PYROLYSIS OF
IN n-OCTANE

MU 5622

Figure 5
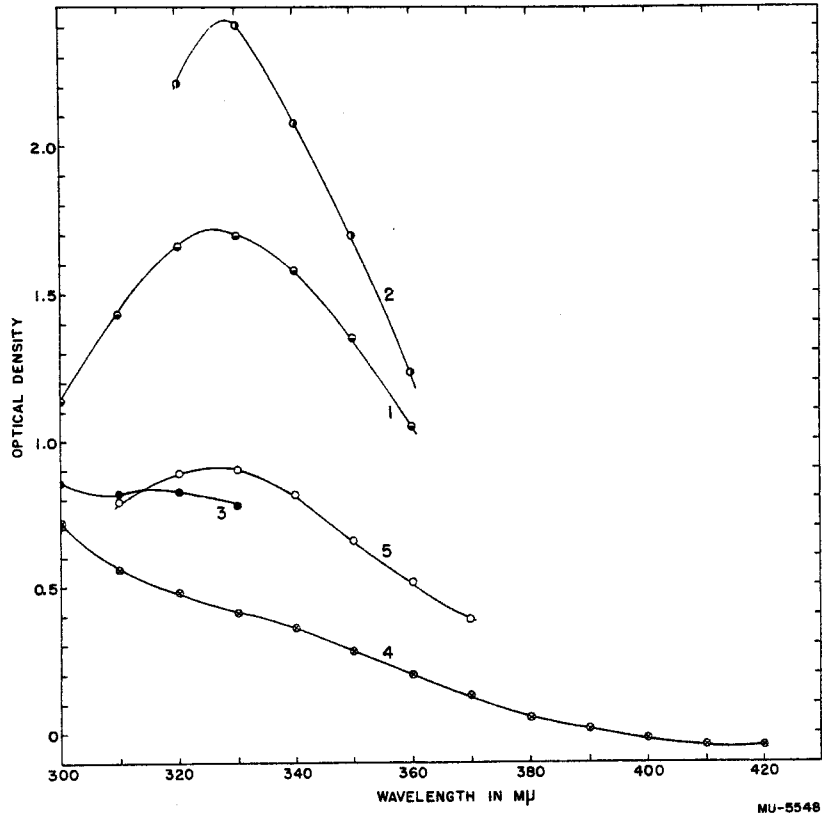
LOW TEMPERATURE PHOTOLYSIS OF TRIMETHYLENE DISULFIDE

1. ⬠ IN E.P.A. BEFORE PHOTOLYSIS (ROOM TEMP.)

2. #1 AT -196° C.

3. #1 AFTER 10 MIN. PHOTOLYSIS

4. #1 AFTER 20 MIN. PHOTOLYSIS

5 #4 AFTER WARMING SLOWLY TO ROOM TEMP.
  AND CENTRIFUGING

Figure 6
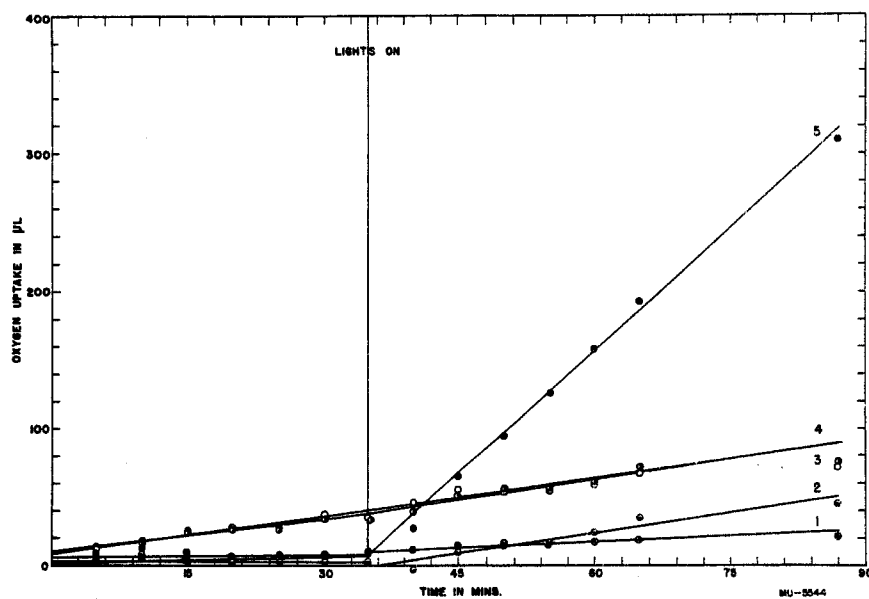
PHOTO-OXIDATION OF ⬠ CATALYZED BY ZINC TETRAPHENYL PORPHIN

1. ⬡ + P.E.

2. PORPHIN ● P.E.

3. PORPHIN + Bz

4. ⬡ + Bz

5. PORPHIN + ⬡ ● Bz + P.E.
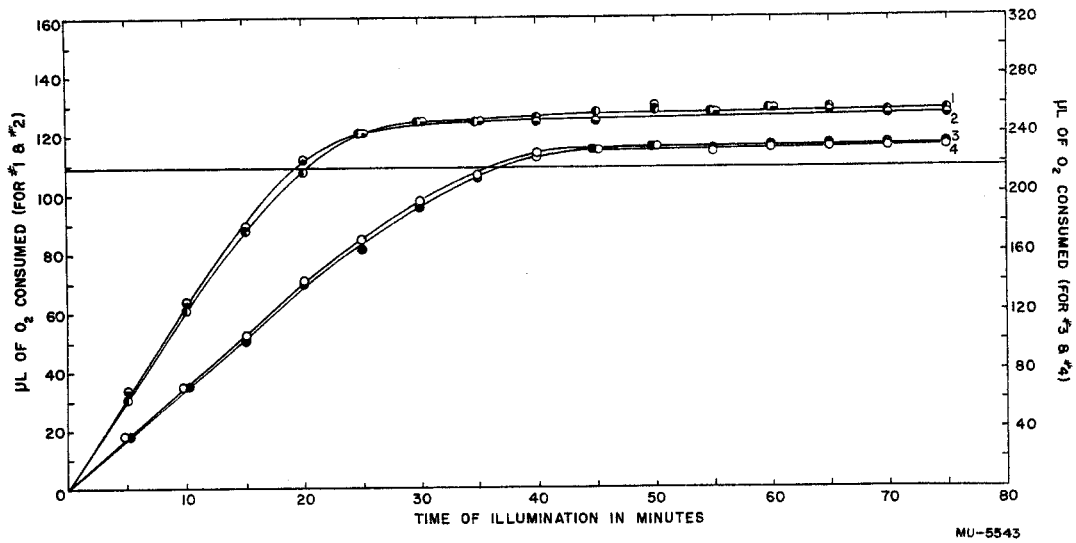
Figure 7

PHOTO-OXIDATION OF ⬠ CATALYZED BY ZINC TETRAPHENYL PORPHIN
   S S

O$_2$ CONSUMPTION CALCULATED FOR ONE ATOM
OF OXYGEN PER MOLE OF ⬠
                        S   S
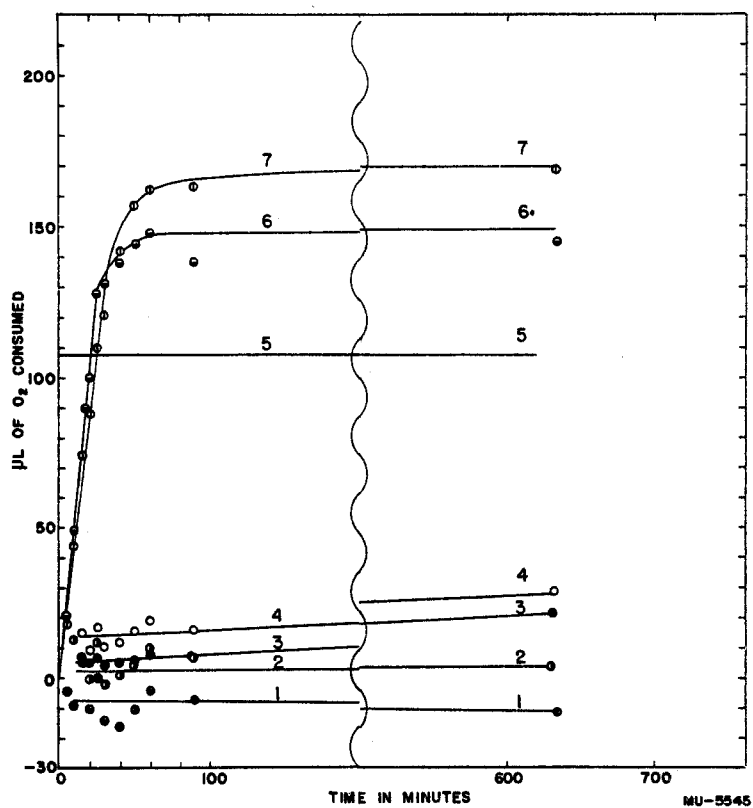
Figure 8

PHOTO-OXIDATION OF ⬠ CATALYZED BY ZINC TETRAPHENYL PORPHIN

I. & 2. ZINC TETRAPHENYL PORPHIN ALONE

3. & 4. ⬠ ALONE

5. CALCULATED O₂ ABSORPTION FOR
ONE ATOM OF O₂ PER MOLE OF ⬠

6. & 7. PORPHIN + ⬠
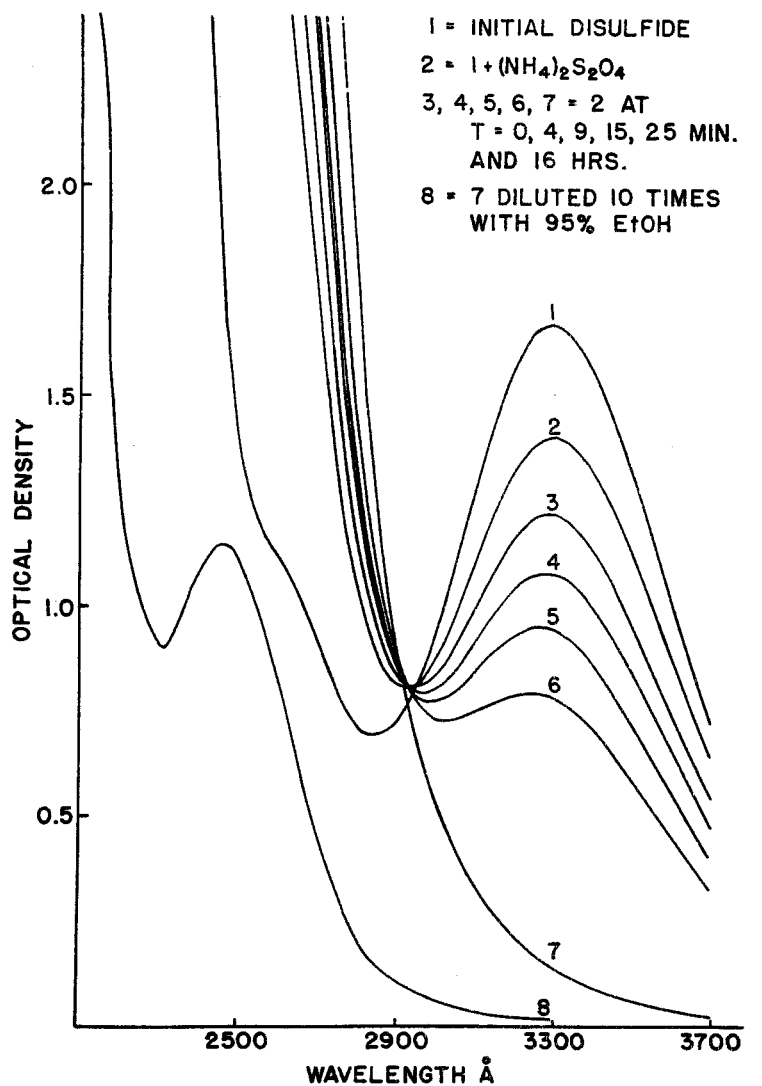
Figure 9

Figure 10

OXIDATION OF TRIMETHYLENE DISULFIDE BY AMMONIUM
PERSULFATE

MU-6990

Figure 11

1. ORIGINAL TMDS PEAK
2. TMDS AFTER 15 MIN. PHOTOLYSIS
3. "   "   60 "   "
4. "   "   105 "   "

OPTICAL DENSITY

WAVELENGTH Å

MU-6991

Figure 12a

1. PHOTOLYSIS SOL'N. (vs. EtOH)
2. PHOTOLYSIS SOL'N. 1×10^{-2} M
   vs. HS⎯OH 1.0×10^{-2} M
3. PHOTOLYSIS SOL'N. 1×10^{-2} M
   vs HS⎯OH 0.5×10^{-2} M
4. PHOTOLYSIS SOL'N. 1×10^{-2} M
   vs HS⎯OH 1×10^{-2} M AFTER
   STANDING OVERNIGHT
5. HS⎯OH 1.0×10^{-2} M vs EtOH

MU-7000
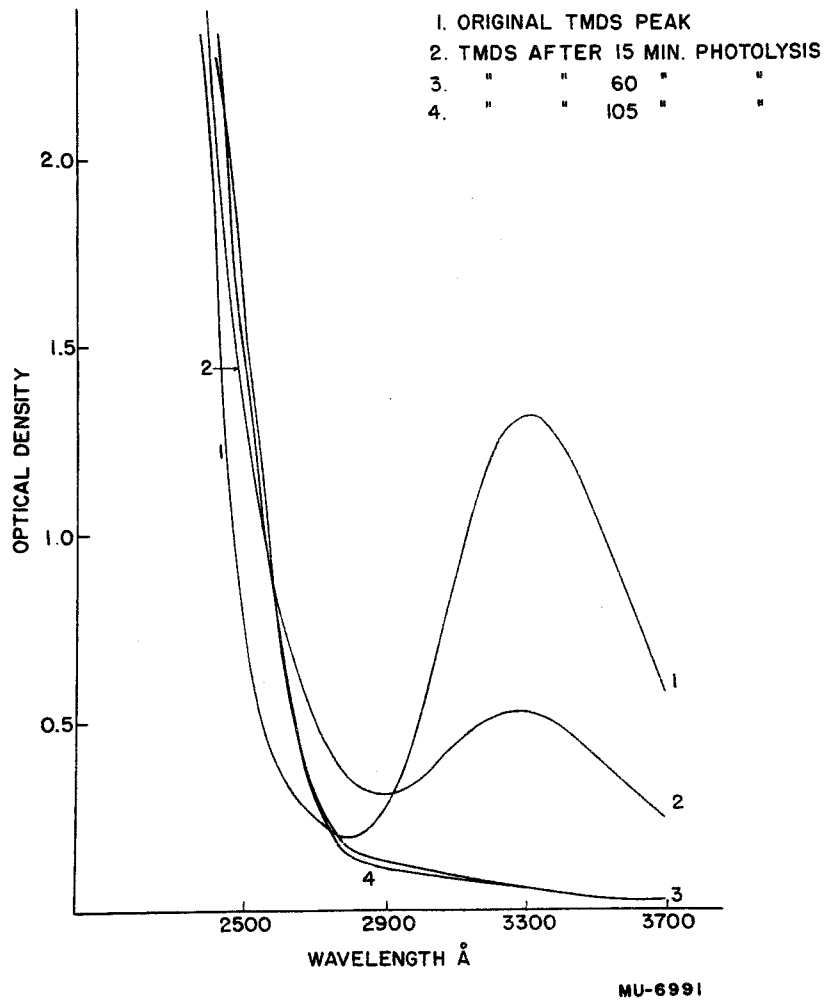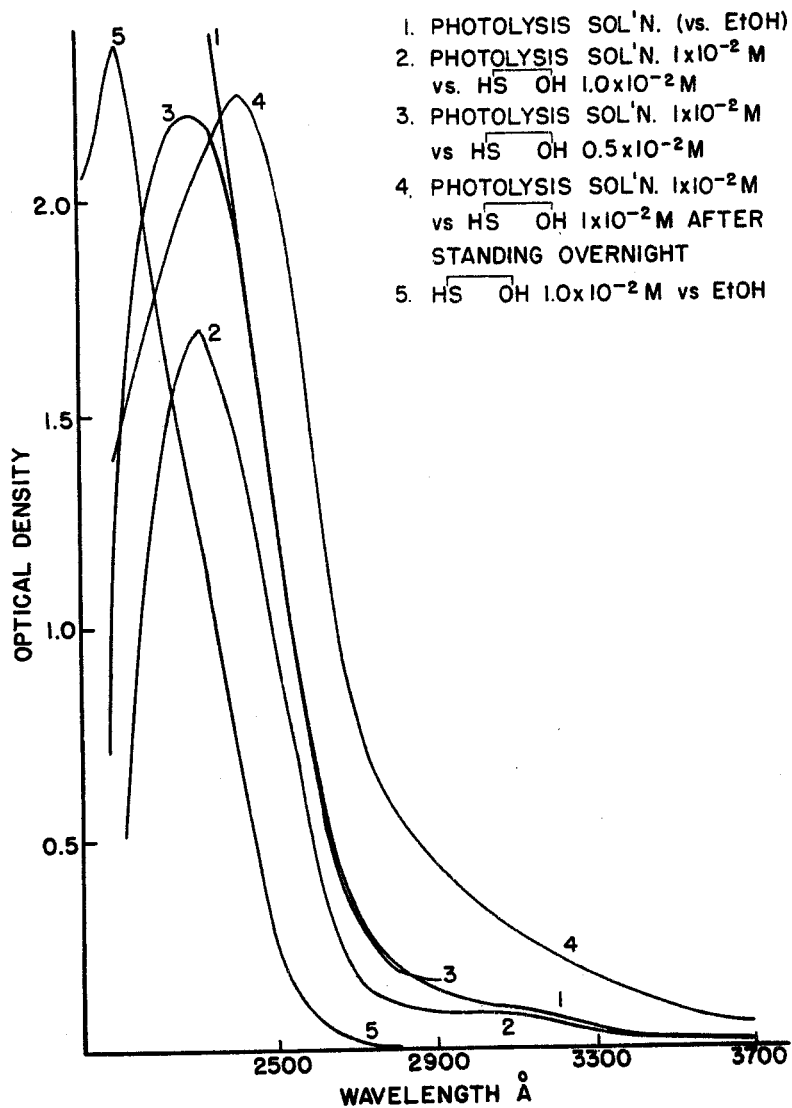
Figure 12b

Figure 13

Figure 14

Figure 15

Figure 16

1. ETHYL †-BUTYLSULFENATE
2.=1+ TRIMETHYLENE DITHIOL
3.=2 AFTER ~10 MIN.
4.=3 DILUTED ~1 VOLUME WITH
   EtOH AND ADDED K$_2$HPO$_4$
5.=4 AFTER 5 MIN.
6.=4 AFTER 15 MIN.
7.=4 AFTER 180 MIN.

MU-6995

Figure 17

1. TRICHLOROMETHYL SULFENYL
   CHLORIDE IN EtOH

2. TRICHLOROMETHYL SULFENYL
   CHLORIDE IN CCl$_4$
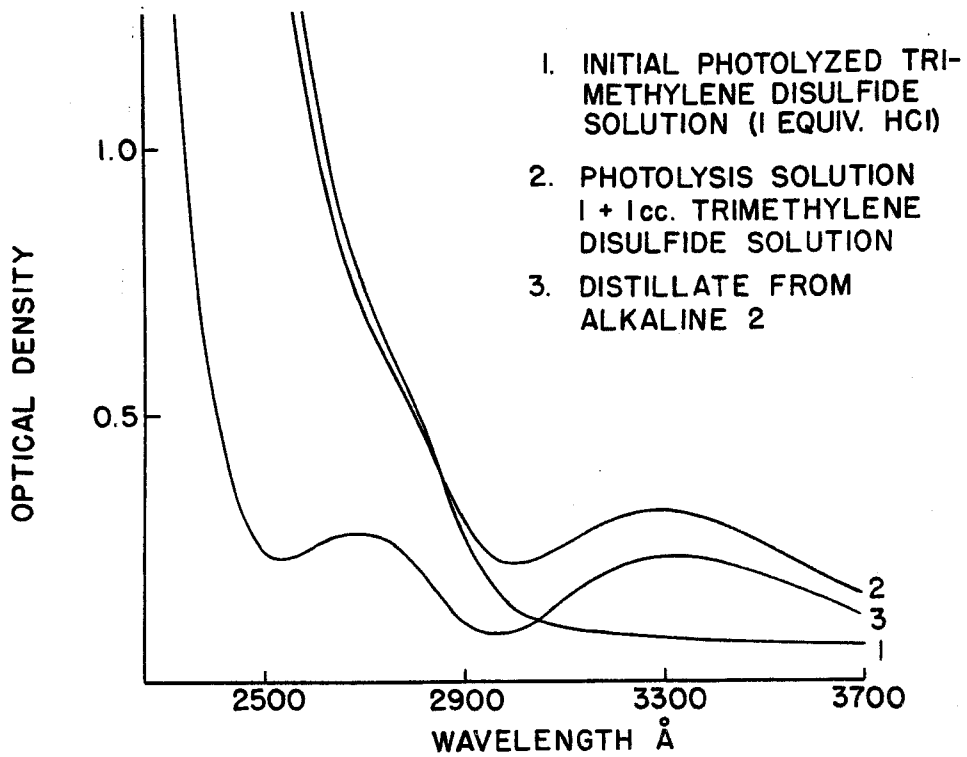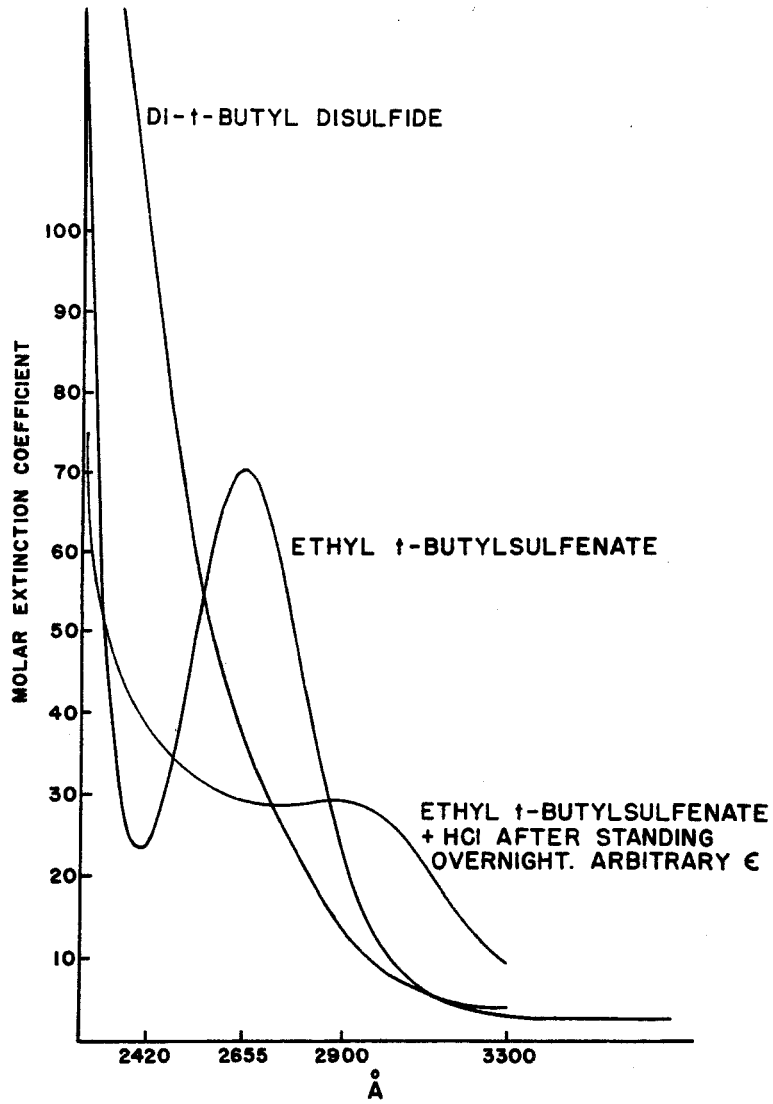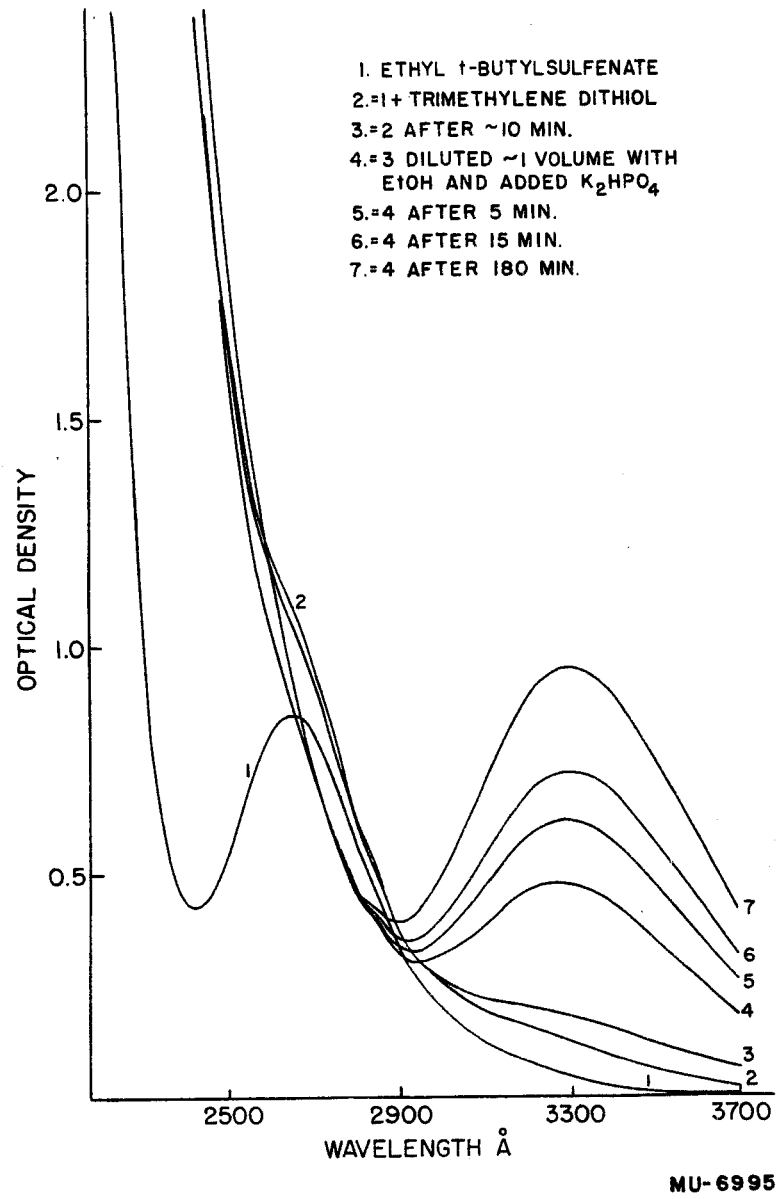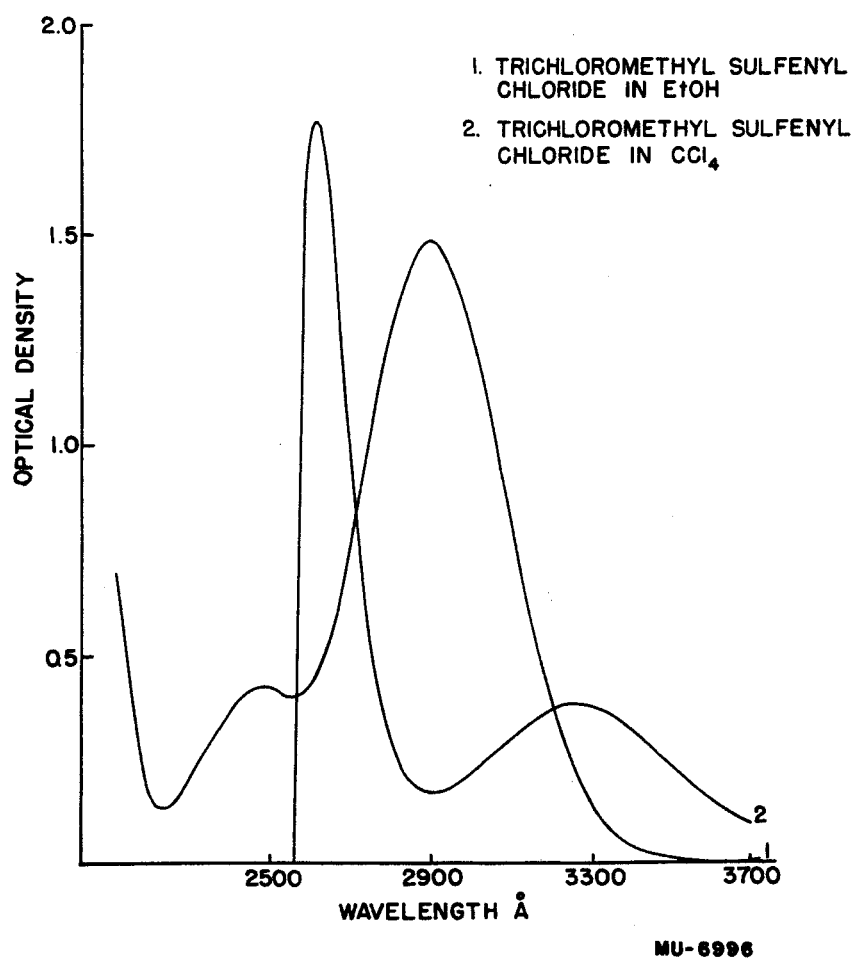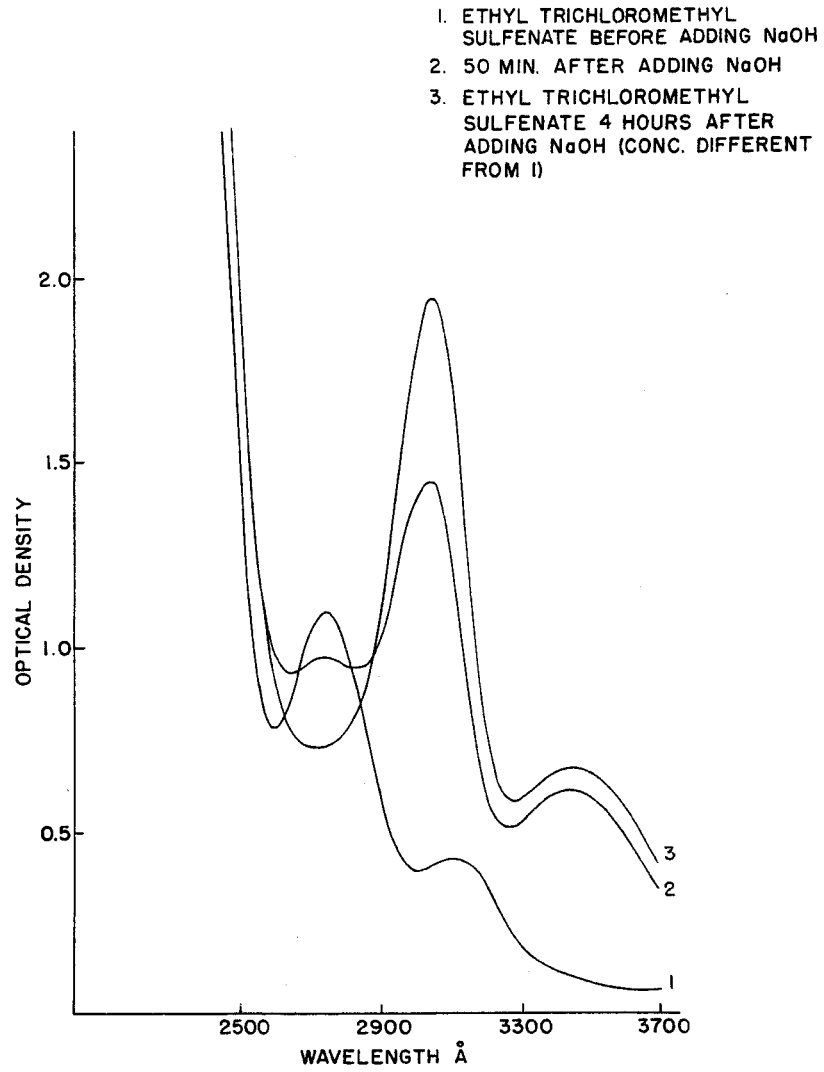
MU-6996

Figure 18

Figure 19

MU-5979

Figure 20

1. TRIMETHYLENE DISULFIDE + NaOH
   AFTER 10 MIN. PHOTOLYSIS
2. AFTER 25 MIN. PHOTOLYSIS
3. AFTER 105 MIN. PHOTOLYSIS
4.=3 AFTER ACIDIFICATION
5=4 WITH THIOL COMPENSATION

Figure 21