

## A Kinetic Study of the Acid Hydrolysis of a Bunte Salt

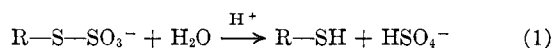
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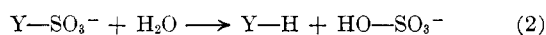
The acid hydrolysis of sodium *S*-ethyl thiosulfate to ethyl mercaptan has been examined kinetically in aqueous perchloric and hydrochloric acids. Although the mechanism is almost certainly given by equation 4, the rate ( $\log k_{\psi}$ ) shows an excellent correlation with  $H_0$ . This presumably occurs because  $\log k_{\psi}$  actually varies as  $w \log a_{\text{H}_2\text{O}} - H_-$ , and for values of  $w$  of about three this function changes with acid concentration in almost exactly the same way as  $H_0$ , at least over the range of acidities employed in the present study.

On heating in aqueous acid Bunte salts (*S*-alkyl thiosulfates) are hydrolyzed to the alkyl mercaptan and bisulfate ion (equation 1).<sup>1-3</sup> Since Bunte salts are



easily prepared by reaction of thiosulfate with alkyl halides, this hydrolysis offers a potentially valuable, but frequently overlooked, alternative to the more usual methods of effecting the transformation of an alkyl halide to a thiol.

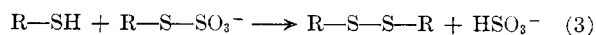
The Bunte salt hydrolysis is related to a number of other hydrolyses of various sulfate derivatives (equation 2), all of which involve substitution of hydroxyl for the group Y originally bonded to the sulfate sulfur.



Pertinent examples, all of which have been the object of at least some mechanistic study, are the hydrolysis of sulfamic acid<sup>4</sup> (equation 2, Y = NH<sub>2</sub>), aryl sulfates<sup>5</sup> (Y = OAr), alkyl sulfates<sup>6</sup> (Y = OR), and thiosulfate<sup>7</sup> (Y = S<sup>-</sup>). In the present study we have investigated, primarily by kinetic methods, the mechanism of the acid hydrolysis of a typical Bunte salt. It was our hope that this might contribute to further understanding of the general mechanism (or mechanisms) important for this type of substitution. As it turns out, the greatest interest centers in the observed dependence of rate on acidity; for here is another reaction, of a rather different type than previous examples,<sup>8</sup> where blind devotion to the Hammett-Zucker hypothesis would appear to lead one to mechanistically incorrect conclusions.

### Results

The specific system chosen for study was the hydrolysis of sodium *S*-ethyl thiosulfate in aqueous hydrochloric or perchloric acid. Several considerations prompted this selection. In particular, the low boiling point and limited water solubility of ethyl mercaptan made it possible, by passing a slow stream of nitrogen through the solution throughout the decomposition, to remove the thiol as rapidly as it was formed, thus preventing any significant incursion of the side reaction shown in equation 3.



- (1) T. S. Price and D. F. Twiss, *J. Chem. Soc.*, **95**, 1725 (1909).
- (2) Z. El-Heweki and E. Taeger, *J. prakt. Chem.*, [4] **7**, 191 (1958).
- (3) (a) H. E. Westlake and G. Dougherty, *J. Am. Chem. Soc.*, **63**, 658 (1941); (b) H. Z. Lecher and E. M. Hardy, *J. Org. Chem.*, **20**, 475 (1955); (c) U. Weiss and S. Sokol, *J. Am. Chem. Soc.*, **72**, 1687 (1950).
- (4) J. P. Candlin and R. G. Wilkins, *J. Chem. Soc.*, 4236 (1960).
- (5) G. N. Burkhardt, W. G. K. Ford, and E. Singleton, *ibid.*, 17 (1936).
- (6) R. L. Burwell, *J. Am. Chem. Soc.*, **74**, 1462 (1952).
- (7) W. A. Pryor, *ibid.*, **82**, 4794 (1960).
- (8) J. F. Bunnett, *ibid.*, **83**, 4956, 4968, 4973, 4978 (1961).

**Stoichiometry of the Reaction.**—To ensure that under our reaction conditions the Bunte salt disappeared almost exclusively by reaction 1, the stoichiometry of the decomposition was examined. The thiol, and any sulfur dioxide formed by equation 3 or other side reactions, were removed and estimated iodometrically by passing the nitrogen stream through an aliquot of standard iodine solution. The actual reaction solution itself was analyzed for sulfate content. The fraction of the Bunte salt consumed by reactions other than equation 1 can then be estimated from the difference between the iodine consumed and the sulfate formed (see Experimental). The results of two duplicate runs indicated that at least 96% of the Bunte salt was consumed by reaction 1 under our conditions. Further confirmation of this result was provided in one of the runs by actual isolation and gas chromatographic estimation of the diethyl disulfide produced by iodine oxidation of the mercaptan. The amount of disulfide formed was identical within the experimental error with that predicted from the yield of sulfate and the amount of iodine consumed.

**Kinetics of the Reaction.**—To follow the kinetics of the hydrolysis a slow stream of nitrogen was passed through the solution throughout the decomposition and thence through a chilled trap containing an aliquot of standard iodine solution. The time required for the iodine color to be discharged was noted, another aliquot of iodine added, and the process repeated as many times as desired. Finally, an excess of iodine was added, and an "infinity" value for the consumption of iodine determined.<sup>9</sup>

Fig. 1 shows that the hydrolysis, as might be expected, is cleanly first order in Bunte salt. Similarly good first-order kinetics were observed in all of the kinetic experiments.

**Dependence of Rate on Acid Concentration.**—The dependence of the hydrolysis rate on acid concentration was studied both with hydrochloric and perchloric acids as the strong acid and at both 63.3 and 77.1°. The results are shown in Table I. The rate is much more strongly dependent on acid concentration than any simple proportionality between  $k_{\psi}$  and (H<sub>3</sub>O<sup>+</sup>) would predict. For example, at 63° doubling the hydrochloric acid concentration from 2.88 to 5.66 *N* leads to greater than a tenfold increase in rate. On the other hand, as Fig. 2 shows, the dependence of rate on acid seems in all cases to be nicely correlated by the Hammett acidity function,  $H_0$ , plots of  $\log k_{\psi}$  vs.  $-H_0$  being linear and showing slopes ranging from 0.99 to 1.04.

(9) By avoiding reaction conditions with very short half-lives there is no difficulty in passing nitrogen through the solution at a sufficiently rapid rate that the time between evolution of the thiol and its absorption in the iodine is short enough not to influence the experimentally determined rate constant.

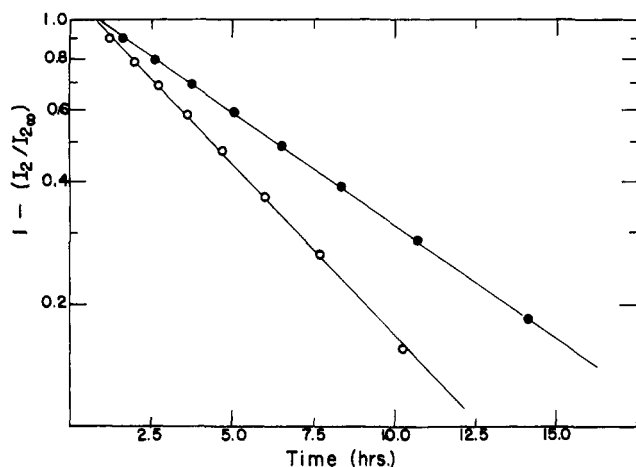


Fig. 1.—Rate of hydrolysis of sodium *S*-ethyl thiosulfate as measured by rate of liberation of ethyl mercaptan. Both runs at 63.3° and  $[\text{EtSSO}_3^-]_0$ , 0.10 *M*: ○,  $\text{HClO}_4$ , 4.68 *M*; ●,  $\text{HCl}$ , 4.89 *M*.

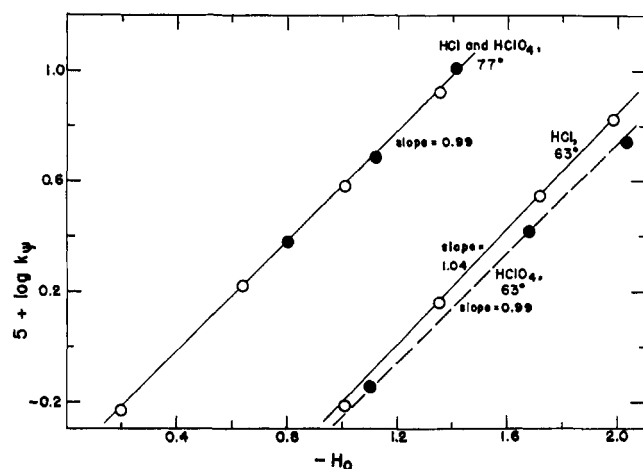


Fig. 2.—Dependence of rate on strong acid concentration:  $\log k_p$  vs.  $-H_0$  for runs in Table I: ○,  $\text{HCl}$  runs; ●,  $\text{HClO}_4$  runs.

TABLE I

KINETICS OF THE HYDROLYSIS OF <i>S</i> -ETHYL THIOSULFATE			
Temp., °C.	$(\text{EtSSO}_3^-)_0$ , <i>M</i>	Acid, concd., <i>M</i>	$k_p \times 10^5$ , sec. <sup>-1</sup>
63.3	0.10	$\text{HCl}$ , 2.88	0.61
		$\text{HCl}$ , 3.81	1.45
		$\text{HCl}$ , 4.89	3.53
		$\text{HCl}$ , 5.66	6.8
63.3	.10	$\text{HClO}_4$ , 2.66	0.72
		$\text{HClO}_4$ , 3.94	2.64
		$\text{HClO}_4$ , 4.68	5.52
		$\text{HCl}$ , 1.00	0.59
77.1	.10	$\text{HCl}$ , 1.92	1.66
		$\text{HCl}$ , 2.88	3.80
		$\text{HCl}$ , 3.81	8.3
		$\text{HCl}$ , 3.81	8.4
		$\text{HClO}_4$ , 2.02	2.38
77.1	.10	$\text{HClO}_4$ , 2.75	4.85
		$\text{HClO}_4$ , 3.35	10.3

**Activation Parameters.**—In systems of this sort the usual activation parameters,  $\Delta H^*$  and  $\Delta S^*$ , must be calculated at some standard state for proton activity, either 1 *M* acid or  $h_0 = 1$ . Despite the excellent linearity of the  $\log k$  vs.  $-H_0$  plot, it is questionable (see Discussion) that this latter standard state is the more appropriate. Consequently  $\Delta S^*$  was calculated for

both cases. The results are:  $\Delta H^* = 31.0 \pm 0.7$  kcal.;  $\Delta S^* = 5.8 \pm 2.0$  e.u. (for 1 *M* acid as standard state),  $4.9 \pm 2.0$  e.u. (for  $h_0 = 1$ ).

**Probable  $pK_a$  of *S*-Ethyl Thiosulfuric Acid ( $\text{EtSSO}_3\text{H}$ ).**—To interpret intelligently the mechanistic significance of the variation of rate with acid concentration one must have at least a rough idea of the  $pK_a$  of the conjugate acid of the Bunte salt. *A priori* it was our opinion this should be a very strong acid; but some data in the literature<sup>10</sup> raised doubts on this point. Thus Page<sup>10</sup> claimed to have measured the first dissociation constant of thiosulfuric acid and reported it to be  $K_1 \cong 0.25$ . If this is true, then, in view of the electron-releasing inductive effect of ethyl relative to hydrogen  $K_a$  for  $\text{EtSSO}_3\text{H}$  should certainly be no larger than 0.25. On that basis, in all the solutions in Table I the Bunte salt should be present at equilibrium *almost entirely in the form of its conjugate acid*. In that event, the only way to explain the pronounced effect of increasing acid, which persists even at high acidities, would be by assuming slight further protonation of  $\text{EtSSO}_3\text{H}$  to such a species as  $\text{EtSSO}_3\text{H}^+$ . Since the  $H_0$

function would be expected to correlate the variation of  $[\text{EtSSO}_3\text{H}]/[\text{EtSSO}_3\text{H}]$  with acid concentration, this possibility has to be considered quite seriously in view of the present results.

However, such a hypothesis has the necessary corollary that at lower acidities than those in Table I the rate should decrease *more rapidly* with decreasing  $h_0$  than would be predicted by extrapolation from runs at acid concentrations 1 *M* or greater. This is because at these lower acidities the Bunte salt at equilibrium would no longer exist entirely as its conjugate acid, and changes in acid concentration would now influence not only the equilibrium  $\text{EtSSO}_3\text{H} + \text{H}^+ \rightleftharpoons \text{EtSHSO}_3\text{H}^+$  but also the equilibrium  $\text{EtSSO}_3^- + \text{H}^+ \rightleftharpoons \text{EtSSO}_3\text{H}$ .

A series of runs (Table II) was therefore made at

TABLE II

RATE OF HYDROLYSIS OF <i>S</i> -ETHYL THIOSULFATE AT 96°			
$(\text{EtSSO}_3^-)_0$ , <i>M</i>	Acid, concd., <i>M</i>	$k_p \times 10^5$ , sec. <sup>-1</sup>	$-H_0$
0.050	$\text{HCl}$ , 2.00	16.0	0.69
	$\text{HCl}$ , 1.00	5.7	.20
	$\text{HCl}$ , 0.10	0.61	(-0.98) <sup>a</sup>
.025	$\text{HCl}$ , .10	.57	
	$\text{HCl}$ , .10	(0.53) <sup>b</sup> obsd.	
(0.00)	$\text{HCl}$ , .10	(0.50) <sup>c</sup> calcd.	

<sup>a</sup> At  $\mu = 0.10$ . <sup>b</sup> Extrapolated to  $\mu = 0.10$  from runs at 0.05 *M* and 0.025 *M*. <sup>c</sup> Calcd. by extrapolation from data at  $\text{HCl} = 2.0$  *M* and 1.0 *M*.

96°, at which temperature the hydrolysis is sufficiently rapid to be measurable even at  $(\text{H}^+)$  as low as 0.1 *M*. Comparison of the observed rate constant for 0.1 *M* hydrochloric acid with that calculated by extrapolation from the data at much higher  $h_0$  values shows that the required increased dependence on acidity is *not* observed. For this to be the case either (1)  $K_a$  for  $\text{EtSSO}_3\text{H}$  must be less than 0.025, or (2)  $\text{EtSSO}_3\text{H}$  must be a very strong acid. The first alternative seems both



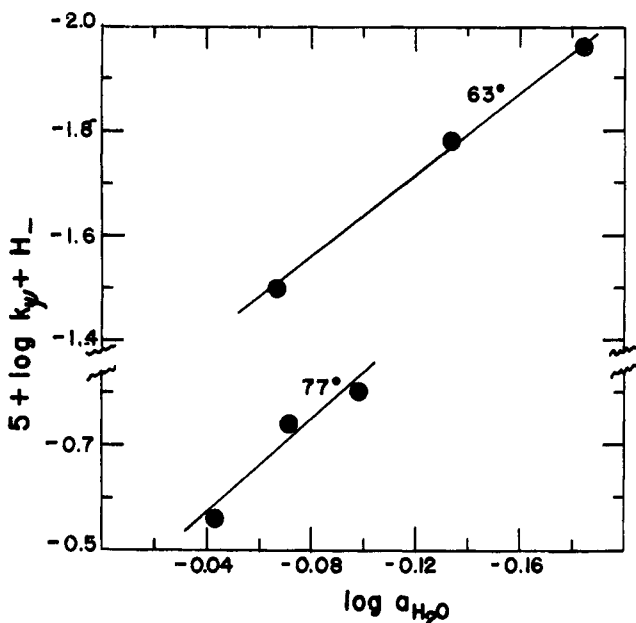


Fig. 3.—Perchloric acid rate data of Table I plotted as  $\log k_{\psi} + H_-$  vs.  $\log a_{H_2O}$ : slope at  $63^\circ$ , +3.9; at  $77^\circ$ , +4.4.

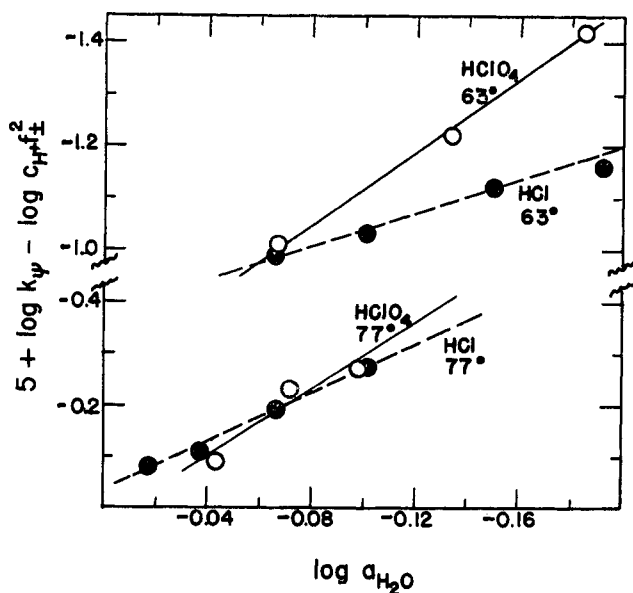


Fig. 4.—Rate data of Table I plotted as  $\log k_{\psi} - \log c_{H+f_{\pm}^2}$  vs.  $\log a_{H_2O}$  (see text):  $\circ$ ,  $HClO_4$  data: slope at  $63^\circ$ , +3.5; at  $77^\circ$ , +3.3.  $\bullet$ ,  $HCl$  data: slope at  $63^\circ$ , +1.7; at  $77^\circ$ , +2.3.

Exactly analogous reasoning predicts for mechanism 4

$$\log k_{\psi} + H_- = w \log a_{H_2O} + \text{constant}$$

with  $w$  again between +1.25 and +7, its exact value dependent on the role of water. Fig. 3 shows a plot of  $\log k_{\psi} + H_-$  (Boyd<sup>14</sup>) vs.  $\log a_{H_2O}$  for the perchloric acid runs, and Fig. 4 a plot of  $\log k_{\psi} - \log c_{H+f_{\pm}^2}$  vs.  $\log a_{H_2O}$  for both the perchloric and hydrochloric acid runs. The plots are satisfactorily linear and the slopes ( $w$  values) are within the range expected.

Because of the uncertainty regarding the exact form of the  $H_-$  function most appropriate for the  $EtSSO_3H-EtSSO_3^-$  equilibrium it would be presumptuous to draw any conclusion about the exact role of water in the rate-determining step of the Bunte salt hydrolysis. However, the above analysis demonstrates that the depend-

ence of rate on acid concentration is definitely consistent with that expected for mechanism 4. Moreover, it also shows that in such a reaction of the conjugate acid of a negatively charged base the combined dependence of rate on  $H_-$  and  $\log a_{H_2O}$  can easily conspire to produce a variation of  $\log k_{\psi}$  with acid concentration exactly paralleling the change in  $H_0$ . The striking constancy of  $\log k_{\psi} + H_0$  has in this case no mechanistic significance.

The present results thus illustrate, with a new type of example, the continual need for caution and common sense, rather than unswerving allegiance to some simple formula, in drawing mechanistic conclusions from kinetic response to changing acid concentration in concentrated aqueous acid.

Finally, a few words about the entropy of activation. For mechanism 4,  $\Delta S^*$  will be equal to  $\Delta S_{4a}^* + \Delta S_{4b}^*$ . The first of these is simply  $-\Delta S_{\text{ionization}}^*$  for the conjugate acid of the Bunte salt. Entropies of ionization are known for a considerable number of monofunctional uncharged acids<sup>17</sup> and lie in the range  $-16$  to  $-24$  e.u. We would, therefore, expect  $\Delta S_{4a}^*$  to have a value of  $+20 \pm 4$  e.u. Since  $\Delta S^*$  is about 6 e.u., this requires  $\Delta S_{4b}^*$  to be  $-14 \pm 4$  e.u. This substantial negative value is in agreement with what we would expect for a reaction such as 4b,<sup>8</sup> where formation of the transition state from  $EtSSO_3H$  requires one or more additional water molecules.

## Experimental

**Preparation and Purification of Materials.**—Sodium *S*-ethyl thiosulfate was prepared by reaction of equimolar quantities of ethyl bromide and sodium thiosulfate in 20% ethanol. The mixture was stirred at room temperature until all the alkyl halide dissolved and then allowed to stand for 12 hr. The solvent was removed under reduced pressure, keeping the bath temperature below  $45^\circ$ , and the residue extracted with hot alcohol. The filtrate was cooled in ice, and the Bunte salt was filtered off and recrystallized two additional times from alcohol. The material so obtained was dried under oil pump vacuum at room temperature. When purified in this way the Bunte salt contains ethanol of crystallization and has approximately the formula  $C_2H_5SSO_3 \cdot Na \cdot \frac{1}{2}C_2H_5OH$ . This can be removed by dissolving the salt in water, evaporating to dryness under reduced pressure, and prolonged drying under high vacuum at room temperature. The compound was stored at  $-20^\circ$  until used. Diethyl disulfide (Eastman Kodak Co.) was fractionally distilled; the fraction boiling  $151-152^\circ$  was retained. The hydrochloric and perchloric acids used to prepare the acid solutions were the best commercial grades.

**Stoichiometry of the Reaction.**—The decomposition was carried out in the same set-up used previously to follow the kinetics of the thermal decomposition of thiosulfonates.<sup>18</sup> A 0.1 *M* solution of the Bunte salt in 5.6 *N* hydrochloric acid was placed in flask A, the flask cooled in ice, and a slow stream of nitrogen passed through the solution for at least 1 hr. An excess of ethanolic standard iodine solution was placed in trap C. The reaction flask was then placed in the constant temperature bath, warm water ( $\sim 50^\circ$ ) was circulated through the condenser of flask A, and the slow passage of nitrogen was continued.

After the decomposition of the Bunte salt was complete the excess iodine in C was titrated with thiosulfate and the contents of the reaction flask were analyzed for sulfate.

Under these conditions the Bunte salt presumably is consumed by reactions 1, 3, and 7. In the strongly acidic medium any bisulfite formed will be converted to sulfur dioxide and presum-



(17) Ref. 16, p. 667.

(18) J. L. Kice, F. M. Parham, and R. M. Simons, *J. Am. Chem. Soc.*, **82**, 834 (1960).

ably is carried over by the nitrogen stream. That being the case, the milliequivalents of iodine consumed will be

$$\text{meq. I}_2 = \text{mmoles RSH} + 2x(\text{mmoles HSO}_3^-)$$

Assuming that all EtSOH formed in equation 7 reacts by 8,<sup>19</sup> the amount of sulfate formed in the original solution will be

$$\text{mmoles SO}_4^{-2} = \text{mmoles RSH} + \text{mmoles HSO}_3^-$$

From this the fraction of the Bunte salt disappearing by reactions 3 and 7,  $\alpha$ , is

$$\alpha = \frac{\text{mmoles HSO}_3^-}{\text{mmoles SO}_4^{-2}} = \frac{\text{meq. I}_2 - \text{mmoles SO}_4^{-2}}{\text{mmoles SO}_4^{-2}}$$

Two runs gave the following results for mmoles  $\text{SO}_4^{-2}$ , meq.  $\text{I}_2$ ,  $\alpha$ : (1) 5.36, 5.65, 0.05; (2) 2.68, 2.76, 0.03.

In run 1, instead of ethanolic iodine, aqueous iodine overlaid with a small amount of pentane was used in trap C.

(19) Alternate assumptions such as disproportionation ( $2\text{EtSOH} \rightarrow \text{EtSH} + \text{EtSO}_2\text{H}$ ) lead to smaller values for  $\alpha$ , the fraction of the Bunte salt not reacting by equation 1.

After titration of the excess iodine the pentane solution was diluted to known volume and its diethyl disulfide content determined by v.p.c. comparison with standard solutions of the disulfide in pentane. The amount of disulfide formed should be

$$\text{mmoles R}_2\text{S}_2 = \frac{\text{moles RSH}}{2} = \frac{2 \text{ mmoles SO}_4^{-2} - \text{meq. I}_2}{2}$$

The value calculated in this way is 2.53 mmoles. The amount actually found in the pentane solution was 2.48 mmoles.

**Procedure for Kinetic Runs.**—The same apparatus and procedure was used for the kinetic runs, except that, instead of an excess of iodine, only a small aliquot of iodine was added initially. The time required for the iodine color to fade was noted, another aliquot was added, and the procedure was repeated as many times as desired. Finally an excess of standard iodine was added, and after a suitable length of time, the "infinity" consumption of iodine was determined. This is the same type of procedure used before to follow the thioisulfonate decompositions.<sup>18</sup>

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## The Metalation and Alkylation of Mercaptals

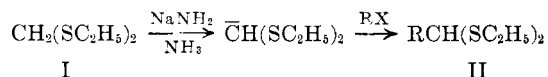
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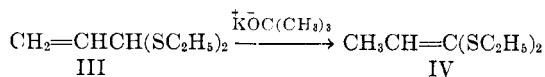
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The formation of a mercaptal anion and its reaction with various alkylating and acylating agents was studied. The conversion of the alkylated and acylated products into carbonyl compounds can serve as a source of mono-ketones,  $\alpha$ -diketones, acylions,  $\alpha$ -keto esters, and  $\beta$ -keto esters. A new desulfurizing system is discussed.

For many years it was generally believed that mercaptals, like their oxygen analogs, were stable to alkali.<sup>1a-c</sup> Recent reports, however, have shown that with a sufficiently powerful base the hydrogen is labile. Arens<sup>2</sup> and co-workers treated formaldehyde diethylmercaptal (I) with sodium amide in liquid am-



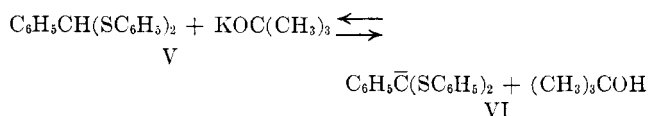
monia followed by an alkyl halide and obtained II. The 3d-orbital overlap between a sulfur atom and the adjacent carbanion and the inductive effect of the two mercapto groups were put forth as explanations for the stability of the carbanion. Furthermore, Oae, Ohno, and Tagaki<sup>3</sup> have shown that acrolein diethylmercaptal (III) was isomerized by potassium *t*-butoxide to the ketene mercaptal IV.



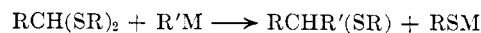
The report<sup>2</sup> that a mercaptal reacts with a strong base to give an anion has opened up some interesting areas of mercaptal chemistry. Besides the inherent interest in metalation reactions, reaction of the anion with alkylating and acylating agents provides a synthetic route to a wide variety of functional groups, *e.g.*, hydrolysis of the products leads to carbonyl compounds, whereas

desulfurization reduces the mercaptole function to a methylene group.

We have studied the metalation of benzaldehyde diphenylmercaptal (V) with various bases and wish to report our results. Alkali metal amides in liquid ammonia proved unsatisfactory for metalation in this system because of undesirable side reactions. Potassium *t*-butoxide in *t*-butyl alcohol apparently did not form the anion, as starting material was recovered after treatment with an alkyl halide. Potassium *t*-butoxide in anhydrous tetrahydrofuran gave a low yield (25%) of alkylated product; the low yield can probably be accounted for by the unfavorable alcohol-mercaptal equilibrium. Organometallic bases such as Grignard



reagents, phenyllithium, or *n*-butyllithium are unsatisfactory metalating agents because they have been reported to cleave mercaptals.<sup>4,5</sup>



Sodium hydride dispersion in ether, tetrahydrofuran or benzene gave very little evolution of hydrogen when treated with benzaldehyde diphenylmercaptal (V). Sodium hydride dispersion in the dimethyl ether of ethylene glycol or the dimethyl ether of diethylene glycol required a high reaction temperature before the evolution of hydrogen began and led to extensive decomposition. However, an excellent solvent for such

(1) (a) O. Behagel and H. Ratz, *Ber.*, **72**, 1257 (1939); (b) D. S. Tarbell and D. P. Harnish, *Chem. Rev.*, **49**, 67 (1951); (c) E. E. Reid, "Organic Chemistry of Bivalent Sulfur," Vol. III, Chemical Publishing Co., Inc., New York, N. Y., 1960, p. 331.

(2) J. F. Arens, M. Froling, and A. Froling, *Rec. trav. chim.*, **78**, 663 (1959).

(3) S. Oae, A. Ohno, and W. Tagaki, *Chem. Ind. (London)*, **7**, 304 (1962).

(4) E. E. Reid, "Organic Chemistry of Bivalent Sulfur," Vol. III, Chemical Publishing Co., Inc., New York, N. Y., 1960, p. 332.

(5) A. Schonberg, *et al.*, *Ber.*, **64**, 2577 (1931).