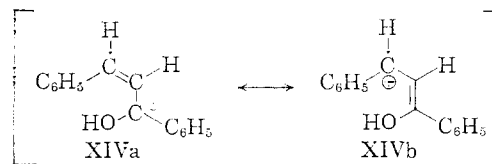


in mechanism for II, when compared with I, III and IV.

Finally, we wish to point out that these data also provide insight into the configurational stability of allylic cations.<sup>37</sup> Among the canonical set of resonance forms which may be written for the salt of *cis*-chalcone are an "allylic pair" (XIVa and XIVb). The conclusions regarding the mechanism of isomerization of I clearly demand configurational stability in this system, since the isomerization is effected only by the attachment of water (eq. 2).

(37) W. G. Young, S. H. Sharman and S. Winstein, *J. Am. Chem. Soc.*, **82**, 1376 (1960).

With sufficient stabilization for a positive charge provided at the ends of the "allylic" system, the



bond order of the  $C_{\alpha}$ - $C_{\beta}$  bond is reduced. The isomerization thus may become unimolecular, as observed in these studies for *cis*-4-methoxy-chalcone (II).

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY OF TEXAS CHRISTIAN UNIVERSITY, FORT WORTH, TEX.]

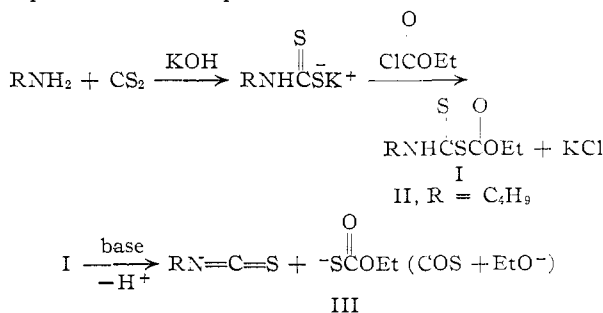
## The Modified Kaluza Synthesis. II.<sup>1</sup> Kinetics and Mechanism

By JOE E. HODGKINS, W. PRESTON REEVES AND YAO-TING GORDON LIU

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The decomposition of intermediate carboethoxy *n*-butyldithiocarbamates is shown to be subject to base catalysis and the evolution of carbon oxysulfide to proceed in three steps. The reaction is unusual in that the rate first increases with increasing concentration of buffer components, then levels and shows no further increase in rate. This behavior as well as the mechanism of the decomposition is discussed.

**Introduction.**—The modified Kaluza synthesis of isothiocyanates, recently introduced,<sup>1</sup> has been applied to the preparation of various aliphatic isothiocyanates from the corresponding amines with generally successful results. The method has been employed in the synthesis of alkyl,<sup>1</sup> alkenyl,<sup>2</sup> benzyl,<sup>3</sup> *o*-, *m*- and *p*-methoxy benzyl,<sup>4,5</sup> 3-hydroxypropyl,<sup>6</sup> various substituted  $\beta$ -hydroxyethyl<sup>7</sup> and  $\omega$ -carbalkoxyalkyl<sup>8</sup> isothiocyanates. The modification of the Kaluza method consists in utilizing base-catalyzed decomposition of intermediate carboethoxy dithiocarbamates, derived from amines by treatment with carbon disulfide and base followed by carboethoxylation with ethyl chlorocarbonate. The reaction can be carried out in aqueous<sup>1</sup> or non-aqueous<sup>1,7,8</sup> solvents.



(1) J. E. Hodgkins and M. G. Ettlinger, *J. Org. Chem.*, **21**, 404 (1956), should be considered the first paper of this series.

(2) M. G. Ettlinger and J. E. Hodgkins, *J. Am. Chem. Soc.*, **77**, 1833 (1955).

(3) M. G. Ettlinger and J. E. Hodgkins, *J. Org. Chem.*, **21**, 204 (1956).

(4) M. G. Ettlinger and A. J. Lundeen, *J. Am. Chem. Soc.*, **78**, 1952 (1956).

(5) A. Kjaer and R. B. Jensen, *Acta Chem. Scand.*, **10**, 141 (1956).

(6) A. Kjaer and R. Gemlin, *ibid.*, **10**, 1193 (1956).

(7) A. Kjaer and R. Gemlin, *ibid.*, **11**, 906 (1957); A. Kjaer, R. Gemlin and R. B. Jensen, *ibid.*, **10**, 432 (1956).

(8) D. L. Garmaise, R. Schwartz and A. F. McKay, *J. Am. Chem. Soc.*, **80**, 3332 (1958).

In the first paper<sup>1</sup> of this series a path was proposed for the base-catalyzed decomposition reaction based on the general chemical behavior of the carboethoxy dithiocarbamate derivatives (I). No evidence was available, however, to ascertain the type of base catalysis or whether the evolution of carbon oxysulfide was concerted with proton removal or preceded by fission to thiocarbonate ion<sup>9</sup> followed by rapid decomposition of this species. We wish now to present evidence that shows the decomposition to be subject to an unusual type of base catalysis and the evolution of carbon oxysulfide to proceed in two steps after proton removal.

**Results.**—The rate of decomposition of carboethoxy *n*-butyldithiocarbamate (II) was followed by measurement of the evolution of carbon oxysulfide and by ultraviolet light absorption measurement of undecomposed II. Gas evolution kinetics were employed only where qualitative comparisons were needed since the rate constants were invariably 20–30% lower than those measured spectrophotometrically, due, presumably, to the solubility of carbon oxysulfide in the solvent. By employing good gas-liquid equilibration techniques (see Experimental), reproducible kinetic results could be obtained from gas evolution measurements, but deviation from first-order kinetics occurred above the half-life and the results indicated only 50–70% completions. In solvents containing water and added salts, which reduces the solubility of carbon oxysulfide, the evolution kinetics closely approached the spectrophotometric measurements and the extent of completions increased to about 90%.

The uncatalyzed rate of decomposition of II at 41° in 83% (by volume) methanol or ethanol is barely detectable. The compound II is perfectly stable to 0.005 *N* HCl in the same solvents.

(9) O. E. Schultz and E. Barthold, *Arch. Pharm.*, **285**, 267 (1952), have isolated sodium *O*-ethyl thiocarbonate in the decomposition of these intermediates in the presence of strong sodium hydroxide.

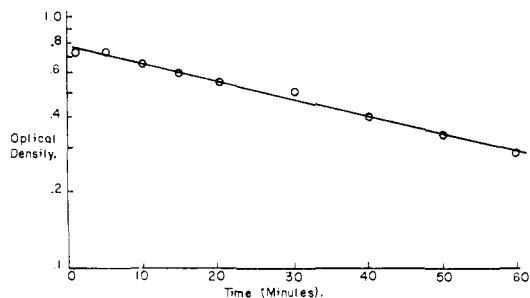


Fig. 1.—Pseudo-first-order plot of the base-catalyzed decomposition of II in "83%" ethanol.

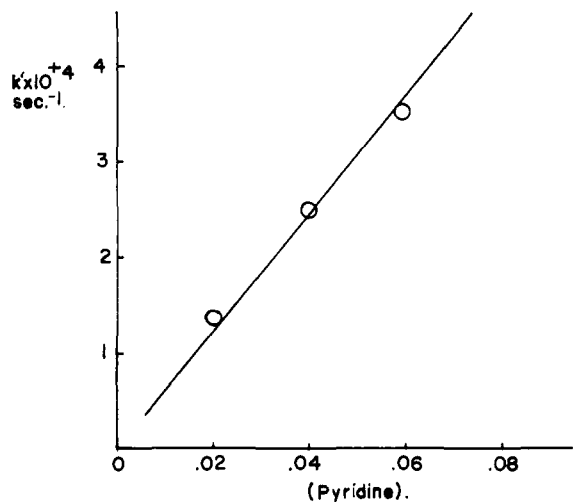


Fig. 2.—Variation of rate constant ( $k'$ ) with pyridine concentration.

The base-catalyzed decomposition of II is pseudo first-order (Fig. 1), and the rate constant increases linearly with initial concentration of added pyridine or pyridine derivatives (Fig. 2). This linear dependence is shown in methanol-water, ethanol-water, acetone-water and in the pure alcohols. The decomposition is faster in solvents of higher dielectric strength.

The decomposition shows a positive salt effect with lithium chloride and with lithium perchlorate (Fig. 3), but the increase in rate with added salt is small (65%). The reaction is subject to catalysis by all bases in solution. This was indicated by observing a regular increase in rate (however; see Discussion) on addition of increasing amounts of pyridine and pyridine hydrochloride in the same ratio (at such a high lithium perchlorate or lithium chloride concentration to minimize changes in ionic strength; Fig. 4). The effect of lyate ion was proved to be negligible in the range of concentrations shown in Fig. 4 by the fact that the buffer curves passed through the origin.

The gas evolution measurements showed a change in rate law from first order to zero order at concentrations of pyridine approaching 1 *N*. Since this suggests a slow secondary reaction involving the loss of COS, the decomposition of potassium O-ethyl thiocarbonate (III, "Bender's salt") was studied briefly in hopes of understanding this behavior. In aqueous solution the decomposition of

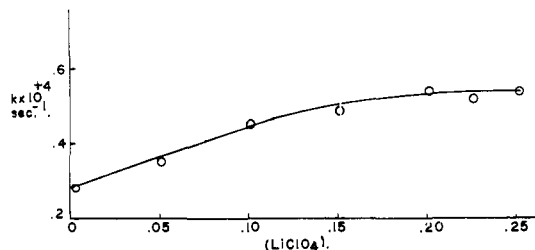


Fig. 3.—Variation of rate constant ( $k'$ ) with lithium perchlorate concentration.

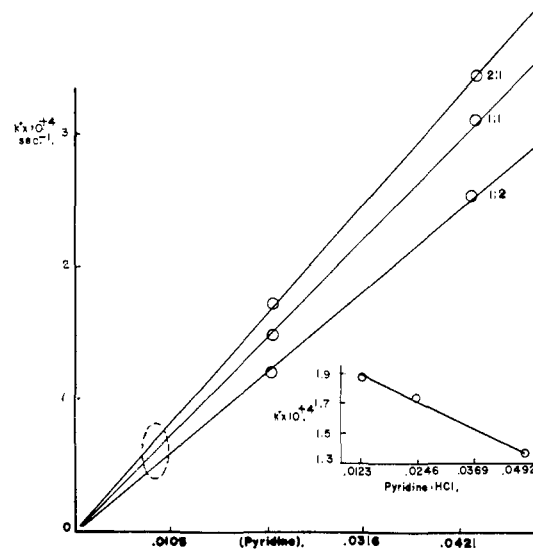


Fig. 4.—Plot of  $k'$  vs. concentration of pyridine-pyridine-HCl. The concentration of pyridine-HCl is given by the ratio term at the top of each curve. The dashed area represents the unusually high experimental error at low concentrations of pyridine-pyridine-HCl. Variation of  $k'$  with pyridine-HCl is shown in the smaller plot.

III in the presence of various buffers is subject to acid catalysis (Fig. 5). Thus, at low base concentrations the thiocarbonate ion produced in the decomposition of III is rapidly decomposed since the

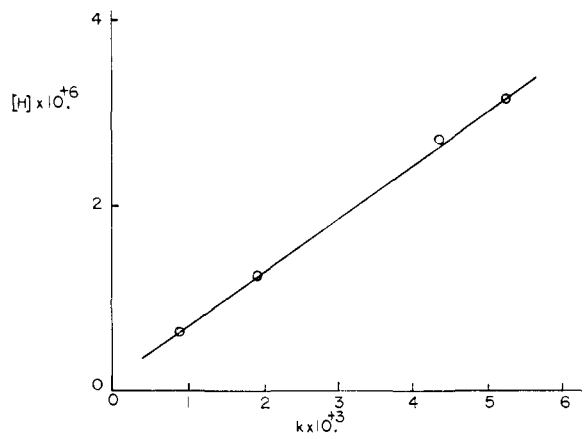


Fig. 5.—Effect of  $[H]^+$  on rate of decomposition of O-ethyl thiocarbonate ion.

concentration of acids in solution is relatively high; but as the concentration of base is increased, the decomposition of III slows and becomes independ-

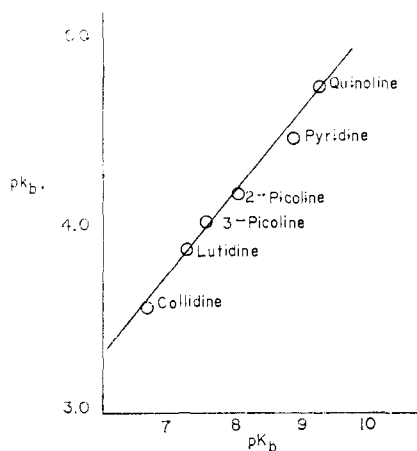
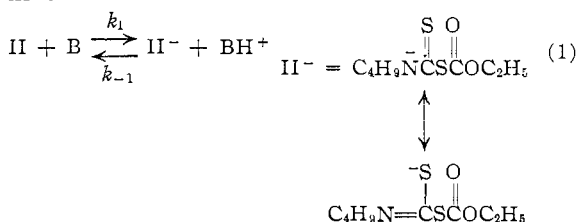


Fig. 6.—“Brønsted plot” of catalytic constants and base strength. The catalytic constants were measured in methanol by evolution kinetics and the  $pK_b$ 's were measured in water.

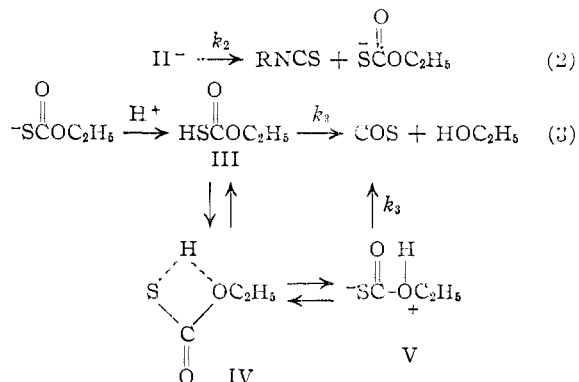
ent of II concentration. Finally, when the concentration of acid species is reduced to a sufficiently low value, III doesn't decompose measurably. In water solution III does not decompose measurably (under conditions described here) at a  $pH$  value greater than 7, and decomposes rapidly at  $pH$  below 6, suggesting the loss of COS *via* a strongly  $pH$  dependent equilibrium reaction. Evidently this  $pH$  dependence is the reason that in strong base the products of decomposition of II are *n*-butyl mustard oil and O-ethyl thiocarbonate salt.<sup>9</sup> A regular kinetic experiment containing 0.03 *M* pyridine in the presence of II, which should be about as acidic as phenol,<sup>1</sup> should have sufficient concentration of acids to rapidly decompose thiocarbonate ion. This is shown by fair agreement of rate constants measured by the evolution method and by the spectrophotometric method, since the rate constants obtained by measurements of evolution of carbon oxysulfide must depend on fast decomposition of thiocarbonate ion.

The variation of catalytic constants with base strength of pyridine derivatives is shown in Fig. 6.

**Discussion.**—A mechanism consistent with the kinetic data is



Equation 1 is indicated by the positive salt effect. The reversibility of eq. 1 is suggested since the addition of the conjugate acid to the base decreases the reaction rate (Fig. 4; the effect of lyate ion is considered below). Equation 2 follows from the reversibility of eq. 1 and from the base catalysis. Equation 2 was shown in this work not to be reversible by showing that *n*-butyl isothiocyanate and thiocarbonate ion do not combine to form II under any circumstances. Equation 3 is reversible in basic media.



By combining eq. 1 and 2 and assuming steady state concentration in  $\text{II}^-$  ion, the following rate expression can be derived.

$$\frac{-d[\text{II}]}{dt} = \frac{k_1 k_2 [\text{II}] [\text{B}]}{k_{-1} [\text{BH}]^+ + k_2} \quad (4)$$

Since the concentrations of  $[\text{B}]$  and  $[\text{BH}]^+$  may be considered as constant within a given reaction, eq. 4 may be integrated to give a first order expression,  $-\ln [\text{II}] = k't + C$ , where

$$k' = \frac{k_2 k_1 [\text{B}]}{k_{-1} [\text{BH}]^+ + k_2} \quad (5)$$

Plots of the specific rate constant  $k'$  against the initial concentrations of  $[\text{B}]$  and  $[\text{BH}]^+$  show the expected behavior in low concentration ranges (Figs. 2 and 4).

The above equations apply directly to the kinetic results obtained by the spectrophotometric method since the rate of cleavage of II is measured; but in order for these equations to apply to the gas evolution kinetic results, it is necessary to conduct the experiments at low basicities where  $k_3$  of eq. 3 is much greater than  $k_2$ .

Equation 5 shows an interesting relationship between possible general and specific base catalysis with this particular mechanism. If the  $k_{-1} [\text{BH}]^+$  term is large compared to  $k_2$ , then by neglecting this term, the equation assumes the form of a typical specific lyate ion catalytic expression (the substitution of  $[\text{OR}]^- = K_b([\text{B}]/[\text{BH}]^+)$ , where R is alkyl or H, in eq. 4 shows the lyate term). If  $k_2$  is large compared to  $k_{-1} [\text{BH}]^+$ , the equation assumes the form of a general base catalytic expression. This behavior was also indicated experimentally (Fig. 7). Thus, at concentrations of buffer components above 0.1 *M* (in always the same ratio) the rate levels to show specific lyate ion catalytic behavior while below this concentration general base catalytic behavior is shown. It is doubtful, however, that the effect of lyate ion is ever observed in the decomposition of II since the concentration of this species is exceedingly low. The concentration of lyate ion in a typical kinetic run employed in this work must range from  $10^{-7}$  to  $10^{-9}$  m./l. (calcd.) and since the lyate ion catalytic constant is probably about  $10^{-2}$  to  $10^{-1}$  (obtained by extrapolation of the “Brønsted plot” Fig. 6), the contribution of lyate ion to the decomposition should be immeasurable. The effect of lyate ion was proved to be negligible by showing the intercepts of the buffer curves, Fig. 4, passed through the origin.

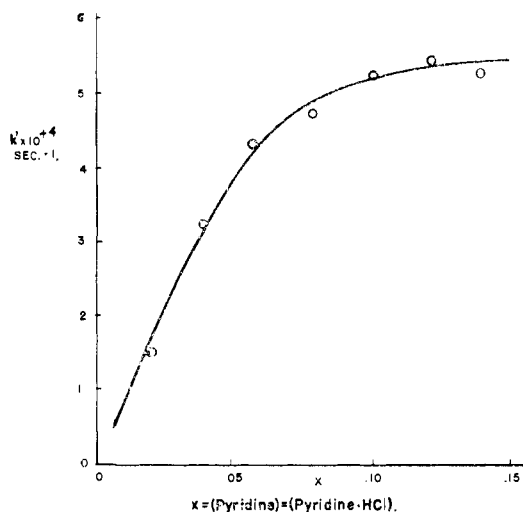


Fig. 7.—Extended plot of rate vs. pyridine-pyridine-HCl concentration in the same ratio.

The data shown by Fig. 7 were further evaluated; eq. 6 can be derived from eq. 5 which is suggested as the equation of the curve shown in Fig. 8.

$$\frac{1}{k'} = \frac{1}{k_1[B]} + \frac{k_{-1}[BH]^+}{k_2k_1[B]} \quad (6)$$

From eq. 6 a plot of  $1/k'$  vs.  $1/[B]$  should be a straight line if the rate equation fits the experimental data. This plot (Fig. 8) verifies the mechanism and indicates that a mechanism of this type can indeed show unchanging rate in certain ranges of concentrations with increasing base concentration and show increasing rate in other concentration ranges.

The form of eq. 5 is similar to the form of the pseudo first-order constant derived for the bromination of acetone<sup>10</sup> except that the second step represents an internal bond-breaking process rather than reaction with another molecule in solution. The equation appears general for acid- and base-catalyzed reactions, but the behavior represented by Figs. 7 and 8 would be observed only if the magnitudes  $k_{-1}[BH]^+$  (or  $k_{-1}[A]^-$  in the case of acid catalysis) and  $k_2$  (the rate of rearrangement or fission) are such that  $k_2$  is larger than  $k_{-1}[BH]^+$  at low conjugate acid concentrations and  $k_{-1}[BH]^+$  is larger at high conjugate acid concentrations. If either term is large compared to the other throughout the concentration range employed, then only general base or specific lyate ion catalysis will be observed depending on which term is the largest, even though the actual effect of lyate ion may be immeasurable. Thus, a paradox arises when applying the kinetically defined "general" or "specific lyate ion" terms to the mechanism presented above.

The fact that thiocarbonate ion V is decomposed rapidly by hydronium ion gives explanation for the retention of COS in high concentration of base or in the presence of strong base,<sup>9</sup> and gives further support to the non-concerted evolution of COS—even in weakly basic solution.<sup>11</sup> These results

(10) A. Lapworth, *J. Chem. Soc.*, **85**, 30 (1904); L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1930, p. 96.

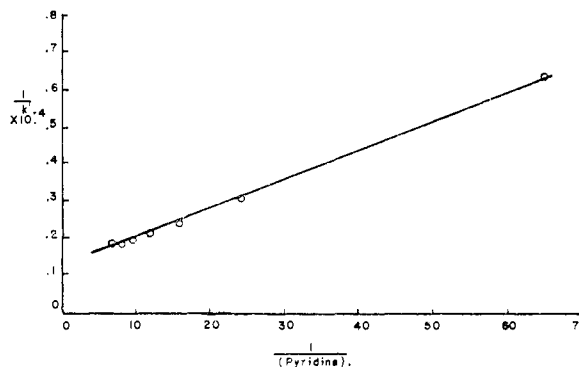


Fig. 8.—Plot of  $1/k'$  vs. reciprocal of the pyridine concentration.

agree with the decomposition of O-ethyl xanthate ion which is subject to specific hydronium ion catalysis.<sup>12</sup> The position of the proton on the acidic atoms<sup>12</sup> in the decomposition of any of these thiocarbonate derivatives is, however, inconsistent with current organic mechanism theory since in this position the proton would tend to localize the electrons on the acid atoms and inhibit the shift of the electrons to the ester oxygen; thus providing no driving force for the reaction. If a simple rearrangement of the acid III is postulated to the corresponding structure IV (or structure V), a driving force for the reaction is provided which is consistent with mechanism theory. A rapid rearrangement of this type would be consistent with previous<sup>12</sup> kinetic data.

**Acknowledgment.**—Support of this work by the Research Corporation and the Robert A. Welch Foundation is hereby gratefully acknowledged. The authors wish to express appreciation of Dr. E. S. Lewis of Rice University for suggesting the plot shown in Fig. 8.

### Experimental

All rate constants have the units  $\text{sec.}^{-1}$ . Melting points and boiling points are uncorrected.

**Kinetic Procedures.**—All decompositions of II were carried out in a constant temperature light oil-bath at  $41.00^\circ$ . The bath was lined with copper screen between the main heater and oil reservoir to give a large surface for heat transfer. The reaction apparatus was a 50-ml. flask shaped like an erlenmeyer flask (to give high liquid surface area and provided for rapid gas-liquid equilibration) except a side tube was equipped with a serum cap and the main outlet at the neck was fitted with a ball joint. The reaction mixtures were vigorously stirred by an iron bar incased in glass. The bar was turned by an AlNiCo magnet underneath the flask, which was connected to an outside motor through a speedometer cable. The ball joint was connected through a condenser to a manometer and a gas buret. Aliquots were taken through the serum cap.

The following was a typical procedure: To the reaction flask was added 3 ml. of water, 14.2 ml. of absolute ethanol and 0.06 ml. of pyridine. The flask was placed in the constant temperature bath and allowed to warm for 10 minutes. A 0.4-ml. (micropipet) sample of II which had been prewarmed to  $41^\circ$  was injected through the serum cap into the vigorously stirring solution. Samples (0.1 ml.) were withdrawn at 15 sec. 5, 10, 15, 20, 30, 40, 50 and 60 min. Each

(11) In fact, pyridinium O-ethyl thiocarbonate will precipitate from kinetic runs in pure methanol employing high concentrations (1 *N*) of pyridine—with a change in reaction rate (evolution kinetics) to zero order. If this salt is dissolved in water it undergoes instantaneous decomposition.

(12) I. Iwasaki and S. R. B. Cooke, *J. Am. Chem. Soc.*, **80**, 285 (1958).

sample was quenched immediately in 10 ml. of 95% alcohol containing 0.09 ml. of 0.03 *N* hydrochloric acid. These samples were stored at 0° until diluted 15:1 with alcohol for spectrophotometric measurement of undecomposed II. The reaction mixture was kept under an atmosphere of carbon oxysulfide at atmospheric pressure—maintained with the gas buret (no reabsorption of COS occurs during the reaction). The ultraviolet spectrum of II shows two peaks, one at 232 *mμ* and the other at 276 *mμ* with the latter having  $\epsilon$  10,100. Although either peak can be used for measurement, somewhat better kinetics are obtained with 276 *mμ* peak. Compound II follows Beer's law from  $1 \times 10^{-5}$  to  $8 \times 10^{-5}$  *M*—the range of concentrations employed in these studies. Infinite time absorption due to pyridine, etc., is very low amounting to  $\epsilon$  ca. 80 at 276 *mμ*.

In experiments where buffer components were varied, the ionic strength was maintained at a high value to minimize changes in ionic strength owing to changes in buffer components. These runs were carried out in 0.25 *M* LiClO<sub>4</sub>, made up in the aqueous phase of the solvent pair before mixing. This concentration of LiClO<sub>4</sub> was arrived at experimentally by showing that changes in salt concentration above this concentration have a negligible effect on rate (see Fig. 3). Lithium chloride was also effective but required almost 1 *M* concentration to accomplish the same purpose.

The measurement of evolved carbon oxysulfide was performed with the gas buret connected to the same system as previously described. The gas buret itself was enclosed in a water jacket held at constant temperature. Ideal gas law computations of gas volume to moles of II reacted give results of sufficient accuracy for these measurements. Gas evolution and sampling kinetics were not performed on the same samples.

The rate of decomposition of potassium O-ethyl thiocarbonate (IV) was studied by gas evolution in the same apparatus in aqueous solution. Gas evolution measurements gave excellent results in aqueous solution (probably because of the low solubility of carbon oxysulfide in water) and showed 98–100% completions. The buffers used to obtain the *pH* values shown in Fig. 6 were NaOH-potassium acid phthalate for the low values of *pH* and NaOH-monopotassium phosphate for the point of highest *pH* employed (6.20). The concentrations of buffer components and lithium chloride were maintained at a sufficiently high level to rule out the possibility of changes in *pH* or ionic strength. The initial concentration of III was 0.077 *M*. The *pH* of the buffer solution was measured before and after each experiment to ascertain if any changes had occurred in *pH*. The following is a typical procedure: A solution composed of 3 ml. of 0.033 *N* HCl and 9 ml. of buffer solution was placed in the reaction flask and allowed to warm to 41.00°. To this warmed solution was added 1 ml. of 1 *M* potassium O-ethyl thiocarbonate solution which had also been prewarmed in the temperature bath. The evolved carbon oxysulfide was measured with the gas buret. Readings were taken each minute for the first 5 minutes, then every 2 minutes for the next 10 minutes and finally every 5 minutes until gas evolution ceased.

**Product Studies.**—The isolated yield of *n*-butyl mustard oil from II varies from 82%, employing pyridine in benzene or triethylamine in benzene to 93% using 20 ml. of 0.5 *M* KOH in aqueous solution.<sup>13</sup> The following is a specific example: To 5 ml. (5.5460 g.) of II in 63 ml. of methanol was added 3 ml. of  $\gamma$ -collidine and the resulting solution was refluxed for 3 hours. The reaction mixture was diluted with 100 cc. of water, extracted with ether, dried and distilled. The yield of *n*-butyl mustard oil was 90%, b.p. 83–84° (35 mm.),  $n_D^{25}$  1.4939 (lit.<sup>13</sup> b.p. (35 mm.) 83.5°,  $n_D^{25}$  1.4933).

**Materials.**—Carboethoxy *n*-butyldithiocarbonate was prepared as follows. A mixture of 20 cc. of water and 10 cc. of carbon disulfide in a pressure bottle was cooled in ice-salt, and 16 cc. of *n*-butylamine was added slowly down the side of the vessel to avoid contact with the carbon disulfide. The closed bottle was shaken and chilled, and 9.5 cc. of 17 *M* potassium hydroxide solution was added. After the carbon disulfide had completely reacted the mixture was chilled, treated dropwise with 15.2 cc. of ethyl chloroformate, and let stand an hour at room temperature. The separated layer of carboethoxy *n*-butyldithiocarbonate was taken up in 100 cc. of ether, washed with three portions of water, one portion of 1 *N* hydrochloric acid and with three additional portions of water and dried over calcium sulfate. The ether was evaporated on a rotary evaporator under vacuum (50 mm.) for 8 hours. The product was a clear yellow oil which can be stored many months at 0° without decomposition. The compound had  $n_D^{25}$  1.5288 and sp. g. 27° 1.1107. The compound showed major infrared bands at 3.1, 5.95, 6.6, 8.8, 10.95 and 14.9  $\mu$  (Beckman IR-5 with NaCl optics). Two batches of the oil were analyzed.

*Anal.* Calcd. for C<sub>8</sub>H<sub>13</sub>O<sub>2</sub>NS<sub>2</sub>: C, 43.5; H, 6.85; N, 6.35. Found (av., of two analyses): C, 43.84; H, 6.51; N, 6.55.

Solvents were AR grade and were used without purification. Reagent grade pyridine and pyridine derivatives were distilled (500 g.) through an efficient column and a center cut (50 g.) having the best physical properties was used for the kinetic procedures. Pyridine hydrochloride was prepared by passing dry HCl gas into an ether-pyridine solution. The precipitate was washed with ether, heated to 70° under vacuum for 30 minutes to decompose occluded HCl and dried *in vacuo*. Lithium perchlorate was prepared by potentiometric titration of LiOH and HClO<sub>4</sub>, collected by evaporating the solution, and used as the trihydrate. Potassium O-ethyl thiocarbonate was prepared by passing carbon oxysulfide gas (Matheson, reagent) into a mixture of ethanol and water saturated with potassium hydroxide at 0° until precipitation ceased. The salt was filtered, washed with ether, recrystallized twice from absolute ethyl alcohol, filtered and dried in a vacuum desiccator; m.p. 206°.

(13) J. E. Hodgkins, The Mustard Oils of Rapaz and Papaya, Thesis, Rice University, 1954.