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Kinetics and Mechanism of Methylthiomethyl Acetate Conversion into Bis(methylthio)methane

A. P. Zaraisky

Litvinenko Institute of Physicoorganic and Coal Chemistry, National Academy of Sciences of Ukraine, Donetsk, Ukraine
e-mail: azar@skif.net

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Abstract—Methylthiomethyl acetate can be rapidly and quantitatively converted into bis(methylthio)methane via an acid-catalyzed reaction under mild conditions. This conversion in the system benzene–aqueous sulfuric acid is interpreted in terms of a complex scheme of consecutive and parallel steps. It begins with the hydrolysis of the ester by the mechanisms $A_{AC}2$ and $A_{AL}1$ (in Ingold's terminology), and the subsequent interactions involve intermediate products. The kinetics of the process is analyzed. A procedure for the synthesis of bis(methylthio)methane from dimethyl sulfoxide is suggested.

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α-Acetoxymethyl sulfides (**I**, X = OAc, Y = H) are readily and smoothly hydrolyzable by water [1]:

$$\underset{\textbf{I}}{\mathrm{RSCH(Y)X}} + \mathrm{H_2O} \longrightarrow \mathrm{RSH} + \mathrm{YCHO} + \mathrm{XH}. \qquad \mathrm{(I)}$$

The corresponding chlorosubstituted compounds (X = Cl) also react readily with water [2–4]. It was reported that the hydrolysis of these substrates (**I**, R = Et, X = Cl, Y = H) in aqueous dioxane is autocatalyzed by HCl being released [3]. However, this conclusion was challenged in a later kinetic study [4], and the S_N1 mechanism was accepted for the hydrolysis of substituted thioethers (**I**; R = Me, Ar; X = Cl; Y = H, Me, Ph) [2]:

$$\begin{aligned}&\mathrm{RSCH(Y)X} \longrightarrow \mathrm{X^-} + \mathrm{RSCH(Y)^+}\\&\xrightarrow[-\mathrm{H^+}]{\mathrm{H_2O}} \underset{\textbf{II}}{\mathrm{RSCH(Y)OH}} \xrightarrow{(\mathrm{H_2O})} \mathrm{RSH} + \mathrm{YCHO}.\end{aligned} \qquad \mathrm{(II)}$$

The first step of this reaction is rate-determining. This mechanism apparently rules out any significant assistance from the acid. At the same time, it was reported that this reaction yields thioacetal **III** [3, 4]:

$$\begin{aligned}&2\mathrm{RSCH(MeS)X} + \mathrm{H_2O}\\&= \underset{\textbf{III}}{\mathrm{(RS)_2CHY}} + \mathrm{YCHO} + 2\mathrm{HX}.\end{aligned} \qquad \mathrm{(III)}$$

Therefore, there can be no rapid and complete conversion of semithioacetal **II** into a thiol and an aldehyde [2], because **II** is an intermediate in the thiol + aldehyde reaction yielding thioacetal **III** (which is usually acid-catalyzed).

Furthermore, it turned out (see Experimental) that methylthiomethyl acetate (**Ia**, R = Me, Y = H, X = OAc) reacts with water rather slowly even under boiling conditions and that the reaction is catalyzed by solutions of strong acids even at room temperature to produce, in quantitative yield, bis(methylthio)methane $(\mathrm{MeS})_2\mathrm{CH_2}$ (**IIIa**), the simplest thioacetal-based carbonyl synthonogen [5]. Use of the corresponding sulfoxides as the starting compounds might open up a new route to more complex compounds of this class, particularly in view of the fact that methylthiomethyl acetate is readily obtainable from DMSO by the Pummerer rearrangement. The purpose of this work is to study the kinetics and mechanism of the acid-catalyzed conversion of methylthiomethyl acetate into bis(methylthio)methane and to develop an efficient synthetic procedure for this reaction. The mechanism of the acid-catalyzed conversion of $\mathrm{MeSCH_2OAc}$ (**Ia**) is obviously different from the mechanism of the same conversion of α-chlorosubstituted sulfides.

EXPERIMENTAL

^{1}H NMR spectra were recorded on a Gemini-200 spectrometer operating at 200 MHz.

The analytical part of the work was carried out by GLC using an LKhM-80 chromatograph (model 6). The chromatographic conditions were as follows: column temperature of 140°C, 3 mm × 3 m stainless steel column, stationary phase QF-1 (5%) on Inerton Super, particle size of 0.16–0.20 mm, helium as the carrier gas, and *n*-tridecane as the internal standard. The correctness of measurements was checked against mixtures containing the standard, methylthiomethyl acetate (**Ia**), and bis(methylthio)methane (**IIIa**) in appropriate proportions.

The conversion of **Ia** was studied in the two-phase system benzene–aqueous sulfuric acid (25°C). A mixture (5–7 ml) of a benzene solution of methylthiomethyl acetate and the standard and an aqueous solution of the acid, in a 20-ml conical flat-bottom flask, was magnetically stirred on a temperature-controlled water bath at 25 ± 1°C. The organic phase or, in the case of a high reaction rate at a high acidity, the averaged reaction mixture was sampled while stirring. The samples (0.05–0.10 ml) were treated with solid sodium bicarbonate, dried with anhydrous magnesium sulfate, and chromatographed. Chromatographic data were analyzed in terms of the height ratios of the peaks from the substrate (**Ia**) and the product (**IIIa**) to the peak of the internal standard.

*Experiment 1: Interaction methylthiomethyl acetate (**Ia**) and water.* Ester **Ia** (1 ml) was vigorously stirred at 32°C with water (1 ml) for 10 min. The composition of the solution did not change. Stirring for 3 min under boiling conditions resulted in the conversion of 7% of the ester.

*Experiment 2: Distribution of the semithioacetal $MeSCH_2OH$ (**IIa**).* A benzene solution (11.4 ml) of *n*-tridecane (0.1177 g) was divided into two equal portions. Ester **Ia** (0.1406 g) was dissolved in the first portion, and the solution was stirred with 53.8% sulfuric acid (5.7 ml) for 9 min. The resulting phases were separated. For the organic phase, the height ratio of the GLC peaks of thioacetal **IIIa** and the standard was 0.381. After the 2-h-long stirring of this phase with a new 5.7-ml portion of 53.8% sulfuric acid, the peak height ratio remained unchanged (0.384). The water–acid phase separated was stirred with the second half of the benzene solution of the standard (5.7 ml) for 90 min. This resulted in the thioacetal/standard peak height ratio increasing from zero to 0.182. The observed peak ratios indicate that the yield of **IIIa** is nearly quantitative. This was confirmed by chromatographing an artificially composed mixture of this product and the standard. Thus, the entire semithioacetal **IIa** was in the polar water–acid phase and did not pass into the organic phase.

*Experiment 3: Distribution of bis(methylthio)methane (**IIIa**) and methylthiomethyl acetate (**Ia**).* A solution of bis(methylthio)methane (0.3369 g) and *n*-tridecane (0.1547 g) in benzene (50 ml) was prepared. The height ratio of the chromatographic peaks was 1.23. This solution (5 ml) and aqueous sulfuric acid with various acid concentrations (5 ml) were combined and stirred for 10 min. Within the accuracy of chromatographic analysis, the thioacetal did not pass into 12.3–63.3% H_2SO_4 and remained entirely in the organic phase: the thioacetal/standard peak height ratio remained equal to 1.20–1.22 and decreased to 1.17 only in the case of 63.3% sulfuric acid. The distribution of methylthiomethyl acetate between benzene and water was studied in the same way. It was found that **Ia** almost does not pass into water and its water solubility is $\lesssim 0.3\%$ at 25°C (the author's data). At the same time, **Ia** is miscible with benzene in all proportions.

Experiment 4: Organic phase containing n-tetradecane at elevated acidities. n-Tridecane (internal standard, 0.0346 g) and **Ia** (0.1179 g) were dissolved in *n*-tetradecane (5 ml) and were combined with 53.8% sulfuric acid (5 ml), and the mixture was stirred at 25°C. After 1 min, the extent of the reaction was 92%. The initial reaction rate in *n*-tetradecane was higher than that in benzene (Table 1, entries 13, 14) by a factor of at least 5–10.

Reaction rates were derived from kinetic curves by graphical differentiation. Tangent lines at given points of a curve were drawn using a flat mirror with straight edges. Reaction orders along the kinetic curves were also calculated using familiar relationships between the logarithms of the reaction rate and the reactant concentrations [6] (see below).

*Methylthiomethyl acetate (**Ia**)* was obtained from DMSO and acetic anhydride via the Pummerer reaction [7]. The product contained 2–4% bis(methylthio)methane (**IIIa**), and this was taken into account in calculations.

*Bis(methylthio)methane (**IIIa**).* Absolute DMSO (25 ml, 25 g, 0.32 mol) and acetic anhydride (41 ml, 43.9 g, 0.43 mol) were combined in a round-bottom flask (300–500 ml) fitted with a backflow condenser and a calcium chloride tube. The mixture was brought to a vigorous boil by heating to 130°C and was held at this temperature for 1 h. Next, the reaction mixture was cooled to room temperature and 51% sulfuric acid (50 ml) was added without allowing a temperature rise. The resulting homogeneous mixture was left to stand for 5 min. Thereafter, water (50 ml) and sodium sulfate (20 g) were added and the mixture was stirred until the salt dissolved. After phase separation, the organic phase was washed with 20 ml of water and was dehydrated with magnesium sulfate. This yielded 17.2 g (98.4%) of bis(methylthio)methane. Found, %: C 32.96, H 7.49, S 55.54. $C_3H_8S_2$. Calculated, %: C 33.29, H 7.46, S 59.23, bp 147°C. ^{1}H NMR (MeCN, TMS), δ, ppm: 12.11 (s, 6H, $2CH_3$), 3.66 (s, 2H, CH_2). $n_D^{20} = 1.5340$; $d_4^{20} = 1.059$. The physicochemical properties of the product are close to those reported in the literature [8, 9].

RESULTS AND DISCUSSION

The kinetic data obtained are presented in Table 1 and in the figures. The data listed in Table 1 indicate that the reaction is acid-catalyzed. The coincidence of the kinetic curves of substrate **Ia** conversion in runs 1 and 2 in the coordinates of Fig. 1 is evidence that the reaction is first-order with respect to the substrate at the low acidities of the system. At higher acidities, the analogous curves (Fig. 2, curves *1*, *2*) diverge. Under these conditions, the effective reaction order calculated in a familiar way [6, 10] from ten data points at pairwise equal conversions of ester **Ia** is 1.83 ± 0.10. The reaction order derived from the initial reaction rates in these

Table 1. Rate constants (k_1) and initial rates of methylthiomethyl acetate (**Ia**) consumption and bis(methylthio)methane (**IIIa**) accumulation in the system benzene–aqueous sulfuric acid (25°C)

Run no.	$[H_2SO_4]$, %	Initial concentrations[a] of ester **Ia**, HCHO, and AcOH, mol/l	$-(d[\mathbf{Ia}]/dt)_0$, mol l^{-1} h^{-1}	$(d[\mathbf{IIIa}]/dt)_0$, mol l^{-1} h^{-1}	$k_1 \times 10^2$, min^{-1}
1	12.3	0.186	0.10	(0)[d]	0.9
2	12.3	0.742	0.50	(0)[d]	1.1
3	12.5	0.154	0.09	(0)[d]	1.0
		HCHO 0.443			
4	12.3	0.190	0.19	(0)[d]	1.7
		AcOH 0.130			
5	12.5	0.187	0.070	(0)[d]	0.62
		HCHO 0.135; AcOH 0.162			
6[b]	12.3	0.210	2.3	(0)[d]	18
7[c]	12.3	0.184	2.2	(0)[d]	20
8	29.3	0.194	0.56	~0.01	4.7
9	44.6	0.0545	0.92	0.14	24
10	44.6	0.184	1.5	0.11	13
11	44.6	0.545	17	2.6	45
12	49.3	0.186	2.35	0.23	19
13	53.8	0.183	4.5	1.1	31
14	53.8	0.187	3.5	0.52	27
15	54.0	0.166	6.4	1.5	49
		HCHO 0.285			
16	54.0	0.168	2.4	0.61	18
		AcOH 0.115			
17	58.5	0.187	12	3.7	74
18	61.5	0.041	5.9	1.2	190
19	61.5	0.182	27	4.1	210
20	63.3	0.053	7.5	1.8	180

[a] In the organic or acid phase.
[b] The organic phase is *n*-tetradecane.
[c] The organic phase is *n*-hexane.
[d] The induction period is up to 40 min (the initial **IIIa** appearance rate is below the detection limit).

two runs is 2.2. Below, these data are substantially corrected. However, even now they provide preliminary insight into the mechanism of the reaction. At low acidities, the reaction is controlled by the slow formation of the reactive species. This gives rise to an induction period in product accumulation (Fig. 1, curves *3*, *4*). As the acidity is raised, the formation of the reactive species increases, the consumption of this species in its reaction with the initial substrate **Ia** begins to affect the substrate consumption rate, and the reaction order with respect to the substrate increases. Simultaneously, the induction period in the formation of **IIIa** vanishes. This situation is illustrated by curve *4* in Fig. 2.

This description is, however, oversimplified. In runs 10 and 11 (Table 1, Fig. 2), the order of the reaction with respect to methylthiomethyl acetate (**Ia**) decreases to 0.5 or below for **Ia** conversions of 50–60%. Thus, although the above result 1.83 ± 0.10 has a narrow confidence interval (small standard deviation), it does not correctly characterize the simple kinetics examined. The limitations of this method of determination of the reaction order were noted by Eremin [10]. The decrease of the reaction order in time is conventionally attributed to the formation of intermediates and the autoacceleration of the process [11]. However, within the accuracy of our measurements, no decrease in the reaction order with respect to ester **Ia** was observed at the lowest and at the highest acidities of the system (12.3 and 61.5% H_2SO_4, respectively). In both case, the initial ("concentrational") and "temporal" [11] orders of the reaction were close to unity (0.9–1.2). The kinetic curves obtained in runs 18 and 19 (Table 1) coincide within the experimental error, as in the runs illustrated by Fig. 1.

The induction period in bis(methylthio)methane accumulation (acceleration of the rate-limiting step, which is the formation of the reactive species) is shortened not only as the acidity is raise, but also as the initial concentration of the starting ester **Ia** is increased (Fig. 1, curve *4*). The observed initial rates of **IIIa** accu-

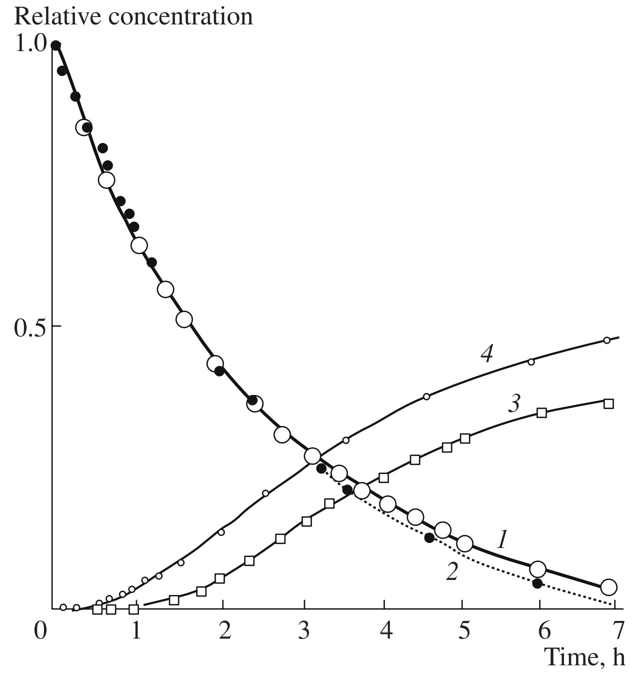

Fig. 1. Kinetics of (*1*, *2*) methylthiomethyl acetate consumption and (*3*, *4*) bis(methylthio)methane accumulation in the system benzene–12.3% sulfuric acid (25°C): runs 1 and 2 (Table 1), respectively.

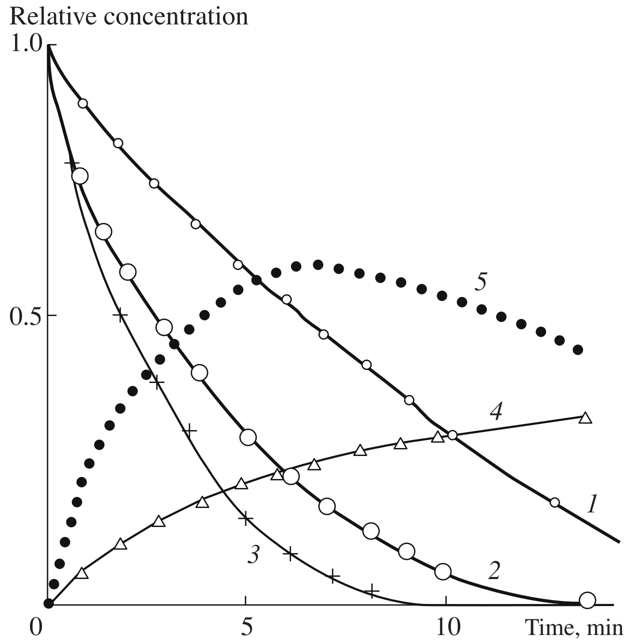

Fig. 2, (*1–3*) Kinetics of methylthiomethyl acetate consumption in runs 10, 11, 14 (Table 1), respectively. (*4*) Bis(methylthio)methane and (*5*) semithioacetal accumulation kinetics in run 14 (Table 1). Benzene–aqueous sulfuric acid system; 25°C.

mulation are listed in Table 1. In Fig. 3, these rates are plotted as a function of the acidity function H_0.

In all runs, the substrate **Ia** consumption rate was higher than the thioacetal **IIIa** formation rate; that is, an intermediate built up in the system. Methanethiol, a product of complete hydrolysis, which is more volatile than benzene under the GLC conditions, was observed in trace amounts and only in low-acidity media some time after the beginning of the reaction.[1] It disappeared in the course of the reaction. Therefore, the decomposition of semithioacetal **IIa**, which is the most likely intermediate (see [2]), is only a minor part of the process. Considerable amounts of semithioacetal can accumulate (Fig. 2). The most likely active species in the reaction is the cation $\mathrm{MeSCH_2^+}$ (**IV**) [7].

Thus, the scheme of the reaction (see [12]) seems to be a complicated set of parallel and consecutive elementary steps. The first part of the set is similar to the acid hydrolysis of organic acetates. The second part, which begins with the formation of the reactive intermediate, accounts for the subsequent transformations in the system.

$$\underset{\textbf{Ia}}{\mathrm{MeSCH_2OAc}} + \mathrm{H^+} \rightleftarrows \underset{\textbf{V}}{\mathrm{MeSCH_2O^+(H)Ac}}, \tag{IV}$$

$$\underset{\textbf{V}}{\mathrm{MeSCH_2O^+(H)Ac}} + \mathrm{H_2O} \rightleftarrows \underset{\textbf{IIa}}{\mathrm{MeSCH_2OH}} + \mathrm{AcOH_2^+}, \tag{V}$$

$$\underset{\textbf{IIa}}{\mathrm{MeSCH_2OH}} + \mathrm{H^+} \rightleftarrows \underset{\textbf{VI}}{\mathrm{MeSCH_2OH_2^+}}, \tag{VI}$$

$$\underset{\textbf{VI}}{\mathrm{MeSCH_2OH_2^+}} \rightleftarrows \underset{\textbf{IV}}{\mathrm{MeSCH_2^+}} + \mathrm{H_2O}, \tag{VII}$$

$$\underset{\textbf{IV}}{\mathrm{MeSCH_2^+}} + \underset{\textbf{Ia}}{\mathrm{MeSCH_2OAc}} \rightleftarrows \underset{\textbf{VII}}{\mathrm{MeS^+(CH_2SMe)CH_2OAc}}, \tag{VIII}$$

$$\underset{\textbf{VII}}{\mathrm{MeS^+(CH_2SMe)CH_2OAc}} \xrightarrow{\mathrm{H_2O}} \underset{\textbf{IIIa}}{\mathrm{(MeS)_2CH_2}} + (\mathrm{AcOH} + \mathrm{HCHO} + \mathrm{H^+}). \tag{IX}$$

Along with reaction (V), which is the key[2] step of the hydrolysis mechanism $A_{\mathrm{AC}}2$, reactions (X) and (XI) should be taken into consideration, which are the key

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[1] Formaldehyde, which resulted from the reaction along with the ultimate product **IIIa**, did not pass into the organic phase and, accordingly, was not detected by GLC.

[2] For this scheme, our approach does not require that step (V), (X), or (XI) be rate-limiting, as in the hydrolysis of acetates. Therefore, the term *key* in the sense of *characteristic* seems correct.

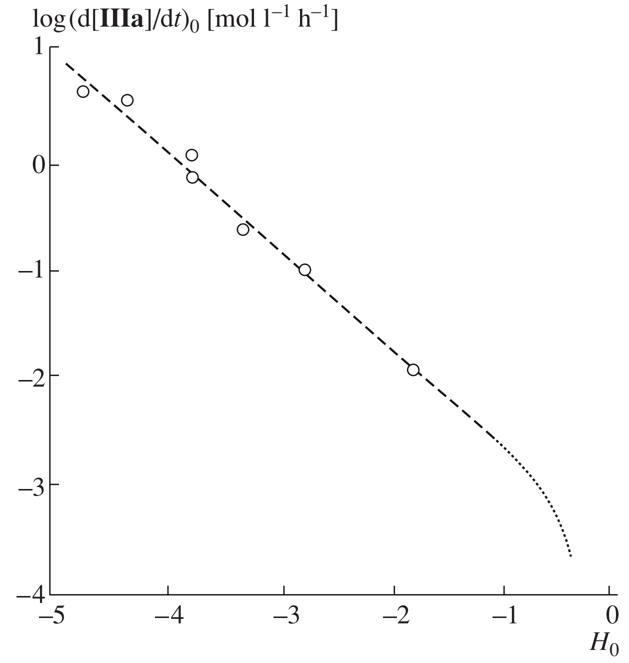

Fig. 3. Initial bis(methylthio)methane (**IIIa**) accumulation rate at the initial methylthiomethyl acetate concentration $[\mathbf{Ia}]_0 = 0.186 \pm 0.004$ mol/l versus the acidity function H_0 for the system benzene–aqueous sulfuric acid at 25°C. The dashed line is the linear regression $\log(\mathrm{d}\,[\mathbf{IIIa}]/\mathrm{d}t)_0 = -(3.66 \pm 0.23) - (0.93 \pm 0.06)H_0$ ($r = -0.988$, $n = 7$).

steps of the mechanisms $A_{\mathrm{Al}}1$ and $A_{\mathrm{AC}}1$, respectively [12].

$$\underset{\mathbf{V}}{\mathrm{MeSCH_2O^+(H)Ac}} \rightleftharpoons \underset{\mathbf{IV}}{\mathrm{MeSCH_2^+}} + \mathrm{AcOH}\,, \qquad \mathrm{(X)}$$

$$\underset{\mathbf{V}}{\mathrm{MeSCH_2O^+(H)Ac}} \rightleftharpoons \underset{\mathbf{IIa}}{\mathrm{MeSCH_2OH}} + \mathrm{Ac}\,. \qquad \mathrm{(XI)}$$

We will analyze our experimental data using Ingold's approach (as represented by Yates and McClelland) [12, 13]. It is necessary to choose among the steps characteristic of the three acid-catalyzed hydrolysis mechanisms. Under the assumption that the concentration of the sulfonium transition complex **VII** is low (possibly quasi-steady-state) as compared to the concentrations of the other components of the system (that is, **VII** does not accumulate or accumulates at a low rate), the substrate **Ia** consumption rate in reaction (V), (X), or (XI) can be taken to be equal to the difference between the overall **Ia** consumption rate and the thioacetal **IIIa** formation rate. Table 1 lists the corresponding rate constants (k_1). The dependence of the Yates–McClelland function on the logarithm of the activity of water for methylthiomethyl acetate (**Ia**) hydrolysis (Fig. 4, light circles) is essentially similar to the same dependence for the hydrolysis of organic acetates in aqueous sulfuric acid [12, 13].

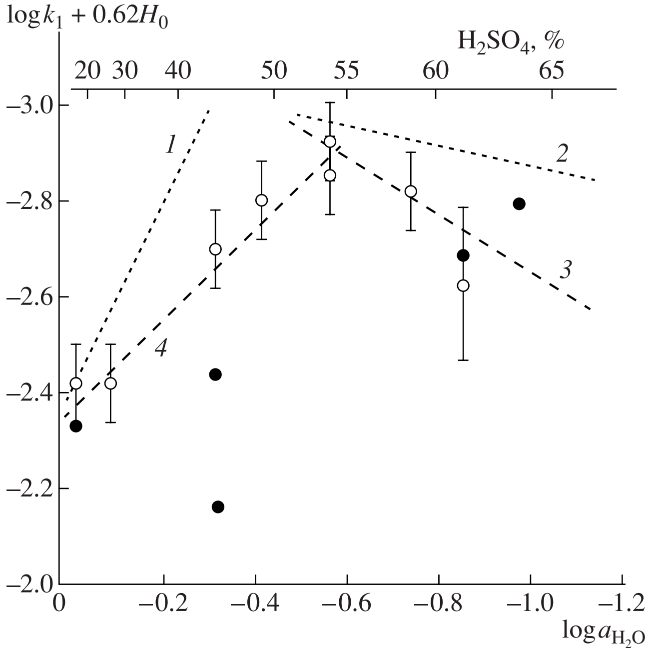

Fig. 4. Yates–McClelland function [13] versus the logarithm of the activity of water in the system benzene–aqueous sulfuric acid at 25°C. Dashed straight lines *1*, *2*, and *3*, with slopes of 2.0, –0.2, and –0.6, refer to the alkyl and aryl acetate hydrolysis mechanism $A_{\mathrm{AC}}2$, $A_{\mathrm{AC}}1$, and $A_{\mathrm{AL}}1$, respectively [12, 13]. Dashed line *4* represents data of this work as the regression line $\log k_1 + 0.62H_0 = (-2.34 \pm 0.03) + (0.99 \pm 0.07)\log a_{\mathrm{H_2O}}$ ($r = 0.988$, $n = 7$). Light circles: $[\mathbf{Ia}]_0 = 0.186 \pm 0.004$ mol/l. The dark circles represent the data obtained without adding any admixtures or at other substrate **Ia** concentrations (Table 1).

These data rule out the mechanism $A_{\mathrm{AC}}1$. It is known to be unnatural for alkyl acetates in dilute sulfuric acid [12]. The right-hand branch of the experimental curve in Fig. 4 testifies to the mechanism $A_{\mathrm{AL}}1$; that is, it is consistent with the reaction proceeding via steps (IV) and (X). So, substrate **Ia** can be viewed as an ester showing a behavior intermediate between *tert*-butyl acetate (whose hydrolysis mechanism is pure $A_{\mathrm{AL}}1$) and benzyl acetate [13] (kinetic data for cyclopropylcarbynyl acetate hydrolysis [14] almost coincide with, or are very similar to, the same data for methylthiomethyl acetate). The fact that the slope of the left branch decreases (extra acceleration takes place) as the acidity is raised, becoming less than 2, a value typical of simple alkyl acetates ($A_{\mathrm{AC}}2$ mechanism), is evidence that either substrate **Ia** is more liophilic or the $A_{\mathrm{AL}}1$ mechanism makes a considerable contribution to hydrolysis at low acidities of the system. This type of behavior of the Yates–McClelland function is also typical of benzyl acetate [13], but to a lesser extent. The local minimum in the k_1 versus acid concentration curve, which is due to the crossover from the $A_{\mathrm{AC}}2$ mechanism to the $A_{\mathrm{AL}}1$ mechanism [12, 13], is not observed here, as in the case of cyclopropylcarbynyl acetate, and possibly for the same reason.

For the mechanism $A_{\mathrm{AL}}1$, the appearance of the reactive species $\mathrm{MeSCH_2^+}$ (**IV**) must be accompanied by its interaction with water, the most abundant and the strongest nucleophile in the system. As a consequence, the equilibrium in step (VII) must be substantially shifted to the left (see [12]) and, irrespective of the mechanisms of the preceding steps, the protonated species of the semithioacetal (**VI**) must serve as a source of cation **IV** in the system.

At low acidities, steps (VIII) and (IX) are retarded, there is an induction period in the formation of the final product **IIIa**, and the consumption of substrate **Ia** in step (VIII) is determined by the slow formation rate of reactive species **IV** via reaction (VII). The overall consumption of **Ia** is a first-order process (Fig. 1) and is due to reaction (V), obeying the law

$$-\mathrm{d}[\mathbf{Ia}]/\mathrm{d}t = k_5\gamma_{\mathbf{V}}(h_0/K_{\mathbf{V}})^{0.62}a_{\mathrm{H_2O}}[\mathbf{Ia}] \qquad (1)$$

(where $K_{\mathbf{V}}$ is the thermodynamic acidity constant of the protonated form of methylthiomethyl acetate, γ is the activity coefficient, h_0 is the nonlogarithmic Hammett acidity function, and a is activity) until the end of the induction period in product **IIIa** accumulation, when step (VIII) becomes significant. Assuming that the concentration of the reactive species **IV** is quasi-steady-state and that step (VII) is the source of this species, we can write the following expression for the substrate **Ia** consumption rate:

$$-\mathrm{d}[\mathbf{Ia}]/\mathrm{d}t = k_8k_{-8}a_{\mathbf{VII}}\gamma_{\mathbf{Ia}}[\mathbf{Ia}]/(k_{-7}a_{\mathrm{H_2O}}). \qquad (2)$$

This expression, takes into account the following plausible conditions:

$$k_{-7}a_{\mathrm{H_2O}} \geqslant k_8\gamma_{\mathbf{Ia}}[\mathbf{Ia}], \qquad (3)$$

$$k_7a_{\mathbf{VI}} \leqslant k_{-8}a_{\mathbf{VII}}. \qquad (4)$$

Condition (3) accounts for the fact that water is more abundant in the system and is more nucleophilic than **Ia**. Condition (4) reflects the low acidity and, accordingly, the low concentration of protonated semithioacetal **VI**. Within one kinetic experiment (Fig. 1), expressions (1) and (2) account for the first order of the reaction with respect to substrate **Ia** before and after the induction period in thioacetal **IIIa** accumulation, but only at a constant concentration of the sulfonium transition complex **VII**, which is possible in a quasi-steady-state regime. For low acidities, this in another argument in favor of the above assumption that the concentration of **VII** is low.

Acetic acid and formaldehyde added to the initial reaction mixture separately or together exert only a slight effect on the reaction (Table 1, runs 3–5, 15, 16) even if their amounts are overstoichiometric. Therefore, the final products have no effect on the determination of reaction rates, particularly for the initial portions of the kinetic curves; that is, steps (V) and (X) can be treated here as unidirectional. Of interest is the effect of acetic acid at low acidities (run 4), which apparently serves as a weak bifunctional catalyst [15].

In Fig. 4, the dark circles show the results obtained for substrate **Ia** initial concentrations other than 0.186 ± 0.004 mol/l. Note that the deviation of the data points from the fitting relationship at intermediate acidities of the system are far beyond the limits of the experimental error (Table 1, runs 9–11), as distinct from the data points for the lowest and the highest acid contents. This deviation is unpredictable and cannot be understood without performing a deeper analysis.

The significant role of the mechanism $A_{\mathrm{AL}}1$ (i.e., step (X)) throughout the acidity range examined will be understandable if the comparatively high stability of the resonance-stabilized methylthiomethyl cation **IV** [7] is taken into account. Furthermore, the comparative stability of **IV** may be the cause of the variability of the rate coefficients (effective constants) in runs 9–11 (Table 1) and of the fact that the reaction order with respect to substrate **Ia** increases as the substrate is consumed at intermediate acidities of the system.

An analysis demonstrated that the second order of the reaction with respect to **Ia** (1.83 and 2.2; see above) and the decrease of the order in some kinetic experiments (Table 1, runs 10, 11; Fig. 2) are inexplicable in the framework of the above reaction scheme in terms of the equilibrium or quasi-steady-state concentrations of the active species.

Figure 5 is illustrative. As the acidity of the system increases, once steps (X) and (VII) have stopped limiting the **Ia** consumption rate in the consecutive reactions (VIII) and (IX), there will be a strong nonlinearity of the logarithm of the initial reaction rate as a function of the logarithm of the substrate **Ia** concentration. It is clear from Fig. 5 that the initial (concentrational, or “true” [11]) order of the reaction in run 11 is close to 4 and increases steeply with an increasing concentration. Therefore, the calculated orders of 2.2 and 1.83 are some intermediate values between 1 (run 10) and 3.9 (run 11, Fig. 5) and depend on the determination method.

This increase of the reaction order with respect to the initial ester is hardly due to a factor other than deviation from the quasi-steady-state concentration of the comparatively stable reactive species **IV** [7] in the system. Here, the concept of reaction order loses the meaning it has in chemical kinetics. It can readily be demonstrated that, in the case of the power-law dependence of the reaction rate on the concentration and a variable (concentrational, or “true” [11]) reaction “order,” a curve like that shown in Fig. 5 must be convex toward the abscissa axis. Here, it is likely more correct to use the term *quasi-order, current order*, or *effective initial order* because a conventional reaction order determination method was used.

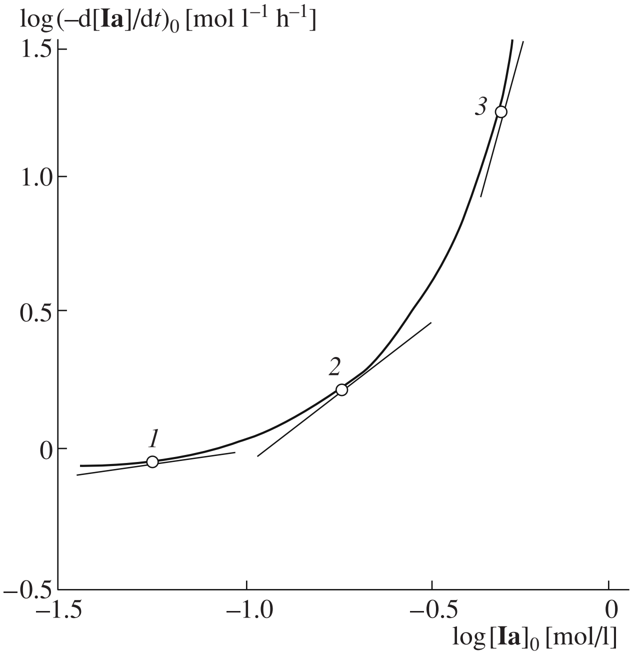

Fig. 5. Initial "true" ("concentrational" [11]) orders of the reaction (*n*) with respect to methylthiomethyl acetate (**Ia**) for its consumption in the system benzene–44.6% sulfuric acid at 25°C: (*1*) $[\mathbf{Ia}]_0 = 0.0545$ mol/l, $n \approx 0.2$; (*2*) $[\mathbf{Ia}]_0 = 0.184$ mol/l, $n = 1.0$; (*3*) $[\mathbf{Ia}]_0 = 0.545$ mol/l, $n = 3.9$ (runs 9, 10, and 11 in Table 1, respectively).

Usually, kinetic data for composite reactions that include consecutive and parallel steps are difficult to interpret and require mathematical simulation, numerical fitting, and computer experimenting. Use of these methods was beyond the scope of this study, so we made an attempt to explain our experimental data by simpler means. Thus, applying the initial rate method to the case of a constant initial concentration of substrate **Ia**, we obtained a "kinetic section" of the system at $[\mathbf{Ia}]_0 = 0.186 \pm 0.004$ mol/l (here, the true and current order of the reaction are equal to unity for all acidities; see light circles in Fig. 4). This allowed us to assign the reaction examined to the well-known acid-catalyzed hydrolysis reactions of acetates.

Here, it is necessary to make the following remark as to the applicability of the initial rate method to the case examined. Even for two consecutive steps with an intermediate, the theoretical initial formation rate of the final product is known to be zero. Near $t = 0$, this rate is proportional to time (and to the concentration of the initial reactant). Therefore, theoretically, there must always be an induction period in such a reaction. However, if the induction period shortens, e.g., with an increasing acidity and becomes shorter than the characteristic time of reactant mixing (~1 s in our case), the experimental data points will fall into the region where the induction period is over (Fig. 2) and the apparent initial reaction rate is quite measurable. This is the region where reactions (V)–(X) develop. The observed initial rates correspond to the concentrations $[\mathbf{Ia}] = [\mathbf{Ia}]_0$ and $[\mathbf{IIIa}] = [\mathbf{IIIa}]_0 = 0$.

As can be seen, the above approach is not limited to consideration of the changes of the key, limiting, and rate-determining steps[3] and the violation of the principle of the quasi-steady-state concentration of the reactive species. However, the above assumption that the concentration of the transition complex **VII** is low (or steady-state) is necessary.

At high acidities (at which the mechanism $A_{\mathrm{AL}}1$ takes place), the reaction proceeds via steps (IV), (X), (±VII), (VIII), and (IX) and returns to the first order with respect to ester **Ia**. This result can be explained in simple terms. As the acidity of the system is raised, the rate of step (X) increases as the function $0.62H_0$, while the rate of step (VII) increases as the more rapidly growing function H_{R}.

$$-\mathrm{d}[\mathbf{Ia}]/\mathrm{d}t = k_{10}\gamma_{\mathbf{V}}(h_0/K_{\mathbf{V}})^{0.62}[\mathbf{Ia}] + \mathbf{r}_8. \tag{5}$$

Step (X) become rate-limiting. In expression (5), the term $\mathbf{r}_8 \equiv \mathrm{d}[\mathbf{IIIa}]/\mathrm{d}t$ can be neglected because this flux is only 15–20% of the overall reaction rate (Table 1) and is, therefore, close to the measurement error.

A Yates–McClelland function coefficient value of $m = 0.52$, which is likely more precise for substrate **Ia**, is obtained from kinetic and thermodynamic data using the following correlation (here, k_1 data for all runs without admixtures are taken into account, in which the initial rate is first-order with respect to methylthiomethyl acetate; see Table 1):

$$\log k_1 = -(2.35 \pm 0.13) - (0.522 \pm 0.026)H_0, \tag{6}$$

$r = 0.989$ (correlation coefficient), $n = 11$.

This value of m modifies Fig. 4 only slightly and does not change its interpretation. Thus, ester **Ia** or its protonated form **V** are somewhat more liophilic than the alkyl acetates [12].

As is demonstrated in Fig. 3, the observed initial rate of the appearance of bis(methylthio)methane **IIIa** is a linear function of H_0 as $\log(\mathrm{d}\,[\mathbf{IIIa}]/\mathrm{d}t)_0$ changes in the range of three orders of magnitude. This kind of dependence, which is unusual for reaction scheme (IV)–(X), can be explained only under arbitrary assumptions. The initial appearance rate of product **IIIa** is determined by step (VIII) (expression (7)), while step (IX) is fast:

$$\begin{aligned}(\mathrm{d}[\mathbf{IIIa}]/\mathrm{d}t)_0 &= k_8(K_7K_{\mathbf{VI}}/K_{\mathbf{V}}^{1.52})[\mathbf{Ia}]^2 \\ &\times \Delta t\{\gamma_{\mathbf{V}}h_0^{0.52}(1 + a_{\mathrm{H_2O}})/a_{\mathrm{H_2O}}\}.\end{aligned} \tag{7}$$

In this expression, k_8 and K_7 are the corresponding rate constant and thermodynamic equilibrium constant (at high acidities, condition (4) is violated and step (VII) appears as an equilibrium being established rather rap-

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[3] The term *limiting step* is not recommended for the cases in which the quasi-steady state approximation is inapplicable [6]. This term is recommended by IUPAC for homogeneous systems [17].

Table 2. Bis(methylthio)methane (**IIIa**) yield with respect to methylthiomethyl acetate (**Ia**) in the system benzene–aqueous sulfuric acid (25°C)

Run no.[a]	$[H_2SO_4]$, %	$t_{\mathbf{Ia}}$,[b] min	$Y_{t(\mathbf{Ia})}$, %	Y^c/t, %/h
1	12.3	420[d]	62	91/21
8	29.3	120	81	87/5
10	44.6	41	94	98.5/6
12	49.3	21	79	100/1.5
13	53.8	9	50	91/3
14	53.8	9	45	95/2
17	58.5	5	48	67/1
19	61.5	~1.5	29	37/1

[a] See Table 1.
[b] Complete **Ia** consumption time.
[c] Product yield in the overall reaction time *t*.
[d] The substrate **Ia** conversion is 95%.

idly); $K_{\mathbf{V}}$ and $K_{\mathbf{VI}}$ are the thermodynamic acidity constants of the respective protonated species; $\gamma_{\mathbf{V}}$ is the activity coefficient of **V**; Δt is the macrokinetic parameter, which is the same (~1 s) for all runs and is the characteristic reactant mixing time (this time is sufficient for all steps of the reaction to come to the kinetic regime, and extrapolation to $t = 0$ does not introduce any errors associated with mixing into the initial reaction rate); the quantities in braces are the parameters depending on the variable composition of the medium. Another possible way of interpreting the above dependence is by taking into account the equilibrium protonation of the oxygen atom of transition complex **VII** (present at a steady-steady concentration) with the transient formation of geminal dication **VIII**, which then decomposes rapidly under the action of water:

$$\mathbf{VII} + H^+ \rightleftarrows \underset{\mathbf{VIII}}{MeS^+(CH_2SMe)CH_2O^+(H)Ac}, \quad \text{(XII)}$$

$$\mathbf{VIII} \xrightarrow{H_2O} \mathbf{IIIa} + (AcOH + HCHO + 2H^+). \quad \text{(XIII)}$$

Thus, we have another argument in favor of the above assumption that the steady-state concentration of intermediate complex **VII** is low.

Reactions (VIII) and (IX) should be augmented with the following steps, parallel to them:

$$\mathbf{IV} + \mathbf{IIa} \rightleftarrows \underset{\mathbf{IX}}{MeS^+(CH_2SMe)CH_2OH} \quad \text{(XIV)}$$

$$\longrightarrow \mathbf{IIIa} + (HCHO + H^+).$$

This is clear from Table 2 and Fig. 2 (curve *4*): the product **IIIa** accumulation rate decreases with time, which can be attributed to the rapidly decreasing concentration of methylthiomethyl acetate **Ia**. The formation of **IIIa** after the disappearance of **Ia**, which accounts for nearly half the overall **IIIa** yield, can be attributed only to reaction (XIV). Relevant data from the cornerstone experiments of this study are presented in Table 2. The low rate of this secondary interaction (reaction (XIV)) can likely be explained in part by the considerable electrophilic solvation of intermediate semithioacetal **IIa**, which is more basic and more hydrophilic than substrate **Ia**. As a consequence, reaction (XIV), which is insignificant at the early stages of the interaction, is responsible for the completion of the process, even though it is slower than the reaction involving ester **Ia**. At the highest acidities examined, the yield of the final product decreases greatly by the time of complete **Ia** conversion (Table 2). Under these conditions, the reaction cannot be brought to completion within a reasonable period of time, because it almost stops. Intermediate acidities are optimal for obtaining bis(methylthio)methane in quantitative yield: nearly quantitative yields of the product are observed immediately after the disappearance of the initial substrate (Table 2), and extra time demands are insignificant. Sulfonium complex **IX**, like **VII**, is the most likely transition structure in this interaction.

The following experiments illustrate the reaction topology:

(1) Under the conditions of run 14 (Table 1, Fig. 2; see Experimental), after the complete disappearance of methylthiomethyl acetate, stirring was stopped and the acid phase and the organic phase were separated. A new portion of 53.8% sulfuric acid was added to the organic phase. Subsequent stirring did not change the amount of **IIIa** in the organic phase. At the same time, stirring the separated acid phase with a new portion of benzene caused the formation of an extra amount of product **IIIa**. The sum of this amount with the amount formed in the primary reaction mixture was close to the theoretical yield. Therefore, the conversion of hydrophilic semithioacetal **IIa** (steps (VI) and (VII)) took place in the water–acid phase of the system.

(2) When benzene in the initial mixture was replaced with *n*-alkane, the initial rate of **Ia** consump-

tion increased by a factor of 22–23 (see runs 1, 6, and 7 in Table 1; it is interesting that the rates of the noncatalytic acetylation of aniline and *meta*-chloroaniline with acetic anhydride increase by the same factor upon the same change of the solvent in the homogeneous system [18]). At high acidities, this replacement of the organic solvent greatly speeds up the consumption of ester **Ia** as well, but it is impossible to make precise measurement here because of the high reaction rate (see Experimental). Thus, firstly, steps (V) and (VIII)–(X) likely occur in the organic phase. Secondly, benzene, as distinct from alkanes, serves as a basic (or nucleophilic) solvent solvating species **V** and/or **IV**. The likely counterion is hydrosulfate. It is likely that the sulfate ion "coordinated" to two organic cationic species insignificantly reduces their lipophilicity (the formation of alkyl or dialkyl derivatives of sulfuric acid as species transferred across the interface cannot be ruled out either).

It is evident that, due to the rather low rate of the reaction, the heterogeneity factors (stirring, component distribution, and interfacial transfer) did not show themselves in this study.

The observed conversion of ester **Ia** in the acid-catalyzed process suggests that the mechanism of arene methylthiomethylation with this reagent [19] should be refined.

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