THE SOLVOLYSIS OF 2-HYDROXYPROPYL p-TOLUENESULFONATE AND ETHYLENE DI-p-TOLUENESULFONATE

by

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ABSTRACT

THE SOLVOLYSIS OF 2-HYDROXYPROPYL p-TOLUENESULFONATE AND ETHYLENE DI-p-TOLUENESULFONATE.

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Master of Science in Chemistry
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Little is known about the neighboring group effects of hydroxyl or sulfonate ester neighboring groups as applied to solvolysis at a primary carbon atom. This research was undertaken to obtain information bearing on these questions.

In the present study, the pseudo-unimolecular rate constants for the solvolysis of 2-hydroxypropyl p-toluenesulfonate were measured in ethanol-water mixtures ranging from 20% ethanol to 90% ethanol. Measurements were made at four different temperatures, and thermodynamic activation parameters were calculated. Correlations were drawn showing the effect of solvent on the rate of reaction. Using pyridine in equimolar concentration did not affect the rate of the reaction.

The rate of solvolysis of 2-hydroxypropyl p-toluenesulfonate was about four times greater than that of ethyl p-toluenesulfonate under the same conditions. After making the appropriate corrections for steric and inductive effects, the magnitude of the hydroxyl neighboring group effect was estimated. By this estimate it would appear that 2-hydroxy-

propyl *p*-toluenesulfonate solvolyzes 3000 times faster than would be expected in the absence of a neighboring group effect. A mechanism was proposed to account for this effect.

The rate constants for the solvolysis of ethylene di-*p*-toluenesulfonate were measured in ethanol-water mixtures ranging from 50% ethanol to 90% ethanol at two different temperatures. The rate of solvolysis of each *p*-toluenesulfonate group was estimated to be about 0.7 times that of ethyl *p*-toluenesulfonate. When considered in the light of steric and inductive effects, this demonstrates that the *p*-toluenesulfonate group has a significant neighboring effect. However, the magnitude of this effect could not be estimated.

It was established that 2-hydroxyethyl *p*-toluenesulfonate undergoes solvolysis faster than does ethylene di-*p*-toluenesulfonate, although the former compound could not be isolated for independent study. Mechanisms were proposed for the solvolysis of these compounds.

ACKNOWLEDGEMENTS

The author wishes to express his grateful appreciation to Dr. James A. Reeder for his suggestion of the problem and for the valuable guidance and encouragement throughout the course of this investigation.

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Finally, the author wishes to thank his friend, Shirish Mody, who helped to make the figures, and special thanks are due to his parents for the assistance and encouragement during the study when it was most needed.

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CHAPTER I

INTRODUCTION

The neighboring group effect in solvolysis reactions is well known in organic chemistry, and has been described in textbooks.[1-4] Basically it is an enhancement in the rate of solvolysis caused by participation of a neighboring group in the transition state for the rate determining step. Many organic substituents, such as thioether groups and acetoxy groups, have been found to give a large neighboring group effect.

The hydroxyl group has not been studied extensively as a neighboring group.[5] The base-catalyzed hydrolysis of halohydrins has been known for a long time and reported to proceed through an intermediate epoxide.[6]

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[1] J. Hine, Physical Organic Chemistry (New York: McGraw-Hill, 1962), pp. 141-151.

[2] E. M. Kosower, Physical Organic Chemistry (New York: John Wiley and Sons, Inc., 1968), pp. 103-143.

[3] E. S. Gould, Mechanism and Structure in Organic Chemistry (Holt, Rinehard and Winston, Inc., 1966), pp. 561-610.

[4] A. Streitwieser, Jr., Solvolytic Displacement Reactions (New York: McGraw-Hill, 1962), pp. 103-156.

[5] Hine, Physical Organic Chemistry, pp. 141-151.

[6] S. Winstein and H. Lucas, J. Am. Chem. Soc., 61, 1575 (1939).

Relatively little is known about the solvolysis of β-hydroxy halides and sulfonate esters in neutral media. Some work has been done on the neutral solvolysis of halohydrins[7-9] and β-methoxy sulfonate esters.[10]

In the neutral solvolysis of halohydrins it has been reported that the rearranged products are favored, owing to the presence of acid during the solvolysis.[11-13] The results were inconclusive regarding the neighboring group participation by hydroxyl groups.

The present research was undertaken to investigate the neighboring group participation by hydroxyl group in the solvolysis of primary sulfonate esters. For this purpose, a kinetic study was made of the solvolysis of 2-hydroxylpropyl p-toluenesulfonate and ethylene di-p-toluenesulfonate in various ethanol-water mixtures and at various temperatures.

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[7] G. J. Harvey, N. V. Riggs and V. R. Stimson, J. Chem. Soc., 3267 (1955).

[8] H. D. Cowan, C. L. McCabe and J. C. Warner, J. Am. Chem. Soc., 72, 1194 (1950).

[9] I. K. Gregor, N. V. Riggs and V. R. Stimson, J. Chem. Soc., 76 (1956).

[10] S. Winstein, C. R. Lindegren and L. L. Ingraham, J. Am. Chem. Soc., 75, 155 (1956).

[11] Harvey, Riggs and Stimson, J. Chem. Soc., 3267, (1955).

[12] Cowan, McCabe and Warner, J. Am. Chem. Soc., 72, 1194 (1950).

[13] Gregor, Riggs and Stimson, J. Chem. Soc., 76 (1956).

CHAPTER II

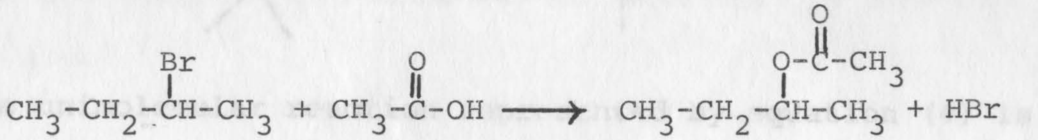
HISTORICAL

Solvolysis reactions are reactions involving solvent as a reagent, or reactions of compound with solvent. A common example of a solvolysis reaction is the hydrolysis of an alkyl halide or sulfonate ester, for example ethyl bromide or ethyl p-toluenesulfonate.

$$CH_3\text{-}CH_2\text{-}X + H_2O \longrightarrow CH_3CH_2OH + HX \qquad (1)$$

where $X = Br^-$ or $C_7H_7SO_3^-$.

Another example would be the acetolysis of 2-butyl bromide,



$$CH_3\text{-}CH_2\text{-}\overset{Br}{\overset{|}{CH}}\text{-}CH_3 + CH_3\text{-}\overset{O}{\overset{\|}{C}}\text{-}OH \longrightarrow CH_3\text{-}CH_2\text{-}\overset{O\text{-}\overset{O}{\overset{\|}{C}}\text{-}CH_3}{\overset{|}{CH}}\text{-}CH_3 + HBr \qquad (2)$$

In terms of their detailed mechanisms, solvolysis reactions are often classed as SN1 or SN2 reactions, a term introduced by Ingold[14] in 1933. The term SN1 means unimolecular nucleophilic substitution and SN2 means bimolecular nucleophilic substitution. These symbols have been used both for experimental properties of reactions and for particular reaction mechanisms. These mechanisms may be described by considering a substitution reaction:

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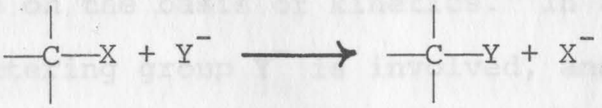
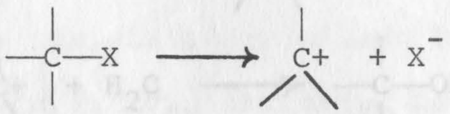
[14] E. D. Hughes, C. K. Ingold and C. S. Patel, J. Chem. Soc., 526 (1933).



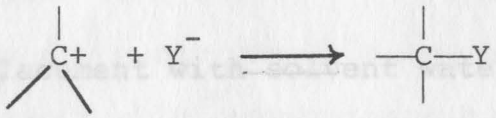
$$-\overset{|}{\underset{|}{C}}-X + Y^- \longrightarrow -\overset{|}{\underset{|}{C}}-Y + X^-, \qquad (3)$$

where X is a leaving group, e.g. halide or sulfonate ester, and Y^- is an entering group such as iodide or hydroxide ions. In general terms, the SN1 reaction may be considered to proceed by unimolecular dissociation of the alkyl halide or sulfonate ester to form a carbonium ion intermediate:



$$-\overset{|}{\underset{|}{C}}-X \longrightarrow C+ + X^- \qquad (4)$$

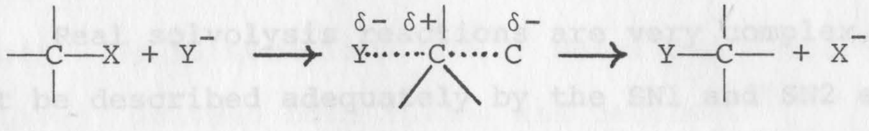
The carbonium ion then reacts with the entering group or with solvent, in a fast step:



$$C+ + Y^- \longrightarrow -\overset{|}{\underset{|}{C}}-Y \qquad (5)$$

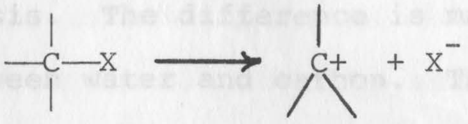
The unimolecular reaction represented by equation (4) is the rate determining step.

The SN2 reaction may be considered to proceed by direct displacement of the leaving group X by the entering group Y:

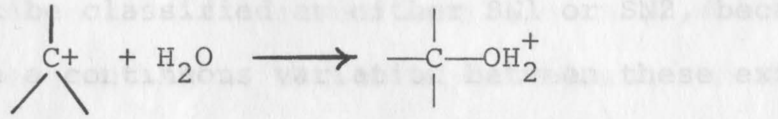


$$-\overset{|}{\underset{|}{C}}-X + Y^- \longrightarrow \overset{\delta-}{Y}\cdots\overset{\delta+}{C}\cdots\overset{\delta-}{C} \longrightarrow Y-\overset{|}{\underset{|}{C}}- + X^- \qquad (6)$$

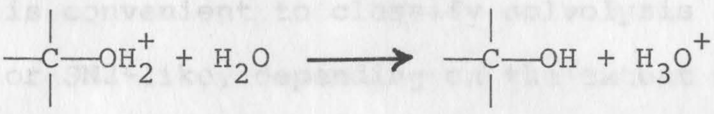
In the above cases, when a substitution reaction involves an entering group Y^-, the SN1 reaction is kinetically first-order and the SN2 reaction is kinetically second-order. Thus it is possible to distinguish between the two mechanisms on the basis of kinetics. In a solvolysis reaction no entering group Y^- is involved, and so it is not possible to distinguish kinetically between the two mechanisms. In an SN1 solvolysis, the intermediate carbonium ion reacts with the solvent. When the solvent is water this reaction may be shown as follows:



$$-\overset{|}{\underset{|}{C}}-X \longrightarrow C^{+} + X^{-} \qquad (7)$$

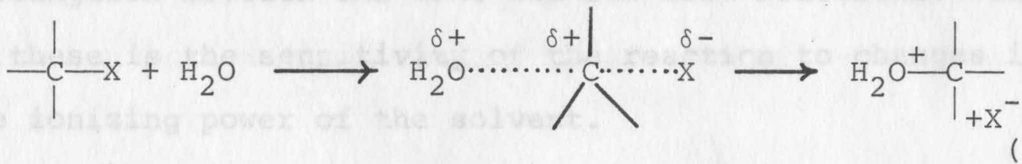


$$C^{+} + H_2O \longrightarrow -\overset{|}{\underset{|}{C}}-OH_2^{+} \qquad (8)$$

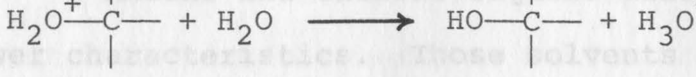


$$-\overset{|}{\underset{|}{C}}-OH_2^{+} + H_2O \longrightarrow -\overset{|}{\underset{|}{C}}-OH + H_3O^{+} \qquad (9)$$

The SN2 displacement with solvent water may be shown as follows:

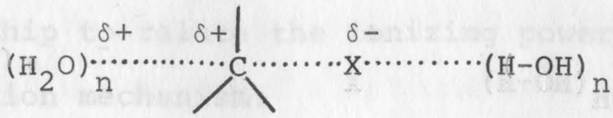


$$-\overset{|}{\underset{|}{C}}-X + H_2O \longrightarrow \overset{\delta+}{H_2O}\cdots\cdots\overset{\delta+}{C}\cdots\cdots\overset{\delta-}{X} \longrightarrow H_2\overset{+}{O}-\overset{|}{\underset{|}{C}}- + X^{-} \qquad (10)$$



$$H_2\overset{+}{O}-\overset{|}{\underset{|}{C}}- + H_2O \longrightarrow HO-\overset{|}{\underset{|}{C}}- + H_3O^{+} \qquad (11)$$

Real solvolysis reactions are very complex, and cannot be described adequately by the SN1 and SN2 approximations shown above. One reason is that carbonium ions must be extensively solvated in order to exist in solution. The anion of the leaving group X must also be extensively solvated. Thus the transition state for an SN1 hydrolysis may be represented crudely by the formula:



$$(H_2O)_n \cdots \overset{\delta+}{} \overset{\delta+}{C} \cdots \overset{\delta-}{X} \cdots (H\text{-}OH)_n$$

where n represents an unknown number of water molecules participating in the solvation of the carbonium ion and the X^- anion. This transition state is similar to that for an SN2 hydrolysis. The difference is mainly in extent of bonding between water and carbon. Thus, solvolysis reactions cannot be classified as either SN1 or SN2, because there can be a continuous variation between these extremes. However, it is convenient to classify solvolysis reactions as SN1-like or SN2-like, depending on the extent of solvent participation.

Since it is not possible to distinguish solvolysis mechanisms kinetically, other tests have been devised to distinguish between SN1-like and SN2-like reactions. One of these is the sensitivity of the reaction to changes in the ionizing power of the solvent.

Anions and cations require different solvating power characteristics. Those solvents having hydrogen bonding solvate anions well; solvents having unshared electron pairs solvate cations well. Water and alcohols can effectively solvate both anions and cations.[15]

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[15] Streitwieser, Solvolytic Displacement Reactions, pp. 561-610.

Winstein and Grunwald[16] have developed the following relationship to relate the ionizing power of the solvent to reaction mechanism:

$$\log \frac{k}{k_o} = mY \qquad (12)$$

where Y is the ionizing power of the solvent, k_o is the rate of reaction in 80% ethanol, and k is the rate of reaction at a particular solvent composition of solvating power Y. The quantity m is a constant characteristic of the compound being solvolyzed. When the m value is large, the reaction tends to proceed by an SN1-like mechanism, and when m value is small reaction tends to proceed by an SN2-like mechanism. Winstein[17] has defined m equal to 1.00 for the solvolysis of tertiary butyl chloride at 25°C in all solvents. The values of Y for particular solvent mixtures were then assigned empirically by studies of t-butyl chloride solvolysis.

It has been reported in many textbooks that β-substituted oxygen, nitrogen, or sulfur containing groups

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[16] E. Grunwald and S. Winstein, J. Am. Chem. Soc., 70, 846 (1946).

[17] S. Winstein and A. H. Fainberg, J. Am. Chem. Soc., 78, 2770 (1956).

increase the rate of solvolysis reactions.[18-21] The increases are due to the participation of neighboring groups. On the basis of the kinds of electrons available for neighboring group participation Kosower[22] classified neighboring groups in three major classes. The nonbonding electrons, or n electrons, are the most effective for neighboring group participation. Examples of nonbonding electrons are those on hydroxy, methoxy, and carboxy groups. The other two classes are π electrons in C=C bonds and σ electrons in C-H or C-C bonds.

There are three principal types of evidence which can indicate neighoring participation. If participation occurs during the rate determing step, the rate of reaction should be greater than that for a model compound which does not contain the neighboring group. It has been reported by Bohme and Sell[23] that the hydrolysis of the β-chlorosulfide $ClCH_2CH_2SC_2H_5$ is 10,000 times faster than the corresponding ether $ClCH_2CH_2OC_2H_5$ in aqueous dioxane. This

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[18] Hine, <u>Physical Organic Chemistry</u>, pp. 141-151.

[19] Kosower, <u>Physical Organic Chemistry</u>, pp. 103-143.

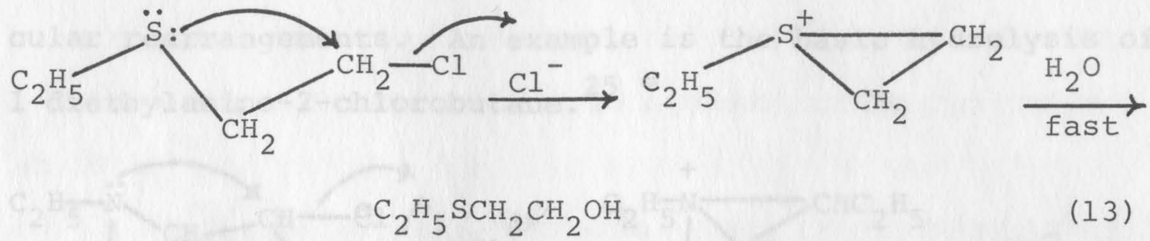
[20] Gould, <u>Mechanism and Structure</u>, pp. 561-610.

[21] Streitwieser, <u>Solvolytic Displacement</u>, pp. 103-156.

[22] Kosower, <u>ibid</u>.

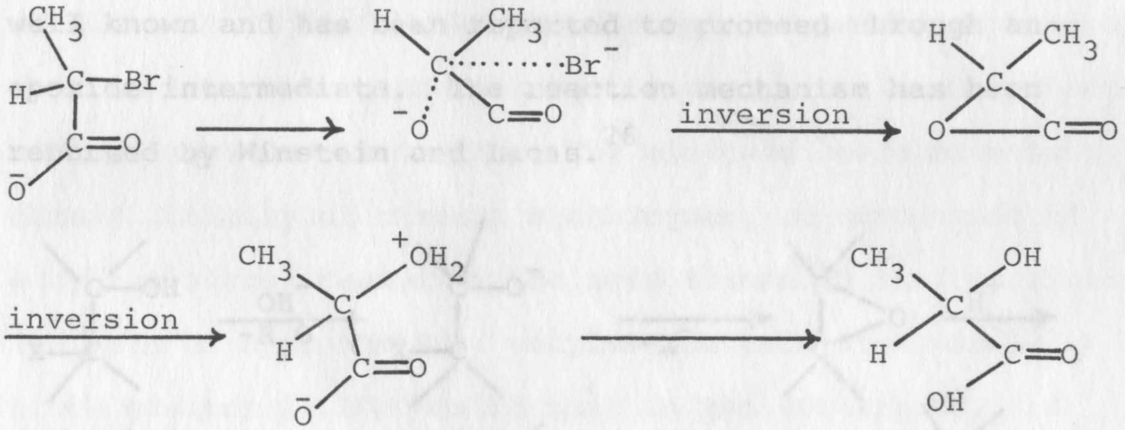
[23] H. Bohme and K. Sell, <u>Ber.</u>, <u>81</u>, 123 (1948).

faster rate was attributed to a cyclic intermediate:



(13)

A second type of evidence for neighboring participation can be obtained from stereochemistry. Hughes and Ingold[24] have reported that the basic hydrolysis of the α-bromopropionate anion proceeds with retention of configuration:



(14)

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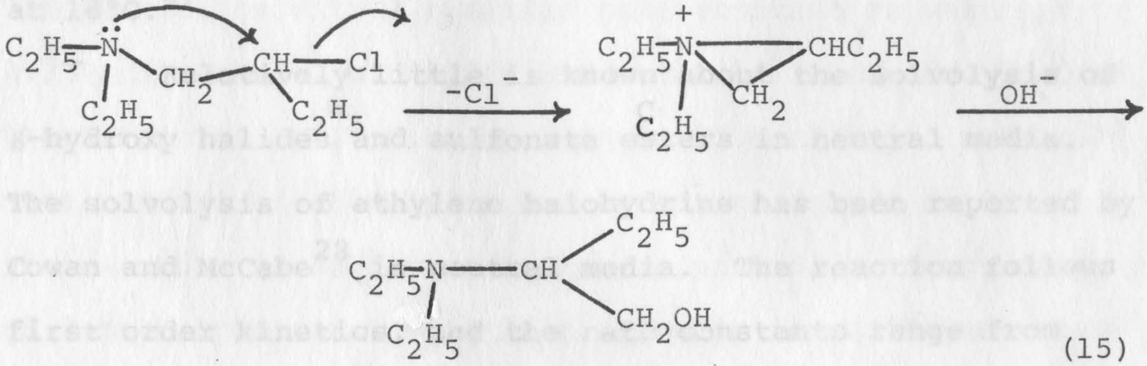
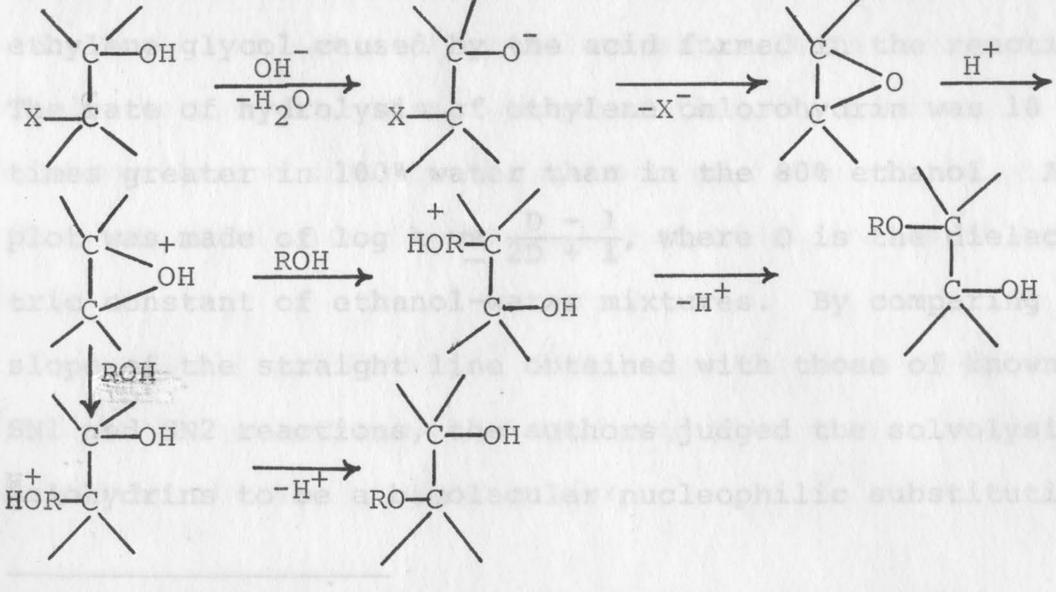
[24] E. D. Hughes, C. K. Ingold and W. A. Cowdrey, J. Chem. Soc., 1208 (1937).

A third type of evidence can be obtained from molecular rearrangements. An example is the basic hydrolysis of 1-diethylamino-2-chlorobutane.[25]



(15)

The base catalyzed hydrolysis of halo-hydrins is well known and has been reported to proceed through an epoxide intermediate. The reaction mechanism has been reported by Winstein and Lucas.[26]



(16)

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[25] S. D. Ross, J. Am. Chem. Soc., 69, 2982 (1947).

[26] Winstein and Lucas, ibid., 61, 1575 (1939).

The basic hydrolysis of ethylene chlorohydrin is 15 times slower than 1-chloropropane-2-ol and 252 times slower than 1-chloro-2-methylpropane-2-ol in aqueous sodium hydroxide at 18°C.[27]

Relatively little is known about the solvolysis of β-hydroxy halides and sulfonate esters in neutral media. The solvolysis of ethylene halohydrins has been reported by Cowan and McCabe[28] in neutral media. The reaction follows first order kinetics, and the rate constants range from 10^{-6} to 10^{-9} depending on the halogen and temperature. Ethylene iodohydrin gave a pseudo-unimolecular rate constant of $4.3 \times 10^{-6}\ \text{sec}^{-1}$ at 80°C. The major product in water at 70°C was ethylene glycol in 90-98% yield, but a small amount of acetaldehyde was formed. The aldehyde could have been formed directly or through a subsequent rearrangement of ethylene glycol caused by the acid formed in the reaction. The rate of hydrolysis of ethylene chlorohydrin was 10 times greater in 100% water than in the 80% ethanol. A plot was made of log k vs $\frac{D - 1}{2D + 1}$, where D is the dielectric constant of ethanol-water mixtures. By comparing the slope of the straight line obtained with those of known SN1 and SN2 reactions, the authors judged the solvolysis of halohydrins to be a bimolecular nucleophilic substitution.

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[27] Streitwieser, <u>Solvolytic Displacement</u>, pp. 103-156.

[28] Cowan, McCabe and Warner, <u>J. Am. Chem. Soc.</u>, <u>72</u>, 1194 (1950).

The solvolysis of 1-chloro-2-methylpropane-2-ol has been reported by Harvey, Riggs and Stimson[29] in aqueous ethanol. The reaction followed first order kinetics up to 75% completion with a specific rate constant in water of 2.33 x 10^{-5} $\sec^{-1}$ at 97.2°C and 6.9 x 10^{-6} $\sec^{-1}$ at 85.5°C. The Arrhenius activation energy was 27 kcal/mole. The ultimate product of solvolysis was isobutyraldehyde in 100% yield. This could have arisen by a pinacol rearrangement of 2-methylpropane-1,2-diol, previously formed in the reaction. The reaction in 100% water was 10 times faster than in 80% ethanol at 97°C. The authors reported that the solvolysis of 1-chloro-2-methylpropane-2-ol is 5.5 times faster than the solvolysis of ethylene chlorohydrin in water at 97°C. In the presence of .02 N hydrochloric acid, 2-methylpropane-1,2-diol was converted into isobutyraldehyde at about half the rate of the solvolysis. In the early stages of the solvolysis, it was observed that the production of acid was faster than the production of isobutyraldehyde. This was attributed to the formation of glycol as an intermediate.

Groll and Kautter[30] have reported that when 1-chloro-2-methyl-propane-2-ol is refluxed with water, isobutyraldehyde is obtained in good yield.

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[29] Harvey, Riggs and Stimson, J. Chem. Soc., 3267 (1955).

[30] H. P. A. Groll and C. T. Kautter, U. S. Patent 2,042,225; Chem. Abs., 30, 4872 (1936).

The solvolysis of 1-chloropropane-2-ol in water at 97°C has been reported by Gregor, Riggs and Stimson.[31] This reaction produced acetone and propylene glycol as the major products. It was observed that the production of glycol was faster than the production of acetone in the early stages of the reaction. The final products were propylene glycol (48%) and acetone (42%). The authors reported that propionaldehyde was not obtained in the solvolysis reaction. In the presence of moderately concentrated acid 1,2-propanediol did not rearrange to aldehyde or ketone. Such rearrangement was reported in the case of 2-methylpropane-1,2-diol. The solvolysis of 1-chloropropane-2-ol was 20 times faster in 100% water than in 80% ethanol at 97°C. This factor was 10 in the case of ethylenechlorohydrin and 1-chloro-2-methylpropane-2-ol. The rate constants for the solvolysis of 2-chloroethanol, 1-chloropropane-2-ol and 1-chloro-2-methylpropane-2-ol in 80% ethanol at 97°C were 3.8×10^{-7} sec^{-1}, 1.6×10^{-7} sec^{-1}, and 21.3×10^{-7} sec^{-1}, respectively.[32-34] The rate constants for the solvolysis of

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[31] Gregor, Riggs and Stimson, J. Chem. Soc., 76 (1956).

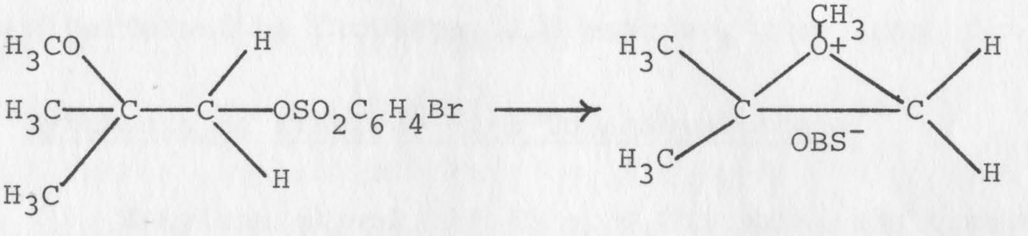
[32] Harvey, Riggs and Stimson, ibid., 3267 (1955).

[33] Cowan, McCabe and Warner, J. Am. Chem. Soc., 72 1194 (1950).

[34] Gregor, Riggs and Stimson, J. Chem. Soc., 76 (1956).

ethyl bromide in 50% ethanol at 95°C has been reported as 5.25×10^{-6} sec^{-1}.[35] Extrapolating the data of Cowan, et.al.[36] to the same conditions gives a rate constant of 1.90×10^{-5} sec^{-1} for 2-bromoethanol. Thus it would appear that the hydroxyl exerts a neighboring group influence over the reaction, although this was not considered by the authors.

Winstein and Lindegren[37] have reported the solvolysis of 2-methyl-2-methoxy-1-propyl-p-bromobenzenesulfonate in 80% dioxane at 99.73°C. The main product obtained from the solvolysis reaction was isobutyraldehyde in 72% yield. The reaction of 2-methyl-2-methoxy-1-propyl-p-toluenesulfonate was 15 times faster than neopentyl-p-toluenesulfonate. This was reported as evidence for methoxy participation in the rate determining ionization step. The suggested reaction mechanism for the formation of isobutyraldehyde was as follows:



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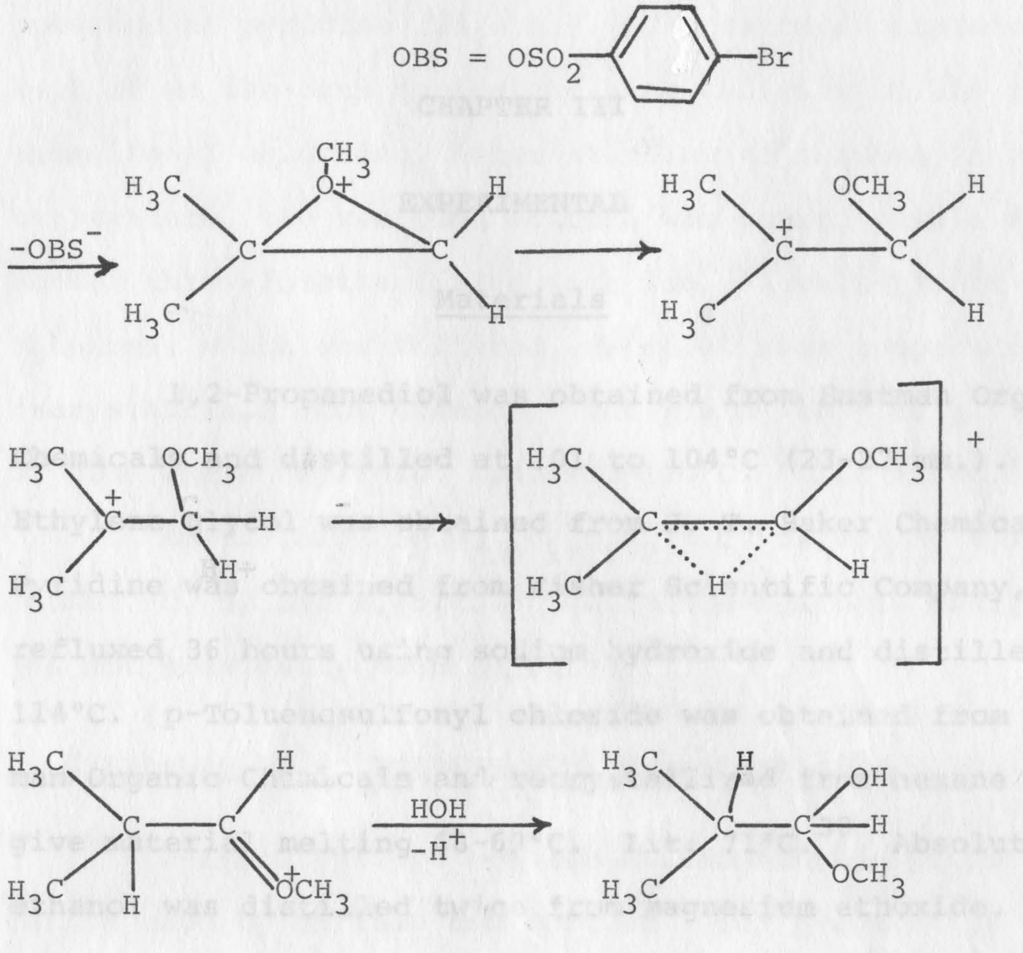
[35] S. Winstein, E. Grunwald and H. W. Jones, J. Am. Chem. Soc., 73, 2700 (1951).

[36] Cowan, McCabe and Warner, ibid., 72, 1194 (1950).

[37] Winstein, Lindegren and Ingraham, ibid., 75, 155 (1953).



$$\longrightarrow (CH_3)_2CH\text{-}CHO. \qquad (17)$$

CHAPTER III

EXPERIMENTAL

Materials

1,2-Propanediol was obtained from Eastman Organic Chemicals and distilled at 101 to 104°C (23-27 mm.). Ethylene glycol was obtained from J. T. Baker Chemicals. Pyridine was obtained from Fisher Scientific Company, refluxed 36 hours using sodium hydroxide and distilled at 114°C. p-Toluenesulfonyl chloride was obtained from Eastman Organic Chemicals and recrystallized from hexane to give material melting 68-69°C. Lit: 71°C.[38] Absolute ethanol was distilled twice from magnesium ethoxide. The initial and final fractions were rejected. The water used in kinetic studies was distilled twice from potassium permanganate in glass apparatus. Carbon-hydrogen analyses were performed by Crobaugh Laboratories, Cleveland, Ohio.

Synthesis of Ethylene di-p-Toluenesulfonate[39]

Ethylene glycol (14.45 g, 0.233 mole) was treated with 37.8 g. (0.198 mole) p-toluenesulfonyl chloride in the

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[38] Handbook of Chemistry and Physics, 1967-1968.

[39] C. L. Buttler, W. L. Nelson, A. G. Renfrew, J. Am. Chem. Soc., 57, 575 (1935).

presence of pyridine (25 c.c.). The reaction mixture was kept in an ice-bath during the slow addition of the p-toluenesulfonyl chloride. After standing 45 minutes at room temperature, the reaction mixture was poured into a 400 ml beaker three-fourths filled with ice. A white solid was obtained, which was filtered, dried at room temperature and recrystallized from ethanol. The yield was 9.56 g. (26.1%) of fine white crystals, m.p. 125-126°C. Lit: 126°C.[40]

Anal. Calc. for $C_{16}H_{18}O_6S_2$: C, 51.88; H, 4.90. Found: C, 52.39; H, 5.25. The carbon-hydrogen analysis has not previously been reported.

Synthesis of 2-Hydroxypropyl p-Toluenesulfonate

2-Hydroxypropyl p-toluenesulfonate was prepared by a method used by Dittmer and Hertler[41] for preparation of 3-hydroxypropyl p-toluenesulfonate. p-Toluenesulfonyl chloride (160.5 g, 0.84 mole) in pyridine (168 g.) was added during 90 minutes to stirred propylene glycol (128.2 g; 1.69 mole) in a 500 ml three-neck flask immersed in a dry ice-acetone bath. The reaction mixture was allowed to reach room temperature and stirring was continued for 30 minutes. The reaction mixture was poured into an 800 ml beaker which was three-fourths filled with ice. The oil thus obtained

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[40] Buttler, Nelson and Renfrew, _J. Am. Chem. Soc._, _57_, 575 (1935).

[41] D. C. Dittmer, W. R. Hertler, _ibid._, _79_, 4431 (1957).

was extracted into ether; washed with one liter of water, then with 500 ml of 3 M hydrochloric acid, and then with one liter of water. The ether solution was dried with magnesium sulfate and the ether was evaporated. The resulting heavy oil was kept in a freezer, and slowly crystallized to give an off-white solid.

The above solid was recrystallized three times from ether-pentane to give 82.5 g (42%) of white crystals, m.p. 50.5-52°C. Lit: 46°C.[42]

Anal. Calc. for $C_{10}H_{14}O_4S_2$: C, 52.16; H, 6.13. Found: C, 52.39; H, 6.44.

In the early part of this work the heavy oil did not give a solid, therefore it was chromatographed on silica gel using chloroform as the eluent. This gave a solid product which was used to seed crystallization.

In another experiment, the very slow addition of propylene glycol to p-toluenesulfonyl chloride in pyridine (reverse addition) below 0°C gave 71.0 g (44.5%) of propylene di-p-toluenesulfonate. This was recrystallized from ethanol-water to give white solid, m.p. 68-70°C. Lit: 68-69°C.[43]

Anal. Calc. for $C_{17}H_{20}O_6S_2$: C, 53.11; H, 5.24. Found: C, 53.42; H, 5.34. The carbon-hydrogen analysis

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[42] A. G. Renfrew, M. H. Green, and C. L. Buttler, J. Am. Chem. Soc., 61, 1783 (1939).

[43] J. H. Chapman and L. N. Owen, J. Chem. Soc., 579 (1950).

has been reported.[44]

Kinetic Measurements

Aqueous ethanol mixtures were prepared by volume at room temperature. Stock solutions were prepared from a quantity of 2-hydroxypropyl p-toluenesulfonate weighed into a 100 ml volumetric flask and made up to the mark with the required aqueous ethanol mixture.

Pipetted portions (10 ml) of the stock solution were added to 5-inch pyrex test tubes. The test tubes were protected by a calcium chloride drying tube, cooled in a dry ice-acetone bath, and sealed with a gas-oxygen flame. The tubes were then placed in a thermostatted oil bath. After a measured time (t), each tube was removed and cooled in a dry ice-acetone bath. The tubes were opened and the liberated acid was titrated with 0.00537 N. sodium hydroxide using phenolphthalein indicator.

Ethylene di-p-toluenesulfonate was not completely soluble in solvent mixture at room temperature, therefore a separate weighing was needed for each tube.

The thermostatted oil bath maintained the indicated temperature within ± 0.025°C. When the cold tubes were added initially at 105°C the bath required a maximum of ten minutes to reach temperature, depending on the number of cold tubes added.

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[44] Chapman and Owen, *J. Chem. Soc.*, 579 (1950).

The rates for both compounds at 84°C were obtained by refluxing 100 ml of .01 M sulfonate ester solutions in 50% ethanol. A separate 100 ml of solution was refluxed for each of the indicated times, and then titrated with 0.00605 N. sodium hydroxide to the phenolphthalein end point.

Calculations

In the case of 2-hydroxypropyl p-toluenesulfonate, the first order rate constants were calculated from the equation:

$$k_1 = \frac{2.303}{t} \log \frac{a}{a-x} \qquad (1)$$

This was done by plotting log (a-x) vs time t, where a is the initial concentration of 2-hydroxypropyl p-toluenesulfonate and x is the decrease in its concentration after time t, as calculated from the titration. Straight lines were obtained with slopes equal to $k_1/2.303$ where k_1 is the pseudo-unimolecular rate constant. The plots are shown in Figures 1-17, and the rate constants are given in Table 1.

The enthalpies of activation $\Delta H^{\ddagger}$ were calculated by plotting log k_1 vs 1/T where k_1 is the rate of reaction at the absolute temperature T. Straight lines were obtained with slopes equal to $\Delta H^{\ddagger}/2.303R$ where R is the gas constant. These are shown in Figures 18-20. The free energies of

TABLE 1.

SOLVOLYSIS OF 2-HYDROXYPROPYL-*p*-TOLUENESULFONATE FIRST-ORDER RATE CONSTANTS

Temp, °C.	% Ethanol	$k_1 \times 10^5 \text{ sec}^{-1}$	LOG k
80	30	3.35	-4.47
"	50	2.08	-4.68
"	70	1.26	-4.90
"	80	1.04	-4.98
84	50[b]	3.16	-4.50
"	50	3.38	-4.47
"	50[a,b]	3.20	-4.49
"	50[a]	3.39	-4.47
95	20	15.35	-3.81
"	40	11.34	-3.94
"	50	9.23	-4.03
"	50	8.39	-4.00
"	60	6.78	-4.17
"	70	5.64	-4.25
"	80	4.10	-4.39
105	50	26.48	-3.58
"	70	12.42	-3.90
"	80	10.70	-3.97
"	90	6.82	-4.17

Initial concentration 0.01 mole/liter unless otherwise noted.

[a] Initial concentration 0.014 mole/liter.

[b] Pyridine was used in equimolar concentration.

activation $\Delta F^{\ddagger}$ were calculated using the equation:

$$k_1 = \frac{kT}{h} e^{\frac{-\Delta F^{\ddagger}}{RT}} \qquad (2)$$

where k is the Boltzman constant, and h is the Planck constant. The entropies of activation $\Delta S^{\ddagger}$ were calculated using the relationship:

$$\Delta F^{\ddagger} = \Delta H^{\ddagger} - T\Delta S^{\ddagger} \qquad (3)$$

The values of the activation parameters are reported in Table 2.

In the case of ethylene di-p-toluenesulfonate, the first order rate constants were calculated by plotting log (2a-x) vs time t; log (a-x) vs time did not give straight line for this compound, as shown in Figure 28. Both tosylate groups in the molecule were solvolyzed at the end of the reaction.

During the study of ethylene di-p-toluenesulfonate at 84°C in 50% aqueous ethanol, the unreacted ethylene di-p-toluenesulfonate was recovered and weighed before each titration. Table 3 shows the percentage recovery.

TABLE 2.

SOLVOLYSIS OF 2-HYDROXYPROPYL *P*-TOLUENESULFONATE
ACTIVATION PARAMETERS

Temp, °C	% Ethanol	$k_1 \times 10^5 \sec^{-1}$	ΔF kcal.	ΔH kcal.	ΔS e.u.
80	50	2.08	28.36	26.95	-3.95
84[b]	"	3.16	28.39	"	-4.03
84	"	3.38	28.34	"	-3.89
95	"	9.23	28.50	"	-4.21
95	"	8.39	28.57	"	-4.40
105	"	26.48	28.50	"	-4.10
80	70	1.26	28.71	26.13	-7.30
95	"	5.64	28.86	"	-7.41
105	"	12.42	29.07	"	-7.77
80	80	1.04	28.85	25.30	-10.05
95	"	4.10	29.10	"	-10.32
105	"	10.70	29.18	"	-10.26

[b] Pyridine was used.

TABLE 3.

RECOVERY OF ETHYLENE DI-p-TOLUENE SULFONATE DURING SOLVOLYSIS IN 50% ETHANOL AT 84°C.

Time Minutes	Recovery Calculated[d]	Grams Found	Percentage Recovery
15	0.362	0.335	92.60
30	0.361	0.313	86.60
60	0.353	0.314	89.00
90	0.345	0.316	91.24
120	0.341	0.301	88.34
180	0.329	0.274	83.31
480	0.261	0.198	76.00
720	0.216	0.159	73.40

[d] From titration.

CHAPTER IV

RESULTS AND DISCUSSION

The Solvolysis of 2-Hydroxypropyl p-Toluenesulfonate

2-Hydroxypropyl p-toluenesulfonate has been synthesized previously from propylene glycol and p-toluenesulfonyl chloride,[45] but not purified sufficiently for kinetic studies. For the present study this compound was prepared by another method and purified by three recrystallizations.

The purpose of the present research was to investigate the neighboring group effect of hydroxyl in the 2-hydroxypropyl tosylate. (In this discussion the word "tosylate" is used to indicate a p-toluenesulfonate group.) The pseudo-unimolecular rate constants are shown in Table 1. The first-order kinetic plots are shown in Figures 1-17 for 0.01 mole/liter initial concentration and in Figures 30 and 31 for 0.0139 mole/liter initial concentration. Both initial concentrations gave the same rate constants within experimental error. The kinetics plots were good straight lines extending in most cases to 60-91% completion.

The addition of pyridine to the reaction mixture in a concentration equimolar to that of the tosylate did not

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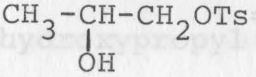
[45] Renfrew, Green and Buttler, J. Am. Chem. Soc., 61, 1783 (1939).

materially affect the rate of solvolysis, as shown by Figures 5 and 31.

The rates of solvolysis were studied in ethanol-water mixtures ranging from 20% ethanol to 90% ethanol and at temperatures of 80°, 84°, 95° and 105°C. The composition of solvent had some effect on the rates but this effect was not as great as that found by previous workers[46-48] for halohydrins. The Winstein-Grunwald mY correlations are shown in Figures 21-23 for temperatures of 80°, 95° and 105°C. The m values are shown in Table 4.

TABLE 4.

WINSTEIN-GRUNWALD m VALUES FOR THE SOLVOLYSIS OF 2-HYDROXYPROPYL P-TOLUENESULFONATE AND ETHYLENE DI-P-TOLUENESULFONATE

Compounds	Temp, °C.	LOG k_o	m
 $CH_3-\underset{OH}{CH}-CH_2OTs=$	80	-4.98	0.205
"	95	-4.39	0.203
"	105	-3.97	0.235
$TsOCH_2-CH_2OTs$	105	-4.42	0.215

Winstein[49] reported an m value of 0.25 for ethyl tosylate,

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[46] Harvey, Riggs and Stimson, J. Chem. Soc., 3267 (1955).

[47] Cowan, McCabe and Warner, J. Am. Chem. Soc., 72, 1194 (1950).

[48] Gregor, Riggs and Stimson, J. Chem. Soc., 76 (1956).

[49] S. Winstein, A. H. Fainberg, and E. Grunwald, J. Am. Chem. Soc., 79, 4146 (1957).

0.226 for ethylene bromohydrin,[50] both of which are assumed to solvolyze by an SN2 mechanism. Since the values in Table 4 are less than that for ethyl tosylate it may be assumed that carbonium ions are not intermediates in the reaction. The solvolysis of 2-hydroxypropyl tosylate therefore seems to be a direct displacement reaction.

The relationship of the rate constants to temperature is shown in Figures 18, 19 and 20. The activation parameters were calculated by standard means and are shown in Table 3. It should be noted that the entropies of activation are negative, as expected for transition states with solvent participation.

The absolute values of the pseudo-unimolecular rate constants are of prime interest in this study. At 95°C in 50% ethanol, the rate constant for the solvolysis of 2-hydroxypropyl tosylate was 8.39 - 9.23 x 10^{-5} $\sec^{-1}$ (Table 1, Figures 9, 10). The rate constants for solvolysis of ethyl tosylate under the same conditions have been reported to be 2.18 x 10^{-5} $\sec^{-1}$.[51] Thus 2-hydroxypropyl tosylate reacts about four times as fast as ethyl tosylate.

A direct study of propyl tosylate under the same conditions is lacking, but n-propyl bromide has been reported to react only 0.66 times as fast as ethyl bromide toward

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[50] Winstein and Fainberg, J. Am. Chem. Soc., 78, 2770 (1956).

[51] Grunwald and Winstein, ibid., 70, 846 (1946).

sodium thiophenolate in an SN2 reaction.[52] Thus it may be assumed that 2-hydroxypropyl tosylate should react about six times faster than propyl tosylate in the solvolysis reaction. Owing to the steric requirement of the hydroxyl group, 2-hydroxypropyl tosylate may be compared to isobutyl tosylate rather than to n-propyl tosylate. Isobutyl tosylate reacts only 0.046 times as fast as ethyl tosylate in 100% ethanol at 75°C.[53] Thus 2-hydroxypropyl tosylate should react about 29.4 times as fast as isobutyl tosylate under the same conditions.

From the above considerations, it may be seen that the hydroxyl group in 2-hydroxypropyl tosylate exerts a neighboring group effect which significantly enhances the rate of solvolysis. This effect is greater than the above figures would suggest, since the hydroxyl group also exhibits an electron-withdrawing inductive effect, which would tend to slow the rate of solvolysis. Winstein[54,55] has reported that a β-hydroxyl group reduces the rate of solvolysis by a

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[52] J. Hine and W. H. Brader, J. Am. Chem. Soc., 75, 3964 (1953).

[53] Streitwieser, Solvolytic Displacement, pp. 103-156.

[54] S. Winstein, E. Grunwald and L. L. Ingraham, J. Am. Chem. Soc., 70, 824 (1948).

[55] S. Winstein and E. Grunwald, J. Am. Chem. Soc., 70, 828 (1948).

factor of 10^2 where the neighboring group effect is eliminated. Thus, the neighboring hydroxyl group in 2-hydroxypropyl tosylate enhances the solvolysis rate by a factor of 3000. The arguments given above are similar to those advanced by Winstein[56] to show a similar neighboring group effect by methoxyl in 2-methyl-2-methoxy-1-propyl p-bromobenzenesulfonate.

Since we have established that the solvolysis of 2-hydroxypropyl p-toluenesulfonate has a solvent sensitivity similar to that of an SN2 reaction, it appears unlikely that carbonium ions are the intermediates. Indeed, Winstein[57-59] has shown that the neighboring group effect of hydroxyl is slight or negative in the case of secondary or tertiary halides where carbonium ions would be expected.

The following mechanism is proposed for the solvolysis of 2-hydroxypropyl tosylate: In the first step, which is rate-determining, the hydroxyl group removes tosylate by a direct displacement.

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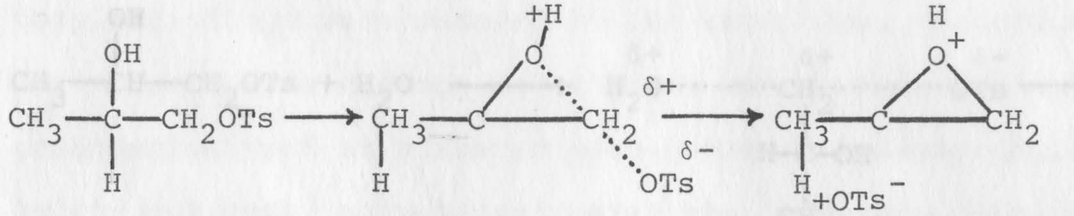
[56] Winstein, Lindegren and Ingraham, _J. Am. Chem. Soc._, _75_, 155 (1953).

[57] Winstein, Grunwald and Ingram, _ibid._, _70_, 824 (1948).

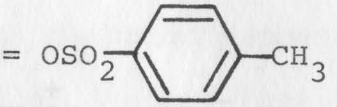
[58] Winstein and Grunwald, _ibid._, 70, 828 (1948).

[59] S. Winstein, E. Grunwald, R. E. Buckles, and C. Hanson, _J. Am. Chem. Soc._, _70_, 816 (1948).



(1)

where OTs =



The intermediate oxonium ion is a protonated epoxide, which may lose a proton to yield epoxide, react with water to yield propylene glycol, or rearrange to give acetone or propionaldehyde. Acetone has been reported as a major product in the hydrolysis of propylene chlorohydrins.[60] This product could have resulted from an acid-catalyzed rearrangement of the propylene glycol, although the authors found no such rearrangement in a model system. In the present work the products from the solvolysis of 2-hydroxypropyl tosylate were not determined, but it is suggested that such studies be made in the future. Since pyridine in equimolar concentration does not affect the rate of solvolysis these studies could be made in the presence of pyridine. If rearranged products were formed, this would confirm the presence of protonated epoxide as an intermediate.

It should be noted that an oxonium ion is also intermediate in a conventional SN2 solvolysis. The conventional solvolysis of 2-hydroxypropyl tosylate would proceed as follows:

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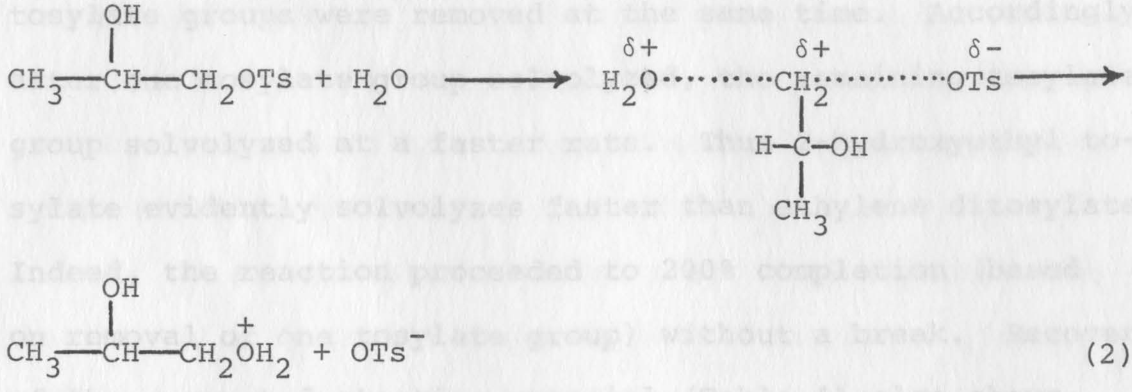
[60] Gregor, Riggs and Stimson, J. Chem. Soc., 76 (1956).



$$CH_3-\overset{\overset{\displaystyle OH}{|}}{CH}-CH_2OTs + H_2O \longrightarrow \overset{\delta+}{H_2O}\cdots\cdots \underset{\underset{\underset{\displaystyle CH_3}{|}}{\underset{\displaystyle H-C-OH}{|}}}{\overset{\delta+}{CH_2}}\cdots\cdots\overset{\delta-}{OTs} \longrightarrow$$

$$CH_3-\overset{\overset{\displaystyle OH}{|}}{CH}-CH_2\overset{+}{O}H_2 + OTs^- \qquad (2)$$

The rate enhancement in the present case may be because reaction (1) competes favorably with reaction (2). Since the transition state for reaction (1) probably involves less solvent than the transition state for reaction (2), one might suppose that the Winstein-Grunwald m value should be less than that for ethyl tosylate. The appears to be the case, as shown above.

The Solvolysis of Ethylene di-p-Toluenesulfonate

It was desired originally to prepare 2-hydroxyethyl tosylate and to compare its solvolysis rates with those of the propyl compound reported above. However, 2-hydroxyethyl tosylate is not reported in the literature and could not be prepared in this laboratory. It has been reported[61] that this compound is very unstable and disproportionates to form ethylene glycol and ethylene ditosylate.

The first-order kinetic plots are shown in Figures 24-27. Note that the plots were made by assuming that both

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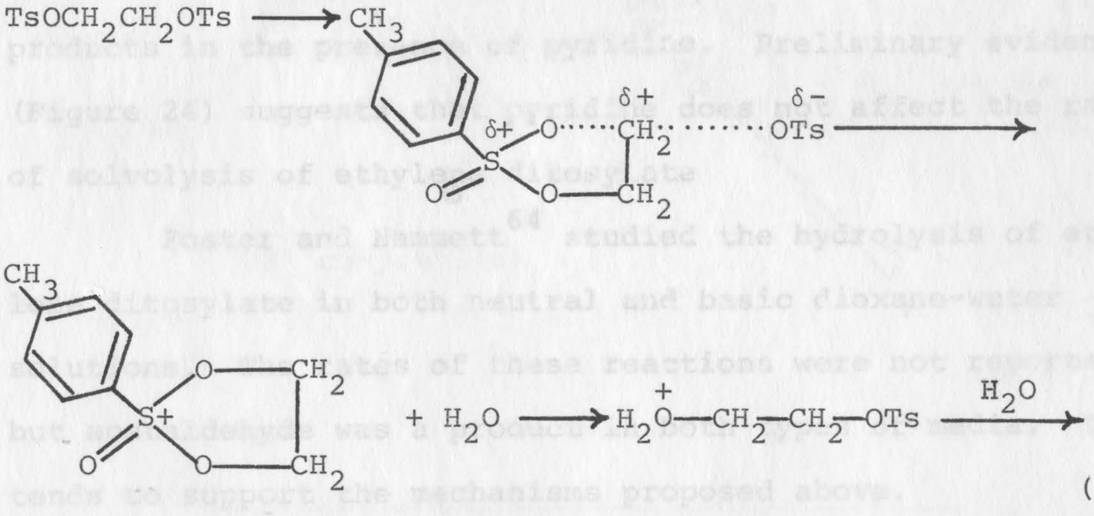
[61] Buttler, Nelson, and Renfrew, J. Am. Chem. Soc., 57, 575 (1935).

tosylate groups were removed at the same time. Accordingly, after one tosylate group solvolyzed, the remaining tosylate group solvolyzed at a faster rate. Thus 2-hydroxyethyl tosylate evidently solvolyzes faster than ethylene ditosylate. Indeed, the reaction proceeded to 200% completion (based on removal of one tosylate group) without a break. Recovery of the unreacted starting material (Table 4) also shows that the second tosylate group reacted faster than the first. The recovery closely matched the percentage reaction calculated by assuming the removal of two tosylate groups. Kinetic plots based on the removal of one tosylate group did not give straight lines (Figure 28).

The Winstein-Grunwald mY correlation for ethylene ditosylate (Figure 29 and Table 2) is consistent with SN2 solvolysis, as was the case for 2-hydroxypropyl tosylate. The activation parameters were not measured for this reaction.

The absolute values of the reaction rates are of primary interest. At 84°C in 50% ethanol, ethylene ditosylate solvolyzed 0.361 times as fast as 2-hydroxypropyl tosylate under the same conditions. We have shown that 2-hydroxypropyl tosylate solvolyzes four times faster than ethyl tosylate at 95°C in 50% ethanol. Thus is appears that ethylene ditosylate would solvolyze about 1.4 times faster than ethyl tosylate under the same conditions. Since two tosylate groups are present, each tosylate would react 0.7 times as fast as ethyl tosylate. This would appear to provide evidence for a neighboring group participation by the tosylate group, especially when one considers that the tosylate group must have a large steric requirement and a large electron-withdrawing inductive effect.

It was established above that 2-hydroxyethyl tosylate solvolyzes faster than ethylene ditosylate. Since ethylene ditosylate solvolyzes faster than ethyl tosylate, the hydroxyl group has a neighboring group effect. However, it is not possible to estimate the magnitude of this effect in 2-hydroxyethyl tosylate. The proposed mechanism for the solvolysis of ethylene ditosylate is as follows:

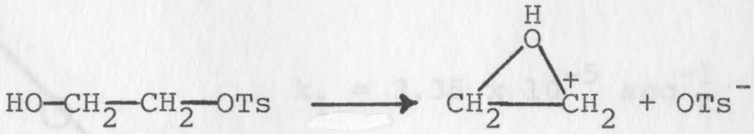


(3)

$$HO{-}CH_2{-}CH_2{-}OTs + H_3O^+ \qquad (4)$$

The postulated intermediate in reaction (3) is an oxosulfonium cation. Stable trimethyloxosulfonium salts have been described by Corey[62] and by Kuhn.[63]

The 2-hydroxyethyl tosylate formed in reaction (4) would then react in a manner similar to that of 2-hydroxypropyl tosylate:


(5)

The postulated intermediate is protonated ethylene oxide, which could yield ethylene oxide, acetaldehyde or ethylene glycol as products. The solvolysis products of ethylene ditosylate have not been investigated in the present study. It is recommended that future studies be made of these products in the presence of pyridine. Preliminary evidence (Figure 24) suggests that pyridine does not affect the rate of solvolysis of ethylene ditosylate

Foster and Hammett[64] studied the hydrolysis of ethylene ditosylate in both neutral and basic dioxane-water solutions. The rates of these reactions were not reported, but acetaldehyde was a product in both types of media. This tends to support the mechanisms proposed above.

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[62] E. J. Corey and M Chaykovsky, J. Am. Chem. Soc., 87, 1353 (1965).

[63] R. Kuhn and H. Trischmann, Ann., 611, 117 (1958).

[64] F. C. Foster and L. P. Hammett, J. Am. Chem. Soc., 68, 1736 (1946).

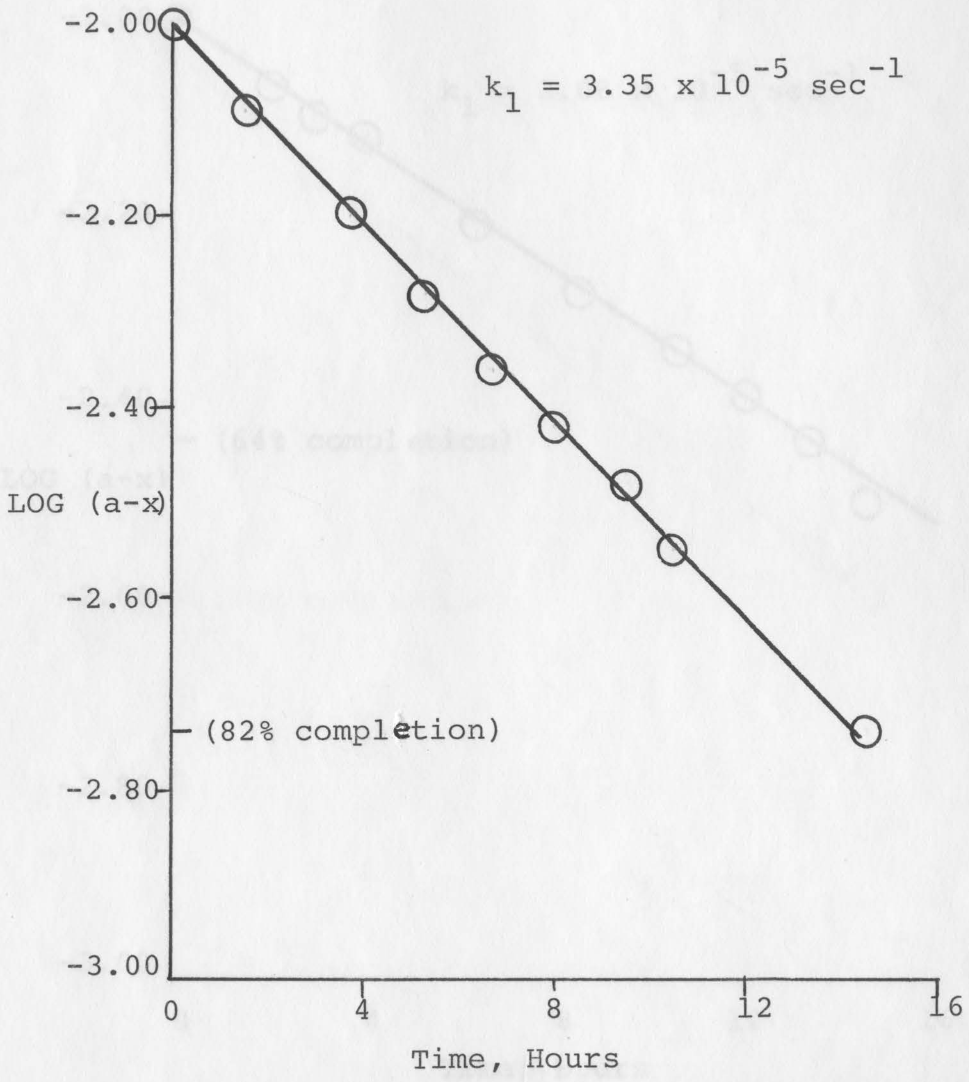

Fig. 1. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 30% ethanol at 80°C. Initial concentration 0.01 mole /liter.

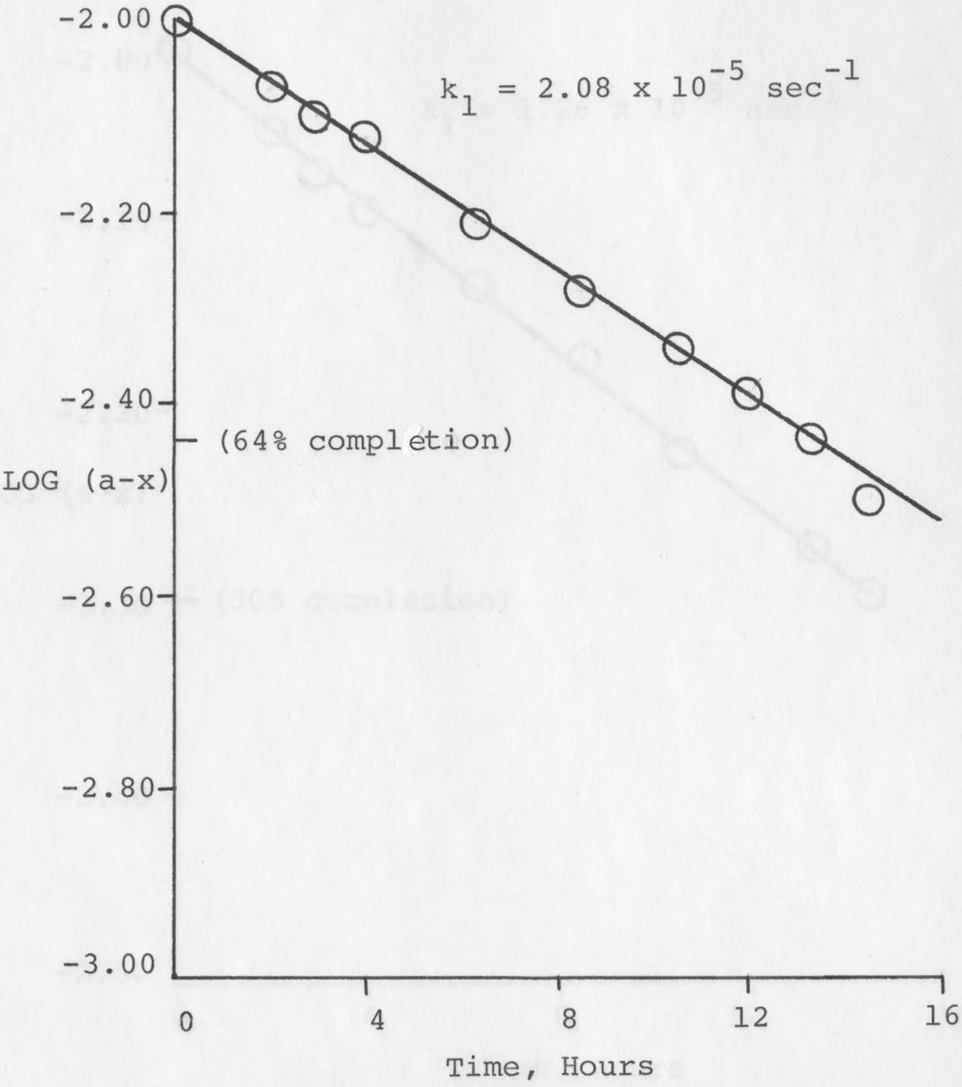

Fig. 2. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 50% ethanol at 80°C. Initial concentration 0.01 mole./liter.

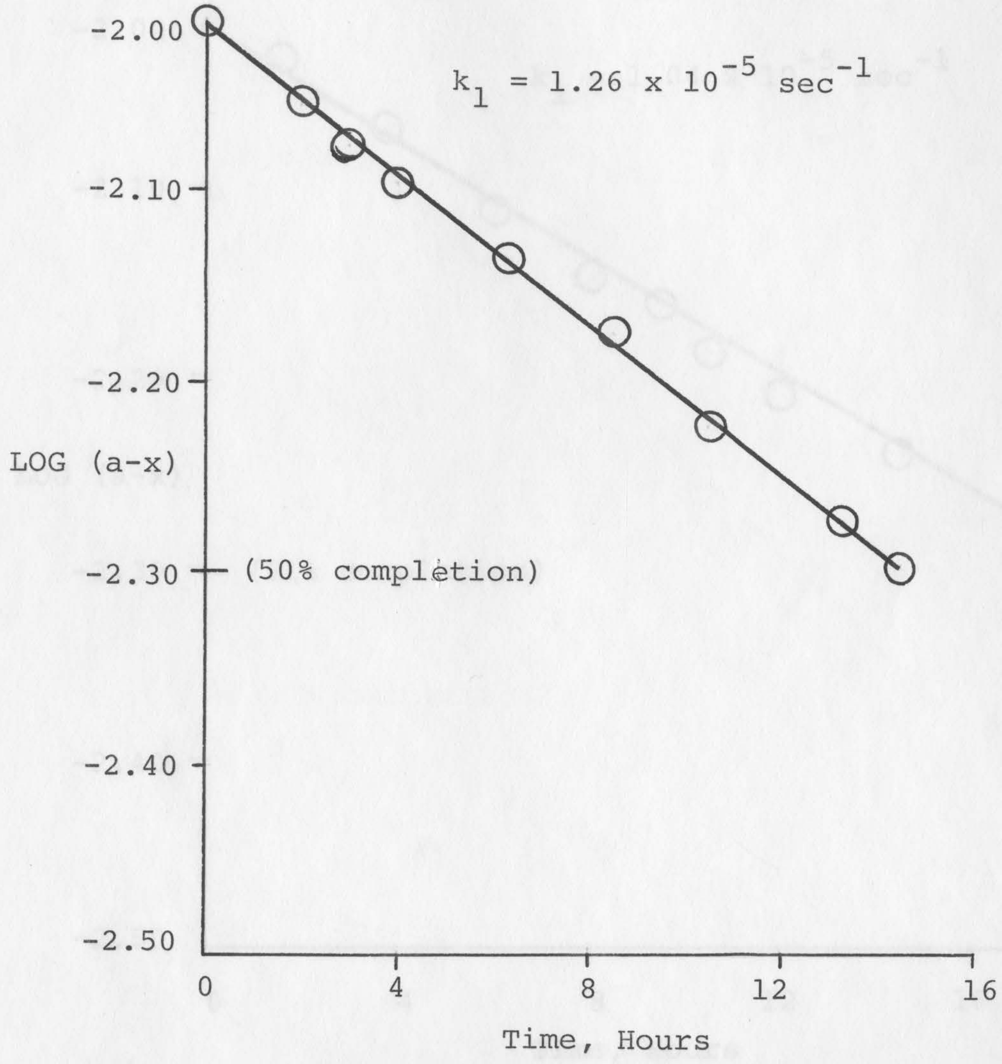

Fig. 3. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 70% ethanol at 80°C. Initial concentration .01 mole/liter.

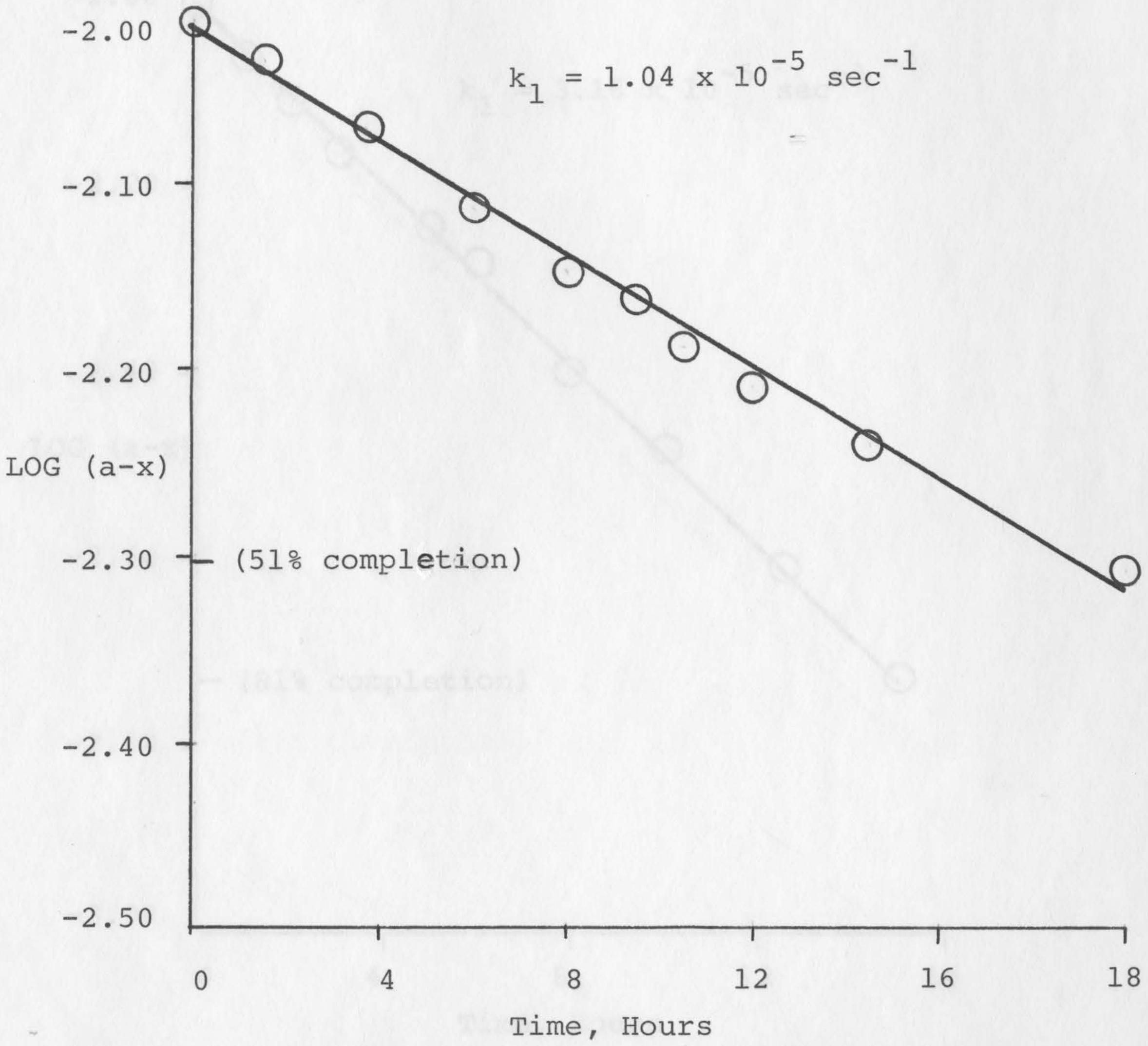

Fig. 4. Solvolysis rate of 2-hydroxypropyl p-toluene sulfonate in 80% ethanol at 80°C. Initial concentration .01 mole/liter.

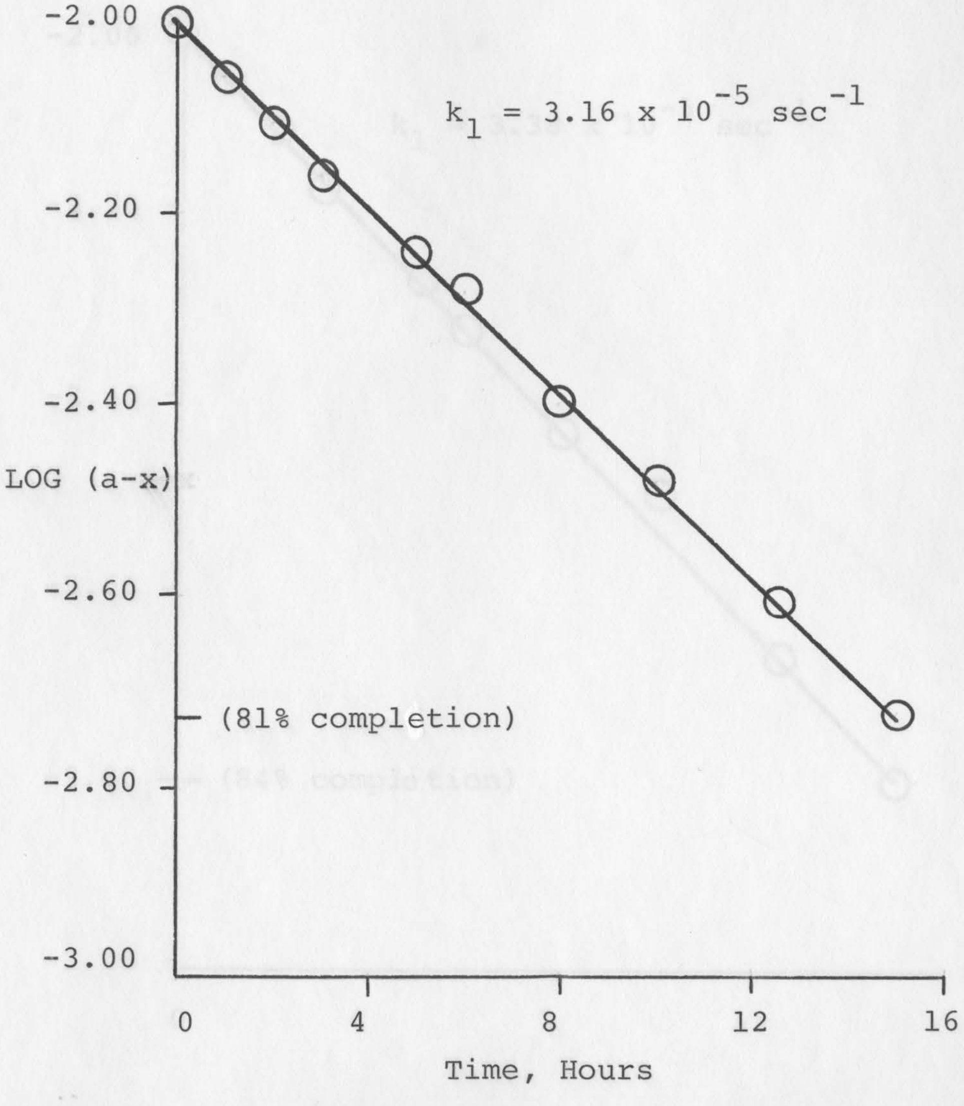

Fig. 5. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 50% ethanol at 84°C. Initial concentration .01 mole /liter. .01 mole /liter pyridine was used.

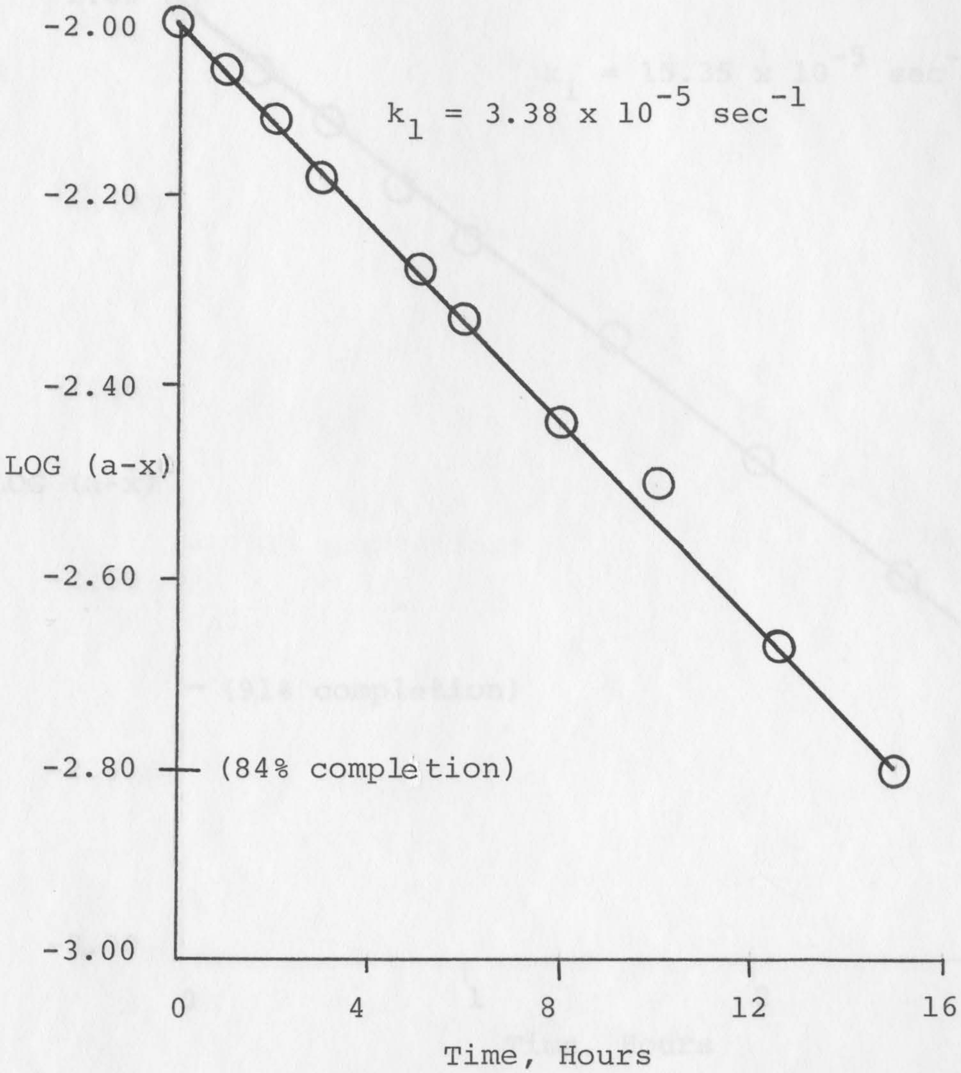

Fig. 6. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 50% ethanol at 84°C. Initial concentration .01 mole/liter.

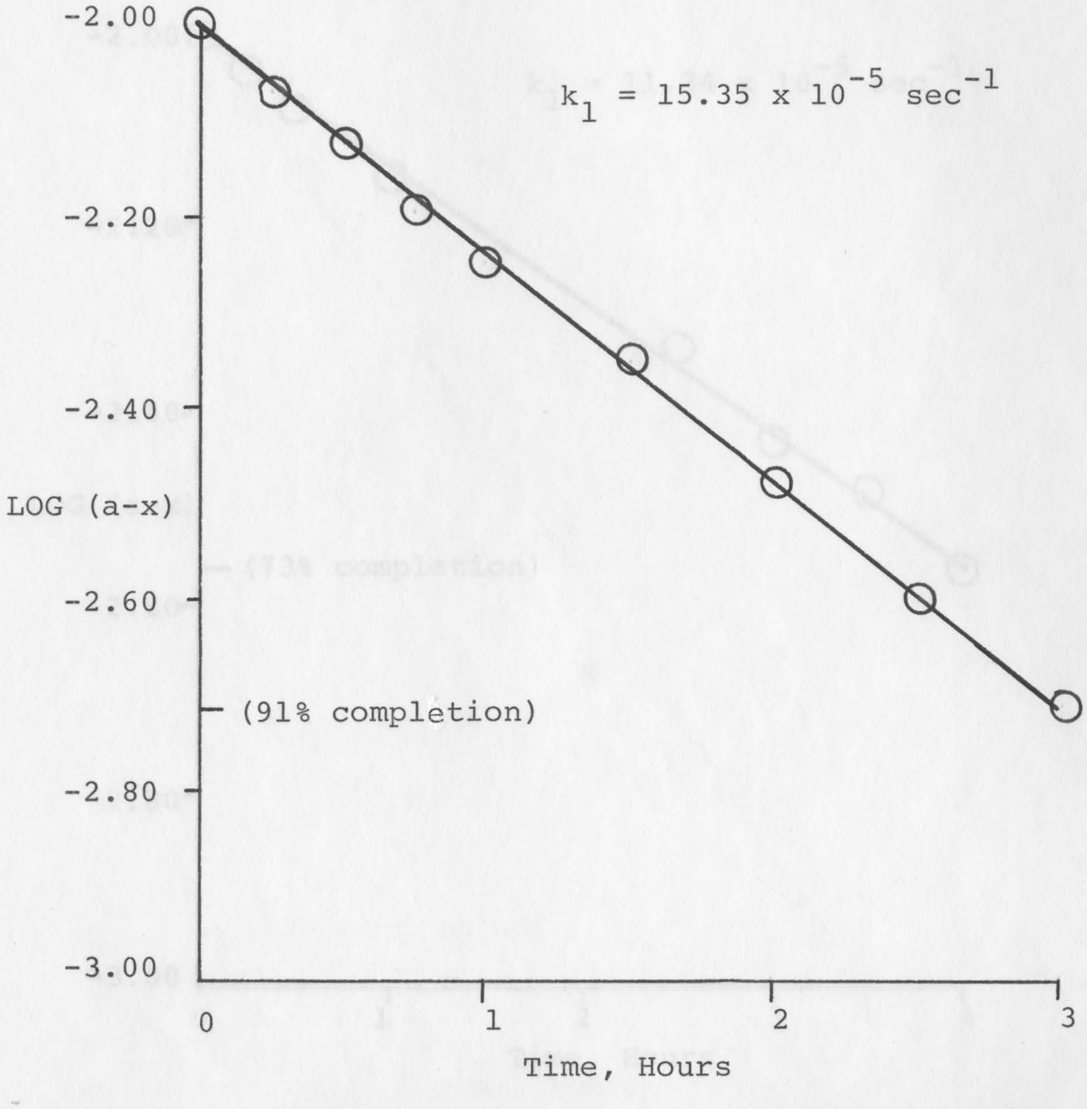

Fig. 7. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 20% ethanol at 95°C. Initial concentration .01 mole/liter.

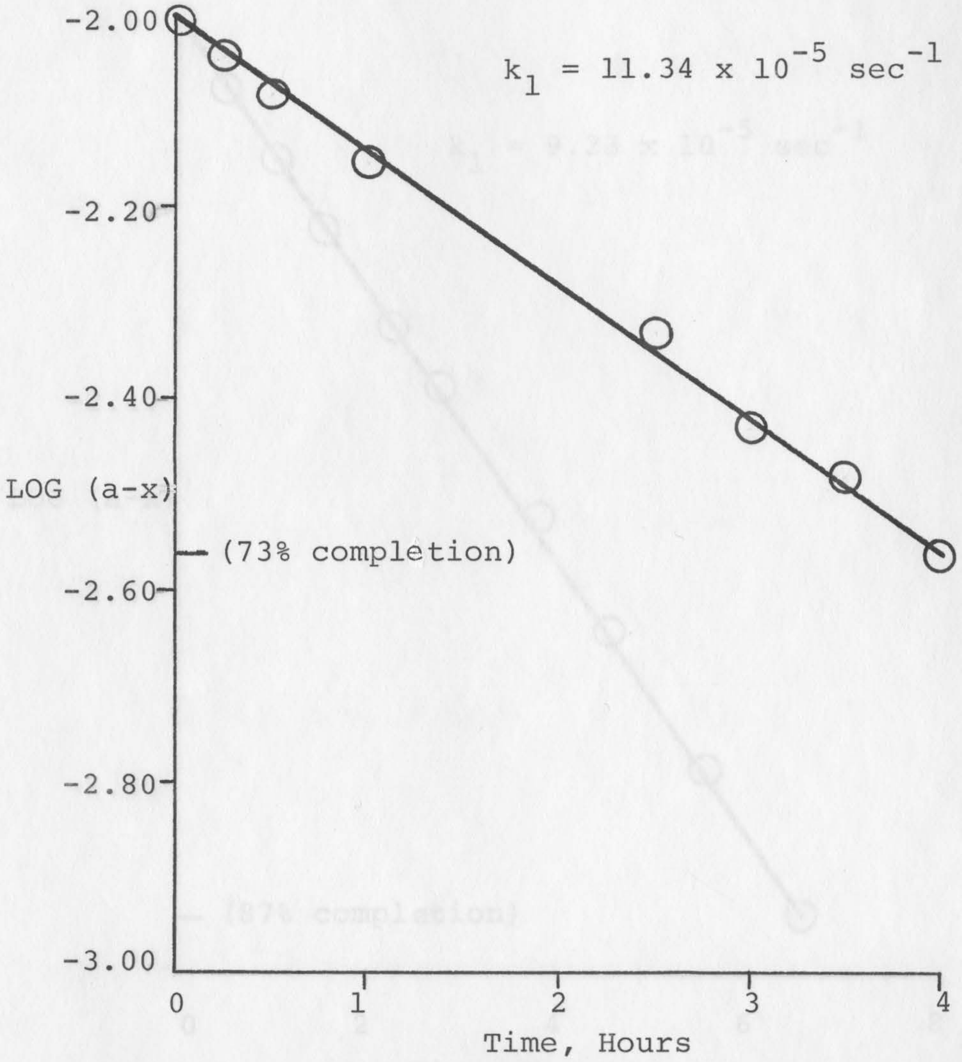

Fig. 8. Solvolysis rate of 2-hydroxypropyl-*p*-toluene sulfonate in 40% ethanol at 95°C. Initial concentration .01 mole /liter.

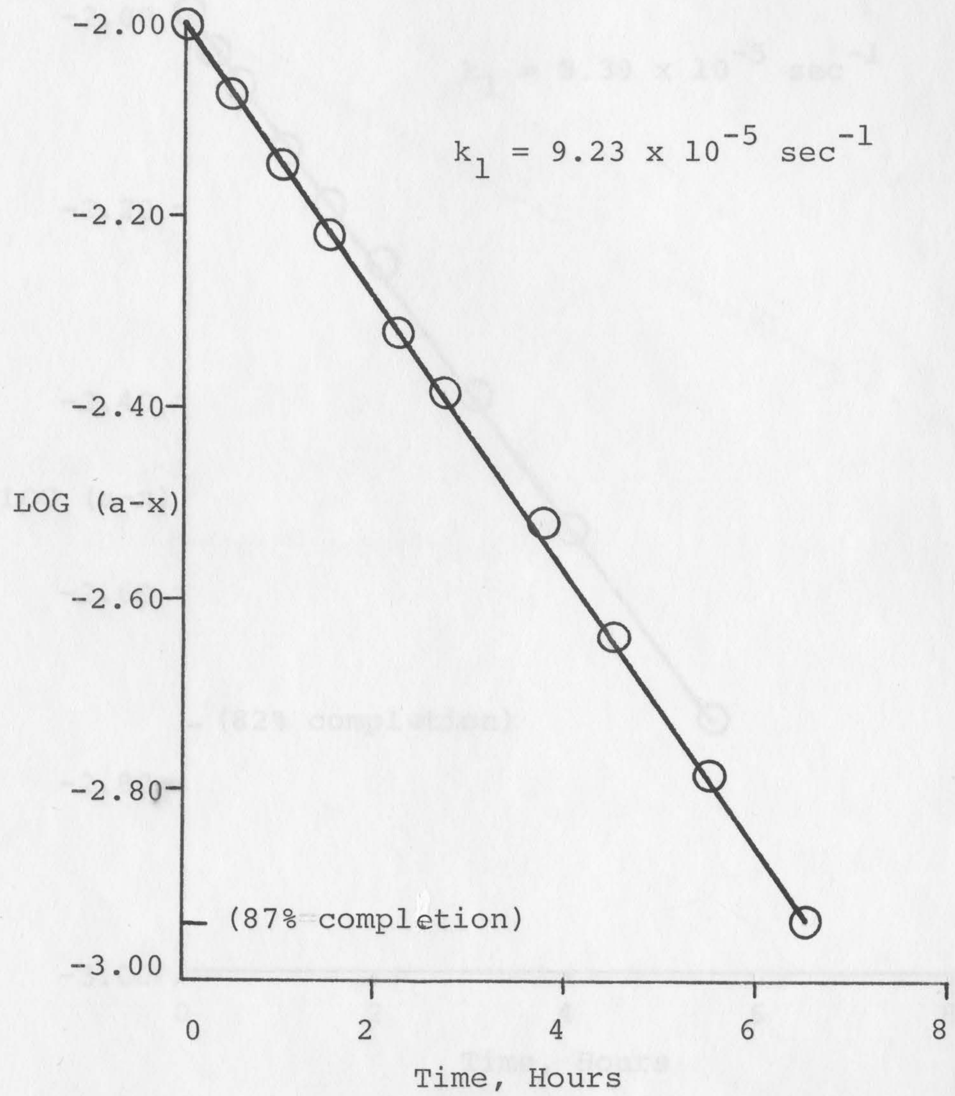

Fig. 9. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 50% ethanol at 95°C. Initial concentration .01 mole/liter.

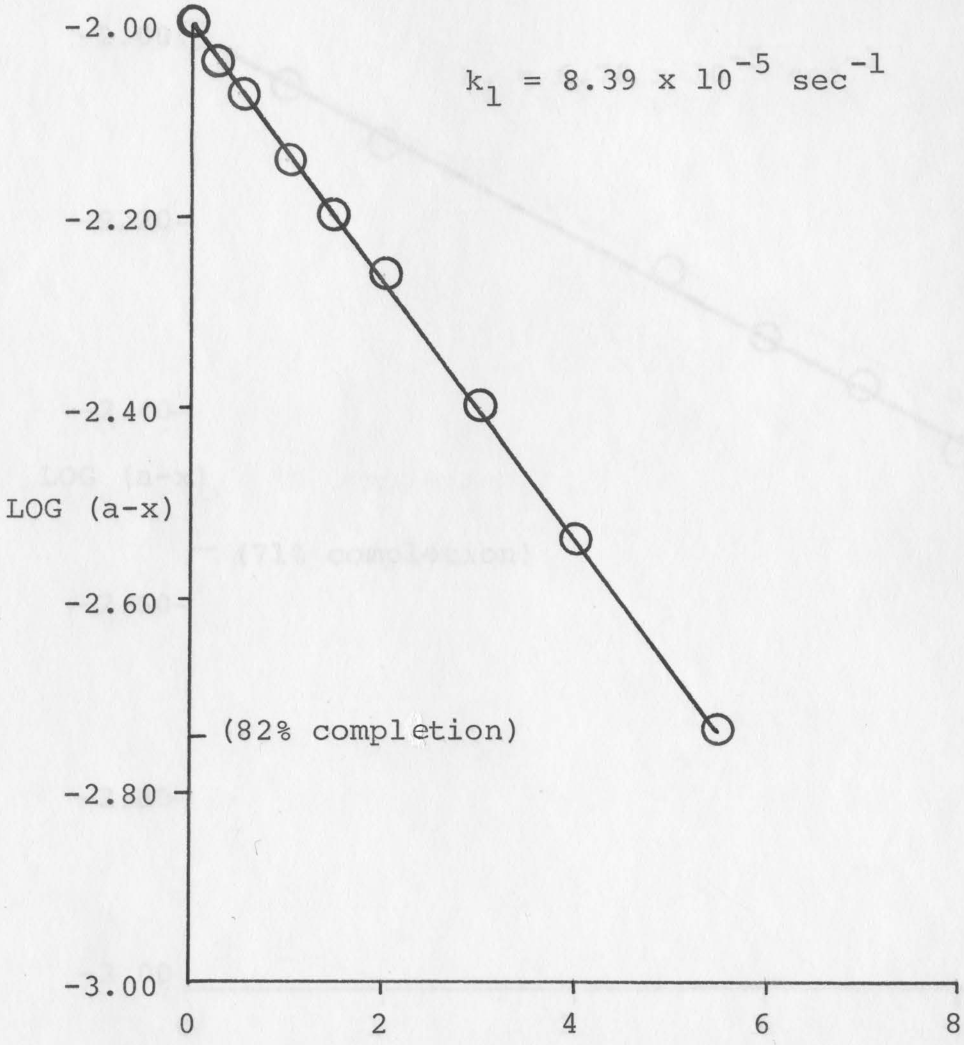

Fig. 10. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 50% ethanol at 95°C. Initial concentration .01 mole/liter.

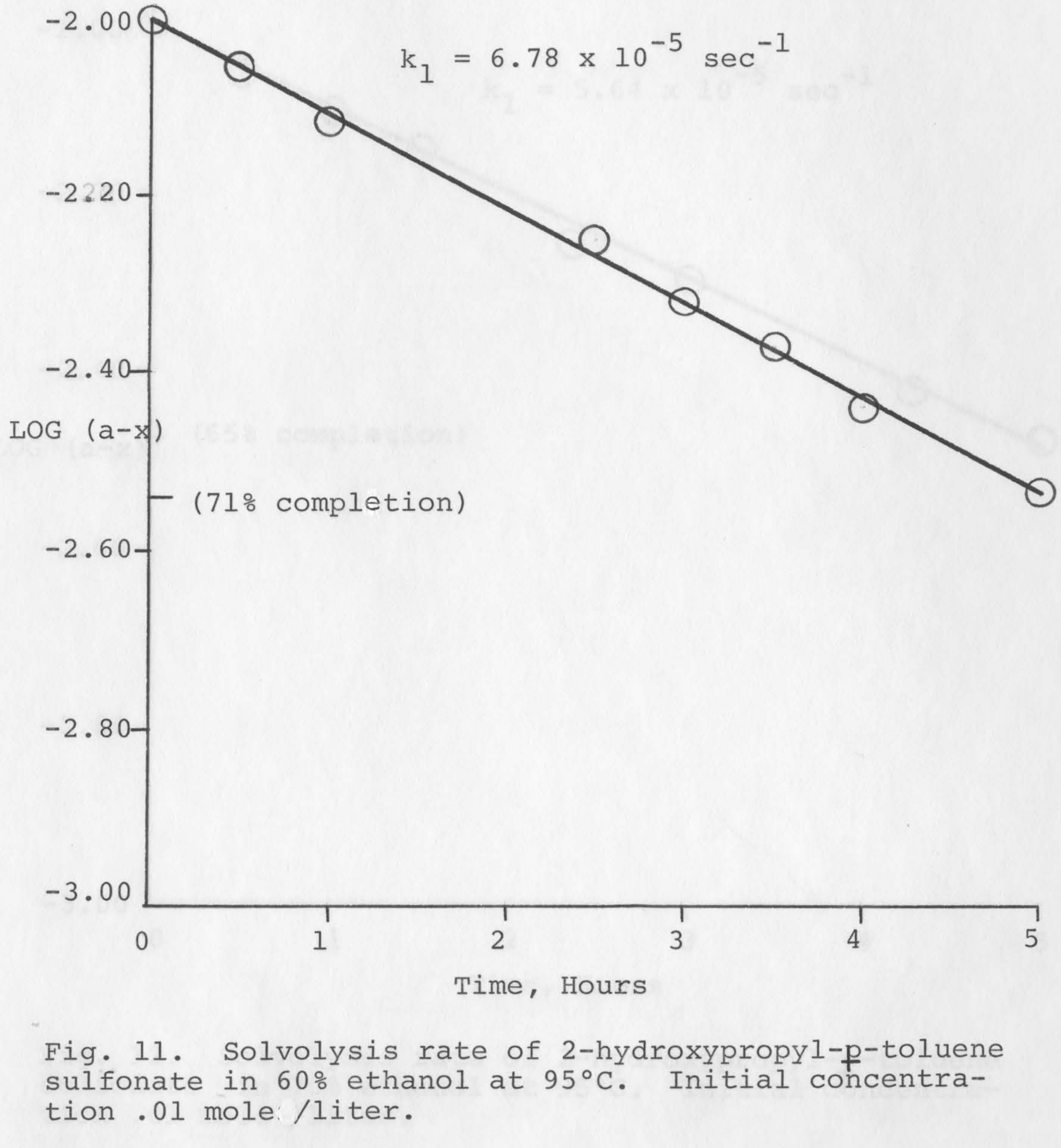

Fig. 11. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 60% ethanol at 95°C. Initial concentration .01 mole /liter.

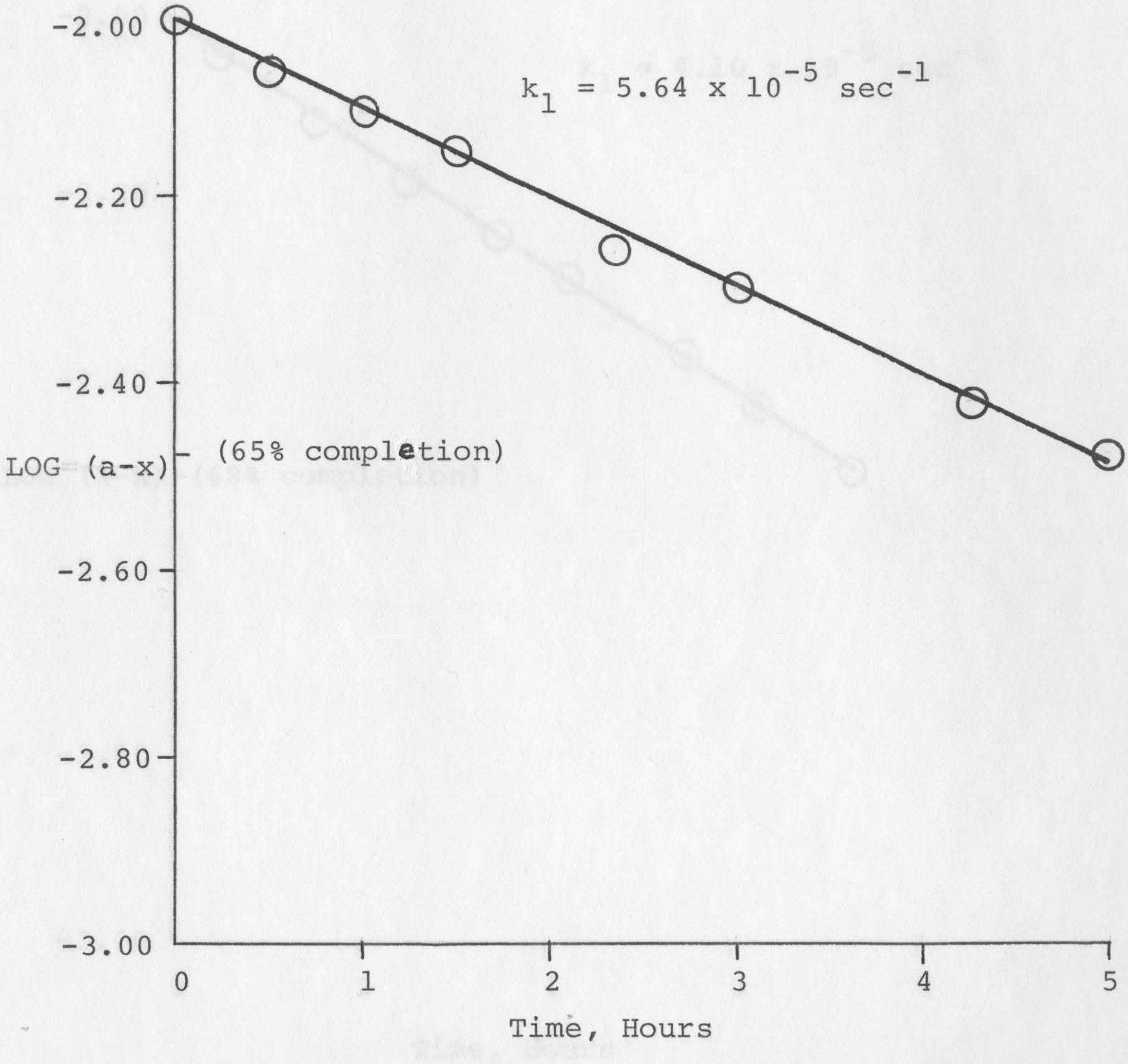

Fig. 12. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 70% ethanol at 95°C. Initial concentration .01 mole/liter.

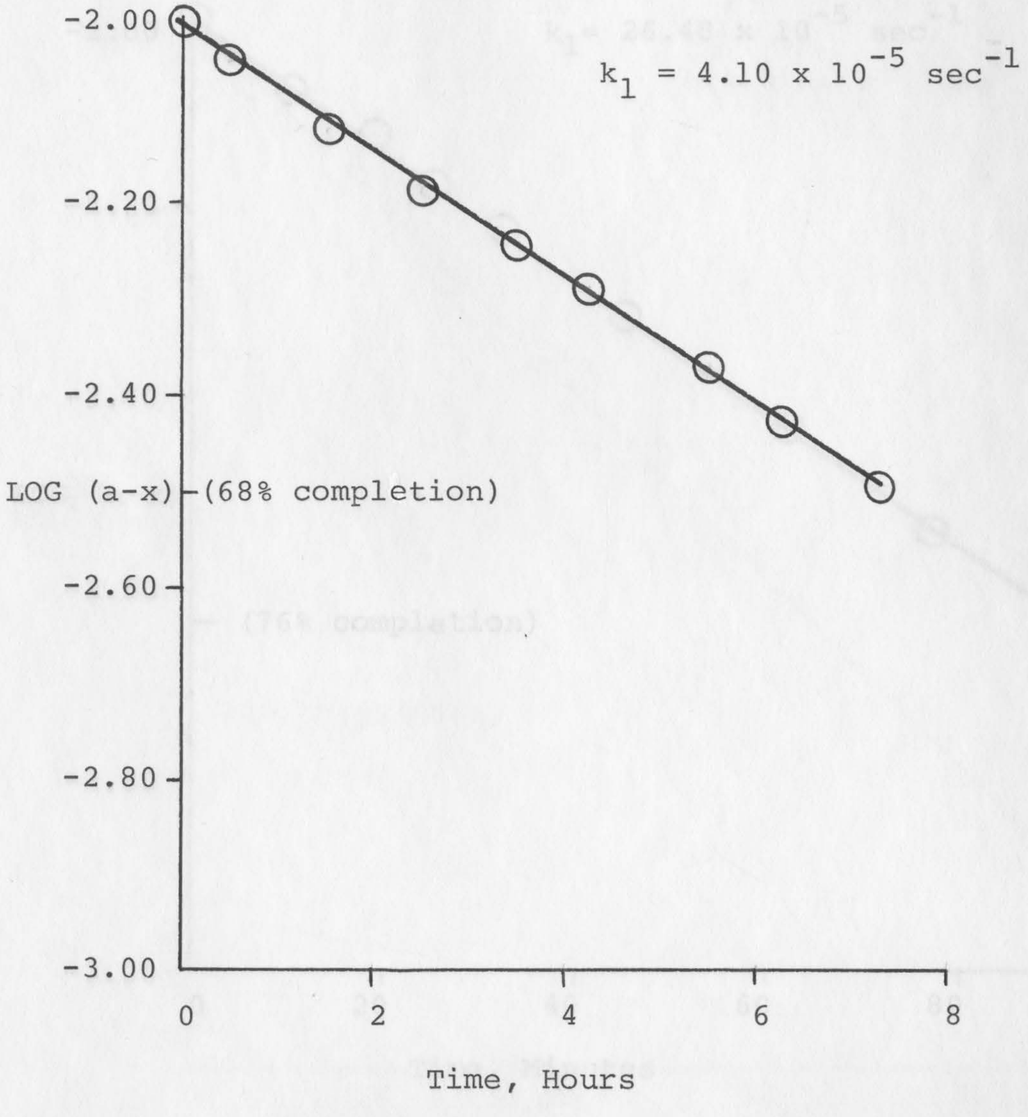

Fig. 13. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 80% ethanol at 95°C. Initial concentration .01 mole /liter.

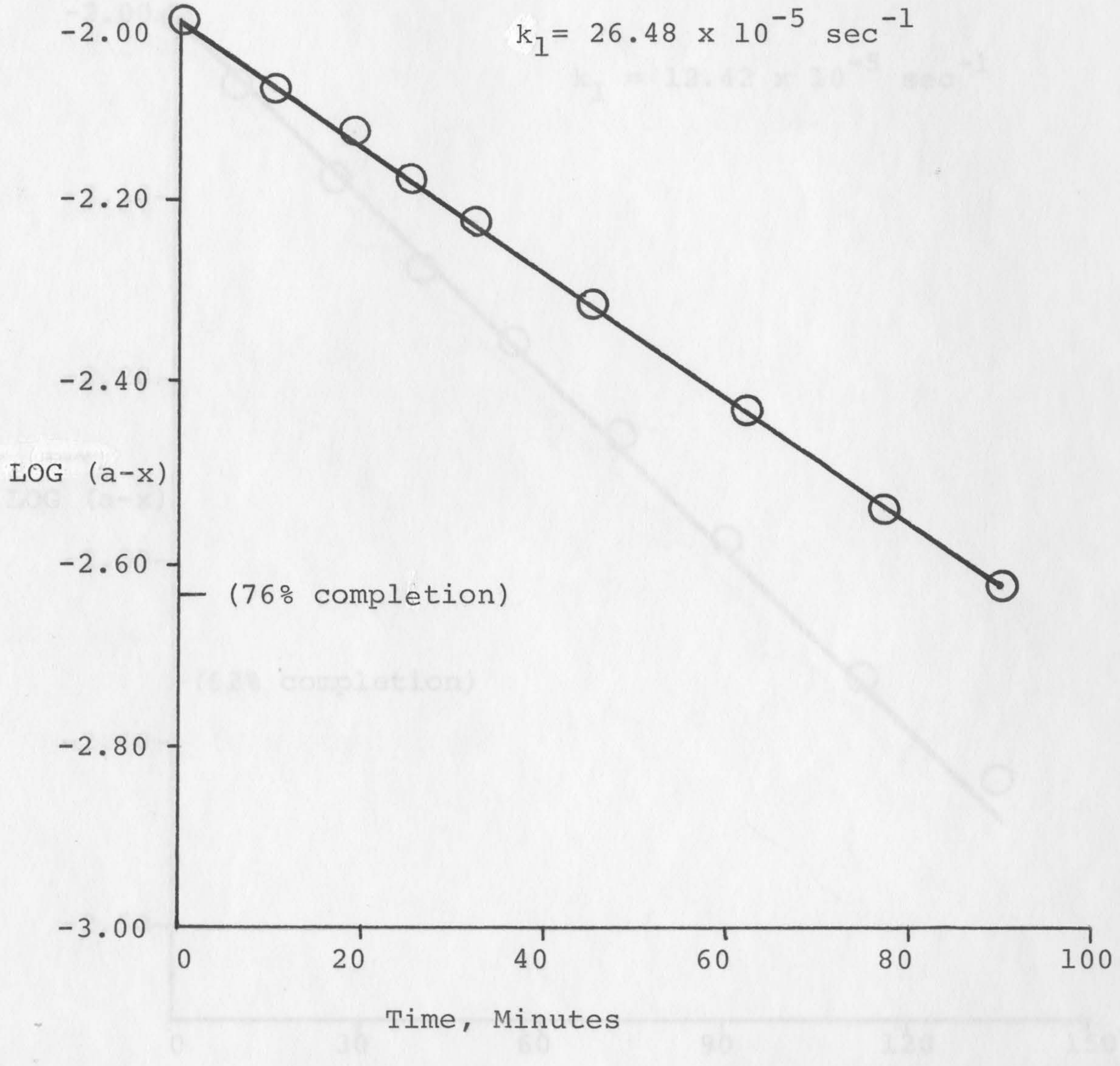

Fig. 14. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 50% ethanol at 105°C. Initial concentration .01 mole /liter.

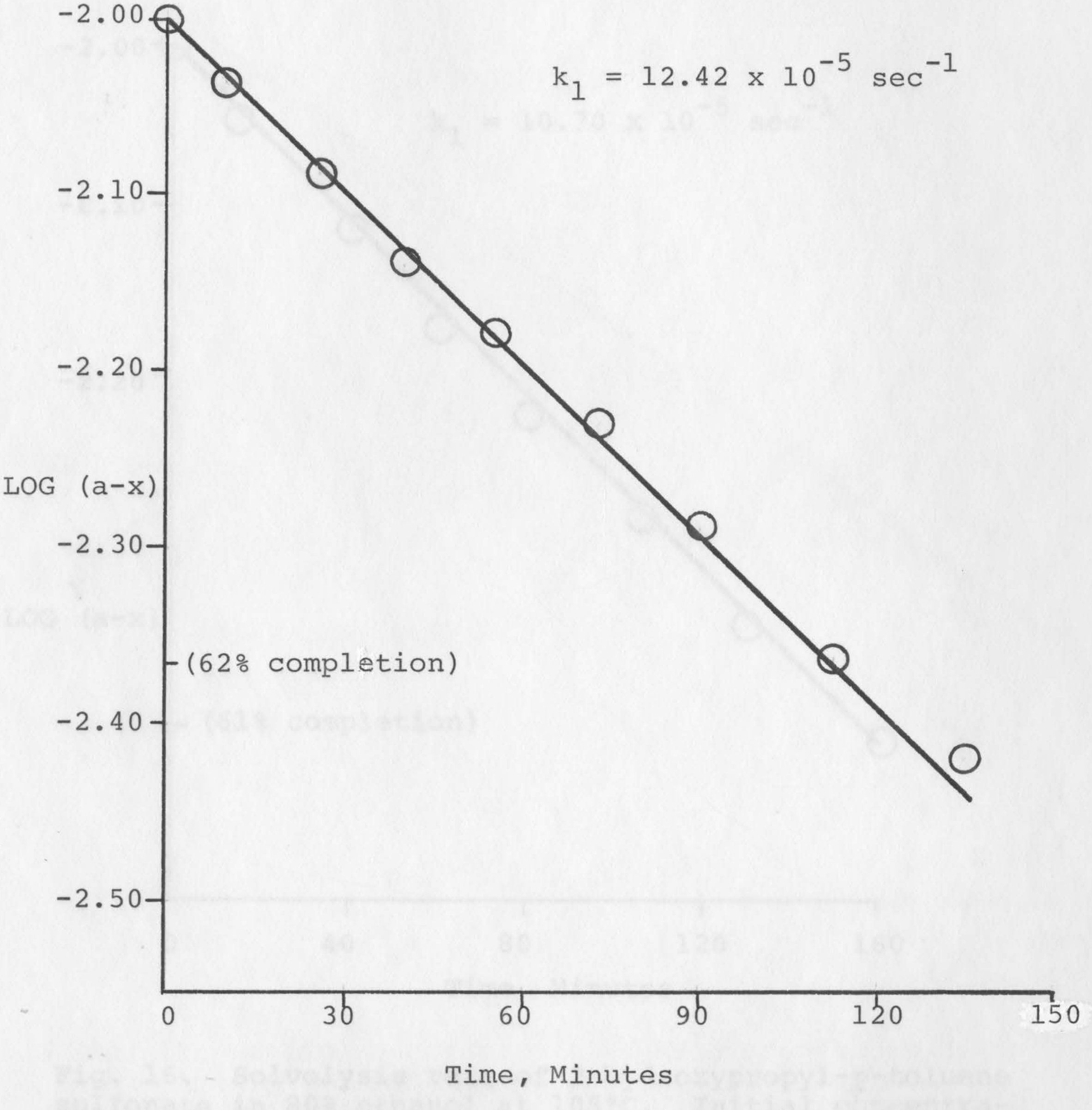

Fig. 15. Solvolysis rate of 2-hydroxypropyl-p-toluene sulfonate in 70% ethanol at 105°C. Initial concentration .01 mole /liter.

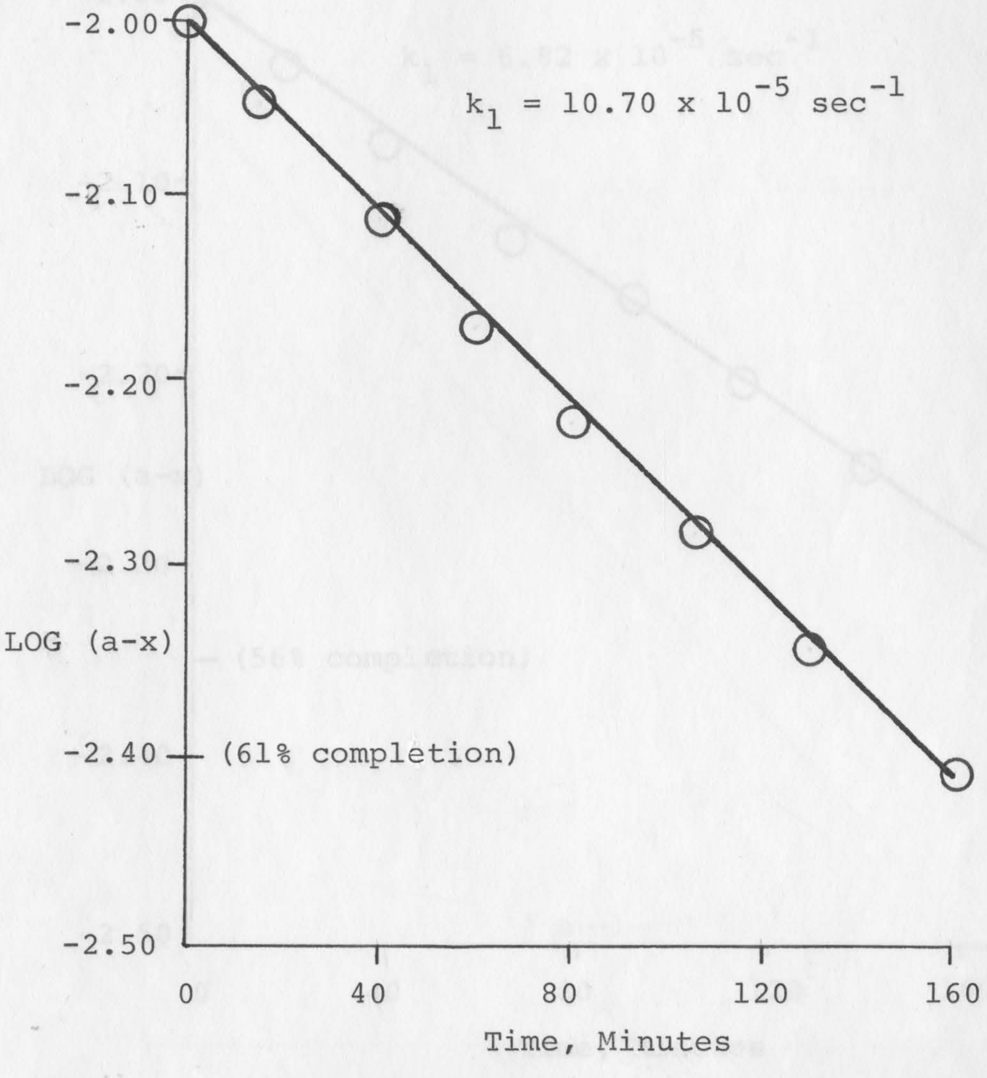

Fig. 16. Solvolysis rate of 2-hydroxypropyl-*p*-toluene sulfonate in 80% ethanol at 105°C. Initial concentration .01 mole /liter.

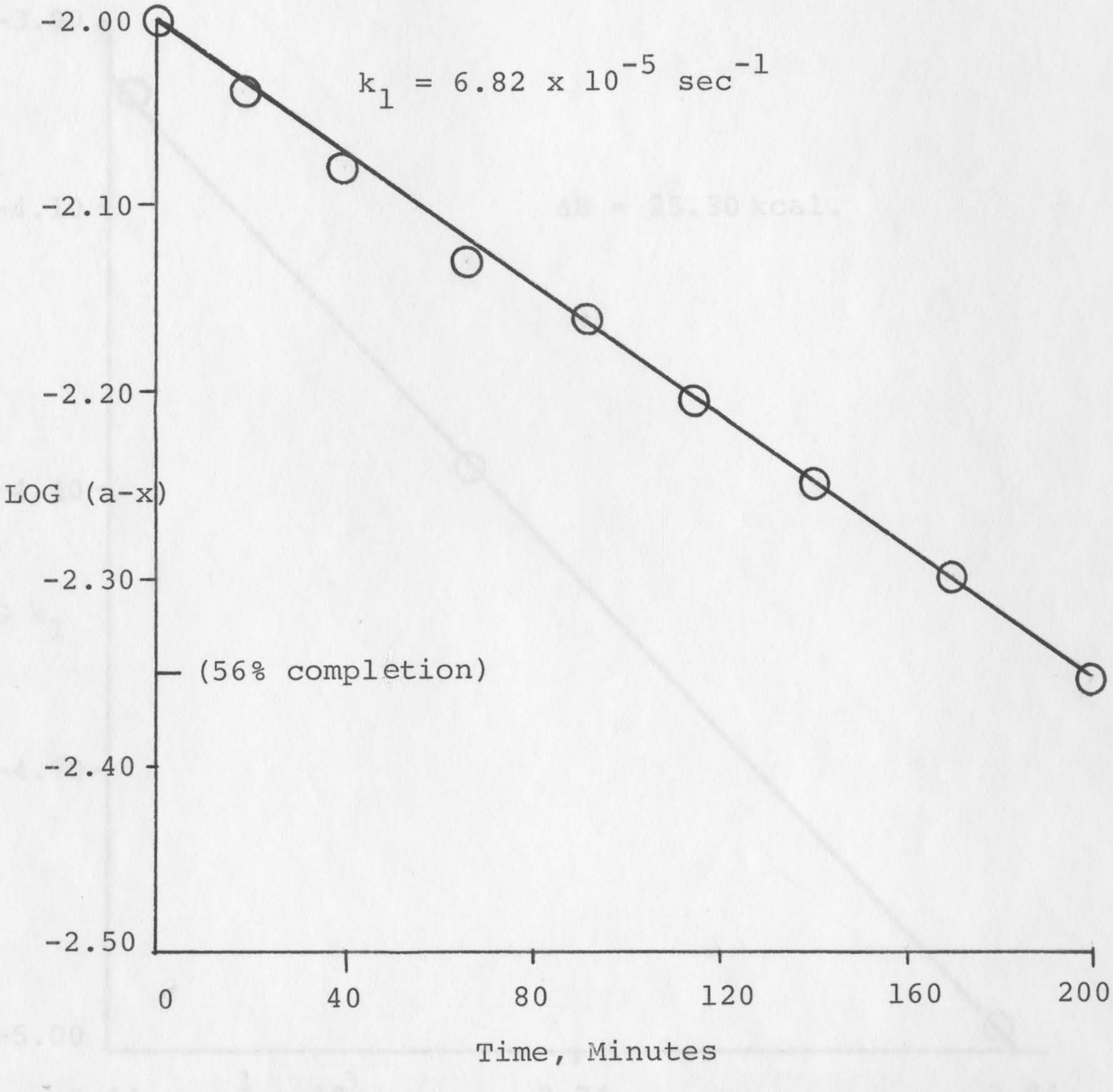

Fig. 17. Solvolysis rate of 2-hydroxypropyl-*p*-toluene sulfonate in 90% ethanol at 105°C. Initial concentration .01 mole/liter.

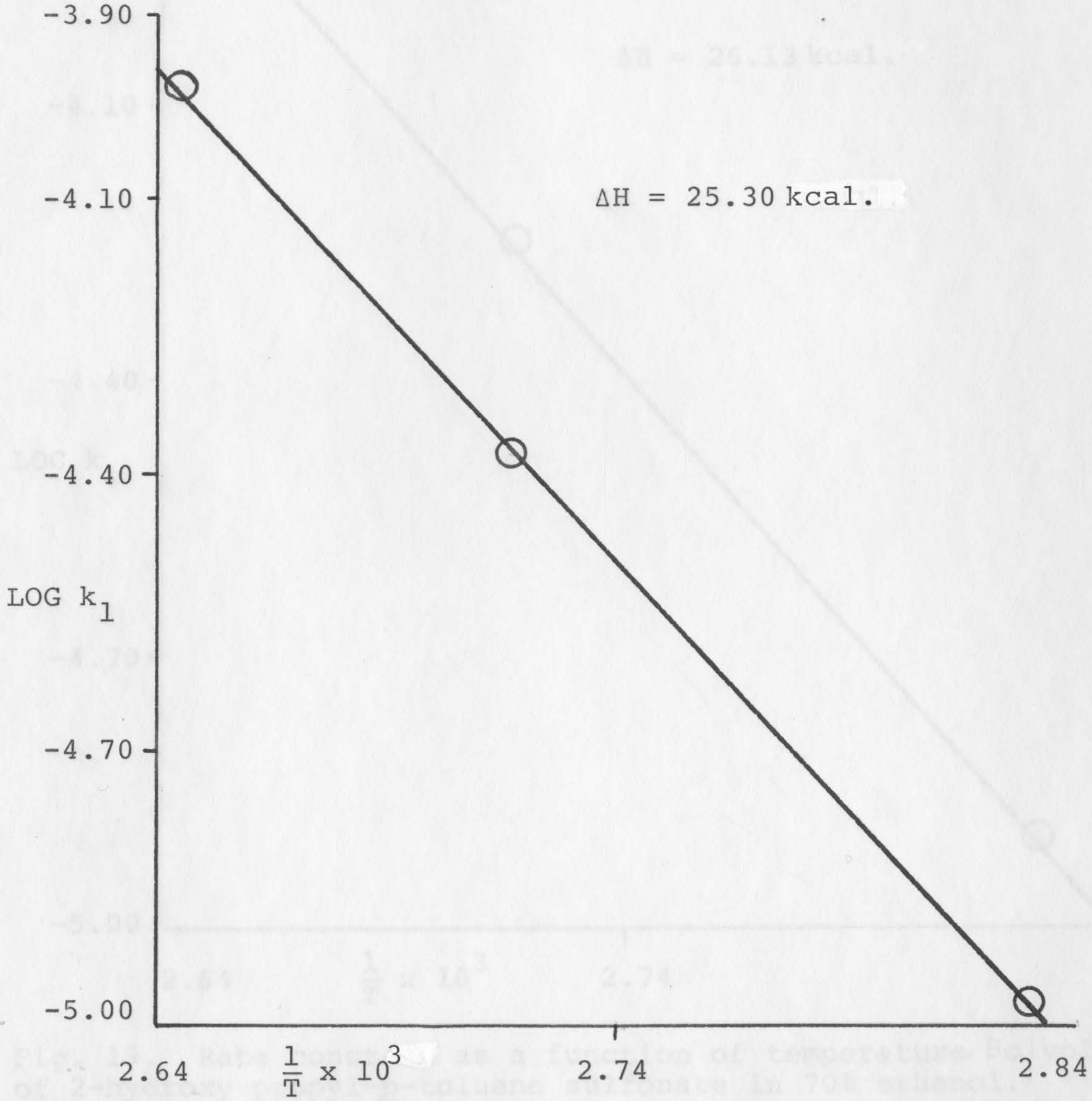

Fig. 18. Rate constant as a function of temperature. Solvolysis of 2-hydroxypropyl-p-toluene sulfonate in 80% ethanol.

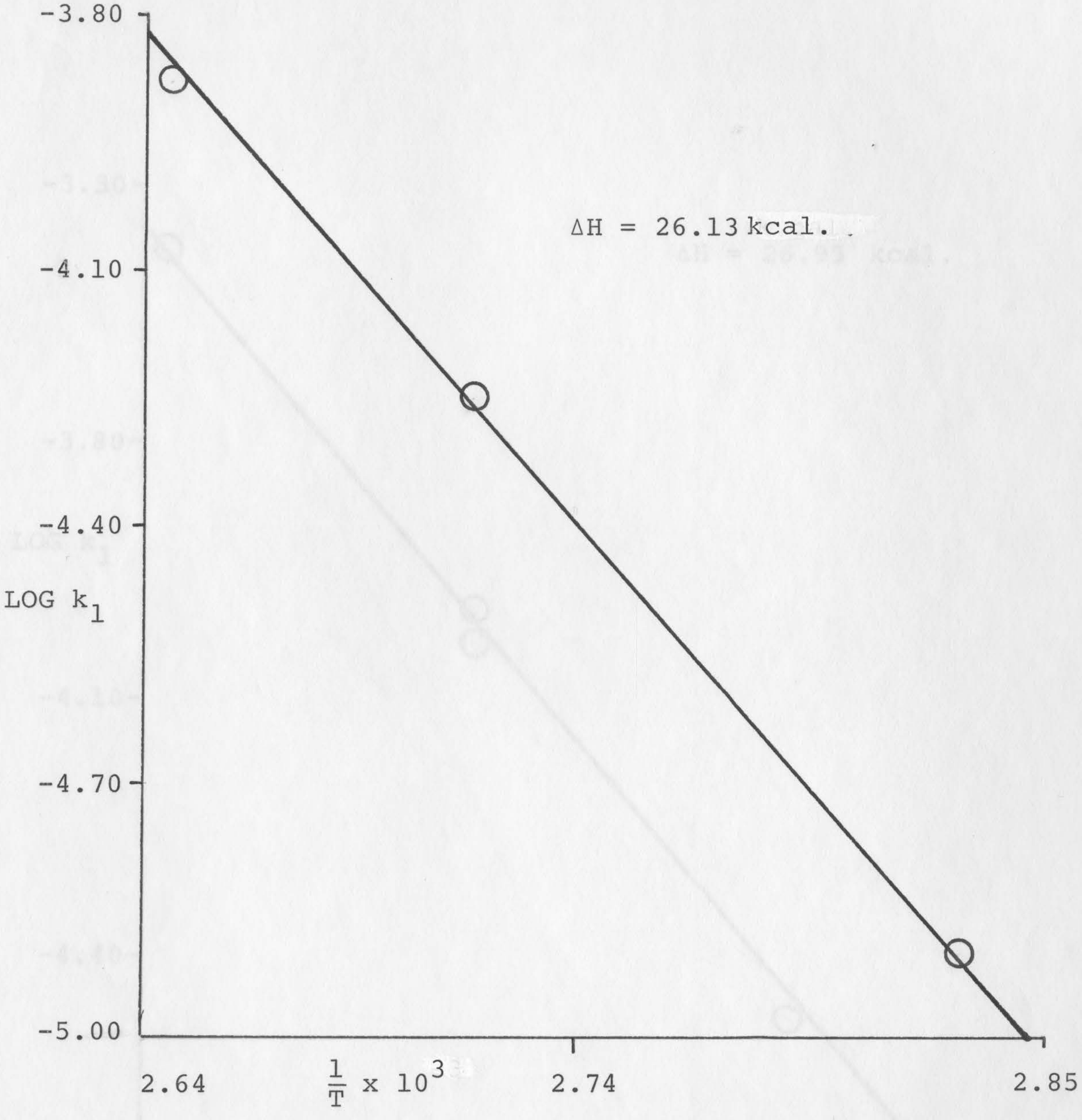

Fig. 19. Rate constant as a function of temperature. Solvolysis of 2-hydroxy propyl-p-toluene sulfonate in 70% ethanol.

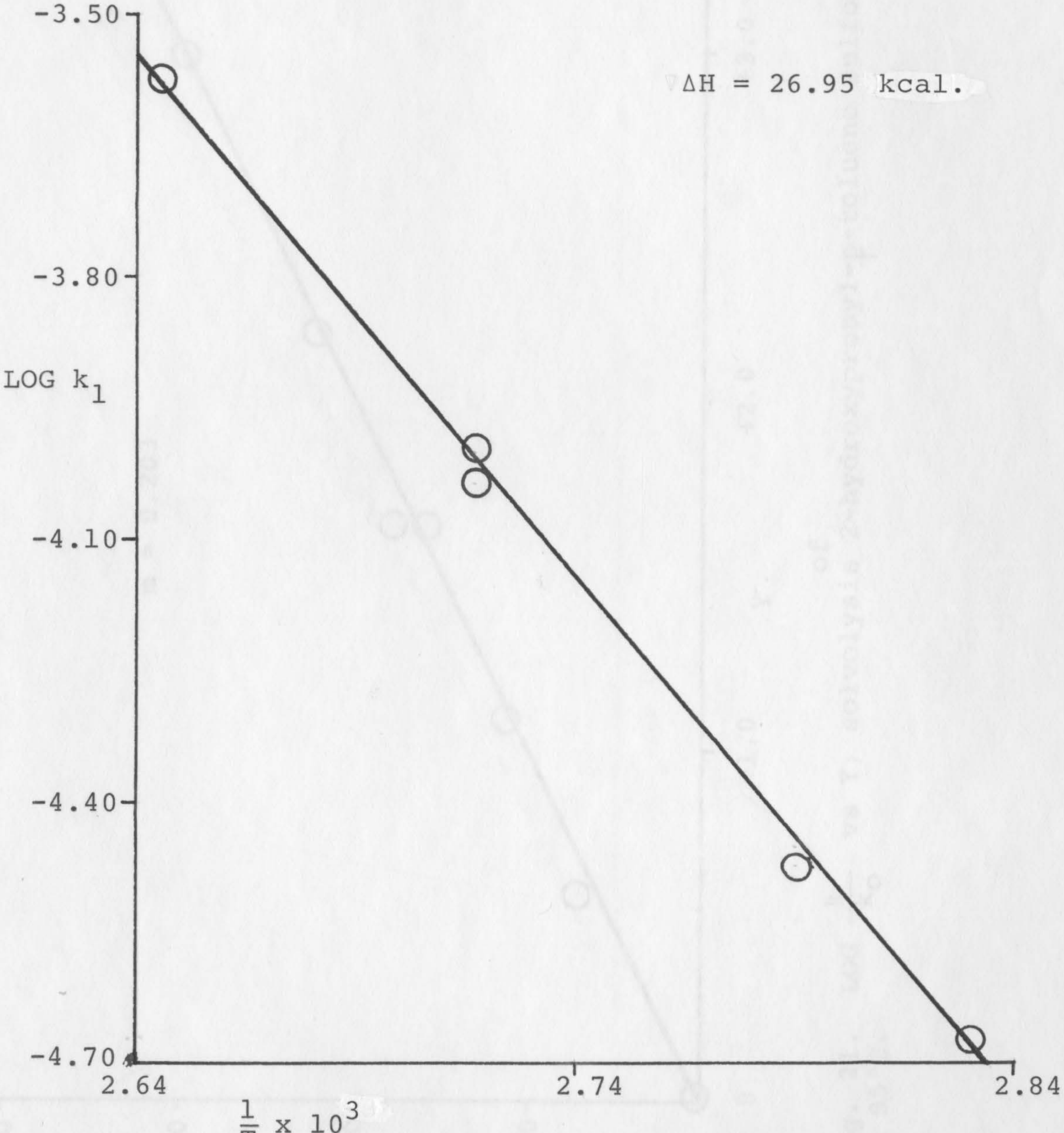

Fig. 20. Rate constant as a function of temperature. Solvolysis of 2-hydroxypropyl-*p*-toluene sulfonate in 50% ethanol.

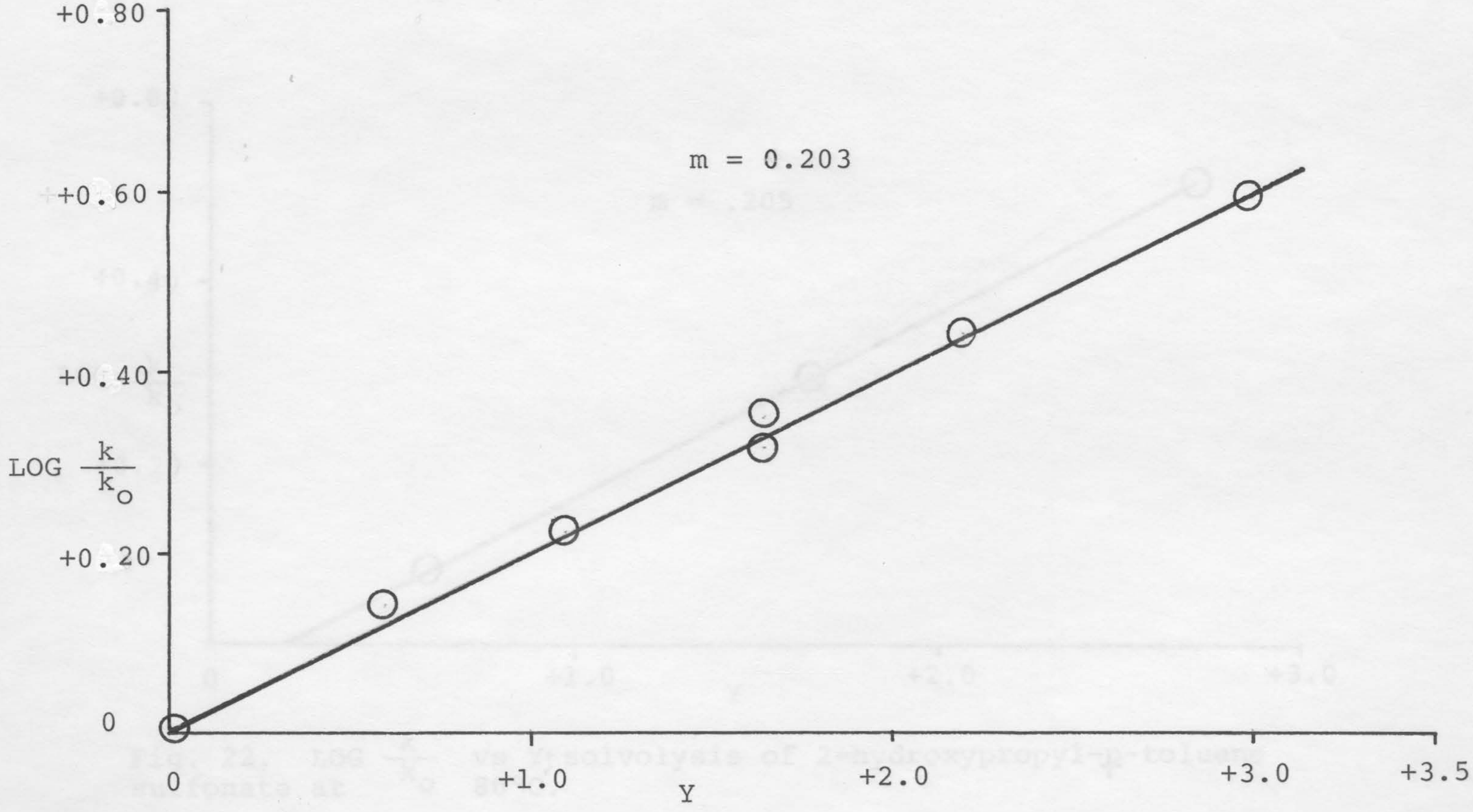

Fig. 21. LOG $\frac{k}{k_o}$ vs Y; solvolysis of 2-hydroxypropyl-*p*-toluene sulfonate at 95°C.

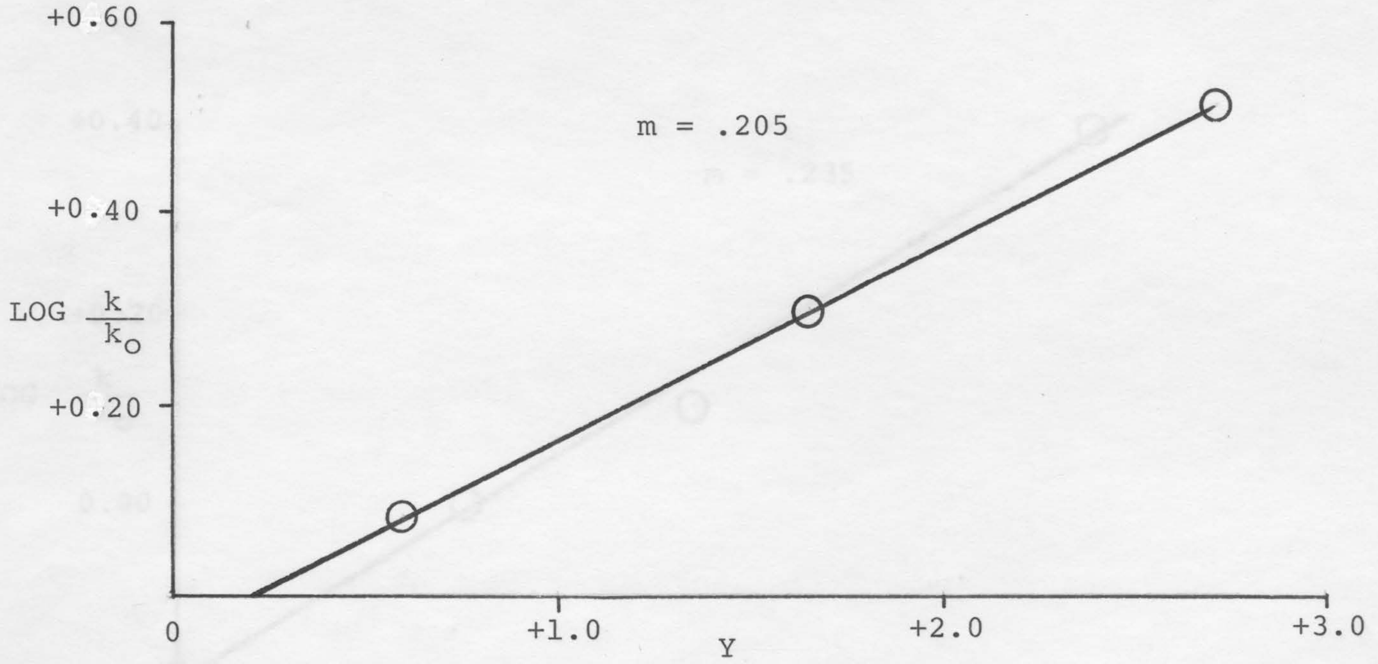

Fig. 22. LOG $\frac{k}{k_o}$ vs Y; solvolysis of 2-hydroxypropyl-*p*-toluene sulfonate at 80°C.

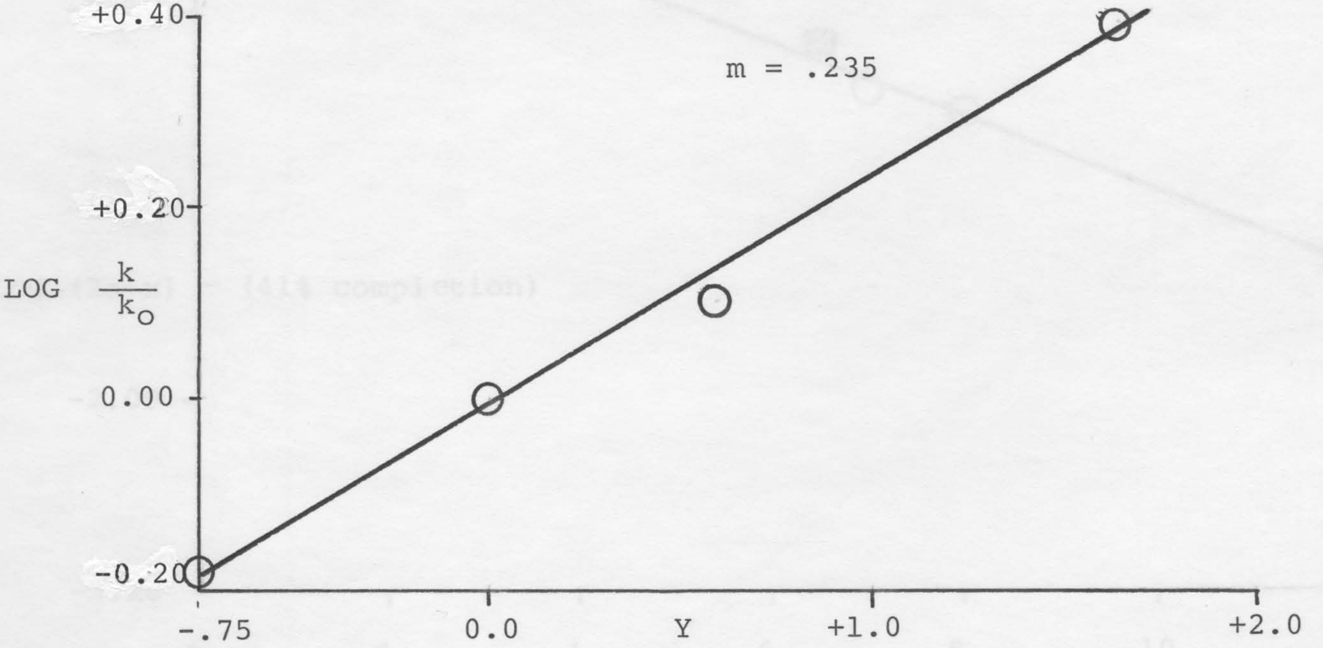

Fig. 23. LOG $\frac{k}{k_o}$ vs Y; solvolysis of 2-hydroxypropyl p-toluene sulfonate at 105°C.

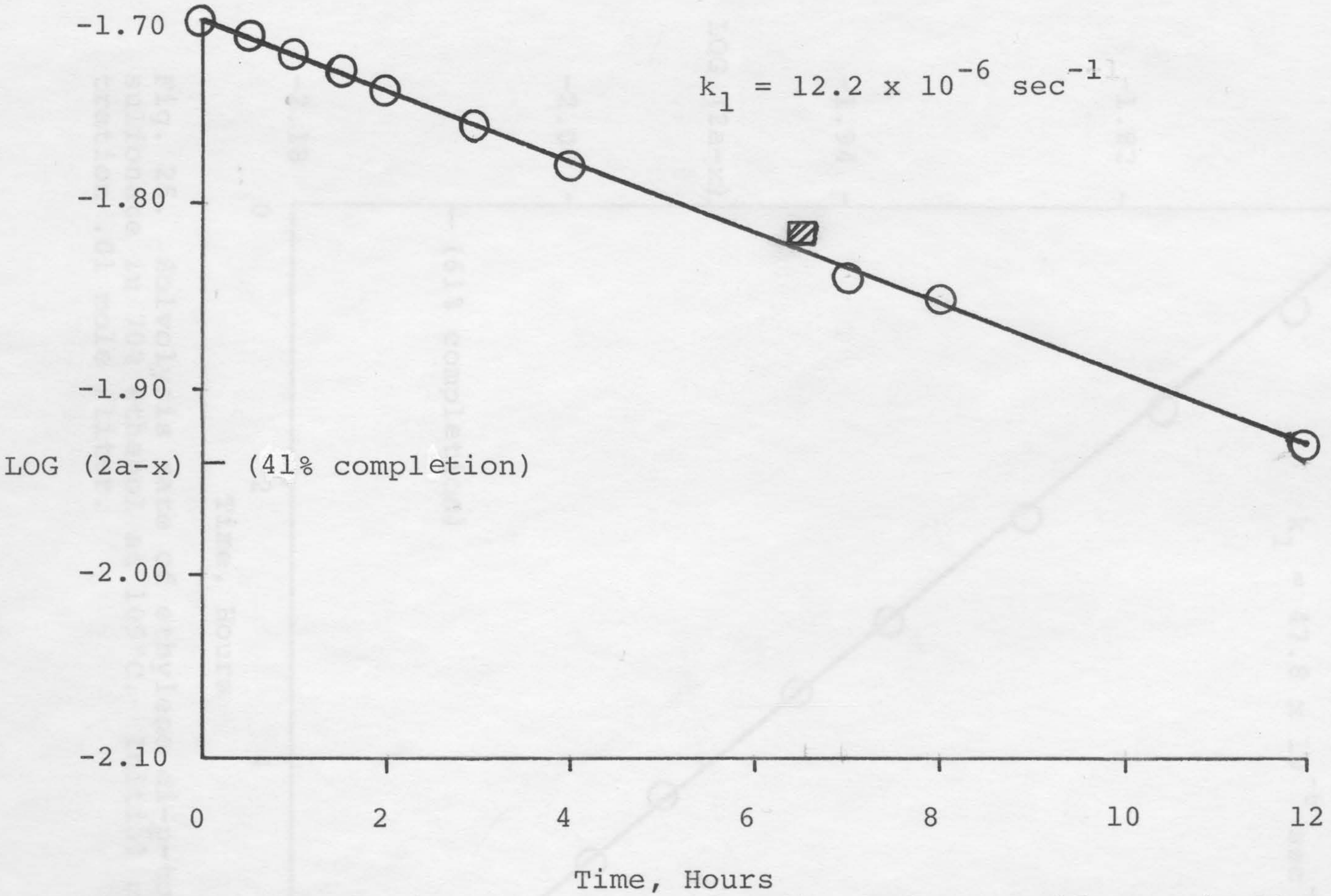

Fig. 24. Solvolysis rate of ethylene di-p-toluene sulfonate in 50% ethanol at 84°C. Initial concentration .01 mole /liter.
▨ Single point with 0.01 mole/liter pyridine present.

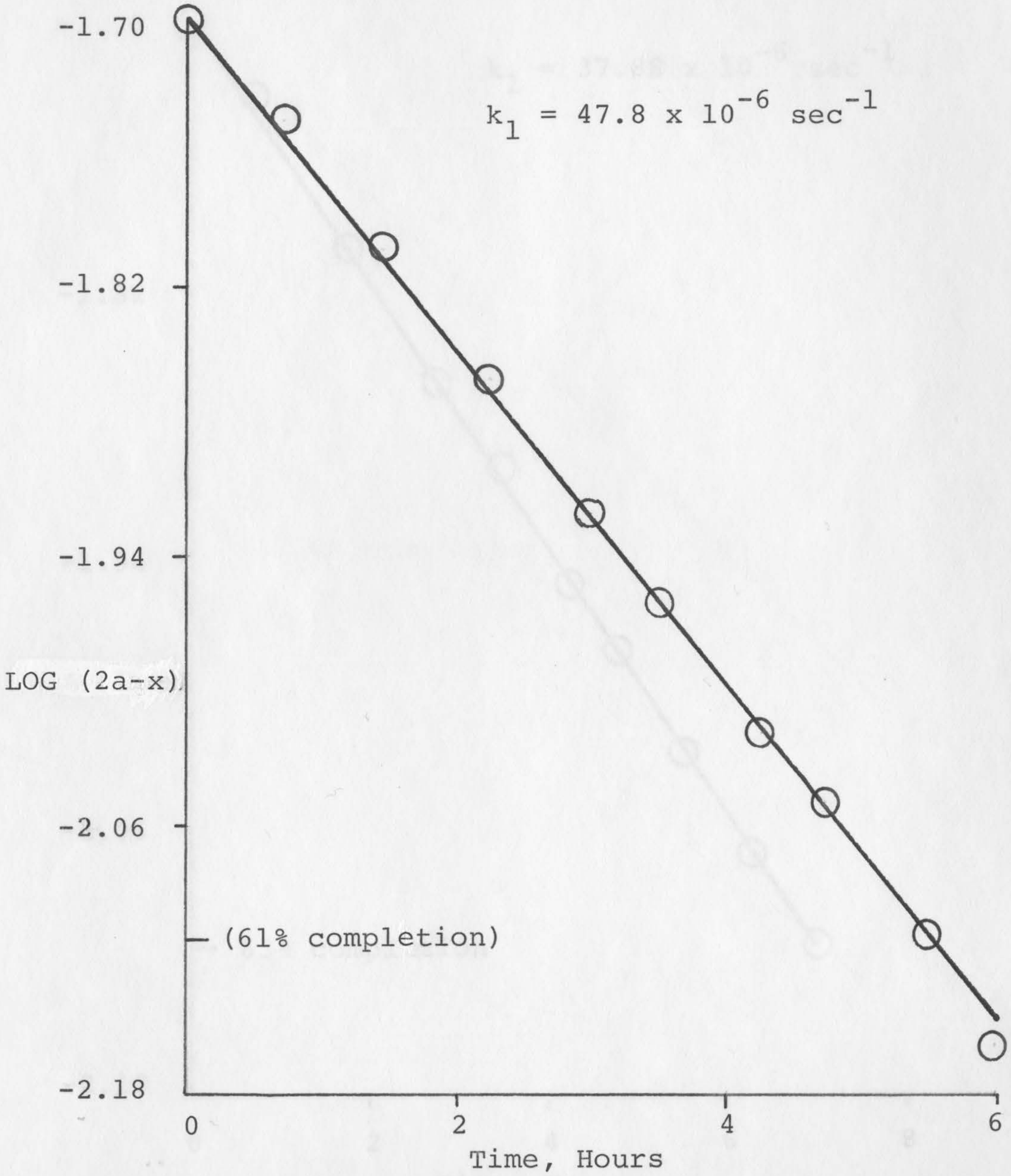

Fig. 25. Solvolysis rate of ethylene di-p-toluene sulfonate in 70% ethanol at 105°C. Initial concentration .01 mole /liter.

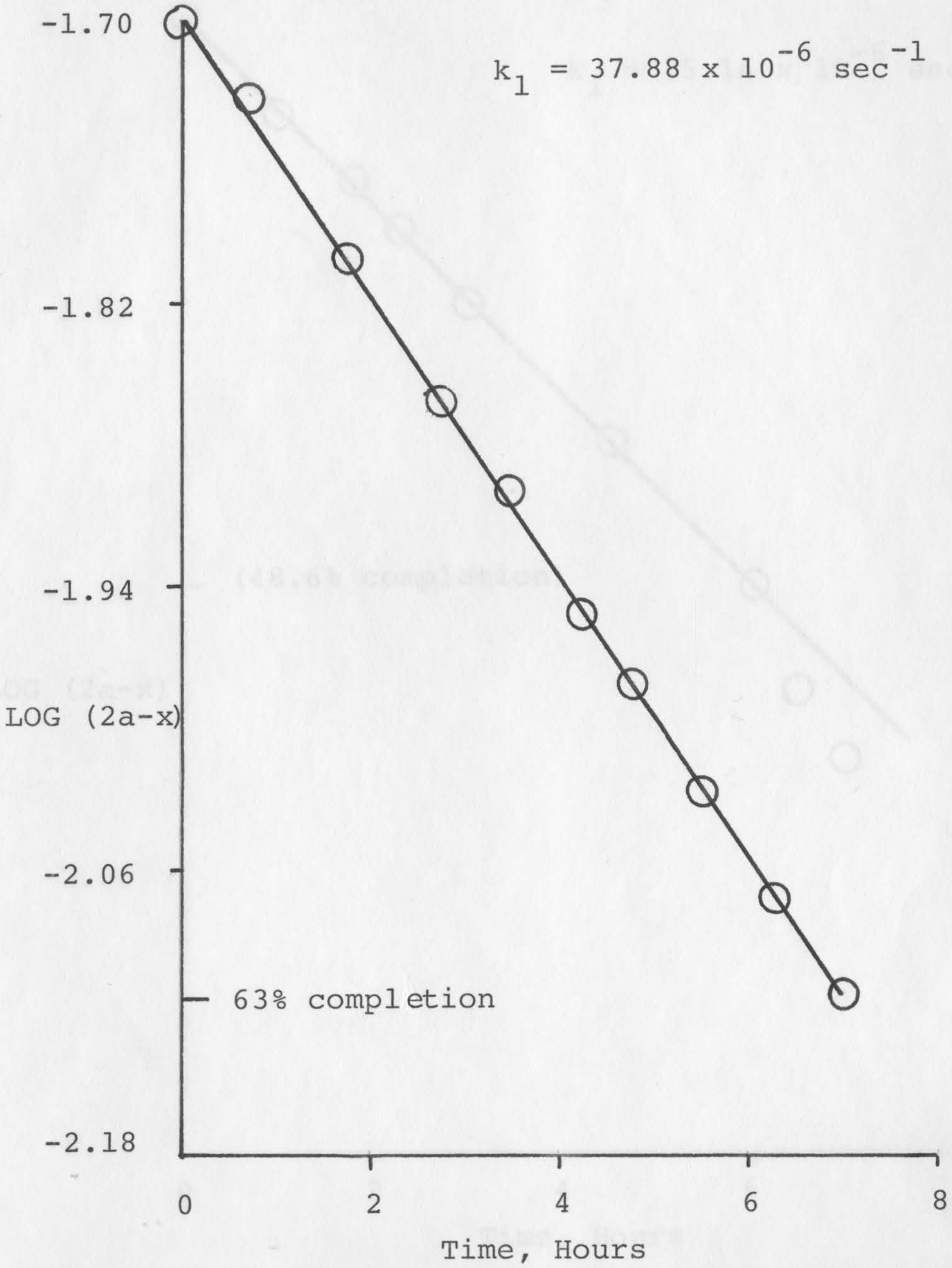

Fig. 26. Solvolysis rate of ethylene di-p-toluene sulfonate in 80% ethanol at 105°C. Initial concentration .01 mole /liter.

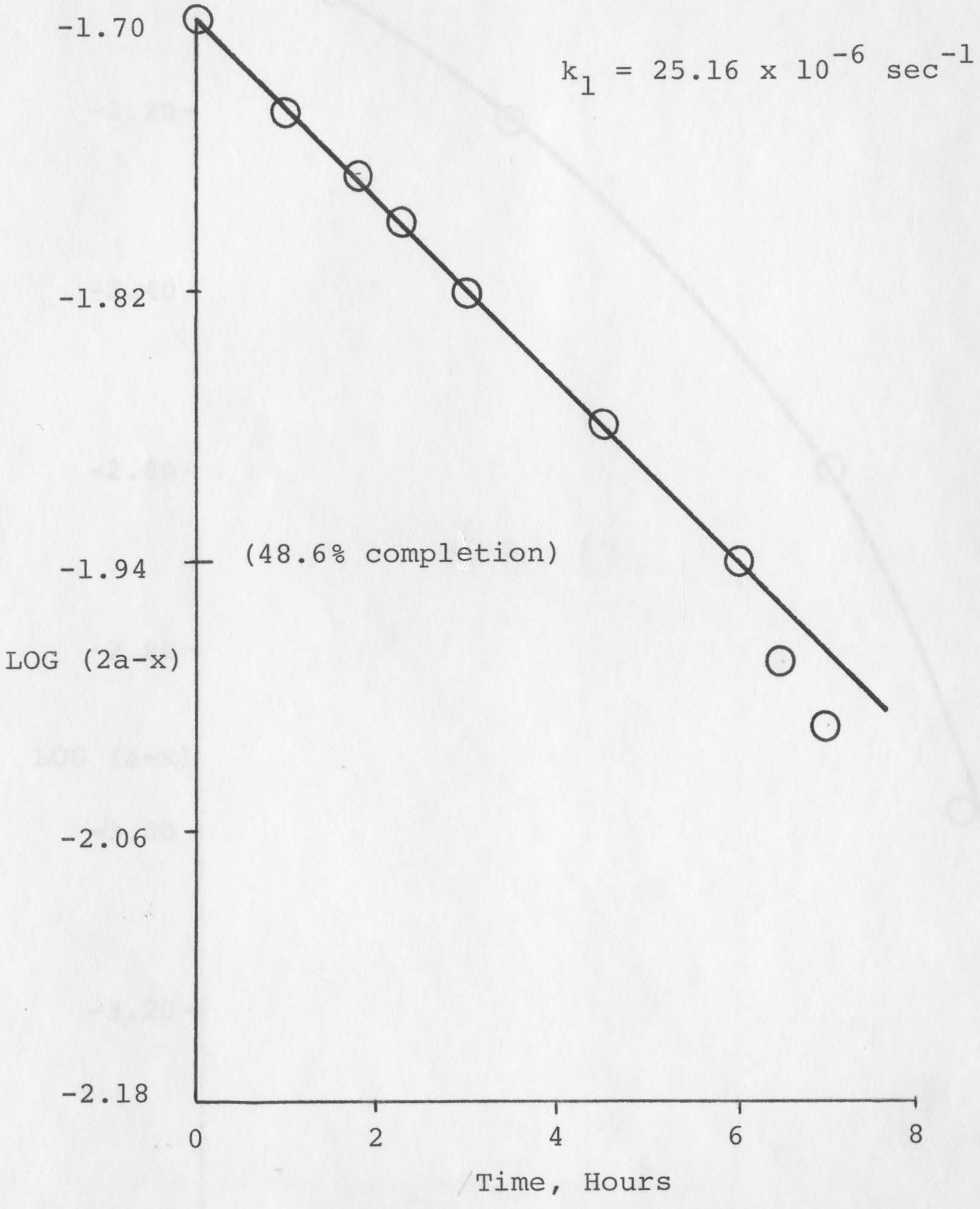

Fig. 27. Solvolysis rate of ethylene di-p-toluene sulfonate in 90% ethanol at 105°C. Initial concentration .01 mole /liter.

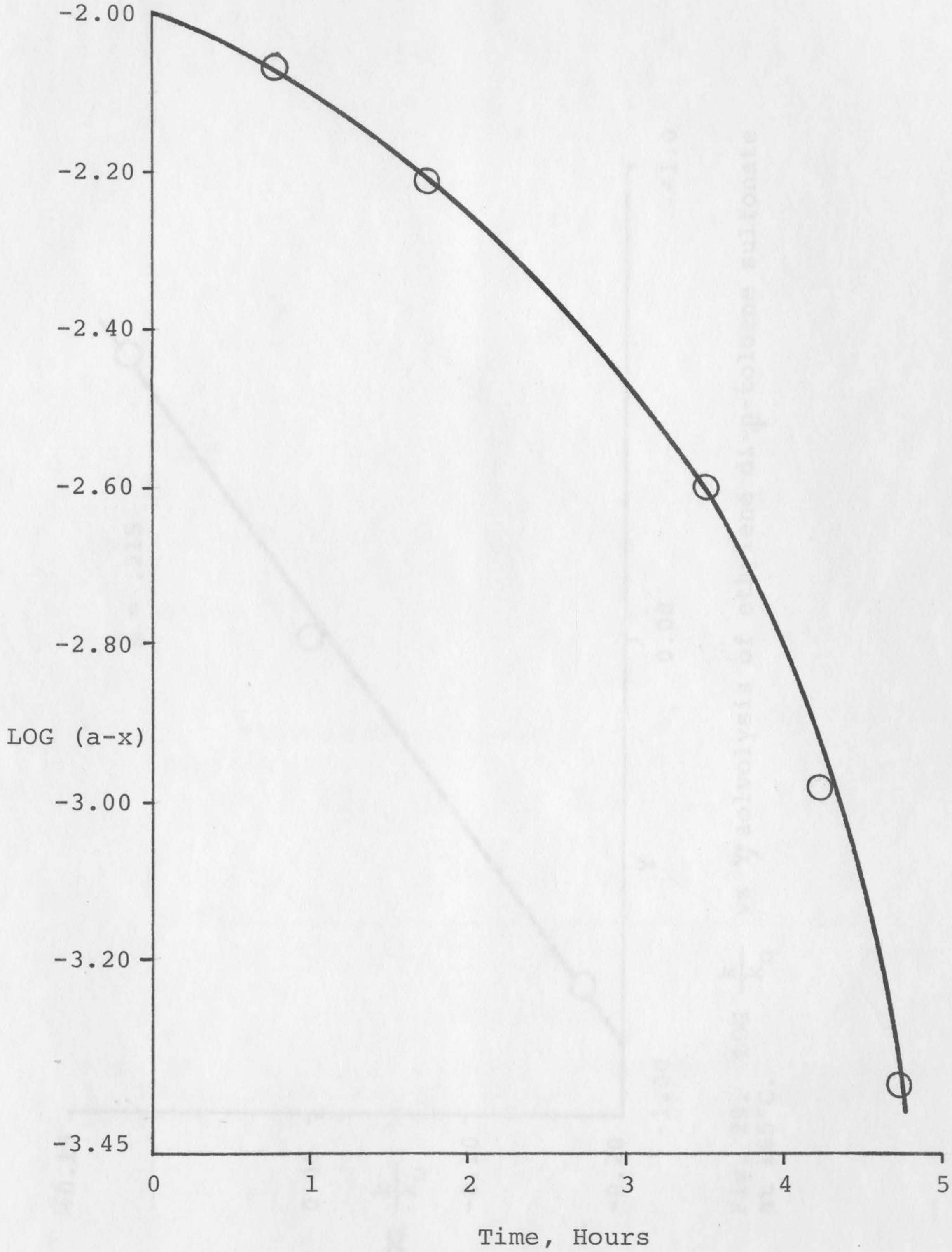

Fig. 28. Solvolysis rate of ethylene di-p-toluene sulfonate in 80% ethanol at 105°C. Initial concentration 0.01 mole/liter.

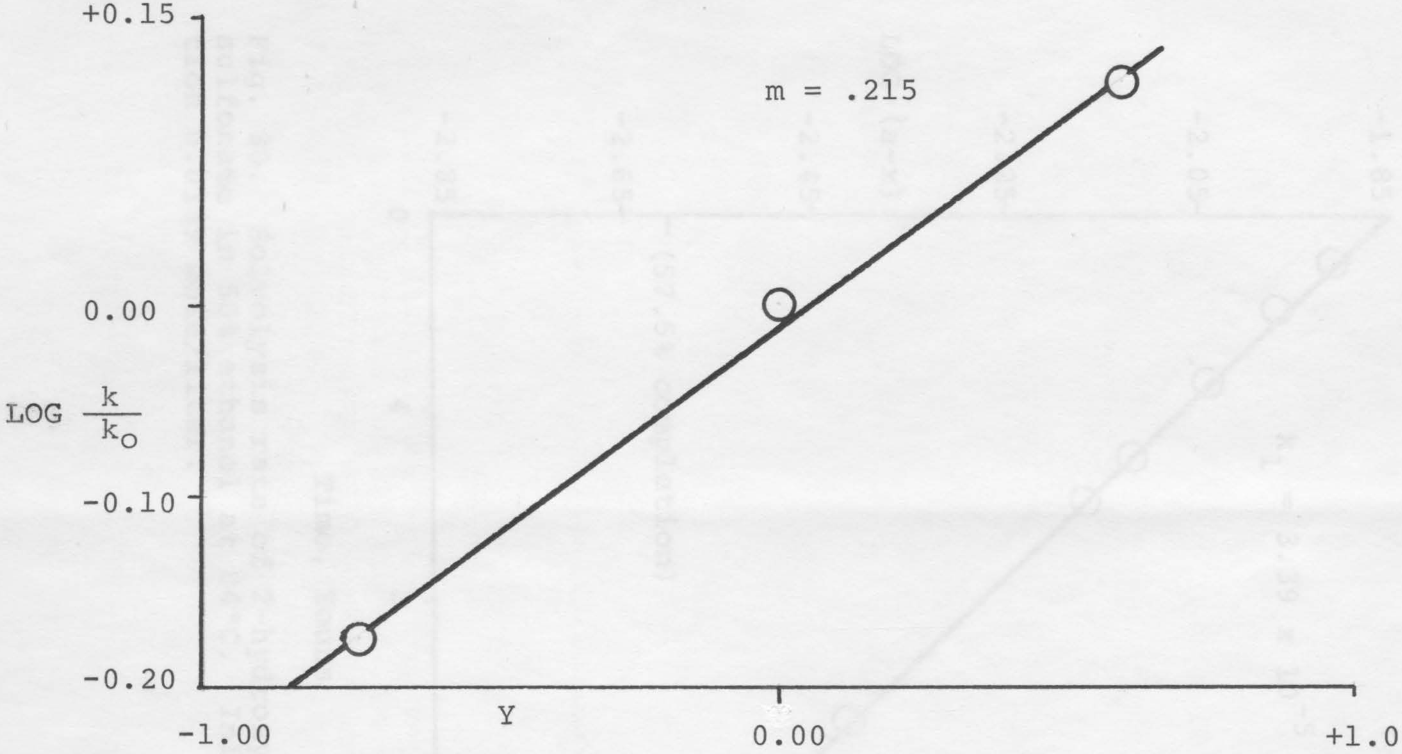

Fig. 29. LOG $\frac{k}{k_o}$ vs Y; solvolysis of ethylene di-p-toluene sulfonate at 105°C.

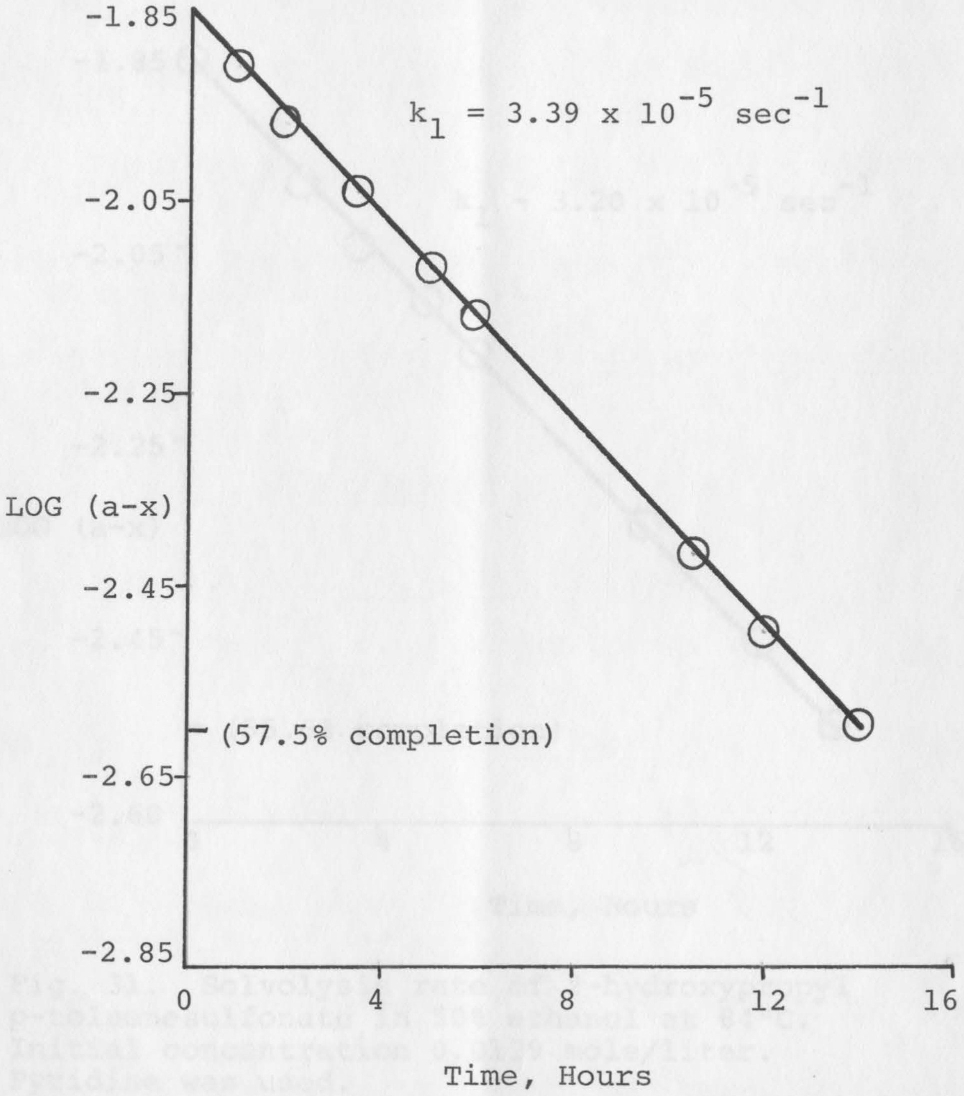

Fig. 30. Solvolysis rate of 2-hydroxypropyl p-toluenesulfonate in 50% ethanol at 84°C. Initial concentration 0.0139 mole/liter.

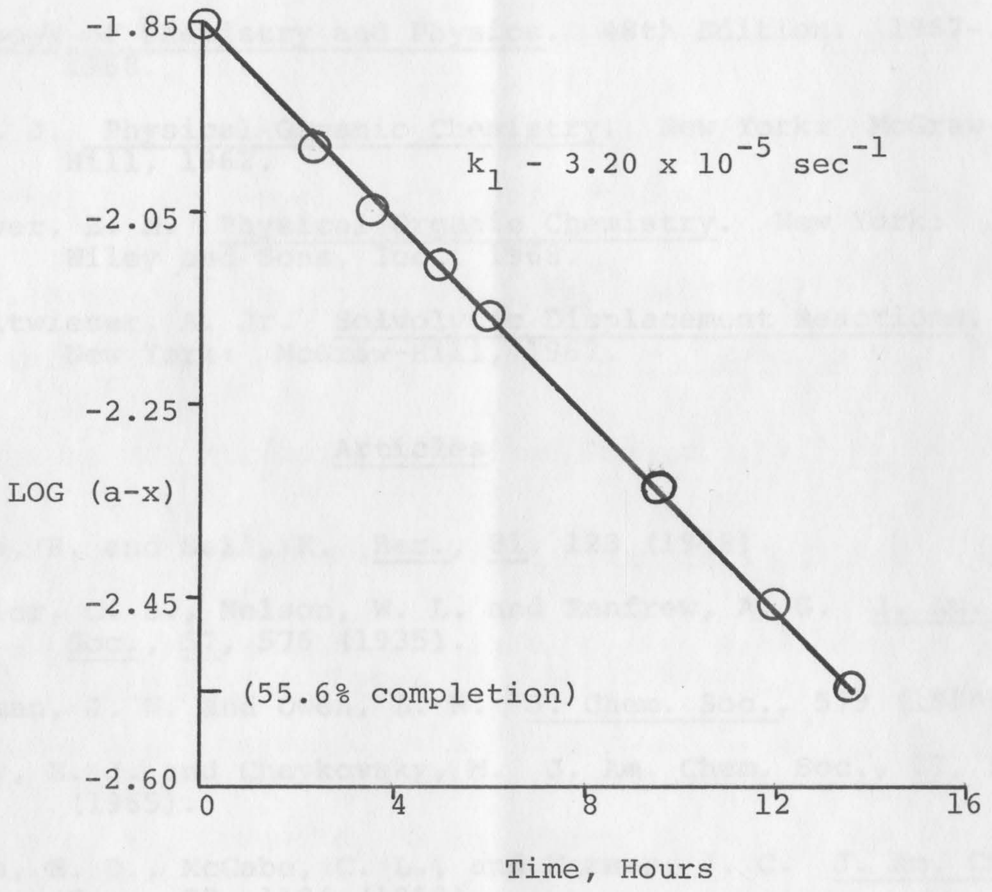

Fig. 31. Solvolysis rate of 2-hydroxypropyl p-toluenesulfonate in 50% ethanol at 84°C. Initial concentration 0.0139 mole/liter. Pyridine was used.

BIBLIOGRAPHY

Books

Gould, E. S. Mechanism and Structure in Organic Chemistry. New York: Holt, Rinehard & Winston, Inc., 1966.

Handbook of Chemistry and Physics. 48th Edition. 1967-1968.

Hine, J. Physical Organic Chemistry. New York: McGraw-Hill, 1962.

Kosower, E. M. Physical Organic Chemistry. New York: John Wiley and Sons, Inc., 1968.

Streitwieser, A. Jr. Solvolytic Displacement Reactions. New York: McGraw-Hill, 1962.

Articles

Bohme, H. and Sell, K. Ber., 81, 123 (1948).

Buttler, C. L., Nelson, W. L. and Renfrew, A. G. J. Am. Chem. Soc., 57, 575 (1935).

Chapman, J. H. and Owen, L. N. J. Chem. Soc., 579 (1950).

Corey, E. J. and Chaykovsky, M. J. Am. Chem. Soc., 87, 1353 (1965).

Cowan, H. D., McCabe, C. L., and Warner, J. C. J. Am. Chem. Soc., 72, 1194 (1950).

Dittmer, D. C. and Hertler, W. R. J. Am. Chem. Soc., 79, 4431 (1957).

Foster, F. C. and Hammett, L. P. J. Am. Chem. Soc., 68, 1736 (1946).

Gregor, I. K., Riggs, N. V. and Stimson, V. R. J. Chem. Soc., 76 (1956).

Groll, H. P. A. and Kautter, C. T. U. S. Patent 2,042,225; Chem. Abs., 30, 4872 (1936).

Grunwald, E. and Winstein, S. J. Am. Chem. Soc., 70, 846 (1946).

Harvey, G. J., Riggs, N. V. and Stimson, V. R. _J. Chem. Soc._, 3267 (1955).

Hine, J. and Brader, W. H. _J. Am. Chem. Soc._, _75_, 3964 (1953).

Hughes, E. D., Ingold, C. K. and Cowdrey, W. A. _J. Chem. Soc._, 1208 (1937).

Hughes, E. D., Ingold, C. K. and Patel, C. S. _J. Chem. Soc._, 526 (1933).

Kuhn, R. and Trischmann, H. _Ann._, _611_, 117 (1958).

Renfrew, A. G., Green, M. H. and Buttler, C. L. _J. Am. Chem. Soc._, _61_, 1783 (1939).

Ross, S. D. _J. Am. Chem. Soc._, _69_, 2982 (1947).

Winstein, S. and Fainberg, A. H. _J. Am. Chem. Soc._, _78_, 2770 (1956).

Winstein, S., Fainberg, A. H. and Grunwald, E. _J. Am. Chem. Soc._, _79_, 4146 (1957).

Winstein, S. and Grunwald, E. _J. Am. Chem. Soc._, _70_, 828 (1948).

Winstein, S., Grunwald, E., Buckles, R. E. and Hanson, C. _J. Am. Chem. Soc._, _70_, 816 (1948).

Winstein, S., Grunwald, E. and Ingraham, L. L. _J. Am. Chem. Soc._, _70_, 824 (1948).

Winstein, S., Grunwald, E. and Jones, H. W. _J. Am. Chem. Soc._, _73_, 2700 (1951).

Winstein, S., Lindegren, C. R. and Ingraham, L. L. _J. Am. Chem. Soc._, _75_, 155 (1953).

Winstein, S. and Lucas, H. _J. Am. Chem. Soc._, _61_, 1575 (1939).