[CONTRIBUTION FROM THE CHEMISTRY DEPARTMENT OF NORTHWESTERN UNIVERSITY]

The Hydrolysis of Chloromethyl Aryl Sulfides[1]

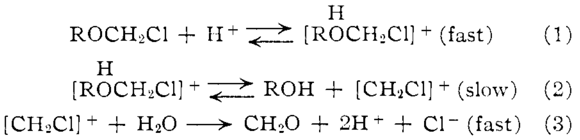
BY F. G. BORDWELL, GLENN D. COOPER AND H. MORITA

RECEIVED AUGUST 21, 1956

The hydrolysis rates for eight *m* and *p*-substituted chloromethyl aryl sulfides, determined conductometrically in 50% aqueous dioxane, were found to follow a first-order law and were independent of hydrogen ion concentration. An earlier suggestion that the mechanism for hydrolysis of α-halo ethers and α-halo sulfides involves the cleavage of a carbon–oxygen (or carbon–sulfur) bond in a complex, such as $[ROHCH_2Cl]^+$, is discarded in light of the kinetic evidence. Although the variation in hydrolysis rates from $p\text{-}NO_2C_6H_4SCH_2Cl$ to $p\text{-}CH_3OC_6H_4SCH_2Cl$ was almost 600-fold, the data fitted the Hammett equation nicely. The lower rate for phenyl as compared to alkyl chloromethyl sulfides and the failure of $p\text{-}CH_3OC_6H_4\text{-}SCH_2Cl$ to deviate appreciably from the value calculated using the Hammett equation are interpreted as indicating only minor resonance contributions involving the aryl groups in the carbonium–sulfonium intermediate, $[ArS\text{–}CH_2^+ \longleftrightarrow ArS^+{=}CH_2]$. This is consistent with other data indicating a reluctance of divalent sulfur to expand its valence shell to ten electrons.

An investigation of the hydrolysis rates of α-halo sulfides and of α-halo ethers has been reported previously by Böhme[2a] and by Böhme, Fisher and Frank.[2b] These authors followed the hydrolysis in dioxane solutions containing 0.2 to 40% of water by titrating the hydrochloric acid produced with a 0.1 *N* solution of tribenzylamine in acetone. The results obtained from Böhme's initial study[2a] may be summarized as follows: (1) the reaction is first order in halide (average deviations of about 5% in first order constants); (2) the reaction is autocatalyzed by the hydrochloric acid produced; (3) $ClCH_2OCH_2CH_3$ hydrolyzes in dioxane containing 2% water at a rate approximately 1000 times that of $ClCH_2SCH_2CH_3$.

Böhme[2a] suggested the following sequence of steps for the hydrolysis of α-chloro ethers and thioethers to account for these experimental results



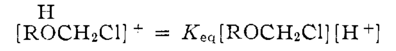
$$RO\overset{H}{C}H_2Cl + H^+ \rightleftarrows [R\overset{H}{O}CH_2Cl]^+ \text{ (fast)} \quad (1)$$

$$[R\overset{H}{O}CH_2Cl]^+ \rightleftarrows ROH + [CH_2Cl]^+ \text{ (slow)} \quad (2)$$

$$[CH_2Cl]^+ + H_2O \longrightarrow CH_2O + 2H^+ + Cl^- \text{ (fast)} \quad (3)$$

The autocatalytic effect observed was ascribed to the necessity of formation of the oxonium or sulfonium salt in step 1. The slower rate of hydrolysis of the α-chloro sulfides than α-chloro ethers was attributed to a lesser tendency of the sulfonium ion than the oxonium ion to undergo ionization in step 2.

Actually, the proposed mechanism does not satisfy the data. Since step 2 is assumed to be rate determining, the rate will be proportional to the oxonium or sulfonium ion concentration, and this depends on the hydrogen ion concentration.



$$[R\overset{H}{O}CH_2Cl]^+ = K_{eq}[ROCH_2Cl][H^+]$$

Inasmuch as $[H^+]$ increases throughout the reaction, the rate expression would not be a simple first-order one but would involve $[H^+]$.

In order to clarify the situation with regard to the mechanism, a kinetic study of the hydrolysis of phenyl chloromethyl sulfide, $C_6H_5SCH_2Cl$, was undertaken. The rates of hydrolysis of seven *m*- and *p*-substituted phenyl chloromethyl sulfides were measured to determine the electrical effects of the *m*- and *p*-substituents.

Böhme, Fisher and Frank[2b] have shown that the over-all reaction involved in the hydrolysis is that given in equation 4.

$$2ArSCH_2Cl + H_2O = (ArS)_2CH_2 + CH_2O + 2H^+ + 2Cl^- \quad (4)$$

Since the reactions are very rapid and are sensitive to solvent when small quantities of water are present, it is much more convenient to follow the rates conductometrically than by a titration method. In this way the rates could be followed in solutions containing more water (50% aqueous dioxane, by volume) than is possible by Böhme's method, even though the half-lives were as low as 23 seconds.

Rate constants were determined graphically from a plot of log $(R/R - R_E)$ *vs.* *t*, where *R* is the resistance at time *t* and R_E is the equilibrium resistance.[3] Straight-line relationships were observed in all instances, often out to 90% completion of the reaction. No indication of auto-catalysis[2a] was evident in any of the runs (see *e.g.* Fig. 1). The rate constants given in Table I are for the most part the average of two or more runs. A plot of log *k* *vs.* $1/T$ gave straight lines for the hydrolysis run at three or more temperatures.

TABLE I

KINETIC DATA FOR THE HYDROLYSIS OF ARYL CHLOROMETHYL SULFIDES, $YC_6H_4SCH_2Cl$, IN 50% AQUEOUS DIOXANE

Y	Rate constants × 10^5 sec.$^{-1}$ 21°	25.1°	30°	34.85°	50°	60.7°	E_α, kcal./mole	$\Delta S^{\ddagger}$, e.u.
p-CH_3O	660		1900	3000			20	−2.8
p-CH_3	260		860	1400			22.5	3.8
m-CH_3				700				
H	110	190		560	2330		20	−6.3
p-Cl				140				
p-Br				90				
m-Cl				49				
p-NO_2				5.7	33	85	21.5	−10.6

Discussion

Since no autocatalytic effect was observed and the hydrolysis rates followed first-order kinetics

---

[1] This investigation was supported by the American Petroleum Institute under Project 48B and by the Office of Naval Research under Contract No. N7onr-45007. A preliminary account of the work was given at the American Chemical Society Meeting at Atlantic City, N. J., September, 1952, Abst. p. 5-M.

[2] (a) H. Böhme, *Ber.*, **74**, 248 (1941); (b) H. Böhme, H. Fisher and R. Frank, *Ann.*, **563**, 54 (1949); (c) see also R. Peters and E. Walker, *Biochem. J.*, **17**, 260 (1923).

[3] A. A. Frost and R. G. Pearson, "Kinetics and Mechanism," John Wiley and Sons, Inc., New York, N. Y., 1953.

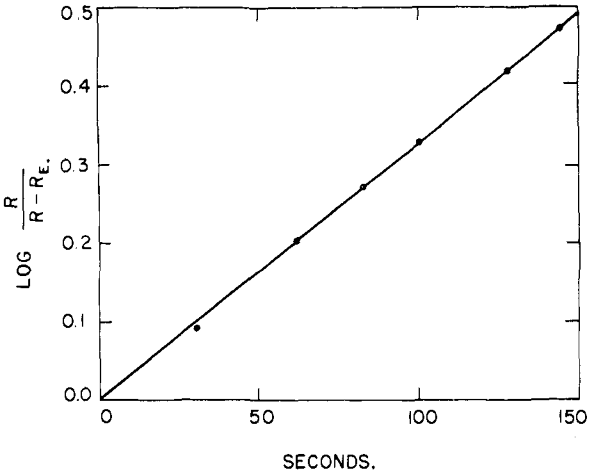
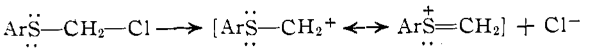


Fig. 1.—Hydrolysis of chloromethyl phenyl sulfide in 50% aqueous dioxane at 34.85°.

despite the large increase in $[H^+]$ during the reaction, the reaction appears to be better represented as an ionization to form a carbonium–sulfonium ion than by the mechanism suggested by Böhme.[2a]



$$\text{Ar}\ddot{\underset{\cdot\cdot}{\text{S}}}\text{—CH}_2\text{—Cl} \longrightarrow [\text{Ar}\ddot{\underset{\cdot\cdot}{\text{S}}}\text{—CH}_2^+ \longleftrightarrow \text{Ar}\overset{+}{\underset{\cdot\cdot}{\text{S}}}\text{=CH}_2] + \text{Cl}^-$$

The importance of the participating effect of sulfur in promoting this hydrolysis is made evident by the fact that $C_6H_5SCH_2Cl$ hydrolyzes 10 times as rapidly under these conditions as does $(CH_3)_3CCl$. Substitution of the C_6H_5S– group for a hydrogen of CH_2Cl therefore has a greater effect than substitution of all three hydrogens by CH_3. The data of Peters and Walker[2c] indicate that the participating effect of sulfur to the release of halide ion is about five times as great in the α-position ($ClCH_2SCH_2Cl$) as in the β-position ($ClCH_2CH_2SCH_2CH_2Cl$).

The hydrolysis rate of a 0.001 M solution of $C_6H_5SCH_2Cl$ was increased only 1.5-fold by the presence of an equal concentration of hydroxide ion (1.1-fold by 0.001 M $KClO_4$). Since the hydroxide ion concentration is increased by 10^4 in this medium, the ionization is virtually unaffected by the presence of alkali.

The greater effect of oxygen than sulfur in promoting the ionization of an α-halogen[2] is understandable since oxygen is known to enter into electron-donor type resonance to a larger extent than does sulfur.[4]

For the hydrolysis of benzyl halides in aqueous acetone ρ of the Hammett equation,[5] $\log k/k_0 = \sigma\rho$, is negative (ρ −1.87) and groups with positive σ values (m or p-NO_2, CN, $COCH_3$, SO_2CH_3, etc.) retard the reaction rate, whereas groups with negative σ-values (p-CH_3O, p-CH_3, m-CH_3, etc.) accelerate it. Similarly it was found that p-NO_2 caused a 100-fold retardation in the rate of hydrolysis of $C_6H_5SCH_2Cl$ and p-CH_3O caused a 5.4-fold acceleration. A plot of $\log k/k_0$ *vs.* the "normal" σ-values[5] for the groups gave a reasonably good straight line (Fig. 2). The ρ-value (slope) was −2.6, showing a somewhat greater effect of substituents than in the hydrolysis of benzyl halides.

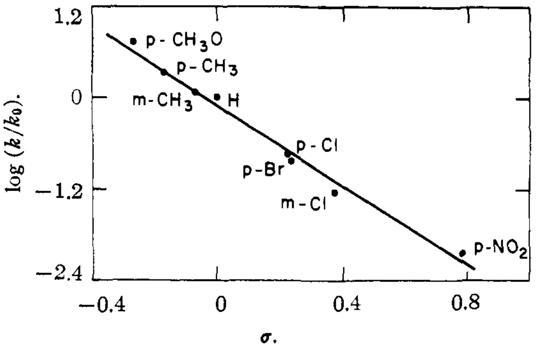
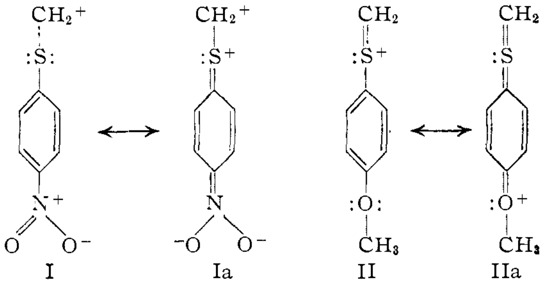


Fig. 2.—Relationship between the rate of hydrolysis of chloromethyl aryl sulfides and Hammett *sigma* values; determination of *rho*.

The fact that the log k/k_0 values for the p-NO_2 and p-CH_3O compounds fell near the line when plotted against the "normal" σ-values (those based on the ionization constants of the benzoic acids[5]) indicates that resonance interactions of these groups with $[-\text{S}\doteq\text{CH}_2]^+$ are no greater than their resonance interactions with the carboxyl group of benzoic acid. This is at first sight somewhat surprising since in other examples of reactions where unshared electrons are available on an atom attached to the ring (reactions of amines, phenols, thiophenols) the p-NO_2 group and similar groups show large deviations[5] (the apparent σ-value for p-NO_2 may be increased by as much as 0.5 unit, which would lead to a rate about 10-fold that calculated on the basis of the "normal value"). Similarly, a deviation for the p-CH_3O group might be possible, since when a positive charge develops on the atom attached to a benzene ring in the transition state of a reaction (*e.g.*, hydrolysis of $ArCH_2Cl$ or $ArCH_2OSO_2C_7H_7$), the p-CH_3O may show a greatly enhanced effect.[6]

In terms of current structural representations the results suggest that neither resonance structures of the type Ia nor IIa make very large contributions to the hybrid structure.



I Ia II IIa

The unimportance of Ia may be rationalized on the basis of Pauling's adjacent charge rule.[7] The failure of IIa to make a substantial contribution is more surprising. However, evidence has been accumulating to indicate that divalent sulfur is reluc-

---

(4) Pertinent data are summarized by F. G. Bordwell and P. J. Boutan, THIS JOURNAL, **78**, 854 (1956).

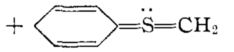
(5) L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, Chapter VII.

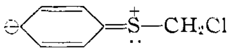
(6) See, for example, G. E. K. Branch and M. Calvin, "Theory of Organic Chemistry," Prentice-Hall, Inc., New York, N. Y., 1941, p. 442; C. J. Swain and W. P. Langsdorf, THIS JOURNAL, **73**, 2813 (1951); J. K. Kochi and G. S. Hammond, *ibid.*, **75**, 3445 (1953).

(7) L. C. Pauling, "The Nature of the Chemical Bond," Oxford University Press, London, 1948.

tant to expand its valence shell to ten electrons,[4,8] which is required in IIa. Evidently the *m*- and *p*-substituents in $ArSCH_2Cl$ do not relay their electrical effects primarily through resonance involving sulfur, and yet ρ has a more negative value here than for the hydrolysis of $ArCH_2Cl$. This further bears out the importance of inductive effects and field effects.[9]

The inability of the positively charged sulfur of $ArS^+{=}CH_2$ to distribute its charge over the phenyl ring by resonance is further emphasized by the observation[2b] that the rate of hydrolysis of $RSCH_2Cl$ (R = Me, Et, Pr, etc.) is about 200 times that of $C_6H_5SCH_2Cl$ in 90% aqueous dioxane. In our more dilute solutions the rate of solvolysis of CH_3SCH_2Cl was found to be too fast to measure. There is no question but what the rate is considerably greater than for the phenyl compound, as reported.[2b] The phenyl group is much superior to alkyl groups in stabilizing the positive charge on an adjacent carbon atom. Thus, substitution of phenyl for an α-hydrogen in ethyl chloride produces an effect comparable to that of *two* methyl groups (in 80% alcohol *k* for solvolysis of $C_6H_5CH(CH_3)Cl$ = 0.59 hr.$^{-1}$ at 50°[10] and *k* for $(CH_3)_3CCl$ = 0.37 hr.$^{-1}$ at 45°[11]). The superiority of phenyl over alkyl in stabilizing a carbonium ion is also apparent from Böhme's work,[2b] which shows that the relative hydrolysis rates of CH_3SCH_2Cl, $CH_3SCH(CH_3)Cl$ and $CH_3SCH(C_6H_5)Cl$ in 90% dioxane at 25° are 1:500:70,000. The relatively slow hydrolysis of $C_6H_5SCH_2Cl$ *vs.* CH_3SCH_2Cl suggests that resonance contributions of structures such as III to the transition state are of less importance than resonance typified by IV in the ground state. Again the reluctance of divalent sulfur to expand its valence shell is made evident.[4,8] Examination of Table I shows that the entropies of activation are not constant in the present series. Agreement with the Hammett relationship is achieved rather by a compensation in the activation energy and entropy terms.[12]


III


IV

Experimental[13]

Preparation of Chloromethyl Aryl Sulfides.—The method used was chlorination of the corresponding methyl aryl sulfides with sulfuryl chloride according to the procedure described previously for phenyl and *p*-methoxyphenyl chloromethyl sulfides.[14] The compounds were purified by vacuum distillation just prior to use except for *p*-nitrophenyl chloromethyl sulfide, which was crystallized from a solution of benzene and pentane. These compounds may cause a persistent and annoying dermatitis so that caution in handling is advisable; *p*-nitrophenyl chloromethyl sulfide appears to be particularly toxic. Table II summarizes the yields obtained and gives the analytical data.

Table II

Chloromethyl Aryl Sulfides

Chloromethyl sulfide	B.p. °C.	Mm.	Yield, %	Analyses, % Calcd. C	Calcd. H	Found C	Found H
m-$CH_3C_6H_4SCH_2Cl$	125	16		55.64	5.25	55.70	5.35
p-$CH_3C_6H_4SCH_2Cl$	125–126	15	83	55.64	5.25	55.98	5.52
m-$ClC_6H_4SCH_2Cl$	115–116	5	75	43.54	3.13	43.45	3.37
p-$ClC_6H_4SCH_2Cl$	128–129	12	68	43.54	3.13	43.76	3.22
p-$BrC_6H_4SCH_2Cl$	150	15		35.40	2.53	35.31	2.76
p-$NO_2C_6H_4SCH_2Cl$[a]	62–64[b]		91	41.28	2.97	41.38	3.00

[a] Calcd.: Cl, 17.48. Found: Cl, 17.30. [b] M.p.

Kinetic Runs.—Approximately 0.0005 mole of the chloromethyl sulfide was weighed into a flask and dissolved in 25 cc. of purified dioxane at 34.85°. To this solution 25 cc. of water (at 34.85°) was added quickly, the solution swirled vigorously to ensure complete mixing and a portion poured into an open conductivity cell. The resistance was measured with a Jones bridge.[15] Plots of log $R/(R - R_E)$ *vs.* time have excellent straight line relationships for all eight compounds. The rate constants were obtained by multiplying the slopes of the lines by 2.3. Activation energies were calculated by multiplying the slopes of the lines from the log *k* *vs.* 1/*T* plots by 4.58. Entropies of activation were calculated using the equation

$$-\Delta S\ddagger = 2.3R\left[\log\frac{k!}{k} - \log k_1 - \frac{E_\alpha - RT}{2.3RT}\right]$$

The rate constants for the hydrolysis of *t*-butyl chloride at 25° and 34.85° in 50% aqueous dioxane were found to be $k = 19 \times 10^{-5}$ and 58×10^{-5} sec.$^{-1}$, respectively. These constants check satisfactorily with the values of 18×10^{-5} and 63×10^{-5} sec.$^{-1}$ obtained previously in a slightly different medium by a titration method.[16]

Evanston, Illinois

---

[8] See A. Mangini and R. Passerini, *Experientia*, **12**, 49 (1956), and references cited therein for recent spectral evidence on this point.

[9] J. D. Roberts and W. T. Moreland, This Journal, **75**, 2167 (1953).

[10] A. M. Ward, *J. Chem. Soc.*, 445 (1927).

[11] E. D. Hughes, *ibid.*, 255 (1935).

[12] See J. E. Leffler, *J. Org. Chem.*, **20**, 1202 (1955), for an interesting discussion of this and related phenomena.

[13] Microanalyses were carried out by Mrs. Malcom White and Miss Joyce Sorenson.

[14] F. G. Bordwell and B. M. Pitt, This Journal, **77**, 572 (1955). We wish to thank Drs. Harry M. Andersen and Burnett M. Pitt for carrying out some of these preparations.

[15] We wish to express our appreciation to Dr. Stanley Langer and Professor Ralph Pearson for assisting us in making some of the initial measurements.

[16] S. Winstein and E. Grunwald, This Journal, **70**, 828 (1948). These values are for 48.61% of water in dioxane; our 50% by volume solution contains about 49% by weight of water.