formate; all of the halogen is hydrolyzed by water. At room temperature, the bis-chloroformate decomposes at the rate of 5% per day, with the evolution of carbon dioxide to form mustard gas, 2,2′-dichlorodiethyl sulfide. At 50°, decomposition occurs at the rate of 50% in 24 hours. When the compound is stored in a steel bomb at 20°, the pressure developed corresponds to a decomposition rate of 4 to 5% per day.

The conversion of the bis-chloroformate to 2,2′-dichlorodiethyl sulfide was demonstrated as follows. A small amount of the completely decomposed material was distilled. All of it boiled between 104-106° under a pressure of 17 mm. The distillate melted at 14°; the recorded melting point of mustard gas is 15°. The identity of the product was further confirmed by converting it to a sulfoxide (m.p. 108-108.5° uncor.); the recorded m.p. of the known sulfoxide is 109.5° cor.

2-Chloro-2′-fluorodiethyl Sulfide.—Anhydrous hydrogen fluoride (90 g., 4.5 moles) was condensed in a copper flask surmounted by a copper reflux condenser equipped at the top with a dropping funnel and an outlet tube. Through the dropping funnel, 110 g. (0.5 mole) of thiodiglycol bis-chloroformate was added while the copper flask was maintained at −80°. After the addition was complete, the acetone-Dry Ice-bath was replaced by an ice-bath and the temperature allowed to rise to 0°. As the mixture warmed, there was a copious evolution of hydrogen chloride. After the mixture had stood for a day at room temperature, the temperature of the water-bath was raised slowly to 75°; during this period there was a vigorous evolution of carbon dioxide and hydrogen fluoride. The flask was then cooled, connected to a copper condenser arranged for distillation, and the reaction product distilled under reduced pressure, using an iron tube as a receiver. The temperature of the flask was raised slowly to 170°; during this slow distillation, there was a continued evolution of carbon dioxide and hydrogen fluoride. The 48 g. of colorless distillate was washed with water, dried over anhydrous sodium sulfate, and distilled under a pressure of 20 mm. The fraction boiling below 100° (25 g.) was refractionated at 30 mm. pressure. Two fractions were obtained. The first (b.p. 62.5 to 91.5°) is probably a mixture of difluorodiethyl sulfide and 2-fluoro-2′-chlorodiethyl sulfide.

Anal. Calcd. for $C_4H_8SF_2$: F, 30.1. Found: F, 18.0. Calcd. for C_4H_8SFCl: Cl, 24.9. Found: Cl, 13.0.

The higher boiling fraction distilled sharply between 91.5° and 92.5° (30 mm.). It weighed 16 g., a yield of 22%, d^{20}_{20} 1.228, n^{25}_D 1.4852.

Anal. Calcd. for C_4H_8SFCl: F, 13.3; Cl, 24.9; S, 22.4. Found: F, 13.9; Cl, 24.4; S. 22.0.

The molecular refractivity calculated from the density and refractive index is 33.1; the value calculated from the sum of the atomic refractivities is 33.4.

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Acid-catalyzed Hydrolysis of Acetal and Chloroacetal[1]

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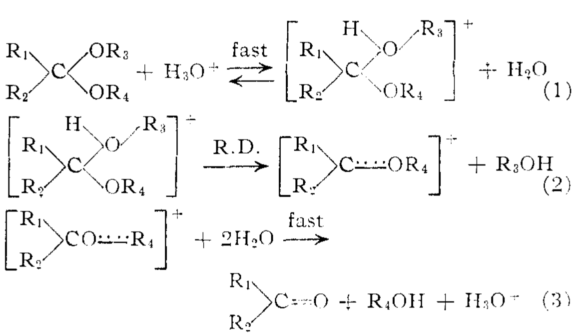
RECEIVED JANUARY 22, 1955

In the course of a study of the relative rates of acetal and ketal hydrolysis,[2] the acid dependence and solvent isotope dependence of the rate has been investigated. The results provide additional support for the mechanism of O'Gorman and Lucas.[3,4] This present work also extends the range of this mechanism, shown in equations 1–3, to a 50% dioxane–water mixture.



Previously reported evidence supporting this mechanism is: (1) there is no racemization at the alcohol carbon[3,5,6]; (2) the reaction exhibits specific hydronium ion catalysis in water[7]; (3) the second-order rate constant is larger by a factor of three in deuterated water[8]; (4) the rate is a linear function of H_0 with unit slope, indicating that no water molecules are firmly bound in the transition state.[9–13]

The present work shows that the rate of hydrolysis of chloroacetal, like that of methylal[9] and ethylal,[11] has the indicated acid dependence, that the hydrolysis of acetal in 50% aqueous dioxane is likewise faster by a factor of three when deuterated water is used, and that acetal hydrolysis in 50% aqueous dioxane is catalyzed only very slightly if at all by molecular formic acid.

Results

H_0 Dependence.—The acid dependence of the rate of hydrolysis of chloroacetal (in a solvent containing 4% dioxane) is shown in Table I. The last column of this table is expected by theory[12,13] to be approximately constant. As with methylal[9] and ethylal[11] the rates in the more concentrated hydrochloric acid solutions are somewhat higher than predicted by theory. With methylal this deviation was shown to be caused by a specific effect of chloride ion on the activity of the substrate,[9b] and the same is very likely true in the present case. Four concentrations of perchloric acid, sulfuric acid, and the most dilute hydrochloric acid give constant values of ($\log k_1 + H_0$) of acceptable precision.

Solvent Isotope Effect and General Acid Catalysis.—The second-order rate constant for the hydronium ion catalyzed hydrolysis of acetal in 50% dioxane-water[14] is 0.248 ± 0.004 l. mole^{-1} sec.$^{-1}$ at 25.0°. Replacing the water by an equal volume of D_2O increases this rate to 0.765 ± 0.050 l. mole^{-1} sec.$^{-1}$, indicating that a prototropic equilibrium

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[1] The work herein reported was carried out on Project NR055-328 between the Office of Naval Research and the Pennsylvania State University.

[2] M. M. Kreevoy and R. W. Taft, Jr., THIS JOURNAL, forthcoming.

[3] J. M. O'Gorman and H. J. Lucas, *ibid.*, **72**, 5489 (1950).

[4] C. K. Ingold, "Structure and Mechanism in Organic Chemistry," Cornell University Press, Ithaca, N. Y., 1953, p. 334.

[5] H. K. Garner and H. J. Lucas, THIS JOURNAL, **72**, 5497 (1950).

[6] E. R. Alexander, H. M. Busch and G. F. Webster, *ibid.*, **74**, 3173 (1952).

[7] J. N. Brönsted and W. F. K. Wynne-Jones, *Trans. Faraday Soc.*, **25**, 59 (1929).

[8] W. J. C. Orr and J. A. V. Butler, *J. Chem. Soc.*, 330 (1937).

[9] (a) D. McIntyre and F. A. Long, THIS JOURNAL, **76**, 3240 (1954); (b) F. A. Long and D. McIntyre, *ibid.*, **76**, 3243 (1954).

[10] Although W. W. Kaeding and L. Andrews, *ibid.*, **74**, 6189 (1952), interpreted their work in terms of another mechanism it is nicely consistent with eq. 1–3 if it is assumed that the intermediate oxo-carbonium ion reacts with ethanol about as readily as with water.

[11] P. M. Leininger and M. Kilpatrick, *ibid.*, **61**, 2510 (1939).

[12] L. P. Hammett and M. A. Paul, *ibid.*, **56**, 830 (1934).

[13] L. Zucker and L. P. Hammett, *ibid.*, **61**, 2791 (1939).

[14] The exact composition of this solvent mixture is described in the Experimental part.

TABLE I

ACID DEPENDENCE OF THE RATE OF HYDROLYSIS OF CHLOROACETAL AT 25°

Acid	Concn., M	$k_1 \times 10^4$, sec.$^{-1}$	Mean[a] dev. $\times 10^4$, sec.$^{-1}$	log k_1	H_0[b]	log $k_1 + H_0$
$HClO_4$	0.911	1.87	0.08	−3.73	−0.07	−3.80
$HClO_4$	1.82	5.35	0.25	−3.27	−0.57	−3.84
$HClO_4$	2.73	18.7	0.8	−2.73	−0.97	−3.70
$HClO_4$	3.64	48.8	4.4	−2.31	−1.42	−3.73
H_2SO_4	3.01	23.3	2.0	−2.63	−1.21	−3.84
HCl	3.68	55.2	1.2	−2.26	−1.17	−3.43
HCl	2.45	13.0	0.2	−2.89	−0.72	−3.61
HCl	1.23	2.20	..[c]	−3.66	−0.22	−3.88

[a] The method of determining rate constants is described in the Experimental section. [b] Determined from the values of L. P. Hammett and A. J. Deyrup, THIS JOURNAL, **54**, 2721 (1932); L. P. Hammett and M. A. Paul, *ibid.*, **56**, 827 (1934); densities taken from "Int. Crit. Tab.," Vol. 3, 1927, pp. 54–56. [c] No mean deviation could be determined because only one run was made under these conditions.

precedes the rate-determining step in this solvent mixture, as well as in pure water. This result also indicates that in this solvent the acid exists mostly as hydronium ion rather than as dioxonium ion.

Despite a thorough search, unambiguous evidence of general acid catalysis by formic acid in the 50% dioxane–water mixture was not obtained. The second-order rate constant for the hydronium ion catalyzed hydrolysis of acetal in 50% dioxane–water, 1.0 M in formic acid and 0.244 M in sodium bromide is 0.710 ± 0.033 l. mole^{-1} sec.$^{-1}$. This value was determined by adding sufficient perchloric acid to suppress the ionization of the formic acid. The resulting second-order constants were constant, within experimental error, from 0.01072 to 0.02143 M perchloric acid. Using conditions under which the ionization of formic acid is not suppressed (<0.001732 M $HClO_4$) the apparent ionization constant of formic acid in this medium was determined from the observed hydrolysis rates, since the hydrogen ion concentration is given by $k^{\text{obs}}/0.710$. The value so obtained is $2.24 \pm 0.05 \times 10^{-5}$. In the same way the apparent dissociation constant of formic acid in 50% dioxane–water[14] containing 1.0 M formic acid and 0.0488 M sodium bromide was found to be $1.03 \pm 0.02 \times 10^{-5}$ [15] and the second-order rate constant was 0.425 ± 0.033 l. mole^{-1} sec.$^{-1}$. The rate of hydrolysis was then determined in formic acid–formate buffers containing 1.0 M formic acid and either 0.244 M sodium bromide or sufficient sodium bromide to maintain an ionic strength of 0.05 M. The results obtained are shown in Table II.

TABLE II

RATE OF HYDROLYSIS OF ACETAL IN FORMIC ACID–FORMATE BUFFERS AT 25° IN DIOXAN–WATER

No.	Concn. formate, M	Concn. formic acid, M	Concn. NaBr, M	k (obs.), sec.$^{-1}$ $\times 10^4$	k (calcd.), sec.$^{-1}$ $\times 10^4$
1	0.05354	0.970	0.244	3.48	2.88
2	.02677	0.977	.244	6.43	5.92
3	.05354	1.012	.000	1.087	0.830
4	.02677	1.039	.024	2.055	1.701

The calculated rate constants shown in Table II were obtained with the assumption that, under conditions (1) and (2), the large excess of sodium bromide swamps out the effect of the sodium formate on the ionic strength and that the reaction exhibits specific hydronium ion catalysis under all conditions studied. The observed rates are consistently higher than the calculated ones. The difference may be due to catalysis by molecular formic acid but it is barely outside the uncertainty of the calculated values.

Irrespective of the value obtained for the ionization constant of formic acid, the ratio of observed rate constants and the value of the buffer quotients, ($HCOOH/HCOO^-$), should be the same if there is only specific hydronium ion catalysis. The quotient is 2.05 in both cases (no. 2/no. 1 and no. 4/no. 3). The ratios of observed rates is somewhat less in both cases. If the difference is ascribed to catalysis by formic acid, a catalytic constant of 0.70×10^{-4} l. mole^{-1} sec.$^{-1}$ is obtained from the rates with 0.244 M sodium bromide (no. 1 and 2) and 0.17×10^{-4} at ionic strength 0.05 (no. 3 and 4). The effects are very small and in examples (1) and (2) the ratio of rate constants may be less than 2.05 because of the small change (from 0.30 to 0.27) in the ionic strength. In any event the ratio of the proton catalytic coefficient to that of molecular formic acid cannot be less than 1.0×10^4 under either set of conditions.

Discussion

General Acid Catalysis and Mechanism.—It should be pointed out that neither in the present case nor in others[16,17] (for which less ambiguous evidence is available) does the finding of a small general acid catalytic coefficient necessarily throw doubt on other criteria (*i.e.*, rate dependence on H_0, or on solvent isotope composition) of the mechanism of the hydronium ion catalyzed reaction (*e.g.*, eq. 1–3).

There is no basic relationship known (although all authors on this subject have implied such a relationship) which demands that the mechanism of the general acid catalyzed portion of the reaction *necessarily* be perfectly analogous with that of the hydronium ion catalyzed portion. The general acid-catalyzed portion, for example, may proceed *via* a coördination complex or a protonated substrate–anion pair (of the Bjerrum type) while the hydronium ion catalyzed portion proceeds *via* the free protonated substrate (solvated only to the extent that ions such as ammonium ions are solvated). Although hydronium ion and formic acid are similar in kind (both may donate a proton) the large difference in degree makes the above possibility a very real one.

On this basis the frequently disregarded finding by Wynne-Jones[18] of a k_D/k_H ratio of 1.40 in the hydronium ion catalyzed hydrolysis of ethyl ortho-

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(15) These values can be compared with 4.45×10^{-6} which is the value obtained for this solvent in the absence of added salt by interpolating the results of Harned and co-workers, H. S. Harned and B. B. Owen, "The Physical Chemistry of Electrolytic Solutions," Reinhold Publ. Corp., New York, N. Y., 1950, p. 508, in the manner which they have suggested.

(16) R. H. DeWolfe and R. M. Roberts, THIS JOURNAL, **76**, 4379 (1954).

(17) C. G. Swain, *ibid.*, **74**, 4108 (1952).

(18) W. F. K. Wynne-Jones, *Trans. Faraday Soc.*, **34**, 245 (1938).

carbonate but a k_D/k_H ratio of 0.60 for the same reaction catalyzed by acetic acid is readily accommodated. This result is nicely consistent with the theories of solvent isotope effect[19] if the proton catalyzed reaction proceeds mostly through the protonated substrate, formed reversibly prior to the rate-determining step, while the acetic acid–catalyzed reaction involves principally a rate-determining proton transfer from acetic acid to the substrate.

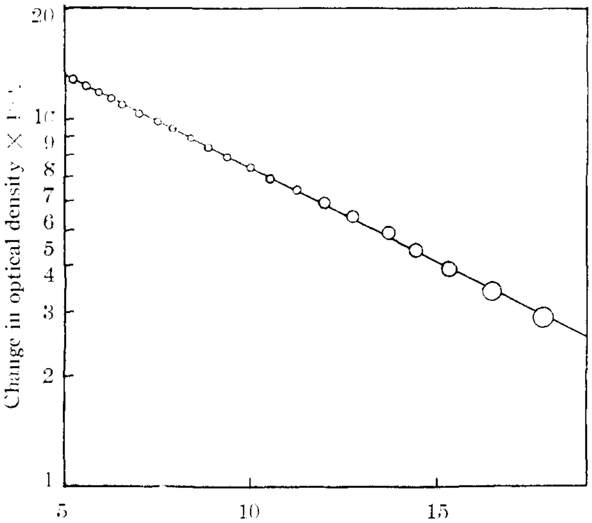

Fig. 1.—A sample rate experiment illustrating the precision of the first-order rate law for chloroacetal in strong acid: chloroacetal in 3.04 *M* $HClO_4$.

Experimental

Materials.—The chloroacetal was Eastman Kodak Co. "White Label," redistilled; b.p. 155.5–157.5°.[20] Acetal was prepared by the method of "Organic Syntheses"[21]; b.p., 104°. Dioxane was purified by the method of Vogel,[22] b.p. 99.5–100.5°, $n^{25}\text{D}$ 1.4227. All other materials were C.P. reagents.

Kinetic Procedure.—All rate measurements were made by following the appearance of the carbonyl peak in the ultraviolet with a Beckman D.U. spectrophotometer. For runs involving chloroacetal 4 cc. of a 10% (by volume) solution of chloroacetal in dioxane was diluted to 100 cc. with the required amount of distilled water and mineral acid. All the components were brought to 25° before mixing. Part of this solution was then transferred to a 10-cm. glass-stoppered quartz cell and intermittent determinations of the optical density were made until at least 75% of the reaction was complete. Except during the periods of measurement, the cell was kept in a thermostatic bath maintained at 25.00 ± 0.05°.

For each of the acetal experiments 4.00 ml. of dioxane, aqueous acid or buffer, and enough distilled water to give a total volume of 10.00 ml. were added to 1.00 ml. of 10% solution of acetal in dioxane. All the components were brought to 25° before mixing. Five ml. of dioxane made up to 10.00 ml. with water was analytically found to correspond to 49.6% dioxane and 50.4% water by weight. Part of the acetal solution was then transferred to a 1.00-cm. quartz cell and placed in a water-jacketed cell compartment for the course of the reaction. Water was pumped through the jackets from the thermostated bath and the cell temperature was found never to vary from the bath temperature by more than 0.1°. The optical density was determined at convenient intervals over a range of at least 75% of its total change.

A number of aldehydes, particularly chloroacetaldehyde, are in equilibrium with hydrates. Since the hydration reaction is relatively fast and reversible in the acidic solutions used[23] the optical density is directly proportional to the sum of the hydrolyzed material throughout the course of the reaction. Thus it is possible to obtain the hydrolysis rates directly from the time change in optical density of the system.[24] Within a given rate run the reactions follow a first-order rate law with very good precision (*ca.* 1%). Constants determined under identical conditions were averaged and the mean deviations are given.

A first-order plot typical of the rate runs is shown in Fig. 1. Under the conditions employed the reaction is essentially irreversible.[25]

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[19] R. P. Bell, "Acid-Base Catalysis," Oxford at the Clarendon Press, 1941, p. 143.

[20] All boiling points are uncorrected.

[21] H. Adkins and B. N. Nissen, "Organic Syntheses," Coll. Vol. I, John Wiley and Sons, Inc., New York, N. Y., 1941, p. 1.

[22] A. I. Vogel; "Practical Organic Chemistry," Longmans, Green and Co., New York, N. Y., 1948, p. 175.

[23] R. P. Bell and B. B. Darwent, *Trans. Faraday Soc.*, **46**, 34 (1950).

[24] A. A. Frost and R. G. Pearson, "Kinetics and Mechanism," John Wiley and Sons, Inc., New York, N. Y., 1953, p. 28.

[25] W. H. Hartung and H. Adkins, THIS JOURNAL, **49**, 2517 (1927), found that the equilibrium constant for the reaction (R = aliphatic)

$$RCHO + 2C_2H_5OH \rightleftharpoons RCH(OC_2H_5)_2 + H_2O$$

is always less than 0.15 in a medium somewhat different from the present one but nevertheless containing about 50% water. This makes it very unlikely that more than 1.5% of any acetal we have discussed remains unhydrolyzed at equilibrium.

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The Preparation and Hydrolysis of a Dineopentyl Acetal[1]

BY C. A. MACKENZIE AND J. H. STOCKER

RECEIVED FEBRUARY 7, 1955

Investigations of the hydrolysis of acetals derived from formaldehyde[2] or acetaldehyde[2–4] and optically active alcohols or glycols have led to a revision of the hydrolytic mechanism proposed earlier by Hammett.[5] The recovery of the non-racemized and non-inverted alcohol in all cases supported the conclusion that the alkyl carbonium ion corresponding to the alcohol cannot be an intermediate in the reaction.

Recent work in this Laboratory has demonstrated that hydrolysis of the dineopentyl acetal of benzaldehyde results in the recovery of neopentyl alcohol; no rearranged products were observed. While it has been pointed out that the scission of a neopentyl–oxygen bond may occur without rearrangement,[6] such reactions are considered to follow an S_N2 route; the existence of a neopentyl fragment with an incomplete octet gives rise to rearranged products.[7] Thus the present results are in agreement with the previously cited work with optically active acetals.

Experimental

Neopentyl alcohol was prepared by the lithium aluminum

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[1] Aided by a grant from the Research Corporation.

[2] H. K. Garner and H. J. Lucas, THIS JOURNAL, **72**, 5497 (1950).

[3] J. M. O'Gorman and H. J. Lucas, *ibid.*, **72**, 5489 (1950).

[4] E. R. Alexander, H. M. Busch and G. L. Webster, *ibid.*, **74**, 3172 (1952).

[5] L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., New York, N. Y., 1950, p. 304.

[6] L. H. Sommer, H. D. Bankmann and P. C. Miller, THIS JOURNAL, **73**, 3542 (1951).

[7] I. Dostrovsky, E. D. Hughes and C. K. Ingold, *J. Chem. Soc.*, 193 (1946).