PHYSICAL AND INORGANIC CHEMISTRY

Ionization and Proton Exchange of Amines in Acetic Acid. I. Kinetic and Equilibrium Properties for Solutions of Methylamine

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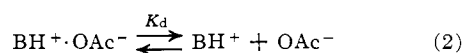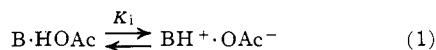
Methylamine in acetic acid is largely converted to methylammonium acetate. The ionization constant, $K_i = [CH_3NH_3{}^+OAc^-]/[CH_3NH_2]$, was estimated as 8×10^5. The degree of association of ion pairs was determined from freezing point data. Equilibrium constants (concentrations in mole/kg.) for the formation of double ion pairs were 0.0 for $CH_3NH_3{}^+OAc^-$, 0.6 for $(CH_3)_4N^+OAc^-$, 2.7 for $CH_3NH_3{}^+Cl^-$, and 1.7 for $CH_3NH_3{}^+OAc^-$ with $CH_3NH_3{}^+Cl^-$. Rates of exchange, R, of NH_3- protons with $COOH-$ protons of the solvent were measured by nuclear magnetic resonance techniques. Kinetic analysis established the rate law as $R = k[CH_3NH_3{}^+OAc^-]$. Kinetic parameters at 25° are: $k = 230$ sec.$^{-1}$, $\Delta H^* = 18.41$ kcal., $\Delta S^* = 14.0$ e.u. A slight decrease of k was observed with increasing concentration and can be interpreted as a viscosity effect. Addition of $(CH_3)_4NOAc$ produced a similar effect on k. Addition of CH_3NH_3Cl resulted in a larger reduction of k, which was ascribed to the formation of unreactive mixed ion pairs, $CH_3NH_3OAc\cdot(CH_3NH_3Cl)_n$.

The kinetics and reaction mechanisms of fast proton transfer reactions involving oxygen and nitrogen acids and bases have been studied extensively in water and also in a few other solvents of relatively high dielectric constant.[1] We now report nuclear magnetic resonance (n.m.r.) measurements of the rate of proton exchange between amines and acetic acid in glacial acetic acid, a solvent of low dielectric constant ($D = 6.22$ at 25°).[2] In part I (this paper) we give a kinetic analysis for methylamine in acetic acid and give quantitative estimates of ionization to methylammonium acetate and of the dissociation and association of the latter. In part II[3] we present kinetic data for a series of amines and discuss the reaction mechanism.

Kolthoff and Bruckenstein[4] have emphasized that it is necessary to represent base dissociation as occurring in two steps, ionization and dissociation, especially in solvents of low dielectric constant.[5] These authors report values of the equilibrium constants K_i and K_d for

$$B\cdot HOAc \underset{}{\overset{K_i}{\rightleftarrows}} BH^+\cdot OAc^- \qquad (1)$$

$$BH^+\cdot OAc^- \underset{}{\overset{K_d}{\rightleftarrows}} BH^+ + OAc^- \qquad (2)$$

several amines in glacial acetic acid. Values for methylamine are not reported, but may be estimated as follows.

The fairly good success of the method of acidity

functions[6a] in predicting pK_B for weak bases in water on the basis of ionization measurements in glacial acetic acid[6b] suggests that K_i in acetic acid is approximately proportional to K_B in water. Direct measurements, summarized in Table I, indicate that the proportionality

TABLE I

K_B IN WATER[a] COMPARED WITH K_i IN GLACIAL ACETIC ACID. DATA AT 25°

Base	K_B in water	K_i in HOAc	K_i/K_B $\times 10^{-9}$
Pyridine	1.7×10^{-9}	5.4^b	3.2
Bis-p-dimethylamino-azobenzene	9×10^{-11}	0.10^b	1.1
2,5-Dichloroaniline	3×10^{-13}	$(4 \times 10^{-4})^{b,c}$	$ca.$ 1
Methylamine	4.2×10^{-4}	8×10^5	$(2)^d$

[a] References for K_B: H. C. Brown, D. H. McDaniel, and O. Häfliger in "Determination of Organic Structures by Physical Methods," E. A. Braude and F. C. Nachod, Ed., Academic Press, Inc., New York, N. Y., 1955, Chapter 14; M. Kilpatrick and C. A. Arenberg, J. Am. Chem. Soc., **75**, 3812 (1953); H. H. Jaffe, Chem. Rev., **53**, 200 (1953). [b] Ref. 7. [c] Assume K_d is equal to K_d for pyridinium acetate. [d] Average for other bases. It is possible that this estimate of K_i/K_B is too small since K_i for methylamine is so large, see footnote 11 of ref. 3.

constant is about 2×10^9. On this basis, K_i for methylamine is about 8×10^5 and we may conclude that methylamine is almost completely ionized in acetic acid.

The available data for K_d of acetate salts in acetic acid[7] suggest that, although K_d varies with the nature of the cation, the variation is relatively small. We

(1) For a recent review, see L. De Maeyer and K. Kustin, Ann. Rev. Phys. Chem., **14**, 5 (1963).

(2) W. Dannhauser and R. H. Cole, J. Am. Chem. Soc., **74**, 6105 (1952).

(3) E. Grunwald and E. Price, ibid., **86**, 2970 (1964).

(4) I. M. Kolthoff and S. Bruckenstein, J. Am. Chem. Soc., **78**, 1 (1956).

(5) In discussions of reaction mechanism it is more useful to represent the un-ionized base realistically as B·HOAc rather than formally as B.

(6) (a) L. P. Hammett, Chem. Rev., **16**, 67 (1935); (b) N. F. Hall and W. F. Spengeman, J. Am. Chem. Soc., **62**, 2487 (1940); H. Lemaire and H. J. Lucas, ibid., **73**, 5198 (1951); T. L. Smith and J. H. Elliott, ibid., **75**, 3566 (1953); S. Bruckenstein, ibid., **82**, 307 (1960).

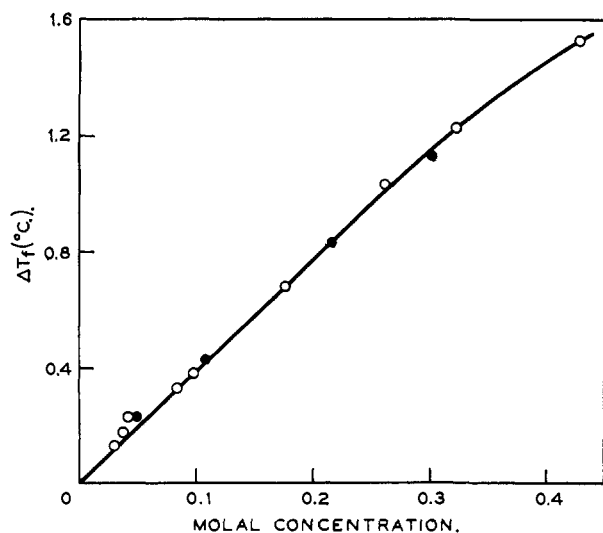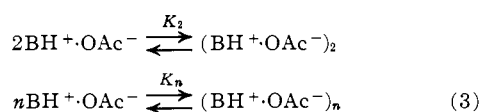(7) S. Bruckenstein and I. M. Kolthoff, ibid., **78**, 10, 2974 (1956).

Fig. 1.—Freezing point lowering in glacial acetic acid: O, naphthalene; ●, benzoic acid.

TABLE II

FREEZING POINT STUDIES OF ASSOCIATION OF ELECTROLYTES IN GLACIAL ACETIC ACID

Solute[a]		m_f[b]	ΔT_f	Σm_i	K_2 or K (m^{-1})
CH₃NH₃⁺OAc⁻ (A)		0.127	0.42	0.11	0.0
		.261	1.03	.27	
		.417	1.57	.43	
(CH₃)₄N⁺OAc⁻		.25	0.77	.20	0.6
		.27	0.91	.24	
		.34	1.17	.31	
Imidazole		.126	0.44	.124	0.0
CH₃NH₃Cl (B)		.048	.18	.042	2.7
		.112	.38	.096	
		.232	.62	.162	
		.392	.87	.230	
		.442	.95	.252	
		.603	1.17	.312	
(A) 0.126	(B) .425	1.32	.353		1.7
(A) .126	(B) .218	0.97	.257		
(A) .261	(B) .256	1.47	.397		
Triphenyl-n-amyl phosphonium iodide[e]					40
Tetraethylammonium chloride[e]					15
Tetramethylammonium thiocyanate[e]					10
Sodium iodide[e]					4
IH⁺OAc⁻ + H₃O⁺OAc⁻[d]					2.0
IH⁺OAc⁻ + C₅H₅NH⁺OAc⁻[d]					1.1

[a] (A) = CH₃NH₃⁺OAc⁻, (B) = CH₃NH₃⁺Cl⁻. [b] m_f = formula weights of solute per kg. of solvent. [e] Ref. 8. [d] Ref. 7; I = bis-p-dimethylaminoazobenzene; concentrations in mole/l.

may therefore expect to predict unknown K_d values, at least to the correct order of magnitude, by analogy with known K_d values for structurally related acetate salts. On this basis, K_d for methylammonium acetate is estimated by analogy with K_d for pyridium acetate to be about 10^{-6} (mole/l.).

In the present work the formal concentrations of amine range from 0.02 to 1 M. Hence the concentrations of the free ion, $CH_3NH_3^+$, are expected to be no more than 1% of the formal concentrations. On the other hand, the association of ion pairs to higher aggregates (eq. 3) might well be considerable in this concentration range

$$2BH^+\cdot OAc^- \overset{K_2}{\rightleftharpoons} (BH^+\cdot OAc^-)_2$$

$$nBH^+\cdot OAc^- \overset{K_n}{\rightleftharpoons} (BH^+\cdot OAc^-)_n \qquad (3)$$

The available data for acetic acid solutions[7,8] show that this phenomenon is highly specific, but are not sufficiently extensive to permit quantitative predictions of association constants. We therefore decided to study ion-pair association directly for the solutes of interest in the present work.

Ion-Pair Association from Freezing Points—In Fig. 1, values of the freezing point lowering, ΔT_f, of acetic acid are plotted vs. the formal solute molality, m_f, for the "normal" solutes naphthalene and benzoic acid. The points define a single line. Up to $m_f \approx 0.3$, ΔT_f is proportional to m_f, the slope being 3.69. At higher concentrations the plot shows slight negative curvature. On the basis of the heat of fusion[9a] and freezing point of pure acetic acid, the thermodynamically predicted molal cryoscopic constant is 3.58.

Assuming that naphthalene and benzoic acid are normal, unassociated solutes, we have used a large-scale plot of Fig. 1 to derive actual solute concentrations, Σm_i, in solutions of electrolytes in acetic acid. The results are listed in Table II. It is seen that ion association is quite insignificant for $CH_3NH_3^+OAc^-$ and imidazole (largely ionized into $C_3H_3N_2H^+OAc^-$), barely significant for $(CH_3)_4N^+OAc^-$, but quite con-

siderable for $CH_3NH_3^+Cl^-$. Values of ΔT_f for mixtures of methylammonium acetate and chloride are smaller than additive for the components, showing that there is some mixed association.

The solute concentration, Σm_i, is expressed as a function of the concentration, m_1, of unassociated ion pairs and of association constants by eq. 4. The formal concentration, m_f, is expressed similarly by eq. 5. Dissociation to free ions has been neglected.

$$\Sigma m_i = m_1 + K_2 m_1^2 + \ldots + K_n m_1^n + \ldots \quad (4)$$

$$m_f = m_1 + 2K_2 m_1^2 + \ldots + nK_n m_1^n + \ldots \quad (5)$$

At moderate concentrations the most important association product is the double ion pair. If we therefore use approximate values for the higher association constants as shown in eq. 6, we obtain eq. 7 and 8 in which K_2 is the only parameter.[9b]

$$K_3 = K_2^2; \quad K_n = K_2^n \qquad (6)$$

$$\Sigma m_i = m_1/(1 - K_2 m_1) \qquad (7)$$

$$m_f = m_1/(1 - K_2 m_1)^2 \qquad (8)$$

Treatment of our data on the basis of eq. 7 and 8 gives adequate fit. Resulting values of K_2 are listed in Table II.

Mixed association of methylammonium acetate and chloride was treated as follows. Because of the low tendency of the acetate toward self-association, we assumed that the only mixed association equilibria of importance are of the type shown in eq. 9.

$$BH^+OAc^- + (BH^+Cl^-)_n \overset{K}{\rightleftharpoons} BH^+\cdot OAc^-\cdot(BH^+Cl^-)_n \quad (9)$$
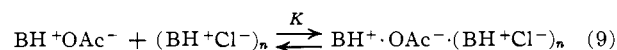
We further assumed that the data can be represented with a single association constant, K, regardless of n. The average value of K calculated from the data for the mixtures on this basis is given in Table II. Although

(8) P. Walden, Z. physik. Chem., **A94**, 310 (1920).
(9) (a) G. S. Parks and K. K. Kelley, J. Am. Chem. Soc., **47**, 2092 (1925);
(b) E. N. Lassettre, Chem. Rev., **20**, 281 (1937).

the fit of the data is adequate, the accuracy of K is low.

Table II also lists some additional values of K_2, calculated by us from the data of Walden,[8] and some values of K as reported by Kolthoff and Bruckenstein.[7] It is seen that the association constants cover a wide range. Judging by this limited sample, ion-pair association is favored by an increase in ion size, particularly of the cation.

The strikingly low values of K_2 for certain acetate ion pairs can be explained on the basis of the following theory. Suppose that the anion in the ion pair forms a stable complex with at least one additional acetic acid molecule, and that the negative charge in the resulting solvated ion pair is dispersed by very rapid proton transfer (eq. 10).

$$BH^+ \cdot AcO^- \cdot HOAc = BH^+ \cdot AcOH \cdot OAc^- \qquad (10)$$
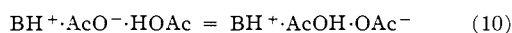
Then, if the rate of charge dispersal is fast compared to the dielectric relaxation time of the ion pairs, the result will be a reduction of the effective dipole moment and, hence, of the tendency toward ion-pair association.

Ionized 2:1 complexes between acetic acid and amines are known to be quite stable in carbon tetrachloride and chloroform.[10] Furthermore, crystal structures have been determined for several acid salts of empirical formula $MA \cdot HA$, where M is an alkali metal and HA a carboxylic acid. The interatomic distances deduced from X-ray diffraction indicate generally that there is a strong hydrogen bond between A^- and HA, and in some of the solid compounds the A-groups in $A^- \cdot HA$ have even been reported to be crystallographically equivalent.[11] Kinetic salt effects observed in certain acetolysis reactions that involve carbonium ion-pair intermediates are also of interest in this connection. "Normal" salt effects of lithium and diphenylguanidinium acetate in these reactions are appreciably smaller than those of the corresponding perchlorates, indicating that the acetate ion pairs are, in effect, less polar.[12,13]

Kinetic Results

Interpretation of N.m.r. Data.—Rates of proton exchange could be measured conveniently by means of n.m.r. data for the carboxyl protons of acetic acid and, in some cases, for the CH_3- protons of methylammonium ion. Values of some relevant chemical shifts and spin-spin interactions are summarized in Table III. The carboxyl n.m.r. line was broadened significantly upon addition of methylammonium acetate, but not upon addition of sodium acetate or tetramethylammonium acetate. The broadening was therefore ascribed to NH–COOH proton exchange. For a given solution the broadening increased with an increase in temperature between 15 and 45°. Since the broadening was sufficiently small, it could be treated[14] as "lifetime broadening," eq. 11.

$$\Delta = \frac{1}{T_2'} - \frac{1}{T_2} = \frac{R}{[HOAc]} \qquad (11)$$

(10) G. M. Barrow and E. A. Yerger, *J. Am. Chem. Soc.*, **77**, 4474 (1955).

(11) H. N. Shrivastava and J. C. Speakman, *J. Chem. Soc.*, 1151 (1961).

(12) S. Winstein and E. Clippinger, *J. Am. Chem. Soc.*, **78**, 2784 (1956).

(13) A. H. Fainberg and S. Winstein, *ibid.*, **78**, 2763 (1956).

(14) The general relationship between Δ and R is given in ref. 3. For the highest values of Δ/p_{NH} reported in this paper, eq. 11 becomes inaccurate by about 2%, and corresponding corrections were made.

TABLE III

N.M.R. CHEMICAL SHIFTS AT 60 Mc./SEC., SPIN–SPIN INTERACTIONS, AND MEASURES OF VISCOSITY FOR SOLUTIONS OF METHYLAMMONIUM SALTS IN GLACIAL ACETIC ACID, 25°

(1) Chemical shift of COOH *vs.* CH_3- protons of acetic acid, c.p.s.

   (a) $CH_3NH_3^+OAc^-$, 0–0.5 M $\delta_{COOH} - \delta_{CH_3} = -576.6 - 22.5c$

   (b) $CH_3NH_3^+Cl^-$, 0–1.2 M $\delta_{COOH} - \delta_{CH_3} = -576.6 + 10.6c$

(2) Chemical shift of COOH– protons *vs.* NH– protons, c.p.s.

   (a) $CH_3NH_3^+OAc^-$, 0–0.5 M $\delta_{COOH} - \delta_{NH} = -279$

   (b) $CH_3NH_3^+Cl^-$, 0–1.2 M $\delta_{COOH} - \delta_{NH} = -270 + 35c - 16c^2$

(3) CH_3-NH_3 proton spin–spin interaction, 6.17 c.p.s. (measured in CH_3NH_3Cl)

(4) $N^{14}-H$ spin–spin interaction, 52 c.p.s. (measured in $NH_4^+OAc^-$)

(5) Bulk viscosity

   (a) $CH_3NH_3^+OAc^-$, 0–2.2 M $\log(\eta/\eta_0) = 0.480c - 0.020c^3$

   (b) $CH_3NH_3^+Cl^-$, 0–1.0 M $\log(\eta/\eta_0) = 0.265c$

(6) T_1 relaxation time of carboxyl protons; $T_1° = 3.12$ sec.

   (a) $CH_3NH_3^+OAc^-$, 0–0.5 M $\log(T_1°/T_1) = 0.462c$

   (b) $(CH_3)_4^+OAc^-$, 0–0.2 M $\log(T_1°/T_1) = 0.382c$

   (c) $CH_3NH_3^+Cl^-$, 0–0.6 M $\log(T_1°/T_1) = 0.253c$

(7) T^1 relaxation time of the N^{14} nucleus in sec.

   (a) $CH_3NH_3^+Cl^-$, 0–1.2 M $\log(1/T^1) = 2.28 + 0.23c$

   (b) $CH_3NH_3^+OAc^-$, 0.465 M $1/T^1 \approx 400$

Here Δ is the exchange broadening, R the rate of NH–COOH proton exchange (in g.-atom/l. sec.), and [HOAc] the moles of acetic acid per liter of solution; Δ was obtained from spin-echo measurements of the transverse (T_2') and of the longitudinal (T_1) relaxation times of the COOH-protons in the solutions. The value of T_2' used in eq. 11 was the directly measured value. The quantity T_2 was taken as $T_1/1.14$, since there was a systematic difference of 14 ± 2% between T_2 and T_1 for the carboxyl protons in a number of reference liquids. The reference liquids were pure acetic acid over a wide temperature range and dilute solutions of perchloric acid and of sodium acetate in acetic acid at 25°. We do not know whether this small difference is real or whether it is a systematic instrumental error. A similar difference between T_1 and T_2, measured on the same instrument, has been noted for water[15a] but not for methanol.[15b]

The rate of NH–COOH exchange has been compared with the *total* rate of exchange of NH– protons. The latter is obtained from the n.m.r. spectrum of the CH_3- protons of methylammonium acetate, which is a spin–spin quadruplet in the absence of exchange due to interaction with the NH_3- protons. Under the present conditions of proton exchange this quadruplet was collapsed into a single broad line, and the rate was calculated as described previously.[16] The results are listed in Table IV. The two rates are equal within experimental error and thus show that the NH–

TABLE IV

RATE OF NH–COOH PROTON EXCHANGE COMPARED WITH TOTAL RATE OF NH– PROTON EXCHANGE FOR METHYLAMMONIUM ACETATE IN ACETIC ACID AT 25°

	Rate of proton exchange, g.-atom/l. sec.	
$[CH_3NH_3^+OAc^-]$	NH and COOH	NH, total
0.175	34.6 ± 1.0	34.1 ± 1.0
0.465	84 ± 8	78 ± 2

(15) (a) Z. Luz and S. Meiboom, *J. Chem. Phys.*, **39**, 366 (1963); (b) E. Grunwald, measurements at 25°.

(16) E. Grunwald, A. Loewenstein, and S. Meiboom, *J. Chem. Phys.*, **27**, 630 (1957).

protons exchange at a significant rate only with COOH–protons.

Kinetic Analysis.—Kinetic results for methylammonium acetate at 25° are summarized in Table V. In this table and throughout the following kinetic analysis we use the symbol k to denote the pseudo-first-order rate constant, $k = R/c$, where c is the molar concentration of the formal species, methylammonium acetate. It is seen that k decreases slightly but significantly with increasing concentration. The theory that this decrease is due to the formation of less reactive ion-pair association products must be ruled out in view of the freezing point data (Table II). Two interpretations therefore suggest themselves: (i) The reaction is precisely first order in BH^+OAc^-, and the decrease of k is a medium effect; (ii) Proton exchange is significant both for the ion pairs, BH^+OAc^-, and for the free ions, BH^+.

TABLE V

KINETIC RESULTS FOR PROTON EXCHANGE OF
METHYLAMMONIUM ACETATE IN ACETIC ACID AT 25°

Molar concn.	k, sec.$^{-1}$	$k\eta/\eta_0$, sec.$^{-1}$	$176 + 8.17c^{-1/2}$
0.0208	230 ± 10	235	233
.0466	217 ± 6	228	214
.0922	202 ± 6	224	206
.175	198 ± 6	240	196
.267	197 ± 10	265	192
.465	180 ± 20	300	188

According to hypothesis ii the rate law is given by eq. 12

$$k = R/c = k' + k''K_d^{1/2}c^{-1/2} \quad (12)$$
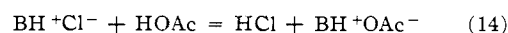
where k' and k'' are rate constants for BH^+OAc^- and BH^+, respectively. Judging by data for structurally related acetate ion pairs,[7] $K_d = (BH^+)(OAc^-)/(BH^+OAc^-)$ is about 10^{-6}.

As a matter of fact, the empirical eq. 13 fits the data well, as shown in the last column of Table V. We
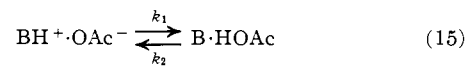
$$k = 176 + 8.17c^{-1/2} \quad (13)$$

doubt, however, that the formal resemblance of eq. 13 to 12 implies the corresponding mechanism. If it did, we would infer from the coefficient of $c^{-1/2}$ in eq. 13 that $k'' = 8.17K_d^{-1/2} \approx 8000$ sec.$^{-1}$. This value seems implausibly high compared to the value 0.9 sec.$^{-1}$ of the first-order rate constant for proton exchange of methylammonium ion with water.[17a] We regard the value as implausible because acetic acid is much less basic than water and methylammonium ion does not have unshared electron pairs.[17b] To obtain further evidence, we have measured the rate of NH proton exchange for 0.8 M CH_3NH_3Cl in acetic acid. Judging by data for model substances,[7] the *lower limit* to K_d for CH_3NH_3Cl is 10^{-7}. If k'' were indeed 8000 sec.$^{-1}$, the *lower limit* to R would therefore be 2.3 g.-atoms/l. sec. The actual rate is just below the threshold of easy measurability by the n.m.r. method. However, the experimental *upper limit* to the rate of NH proton exchange is 1.0 g.-atom/l. sec. Furthermore, a rate of the order of 0.1 to 0.3 g.-atom/l. sec. is to be expected under these conditions due to $CH_3NH_3^+OAc^-$

(17) (a) E. Grunwald, P. J. Karabatsos, R. A. Kromhout, and E. L. Purlee, *J. Chem. Phys.*, **33**, 556 (1960); (b) A. I. Brodskii, *Zh. Obshch. Khim.*, **24**, 413 (1954).

ion pairs produced by acetolysis.[18] It seems safe

$$BH^+Cl^- + HOAc = HCl + BH^+OAc^- \quad (14)$$

to conclude that hypothesis ii is inconsistent with the facts.

Assuming, therefore, that the NH–COOH proton exchange rate is precisely first order in the ion pair $CH_3NH_3^+OAc^-$, we regard the two-step kinetic scheme, eq. 15 and 16, as the most plausible representation of the reaction mechanism.[19a]

$$BH^+ \cdot OAc^- \underset{k_2}{\overset{k_1}{\rightleftarrows}} B \cdot HOAc \quad (15)$$

$$AcOH^* + B \cdot HOAc \overset{k_H}{\longrightarrow} AcOH^* \cdot B + HOAc \quad (16)$$

The reaction rate is then given by eq. 17. If k_H

$$R = k_1[BH^+ \cdot OAc^-] \frac{k_H}{k_2 + k_H} \quad (17)$$

$\gg k_2$, the observed rate constant k is equal to k_1. If $k_2 \gg k_H$, $k = k_1k_H/k_2 = k_H/K_i$, where K_i is the ionization constant (see eq. 1). We believe the second inequality to be correct but shall postpone discussion to the next paper.[3]

Dependence on Viscosity.—In this section we shall suggest that a theory in which $k = k_H/K_i$ could be correct by showing that the decrease in k parallels that of the viscosity, η/η_0, up to moderate concentrations.[19b] Furthermore, we shall show that the ion-pair aggregates, $BH^+OAc^- \cdot (BH^+Cl^-)_n$, are apparently unreactive.

We have obtained three separate measures of the viscosity: (i) η, the macroscopic viscosity of the solution; (ii) T_1, the longitudinal relaxation time of the carboxyl protons, a rather complex measure of the microscopic viscosity opposing the tumbling and diffusion of the acetic acid molecules in their various microscopic environments; and (iii) T^1, the longitudinal relaxation time of the N^{14} nucleus. $1/T^1$ is a rather good index of the microscopic viscosity opposing the tumbling of the molecular species containing the N^{14} nucleus.[20] In the concentration range of interest, plots of log (η/η_0), log $(T_1°/T_1)$, or log $(1/T^1)$ vs. c are linear, and the empirical slopes are listed in Table III. For both CH_3NH_3Cl and CH_3NH_3OAc, the slope obtained for the concentration dependence of log (η/η_0) is very nearly equal to that obtained for the logarithm of the microscopic viscosity[21] as measured either by $1/T_1$ or by $1/T^1$. We therefore believe that the bulk viscosity is also a fairly good index of the kind of micro-

(18) The equilibrium constant for eq. 14 is of the order of 10^{-6} or 10^{-7} on the basis of data in ref. 4 and 7.

(19) (a) Bimolecular reaction between free BH^+ and free OAc^- ion is kinetically indistinguishable from a first-order reaction of the ion pair. However, it is necessary to make a distinction between the two processes only if the transition state for the reaction cannot be reached directly from the ion pairs but must be reached by way of the free ions. This is evidently not the case here, since in the ion pairs the functional groups CO_2^- and HN^+ are already connected by a hydrogen bond. Furthermore, we can show that the assumption that the transition state can be reached only from the free ions leads to an improbably high value for the rate constant of the reverse reaction, $B + HOAc \rightarrow BH^+ + OAc^-$. (b) See, for example, M. T. Emerson, E. Grunwald, M. L. Kaplan, and R. A. Kromhout, *J. Am. Chem. Soc.*, **82**, 6307 (1960).

(20) W. B. Moniz and H. S. Gutowsky, *J. Chem. Phys.*, **38**, 1155 (1963).

(21) The similarity is not so good for the temperature dependence of η and of T_1 for pure acetic acid. Between 15 and 45°

$$\log \eta \text{ (poise)} = 568/T - 3.858$$

$$-\log T_1 \text{ (sec.)} = 417/T - 1.896$$

scopic viscosity opposing the replacement of one acetic acid molecule by another in the solvation shell of B as shown in eq. 16.

Values of $k\eta/\eta_0$ for methylammonium acetate are listed in column 3 of Table V. It is seen that the factor η/η_0 compensates quite well for the decreasing trend in k, up to $c \sim 0.2\ M$. At higher concentrations, values of $k\eta/\eta_0$ seem to increase.

Preliminary results indicate that this increase might be due to an additional reaction of higher than first order in methylammonium acetate. For example, at $c = 5\ M$, the NH– and COOH– proton resonance is collapsed into a single sharp line with $T_2' \approx T_2$, and $k > 1 \times 10^6\ \mathrm{sec.}^{-1}$ at 25°.

Kinetic data for mixtures of electrolytes are given in Table VI. Addition of $(CH_3)_4N^+OAc^-$ decreases the value observed for k, but $k\eta/\eta_0$ is again quite constant, being independent of the concentration of added salt. It will be recalled that, on the basis of the freezing point data, $(CH_3)_4N^+OAc^-$ does not have a strong tendency to form ion-pair aggregates.

TABLE VI

SALT EFFECTS ON PROTON EXCHANGE RATES FOR METHYLAMMONIUM ACETATE IN ACETIC ACID AT 25°

Concentration of		k,	$k\eta/\eta_0$,	$\alpha^{-1}k\eta/\eta_0$,
[BH+OAc–]	[Added salt]	sec.$^{-1}$	sec.$^{-1a}$	sec.$^{-1b}$
1. Added salt = $(CH_3)_4N^+OAc^-$				
0.0963	....	207	230	...
.0963	0.0328	204	234	...
.0963	.0721	188	223	...
.0963	.1377	179	225	...
.267	....	197	265	...
.267	0.0372	191	265	...
.267	.0881	185	269	...
.267	.163	177	274	...
2. Added salt = $CH_3NH_3^+Cl^-$				
0.0934	0.0544	195	223	239
.0934	.1250	168	201	232
.0934	.274	123	165	213
.175	.63	76	136	205

[a] Calculated from log $\eta/\eta_0 = a_1c_1 + a_2c_2$. Values of the slopes a_1, a_2 from Table III. $a = 0.382$ for $(CH_3)_4N^+OAc^-$. [b] Calculation of α (= fraction of BH+OAc– not complexed to BH+Cl–)

$$\alpha = \left(1 + \frac{1.7m_1}{1 - 2.7m_1}\right)^{-1}$$

where m_1 = molal concentration of uncomplexed BH+Cl–, 1.7 = K, 2.7 = K_2 (Table II). Calculate m_1 by successive approximations from: $m_1(1 + 1.7\alpha m_{\mathrm{BHOAc, formal}}) = m_{\mathrm{BHCl, formal}}(1 - 2.7m_1)^2$.

Addition of CH_3NH_3Cl causes a marked reduction in k, which is only partially compensated by the factor η/η_0. Since methylammonium chloride has been shown to form complexes with methylammonium acetate, with the assumed formula BH+OAc–·(BH+Cl–)$_n$, we attempted to account for the additional decrease in k on the basis that the mixed complexes are completely unreactive. Accordingly, the fraction, α, of uncomplexed BH+OAc– was calculated for each solution, using the association constants in Table II. As shown by the last column in Table VI, the values of $k\eta/\alpha\eta_0$ are indeed quite constant.

Activation Parameters.—For 0.09 M methylammonium acetate, k was measured at 15, 25, 35, and 45°. The plot of log (k/T) vs. T^{-1} was nicely linear, T being the absolute temperature. Values of the activation

parameters, calculated according to the transition state theory, are: $\Delta H^* = 18.41$ kcal., $\Delta S^* = 14.0$ e.u.

Experimental

Materials.—Glacial acetic acid (Merck) was dried and purified in an all-glass apparatus by refluxing for 12 hr. over chromium trioxide (2 g./1000 ml.) and acetic anhydride, followed by distillation.[22] The acetic anhydride was just sufficient to react with the water in the unpurified reagent. The first fraction of 20% was rejected, and the next fraction (20%) was used as solvent to prepare solutions.

The water content of the acid was determined by titration with Karl Fischer reagent in the presence of excess methanol (5:1).[23] The methanol was titrated first; then acetic acid was added and the mixture titrated. Acetic acid, taken directly from the bottle, was usually between 0.03 and 0.04 M in water. However, after purification, 1 drop of Karl Fischer reagent was sufficient to produce an end point ($H_2O < 0.005\ M$).

Reagent grade methylamine hydrochloride was recrystallized twice from pure ethanol and dried *in vacuo* at 25°.

Methylamine was obtained from Matheson and Co. in the form of a compressed gas.

Tetramethylammonium acetate was prepared[24] by neutralizing a solution of the hydroxide with acetic acid. Isolation of dry solid product was difficult, however, because of the great hygroscopicity of this substance. After evaporation of the solution on a steam bath, the product was dried for 2 weeks over magnesium perchlorate at a pressure of 0.01 mm., at which time it still had not reached constant weight. The material was then dried in 37 hr. to constant weight in an Abderhalden pistol at 80° over phosphorus pentoxide at pressures below 0.01 mm.; equivalent weight, determined by acid–base titration in glacial acetic acid: obsd. 140 ± 14, calcd. 133.

Solutions Preparation.—A stock solution of 3 M methylammonium acetate in glacial acetic acid was prepared by bubbling methylamine into purified acetic acid in a special all-glass apparatus designed to protect the solution from atmospheric water. Subsequent solutions of the desired concentration were prepared from this stock solution by dilution. The solutions were then standardized by titration with 0.5 N HClO$_4$ in acetic acid, using bromophenol blue indicator.[25]

Mixtures of methylammonium acetate and methylammonium chloride of the desired concentrations were prepared by adding the required volume of the acetate of known concentration to a weighed quantity of the hydrochloride. In the same manner, solutions of methylammonium acetate and tetramethylammonium acetate were prepared, but tetramethylammonium acetate was dried *in vacuo* until a constant weight (within 0.2 mg.) was obtained. The total acetate concentration of the final solutions was checked by titration with 0.5 N HClO$_4$ in glacial acetic acid.

All experiments were done with air-saturated solutions.

Freezing Point Determinations.—Freezing points were determined in an apparatus of the Beckmann type. The thermometer was graduated in 0.1° intervals and could be read to 0.02°. The freezing point of each solution and of pure acetic acid was determined from a temperature vs. time plot.[26] The molal concentration, m, of each solution was corrected for the fraction, f, of acetic acid that had solidified (eq. 18, 19).

$$f = \Delta T \cdot C_p/L_f = 0.01\Delta T \text{ for HOAc} \qquad (18)$$

$$m_{\mathrm{cor}} = m_{\mathrm{uncor}}/(1 - f) \qquad (19)$$

In eq. 18, ΔT is the amount of supercooling (in °C.) of the liquid prior to freezing, C_p is the molar heat capacity of liquid acetic acid, and L_f is the molar heat of fusion. Values of ΔT were usually less than 1°. The freezing point of pure acetic acid was found to be 16.56° (reported[9] 16.6°).

Viscosities.—The viscosities of solutions of methylammonium acetate and methylammonium chloride relative to that of pure acetic acid were determined in the standard way from densities and flow times, using an Ostwald modification of the Poiseuille

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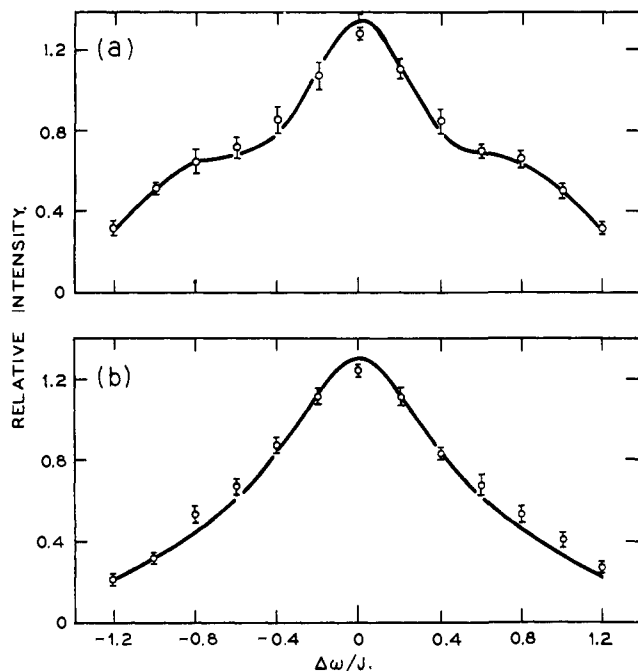(26) A. R. Glasgow, Jr., A. J. Streiff, and F. D. Rossini, J. Res. Natl. Bur. Std., 35, 355 (1945).

Fig. 2.—NH– proton resonance spectrum for (a) 0.388 M and (b) 1.19 M methylammonium chloride in acetic acid at 25°. Solid curves are calculated for (a) $1/T^1 = 230$ sec.$^{-1}$; (b) $1/T^1 = 360$ sec.$^{-1}$.

viscosimeter. The measurements were made at $25 \pm 0.05°$ and were precise to $\pm 0.5\%$.

N.m.r. Measurements.—All n.m.r. measurements were made on the high resolution n.m.r. spectrometer of Meiboom, operating at 60 Mc./sec. A brief description of the instrument and of methods of frequency stabilization, temperature control, and spin-echo determination of T_1 and T_2' relaxation times has been

given elsewhere.[27] In measurements of T_1 and T_2' of the carboxyl protons, the radiofrequency field was adjusted according to the method of Alexander[28] so as to eliminate interference from the CH_3– protons of acetic acid. Values of $1/T_2'$ were precise to ± 0.012 sec.$^{-1}$, up to about 4 sec.$^{-1}$, but the precision decreased rapidly above this value. Values of $1/T_1$ were usually precise to ± 0.005 sec.$^{-1}$. Typical values of Δ ranged from 0.3 to 3 sec.$^{-1}$.

The CH_3– proton resonance of methylammonium ion (present largely in the form of ion pairs) was recorded under conditions of slow passage and negligible saturation. $1/T_2$, the effective line width in the absence of exchange, was taken as equal to the actual line width of the $C^{13}H_3$– proton satellite of acetic acid.

The T^1 relaxation time of N^{14} in the methylammonium salts was obtained from slow passage records of the NH– proton resonance under conditions of negligible saturation. In spite of the low signal-to-noise ratio, results were sufficiently precise even at concentrations as low as 0.2 M. Typical data are shown in Fig. 2. The interpretation of the NH– proton spectra was based on Pople's theory,[29] and allowance was made for spin–spin interaction with the CH_3– protons and for chemical exchange. Pople assumes in his theory that in the absence of T^1 relaxation, the NH– proton resonance is a $1:1:1$ triplet of extremely sharp lines. (Line width approaches zero as $1/T^1$ approaches zero.) Chemical exchange can be introduced into the theory by assuming that in the absence of T^1 relaxation the width of each triplet line equals $R/(NH)$, where R is the known rate of exchange and (NH) is the number of gram-atoms of NH– protons per liter. For any given value of T^1 and $R/(NH)$, a theoretical curve that allows for spin–spin interaction with the CH_3– protons can then be constructed by adding a $1:3:3:1$ quadruplet of curves, calculated as described above, and separated by the known CH-NH spin–spin interaction of 6.17 c.p.s. The fit of theoretical curves calculated in this way to some data for methylammonium chloride is illustrated in Fig. 2. In this case, $R/(NH) = 0$, and T^1 is defined by the data with a precision of about 10%. However, for methylammonium acetate the precision is less (about ± 25) because of the relatively large value of $R/(NH)$.

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[CONTRIBUTION FROM THE BELL TELEPHONE LABORATORIES, INC., MURRAY HILL, NEW JERSEY]

Ionization and Proton Exchange of Amines in Acetic Acid. II. Effect of Structure on Reactivity and the Reaction Mechanism

BY ERNEST GRUNWALD AND ELTON PRICE

Rates of exchange of NH– with COOH– protons have been measured over a wide temperature range for a series of amines in dilute solution in acetic acid by precise nuclear magnetic resonance techniques. The ionization of these amines is nearly complete in acetic acid, the actual reactant being $R_3NH^+OAc^-$. Reaction rates were calculated from the n.m.r. data by a rigorous method that requires the data: (i) the exchange broadening of the dominant line in the NH–COOH proton system, (ii) the chemical shift of NH– relative to COOH– protons, (iii) the N^{14}-H spin–spin coupling constant of $R_3NH^+OAc^-$, and (iv) the T^1 relaxation time of N^{14} in R_3NH^+-OAc^-. The rates of proton exchange in all cases were first order in $R_3NH^+OAc^-$, and the rate constant, k, appeared to decrease slightly with increasing concentration. The following kinetic results are reported: for $NH_4^+OAc^-$, k (25°) = 6080 sec.$^{-1}$, $\Delta H^* = 19.75$ kcal., $\Delta S^* = 25.0$ e.u.; for $(CH_3)_3NH^+OAc^-$, k (25°) = 945 sec.$^{-1}$, $\Delta H^* = 16.95$ kcal., $\Delta S^* = 11.9$ e.u.; for $(HOCH_2)_3CNH_3^+OAc^-$, k (25°) = 41,000 sec.$^{-1}$, $\Delta H^* \approx 20.7$ kcal., $\Delta S^* \approx 32$ e.u. Previously reported values for $CH_3NH_3^+OAc^-$, k (25°) = 230 sec.$^{-1}$, $\Delta H^* = 18.41$ kcal., $\Delta S^* = 14.0$ e.u., were also included in the analysis of the effect of structure on reactivity. The preceding rate constants for proton exchange of $R_3NH^+OAc^-$ in acetic acid are very nearly proportional to both rate and equilibrium constants for the acid dissociation of R_3NH^+ in water. Moreover, the variations with structure of ΔH and ΔS for the two processes follow similar patterns. These remarkably simple relationships suggest a close analogy of reaction mechanism. Thus in acetic acid the mechanism 13, with $B = R_3N$ and with $k_2 > k_H$, seems to fit these facts as well as some data concerning diffusion control of the reaction of acetic acid with strongly basic amines in water. The order of magnitude of k_H is estimated as 10^8 sec.$^{-1}$; k_2 appears to be so large that the interactions with the surrounding dielectric in the activation step of this reaction cannot be described by a reversible electrostatic model. As a result, it is probable that k_2 is greater in acetic acid than in water, in spite of the smaller static dielectric constant of acetic acid.

In water, equilibrium and rate constants for the acid dissociation of the methyl-substituted ammonium ions

are in the well-known and peculiar sequence, $NH_4^+ \gg CH_3NH_3^+ \approx (CH_3)_2NH_2^+ < (CH_3)_3NH^+$, which