Nuclear Magnetic Resonance Study of Proton Exchange Involving Ion Pairs of Ammonium Salts in *t*-Butyl Alcohol

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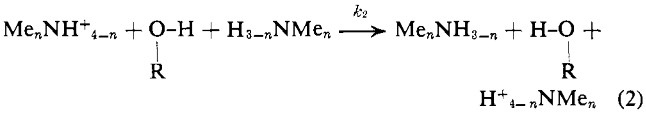
Abstract: The rates of proton transfer (*via* solvent) between methyl-substituted ammonium salts and their conjugate amines in *t*-butyl alcohol have been measured by means of the nmr technique. These rates were found to follow second-order kinetics, first order each in the ammonium salt and the amine. On the basis of the kinetics and other data, it is concluded that the proton transfer involves the ion pair of the salt rather than the dissociated ammonium ion. The rate of this reaction is much slower than the corresponding rate in water, the ratio $k(\text{in H}_2\text{O}) : k(\text{in } t\text{-BuOH})$ ranging from 23 for $MeNH_3^+Cl^-$ to 3500 for $Me_3NH^+Cl^-$. Furthermore, the rate depends on the degree of methyl substitution, the mono- and dimethylammonium salts reacting one to two orders of magnitude faster than the trimethylammonium salts. In addition, for the trimethylammonium salts, the rate depends on the anion and increases in the order $Cl^- < TFA^- < Br^- < OTs$. These results are explained by an ion-pair model in which the anion is hydrogen bonded to only one NH proton. Other evidence is presented in support of this model.

Nuclear magnetic resonance has been used to study proton-transfer reactions between acids and bases in a variety of solvents. In a hydroxylic solvent with a large dielectric constant, such as water or methanol, proton transfer between methyl-substituted ammonium salts and their conjugate bases has been found to occur by two paths: a direct proton transfer as in eq 1, and a proton transfer which involves a solvent molecule, ROH, as in eq 2.[1-5] As indicated by these equations,

$$Me_nNH^+_{4-n} + H_{3-n}NMe_n \xrightarrow{k_1} Me_nNH_{3-n} + H^+_{4-n}NMe_n \quad (1)$$



$$Me_nNH^+_{4-n} + \underset{\displaystyle R}{\underset{|}{O}}\text{-H} + H_{3-n}NMe_n \xrightarrow{k_2} Me_nNH_{3-n} + \text{H-}\underset{\displaystyle R}{\underset{|}{O}} + H^+_{4-n}NMe_n \quad (2)$$

in solvents with large dielectric constants, the proton transfer involves the dissociated ammonium ions.

In this paper, the results of an nmr study of the rate of the exchange reaction given by eq 2 using *t*-butyl alcohol as the solvent will be presented.[6] The dielectric constant of *t*-butyl alcohol is fairly low ($D = 12.47$ at 25°),[7] and ion pairing of the ammonium salt is much more pronounced than in water or methanol.

The effect of ion pairing on the exchange reaction in *t*-butyl alcohol can be seen clearly when the second-order rates in this solvent are compared with those in water or methanol. First, the rates in *t*-butyl alcohol are significantly slower than those in water or methanol. For example, the rates for the mono-, di-, and trimethylammonium chlorides are decreased by factors of 23, 38, and 3500, respectively. Second, in contrast to the rates in water, those in *t*-butyl alcohol are strongly dependent on the degree of methyl substitution; the rates for the mono- and dimethylammonium salts are as much as two orders of magnitude larger than those for the trimethylammonium salts. Third, the rates for the trimethylammonium salts depend on the nature of the anion while those for the mono- and dimethylammonium salts do not.

These results are explained by a model in which the proton exchange involves the ion pair of the ammonium salt. In this ion pair, the anion is hydrogen bonded to only one NH proton. Any other NH protons are hydrogen bonded to solvent and are more reactive.

Experimental Section

Materials. *t*-Butyl alcohol (Fisher Certified Reagent) was refluxed over calcium hydride (1 g/l.) for 24 hr and then distilled. The center fraction was redistilled after addition of a small amount of benzoic acid (about 25 mg/l.). The center fraction was employed for the kinetic measurements. From potentiometric measurements it was found to contain less than 10^{-5} *M* acid. In addition the solvent contained less than 0.002 *M* water (determined by Karl Fischer titration).

p-Toluenesulfonic acid (Eastman Kodak White Label) was recrystallized twice from ethylene dichloride and dried *in vacuo* over $Mg(ClO_4)_2$.

Trifluoroacetic acid (Eastman Kodak White Label) was refluxed over powdered molecular sieves for 3 hr and then distilled. The fraction boiling at 72° was used in the kinetic measurements.

The methyl-substituted ammonium chlorides were Eastman Kodak White Label compounds. Each was recrystallized twice, and dried *in vacuo* over $Mg(ClO_4)_2$. The equivalent weight as determined by the Volhard method always was within 0.5% of the calculated value. In addition, these salts were found to contain no measurable excess acid or base.

The methyl-substituted ammonium *p*-toluenesulfonates were prepared by passing the appropriate gaseous amine (the Matheson Company) through a methanol solution of *p*-toluenesulfonic acid. The methanol was stripped off and the salt was recrystallized twice from *t*-butyl alcohol or *t*-butyl alcohol–benzene. The equivalent weight was determined by neutralizing the salt with excess sodium methoxide, distilling away the amine, and then titrating the remaining base with hydrochloric acid. The titrated equivalent weight including *p*-toluenesulfonic acid was within 1% of the calculated value except for the monomethylammonium salt, for which it differed by 1.3%. The salts were tested for excess acid by potentiometric titration with dilute sodium methoxide. The trimethylammonium salt contained 1.89% *p*-toluenesulfonic acid. The dimethyl salt contained less than 0.08%, and the monomethyl salt contained 0.31%.

Trimethylammonium bromide was prepared by passing gaseous trimethylamine through 48% aqueous hydrobromic acid (Matheson Coleman and Bell). The water was stripped off, and the salt was recrystallized twice from methanol. The equivalent weight was determined by a potentiometric titration of the salt with sodium methoxide and was within 0.2% of the calculated value.

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(1) E. Grunwald, P. Karabatsos, R. Kromhout, and E. Purlee, *J. Chem. Phys.*, 33, 556 (1960).

(2) A. Loewenstein and S. Meiboom, *ibid.*, 27, 1067 (1957).

(3) Z. Luz and S. Meiboom, *ibid.*, 39, 366 (1963).

(4) E. Grunwald, unpublished work.

(5) E. Grunwald, *J. Phys. Chem.*, 67, 2208 (1963).

(6) Preliminary results of this study were presented earlier: M. Cocivera and E. Grunwald, *J. Am. Chem. Soc.*, 87, 2070 (1965).

(7) W. Dannhauser and L. Bahe, *J. Chem. Phys.*, 40, 3058 (1964).

The methyl-substituted ammonium trifluoroacetates were prepared *in situ* from standardized solutions of the amine and trifluoroacetic acid. Preparation of the standard amine solutions will be discussed below.

Solutions. The amine solutions were prepared in the following manner. The recrystallized ammonium salt, potassium *t*-butoxide, and *t*-butyl alcohol were mixed. The *t*-butyl alcohol–amine solution was distilled from the mixture, and the concentration of amine was determined by potentiometric titration using hydrochloric acid. Using this solution as a stock solution, the other solutions were prepared gravimetrically. For a given buffer ratio, a number of solutions were prepared by dilution of an initial solution. As a double check on the concentrations, the concentration of the amine in the initial solution was determined by a potentiometric titration with HCl.

Chemical Shifts. For most of the salts, the chemical shift, δ, between the OH and NH proton resonances was obtained from the slow passage spectrum of a *t*-butyl alcohol solution containing the ammonium salt and approximately 0.05 M acid. However, the chemical shifts for the mono- and trimethylammonium chlorides could not be measured directly because these salts were not sufficiently soluble in *t*-butyl alcohol to permit detection of the NH proton resonance.

For these salts, the chemical shift, δ, between the NH and OH lines was obtained indirectly in the following manner. The slow passage spectrum was obtained for a solution of the salt and HCl in *t*-butyl alcohol containing approximately 5 wt % of water to increase the solubility of the salt. From this spectrum, the chemical shift between the NH resonance and the CH_3 resonance of *t*-butyl alcohol was obtained. The chemical shift between the OH and CH_3 resonances of *t*-butyl alcohol was obtained from a saturated solution of the ammonium salt in dry *t*-butyl alcohol. From these two chemical shifts, the chemical shift, δ, between the NH and OH lines was obtained by assuming that the NH *vs.* CH_3 chemical shift in dry *t*-butyl alcohol is the same as that in 95% *t*-butyl alcohol. As an added check for trimethylammonium chloride, δ was determined from rates which were in the region of maximum broadening by means of the expression, $\Delta_{max}/p\delta = 0.5$, where Δ_{max} is the maximum broadening of the OH resonance caused by exchange and p is the NH proton fraction. This expression for the maximum value of Δ was derived from eq 3. The values of δ at 25 and 35° obtained in this manner are 324 and 343 (cps at 60 Mc/sec), respectively. These values agree fairly well with those determined by the previous method: 315 and 327 cps at 25 and 35°, respectively. For the rate measurements, the average value from the two methods was used.

The chemical shifts relative to the CH_3 protons of *t*-butyl alcohol are given in Table I for all of the salts at 35°. All of the resonances listed are downfield from the *t*-butyl alcohol CH_3 resonance. From Table I, it is clear that the NH chemical shift depends on the degree of methyl substitution on the nitrogen. Such an effect is not unexpected since it has been found to occur in water for the same ammonium ions.[1,5] However, the NH chemical shift also is slightly dependent on the nature of the anion of the salt. For both the di- and trimethylammonium salts, it decreases in the order $TFA^- > Cl^- > OTs^-$.

Table I. Chemical Shifts (Relative to CH_3 of *t*-Butyl Alcohol)[a] for Various Ammonium Salts at 35°

Salt	Concn, M	CH_3N	OH	NH
$MeNH_3Cl$	0.156	80	207	406
$MeNH_3OTs$[b]	0.198	78	201	403
Me_2NH_2TFA[c]	0.821	82	218	475
Me_2NH_2Cl	0.408	87	216	469
Me_2NH_2OTs	0.412	86	204	434
Me_3NHTFA	0.801	99	211	541
Me_3NHCl	0.289	101	204	531
Me_3NHOTs	0.439	96	196	494

[a] In cps at 60 Mc. [b] *p*-Toluenesulfonate. [c] Trifluoroacetate.

At 50, 65, and 80°, the chemical shift, δ, between the NH and OH proton resonances was obtained indirectly by measuring the OH *vs.* CH_3–C chemical shift and assuming that the chemical shift between the NH and CH_3–C resonances is temperature independent over the temperature range employed. This assumption was found to be valid at 25 and 35° and probably is acceptable at the higher temperatures.

Except for the trifluoroacetates, the CH_3N proton resonances for the ammonium salts at 35° are multiplets under the conditions used for the slow passage experiments, *i.e.*, a quadruplet for the monomethyl-, a triplet for the dimethyl-, and a doublet for the trimethylammonium salts. For both the di- and trimethylammonium trifluoroacetates, the CH_3–N resonances are broadened singlets, indicating that exchange is occurring under these conditions.

Of more interest is the fine structure observed for the NH resonances of mono- and trimethylammonium chlorides and *p*-toluenesulfonates. For both the trimethylammonium chloride and *p*-toluenesulfonate, the NH has definite fine structure which indicates that it is a decuplet. The NH resonance for both the monomethylammonium chloride and *p*-toluenesulfonate is a very well-resolved quadruplet whose lines are somewhat broader than those in the CH_3N quadruplet. Only a broad singlet was observed for the NH resonance of the dimethylammonium salts. The significance of these results will be discussed later in this paper.

Rate Measurements. The rates of proton exchange were measured using the nmr method.[8] For the solutions studied, the ammonium salt concentration was never larger than $\sim 10^{-1}$ M. Consequently, the spectra consisted of strong singlet OH and CH_3 proton resonances of *t*-butyl alcohol, a weak singlet CH_3 line of the ammonium ion, and other weak lines due to the anion of the salt. Because of proton exchange between the NH and OH protons, the weak NH line could not be detected, and the OH line was broadened. In this case of exchange between a weak line and a strong one, Meiboom[9] has shown that the excess width, Δ, of the strong line is given by eq 3, which is valid over the entire range of exchange rates. In

$$\Delta = \frac{1}{T'_2} - \frac{1}{T_2} = \frac{p\delta^2\tau}{1 + \delta^2\tau^2} \tag{3}$$

this equation, τ is the lifetime for proton exchange, p is the NH proton fraction of the exchanging protons (*i.e.*, NH and OH protons), δ is the chemical shift between the NH and OH lines, T'_2 is the measured transverse relaxation time of the exchange broadened line, and T_2 is the relaxation time due to all processes except exchange. If the proton exchange is too slow to affect the line width, the transverse relaxation time, T_2, in eq 3 can be set equal to the longitudinal relaxation time, T_1. To check the validity of this relation, T_2 and T_1 of the OH line were measured for *t*-butyl alcohol containing no ammonium salt. T_2 was found to be slightly shorter than T_1, and the relation between them is

$$\frac{1}{T_2} = \frac{1}{T_1} + 0.02$$

This relation holds when the *t*-butyl alcohol is neutral and exchange is slow, and when the alcohol is acidified and the exchange is fast. Although the cause for this difference between T_1 and T_2 is unclear,[10] application of this relation to obtain T_2 from the measured T_1 seemed to be the best approach. The transverse relaxation times were measured by a modified Carr–Purcell spin–echo scheme.[13,14] The longitudinal relaxation time, T_1, was measured using the Carr–Purcell 180°–90° pulsing scheme.[15] A Varian DP60 spectrometer operating at 60 Mc was used for these measurements.

Once τ was determined, the rate, R, for proton exchange was obtained from eq 4. In this equation [BHX] is the molar concentration of the ammonium salt, and τ_{NH} is the average lifetime of the NH proton and is calculated from the following equations.

$$-\frac{1}{[BHX]}\frac{d[BHX]}{dt} = \frac{R}{[BHX]} = \frac{1}{\tau_{NH}} \tag{4}$$

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(8) J. A. Pople, W. G. Schneider, and H. J. Bernstein, "High Resolution Nuclear Magnetic Resonance," McGraw-Hill Book Co., Inc., New York, N. Y., p 222.

(9) S. Meiboom, *J. Chem. Phys.*, **34**, 375 (1961).

(10) Similar differences have been found in other compounds.[3,11,12] At the present time, it is unclear whether this effect is real or merely due to an instrumental defect in the T_1 and T_2 measurements.

(11) Z. Luz and S. Meiboom, *J. Chem. Phys.*, **40**, 2686 (1964).

(12) M. Cocivera and E. Grunwald, *J. Am. Chem. Soc.*, **87**, 2551, (1965).

(13) S. Meiboom and D. Gill, *Rev. Sci. Instr.*, **29**, 688 (1958).

(14) S. Alexander, *ibid.*, **32**, 1066 (1961).

(15) H. Y. Carr and E. M. Purcell, *Phys. Rev.*, **94**, 630 (1954).

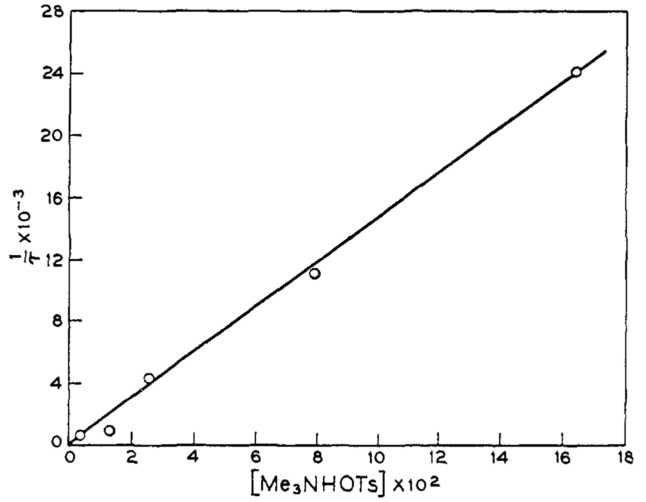


Figure 1. A plot of $1/\tau_{NH}$ *vs.* [BHX] for trimethylammonium *p*-toluenesulfonate at 35° and a buffer ratio of 5.06.

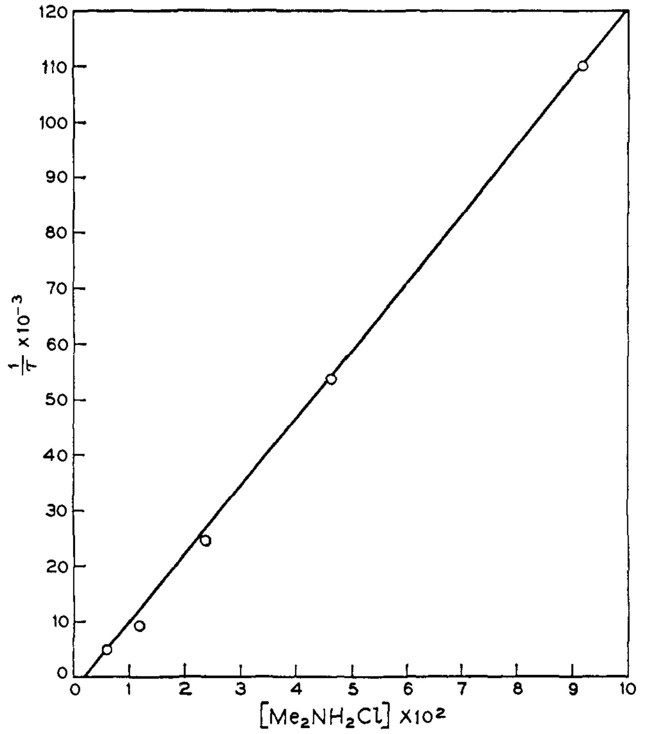


Figure 2. A plot of $1/\tau_{NH}$ *vs.* [BHX] for dimethylammonium *p*-toluenesulfonate at 35° and a buffer ratio of 24.75.

$$\frac{1}{\tau} = \frac{1}{\tau_{NH}} + \frac{1}{\tau_{OH}}$$

$$\frac{p_{OH}}{\tau_{OH}} = \frac{p_{NH}}{\tau_{NH}}$$

In the latter equation, p is the proton fraction, and $p_{OH} + p_{NH} = 1$.

Results

The rate, R, was found to follow second-order kinetics, first-order each in the ammonium salt and the amine. This kinetic order was deduced from the fact that plots of $1/\tau_{NH}$ *vs.* [BHX] were linear. If the kinetics is second order, from eq 4 one obtains eq 5.

$$\frac{1}{\tau_{NH}} = \frac{k'_2}{r}[BHX] \tag{5}$$

In this equation, r is the buffer ratio, [BHX]/[B], and k'_2 is the second-order rate constant. Thus when the buffer ratio is kept constant, a plot of $1/\tau_{NH}$ *vs.* [BHX] should be linear if the rate follows second-order kinetics. Some plots of this type are illustrated in Figures 1, 2, and 3 for various salts. As can be seen from these figures, the plots are indeed linear. The slopes of such plots give k'_2/r, and since r is known, k'_2 can be obtained.

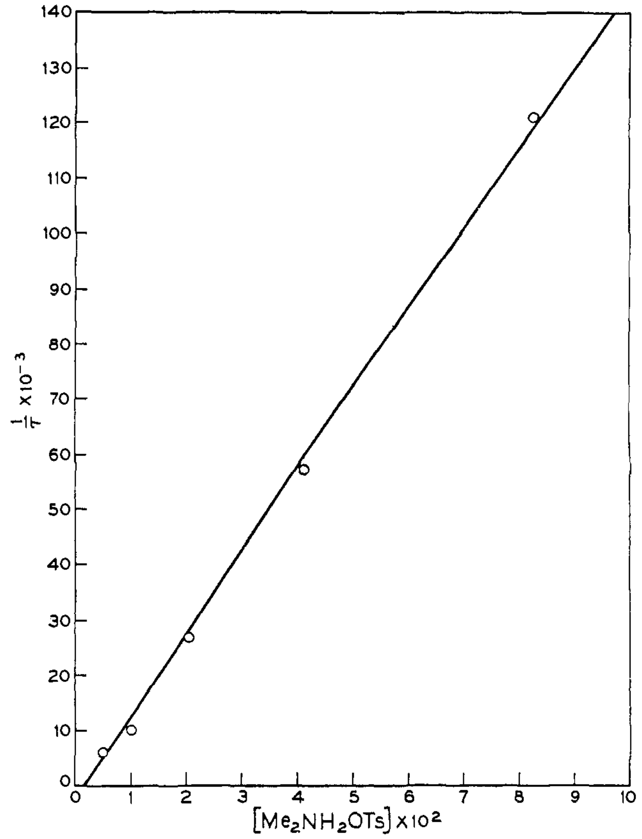


Figure 3. A plot of $1/\tau_{NH}$ *vs.* [BHX] for dimethylammonium chloride at 35° and a buffer ratio of 25.52.

For the dimethylammonium *p*-toluenesulfonate and trifluoroacetate salts, k'_2 was determined for only one buffer ratio. For all other salts, k'_2 was determined at three or more buffer ratios, and for each salt k'_2 is independent of the buffer ratio. To illustrate this independence, k'_2 values at 35° are listed in Table II at various buffer ratios for mono- and trimethylammonium *p*-toluenesulfonate salts. Because k'_2 is independent of the buffer ratio and plots of $1/\tau_{NH}$ *vs.* [BHX] are linear, one may conclude that the rate follows second-order kinetics.

The average second-order rate constants for the various salts at 35° are listed in Table III. As can be seen from these results, the rate of proton exchange is strongly dependent upon the nature of the reactant. For the trimethylammonium salts, the value of the second-order rate constant, k'_2, depends on the nature of the anion, increasing in the order $Cl^- < TFA^- < Br^- < OTs^-$. For the dimethylammonium salts, k'_2 does not depend on the nature of the anion, and its values are up to two orders of magnitude larger than those for the trimethylammonium salts. The values of k'_2 for the monomethylammonium salts are approximately the same as those for the dimethyl salts and

Table II. Values of k'_2 at 35° for Mono- and Trimethylammonium *p*-Toluenesulfonates at Various Buffer Ratios

Salt	$\frac{[BH^+X^-]}{[B]}$	$k'_2 \times 10^{-5}$, M^{-1} sec^{-1}
$MeNH_3OTs$	7.25	200
	8.23	260
	17.6	290
	22.8	240
	36.3	220
Me_3NHOTs	0.462	5.5
	1.58	7.0
	1.73	7.0
	2.01	6.2
	5.06	7.4
	10.37	8.2
	11.02	7.5
	69.2	7.3

Table III. Values of k'_2 for Methyl-Substituted Ammonium Salts in *t*-BuOH at 35°

Cation	Anion	$k'_2 \times 10^{-5}$, M^{-1} sec^{-1}	$\frac{k_2(H_2O)}{k'_2(t\text{-BuOH})}$
$(CH_3)_3NH^+$	Cl^-	1.1 ± 0.12	3500
$(CH_3)_3NH^+$	TFA^-	1.8 ± 0.4	2200
$(CH_3)_3NH^+$	Br^-	5.3 ± 0.7	740
$(CH_3)_3NH^+$	OTs^-	7.0 ± 0.7	560
$(CH_3)_2NH_2^+$	Cl^-	270 ± 40	38
$(CH_3)_2NH_2^+$	TFA^-	220 ± 20	46
$(CH_3)_2NH_2^+$	OTs^-	310 ± 30	33
$CH_3NH_3^+$	Cl^-	260 ± 30	23
$CH_3NH_3^+$	OTs^-	240 ± 40	25

again do not change significantly when the anion is changed.

In addition, the rates were measured at 25, 50, 65, and 80° for some of the salts. From the temperature dependence of the second-order rate constants, the activation parameters, $\Delta H^{\ddagger}$ and $\Delta S^{\ddagger}$, were calculated, and they are presented in Table IV.

Table IV. Enthalpy[a] and Entropy of Activation at 35° for Various Methyl-Substituted Ammonium Salts

Salt	$k'_2 \times 10^{-5}$, M^{-1} sec^{-1}	$\Delta H^{\ddagger}$, kcal	$\Delta S^{\ddagger}$, eu
Me_3NHOTs	7.0	1.6 ± 1.0	−26
Me_3NHCl	1.1	1.7 ± 1.0	−29
Me_2NH_2OTs	310	2.8 ± 0.7	−15
Me_2NH_2Cl	270	2.7 ± 0.8	−15
$MeNH_3OTs$	240	2.4 ± 0.8	−16

[a] The enthalpy of activation was calculated using rate constants measured at 35, 50, 65, and 80°.

For trimethylammonium chloride in water, the value of $\Delta H^{\ddagger}$ reported by Grunwald[5] is 3.3 kcal. Loewenstein[16] has found a value of 2.0 ± 1.2 kcal for dimethylammonium chloride. In view of the experimental error and the indirect manner in which the chemical shift, δ, is obtained at the higher temperatures, the differences and similarities in the values of the activation parameters should not be considered very meaningful.

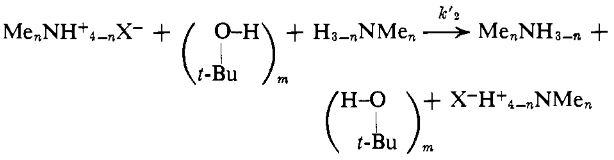
Discussion

In contrast with the proton-exchange reaction in water and methanol, the results for exchange involving methyl-substituted ammonium salts in *t*-butyl alcohol are not consistent with a mechanism involving the dissociated ammonium ion as given in eq 2. The dielectric constant of *t*-butyl alcohol is small, and one would expect the dissolved salts to exist largely in the form of ion pairs. In fact, Marple and Fritz[17] have determined ion-pair dissociation constants for a number of tetrabutylammonium salts in *t*-butyl alcohol at 26° and find values between 10^{-5} and 10^{-6}. For the methyl-substituted ammonium salts, one would expect the ion-pair dissociation constants to be in the same range, if not smaller,[18] The rates of proton exchange reported here were measured using solutions with salt concentrations between 10^{-3} and 10^{-1} M, and under these conditions the salts should exist largely in the form of ion pairs. Therefore, if the proton exchange involves the dissociated ammonium ion, a prior dissociation step (eq 6) must be added to the

$$Me_nNH^+_{4-n}X^- \overset{K}{\rightleftharpoons} Me_nNH^+_{4-n} + X^- \qquad (6)$$

mechanism given by eq 2. For this mechanism, the rate of proton exchange depends on the square root of the salt concentration. Thus, this mechanism is not consistent with the experimental results, which indicate that the rate depends on the first power of the salt concentration, rather than the square root.

A mechanism which is consistent with the data is given in eq 7. In this mechanism, the ion pair (BH^+X^-),



$$Me_nNH^+_{4-n}X^- + \left(\begin{matrix}O\text{–}H\\ |\\ t\text{-Bu}\end{matrix}\right)_m + H_{3-n}NMe_n \xrightarrow{k'_2} Me_nNH_{3-n} + \left(\begin{matrix}H\text{–}O\\ |\\ t\text{-Bu}\end{matrix}\right)_m + X^-H^+_{4-n}NMe_n \qquad (7)$$

rather than the dissociated ion, is involved in the proton exchange. The number of solvent molecules involved in this reaction has not been determined. Consequently, in eq 7, the number of solvent molecules has been denoted by the subscript $m \geq 1$. However, it should be noted that for the reaction given by eq 2, only one solvent molecule is involved when the solvent is water[3] or methanol.[19]

Comparing the values of k'_2 obtained for each salt will furnish more information about the details of the mechanism; however, before discussing them it is instructive to discuss, for the purpose of comparison, the data for the reaction (eq 2) involving the dissociated ammonium ions. This reaction occurs in water[1,2,6] and in methanol.[4] The data are listed in Table V, and the k_2 is for reaction of the whole cation, not for one proton. After making the statistical correction to obtain the value of k_2 for one proton, one finds that the value still depends on the degree of methyl substitution; however, the differences are not large (the smallest and the largest values differ by a factor of only 2.5). Furthermore, for the trimethylammonium

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[16] A. Loewenstein, *J. Phys. Chem.*, **67**, 1728 (1963).

[17] L. Marple and J. Fritz, *Anal. Chem.*, **35**, 1223 (1963).

[18] The smaller size of the cation and the availability of NH protons for hydrogen bonding to the anion would tend to decrease the value of the dissociation constant.

[19] M. Cocivera, E. Grunwald, and C. F. Jumper, *J. Phys. Chem.*, **68**, 3234 (1964).

Table V. Values of k_2 for Methyl-Substituted Ammonium Ions in Water and Methanol at 35°[a]

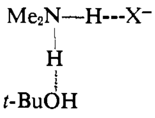
Cation	$k_2 \times 10^{-8}$, M^{-1} sec^{-1}	Solvent	Ref
$MeNH_3^+$	6.1	H_2O	1
$Me_2NH_2^+$	10.3	H_2O	2
Me_3NH^+	3.9	H_2O	5
Me_3NH^+	3.7	MeOH	4

[a] Extrapolated from 25° using the activation energies given in ref 5 and 16.

ion, the value of k_2 does not change significantly when the solvent is changed from water to methanol.

On the other hand, when the solvent is changed from water to *t*-butyl alcohol, the behavior is quite different. The value of the second-order rate constant for each trimethylammonium salt in *t*-butyl alcohol is very much smaller than the value in water. The extent of the decrease is shown in the fourth column of Table III; the ratio of $k(H_2O)$ to $k(t\text{-BuOH})$ ranges from 560 to 3500. Furthermore, the value of k'_2 is strongly dependent on the degree of methyl substitution, the mono- or dimethylammonium salts reacting 40 to 300 times faster than the trimethylammonium salts. Finally, for the trimethylammonium salts, the value of k'_2 depends on the anion.

Not only are these results consistent with the ion-pair mechanism, but they also lead to further deductions about the details of the mechanism. The dependence of the rate on the nature of the anion exhibited by the trimethyl salts indicates that the anion is interacting strongly with the NH proton of the ammonium ion, probably through a hydrogen bond. The fact that the dimethylammonium salts react much faster than the trimethyl salts and exhibit no anion dependence indicates that only one NH proton is hydrogen bonded to the anion and that the other is hydrogen bonded to solvent and is much more reactive. A possible structure is illustrated as



$$\begin{array}{l} Me_2N\text{—}H\text{---}X^- \\ \quad\;\;| \\ \quad\;\;H \\ \quad\;\;\vdots \\ t\text{-BuOH} \end{array}$$

The results for the monomethyl salts lead to the same conclusion. In this case, two NH protons are hydrogen bonded to solvent.

In addition to the kinetic data, certain features of the nmr spectra indicate that the salts in *t*-butyl alcohol exist in the form of ion pairs in which the anion is in intimate contact with the ammonium ion. First, for a given ammonium ion, the chemical shift of the NH proton resonance listed in Table I depends on the nature of the anion.[20] Such a dependence suggests that the anion is interacting with the NH protons. Second, for the mono- and trimethylammonium salts in *t*-butyl alcohol, the NH proton resonance exhibits a fine structure not found in water solutions. In water, where the degree of ion pairing is small at the concentrations used in this study, the NH proton resonance of the monomethylammonium ion is a triplet due to spin–spin coupling with N^{14} (spin of 1).[21] Theoretically, each line of the triplet is split into a quadruplet because of spin–spin interaction with the CH_3N protons; however, the spin-lattice relaxation of the N^{14} (due to quadrupole relaxation caused by fluctuating electric-field gradients)[22] broadens these lines so that the quadruplet structure can barely be detected.[21] For all of the ammonium salts studied in *t*-butyl alcohol, the NH proton resonance is not split by N^{14}. In fact, fine structure due to spin–spin interaction with the CH_3N protons is observed for the mono- and trimethylammonium salts. These results indicate that the N^{14} is undergoing very rapid quadrupole relaxation and, therefore, that the electric-field gradient is much larger than when water is the solvent. This large increase in the electric-field gradient is consistent with the theory that the anion is hydrogen bonded to an NH proton.

Let us return to the kinetic data for the mono- and dimethylammonium salts. As shown in Table III, the rate of proton exchange for these salts is reduced significantly when the solvent is changed from water to *t*-butyl alcohol. However, before comparing the rates in the two solvents, one should correct statistically for the number of ammonium protons which can react in each solvent. If one assumes that one NH proton which is reactive in water becomes unreactive in *t*-butyl alcohol because of ion pair formation, the ratio $k(\text{in } H_2O):k(\text{in } t\text{-BuOH})$ is 2:1 for the dimethyl salts and 3:2 for the monomethyl salts. On making these statistical corrections to the ratios given in Table III, the value for each of these salts becomes ~17. Consequently, the tying up of one NH proton cannot be the sole cause of the decrease in rate when the solvent is changed from water to *t*-butyl alcohol. Other factors which also could be effective in reducing the rate are the low dielectric constant of *t*-butyl alcohol, the bulkiness of the *t*-butyl group, and the interaction of the anion with the ammonium ion to cause a decrease in the acidity of all of the NH protons.

Conclusions

In *t*-butyl alcohol, the proton transfer (*via* solvent) between a methyl-substituted ammonium salt and its conjugate amine involves the ion pair of the salt rather than the dissociated ammonium ion. In this ion pair, only one NH proton is interacting strongly with the anion (probably through a hydrogen bond); any others are hydrogen bonded to solvent and are much more reactive.

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(20) A similar effect was found for trimethylammonium chloride and acetate in acetic acid: E. Grunwald and E. Price, *J. Am. Chem. Soc.*, **86**, 2970 (1964).

(21) E. Grunwald, A. Loewenstein, and S. Meiboom, *J. Chem. Phys.*, **25**, 382 (1956).

(22) Reference 8, p 227.