Proton Exchange and Ion-Pair Exchange of N,N-Dialkylanilinium Salts and Basicity of N,N-Dialkylanilines in Acetic Acid[1]

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Abstract: Analysis of kinetic measurements of proton exchange and equilibrium measurements of basicity in glacial acetic acid has led to rate constants for the following fast reactions (B = N,N-dimethyl, -diethyl-, or -di-*n*-propylaniline; HTs = *p*-toluenesulfonic acid; HAc = acetic acid). (1) $BH^+Ts^- + HAc \rightleftharpoons BH^+Ac^- + HTs$ (k_+, k_-). In this reversible acid–base reaction, the basic anions exist in combination with BH^+ in ion pairs. The N–H proton of BH^+ does not exchange in this process. Typical results, for N,N-diethylaniline at 26°, are $k_+ = 1040\ \text{sec}^{-1}$ and $k_- = 11 \times 10^9\ \text{sec}^{-1}\ M^{-1}$; k_- is of the appropriate magnitude for a diffusion-controlled reaction. (2) $BH^+Ts^- + BH^+Ac \longrightarrow BH^+Ac^- + BH^+Ts^-$ (k). This symmetrical metathesis, or ion-pair exchange, is not a diffusion-controlled reaction. A typical result for N,N-diethylaniline at 26° is $k = 6.5 \times 10^7\ \text{sec}^{-1}\ M^{-1}$. (3) $BH^+Ac^- \rightleftharpoons B \cdot HAc$ (k_{-i}, k_i). (4) HAc (solvent) + $B \cdot HAc \longrightarrow$ HAc(solvent) + $B \cdot HAc$ (k_H). Reactions 3 and 4 accomplish the transfer of an N–H proton to a carboxyl group of acetic acid in the bulk solvent. In (3) the proton is transferred without dissociation of the hydrogen-bonded complex; in (4) the $B \cdot HAc$ hydrogen bond breaks. The new data confirm previous evidence[3] that reaction 4 is rate determining, that is, $k_i > k_H$. Typical results for N,N-diethylaniline at 26° are $k_H = 6.0 \times 10^8\ \text{sec}^{-1}$; $k_i > 3 \times 10^9\ \text{sec}^{-1}$; $k_i/k_{-i} = 109$. The basicity of N,N-dialkylanilines and other bases was measured in acetic acid at 26° by means of $K_i = [BH^+Ac^-]/[B]$ and of $K_i^{B \cdot HTs} = [BH^+Ts^-]/[B][HTs]$. For N,N-dialkylaniline, K_i and $K_i^{B \cdot HTs}$ both show the peculiar sequence $Me_2 \ll Et_2 > n\text{-}Pr_2$ that had previously been noted for K_B in water.[4] The rate constants, k_+ and k_-, also show a maximum at Et_2, but k_H decreases monotonically with increasing alkyl size.

We have measured the rates of $NH\text{–}CO_2H$ proton exchange between N,N-dialkylanilinium *p*-toluenesulfonate salts (BH^+Ts^-, 0.05–0.12 *M*) and glacial acetic acid (HAc) in the presence of *p*-toluenesulfonic acid (HTs, 10^{-4}–0.5 *M*). Under these experimental conditions the salts exist largely in the form of undissociated ion pairs, and *p*-toluenesulfonic acid is known to be largely un-ionized in acetic acid.[2] The N,N-dialkyl substituents are dimethyl, diethyl, and di-*n*-propyl. Rates were measured by nuclear magnetic resonance (nmr) techniques.[3]

We find that proton exchange proceeds by a fairly complicated mechanism, involving three consecutive steps: (1 or 2) reversible formation of the N,N-dialkylanilinium acetate ion pair by acid–base reaction or ion-pair exchange

$$BH^+Ts^- + HAc(\text{solvent}) \xrightleftharpoons[k_-]{k_+} BH^+Ac^- + HTs \qquad (1)$$

$$BH^+Ts^- + BH^+Ac^- \xrightleftharpoons{k} BH^+Ac^- + BH^+Ts^- \qquad (2)$$

(3) reversible formation of solvated N,N-dialkylaniline by proton transfer

$$BH^+Ac^- \xrightleftharpoons[k_i]{k_{-i}} B \cdot HAc \qquad (3)$$

(4) breaking of the $B \cdot HAc$ hydrogen bond and exchange of HAc with a molecule in the bulk solvent

$$HAc(\text{solvent}) + B \cdot HAc \xrightarrow{k_H} HAc(\text{solvent}) + B \cdot HAc \qquad (4)$$

In order to complete the kinetic analysis, we have also measured the equilibrium constants for reactions 1 and 3. As a result, we are able to report absolute values for the rate constants k_+, k_-, k, and k_H. Results obtained at 26° are listed in Table I.

Table I. Rate and Equilibrium Constants for Reactions 1–4 of N,N-Dialkylanilines in Acetic Acid at 26°

Constant (eq 1–4)	N,N-Dialkyl groups Me_2	Et_2	$n\text{-}Pr_2$
(1) $BH^+Ts^- + HAc(\text{solvent}) \rightleftharpoons BH^+Ac^- + HTs$			
k_+, sec^{-1}	810	1040	780
k_-, sec^{-1} M^{-1}	5.0×10^9	11×10^9	8.5×10^9
$K = k_+/k_-$ [a]	1.61×10^{-7}	0.94×10^{-7}	0.92×10^{-7}
(2) $BH^+Ts^- + BH^+Ac^- \rightleftharpoons BH^+Ac^- + BH^+Ts^-$			
k, sec^{-1} M^{-1}	...	6.5×10^7	5.0×10^7
K_{equil}	Unity	Unity	Unity
(3) $BH^+Ac^- \rightleftharpoons B \cdot HAc$			
k_{-i}/k_i [b]	0.079	0.0092	0.033
(4) $HAc(\text{solvent}) + B \cdot HAc \rightarrow HAc(\text{solvent}) + B \cdot HAc$			
k_H, sec^{-1} [c]	1.3×10^9	6.0×10^8	2.7×10^8

[a] $K = K_i/K_i^{B \cdot HTs}$; data in Table III. [b] $k_{-i}/k_i = (K_i)^{-1}$; data in Table III. [c] $k_H = k_e K_i$; k_e defined just after eq 16.

Before discussing these data, we wish to examine the basicity of N,N-dialkylanilines, which has been a puzzle for a long time.[4] Three measures of basicity are listed in Table II: the base dissociation constant K_B in water,[4-6] the base ionization constant K_i in acetic acid,[7] and the equilibrium constant K_i^{BHTs} for acid–base reaction with *p*-toluenesulfonic acid in acetic acid.

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(1) This work was supported by the National Science Foundation under Grant GP-3921 and by the Petroleum Research Fund of the American Chemical Society. Grateful acknowledgment is made to the donors of that fund.

(2) I. M. Kolthoff and S. Bruckenstein, *J. Am. Chem. Soc.*, **78**, 1 (1956). Our notation for equilibrium constants follows that of these authors.

(3) E. Grunwald and E. Price, *ibid.*, **86**, 2965, 2970 (1964).

(4) N. F. Hall and M. R. Sprinkle, *ibid.*, **54**, 3469 (1932).

(5) A. L. Bacarella, E. Grunwald, H. P. Marshall, and E. L. Purlee, *J. Org. Chem.*, **20**, 747 (1955).

(6) (a) G. Girault-Vexlearschi, *Bull. Soc. Chim. France*, 589 (1956); (b) *ibid.*, 606 (1956).

(7) G. W. Ceska and E. Grunwald, *J. Am. Chem. Soc.*, **89**, 1371 (1967).

$$K_i = [BH^+Ac^-]/[B] \qquad (5)$$

$$K_i^{B \cdot HTs} = [BH^+Ts^-]/[B][HTs] \qquad (6)$$

In (5) and (6), [B] denotes the formal concentration of solvated, unprotonated N,N-dialkylaniline. We shall assume that [B] consists largely of molecules of the hydrogen-bonded complex $B \cdot HAc$, so that K_i in eq 5 may be equated to k_i/k_{-i} in eq 3 of the kinetic scheme.

Table II shows that as the size of the *n*-alkyl groups increases basicity increases to a pronounced maximum at N,N-diethyl, and then appears to oscillate with diminishing amplitude. The maximum at N,N-diethyl is evident in all three measures of basicity and therefore cannot be ascribed to any particular solvent or molecular environment: it is an intrinsic effect. The sequence $H_2 < Me_2 \ll Et_2$ as been explained as resulting from a combination of polar effects and steric inhibition of resonance with the benzene ring,[6,8] but the variation beyond Et_2 cannot be explained in this way.

Table II. Measures of Basicity and Optical Data for N,N-Dialkylanilines

			Log	"D band"	
Aniline	Log K_B (H_2O, 25°)	Log K_i (HAc, 26°)	$K_i^{B \cdot HTs}$ (HAc, 26°)	λ_{max} (HAc), mμ	ϵ_{max}
$C_6H_5NH_2$	−9.38[a]	0.90[d]	7.58[e]	280[d]	1412[f]
$C_6H_5NMe_2$	−8.85[a]	1.10[d]	7.90[h]	299[d]	1950[g]
$C_6H_5NEt_2$	−7.44[b]	2.04[h]	9.06[h]	302[h]	2080[g]
C_6H_5N-*n*-Pr_2	−8.41[b]	1.48[d]	8.52[h]	305[d]	2310[g]
C_6H_5N-*n*-Bu_2	−7.79[c]	...	...	...	...

[a] Reference 5. [b] Reference 4. [c] Reference 6. [d] Reference 7. [e] Unpublished observation by Dr. C. S. Leung in this laboratory. [f] In 95% H_2O–5% MeOH; λ_{max} in this solvent is equal to λ_{max} in HAc. [g] In methanol; λ_{max} in this solvent is equal to λ_{max} in HAc. [h] Present work.

We can rationalize our kinetic results best if we assume that the variation in the measures of basicity results from some peculiarity of structure of the N,N-dialkylanilinium ions, and that the properties of the anilines vary monotonically. Certain optical data strongly support this assumption. Table II lists wavelengths and extinction coefficients at the maximum of the D band (in the terminology of Klevens and Platt)[9] in the absorption spectrum of unprotonated anilines. This band is characteristic of anilines[9,10] and is quite sensitive to two kinds of substituent effects: (a) N-alkyl substitution raises λ_{max} and ϵ_{max},[6b] probably by a polar effect. (b) Steric inhibition of resonance with the benzene ring lowers λ_{max} and reduces ϵ_{max} dramatically.[6b,9,11,12] It is clear from Table II that λ_{max} and ϵ_{max} both increase monotonically with increasing size of the *n*-alkyl groups. On the other hand, the NH-proton chemical shift of BH^+Ts^-, which provides an index of the structure of BH^+, goes through a maximum at N,N-diethyl and varies in the same sequence as does pK_A. Values of δ (defined in the Experimental Part) are (in parts per million): 0.79 for Me_2, 1.33 for Et_2, and 1.23 for *n*-Pr_2.

We now wish to return to the rate constants in Table I. At this point we shall consider only the actual numbers, leaving details of measurement and kinetic analysis for examination in later sections.

Reaction 1 is an acid–base reaction, $Ts^- + HAc \rightleftharpoons HTs + Ac^-$, in which the anions are parts of ion pairs. The precision (standard indeterminate error) of k_+ and k_- is estimated at ±30%. k_- is of the order of magnitude expected for diffusion-controlled reaction.[13] The rate constants seem to vary inversely with the acid strength of BH^+, going through a maximum at $(C_2H_5)_2$. Although this relationship is not securely outside of experimental error, it is plausible, especially if Ts^- or Ac^- in the ion pairs are connected to BH^+ by a fairly strong hydrogen bond.

Our data leave little doubt that the N–H bond of BH^+ remains intact in reaction 1. If we make the plausible guess that the reversal of (1) involves direct proton transfer from HTs to the carboxylate group of BH^+Ac^-, then microscopic reversibility requires that the forward reaction takes place in two steps, as follows.

$$BH^+Ts^- + HAc \rightleftharpoons BH^+ \cdot AcH \cdot Ts^- \rightleftharpoons BH^+Ac^- + HTs$$

In the second step of this mechanism, the considerable work that would normally be required to transfer a proton from HAc to Ts^- is reduced by the gain in Coulomb energy as the negative charge moves closer to BH^+.

Reaction 2 represents symmetrical ion-pair exchange with an equilibrium constant of unity. This process makes a highly significant contribution to the measured rate of proton exchange at low HTs concentrations. The rate constants, k, in Table I are precise to ±30% and should be accurate within a factor of 2. As far as we know, they represent the first measurements of an absolute rate constant for ion-pair exchange in liquid solution. For this reason, special interest attaches to their order of magnitude.

As a basis for discussion we shall assume that reaction 2 requires the formation of a double-ion pair, which can break up with or without metathesis (eq 7).

$$\begin{array}{c} BH^+Ts^- \\ +Ac^-HB^+ \end{array} \underset{k_d}{\overset{k_a}{\rightleftharpoons}} \begin{array}{c|c} BH^+ & Ts^- \\ \hline Ac^- & HB^+ \end{array} \xrightarrow{k'_d} BH^+Ac^- + BH^+Ts^- \qquad (7)$$

The concentration of the double-ion pair would be small in our experiments.[14] On making the steady-state approximation, we obtain

$$k = k_a k_d'/(k_d' + k_d) \qquad (8)$$

k_a would be at least as large as the rate constant, k_-, described above, because the encounter of two ion pairs is facilitated by the electrostatic attraction of their dipoles.[13] If we adopt a value of 10^{10} sec^{-1} M^{-1}, then it follows from (8) and k values in Table I that k_d'/k_d is less than 10^{-2}. We conclude that reaction 2 should be classified as activation controlled rather than as diffusion controlled, that it is more like a chemical reaction than like an exchange of independent ionic partners, and that these ion pairs are genuine molecular

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(8) 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, p 596.

(9) H. B. Klevens and J. R. Platt, *J. Am. Chem. Soc.*, 71, 1714 (1949).

(10) J. N. Murrell, *J. Chem. Soc.*, 3779 (1956).

(11) W. R. Remington, *J. Am. Chem. Soc.*, 67, 1838 (1945).

(12) B. M. Wepster, *Rec. Trav. Chim.*, 71, 1159 (1952).

(13) See, for example, P. Debye, *Trans. Am. Electrochem. Soc.*, **82**, 265 (1942).

(14) Measurements of ion-pair association in similar systems have been reported by Grunwald and Price.[3]

complexes. The view that ion pairs partake of the properties of molecular complexes and can show weak electron delocalization, specific reactivity, and even isomerism is of course consistent with a variety of evidence, including optical[15] and electron spin resonance[16,17] spectra and salt effects in solvolysis.[18]

Reaction 3 transfers the N–H proton to the carboxyl group, and reaction 4 transfers the acetic acid molecule from the solvation shell of the aniline to bulk solvent, thus completing the half-cycle of exchange. We shall show with high probability that reaction 3 is a rapid reversible equilibrium in the kinetic scheme; that is, $k_i \gg k_H$. If this be granted, then the accuracy of the k_H values in Table I is measured by their precision of about $\pm 20\%$.

As far as we know, our results for k_H are the first measurements of the mean lifetimes of hydrogen-bonded solvation complexes in acetic acid. The rate constant for the dissociation of acetic acid dimer in the glacial solvent has been inferred from ultrasonic measurements and is thought to be 1.6×10^5 sec^{-1} at 25°,[19] considerably smaller than k_H, but this process requires the breaking of two hydrogen bonds. In water, k_H for amines of comparable molar volume and somewhat higher basicity would be in the range 10^9–10^{10} sec^{-1}. For example, $k_H = 2\ 7 \times 10^9$ sec^{-1} for dibenzylmethylamine at 30°.[20]

Table I shows that k_H decreases monotonically with increasing chain length of the N,N-dialkyl group: $\delta \log k_H \approx -0.34$ for each increase in chain length. This behavior is quite different from that of the measures of basicity listed in Table II, which go through a sharp maximum at Et_2, but resembles that of λ_{max} and ϵ_{max} of the "D band." The data permit the interpretation that the decrease in k_H with increasing chain length results in substantial part from London dispersion interactions.[20] Molecular models indicate that these interactions would not reach saturation until the chain length is five carbon atoms, provided that the acetic acid molecule that is hydrogen-bonded to nitrogen orients itself with its methyl group adjacent to the N,N-dialkyl group.

Experimental Part

Materials. Acetic acid (reagent grade) was dried by the addition of acetic anhydride equivalent to the water present. After standing for several days, the dry acid was purified further by partial crystallization.[7] The final product contained less than 0.003 *M* water (by Karl Fischer titration) and less than 0.007 *M* anhydride (by anthranilic acid titration[21]).

p-Toluenesulfonic acid monohydrate (Eastman White Label) was recrystallized twice from chloroform and dried in a vacuum oven at 40°, mp 102–103°. Comparison of the acid equivalent weight with Karl Fischer analysis for water established that the product was the monohydrate.

The N,N-dialkylanilines (Eastman White Label) were freshly redistilled at reduced pressure (1–2 mm) before each series of measurements. Acetic acid solutions of the anilines reacted with theoretical amounts (to better than $\pm 0.25\%$) of *p*-toluenesulfonic acid monohydrate.

p-Nitro-N,N-dimethylaniline (Eastman White Label) was recrystallized from alcohol–water and dried *in vacuo*, mp 164.5°. *p*-Cyano-N,N-dimethylaniline [*p*-(N,N-dimethylamino)benzonitrile, Aldrich] was recrystallized twice from alcohol–water, mp 74.5–75.5°. Butter yellow was recrystallized by Ceska,[7] mp 117°.

Solutions. Solutions were prepared from these materials by standard quantitative techniques. In some experiments, the water introduced with *p*-toluenesulfonic acid was removed by addition of an equivalent amount of acetic anhydride. Pure dry sodium carbonate served as primary standard for acid–base titration in glacial acetic acid. N,N-Dialkylanilinium *p*-toluenesulfonate was prepared in acetic acid solution by mixing equivalent quantities of solutions of N,N-dialkylaniline and *p*-toluenesulfonic acid.

In many of the reaction mixtures the HTs concentration was less than 10^{-3} while the BH^+Ts^- concentration was 0.05 to 0.12 *M*. In such cases the HTs concentration was determined spectrophotometrically, using *p*-nitro-N,N-dimethylaniline (N) as indicator. The basicity of N is such that N is half-converted to NH^+Ts^- when HTs = 6.8×10^{-4} *M*; precise data are given in the next section. In a typical determination, 0.200 ml of 0.002 *M* indicator was added to 7.00 ml of reaction mixture and the concentration of unprotonated N determined spectrophotometrically at 390 mμ. The concentration of HTs in the reaction mixtures (before N was added) can then be calculated from standard stoichiometric and mass-action expressions. The accuracy of this method is 10% or better for HTs concentrations above 1×10^{-4} *M* but deteriorates rapidly below this figure. All but two of the reaction mixtures used in the kinetic experiments had HTs concentrations above 1×10^{-4} *M*.

When the water introduced with *p*-toluenesulfonic acid monohydrate is not removed with acetic anhydride, the actual HTs concentration is smaller than the formal acid concentration because of reaction 9. The equilibrium constant for (9), $K_i^{H_2O \cdot HTs}$, was

$$H_2O + HTs \rightleftharpoons H_3O^+Ts^- \tag{9}$$

found to be 2.3 at 26°. All HTs concentrations reported in this paper are *actual* concentrations and have been corrected for the formation of $H_3O^+Ts^-$ when necessary. The spectrophotometric method described above, of course, measures the *actual* HTs concentration. The value of $K_i^{H_2O \cdot HTs}$ was derived from an analysis of the effect of water on the indicator ratio, $[NH^+Ts^-]/[N]$, at given acid concentrations.

Equilibrium Constants. K_i (eq 5) for N,N-diethylaniline was measured by the method of Ceska and Grunwald.[7]

Basicity toward *p*-toluenesulfonic acid (eq 6) was measured spectrophotometrically by a stepwise procedure using indicators. This procedure turned out to be accurate and convenient and can easily be adapted to other systems.

Since [B] and $[BH^+Ac^-]$ are in a constant ratio, K_i, independent of [HTs], it is convenient to define a formal base species whose concentration $[B_f]$ is equal to [B] + $[BH^+Ac^-]$. Alternative expressions for $[B_f]$ are given in (10). Furthermore, it is convenient to

$$[B_f] = [B] + [BH^+Ac^-] = [B](1 + K_i) = [BH^+Ac^-](1 + K_i)/K_i \tag{10}$$

define a formal equilibrium constant in terms of $[B_f]$, as in (11).

$$K_f^{B \cdot HTs} = [BH^+Ts^-]/[B_f][HTs] = K_i^{B \cdot HTs}/(1 + K_i) \tag{11}$$

For weak bases ($K_i << 1$) $K_f^{B \cdot HTs} \approx K_i^{B \cdot HTs}$; for strong bases ($K_i >> 1$), $K_f^{B \cdot HTs} \approx K_i^{B \cdot HTs}/K_i$; that is, $1/K_f^{B \cdot HTs} \approx$ the equilibrium constant for reaction 1.

In the first step of the procedure a dilute solution of a weak base is treated with a known excess of HTs, and the concentration $[B_f]$ of unreacted base is measured spectrophotometrically. The relevant extinction coefficient ϵ = OD/$[B_f]$ (where OD denotes optical density per unit path length) is obtained in measurements without HTs but in the presence of *ca.* 0.001 *M* sodium acetate to repress ion-pair dissociation of BH^+Ac^-. The weak base was either *p*-nitro-N,N-dimethylaniline (N) or *p*-cyano-N,N-dimethylaniline (C).[22] For N, we measured optical density at the band maximum at 390 mμ, where $\epsilon_N = 19.86 \times 10^3$. For C we measured optical density at 290 and 300 mμ; the result needed to be corrected for

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[15] T. E. Hogen-Esch and J. Smid, *J. Am. Chem. Soc.*, **87**, 669 (1965); **88**, 307, 318 (1966).

[16] N. M. Atherton and S. I. Weissman, *ibid.*, **83**, 1330 (1961).

[17] N. Hirota, *J. Phys. Chem.*, **71**, 127 (1967).

[18] S. Winstein, E. Clippinger, A. H. Fainberg, R. Heck, and G. C. Robinson, *J. Am. Chem. Soc.*, **78**, 328 (1956).

[19] E. Freedman, *J. Chem. Phys.*, **21**, 1784 (1953).

[20] E. Grunwald and E. K. Ralph, III, *J. Am. Chem. Soc.*, **89**, 4405 (1967).

[21] S. Kilpi, *Suomen Kemistilehti*, **13B**, 19 (1940).

[22] For a spectrophotometric study of basicity of these substrates in water, see M. M. Fickling, A. Fischer, B. R. Mann, J. Packer, and J. Vaughan, *J. Am. Chem. Soc.*, **81**, 4226 (1959).

Table III. Equilibrium Constants in Acetic Acid at 26°

Base	$K_f^{B\cdot HTs}$	$K_i^{B\cdot HTs}$	K_i
H_2O	2.3 ± 1	2.3	...
Aniline, *p*-NO_2-N,N-Me_2	$(1.48 \pm 0.10) \times 10^3$	1.48×10^3	...
Aniline, *p*-CN-N,N-Me_2	$(8.2 \pm 0.3) \times 10^3$	8.2×10^3	...
Aniline	$(4.2 \pm 0.4) \times 10^6$	3.8×10^7	8.04
Aniline, N,N-Me_2	$(5.8 \pm 0.5) \times 10^6$	7.9×10^7	12.7
Aniline, N,N-Et_2	$(1.05 \pm 0.10) \times 10^7$	11.6×10^8	109
Aniline, N,N-*n*-Pr_2	$(1.04 \pm 0.10) \times 10^7$	3.3×10^8	30.5
Butter yellow	...	1.40×10^6	0.105

CH^+Ts^- absorption but not for HTs absorption. $\epsilon_C = 25.3 \times 10^3$ at 290 mμ and 24.0×10^3 at 300 mμ; $\epsilon_{CH^+Ts^-} = 18$ at 290 mμ and 5.5 at 300 mμ. In a typical experiment, [HTs] is 0.05 *M*, $[BH^+Ts^-]$ is 0.002 *M*, and $[B_f]$ is between 10^{-4} and 10^{-5} *M* at equilibrium. Ion-pair dissociation could be neglected. The concentrations of the solute components were varied by factors of about 4. Results for $K_f^{B\cdot HTs}$ and their precision measures are listed in Table III.

In the second step of the procedure we measured the base strength of butter yellow (I), a color indicator, by measuring relative base strength with respect to N and C. Solutions were prepared from acetic acid (the solvent) and three solute components: N or C (about 0.01 *F*), HTs (about 10^{-4} *F*), and I (about 10^{-5} *F*). At equilibrium the HTs is largely converted to NH^+Ts^- or CH^+Ts^-; there is some IH^+Ts^-, some unreacted HTs, and some dissociated Ts^-, but the concentrations of these species are relatively small and can be estimated adequately from known values or reasonable estimates of the relevant equilibrium constants. The ratio $[BH^+Ts^-]/[B_f]$ can therefore be obtained with good accuracy and is in any case not very different from the stoichiometric ratio of the solution components, (acid)/(base-acid).

To evaluate the acid-base equilibrium for butter yellow, we measured optical density at 534 mμ where the protonated forms of this indicator (IH^+Ts^-, IH^+Ac^-, and free IH^+) absorb with $\epsilon = 40.4 \times 10^3$. The concentration of free IH^+ is relatively small and can be estimated from known values or reasonable estimates of the relevant dissociation constants.[23] The optical density data therefore yield accurate approximations to the "undissociated indicator" ratio, $([IH^+Ac^-] + [IH^+Ts^-])/[I] = K_i^I + ([IH^+Ts^-]/[I])$. At equilibrium $[IH^+Ts^-]/[I]$ is proportional to $[BH^+Ts^-]/[B_f]$. On expressing the proportionality constant in terms of defined equilibrium constants (eq 6 and 11), we obtain the following relationship (12) among quantities that differ only slightly from primary experimental quantities.

$$\frac{[IH^+Ac^-] + [IH^+Ts^-]}{[I]} = K_i^I + \frac{K_i^{I\cdot HTs}}{K_f^{B\cdot HTs}}\frac{[BH^+Ts^-]}{[B_f]} \tag{12}$$

In making the small corrections described above, we used $K_i^I = 0.105$ at 26°, K_d for acetate salts as reported by Ceska and Grunwald,[7] and assumed that for any given BH^+, K_d for the tosylate salt equals K_d for the acetate salt. $K_i^{I\cdot HTs}$ was then calculated from (12), using the known values of $K_f^{B\cdot HTs}$. Six experiments with N yielded $K_i^{I\cdot HTs} = (1.46 \pm 0.15) \times 10^6$; four experiments with C yielded $K_i^{I\cdot HTs} = (1.38 \pm 0.08) \times 10^6$. Table III lists an average value of 1.40×10^6.

In the third and final step of the procedure, $K_f^{B\cdot HTs}$ for the given aniline was measured by comparing its base strength with that of butter yellow. This experiment is similar to that described above, and eq 12 is the basic equation. Because of the greater base strength of N,N-dialkylanilines relative to butter yellow, it is possible to design this experiment so that both slope and intercept of the linear relationship (12) can be obtained with good accuracy. The results are summarized in Table IV and are in good agreement with theory. The intercepts, which should be equal to K_i^I, are indeed independent of the aniline and agree well with independent measurements of K_i^I.[7,24] The linearity of the relationship is confirmed by the close approach of the correlation coefficients to unity. Values of $K_f^{B\cdot HTs}$ were calculated from the slopes and are listed in Table III.

Table IV. Fit of Eq 12[a]

Base	Intercept	Slope	Correlation coeff
Aniline[b]	0.104	0.334	0.9985
Aniline, N,N-Me_2	0.110	0.243	0.9996
Aniline, N,N-Et_2	0.104	0.133	0.9997
Aniline, N,N-*n*-Pr_2	0.102	0.134	0.9993

[a] Indicator: butter yellow; temperature 26°. [b] Measurement by Dr. C. S. Leung in this laboratory.

Proton-Exchange Rates. We used the nmr technique to measure rates of NH-CO_2H proton exchange between BH^+Ts^- and acetic acid. Since the proton fraction of NH protons was small ($p_{NH} < 10^{-2}$), the procedures described by Grunwald and Price could be used.[3] Briefly, we measured T_2 and T_1 of the carboxyl (or of the exchange-averaged CO_2H-NH) proton resonance at 56.4 MHz by spin-echo techniques and deduced the exchange broadening Δ, which in turn permitted calculation of the rate of proton exchange.[3] Proton exchange between carboxyl protons of HAc and OH protons of HTs, $H_3O^+Ts^-$, or H_2O was very fast under the conditions of our rate measurements, and we were unable to detect any exchange broadening in the absence of BH^+Ts^-. As a result, the "carboxyl" resonance is really a fully exchange-averaged resonance of all OH protons. However, the NH-"CO_2H" chemical shift varied only slightly with acid and water concentration at the low concentrations employed, and in the rate calculations we allowed for fast exchange with these solutes only to the extent of defining the proton fraction p_{NH} with respect to *all* OH protons.

Figure 1 shows a plot of Δ/p_{NH} *vs.* log [HTs] for dimethylanilinium *p*-toluenesulfonate at 26°. The figure is a composite of three series of experiments, two of them done early in this research and of less than typical precision. We selected these data for illustration because Δ/p_{NH} *in this case* is independent of the substrate concentration; that is, proton exchange is first order in $[BH^+Ts^-]$. Moreover, the concentration of water (introduced by using HTs·H_2O for the acid in two of the series) goes to higher values than in any experiments on the other substrates. Two facts stand out: the plot is unsymmetrical about a well-defined maximum; and Δ/p_{NH} is independent of H_2O and $H_3O^+Ts^-$ and depends only on the concentration of un-ionized HTs.

The following semiquantitative experiments helped us in the interpretation of Δ/p_{NH}. We measured the full proton nmr spectrum in slow passage for each substrate in the presence of fairly strong acid at 26°. Solute components in these experiments were: HTs·H_2O, about 2.1 *F;* N,N-dialkylaniline about 1.4 *F*. In each case we observed the NH-proton resonance of BH^+Ts^- directly as a broad singlet about 1.2–1.3 ppm upfield from the "CO_2H" resonance. We also measured the CH_3-proton resonance of 0.265 *M* N,N-dimethylanilinium *p*-toluenesulfonate in slow passage at 0° in supercooled liquid solution over a 15-fold variation of [HTs]. At the highest HTs concentration, 0.4 *M*, the CH_3-proton resonance is a fairly well-resolved doublet (*J* = 6.1 cps) due to spin-spin interaction with the NH proton, which evidently exchanges relatively slowly under these conditions. As the acid concentration is decreased, the CH_3 resonance shows the characteristic changes associated with NH-proton exchange at progressively increasing

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(23) Dissociation of IH^+Ts^- is repressed by the presence of BH^+Ts^- at much higher concentration.

(24) S. Bruckenstein and I. M. Kolthoff, *J. Am. Chem. Soc.*, **78**, 10 (1956).

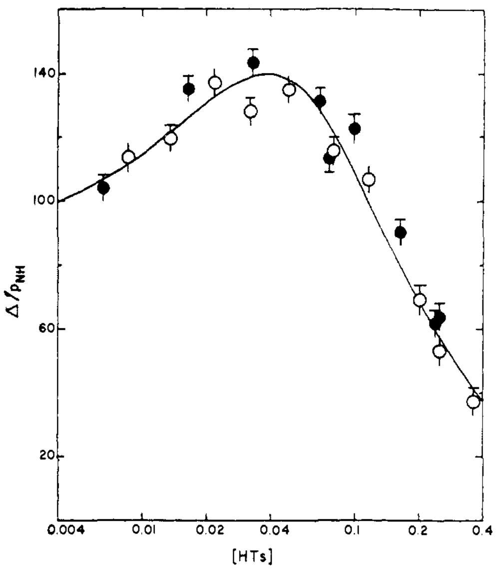

Figure 1. Summary of Δ/p_{NH} for N,N-dimethylanilinium *p*-toluenesulfonate at 26° in the presence of HTs·H_2O (solid circles) or anhydrous HTs (open circles). The abscissa (labeled [HTs]) is the concentration of un-ionized *p*-toluenesulfonic acid. The concentration of BH^+Ts^- is 0.05 *M* (cross-bar at top), 0.10 *M* (no cross-bar), or 0.15 *M* (cross-bar at bottom). The smooth curve is calculated from eq 13, 14, and 19, and from rate constants as listed in Tables I and V.

rate:[25] it collapses into a singlet at [HTs] ≈ 0.1 *M*, which sharpens progressively as the HTs concentration is further reduced in several steps to 0.03 *M*, the lowest value tried.

Returning to the interpretation of Δ/p_{NH}, we must therefore assume that the rate of proton exchange is slow in strong acid and increases monotonically with decreasing [HTs]. A rigorous interpretation of the data is made difficult by the complexity of the spin-spin interaction of the NH proton with (i) α protons of the N,N-dialkyl groups, and (ii) with the N^{14} nucleus. *J* for the first interaction is about 6 Hz. The effective *J* for the second interaction is unknown, but it is known to decrease with the T_1 relaxation time of the N^{14} nucleus.[3] The fact that the NH-proton resonance in strong acid is a broad singlet suggests that T_1 is small.

We did not think it worthwhile to try to allow explicitly for spin-spin interaction. Instead, we treated our data simply as proton exchange between two nmr lines, NH and COOH, separated by an effective chemical shift parameter, δ, which we evaluated for each substrate from the maximum in the plot of Δ/p_{NH} (see Figure 1). Since p_{NH} is very small, the relevant equations[3] reduce to eq 13–15, where *R* denotes the rate of proton exchange.

$$\Delta/p_{NH} = \delta^2\tau/(1 + \delta^2\tau^2) \tag{13}$$

$$\delta = 2(\Delta/p_{NH})_{max} \tag{14}$$

$$1/\tau = R/[BH^+Ts^-] \tag{15}$$

According to our error analysis, this approximate treatment should not raise the error in τ significantly for $\delta\tau < 0.4$, which range includes more than three-quarters of our data. Below this range systematic error may become significant, but should not exceed 20% even at the highest rates. δ parameters obtained at 26° and 56.4 MHz are as follows: for N,N-Me_2, 0.79 ppm; for N,N-Et_2, 1.33 ppm; for N,N-*n*-Pr_2, 1.23 ppm.

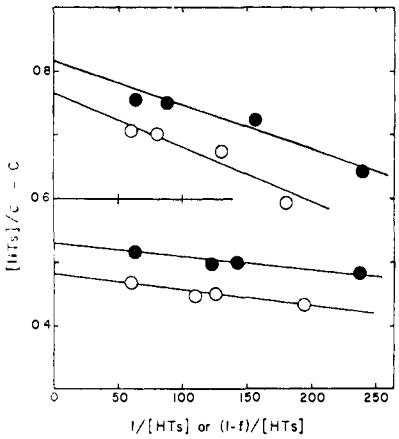

Figure 2. Kinetic plots according to eq 21: upper data, N,N-di-*n*-propylanilinium *p*-toluenesulfonate; lower data, N,N-diethylanilinium *p*-toluenesulfonate; solid circles, abscissa is 1/[HTs], *C* = 0.0; open circles, abscissa is (1 − *f*)/[HTs], *C* = 0.05.

Kinetic Analysis

The data presented in the Experimental Part consistently suggest that we have measured the rate of exchange of NH protons in BH^+Ts^- with carboxyl protons in acetic acid molecules in the bulk solvent. These rates are measured at dynamic equilibrium. The concentrations of other molecular forms of BH^+, such as BH^+Ac^- or free BH^+ ions, are too small relative to that of BH^+Ts^- to contribute significantly to the measured exchange broadening under these conditions. However, such species may enter the kinetic analysis as reactive intermediates.

BH^+Ac^- As a Reactive Intermediate. We find that we cannot rationalize the kinetic observations unless we assume that BH^+Ac^- is a reactive intermediate, that is, the rate of proton exchange must be expressed as in (16). If we let k_e = probability per second that a

$$R = (\text{rate of conversion of } BH^+Ts^- \text{ to } BH^+Ac^-) \times (\text{probability that } BH^+Ac^- \text{ exchanges its NH proton before reverting to } BH^+Ts^-) \tag{16}$$

BH^+Ac^- ion pair exchanges its NH proton, and if the conversion of BH^+Ts^- to BH^+Ac^- involves only reactions 1 and 2, we can rewrite (16) in the form

$$R = [BH^+Ts^-](k_+ + k[BH^+Ac^-]) \times \{k_e/(k_e + k_-[HTs] + k[BH^+Ts^-]\} \tag{17}$$

Finally, we make the substitutions, $R/[BH^+Ts^-] = 1/\tau$ (eq 15), and $k_+/k_- = K = [BH^+Ac^-][HTs]/[BH^+Ts^-]$, to obtain (18). Note that k_e would be identical with

$$\frac{1}{\tau} = \frac{k_eK}{[HTs]}\,\frac{k[BH^+Ts^-] + k_-[HTs]}{k_e + k_-[HTs] + k[BH^+Ts^-]} \tag{18}$$

the first-order rate constant for proton exchange if the substrate were BH^+Ac^- rather than BH^+Ts^-.

According to (18) or any other conceivable mechanism that involves BH^+Ac^- as a reactive intermediate, as [HTs] becomes large, [HTs]/τ approaches a constant value equal to k_eK. As shown in Figure 2, [HTs]/τ is in fact nearly constant at experimentally accessible HTs concentrations. A short extrapolation to 1/[HTs] = 0 yields k_eK and, because *K* is known, k_e. Results are shown in Table V.

We now wish to show that k_e is of precisely the right magnitude to be identified with the rate constant for proton exchange of BH^+Ac^-. The latter rate constant has been measured in acetic acid for a number of acetate salts of aliphatic amines[3] and increases in almost direct proportion with the acid dissociation constant, K_A, of the aliphatic BH^+ in water.[3] This relationship is shown

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[25] A. Loewenstein and S. Meiboom, *J. Chem. Phys.*, 27, 1067 (1957).

Table V. Kinetic Constants at 26°

Anilinium *p*-toluenesulfonate	k_eK	k_e	1000 k_e/k_-	kK
N,N-Dimethyl, 0.05–0.15 *M*	16.2[a]	1.01×10^8	20.0	...
N,N-Diethyl, 0.056 *M*	0.48	(av)	0.46	6.3 ± 1.8
0.11 *M*	0.53	5.5×10^6	0.57	6.0 ± 0.8
N,N-Di-*n*-propyl, 0.10–0.12 *M*	0.815	8.8×10^6	1.04	4.6 ± 1.4

[a] At 0°, $k_eK \approx 4$ on the basis of rate measurements of NH-proton exchange by means of the N,N-dimethyl proton resonance. In this series, $[BH^+Ts^-] = 0.265$ *M* and $[HTs] = 0.4$ to 0.03 *M*. The solutions were liquid but were probably supercooled at 0°.

in logarithmic form in Figure 3, where the straight line has been drawn with unit slope. When the relationship is extrapolated to the pK_A values of the present substrates, the k_e values are reproduced within a factor of 2 (Figure 3, points 5–7). We think that agreement of this quality is very pretty, especially when the variables change by almost six orders of magnitude.

Proton Exchange of BH^+Ac^-. Proton exchange between BH^+Ac^- and acetic acid has been studied by Grunwald and Price.[3] Their evidence is strongly in favor of reaction in two steps, (3) and (4). On that basis, $k_e = k_{-i}k_H/(k_i + k_H)$. They considered two limiting cases. If step 3 is rate determining, then $k_e = k_{-i}$ and $k_eK_i = k_i$. If step 4 is rate determining ($k_i \gg k_H$), then $k_e = k_Hk_{-i}/k_i = k_H/K_i$, and $k_eK_i = k_H$. They did not have direct experimental values for K_i but argued on other grounds that step 4 is rate determining.

In the present case experimental values for K_i are available, and the variation of k_eK_i with molecular structure of the substrate can be examined. This is done in Table I, section 4; the numbers labeled k_H are really k_eK_i. If step 3 were rate determining, k_eK_i would be equal to k_i, and we would expect a Brønsted relationship with basicity,[26] that is, k_eK_i should go through a maximum when the substrate is N,N-diethylaniline (see Table II). In fact, no such maximum exists. Instead, k_eK_i decreases monotonically with increasing alkyl size, providing strong support for the theory that k_eK_i is a close approximation to k_H. If this be granted, it follows that $k_i > k_H$, and since k_H is already large (*e.g.*, 1.3×10^9 sec^{-1} for N,N-dimethylaniline), k_i would be on the order of 10^{10} or greater. Proton transfer with a rate constant of this magnitude would be considered "ultrafast."[27]

Formation of BH^+Ac^- from BH^+Ts^-. For N,N-dimethylanilinium *p*-toluenesulfonate, ion-pair exchange could be neglected. When the corresponding term, $k[BH^+Ts^-]$, is dropped from (18), the rate law reduces to (19). The plot of $[HTs]/\tau$ *vs.* $1/\tau$ was linear, as required by (19), and led to the rate constants listed in

$$\frac{1}{\tau} = \frac{Kk_ek_-}{k_e + k_-[HTs]} = \frac{k_+}{1 + (k_-/k_e)[HTs]} \quad (19)$$

Tables I and V. The ability of eq 19 to reproduce the experimental data is illustrated in Figure 1 and is quite satisfactory. The validity of the rate law is confirmed by the plausible values obtained for k_e and k_-.

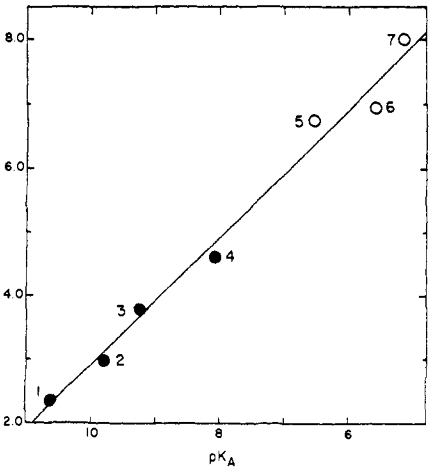

Figure 3. Rate constants for proton exchange of BH^+Ac^- in acetic acid (k_e) compared with acid dissociation constants of BH^+ in water (K_A) (data at 25 ± 1°). Points correspond to the following substrates: B = (1) CH_3NH_2; (2) $(CH_3)_3N$; (3) NH_3; (4) $(HOCH_2)_3CNH_2$; (5) $C_6H_5NEt_2$; (6) $C_6H_5N\text{-}n\text{-}Pr_2$; (7) $C_6H_5N(CH_3)_2$. The straight line is drawn with unit slope.

For N,N-diethyl- and N,N-di-*n*-propylanilinium *p*-toluenesulfonate, eq 19 would not fit. Rates measured at the higher acid concentrations were consistent with plausible values for k_-, but rates measured at the lower acid concentrations were far higher than expected. These facts can be given a plausible explanation by a kinetic scheme that includes ion-pair exchange.

We begin by showing that ion-pair exchange becomes insignificant at the higher acid concentrations. We suppose that [HTs] is high enough so that $k_-[HTs] > k_e + k[BH^+Ts^-]$. We may then expand the right-hand side of (18) as a power series in 1/[HTs] to obtain (20). Successive terms in this series differ by factors

$$\frac{1}{\tau} = \frac{k_eK}{[HTs]} - \frac{1}{[HTs]^2}\frac{k_e^2K}{k_-} + \frac{1}{[HTs]^3}\frac{k_e^2K}{k_-^2}(k[BH^+Ts^-] + k_e) \quad (20)$$

that are less than unity, by hypothesis, and the series converges. The first term to show the effect of ion-pair exchange is the cubic term, and this term would be small under conditions where the second term begins to be significant.

Equation 20 suggests the following method for evaluating k_-. On rearranging and neglecting terms beyond the cubic term, we obtain

$$[HTs]/\tau = k_eK\left(1 - \frac{k_e}{k_-}\frac{(1 - f)}{[HTs]}\right) \quad (21a)$$

$$f = (k[BH^+Ts^-] + k_e)/k_-[HTs] < 1 \quad (21b)$$

Thus a plot of $[HTs]/\tau$ *vs.* $(1 - f)/[HTs]$ should be a straight line with intercept = k_eK, and slope divided by intercept = k_e/k_-. Since f is small, $(1 - f)$ will be sufficiently accurate even if a fairly crude approximation is used for k/k_-. Plots according to eq 21 are shown in Figure 2, and parameters derived from them are listed in Table V. Within the experimental error, k_e/k_- is found to be independent of $[BH^+Ts^-]$, as expected. The method leads to plausible values for k_- (see Table I).

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(26) See, for example, L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, Chapter 7; J. E. Leffler, *Science*, 117, 340 (1953).

(27) E. Grunwald, *Progr. Phys. Org. Chem.*, 3, 317 (1965).

The failure of eq 19 at the lower acid concentrations is illustrated in Table VI. In the first two columns of that table, experimental values of $1/\tau$ are compared with calculations of $k_+/(1 + k_-[\mathrm{HTS}]/k_e)$, using rate constants obtained at the higher acid concentrations according to (21). The discrepancy between the two columns increases with decreasing [HTs] and becomes highly significant below $5 \times 10^{-4} M$.

Table VI. Fit of Eq 18 to Data for N,N-Diethylanilinium *p*-Toluenesulfonate at 26°

$1/\tau$, sec^{-1}	$k_+/(1 + k_-[\mathrm{HTs}]/k_e)$ sec^{-1}	10^3[HTs], *M*	$[\mathrm{BH^+Ts^-}]$, *M*	*kK*
281	257	1.522	0.104	7.9
744	492	0.556	0.110	(13)[a]
932	581	0.395	0.110	7.7
1102	625	0.332	0.112	7.9
1243	677	0.268	0.110	5.9
1735	768	0.178	0.107	5.3
2150	828	0.128	0.105	4.3
2980	902	(0.076)[b]	0.103	(3.2)[a]
750	596	0.372	0.056	3.8
1150	694	0.249	0.056	6.7
1460	742	0.201	0.056	8.3

[a] Omitted from average. [b] Outside the range in which [HTs] can be measured with good accuracy.

The final column of Table VI shows the extent to which inclusion of ion-pair exchange in the rate law succeeds in fitting the data. Using the values of k_e, k_+, and k_- obtained at the higher acid concentrations, we have calculated the parameter *kK* for each experimental point, using eq 18. The values obtained in this way are fairly constant and the scatter could be random. What is more significant, Table V shows that average values obtained in the series with 0.11 *M* $\mathrm{BH^+Ts^-}$ and with 0.056 *M* $\mathrm{BH^+Ts^-}$ are almost exactly equal. This is to say, the kinetics of interconversion of $\mathrm{BH^+Ts^-}$ and $\mathrm{BH^+Ac^-}$ by this mechanism is, on the average, consistent with that for ion-pair exchange.

Alternatives to $\mathrm{BH^+Ac^-}$ As the Reactive Intermediate. We have seen that a reaction mechanism involving $\mathrm{BH^+Ac^-}$ as the reactive intermediate can accommodate all facts. We now wish to show that there are few if any alternatives.

One constraint is imposed by the kinetics. The fact that the rate becomes proportional to $[\mathrm{BH^+Ts^-}]/[\mathrm{HTs}]$ at the higher acid concentrations implies that the molecular formula of the intermediate is $[\mathrm{BH^+Ts^-} - \mathrm{HTs} + n\mathrm{HAc}]$; that is, the intermediate must be either solvated $\mathrm{BH^+Ac^-}$ or solvated B or a subspecies or isomer of these species. The other constraint is imposed by the rate itself. For definiteness, consider the rate constant, k_-, for the reaction of the intermediate with HTs to form $\mathrm{BH^+Ts^-}$, and suppose that k_- is obtained by a graphical method based on eq 21, that is, $k_- = (\mathrm{intercept})^2/(\mathrm{slope} \cdot K)$. Intercept and slope are defined by the kinetic data, but *K* is the equilibrium constant for the formation of the intermediate, and its value will depend on the particular species we propose for the intermediate. If we modify our theory so as to make the intermediate less stable, *K* will decrease and k_- will increase. But k_- cannot be larger than the encounter frequency of the intermediate with HTs and should certainly be less than $10^{11}\ \mathrm{sec}^{-1}\ M^{-1}$. If we assume that the intermediate is $\mathrm{BH^+Ac^-}$, k_- turns out to be $10^{10}\ \mathrm{sec}^{-1}\ M^{-1}$. We conclude, therefore, that any intermediate we propose must be so stable that its concentration is at least one-tenth as great as that of solvated $\mathrm{BH^+Ac^-}$. This constraints rules out all mechanisms that propose solvated B or any of its subspecies for the intermediate (see K_i values in Table I).

Reactions of $\mathrm{He^3H^+}$ Ions with Gaseous Hydrocarbons.
I. Toluene

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Abstract: The gas-phase reaction of the $\mathrm{He^3H^+}$ ions from the decay of molecular tritium with toluene was investigated under conditions ensuring that the labeled decay fragments were the only source of tritiated products, with no significant interference from radiolytic processes. The reaction responsible for the formation of tritiated toluene, the major product isolated with a yield of 50 to 60%, appears to be similar to the electrophilic aromatic substitutions occurring in solution, as indicated by experiments involving the competition for the $\mathrm{He^3H^+}$ ions between toluene and benzene, and benzene and α-trifluorotoluene, and by the analysis of the isomeric distribution of the tritiated toluenes formed. The mechanisms responsible for the formation of the two other products identified, *i.e.*, labeled methane and benzene, are also discussed.

The existence of the singly charged helium hydride molecular ion $\mathrm{HeH^+}$ has been known for many years by mass spectrometric investigations,[1] and several quantum mechanical calculations for the ground state of this very simple two-electrons heteronuclear system were

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[1] T. R. Hogness and E. C. Lunn, *Phys. Rev.*, 26, 44 (1925).