is responsible for the unusual reaction with deoxyguanosine is under investigation in this laboratory.

It is also possible that electrophilic species such as SO_3 could play a role in the in vivo chemistry of the carcinogenic sulfate esters. In fact, as a result of this investigation it is obvious that the chemistry of the sulfuric acid esters of *N*-hydroxy-*N*-arylamides is considerably more complicated than previously appreciated. It is our intention to continue to investigate the chemistry of these species with special emphasis on the points mentioned above.

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Registry No. 1a, 91631-50-4; **1b**, 91631-52-6; **1c**, 91631-54-8; **1d**, 91631-56-0; **1e**, 91631-58-2; **1f**, 91631-60-6; *N*-hydroxy-*p*-bromoacetanilide, 67274-48-0; *N*-hydroxy-3,4-dichloroacetanilide, 86412-49-9.

Supplementary Material Available: Table of pseudo-first-order rate constants for the solvolysis of *N*-sulfonoxyacetanilides in aqueous solution (2 pages). Ordering information is given on any current masthead page.

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Deuterium Fractionation Factor for Unhydrated Hydronium Ion. Deuterium Isotope Effects on Proton-Transfer Equilibria in Acetonitrile

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Abstract: Spectrophotometric measurements of pK_a values for three *N,N*-dimethylanilinium ion indicators have been made in CH_3CN-H_2O and CH_3CN-D_2O solvent mixtures throughout the range of water content from a mole fraction of 1.00 down to a mole fraction of 3×10^{-4}. The three indicator bases were *N,N*-dimethyl-*p*-nitroaniline, *N,N*-dimethyl-3,5-dinitro-4-toluidine, and 4-chloro-*N,N*-dimethyl-2,6-dinitroaniline. The observed H_2O/D_2O isotope effects on the pK_a values allow estimation of the value of ℓ_1, the deuterium fractionation factor relative to L_2O for unhydrated L_3O^+ (where L = H or D) in CH_3CN; that value is near 0.79. (Compare $\ell = 0.69$ for L_3O^+ in L_2O.) This observation implies that if, as has commonly been assumed, the low value of ℓ in liquid L_2O results from strong hydrogen bonding of L_3O^+ to three L_2O molecules, then hydrogen bonding of L_3O^+ to CH_3CN molecules in liquid CH_3CN is almost as effective in weakening the force field experienced by the three fractionated L's in L_3O^+ as is hydrogen bonding to L_2O molecules in liquid L_2O. The observation that ℓ_1 is significantly less than 1.0 also supports our contention that the absence of an H_2O/D_2O *kinetic* isotope effect on methyl transfer to L_2O in CH_3CN implies that no significant L_2O-CH_3 bond is present in the transition state and thus that the activation process for those methyl transfers is predominantly a fluctuation in solvent polarization.

It has been known for almost a half a century[1] that in aqueous solutions the D/H ratio in "hydrogen ion" differs from the D/H ratio in the water in which that hydrogen ion is dissolved. It also has been known for about two decades[2-6] that this D/H fractionation occurs at three sites in the hydrogen ion (thus confirming its formulation as hydronium ion, L_3O^+, where L denotes either H or D) and that the value of the fractionation factor relative to L_2O, ℓ, is 0.69 ± 0.02 near 25 °C. However, the molecular origin of this fractionation remains uncertain. Most attempts to account for this value of ℓ have emphasized the importance of hydrogen bonding of the L_3O^+ moiety to three L_2O molecules, giving $(L_2O\text{--}L)_3O^+$, or of the effect of such hydrogen bonding on the libration frequencies of the L_3O^+ unit.[7-10] Yet others have qualitatively attributed this fractionation to the effect of the positive charge on the internal vibrations of L_3O^+.[11]

Knowledge of the true origin of the value of ℓ in aqueous solution would further our understanding of the structure of aqueous solutions, obviously an important topic, and also would allow a more certain interpretation of kinetic deuterium isotope effects on alkyl transfers to water (eq 1) in aprotic solvents. There

$$L_2O + RX \rightarrow ROL_2^+ + X^- \qquad (1)$$

exists strong evidence[12-15] that solvent repolarization is rate determining in these reactions, a controversial conclusion. That evidence is based in part on our observation[12,13] that values of k_{H_2O}/k_{D_2O} for methyl transfers from CH_3OClO_3, $CH_3OSO_2CF_3$, and $CH_3S^+C_4H_4$ to L_2O which is a dilute solute in acetonitrile

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(1) For example: (a) Korman, S.; LaMer, V. K. *J. Am. Chem. Soc.* **1936**, *58*, 1396-1403. (b) Gross, P.; Steiner, H.; Suess, H. *Trans. Faraday Soc.* **1936**, *32*, 883-889. (c) Abel, E.; Bratu, E.; Redlich, O. *Z. Phys. Chem., Abt. A* **1935**, *173A*, 353-364.

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(3) Gold, V. *Proc. Chem. Soc., London* **1963**, 141-143.

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(6) Salomaa, P.; Aalto, V. *Acta Chem. Scand.* **1966**, *20*, 2035-2042.

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(10) More O'Ferrall, R. A.; Koeppl, G. W.; Kresge, A. J. *J. Am. Chem..Soc.* **1971**, *93*, 1-8.

(11) For example: Schowen, R. L. In "Isotope Effects on Enzyme-Catalyzed Reaction"; Cleland, W. W., O'Leary, M. H., Northrop, D. B., Eds.; University Park Press: Baltimore, 1977; p 69.

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(14) Kurz, J. L. *Acc. Chem. Res.* **1972**, *5*, 1-9.

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or sulfolane are all equal to 1.00 within experimental error. If ℓ for L_3O^+ in these aprotic solvents is significantly less than unity (as it is in L_2O), then the observed absence of deuterium isotope effects on the rates of those methyl transfers implies that little or no L_2O–CH_3 bonding is present in their transition states. On the other hand, if the low value of ℓ for L_3O^+ in L_2O results from strong hydrogen bonding of L_3O^+ to L_2O, it is possible that the corresponding hydrogen bonding of L_3O^+ to CH_3CN or $(CH_2)_4SO_2$ could be sufficiently weak to allow ℓ to be near 1.00 in those solvents; in that case, no conclusion concerning the L_2O–CH_3 bonding in those transition states could be drawn from the observed absence of an isotope effect.

This paper describes measurements of deuterium isotope effects on *equilibrium* constants for proton transfer from L_3O^+ in CH_3CN. These measurements imply that ℓ is indeed significantly less than 1.00 when the L_3O^+ is solvated by CH_3CN rather than by L_2O and thus that the above-mentioned *kinetic* isotope effect measurements do imply that significant L_2O–CH_3 bonding is absent in those methyl-transfer transition states and that the activation process is primarily a fluctuation in solvent structure.

Results and Discussion

The objective of these measurements was the evaluation of ℓ_1, the fractionation factor relative to L_2O for the *mono*hydrated proton (or deuteron), L_3O^+, in CH_3CN. The approach used to evaluate ℓ_1 was via the measurement of K_a^H and K_a^D for indicators B. Three substituted dimethylanilines (*N*,*N*-dimethyl-*p*-nitroaniline, **1**; *N*,*N*-dimethyl-3,5-dinitro-4-toluidine, **2**; 4-chloro-*N*,*N*-dimethyl-2,6-dinitroaniline, **3**) whose pK_a values could be measured in overlapping ranges of solvent composition were used as indicators. The dissociation constants K_a^H and K_a^D for each indicator are defined by eq 2 and 3 where L = H or D, L^+ is the

$$BL^+ \rightleftharpoons L^+ + B \tag{2}$$

$$K_a^L = \frac{[L^+][B]}{[BL^+]}\frac{\gamma_L\gamma_B}{\gamma_{BL}} \tag{3}$$

conjugate acid of the solvent(s), and the γ's are activity coefficients relative to a reference state of infinite dilution for all solute species in a CH_3CN–L_2O mixed solvent with the specified composition to which K_a refers.

If L_3O^+ were the only solvent conjugate acid present, then the observed value of K_a^H/K_a^D for an indicator, B, would depend only on ℓ_1 and ϕ_{BL} (the fractionation factor for BL^+), and ℓ_1 would equal $(\phi_{BL}K_a^D/K_a^H)^{1/3}$.[16] However, in aqueous CH_3CN there are five possible L^+ species: CH_3CNL^+ and $(L_2O)_iL^+$ for $1 \leq i \leq 4$. Let the fractionation factors relative to L_2O for these five species be denoted as ϕ_{AL} and ℓ_i, respectively. The observed value of K_a^H/K_a^D can thus depend on the values of six fractionation factors, ϕ_{AL}, ϕ_{BL}, and the four ℓ_i's. Since the activity coefficient ratio, $\gamma_L\gamma_B/\gamma_{BL}$, in eq 3 should have very nearly the same value for L = H and L = D when the total amounts of $LClO_4$, L_2O, B, and CH_3CN present in the H and D systems are equal, K_a^H/K_a^D was calculated as the quotient of the concentration ratios, $[H^+][B]/[BH^+]$ and $[D^+][B]/[BD^+]$.

The values of the equilibrium constants, K^H_i where $1 \leq i \leq 4$, governing the distribution of the five L^+ species when L = H (eq 4, 5) have been determined by Chantooni and Kolthoff.[17] The

$$iH_2O + CH_3CNH^+ \rightleftharpoons (H_2O)_iH^+ + CH_3CN \tag{4}$$

$$K^H_i = [(H_2O)_iH^+]/[CH_3CNH^+][H_2O]^i \tag{5}$$

distribution of these species calculated from those K^H_i is shown in Figure 1. The abcissa in Figure 1 and later figures, X_{H_2O}, is the mole fraction of *free* [i.e., not bound to H^+ in $(H_2O)_iH^+$] H_2O in the solution. Similarly, $[H_2O]$ in eq 5 is the molarity of *free* H_2O.

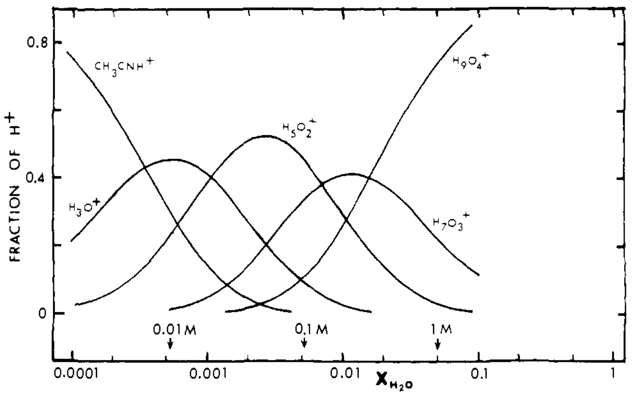

Figure 1. Distribution of the five species of "H⁺" in acidic aqueous acetonitrile. The curves are calculated from the K^H_i values given in ref 17. The logarithmic abscissa in all three figures is graduated both in the mole fraction (X) and in the molarity (M) of H_2O.

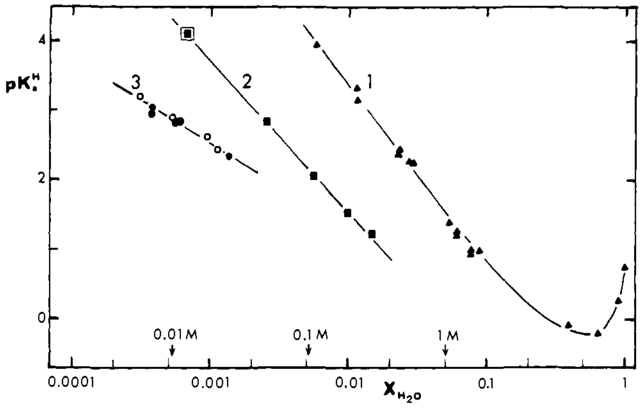

Figure 2. Values of pK_a^H for **1** (triangles), **2** (squares), and **3** (circles) in aqueous acetonitrile. Closed symbols represent values measured using acetonitrile, which either was dried as described in the Experimental Section (for $[H_2O] < 1$ M) or used as received (for $[H_2O] > 1$ M). Open symbols represent values measured by using acetonitrile which had been made 1 M in D_2O before being dried. The concentric pair of squares for **2** represent identical values obtained by using the two kinds of dried acetonitrile.

Figure 2 shows the measured values of pK_a^H for the three overlapping indicators in the protio system as functions of X_{H_2O}. Compare the different slopes in Figure 2 to the distributions of the $(H_2O)_iH^+$ in Figure 1: the average values of i in the ranges of X_{H_2O} spanned by the plots for indicators **3** and **2** and by the linear part of the plot for indicator **1** are close to 1, 2, and 3, while the observed values of $\partial pK_a/\partial \log X_{H_2O}$ for these ranges are −1.2, −2.2, and −2.6, suggesting that the dominant factor controlling the dependence of pK_a^H on X_{H_2O} when X_{H_2O} is small is the number of H_2O's released when a proton is transferred from $(H_2O)_iH^+$ to B.

Following Chantooni and Kolthoff,[17] we can write each of the observed dissociation constants K_a^H and K_a^D as

$$K_a^L = K_B^L y^L \tag{6}$$

where

$$y^L = 1 + \sum_{i=1}^{4} K^L_i [L_2O]^i \tag{7}$$

and K_B^L is the equilibrium constant for $BL^+ + CH_3CN \rightleftharpoons CH_3CNL^+ + B$.

$$K_B^L = [CH_3CNL^+][B]/[BL^+] \tag{8}$$

Since $K_B^H/K_B^D = \phi_{BL}/\phi_{AL}$, the observed equilibrium isotope effect becomes

$$K_a^H/K_a^D = (\phi_{BL}/\phi_{AL})(y^H/y^D) \tag{9}$$

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[16] The fractionation factor relative to water for any exchangeable site, XL, is defined as the equilibrium constant for LOD + XH ⇌ LOH + XD. For a recent review of the formulation of isotope effects in terms of fractionation factors, see: Schowen, K. B. J. In "Transition States of Biochemical Processes"; Gandour, R. D., Schowen, R. L., Eds.; Plenum Press: New York, 1978; Chapter 6.

[17] Chantooni, M. K., Jr.; Kolthoff, I. M. *J. Am. Chem. Soc.* **1970**, *92*, 2236–2239.

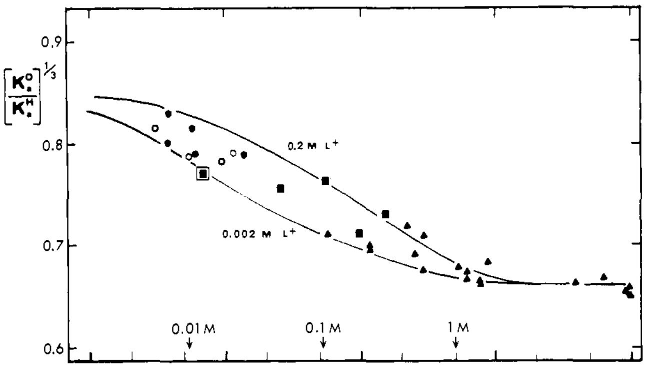

Figure 3. Values of $(K_a{}^D/K_a{}^H)^{1/3}$ for **1**, **2**, and **3** in aqueous acetonitrile. Symbols have the meaning defined in Figure 2. Comparison to Figure 1 shows that a significant number of points fall in the range of X_{H_2O} in which the predominent forms of H^+ are H_3O^+ and CH_3CNH^+. The curves are calculated from $\phi_{AL} = 0.70$, $\phi_{BL} = 1.14$, $\ell_1 = 0.79$, $\ell_2 = 0.76$, $\ell_3 = 0.72$, $\ell_4 = 0.69$, and $[L^+]$ equals either 0.002 or 0.2 M.

where y^H can be evaluated from eq 7 by using the known[17] values of the $K^H{}_i$'s together with the value of $[H_2O]$ calculated from those $K^H{}_i$'s as described.[18]

Since the value of $[D_2O]$ in each deuterio system could be calculated analogously from the values of the $K^D{}_i$'s, knowledge of those $K^D{}_i$ values would be sufficient to allow evaluation of y^D from eq 7. The needed relationship between $K^D{}_i$ and $K^H{}_i$ can be written in terms of the ℓ_i's defined above. Since even in aqueous solution, only the inner three L's in $(L_2O)_4L^+$ [more revealingly written as $(L_2O\cdot\cdot L)_3O^+$] undergo significant D/H fractionation,[2-6] we can assume that this remains true as CH_3CN molecules are successively substituted for the three hydrogen-bonded L_2O molecules in $(L_2O\cdot\cdot L)_3O^+$, and thus that each $K^D{}_i$ is related to the corresponding known value of $K^H{}_i$ by eq 10.

$$K^D{}_i = K^H{}_i(\ell_i{}^3/\phi_{AL}) \qquad (10)$$

Equations 7, 9, and 10 allow the theoretical dependence of $K_a{}^H/K_a{}^D$ on the molarity (or mole fraction) of free H_2O in the protio system to be calculated from any set of assumed values for ϕ_{AL}, ϕ_{BL}, and the four ℓ_i's. Because of the limiting behavior of $K_a{}^H/K_a{}^D$ as $X_{H_2O} \rightarrow 1$ and $X_{H_2O} \rightarrow 0$ (vide infra), plots of that dependence are conveniently displayed as $(K_a{}^D/K_a{}^H)^{1/3}$ vs. X_{H_2O}. Figure 3 shows two such theoretical curves along with the experimental values. Both curves correspond to the same set of values for ϕ_{AL}, ϕ_{BL}, and the ℓ_i's but differ in the assumed value of $[L^+]$. The dependence on $[L^+]$ results from the isotope effect on the values of the $K^L{}_i$'s. Each experimental value of $(K_a{}^D/K_a{}^H)^{1/3}$ is calculated from measurements on a pair of systems which had been prepared from the same *total* numbers of formula weights of H_2O and D_2O and of $HClO_4$ and $DClO_4$. Because of the isotope effect on $K^L{}_i$, the amount of D_2O which is bound into $(D_2O)_iD^+$ in the deuterio system will be less than the amount of H_2O which is bound into $(H_2O)_iH^+$ in the protio system with which it is paired. Thus $K_a{}^D$ corresponds to a higher value of $[L_2O]$ than does $K_a{}^{H+}$, and the observed value of $(K_a{}^D/K_a{}^H)^{1/3}$ will be larger (i.e., closer to 1.00) than the true value which would be observed if the concentration of free D_2O were equal to that of free H_2O. Only in the limit as $[L^+] \rightarrow 0$ does the observed value of $(K_a{}^D/K_a{}^H)^{1/3}$ become equal to the true value, so that points corresponding to higher values of $[L^+]$ are expected to scatter above those corresponding to low values of $[L^+]$. Almost all of the observations were of systems with $0.002 \leq [L^+] \leq 0.2$ M, and the two theoretical curves in Figure 3 correspond to those limiting values. (The curve for $[L^+] = 0.002$ M is indistinguishable on the scale of Figure 3 from that for the limit as $[L^+] \rightarrow 0$). In addition to that which results from the varying $[L^+]$, some of the scatter of the experimental points possibly results from isotope effects on the hydration of ClO_4^- and/or BH^+.

Values of the six fractionation factors used to calculate the theoretical curves in Figure 3 were chosen as follows. It is known[2-6] that $\ell^4 = 0.69$. In the limit as $X_{H_2O} \rightarrow 1$, $(K_a{}^D/K_a{}^H)^{1/3} \rightarrow \ell_4/(\phi_{BL})^{1/3}$ and the observed value of this limit is 0.66, so $\phi_{BL} = 1.14$.[19] In the limit as $X_{H_2O} \rightarrow 0$, $(K_a{}^D/K_a{}^H)^{1/3} \rightarrow (\phi_{AL}/\phi_{BL})^{1/3}$ and the observed value of this is 0.85; assuming ϕ_{BL} to have approximately the same values for the conjugate acids of indicators **1** and **3**[22] and to be independent of X_{L_2O}[23] gives $\phi_{AL} = 0.70$.[24] Of the remaining three fractionation factors (ℓ_1, ℓ_2, ℓ_3), the fit of the curve to the experimental points is quite sensitive to the value of ℓ_1, but very much less sensitive to the values of ℓ_2 and ℓ_3 so long as they fall in the expected order $\ell_1 > \ell_2 > \ell_3 > \ell_4 = 0.69$; for the curves shown, it was assumed that $\ell_1/\ell_2 = \ell_2/\ell_3 = \ell_3/\ell_4$. Those curves correspond to the value of ℓ_1 which gives the visually judged best fit ($\ell_1 = 0.79$, so that $\ell_2 = 0.76$ and $\ell_3 = 0.72$); other choices of ℓ_2 and ℓ_3 which are consistent with the expected order change the best fit value of ℓ_1 no more than ±0.02.

Conclusions

It is clear that almost certainly the value of ℓ_1 for L_3O^+ in CH_3CN is significantly less than 1.00,[26] and it appears likely that its value is near 0.79; the major uncertainty in ℓ_1 results from the uncertainty in the value of ϕ_{BL} for indicator **3** in CH_3CN. If it is true that the low value (0.69) of ℓ_4 in L_2O results from strong hydrogen bonding of three L_2O molecules to each L_3O^+ ion, then it is necessary that hydrogen bonding of L_3O^+ to CH_3CN molecules in liquid CH_3CN be almost as effective in weakening the force field experienced by the three fractionated L's in L_3O^+ as is hydrogen bonding to L_2O molecules in liquid L_2O. For comparison, it may be noted that, in the gas phase, $R_2O^+H\cdot\cdot NCCH_3$ hydrogen bonds are known[27] to be particularly strong.

This evidence that ℓ_1 is significantly less than 1.00 supports our interpretation[12,13] of the absence of an H_2O/D_2O kinetic isotope effect on methyl transfers to L_2O in liquid CH_3CN (eq 1). If the nucleophilic L_2O in the transition state had significant bonding to the CH_3 group, then the observed k_{H_2O}/k_{D_2O} should have been experimentally distinguishable from 1.00.[28] The H_2O/D_2O *solvent* isotope effect (from changing *both* solvent and nucleophile from H_2O to D_2O), which ranges from ca. 1.1 to 1.3 for various CH_3X, thus apparently reflects the isotope effect on the rotational correlation time of the liquid L_2O and is consistent with an activation process which is predominantly a fluctuation in solvent polarization. Such an interpretation of the solvent

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(18) For solutions with $[H_2O] < 2$ M, the molarity and mole fraction of free H_2O were calculated from the known total amounts of CH_3CN, H_2O, $HClO_4$, and B by using the values of $K^H{}_i$'s from ref 17; BH^+ was assumed to be unhydrated (assuming BH^+ to be monohydrated would change the calculated values of $[H_2O]$ by less than 5%), and $HClO_4$ was assumed to be completely dissociated.[17] For solutions with $[H_2O] > 2$ M, "H^+" was assumed to exist entirely as $(H_2O)_4H^+$.

(19) Few ϕ values for conjugate acids of tertiary amines are available for comparison. This value of 1.14 falls within the range of reported values. Schowen[20] quotes an average of 0.97, while the only two acids of this type which appear in the compilation of $K_a{}^H/K_a{}^D$ values by Laughton and Robertson[21] are imidazolium and trimethylammonium ions for whch $\phi_{BL} = 1.19$ and 1.28, respectively.

(20) Schowen, R. L. *Prog. Phys. Org. Chem.* **1972**, *9*, 281.

(21) Laughton, P. M.; Robertson, R. E. In "Solute-Solvent Interactions"; Coetzee, J. F., Ritchie. C. D., Eds.; Marcel Dekker: New York, 1969; pp 407-412.

(22) The consistency of the $(K_a{}^D/K_a{}^H)^{1/3}$ values for indicators **1** and **2** and for indicators **2** and **3** in the two overlap regions justifies this assumption.

(23) Even if ϕ_{BL} were to fall to 1.0 or rise to 1.3, the resulting value of ℓ_1 would change only to 0.76 or 0.83, respectively (and ϕ_{AL} would change to 0.61 or 0.80); also since BL^+ should have much weaker interactions with the medium than does L_3O^+, it is expected that the value of ϕ_{BL} should be less sensitive to changes in the medium than is the value of ℓ.

(24) Such a low value of the fractionation factor for $RC{\equiv}NL^+$ is in agreement with expectation based on the known value[25] of 0.64 for the isoelectronic $RC{\equiv}CL$.

(25) Reference 11, p 70.

(26) In order to have $\ell_1 = 1.00$ and $0.6 < \phi_{AL} < 1.0$, the values of $(K_a{}^D/K_a{}^H)^{1/3}$ observed for indicator **3** would require $1.6 < \phi_{BL} < 1.9$ for indicator **3** in CH_3CN. Contrast this to $\phi_{BL} = 1.1$ for indicator **1** in L_2O.

(27) Bromilow, J.; Abboud, J. L. M.; Lebrialla, C. B.; Taft, R. W.; Scorrano, G.; Lucchini, V. *J. Am. Chem. Soc.* **1981**, *103*, 5448-5453.

(28) For example, if 0.79 is the correct value for ℓ_1, then an H_2O-CH_3 bond order in the transition state of 0.1 would have given $k_{H_2O}/k_{D_2O} \approx 0.79^{-2(0.1)} = 1.05$; in contrast, the reproducibility of the k_{H_2O}/k_{D_2O} measurements is ca. ±0.01.

isotope effect was first advocated by Robertson and co-workers about one and one-half decades ago,[15] and very recently has received support from theoretical calculations by van der Zwan and Hynes.[29]

Experimental Section

Materials. Aldrich spectrophotometric grade CH_3CN [$(H_2O] \approx 0.01$ M by Karl Fischer titration) was used without further purification in preparing solutions with high water concentrations ($[L_2O] > 1$ M). For use in preparing solutions with lower water concentrations, 100-mL samples of this CH_3CN were dried by two successive bulb-to-bulb distillations on a vacuum line from 12-g portions of freshly activated (>3 h at $T > 300$ °C and $P < 10^{-5}$ torr) 3-Å molecular sieve. Karl Fischer titration of this dried CH_3CN indicated $[H_2O] \ll 10^{-3}$ M. This low water content was confirmed by drying a sample of CH_3CN which contained 1 M D_2O (three successive distillations from portions of 3-Å sieve which had been wetted with D_2O before being activated); the agreement (see Figures 2 and 3) of K_a^H/K_a^D values measured by using this CH_3CN as solvent with those measured by using other dried samples of CH_3CN indicates $[L_2O] \leq 10^{-4}$ M in both solvents. The H_2O used was ordinary distilled water which had been passed through a Barnsted 0808 mixed-bed ion-exchange column. Aldrich D_2O (99.8 atom % D) was used without further purification. A stock solution of $DClO_4$ was prepared from 70% $HClO_4$ by repeatedly adding D_2O and distilling off the excess L_2O in vacuo; the residual H content was measured by ^1H NMR spectroscopy. A corresponding stock solution of $HClO_4$ was prepared by distilling the excess H_2O from 70% $HClO_4$ in vacuo; the compositions of the constant-boiling deuterio and protio acids whch remained in the pot were both very close to $LClO_4{\cdot}2L_2O$; these compositions were checked by titration. Kodak *N*,*N*-dimethyl-*p*-nitroaniline (**1**) was recrystallized from ethanol. *N*,*N*-Dimethyl-3,5-dinitro-4-toluidine (**2**) was prepared as described[30] and recrystallized from heptane: mp 90.5–92 °C. 4-Chloro-*N*,*N*-dimethyl-2,6-dinitroaniline (**3**) was prepared by the reaction of 1,4-dichloro-2,6-dinitrobenzene[31] with dimethylamine and was recrystallized from methanol: mp 110.5–111 °C.

p*K*$_a$ Measurements. The dissociation constants (K_a^H and K_a^D) for the conjugate acids of indicators **1**, **2**, and **3** were measured by UV spectrometry. Absorbance measurements of **1**, **2**, and **3** were made at 395, 430, and 430 nm, respectively; the conjugate acids of **1**, **2**, and **3** do not absorb at these wavelengths. Absorbances were measured for sets of four cells in the thermostated (25.0 ± 0.1 °C) carrousel of a Varian Cary 219 spectrophotometer. Two cells in each set contained CH_3CN, H_2O, and indicator; one of these two also contained $HClO_4$. The other two cells were set up analogously, D_2O and $DClO_4$ being substituted for H_2O and $HClO_4$; the total numbers of formula weights of CH_3CN, D_2O, $DClO_4$, and indicator in each cell in the deuterio pair were equal to those of CH_3CN, H_2O, $HClO_4$, and indicator in the protio pair. Within each pair, the value of the extinction coefficient of the indicator base at that L_2O concentration was calculated from the absorbance of the cell which lacked $LClO_4$, and this value was used, together with the absorbance of the acidic solution and the known total amounts of indicator and $LClO_4$, to evaluate the molarities of L^+, indicator base (B), and indicator conjugate acid (BL^+) needed for the evaluation of K_a^L via eq 3. Small amounts of H (<10 atom %) were present in the D solutions (mostly from the $DClO_4$ stock which was 91 atom % D). The apparent value of $(K_a^D/K_a^H)^{1/3}$ was corrected for the presence of that H by the approximate equation $(K_a^D/K_a^H)^{1/3} \approx [n - 1 + (K_a^D/K_A^H)^{1/3}\mathrm{app}]/n$, where n is the atom fraction of D.

Special precautions were taken in preparing samples with total $[L_2O] \leq 0.2$ M: glassware and cells were dried in a vacuum oven and allowed to cool in Ar; all transfers were by syringe through serum caps, and vapor spaces above solutions were maintained at 1 atm of Ar. Samples and intermediate dilutions were used immediately after preparation to avoid errors arising from a slow decrease in acidity (presumably caused by the hydrolysis of CH_3CN in these extremely acidic solutions).

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Registry No. CH_3CN, 75-05-8; H_3O^+, 13968-08-6; D_2, 7782-39-0; *N*,*N*-dimethyl-*p*-nitroaniline, 100-23-2; *N*,*N*-dimethyl-3,5-dinitro-4-toluidine, 91266-70-5; 4-chloro-*N*,*N*-dimethyl-2,6-dinitroaniline, 10156-68-0.

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On the Magnitude of Primary Isotope Effects for Proton Abstraction from Carbon

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Abstract: Primary deuterium isotope effects have been measured for the loss of an α proton from 4-(4-nitrophenoxy)-2-butanone. The isotope effect data showed that no significant change in transition-state structure occurred in the same region of oxyanion catalyst pK_a range that does exhibit a large change in β value. This confirms the previous assertion that the dramatic curvature in the Brønsted plot for proton abstraction by oxyanions from carbon is due to solvation. The effect upon k_H/k_D of varying the identity of the atom acting as the proton acceptor was examined. Catalysts having oxygen, nitrogen, carbon, and sulfur as the acceptor atom were included. An inverse correlation was found between the magnitude of the isotope effect for proton abstraction by X^- and the energy change upon conversion of HX to DX. With use of the Marcus equation as a basis, an equation was derived for predicting values of k_H/k_D which not only relies on the ΔpK_a value but also incorporates the zero-point energy difference of both the reactants and the products as measured by infrared stretching frequency shifts upon deuteration. This equation gives an improved correlation between calculated and experimental values of k_H/k_D taken from the literature.

We have previously demonstrated[1] that Brønsted plots for proton abstraction from carbon acids by oxyanions in aqueous solution exhibit pronounced curvature, such that highly basic oxyanions are nonselective ($\beta = 0.2$) whereas weakly basic oxyanions are very selective ($\beta = 0.8$). We have proposed that this curvature is caused by an effect due to solvation rather than by a Hammond postulate type of change in transition-state structure. Consistent with this argument is the fact that when thiol anions abstract protons from carbon they do not exhibit the Brønsted plot curvature or rate enhancement due to solvation that is found

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