Proton Transfer in Mixed Water–Organic Solvent Solutions: Correlation between Rate, Equilibrium Constant, and the Proton Free Energy of Transfer

Ehud Pines* and Graham R. Fleming*

Department of Chemistry and The James Franck Institute, The University of Chicago, 5735 S. Ellis Ave., Chicago, Illinois 60637 (Received: April 15, 1991; In Final Form: June 24, 1991)

We demonstrate that the proton dissociation reaction of excited 1-aminopyrene follows a simple structure–reactivity correlation in water–methanol, water–DMSO, water–dioxane, and water–acetonitrile binary mixtures. The role of the proton free energy of transfer from pure water to the binary mixtures, $\Delta G_t^{w,s}(H^+)$ in determining the rate of transfer is also demonstrated. We have extended our analysis to include neutral ROH type photoacids where the simultaneous creation of the anion during proton transfer affects the energetics of the dissociation reaction in the mixed solvents as compared to pure water.

I. Introduction

From the very beginning of human interest in natural sciences water has held a special place.[1-3] The intensity of water research in modern times can be best demonstrated by the fact that some 3600 publications directly dealing with water properties are listed in a bibliography covering 18 years, between 1957 and 1974, a period which witnessed a revived interest in water research.[4]

It is thus very intriguing to note that many aspects of the properties of pure water and of water solvation are not yet clearly understood.[5,6] If one tries the impossible task of summarizing the current view of water, it emerges that the idea that liquid water is made of well-defined icelike microstructures floating in a sea of broken hydrogen bonds has been abandoned in favor of a less well-defined three-dimensional structure of ever-changing hydrogen bonded networks which locally have a tendency toward tetrahedral geometry. Current views have been greatly influenced by molecular dynamics[7] and Monte Carlo[8,9] simulations of liquid water. The average water molecule hydrogen-bond connectivity found in these calculations is only 2.3–2.4 per water molecule at 10 °C.[7,8] This is far less then the connectivity of 4 found in hexagonal ice. In fact, this degree of connectivity is only slightly higher than that of a one-dimensional infinite chain of water molecules, i.e., 2. This implies that liquid water forms a more open structure than was previously thought. Keeping this in mind, we must still regard the icelike picture of liquid water as a possible zero-order approximation for liquid water structure at least near the melting point.

Although the exact nature of proton solvation in water is as controversial as the notion of "water structure" itself,[10] it is generally recognized that proton solvation and diffusion in pure water involves some kind of a cooperative participation of water molecules beyond the simple hydroxonium ion, H_3^+O. In view of the classic picture of water structure[11] and water solvation[12]

it was natural to assume that the proton is solvated by a four-water-molecule cluster of tetrahedral geometry.[13-15] In other words, the proton was assumed to primarily reside on a single water molecule which is further hydrogen bonded to three "first solvation shell" water molecules.[13] Recently, Robinson et al.[16] have suggested a similar proton solvation number (i.e., 4 ± 1) as a result of their kinetic analysis of proton-transfer rates in aqueous mixtures (see below). Traditionally, the proton–water cluster solvation model was most convincingly pursued by Eigen, who mainly added kinetic arguments in favor of a relatively stable $H_9^+O_4$ complex.[17] Although many of Eigen's kinetic arguments are probably based upon imprecise measurements,[18-20] the existence of the $H_9^+O_4$ entity in water is still considered to be quite well established.[16] Quantum mechanical calculations[21] and gas-phase distributions of H^+ hydrates[22,23] have lent support to the four-water-molecular proton hydration model but have not proven its exclusiveness.

Recently both Robinson et al.[16,26,27] and Huppert et al.[24,25] utilized direct time-resolved measurements of proton dissociation rates from naphthol type photoacids in various water–alcohol mixtures as a means by which the concentration of the proton acceptor entity in water could be modified and probed.[24,26] In a manner similar to Conway's interpretation of the proton conductivity in water–organic solvent mixtures,[28] both groups interpreted the decrease in the proton dissociation rates as a manifestation of a water structure breaking effect caused by the

(1) Horne, R. A. In *Water and Aqueous Solutions;* Horne, R. A., Ed.; Wiley-Interscience: New York, 1971; p 1. Kell, G. S. *Ibid.;* Chapter 9. Davis, C. M.; Jarzynski, J. *Ibid.;* Chapter 10.
(2) Eisenberg, D.; Kauzman, W. In *The Structure and Properties of Liquid Water;* Oxford University Press: New York, 1969.
(3) Franks, F. In *Water, a Comprehensive Treatise;* Franks, F., Ed.; Plenum Press: New York, London, 1973; Vol. 1, p 1.
(4) *Physical and Chemical Properties of Water, a Bibliography; 1957–1974;* Hawkins, D. T., Ed.; IFI/Plenum Data Company: New York, 1976.
(5) Kabisch, G.; Kalman, E. In *The Chemical Physics of Solvation;* Dogonadze, R. R., Kalman, E., Kornyshev, A. A., Ulstrup, J., Eds.; Elsevier: Amsterdam, 1986; Chapter 1. Brooker, M. H. *Ibid.;* Chapter 4.
(6) Lesk, A. M. An *Introduction to Physical Chemistry;* Prentice-Hall, Inc.: Englewood Cliffs, NJ, 1982; pp 497–502.
(7) Stillinger, F. H. *Science* **1980,** *209,* 451.
(8) Mezei, M.; Bereridge, D. L. *J. Chem. Phys.* **1981,** *74,* 622.
(9) Jorgensen, W. L. *Chem. Phys. Lett.* **1980,** *70,* 326.
(10) Erdey-Gruz, T.; Lengyel, S. In *Modern Aspects of Electrochemistry;* Bockris, J. O.'M., Conway, B. E., Eds.; Plenum Press: New York, 1977; No. 12, Chapter 1.
(11) Bernal, J. D.; Fowler, R. H. *J. Chem. Phys.* **1933,** *1,* 515.

(12) Frank, H. S.; Evans, M. W. *J. Chem. Phys.* **1945,** *13,* 507.
(13) Wicke, E.; Eigen, M.; Ackermann, T. *Z. Phys. Chem. (Munich)* **1954,** *1,* 340.
(14) Robinson, R. A.; Stokes, R. H. *Electrolyte Solutions,* 2nd ed.; Butterworths Scientific Publications: London, 1959; Chapter 1.
(15) Bell, R. P. *The Proton in Chemistry,* 2nd ed.; Cornell University Press: Ithaca, NY, 1972.
(16) Robinson, G. W.; Thistlethwaite, P. J.; Lee, J. *J. Phys. Chem.* **1986,** *90,* 4224.
(17) Eigen, M. *Angew. Chem., Int. Ed. Engl.* **1964,** *3,* 1.
(18) For revised smaller rates values of the recombination reaction: $H^+ + OH^- \rightarrow H_2O$, see discussion in ref 15, p 129, and ref 19.
(19) Phillips, G. R.; Eyring, E. M. *J. Phys. Chem.* **1986,** *90,* 316.
(20) For a drastic downward revision of Eigen's proton mobility values in ice, see: *Physics and Chemistry of Ice;* Walley, E., Jones, S. J., Gold, L. W., Eds.; Royal Society of Canada: Ottawa; 1973. Of special interest is the introductory lecture by Önsager, p 7.
(21) Newton, M. D. *J. Chem. Phys.* **1978,** *67,* 5535.
(22) Lau, Y. K.; Ikuta, S.; Kebarle, P. *J. Am. Chem. Soc.* **1982,** *104,* 1462.
(23) Searcy, J. Q.; Fenn, J. B. *J. Chem. Phys.* **1974,** *61,* 5282.
(24) Huppert, D.; Kolodney, E. *Chem. Phys.* **1981,** *63,* 401.
(25) Huppert, D.; Kolodney, E.; Gutman, M.; Nachliel, E. *J. Am. Chem. Soc.* **1982,** *104,* 6949.
(26) Moore, R. A.; Lee, J.; Robinson, G. W. *J. Phys. Chem.* **1985,** *89,* 3648.
(27) Krishnan, R.; Fillingim, T. G.; Lee, J.; Robinson, G. W. *J. Am. Chem. Soc.* **1990,** *112,* 1353.
(28) Conway, B. E. In *Modern Aspects of Electrochemistry;* Bockris, J. O'M., Conway, B. E., Eds.; Butterworths: London, 1964; No. 3, Chapter 2.

"dilution" of water with the organic solvent. Using slightly different arguments, the two groups concluded, separately, that the dilution of high-order water clusters ($n \geq 2$) is indeed the cause for the apparent decrease in the proton-transfer rate as water is diluted with the organic solvent. However, while Robinson et al. have concluded that the four-water-molecule cluster is mainly responsible for proton solvation in water[16,27] (the 4WM model), Huppert et al. have concluded that higher than four water clusters are important ($6 \leq n \leq 10$, the n-cluster model).[24,25]

The aim of the present work is to investigate the generality of these ideas. An obvious place to start is with the question of whether the rate of this type of proton dissociation reaction really correlates with some gradual structural change brought upon pure water by the addition of the organic solvent. For that reason we carried out the same type of experiments on a different class of acids, i.e., cationic acids of the type RNH^+_3. Such acids dissociate via a "net" proton-transfer reaction, while the dissociations of ROH type acids create an ion pair consisting of the solvated proton and a solvated large molecular anion. In fact, the two classes of acids are known to differ substantially in their thermodynamic behavior in water–organic solvent mixtures.[29,30] It was thus of interest to see if the dissociation of cationic acids conforms to any type of water cluster model.

II. Proton Dissociation from Excited Photoacids

Photoacids are large organic molecules that become more acidic upon electronic excitation.[31] Their thermodynamic behavior was first characterized by Förster using a thermodynamic cycle.[32] These photoacids are well described kinetically by a simple ground-state acid–base equilibrium.[31,33] In many cases the proton dissociation reaction is found to be completely reversible on the excited-state potential surface,[34,35] the dissociation and recombination taking place from a vibronically relaxed electronic excited state.[35,36] While the actual bond breaking and remaking at the reaction site is considered to be fully adiabatic,[37] in some cases the proton attachment to the aromatic residue of the photoacid results in quenching of the excited state.[38,39]

The protonated 1-aminopyrene molecule (1-amp) and other protonated aminopyrene derivatives were among the first photoacids investigated by Förster[32] and subsequently by Weller.[40-42] More recently, Shizuka et al.[43] investigated the 1-amp molecule by time-resolved techniques. In this case one starts with a charged molecule and ends up with the charge relocated within the solvent. One should emphasize that from a purely kinetic point of view dissociation of protonated nitrogen acids and unprotonated oxygen acids that form ion pairs have always been considered examples of the same type of proton-transfer reactions.[15,42] It is for that reason that a comparison between the two types of photoacids should be suitable for a critical evaluation of the role of the proton solvation in determining the dissociation rates of these acids.

Figure 1. (a) Proton dissociation lifetimes of protonated 1-amp in various water–organic solvent binary mixtures as a function of mass fraction of the organic solvent. From top to bottom: ×, water–methanol; +, water–CH_3CN; O, water–dioxane; *, water–DMSO solutions. (b) Same as in Figure 1a but on the mole fraction scale of the organic solvents. The y axis is expanded relative to that of Figure 1a.

Furthermore, one could check the basic assumption made by Robinson et al.[16] and Huppert et al.[24] that, in the phenol type photoacids, dissociation merely reflects the availability of the proton acceptor within the solution, without consideration of the simultaneous creation of a counterion.

III. Experimental Section

1-Aminopyrene (1-amp, Aldrich, 97%) was doubly recrystallized from cyclohexane, and its steady-state excitation and fluorescence spectra were identical with the spectra published in the literature.[43,44] All organic solvents were used without further purification (acetonitrile was Baker, HPLC reagent with 5×10^{-3} mass % of water (2×10^{-3} M), dioxane was Baker AR, methanol was Baker HPLC reagent with typically less than 10^{-3} mass % of water, and DMSO was either Mallinckrodt A.C.S. reagent or Aldrich spectrophotometric grade). All solvents were used from freshly opened bottles. $HClO_4$ was 60% concentrated, and concentrations of typically 2×10^{-3}–2×10^{-2} M of the acid were used in order to protonate the 1-amp in the ground state. Steady-state fluorescence spectra were measured with a Perkin-Elmer MPF-6 spectrometer.

The time-resolved fluorescence intensities were measured with a time-correlated single-photon-counting system described in detail elsewhere.[45] The instrument response function was 55 ps full

(29) There is an enormous amount of data about the pK_a values of various acids in aqueous binary mixtures. As representative examples see ref 14, pp 541–542.

(30) Bates, R. G.; Robinson, R. A. In *Chemical Physics of Ionic Solutions*; Conway, B. E., Barradas, R. G., Eds.; Wiley: New York, 1966; Chapter 12, p 211.

(31) Ireland, J. F.; Wyatt, P. A. H. *Adv. Phys. Org. Chem.* **1976**, *12*, 131. Schulman, S. G. In *Fluorescence and Phosphorescence Spectroscopy*; Pergamon Press: London, 1977.

(32) Förster, Th. *Z. Naturwissen.* **1949**, *36*, 186; *Z. Electrochem.* **1950**, *54*, 42.

(33) Weller, A. *Prog. Reac. Kinet.* **1961**, *1*, 189.

(34) Schulman, S. G.; Rosenberg, L. S.; Vincent, W. R. Jr. *J. Am. Chem. Soc.* **1979**, *101*, 139.

(35) Pines, E.; Huppert, D.; Agmon, N. *J. Chem. Phys.* **1988**, *88*, 5620.

(36) Demjachkewitch, A. B.; Zaitsev, N. K.; Kuzmin, M. G. *Chem. Phys. Lett.* **1978**, *55*, 80.

(37) Förster, Th. *Pure Appl. Chem.* **1970**, *24*, 443.

(38) (a) For an early discussion see: Förster, Th. *Chem. Phys. Lett.* **1972**, *17*, 309. (b) For a recent view: Shizuka, H. *Acc. Chem. Res.* **1985**, *18*, 141.

(39) Webb, S. P.; Philips, L. A.; Yeh, S. W.; Tolbert, L. M.; Clark, J. H. *J. Phys. Chem.* **1986**, *90*, 5154.

(40) Urban, W.; Weller, A. *Ber. Bunsen.-Ges. Phys. Chem.* **1963**, *67*, 787.

(41) Grunwald, E. *Prog. Phys. Org. Chem.* **1965**, *3*, 317.

(42) Albery, W. J. *Annu. Rev. Phys. Chem.* **1980**, *31*, 227.

(43) Shizuka, H.; Tsutsumi, K.; Takeuchi, H.; Tanaka, I. *Chem. Phys.* **1981**, *59*, 183; *Chem. Phys. Lett.* **1979**, *62*, 408.
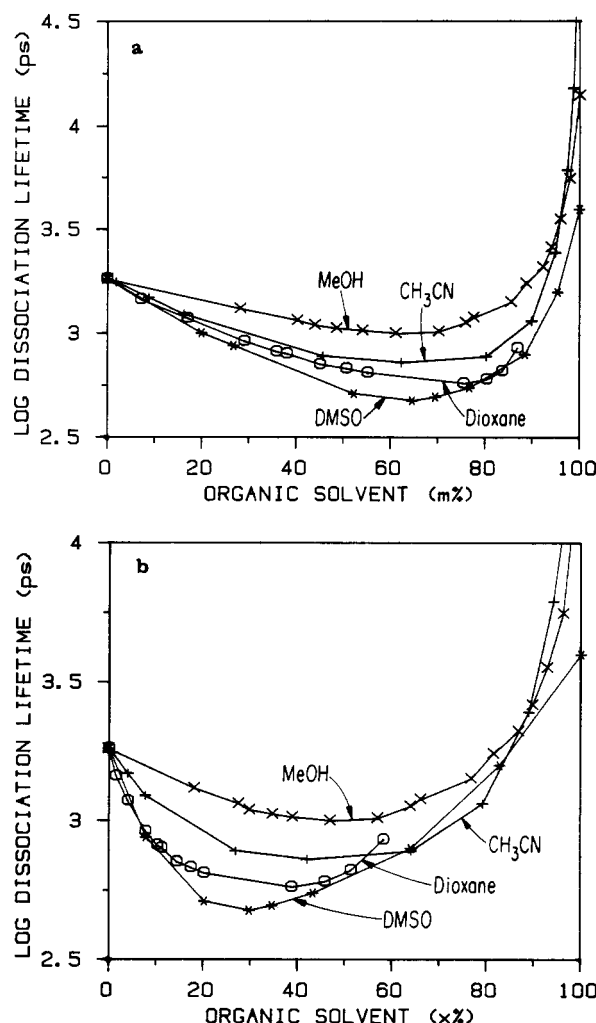
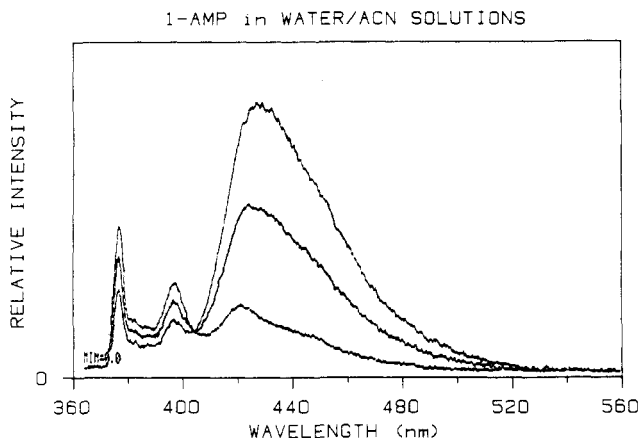(44) Sadtler Research Labs., Inc. 1964. Jones, R. N. *J. Am. Chem. Soc.* **1945**, *67*, 2127.

Figure 2. Fluorescence spectra of protonated 1-amp in CH_3CN–water binary mixtures. The fluorescence band at 374 nm corresponds to that of the protonated species (BH^+), and the band at 425 nm of the unprotonated molecule (B). The isoemissive point is at 405 nm. From top to bottom at the 425-nm band: c = 0.27, 0.54, 0.81 M of H_2O.

width at half-maximum, and the typical counting rate was 1 kHz. The resolution was varied between ca. 5 and ca. 60 ps/channel. Varying the excitation wavelength between 297 and 310 nm did not change the measured lifetime of the protonated 1-amp. The fluorescence data were collected at either 374 nm for the protonated 1-amp detection or 435 ± 10 nm for the unprotonated form detection. The overall concentration of the 1-amp was typically 3 × 10^{-4} M. No change in the lifetime was detected when the 1-amp concentration was varied between 10^{-4} and 2 × 10^{-3} M. In most cases the fluorescence spectra were collected when some unprotonated 1-amp was present in the ground state. In this case the overlap of the fluorescence from the unprotonated and protonated forms results in a biexponential fluorescence response at 374 nm (see below). Even in the extreme case of a total 1-amp concentration of 2 × 10^{-3} M and a protonated to unprotonated ratio of 1:10. The presence of the ground-state unprotonated 1-amp did not affect the lifetime of the protonated form. All experimental data were collected at 22 ± 1 °C and analyzed by an iterative fit to an exponential decay. A biexponential response was used to fit overlapped spectra. A fit was then judged successful only when it reproduced the lifetime of the unprotonated form within ±5%. This lifetime was measured independently prior to acidification when only the unprotonated form is present. The samples were saturated with N_2 prior to the measurement.

IV. Results

Figure 1a shows the lifetime of the protonated 1-amp as a function of the mass fraction (m) of CH_3OH, CH_3CN, dioxane, and DMSO cosolvents. In contrast to the previous measurements on ROH type photoacids,[16,24-27] the lifetime decreased and the proton dissociation rates increased in all water–organic solvent mixtures up to 60–75% (m = mass) of the organic solvent. We will refer to this region as the water rich region. Further increase in the cosolvent concentration results in a sharp decrease in the proton-transfer rate. We will call this region the organic-solvent-rich region. The organic-solvent-rich region could not be reached in dioxane due to insufficient protonation of 1-amp in the ground state. The behavior in the water-rich region is only slightly dependent on the cosolvent, while in the organic-solvent-rich mixtures the proton dissociation rates differ substantially from one cosolvent to the other. In Figure 1b the same data are shown as a function of the mole fraction (x) of the organic cosolvent. The minimum lifetimes are found to lie between 30% and 50% mole fraction of the cosolvents.

Figure 2 shows typical steady-state fluorescence spectra of 1-amp in CH_3CN with various small concentrations of H_2O. An
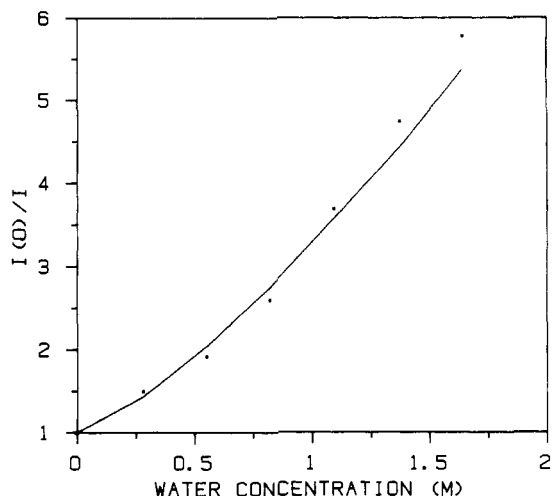

Figure 3. Stern–Volmer plot of the 374-nm band of protonated 1-amp in CH_3CN–H_2O solutions. The dots are the steady-state fluorescence data, and the solid line is a reconstructed Stern–Volmer plot from the fluorescence lifetime data.

isoemissive point is clearly evident at 405 nm. Figure 3 shows a Stern–Volmer analysis of such 1-amp fluorescence spectra in acetonitrile together with the Stern–Volmer curve as predicted from the directly measured proton-transfer rates extracted from the data in Figure 1.

V. Discussion

(1) Structure–Reactivity Correlations. Acid–base kinetics involving proton transfer between a donor (acid, A) and an acceptor (base, B) are traditionally analyzed and catalogued by the Brønsted relation,[15,46]

$$\log k_p = \log \alpha + \beta \Delta p K_{AB} \tag{1}$$

where k_p is the proton-transfer rate, α and β are constants for a series of similar reactions, and $\Delta p K_{AB}$ is the difference in the acidity strength (equilibrium constant) between the proton donor and proton acceptor.

If the donor and acceptor acidities are both defined by the same acidity scale, it follows that

$$\Delta p K_{AB} = p K_B - p K_A = \log K \tag{2}$$

where $p K_B$ and $p K_A$ are the acidity equilibrium constants of the base and the acid and K is the equilibrium constant of the direct proton transfer between them, $AH + B \leftrightarrow A + BH$:

$$K = [BH][A]/[AH][B] \tag{3}$$

The Brønsted relation thus generally implies that there exists a linear relation between the rate of a proton-transfer process and the equilibrium constant for this process. It follows that one should expect a linear dependence between the dissociation rate of an acid and its equilibrium constant. Unfortunately in many cases this statement becomes trivial for proton-transfer reactions in aqueous solutions for the following reason: For any reaction K = k_{off}/k_{on}, where k_{on} and k_{off} are the overall kinetic recombination and dissociation rate constants. In water the overall proton recombination constant of weak and moderately strong acids is diffusion controlled. This makes k_{on} practically a constant for most acid–base reactions in water. It then follows that in this case
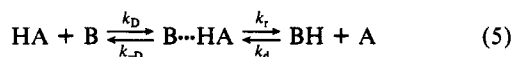
$$k_{off} = \alpha K \tag{4}$$

where α is a constant given by the diffusion controlled recombination rate. Equation 4 has the form of a Brønsted relation (eq 1) with β being unity. This purely kinetic controlled phenomenon is some time considered as a structure–reactivity correlation.[47] A similar problem arises with bimolecular recom-

(45) Chang, M. C.; Courtney, S. H.; Cross, A. J.; Gulotty, R. J.; Petrich, J. W.; Fleming, G. R. *Anal. Instrum.* **1985**, *14*, 433.

(46) Brönsted, J. N.; Pedersen, K. J. *Z. Phys. Chem.* **1924**, *108*, 185.
(47) Gutman, M. *Methods Biochem. Anal.* **1984**, *30*, 1.

bination rates which approach the diffusion-controlled limit.[17,42] It follows that one should verify that diffusion does not mask the intrinsic reaction rate. Following the definition of a general acid–base reaction due to Eigen

$$HA + B \underset{k_{-D}}{\overset{k_D}{\rightleftharpoons}} B\cdots HA \underset{k_d}{\overset{k_r}{\rightleftharpoons}} BH + A \qquad (5)$$

one has for the equilibrium and the kinetic constants

$$k = \frac{k_{off}}{k_{on}} \qquad k_{on} = \frac{k_D k_r}{k_r + k_{-D}} \qquad k_{off} = \frac{k_d k_{-D}}{k_r + k_{-D}} \qquad (5a)$$

where k_r and k_d are the chemical rate constants for recombination from and dissociation to an intermediate ion-pair state, k_D and k_{-D} are the diffusion-controlled rate constants from the bulk solution to the intermediate ion-pair state and vice versa. From eq 5a it follows that only when $k_{-D} \gg k_r$ does $k_{off} = k_d$. In other words, essentially irreversible dissociation rates can be directly used in a structure–reactivity relation. When this is not the case the overall dissociation rate must be first broken down into its chemical and diffusional contributions.[48]

A more general view of the structure–reactivity law is as a correlation between the total free energy of a reaction and its activation free energy. Realizing that barrier crossing is a multidimensional process, these approaches seek the most probable route for a reaction to take place. They designate this route as the reaction coordinate and approximate the overall process as a one-dimensional barrier crossing along this reaction coordinate.[49,50] For example, the well-known approach of Marcus[51] is very useful in electron-transfer reactions.[52] In the case of proton transfer, the proton is pictured as being transferred along a linear hydrogen bond A–H···B, which exists between the donor and the acceptor.[53] The position of the transition state along the A–H···B complex is determined by the exothermicity of the reaction, i.e., nearer to the acceptor when the reaction is endothermic and nearer to the donor when it is exothermic. These theories were recently reviewed by Yates.[54] Although structure–reactivity relations in proton-transfer reactions were the first to be realized empirically in chemical kinetics by Brønsted and by Hammett,[55,56] the utilization of more sophisticated theoretical models, based on the Marcus approach, to proton-transfer reactions have met with considerable difficulties. While slow proton-transfer reactions and proton-catalyzed reactions are usually well characterized by simple Brønsted type structure–reactivity correlations, where correlations spanning well over 20 pK units are found[15,57] moderately fast proton-transfer reactions can be analyzed by the basic Marcus theory.[15,42,58] However, exothermic proton-transfer reactions which are almost barrierless are poorly described by both the Brønsted relation and the basic Marcus theory which predicts an "inverted region"[59] (i.e., a region where the reaction slows down as it becomes more exothermic).[52] An effort by Formosinho to analyze photoinitiated proton-transfer reactions by the ISM model[60] resulted in only rough correlations between proton dissociation rates and equilibria in aqueous solutions.[61]
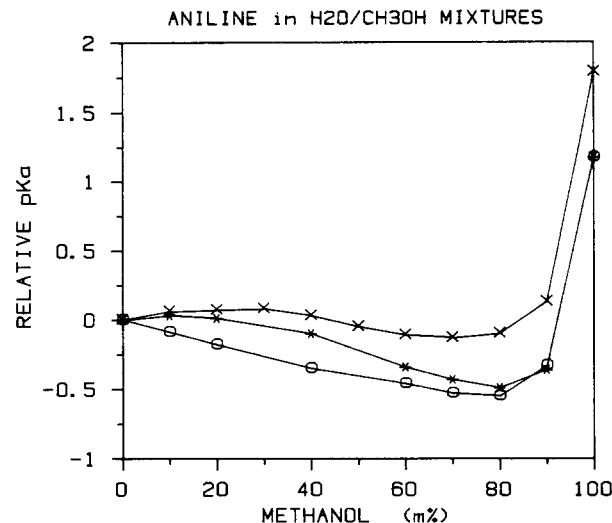
As further described below, we believe that the somewhat

Figure 4. Acid–base equilibrium constant of protonated aniline in H_2O/CH_3OH solutions as a function of the mass fraction of CH_3OH: ($\times$) pK_a values calculated from eq 15 without the Born energy; ($*$) same but with the Born energy; (O) measured pK_a values.[78]

blurred nature of fast proton transfer from a photoacid can now be systematically resolved and catalogued alongside the classic ground-state acid–base processes.

(2) Excited 1-amp and Hammett Indicators. Anilinium derivative ions are probably the most common Hammett indicators used to measure the acidity function of a nonaqueous solvent:[15,55,56]

$$H_0 = pK_{BH^+} - \log (C_{BH^+}/C_B) \qquad (6)$$

where H_0 is the Hammett acidity function and C_{BH^+} and C_B are the directly observable concentration of the acid and base form of the Hammett indicator. For that reason pK_{BH^+} values of anilinium ions in various pure organic solvents and water mixtures have been measured with great precision.[62] On going from pure water to pure organic solvent composition all primary aromatic amines show remarkably similar behavior in their acid–base properties.[15,63-66] The value of $\Delta pK^{s,w} = pK^s - pK^w$ is found to be almost a constant for all members of the same family of anilinium ions, where pK^s and pK^w are the pK_{BH^+} values in the solvent s and in pure water, respectively. If one assumes that on going from the ground state to the excited state $\Delta pK^{s,w}$ does not change significantly, one can apply directly the readily available $\Delta pK^{s,w}$ values of the anilinium ion to the acid–base equilibria of the related excited 1-amp molecule.

(3) Proton Reaction Coordinate and Choice of a Suitable Structure–Reactivity Model. The free energy of transfer of single ions (i) from pure water to any other solvent has been evaluated quite accurately.[67-69] Using recently recommended values for the transfer of H^+ from pure water to water–methanol mixtures,[67b] one can demonstrate the correlation between the free energy of transfer of the proton, $\Delta G_t^{w,s}(H^+)$ and the pK_a = pK_{BH^+} values of the anilinium ion in water–methanol solutions (Figure 4). One of the main features in Figure 4 is the shallow minimum at 70–80% by mass (m) of methanol and a sharp increase in the pK_a^s values as the pure CH_3OH composition is approached.

The change in the acidity constant, pK_a, of an acid resulting from a change in the solvent composition is given by[62]

$$\ln (10) RT \Delta pK_a^{s,w} = \sum_{p=1}^{n} \Delta G_t^{w,s}(P) - \sum_{r=1}^{n} \Delta G_t^{w,s}(R) \qquad (7)$$

(48) Agmon, N.; Pines, E.; Huppert, D. *J. Chem. Phys.* **1988**, *88*, 5631.

(49) Agmon, N. *Int. J. Chem. Kinet.* **1981**, *13*, 333.

(50) Grunwald, E. *Prog. Phys. Org. Chem.* Taft, R. N., Ed.; Wiley: New York, 1990; Vol. 47, p 55.

(51) Marcus, R. A. *J. Chem. Phys.* **1956**, *24*, 966.

(52) Closs, G. L.; Miller, J. R. *Science* **1988**, *240*, 440.

(53) More O'Ferrall, R. A. In *Proton Transfer Reactions*; Caldin, E. F., Gold, V., Eds.; Chapman and Hall: London, 1975; p 201.

(54) Yates, K. J. *J. Phys. Org. Chem.* **1989**, *2*, 300.

(55) Hammett, L. P. *J. Am. Chem. Soc.* **1937**, *59*, 96.

(56) Shorter, J. In *Prog. Phys. Org. Chem.* **1990**, *17*, 1.

(57) Bell, R. P.; Hillier, G. R.; Mansfield, J. W.; Street, D. G. *J. Chem. Soc. B* **1967**, 827. Bell, R. P.; Higginson, W. C. E. *Proc. R. Soc. London, A* **1979**, *197*, 141.

(58) Sühnell, J. *J. Phys. Org. Chem.* **1990**, *3*, 62.

(59) Marcus, R. A. *Faraday Discuss. Chem. Soc.* **1982**, *74*, 7. Marcus, R. A. *J. Phys. Chem.* **1989**, *93*, 3078.

(60) Varandas, A. J. C.; Formosinho, S. J. *J. Chem. Soc., Faraday Trans. 2* **1986**, *82*, 953. Arnaut, L. G.; Formosinho, S. J. *J. Phys. Org. Chem.* **1990**, *3*, 95 and references therein.

(61) Arnaut, L. G.; Formosinho, S. J. *J. Phys. Chem.* **1988**, *82*, 685.

(62) Bates, R. G. *J. Electroanal. Chem.* **1971**, *29*, 1.

(63) Paul, M. A.; Long, F. A. *Chem. Rev.* **1957**, *57*, 1.

(64) Sager, E. E.; Robinson, R. A.; Bates, R. G. *J. Res. Nat. Bur. Stand.* **1964**, *68A*, 305.

(65) De Ligny, C. L.; Loriaux, H.; Ruiter, A. *Recueil* **1961**, *80*, 725.

(66) Coetzee, J. F.; Padmanabhan, G. R. *J. Am. Chem. Soc.* **1965**, *87*, 5005.

(67) (a) Marcus, Y. In *Ion Solvation*; Wiley: New York, 1985. (b) Marcus, Y. *Pure Appl. Chem.* **1990**, *5*, 899.

(68) Popovych, O. *Anal. Chem.* **1974**, *46*, 2009.

(69) Strehlow, H. *Z. Phys. Chem. (Munich)* **1986**, *148*, 1.

where the summation is carried over all products (P) and reactants (R) involved in the dissociation reaction, $\Delta G_t^{w,s}(P,R)$ is the free energy change in transferring the species from water (w) to the solvent (s). For cationic acid dissociations which are formally described by the Hammet equilibrium

$$BH^+ \rightleftharpoons B + H^+ \tag{8}$$

one arrives at the following expression for the total free energy change in transferring the reaction from water to s:

$$\sum_{r,p}^{n}\Delta G_t^{w,s}(P,R) \equiv \Delta G_t^{w,s} =$$

$$\Delta G_t^{w,s}(H^+) + \Delta G_t^{w,s}(B) - \Delta G_t^{w,s}(BH^+) \tag{9}$$

where $\Delta G_t^{w,s}(B)$ and $\Delta G_t^{w,s}(BH^+)$ are the free energy of transfer of the unprotonated and protonated form of the Hammet indicator. Following Popovych,[68] it is customary to break up the $\Delta G_t^{w,s}(i)$ of the ions into purely electrostatic (el) and nonelectrostatic (nel) parts:

$$\Delta G_t^{w,s}(i) = \Delta G_{t,el}^{w,s} + \Delta G_{t,nel}^{w,s} \tag{10}$$

The additional assumption that

$$\Delta G_{t,nel}^{w,s}(BH^+) = \Delta G_{t,nel}^{w,s}(B) \tag{11}$$

reduces eq 9 to

$$\Delta G_t^{w,s} = \Delta G_t^{w,s}(H^+) - \Delta G_{t,el}^{w,s}(BH^+) \tag{12}$$

The simplest approximation to $\Delta G_{t,el}^{w,s}(BH^+)$ is the continuum solvation model of Born:[70]

$$\Delta G_{t,el}^{w,s} = \frac{Z_i^2}{2}\left(\frac{1}{\epsilon_w} - \frac{1}{\epsilon_s}\right)\frac{1}{r_i} \tag{13}$$

where $\Delta G_{t,el}^{w,s}(i)$ is the free energy change due to the change in the static dielectric constant of the solution, Z_i and r_i are the charge and the radius of the ion i, and ϵ_w and ϵ_s are the static dielectric constants of pure water and the organic solvent. Combining eqs 13 and 7, one gets[64] when ΔpK is on the molar scale, r_i in angstrom units, and at 22 °C

$$\Delta pK_{t,el}^{s,w} = 122.8\left(\frac{1}{\epsilon_s} - \frac{1}{\epsilon_w}\right)\frac{1}{r_i} \tag{14}$$

where $\Delta pK_{t,el}^{s,w}$ is the purely electrostatic contribution to the pK_a change. For anilinium type acids eq 12 can now be rewritten as

$$\Delta pK_a^{s,w} = \Delta pK_t^{s,w}(H^+) - \Delta pK_{t,el}^{s,w}(BH^+) \tag{15}$$

Equation 15 implies, as Figure 4 indeed shows, that for cationic acids the effect of the Born electrostatic free energy is to lower the pK_a values of the acids, as compared to what was predicted from the $\Delta pK_t^{s,w}(H^+)$ values in solvents having a lower static dielectric constant than pure water. Although the prediction of eq 15 should be regarded as only qualitative,[68] it correctly reproduces the trend found in Figure 4. The fit of the data points at 100% methanol gives a reasonable figure for the radius of the anilinium ion $r_{BH^+} = 3.7$ Å.[71] Keeping this value of r_{BH^+} a constant reproduces the $\Delta pK_a^{s,w}$ values to within 0.1 pK units down to 60% by mass of methanol. Below this composition range the data are reproduced to within 0.2 pK units. Although not totally satisfying, the above deviation is within the upper limit of the total experimental error in evaluating $\Delta pK_a^{s,w}$ and $\Delta pK_t^{s,w}(H^+)$ values.

A successful application of eqs 12 and 13 suggests that specific solvation of the aromatic residue which is distant from the reaction site does not play an important role in determining $\Delta pK_a^{s,w}$. It further implies that the solvation change brought upon the precursor molecule at the reaction site as a result of the proton dissociation is relatively insensitive to the solution composition. This points to a preferential water solvation at the reaction site,

so that the composition of the solution could change considerably while the solvation of the amine group remains largely the same. We thus suggest that the composition sensitive part of the precursor solvation is largely due to the change in the continuum dielectric response of the medium, i.e., the Born charging free energy. Recent studies have shown that the time scale of polar solvation at room temperature is much faster than the characteristic time scale of a bimolecular proton-transfer reaction.[72] This conclusion is in accord with that made by Robinson et al.[16] in the case of ROH acid dissociations. Summarizing, we conclude that the dynamics of the overall dissociation process of BH$^+$ type acids is largely determined by the proton solvation dynamics along the hydrogen bond which connects the proton donor and the proton acceptor. We further suggest that eq 15 could serve as a general test for the applicability of a structure–reactivity correlation to a series of proton dissociation reactions in binary solvent mixtures for any type of acids. This suggestion is further discussed in subsection 7.

We proceed by selecting a structure–reactivity correlation model first suggested by Agmon and Levine[73-75] and later successfully utilized by Agmon[75] to analyze several proton-transfer reaction series. Although not as widely used as the Marcus theory, the Agmon–Levine (AL) model is capable of faithfully reproducing the behavior of a number of proton-transfer reaction series in the ground state. The model also had some encouraging initial results in predicting the behavior of ROH type acids in the excited state.[73] The AL model starts from the Pauling definition of the bond order:[76]

$$n \equiv \exp(-r/a) \tag{16}$$

where $r = R - R_e$ is the displacement of the bond length, R, from its equilibrium value, and R_e and a are empirical parameters. One also assumes that the breaking and forming bonds are coupled to each other on the reaction coordinate, so $n_1 + n_2 = 1$. This is the basic assumption for any empirical bond energy bond-order (BEBO) model.[77] The location of the barrier along the bond-order reaction coordinate was found by AL using the entropy function and is given by

$$n^* = [1 + \exp(-\Delta G \ln (2)/G_a^\circ)]^{-1} \tag{17}$$

where n^* is the location of the barrier in the normalized reaction coordinate, ΔG is the overall free energy change in the reaction, and G_a° is the free energy barrier which exists when $\Delta G = 0$. In any case other than $\Delta G = 0$, $G_a \neq G_a^\circ$ and is given by

$$G_a = \Delta G - G_a^\circ \ln (n^*)/\ln 2 \tag{18}$$

where G_a is the free energy of activation of the reaction. Evidently when $\Delta G \rightarrow \infty$ (very endothermic reactions) $n^* \rightarrow 1$ and $G_a \rightarrow \Delta G$, while for the very exothermic reactions ($\Delta G \rightarrow -\infty$) both n^* and G_a approach zero. Finally, the rates of a series of reactions which belong to the same family are related by

$$k = k^\circ \exp[-G_a/RT] \tag{19}$$

where k° is the maximal reaction rate which is encountered in a given reaction series when $G_a = 0$. Equation 19 can be directly utilized to analyze proton dissociation reactions. Thus, k_d can now be written as

$$k_d = k_d^\circ \exp[-G_a/RT] \tag{20}$$

ΔG in these reactions is given by the pK_a:

$$\Delta G = RT \ln (10)pK_a \tag{21}$$

To fit a given reaction series according to eqs 18 and 20, one needs to find two (adjustable) parameters, namely, k_d° and G_a°. However, the lower bound for k_d° can be independently estimated from the most exothermic reaction in the reaction series. Likewise,

(70) Born, M. *Z. Phys.* **1920**, *1*, 45.
(71) Meot-Ner, M. *J. Phys. Chem.* **1987**, *91*, 417.

(72) Bagchi, B. *Annu. Rev. Phys. Chem.* **1989**, *40*, 115 and references therein. Barbara, P. F.; Jarzeba, W. *Adv. Photochem.* **1990**, *15*, 1.
(73) Agmon, N., unpublished results.
(74) Agmon, N.; Levine, R. D. *Chem. Phys. Lett.* **1977**, *52*, 197.
(75) Agmon, N.; Levine, R. D. *Isr. J. Chem.* **1980**, *19*, 330.
(76) Pauling, L. *J. Am. Chem. Soc.* **1947**, *69*, 542.
(77) Johnston, H. S. *Adv. Chem. Phys.* **1960**, *3*, 131. Johnston, H. S.; Parr, C. A. *J. Am. Chem. Soc.* **1963**, *85*, 2544.
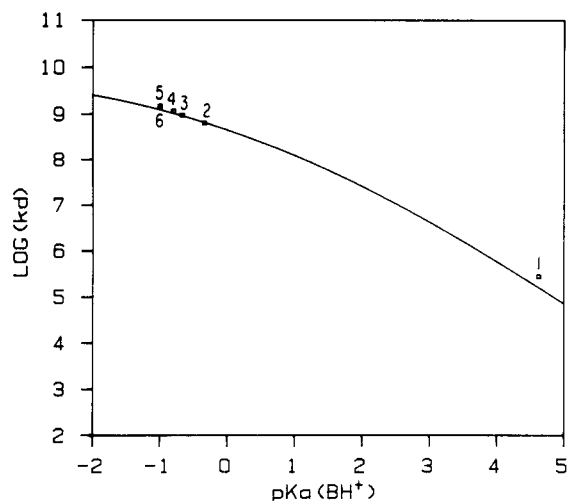
Figure 5. Correlation between proton dissociation rate and pK_a values for several aromatic amines: (1) ground-state aniline; (2) excited 1-amp in pure water; (3, 4) excited 1-amp in 20% and 50% (vol) of CH_3CN (corrected as described in the text); (5, 6) excited 1-naphthylamine and 2-naphthylamine in pure water.[38b]

for a given reaction, G_a and hence $G_a{}^\circ$ can be estimated from temperature-dependent experiments. One final note is that eqs 17–20 do not predict an inverted region but rather an asymptotic approach to the maximum activationless rate attainable in a reaction series. This behavior is in accord with the modified Marcus equation for proton-transfer reactions. Although the Marcus expression is of a very different form, the two models predict a similar general dependence of rates on equilibria.[54,59]

(4) Analysis of Proton Dissociation Rates from Protonated Excited Aromatic Amines by the Agmon–Levine Structure–Reactivity Model. Figure 5 shows the correlation between $pK_a{}^*$ values of excited aromatic amines and their respective proton dissociation rates. The analysis was carried out according to the AL model with $k_d{}^\circ = 9.3 \times 10^9$ s^{-1} and $G_a{}^\circ = 1.84$ kcal/mol for the maximum proton dissociation rate achievable in the amine series and the free energy activation barrier when $pK_a = 0$, respectively. The kinetic data in Figure 5 were collected from the literature. $pK_a{}^*$ values of 1-amp were taken from Shizuka et al.[43] and were corrected according to our own time-resolved measurements taken at the same acetonitrile–water compositions. The $pK_a{}^*$ value of 1-amp in pure water was found by combining our own value of $k_d = 5.5 \times 10^8$ s^{-1} with k_r values extracted from Shizuka et al.[43] at 20% and 50% acetonitrile compositions and extrapolated to pure water $k_r \approx 2.5 \times 10^8$ M^{-1} s^{-1}. The $pK_a{}^*$ of 1-amp was thus determined to be $pK_a{}^* = -\log (k_d/k_r) = -0.34$. This is the pivotal value for the determination of the absolute $pK_a{}^*$ values of 1-amp in the various water–organic solvent compositions. These values were extracted according to the following relation:

$$pK_a{}^s(1\text{-amp}) = -0.34 + \Delta pK_a{}^{w,s}(\text{anilinium}) \qquad (22)$$

where $pK_a{}^s$ is the acid equilibrium constant in the solvent s and $\Delta pK_a{}^{w,s}$ is the change in the pK_a of the anilinium ion on going from pure water to the solvent s. It is reassuring to find that deprotonation rate of the ground-state anilinium in pure water[17] falls on the same structure–reactivity curve shown in Figure 5 although it involves a nonlinear extrapolation of the correlation by some 5 pK units. In Figure 6 an expanded portion of the same structure–reactivity curve is shown with additional data points for 1-amp (this work) in several water–organic solvent mixtures and correlated by their $pK_a{}^s$ values as found from eq 22.

The most extensive thermodynamic $pK_a{}^s$ data available for the anilinium ion were in H_2O–methanol mixtures.[78] Other additional points in Figure 6 are for 70% (m) DMSO[79] and water–dioxane mixtures.[80] In the case of the $pK_a{}^s$(anilinium) in water–methanol

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

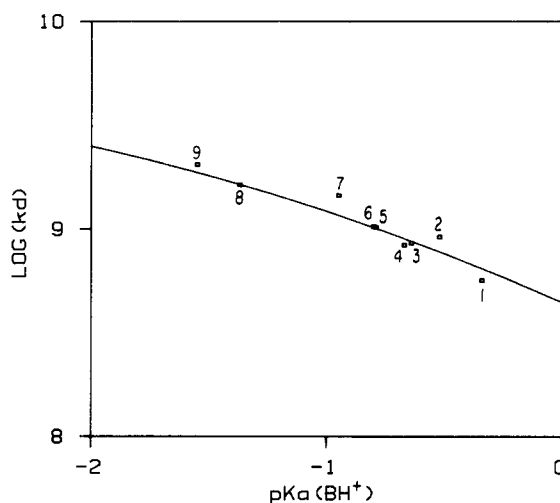(79) Yates, K.; Welch, G. *Can. J. Chem.* **1972**, *50*, 474.

Figure 6. Correlation between proton dissociation rate and pK_a of 1-amp in various water–organic solvent mixtures. The pK_a values were calculated according to the procedure described in the text: (1) pure H_2O; (2) 20% (vol) CH_3OH; (3) 40% (vol) CH_3OH; (4) 20% (vol) CH_3CN; (5) 60% (vol) CH_3OH; (6) 2-naphthylamine in pure water; (7) 45%(m) dioxane; (8) 70%(m) dioxane; (9) 70%(m) DMSO.
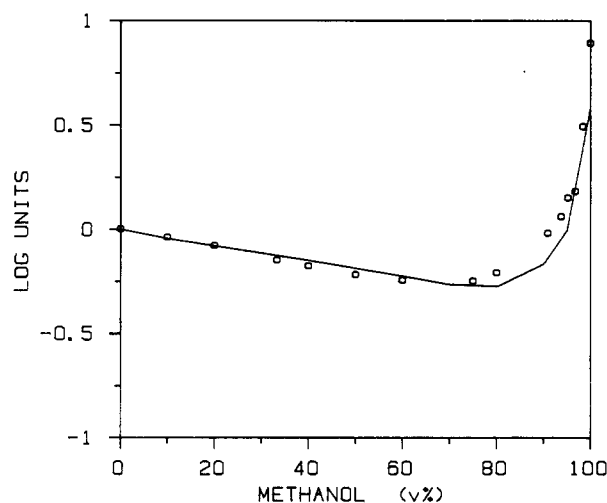

Figure 7. Correlation between the relative proton dissociation lifetime and the relative pK_a values of excited 1-amp over the full (0–100%) water/methanol composition range. The pK_a values (O) were taken from ref 78, and the full line is the kinetic data obtained from the fluorescence decay experiments.

mixtures the experimental $pK_a{}^s$ values which were measured as a function of the volume fraction of methanol[78] could be reproduced within ± 0.01 pK units over the full composition range according to the following formula:

$$pK_a{}^s = pK_a{}^w - 0.75v + 1.93 \exp[-c/2.77] \qquad (23)$$

where v is the volume fraction of methanol and c the water concentrations in moles per liter.

Figure 7 shows the correlation of the proton dissociation rate and $0.5\Delta pK_a$ for excited 1-amp over the full composition range of water–methanol solutions. As can be seen from Figure 7 the correlation between rates and equilibria is remarkably good, spanning the composition range from pure water to pure methanol. The linear dependence of rates on equilibria means that in this case the correlation also obeys the basic Brønsted relation (eq 1) with α being equal to 0.5 and β being close to unity. A linear Brønsted type dependence is predicted by both the AL model and Marcus equation over any small range of pK_a's. Furthermore, when $pK_a \approx 0$ eqs 17–21 predict that α should be equal to 0.5. In our case $pK_a = -0.34$ and $\Delta pK_a = \pm 1$, so the experimental

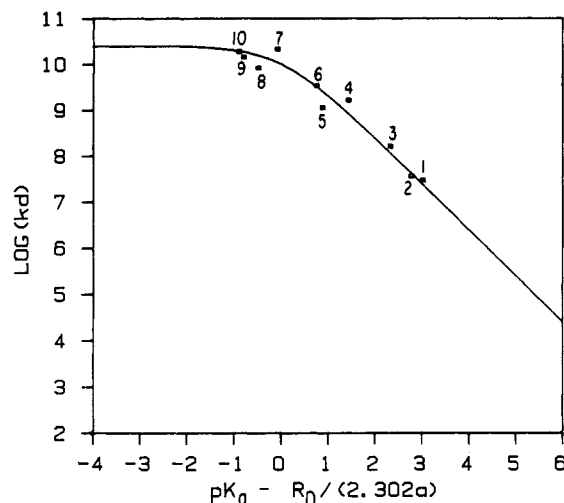(80) James, J. C.; Knox, J. G. *Trans. Faraday Soc.* **1950**, *46*, 254.

Figure 8. Correlation between $pK_{a,chem}$ of various naphthol type photoacids and their proton dissociation rate in pure water according to eqs 24–26 in the text: (1) 2-naphthol 6-methyl;[85] (2) 2-naphthol 7-methyl;[85] (3) 2-naphthol;[25,82] (4) 2-naphthol 6-bromo;[85] (5) 2-naphthol 1-chloro;[85] (6) 2-naphthol 6-sulfonate;[25,83,100] (7) 1-naphthol;[81] (8) 8-hydroxypyrene-1,3,6-trisulfonate;[35,100] (9) 2-naphthol 6,8-disulfonate;[34,25,84] (10) 2-naphthol 3,6-disulfonate.[25,84]

findings are in accord with the theoretical prediction.

(5) Structure–Reactivity Correlations in Naphthol Type Photoacids. The 1-naphthol and 2-naphthol molecules have been extensively studied in the past 30 years.[33,39,81] These molecules are considered to be model photoacids and were used by Robinson et al. to probe the structure of the proton solvation shell in water–alcohol mixtures.[16,82] In Figure 8 these two photoacids together with seven more naphthol derivatives and the closely related 8-hydroxypyrene-1,3,6-trisulfonate (HPTS) molecule are correlated according to the AL model. The dissociation rates and equilibrium constants were all taken from the literature[27,48,82–85] and are all in pure water. The purely electrostatic contribution to the pK_a was eliminated from the reported pK_a values via

$$pK_a(\text{ROH}) = pK_{a,chem} + pK_{a,el} \qquad (24)$$

where $pK_{a,chem}$ is the chemical part of the equilibrium and $pK_{a,el}$ is the purely electrostatic part given by

$$pK_{a,el} = R_D/2.303a \qquad (25)$$

where a is the initial separation between the ion pair following the complete dissociation of the bound state and R_D is the "Debye radius" which scales the electrostatic interaction given by

$$R_D = |Z_1 Z_2| e^2 / \epsilon_S k_B T \qquad (26)$$

where Z_1 and Z_2 are the charges of the ions, and ϵ_S is the static dielectric constant of the media. For univalent ion pairs at 22 °C in water $pK_{a,el} = 0.47$ for a equal to 6.5 Å.[33,35,48] The corrected pK_a's were correlated with the reported proton dissociation rates using $k_d^\circ = 2.5 \times 10^{10}$ s^{-1} and $G_a^\circ = 0.55$ kcal/mol. As discussed in section I the correlation of photoacids having $pK_a > 2$ is rather trivial when one correlates between the overall equilibrium constant and the overall dissociation rate constants. However, when one attempts to correlate between $pK_{a,chem}$ and k_d, one should also take into account the contribution of k_r to the overall dissociation rate (eq 5). In a procedure similar to that recommended by Albery[42]
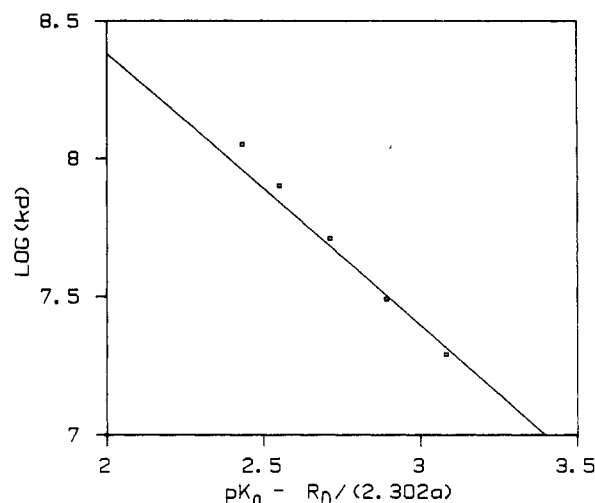

Figure 9. Correlation between the pK_a and the proton dissociation rate of 2-naphthol in water/methanol solutions as measured by Trieff and Sundheim.[89] The $pK_{a,chem}$ values were found according to eqs 24–26 in the text.

we assumed a constant correction factor in cases where only the overall dissociation rate constants were reported.[82–85] Assuming $k_r \approx k_{-D}$, one has from eq 5 that $k_d \approx 2k_{off}$. It should be noticed that such corrections are minor on a logarithmic scale and do not affect the general trend found in Figure 8.

The structure–reactivity correlation of a ROH type acid in water–organic solvent mixtures is based on a less rigorous procedure than that for the BH$^+$ type acids. A change of solvent would be expected to change both k_d° and G_a° and thus render the structure–reactivity correlation useless.[86] Indeed in ROH type dissociations the total free energy change on transferring the reaction from pure water to mixtures is in most cases strongly dependent on the anion free energy of the solvation. However, in cases when the analogous criteria of eq 15 are roughly fulfilled, i.e., when

$$\Delta G_t^{w,s} - \Delta G_t^{w,s}(\text{H}^+) \approx \Delta G_{t,el}^{w,s}(\text{RO}^-) \qquad (27)$$

one could still hope to find a suitable structure–reactivity correlation between rates and equilibria. In particular, the pK_a of the phenol derivatives in methanol and ethanol water mixtures is known to roughly scale as $(\epsilon_S)^{-1}$ up to 60–80%(m) of the alcohol.[87,88] In view of the relatively flat behavior of $\Delta G_t(\text{H}^+)$ in water–methanol solutions up to 80%(m) of methanol, it follows that $\Delta G_t^{w,s} - \Delta G_t^{w,s}(\text{H}^+)$ also scales roughly according to (ϵ_S^{-1}) in these mixtures. Indeed, in Figure 9 the data of Trief and Sundheim[89] for 2-naphthol in water–methanol solutions obey the same general structure–reactivity curve found in pure water for the naphthol derivative series (Figure 8). Recently, the pK_a^* and k_d values of HPTS in water–methanol solutions were extracted by Agmon et al. from a precise analysis of the proton dissociation curve profile.[90] Similar to the data shown in Figure 8 there is an almost linear correlation between the pK_a^* and k_d values of HPTS in water–methanol solutions up to 80%(m) of methanol. This observation confirms our suggestion that photoacids such as included in Figure 8 should conform in the water-rich region of alcoholic solvents to the same structure–reactivity relation which correlates them in pure H$_2$O.

(6) Proton Dissociation Rates in the Organic-Solvent-Rich Region. Figure 10 shows the proton dissociation rates from 1-amp in the methanol-rich region as a function of the water molar concentration. The proton dissociation rates were extracted from

(81) Shizuka, H.; Ogiwara, T.; Narita, A.; Sumitani, M.; Yoshihara, K. *J. Chem. Phys.* **1986**, *90*, 6708.

(82) Lee, J.; Griffin, R. D.; Robinson, G. W. *J. Chem. Phys.* **1985**, *82*, 4920.

(83) Kosower, E. M.; Huppert, D. *Annu. Rev. Phys. Chem.* **1986**, *37*, 127. Smith, K. K.; Huppert, D.; Gutman, M.; Kaufmann, K. J. *Chem. Phys. Lett.* **1979**, *64*, 522.

(84) Huppert, D.; Pines, E., unpublished results.

(85) Rosenberg, J. L.; Brinn, I. *J. Phys. Chem.* **1972**, *76*, 3558.

(86) Buncel, E.; Wilson, H. *Adv. Phys. Org. Chem.* **1977**, *14*, 133.

(87) Kortüm, G.; Buch, M. *Z. Elektrochem.* **1958**, *62*, 1083.

(88) Parsons, G. H.; Rochester, C. H. *J. Chem. Soc., Faraday Trans. 1* **1975**, *71*, 991.

(89) Trieff, N. M.; Sundheim, B. R. *J. Phys. Chem.* **1965**, *69*, 2044.

(90) Agmon, N.; Huppert, D.; Masad, A.; Pines, E. *Phys. Chem.*, submitted.
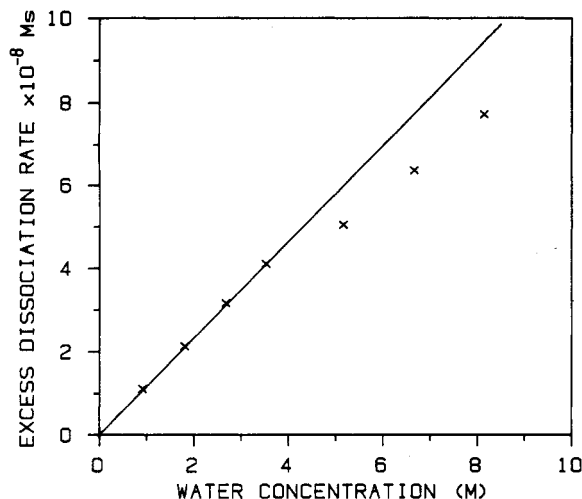
Figure 10. Dependence of the excess (over pure methanol) proton dissociation rate of 1-amp in the methanol-rich region of water/methanol mixtures on water concentration. The calculated bimolecular proton-transfer rate from 1-amp to a water molecule is $k_d = (1.0 \pm 0.1) \times 10^8$ M^{-1} s^{-1}.

the data of Figure 1 assuming that the "radiative" lifetime of protonated 1-amp is that which is found in pure methanol (~ 15 ns). In other words, we treat the proton-transfer rate to the solvent (methanol) as a constant contribution to the 1-amp fluorescence lifetime. The linear dependence of the proton-transfer rate on the water concentration below 4 M concentration of CH_3OH clearly suggests that in this region water behaves kinetically as a monomer and reacts with protonated 1-amp with bimolecular kinetics. It further demonstrates that the proton transfer to the solvent (CH_3OH) can be effectively regarded as a parallel independent process to the bimolecular reaction. Such an analysis was also carried for the other cosolvents shown in Figure 1, and similar trends were found. CH_3CN mixtures exhibit somewhat higher than linear dependence on water concentration. In all cosolvents the bimolecular rate for proton abstraction by water is in the range 7×10^7 M^{-1} s^{-1} < k_d <1.5×10^8 M^{-1} s^{-1}. This rate is an order of magnitude less than the diffusion-controlled rate limit as calculated from the self-diffusion coefficient of water in these solvents.[91] This evidently points to an activation controlled process. An additional conclusion is that in the organic-rich region of methanol, dioxane, and acetonitrile, the proton is still preferably transferred to the much less abundant water molecules. This observation is in accord with the known proton-exchange equilibrium between water and these solvents:

$$H_3{}^+O + S \rightleftharpoons SH^+ + H_2O \qquad (28)$$
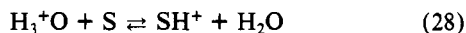
$$K_{ex} = \frac{[SH^+][H_2O]}{[S][H_3{}^+O]} \qquad (29)$$

where S is the organic molecule and K_{ex} is the proton partition equilibrium constant between s and water on the molar scale. pK_{ex} values for methanol and CH_3CN are 2.3[92] and ~ 3.5,[93a] respectively. So even with 1 M H_2O in 24 M CH_3OH (the pure solvent molarity) there is almost an order of magnitude preference of water over methanol as the proton acceptor. In fact in CH_3CN there is some evidence which implies that even with water concentrations

(91) The self-diffusion coefficients of water in CH_3OH, dioxane, CH_3CN, and DMSO are $\sim 1.6 \times 10^{-5}$, $\sim 2 \times 10^{-5}$, $\sim 4.6 \times 10^{-5}$, and $\sim 0.8 \times 10^{-5}$ cm^2 s^{-1}, respectively. Data from: Kida, J.; Hedaira, H. *J. Magn. Reson.* **1977**, *27*, 253 (methanol). Packer, K. J.; Tomilinson, D. J. *Trans. Faraday Soc.* **1971**, *67*, 1302 (DMSO). Clemett, C. J. *J. Chem. Soc. A* **1969**, 458 (dioxane). Eastel, A. J. *Aust. J. Chem.* **1980**, *33*, 1667 (CH_3CN).

(92) Kondo, Y.; Tokura, N. *Bull. Chem. Soc. Jpn.* **1972**, *45*, 818. Covington, A. K.; Dickinson, T. In *Physical Chemistry of Organic Solvent Systems*; Plenum Press: New York, 1973; p 389.

(93) (a) Kolthoff, I. M.; Ikeda, S. *J. Phys. Chem.* **1961**, *65*, 1020. (b) Kolthoff, I. M.; Chantoni, J. K., Jr.; Bhowmik, S. *J. Am. Chem. Soc.* **1968**, *90*, 23.
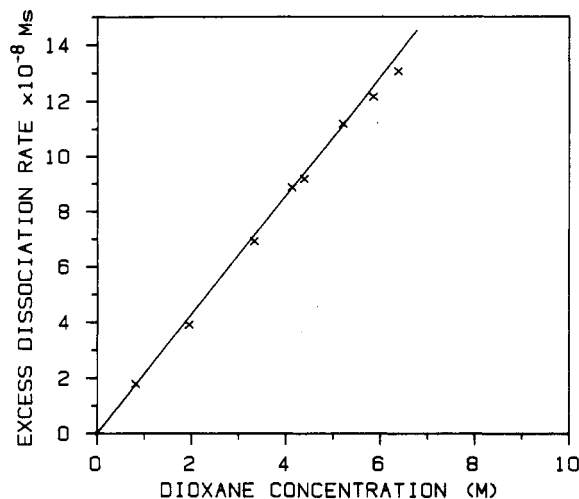

Figure 11. Dependence of the excess (over pure water) proton dissociation rate on the molar concentration of dioxane in the water-rich region of the water/dioxane mixtures. The pseudo-bimolecular proton-transfer rate (see text) is $k_d = (1.7 \pm 0.1) \times 10^8$ M^{-1} s^{-1}.

as low as 10^{-2} M the proton still prefers to be exclusively transferred to water.[94]

(7) Proton Dissociation Rates in the Water-Rich Region. Figure 11 shows the proton dissociation rates in the water-rich region of water–dioxane mixtures as a function of the dioxane concentration. Assuming that the "radiative" lifetime of protonated 1-amp is that in pure water (1.8 ns), i.e., regarding the proton-transfer rate to bulk water as a constant contribution to the 1-amp fluorescence lifetime, the proton dissociation rates linearly scale with the dioxane concentration. Such a trend is also found in the methanol, CH_3CN, and DMSO water-rich region. Although it is tempting to conclude on the basis of this observation that the proton is transferred in this region to the organic solvent the following additional observations suggest otherwise:

(a) In this region we have demonstrated a free energy relationship between rate and equilibria of the proton transfer. It is unlikely that such a free energy relationship would still hold when the proton acceptor is changed from, say, CH_3CN to DMSO. In fact, the structure reactivity correlation which is shown in Figure 6 breaks down for pure DMSO where pK_a(aniline) is 1 pK_a unit less than that in water[93b] but the proton dissociation rate was found to be half of that in pure water instead of being much larger (Figure 1). This breakdown is most probably caused by the change in the entity of the proton acceptor from H_2O to DMSO. However, even at 70%(m) of DMSO the proton transfer rates still correlate with the $pK_a{}^s$ of aniline.

(b) In water–methanol solutions the situation is more complex. As shown in Figure 7 the structure–reactivity correlation holds over the full composition range. It was also found from gas-phase measurements that there is little free energy change from methanolic solvation core to an aqueous solvation core, although the latter seems to be less favorable.[95] However, proton dissociation equilibria of hexamethylbenzenium and ammonium ions with a series of oxygen bases in aqueous solutions showed that the proton transfer to water is favorable by 1.2 kcal/mol over that to methanol,[96] while in the gas phase methanol is the preferable acceptor by 13 kcal/mol.

(c) A series of water-catalyzed deprtonation reactions of malononitrile and *tert*-butylmalononitrile in water–ethanol, water–DMSO, and water–dioxane mixture exhibit the same general pattern of proton transfer rates vs composition as found in Figure 1.[97-99] However, the proton-transfer rate decreases

(94) Desbarres, J. *Bull. Soc. Chim. Fr.* **1965**, 3240.

(95) Meot-Ner (Mautner), M. *J. Am. Chem. Soc.* **1986**, *108*, 6189.

(96) Taft, R. W. In *Prog. Phys. Org. Chem.* **1983**, *14*, 247.

(97) Hibbert, F.; Long, F. A. *J. Am. Chem. Soc.* **1972**, *94*, 7637.

(98) Menninga, L.; Engberts, J. B. F. N. *J. Org. Chem.* **1976**, *41*, 3101.

(99) Engberts, J. B. F. N. In *Water, a Comprehensive Treatise*; Franks, F., Ed.; Plenum Press: New York, 1979; Vol. 6, Chapter 4.

TABLE I: Water/Organic Solvent Composition at Which the Maximum Proton Dissociation Rate from Excited Protonated 1-amp Is Observed

water cosolvent	proton affinity,[a] kcal/mol	proton dissociation lifetime, ps	H_2O/cosolvent molecular ratio
H_2O	159.4	1800	
CH_3OH	174.9	995	1:1
CH_3CN	180.1	725	1:1
dioxane	185.1	576	3:2
DMSO	203.3	476	2:1

[a] From the recently published uncorrected values of Taft et al.[105]

TABLE II: Single-Ion Free Energy, Enthalpy, and Entropy of the Proton Transfer from Pure Water to Aqueous Methanol at 298 K[a]

MeOH, m%	ΔG, kcal/mol	ΔH, kcal/mol	$T\Delta S$, kcal/mol
0	0	0	0
20	0.10	0.38	0.28
40	0.04	1.50	1.46
60	−0.15	1.00	1.15
80	−0.13	−0.63	−0.50
90	0.18	−2.29	−2.47
100	2.45	−3.80	−6.25

[a] Data taken from Abraham et al.[103]

monotonically when dioxane is added to pure ethanol solutions, thus proving that dioxane is not directly participating in the proton-transfer reaction, although its presence in the water-rich region clearly accelerates the process.

The similarity in the behavior of all cosolvent mixtures in the water-rich region and the observation that the magnitude of the minima of Figure 1 scale directly with the proton affinities (PA) of the cosolvents in the gas phase, DMSO > dioxane > CH_3CN > CH_3OH > H_2O,[71,96,97,99] lead us to suggest that as with the organic-solvent-rich region the proton acceptor is the water molecule over the full water–organic solvent composition range (see below). The aqueous core of the proton solvation shell interacts with the organic molecule via hydrogen bonds which enhance the basicity of the aqueous core by stripping away some of its proton density. For that reason we still monitor a linear dependence of the excess (over pure water) proton-transfer rate on the organic solvent concentration. Each organic solute affects its immediate water molecules and makes them more basic so the concentration of these more basic water molecules scales linearly with the organic solute concentration. As implied from the solution compositions at which the maximum proton dissociation rate is found (Figure 1), this interaction is essentially a pair interaction between water and the organic solvent molecule (see Table I). Further increase in the organic molecule/water ratio has much less effect on the basicity of the water over that is found in the 1:1 complex. This further effect does not compensate for the overall decrease in the proton acceptor concentration (water) as its dilution is increased. Thus, in this picture the proton affinity of the single water molecule is usually enhanced with its "dilution". For that reason, it may be oversimplified to analyze experimental data in terms of pure water proton solvation clusters having fixed proton affinities which depend only on the water cluster size.[16]

Finally, in the case of cosolvents with large proton affinities there is no a priori reason to assume that the proton is preferentially transferred to water rather than directly to these solvents so as to form an organic core for the proton solvation. The latter process probably occurs with organic amine bases,[100] which are excellent proton acceptors both in the gas and in the liquid phase due to their large intrinsic proton affinity and their ability to form a high-symmetry protonic core.

(8) Photoacid Dissociation in Water–Organic Solvent Mixtures as a Probe for the Structure of Liquid Water and the Proton Solvation Shell. In recent years it has become routine to assume that photoacid dissociation rates directly correlate with the kinetic availability of large water clusters in the medium.[16,24,26,101,102] Furthermore, it was assumed that the apparent decrease in the proton-transfer rate of ROH type photoacids in water–alcohol mixtures implies a decrease in the proton affinity of the medium as compared to pure water. This change in the proton affinity was attributed to the "dilution" of high-order water clusters by the lesser proton acceptor.

In view of the recently established $\Delta G_t^{w,s}(H^+)$ values,[67b] which are in many cases negative on going from pure water to its organic

solvent mixtures, it is obvious that this interpretation of the apparent decrease in the proton dissociation rates is not correct. The kinetic evidence presented in our work implies that many of these water–organic solvent mixtures are indeed stronger basic media then pure water. One should try to understand the meaning of these new observations without the perspective that pure liquid water is the ideal proton acceptor. We start by observing the anomaly that pure water is an exceptional proton acceptor not by being a strong base on the molecular level but rather by its ability to accommodate the proton without altering its own equilibrium structure in a major way (a proton "sponge"). Furthermore, compared to most other pure solvents, water is capable of solvating the proton with the least entropy loss. It is for that reason that many organic molecules which are stronger bases than water in the gas phase become weaker bases in the liquid as was shown by Taft.[96] This is clearly demonstrated in Table II, which is composed from the data of Abraham et al.[103] In accordance with its greater proton affinity, pure methanol solvates the proton with a larger enthalpy gain than pure water. However, the solvation process inflicts much greater entropy loss than that of the equivalent aqueous process. This entropy loss more than compensates for the enthalpy gain which results in a net large positive $\Delta G_t^{w,s}(H^+)$.

Turning to the situation in water–organic solvent mixtures, one should first note the large negative excess entropy of mixing (ΔS_m^E) found in almost all of these mixtures.[67a] With the exception of water acetonitrile mixtures, the negative ΔS_m^E is accompanied by a negative excess enthalpy of mixing (ΔH_m^E). This points out that solvents capable of forming strong hydrogen bonds with water mix with water via water–solute interactions which are stronger than the average interactions found in the pure solvents. In the case of acetonitrile water mixtures where the ΔH_m^E is positive, the structure of the mixture may not be completely homogeneous.[104] However, positive ΔH_m^E values do not necessarily mean a weaker pair interaction between the water and acetonitrile as compared to the average interaction in the pure solvents but may well be due to fewer overall hydrogen bond interactions. When a proton is introduced to a water–organic solvent binary mixture, it affects both the entropy and the enthalpy of the system (Table II). We also note that in water–methanol solutions, the proton is more stable than in pure water because of the large gain in entropy. It follows that these entropy-depleted systems are able to significantly increase their entropy by accommodating the proton. The large increase in the entropy on going from 90%(m) to 100%(m) methanol which is accompanied by a much smaller enthalpy gain may signal the transition from an aqueous protonic core (which is fully solvated by methanol) to a methanolic protonic core which is less symmetric and more rigid. The large 2 orders of magnitude drop in the bimolecular proton-transfer rate to pure methanol supports this solvation model. As for the precise composition of the proton "solvation shell", one also has to consider the possibility that one or two water molecules accompany the proton when it dissociates in pure water,[15,96] so

(100) Politi, M. J.; Chaimovich, H. *J. Solution Chem.* **1989**, *18*, 1055.

(101) Schulman, S. G.; Kelly, R. N.; Gonzalez, N. J. *Pure Appl. Chem.* **1987**, *59*, 655.

(102) Bardez, E.; Goguillon, B. T.; Keh, E.; Valleur, B. *J. Phys. Chem.* **1984**, *88*, 1909. Bardez, E.; Monnier, E.; Valeur, B. *J. Phys. Chem.* **1985**, *89*, 5031.

(103) Abraham, M. H.; Hill, T.; Ling, H. C.; Schulz, R. A.; Watt, R. A. C. *J. Chem. Soc., Faraday Trans. 1* **1984**, *80*, 489.

(104) Balakrishnan, S.; Easteal, A. J. *Aust. J. Chem.* **1981**, *34*, 943 and references therein.

(105) Taft, R. W.; Anvia, F.; Gal, J.-F.; Walsh, S.; Capon, M.; Holmes, M. C.; Hosn, K.; Oloumi, G.; Vasanwala, R.; Yazdani, S. *Pure Appl. Chem.* **1990**, *62*, 17.

the extent of the preferential water solvation of the proton is yet to be determined. Finally, we point out that the generality of our conclusions can be tested even in cases where the water cluster models fail to predict the behavior of the ROH type acids. In particular, it was found that the dissociation rates from various naphthol derivatives decrease in the series water–MeOH, water–EtOH, and water–PrOH in a way which parallels the decrease in their ϵ_s values.[24] A similar trend in the same cosolvent series was found in the dissociation rate of the H_2O molecule which dissociates according to $H_2O \leftrightarrow OH^- + H^+$.[79] In addition, the proton dissociation rates of several naphthol derivatives were found to be unchanged or even to increase in water–urea solutions,[100] which is a strong gas-phase base, and its aqueous mixtures have a very large static dielectric constant. All these observations are consistent with a general structure–reactivity law which dictates the proton dissociation dynamics and point out the importance of the water cosolvent in determining the proton-transfer rate.

VI. Summary

In this study we have demonstrated that photoacid dissociation rates obey a general structure–reactivity law which correlates the proton dissociation rate with the pK_a values of the precursor acid. By doing so we have shown that photoacids conform to the same empirical general reactivity laws which dictate the behavior of the overwhelming majority of ground-state acids. At the same time we do not find any support to the idea that these rates are determined in water–organic solutions by the kinetic availability

of large water molecule clusters.

An alternative view has been expressed in recent correspondence with Professor G. Robinson. He proposes that cationic acids prepolarize the water structure around them and as such they are suitable neither for a direct comparison with the ROH type acids nor for the critical test of the validity of the various water-cluster models. Clearly more work is required to establish the generality of the two types of approaches.

Our observations also indicate that the proton-transfer rate is greatly influenced by the chemical and physical properties of the water cosolvent. In particular, the basicity of a water molecule which is hydrogen bonded to an organic molecule increases when the gas-phase proton affinity of the organic molecule is larger than that of water. Such induced basicity in water was very recently reported for the water–DMSO system from direct spectrophotometric measurements of the coordination constant of 1:1 water–DMSO dimer to a Ni^{2+} metal complex in nitrobenzene.[106] As for the proton solvation shell, we consider the question of its precise nature in binary solution to be largely open. Both direct cosolvent participation and inductive modification of the water basicity should be considered in formulating a complete model.

(106) Iwamoto, E.; Nishimoto, J.; Yakoyama, T.; Yamamoto, K.; Kumamaru, T. *J. Chem. Soc., Faraday Trans.* **1991,** *87,* 1537.

Tunneling Paths in Intramolecular Proton Transfer

Norihiro Shida,[†] Jan Almlöf,* and Paul F. Barbara*

Department of Chemistry, University of Minnesota, 207 Pleasant Street S.E., Minneapolis, Minnesota 55455 (Received: April 16, 1991)

We recently introduced a procedure for determining effective tunneling paths in nonthermal symmetric proton-transfer reactions. The procedure involves a "maximum probability path" (MPP) based on the ground-state vibrational wave function, which can be determined within the reaction surface Hamiltonian framework by ab initio electronic structure methods. In this paper we further define the MPP concept and explore this path using the malonaldehyde molecule and the formic acid dimer as examples, as well as calculations on simpler model systems. The results verify our qualitative expectations for tunneling in heavy–light–heavy systems. Furthermore, a comparison of the MPP to other more conventional paths, such as the IRC, helps elucidate certain physical aspects of multidimensional tunneling. Finally, simple extensions of the MPP procedure suggest possible directions for new methods for calculating tunneling splitting in complex molecules.

1. Introduction

Nuclear tunneling between isomers in polyatomic molecules plays an important role in the static and dynamic spectroscopy of inter- and intramolecular hydrogen bonded molecules in their ground and excited electronic states. Tunneling, for example, can be the dominant mechanism of irreversible excited-state intramolecular proton transfer, and tunneling is thus essential to the ultrafast kinetics of this reaction type.[1,2] A less common but simpler role for nuclear tunneling of hydrogen-bonded systems is observed in smaller symmetrical systems like malonaldehyde (Figure 1), for which nuclear tunneling is responsible for the splitting of the lowest energy vibrational wave function of these systems into two levels that can be detected spectroscopically,[3,4] in analogy with the splitting observed in an one-dimensional double-minimum potential.

For polyatomic systems the tunneling process is complex due to the large number of vibrational degrees of freedom. Tunneling

splittings of molecules such as, e.g., malonaldehyde or ammonia, are analogous to the tunneling splittings of carboxylic acid dimers,[5] which involve a double proton transfer, as shown in Figure 2.

One characteristic feature of all of these systems is that several vibrational degrees of freedom with very different effective masses are involved in the tunneling motion. In addition to the obvious light hydrogenic coordinate, more massive degrees of freedom are strongly coupled to the tunneling motion. The stretching displacement of the oxygen atoms bridged by the tunneling proton

(1) Many of the papers in this special issue deal with the role of nuclear tunneling in kinetics and spectroscopy. Also, see the special issue of *Chemical Physics* (**1989,** *136,* No. 2) devoted to "Spectroscopy and Dynamics of Elementary Proton Transfer in Polyatomic Systems".

(2) For a recent review of excited-state proton-transfer research, see: Barbara, P. F.; Walsh, P. K.; Brus, L. E. *J. Phys. Chem.* **1989,** *93,* 29.

(3) Rowe, W. F.; Durest, R. W.; Wilson, E. B. *J. Am. Chem. Soc.* **1981,** *103,* 6292.

(4) Firth, D. W.; Beyer, K.; Dvorak, M. A.; Reeve, S. V.; Grushow, A.; Leopole, K. R. *J. Chem. Phys.* **1991,** *94,* 1812.

(5) Rambaud, C.; Oppenländer, A.; Pierre, M.; Trommsdorff, H. P.; Vial, J. C. *Chem. Phys.* **1989,** *136,* 335.

† Present address: Division of Theoretical Studies, Institute for Molecular Science, Myodaiji, Okazaki, 44 Japan.
* To whom correspondence should be addressed.