16 The Kinetic Isotope Effect in the Photo-Dissociation Reaction of Excited-State Acids in Aqueous Solutions

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CONTENTS

I. INTRODUCTION

The effect of isotope substitution (H/D) on the kinetics of acid–base reactions has long been a primary mechanistic tool in proton-transfer research.[1–7] The relative ease with which such experiments may be carried out in aqueous solutions of strong acids has made this mechanistic approach very appealing to experimentalists. From the standpoint of theory, the isotope effect on proton transfer dynamics gives access to the microscopic mechanistic details of the process.

Laser-induced photoacidity has been recognized for decades as the major experimental tool in the research of bimolecular proton transfer reactions in aqueous solutions. In such experiments, which were pioneered by Förster and Weller,[8,9] one uses suitable organic molecules which are weak acids in the electronic ground state. Upon optical excitation the acidity of the molecules increases considerably and rivals the acidity of strong mineral acids. Molecules which undergo such a transition in their acidity are usually termed *photoacids*.[10] The most commonly used photoacids are simple hydroxyarenes (phenol-like photoacids) such as the 1- and 2-naphthols which increase their acidity by more than a factor of a million when excited to their first electronic singlet state.[11–16]

Combining isotope substitution with photoacid research is potentially a very promising avenue toward unveiling the important mechanistic details of bimolecular proton transfer reactions. Despite extensive experimental and theoretical efforts many principal questions about bimolecular proton-transfer reactions in solutions are still open: in particular, is the reaction coordinate in the transition-state region the hydrogen-bond coordinate or, as expected for sufficiently highly exothermic and thereby almost barrierless reaction, is it the coordinate of solvent motion? Some related questions are: (a) what is the rate limiting step of proton transfer in solution; (b) how is proton motion coupled to the solvation process; (c) how important is proton tunneling; (d) are bimolecular

proton-transfer reactions between acids and bases in aqueous solutions through-solvents or inner-sphere processes?

Particularly valuable information about molecular mechanism can be obtained by determining the influence of the isotopic substitution on the rate constants and the activation energy of proton transfer. Below we consider the basic kinetic models for acid–base reactions in solutions and apply them to the proton-dissociation reaction of photoacids.

II. GENERAL KINETIC MODELS FOR ACID–BASE REACTIONS IN SOLUTIONS

A. THE TWO-STATE PROTON-TRANSFER REACTION MODEL (THE EIGEN–WELLER MODEL)[9,17]

To describe the photoacid dissociation rate in aqueous solution one may apply a general two-state reaction model first explored by Eigen[17] and Weller[9] in acid–base reactions in the ground state (Scheme 16.1).

where ROH is the photoacid, k_d, k_r is dissociation and recombination intrinsic rate constants, respectively, k_s, k_{diff} are the diffusion controlled rate constants for forming and separating the reactive (acid–base) complex.

The overall dissociation and the recombination rate constants of an acid (k_{dis} and k_{rec}, respectively) may be found using the steady-state approximation and are given by:

$$k_{dis} = \frac{k_s k_d}{k_s + k_r} \tag{16.1}$$

and

$$k_{rec} = \frac{k_{diff} k_r}{k_s + k_r} \tag{16.2}$$

$$k_{diff} = \frac{4\pi D R_D N}{\exp(R_D/a) - 1} \tag{16.3}$$

is the diffusion-limited rate constant of the recombination reaction[18,19] which is approached by k_{rec} when $k_r \gg k_s$. R_D is the Debye radius, given by

$$\frac{z_1 z_2 e^2}{\varepsilon k_B T}$$

z_1 and z_2 are the charge numbers of the proton and the base, e is the electron charge, k_B is the Boltzmann constant, T is the temperature, D is the mutual diffusion coefficient, a is the reaction radius (contact distance), and
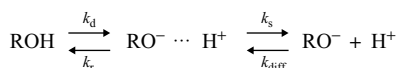
$$k_s = \frac{3D R_D}{a^3(1 - \exp(-R_D/a))} \tag{16.4}$$

is the diffusion limited rate constant of separation of the pair of reactants from their contact distance to infinite separation.

$$\text{ROH} \underset{k_r}{\overset{k_d}{\rightleftarrows}} \text{RO}^- \cdots \text{H}^+ \underset{k_{diff}}{\overset{k_s}{\rightleftarrows}} \text{RO}^- + \text{H}^+$$

SCHEME 16.1

The overall equilibrium constant of the proton-transfer reaction is given by the multiplication of the intrinsic (chemical) equilibrium constant, K_i and K_F which is the stability (Fuoss) constant of the pair at contact separation:[20]

$$K_a = \frac{k_{dis}}{k_{rec}} = \frac{k_d}{k_r}\frac{k_s}{k_{diff}} = K_i/K_F \qquad (16.5)$$

here,

$$K_i = \frac{k_d}{k_r}, \qquad K_F = \frac{k_{diff}}{k_s} = 4\pi Na^3 \exp(-R_D/a)/3 \qquad (16.6)$$

so one can write for K_i and k_r,

$$K_i = K_a K_F = 10^{-pK_a} K_F \qquad (16.7)$$

$$k_r = \frac{k_d 10^{pK_a}}{K_F} \qquad (16.8)$$

For strong acids in aqueous solutions at room temperature $k_r \ll k_s$, and by Equation 16.1

$$k_{dis} \approx k_d \qquad (16.9)$$

so for strong acids the apparent dissociation constant is practically equal to the intrinsic proton-dissociation rate.

The situation is different for weak acids. In this case $k_s \ll k_r$ and

$$k_{dis} = \frac{k_s k_d}{k_s + k_r} \cong \frac{k_s k_d}{k_r} \qquad (16.10)$$

so the apparent dissociation rate of weak acids depends on both the intrinsic equilibrium constant K_i and k_s.

Similarly, for deuterated water one can write,

$$k_{dis}^D = \frac{k_s^D k_d^D}{k_r^D} \qquad (16.10a)$$

and the apparent kinetic isotope effect (KIE) of the acid dissociation may be estimated using Equation 16.11

$$k_{dis}^H/k_{dis}^D = \frac{k_s^H k_d^H}{k_r^H} \Big/ \frac{k_s^D k_d^D}{k_r^D} = \frac{k_s^H K_i^H}{k_s^D K_i^D} = \frac{k_s^H K_a^H}{k_s^D K_a^D} = \frac{D_H}{D_D} 10^{\Delta pK_a} = 1.4 \times 10^{\Delta pK_a} \qquad (16.11)$$

here $\Delta pK_a = pK_a^D - pK_a^H$ and the ratio D_H/D_D is calculated for room temperature (≈ 1.4 in aqueous solutions) and is assumed the only reaction parameter in k_s^H/k_s^D which is isotope dependent.

The important conclusion from this simple kinetic analysis is that there is a fundamental difference in the observed dissociation rates of strong and weak acids. While the actual proton dissociation stage may be observed directly in strong-acid dissociation, it is impossible to do so in weak-acid dissociation. In the latter case a multistage reaction rate is the directly accessed observable which is made of both the chemical and the diffusive stages of the dissociation reaction.

154 **B. FREE-ENERGY CORRELATIONS OF THE PROTON (DEUTERON) TRANSFER RATES**

155 Brønsted relation[21,22] represents the earliest example of a linear free-energy relation between
156 reaction rates and equilibrium constants:
157

$$\Delta G^{\#} = \alpha(\Delta G^0) \qquad (16.12)$$

159
160 where $\Delta G^{\#}$ is the free energy of activation and ΔG^0 is the standard free-energy change in the
161 proton-transfer reaction, $\Delta G^0 = RT \ln(10) pK_a$. The curvature of the Brønsted plot, α, determines
162 how the potential surface of the reactions varies with ΔG^0. Most modern theories treat proton
163 transfer in aqueous solution as a dynamic process closely coupled with solvent motion and solvent
164 relaxation.[23−30] These theories assume that the proton-transfer rate between a proton donor and a
165 proton acceptor has a general form akin to a transition-state rate constant which may be written as:
166

$$k_d = \frac{w}{2\pi} \exp\left(-\frac{\Delta G^{\#}}{RT}\right) \qquad (16.13)$$

169 where w is the frequency factor of the proton transfer reaction which may also depend on solvent
170 relaxation frequencies.
171   As early as 1924 Brønsted and Pedersen[22] argued that for a highly exothermic proton transfer
172 reaction $\Delta G^{\#} = 0$, so that at this limit the curvature of the Brønsted plot approaches zero and k_d
173 becomes a constant independent of ΔG^0. These ideas were refined by Marcus using a reaction
174 model originally developed for outer-sphere (nonadiabatic) electron-transfer reactions. According
175 to the Marcus theory,[23] charge-transfer reactions proceed along the solvent coordinate with an
176 intrinsic activation energy equal to 1/4 of the total solvent reorganization energy. The potential
177 energy of such a reaction along the reaction coordinate involves a pair of intersecting parabolas
178 which gives at their intersection point the activation energy needed for the transfer from the reactant
179 to the product state:
180

$$\Delta G^{\#} = \left(1 + \frac{\Delta G^0}{4\Delta G_0^{\#}}\right)^2 \Delta G_0^{\#} \qquad (16.14)$$

183
184 where $\Delta G_0^{\#}$ is the intrinsic activation energy of a symmetric transfer where the total free-energy
185 change (ΔG^0) following the charge transfer is equal to zero. Marcus theory gives solvent
186 reorganization central role in charge-transfer reactions in polar solvents where solvent
187 reorganization energy is large. After being developed for weak-overlap electron transfers Marcus
188 and Cohen,[24] subsequently applied it in a semiempiric way to proton transfer reactions by using a
189 Bond-Energy–Bond-Order (BEBO) model for the proton-transfer coordinate along a preexisting
190 hydrogen bond. The semiempiric BEBO treatment represents the opposite extreme to nonadiabatic
191 electron transfer. Here, bond rupture–bond formation is the principal contributor to the reaction
192 coordinate. In this case the potential energy along the reaction coordinate is initially constant and
193 then rises to a maximum at the transition state and than falls to another constant value. The
194 activation energy is given by:
195

$$\Delta G^{\#} = \frac{\Delta G^0}{2} + \Delta G_0^{\#} + \frac{\Delta G_0^{\#}}{\ln 2} \ln \cosh\left(\frac{\Delta G^0 \ln 2}{2\Delta G_0^{\#}}\right) \qquad (16.15)$$

198
199 Remarkably, the two different approaches (i.e., Equation 16.14 and Equation 16.15) yield almost
200 identical results in the endothermic branch of the reactions. However, at the exothermic branch the
201 approaches differ considerably. While Marcus theory predicts an "inverted region" where, because
202 of the quadratic nature of Equation 16.14, activation energy "reappears" when the driving force of
203 the reaction becomes very large, it is not so with the BEBO model, Equation 16.15, where the
204 reaction rate assumes a constant (maximal) value in this limit.

A nontraditional picture for adiabatic proton-transfer reaction in polar environment was recently developed by Kiefer and Hynes.[29] Proton nuclear motion was treated quantum mechanically but the proton did not tunnel as the reaction barrier was found to fluctuate because of solvent rearrangements to below the zero-point energy of the proton. This description of nonadiabatic proton-transfer strongly differs from Marcus semiempiric treatment of proton transfer in quantization of the proton and consideration of solvent coordinate as the reaction coordinate. Remarkably, despite the different treatment, Kiefer and Hynes showed that the reaction path of the proton transfer characterized by the values of the quantum-averaged proton coordinate is very similar to the BEBO pathway.

A quadratic free energy relationship resembling that of Marcus was found between the activation free energy $\Delta G^{\#}$ and the reaction asymmetry ΔG^0:

$$\Delta G^{\#} = \Delta G_0^{\#} + \alpha \Delta G^0 + \alpha' \frac{(\Delta G^0)^2}{2} \tag{16.16}$$

With $\alpha_0 = 1/2$, the Brønstead coefficient evaluated at $\Delta G^0 = 0$ and α' the Brønstead coefficient slope evaluated at $\Delta G^0 = 0$.

C. THE ISOTOPE EFFECT IN A SERIES OF SIMILAR REACTIONS

It was first suggested by Westheimer and Melander[31,32] to correlate KIE with the free-energy change in reaction. They have argued that the KIE, k_H/k_D, given by Equation 16.17, would have a maximum value for a symmetrical transition state ($\Delta G^0 = 0$).

$$\frac{k_H}{k_D} = \exp(-(\Delta G_H^{\#} - \Delta G_D^{\#})/RT) \tag{16.17}$$

Marcus, assuming that $\Delta G_0^{\#}$ is the only isotope-dependent quantity in Equation 16.14 and Equation 16.15 derived a relationship between $\ln(k_H/k_D)$ and the free-energy change of the proton transfer. By Equation 16.14, the free-energy dependence of the KIE is given by a symmetric curve having a maximum at $\Delta G^0 = 0$:

$$\frac{k_H}{k_D} \cong \exp\left(\left(\frac{\Delta G_{0H}^{\#} - \Delta G_{0D}^{\#}}{RT}\right)\left(1 - \frac{\Delta G^{0^2}}{16\Delta G_{0H}^{\#}\Delta G_{0D}^{\#}}\right)\right) \tag{16.18}$$

or assuming $(\Delta G_0^{\#})^2 = \Delta G_{0H}^{\#}\Delta G_{0D}^{\#}$ in term of maximal KIE for the symmetric reaction

$$\ln\left(\frac{k_H}{k_D}\right) \cong \left[\ln\left(\frac{k_H}{k_D}\right)\right]_{max}\left[1 - \left(\frac{\Delta G^0}{4\Delta G_0^{\#}}\right)^2\right] \tag{16.18a}$$

The BEBO equation[24] (Equation 16.15) results with a more complex relationship:

$$\ln\left(\frac{k_H}{k_D}\right) \cong \left(\frac{\Delta G_{0H}^{\#} - \Delta G_{0D}^{\#}}{RT}\right)\left(1 - \frac{\dfrac{\Delta G^0 \ln 2}{2\Delta G_0^{\#}}\tanh\left(\dfrac{\Delta G^0 \ln 2}{2\Delta G_0^{\#}}\right) + \ln\cosh\left(\dfrac{\Delta G^0 \ln 2}{2\Delta G_0^{\#}}\right)}{\ln 2}\right)$$

$$= \left[\ln\left(\frac{k_H}{k_D}\right)\right]_{max}\left(1 - \frac{\dfrac{\Delta G^0 \ln 2}{2\Delta G_0^{\#}}\tanh\left(\dfrac{\Delta G^0 \ln 2}{2\Delta G_0^{\#}}\right) + \ln\cosh\left(\dfrac{\Delta G^0 \ln 2}{2\Delta G_0^{\#}}\right)}{\ln 2}\right) \tag{16.19}$$

256 This yields a KIE curve with maximum at $\Delta G^0 = 0$ which decreases in a bell-like fashion for
257 large $|\Delta G^0|$

258   Kiefer and Hynes[30] using the Equation 16.16 have obtained:

$$\frac{k_H}{k_D} = \exp(-(\Delta G_{0H}^{\#} - \Delta G_{0D}^{\#})/RT)\exp\left((\alpha'_{0H} - \alpha'_{0D})\frac{(\Delta G^0)^2}{2RT}\right) \qquad (16.20)$$

263 and in term of maximal KIE for the symmetric reaction

$$\frac{k_H}{k_D} = \left(\frac{k_H}{k_D}\right)_{max}\exp\left((\alpha'_{0H} - \alpha'_{0D})\frac{(\Delta G^0)^2}{2RT}\right) \qquad (16.20a)$$

269 Q1 The decrease from the maximum KIE value is characterized in Equation 16.22 by a symmetric
270 Gaussian falloff.

271   Although Equation 16.20 results from quantum mechanical considerations of the proton-
272 transfer reactions it has a similar functional form to the empiric Marcus treatment of the KIE in
273 proton transfer. I thus review the available KIE data of photoacids proton-dissociation using the
274 functional form of Marcus treatment Equation 16.18.

III. THE ISOTOPE EFFECT IN THE EQUILIBRIUM CONSTANT
OF PHOTOACIDS

279 The isotope effect on the equilibrium constant of acids, K_a^H/K_a^D, is directly related to the KIE in
280 the acid-dissociation rate constant, k_H/k_D, (Equation 16.11). R.P. Bell summarized in 1959 the
281 available data at that time on the isotope effect in acids equilibria and concluded in the first edition
282 of his classic book *The Proton in Chemistry*[1a] that a linear correlation exists between the acid
283 strength (K_a) and K_a^H/K_a^D. Bell retreated from this observation in the second edition of his
284 book published some 14 years later.[1b] In Bell's words "Early results for a number of weak acids
285 suggested that the value of K_a^H/K_a^D decreased regularly with increasing acid strength, but when
286 more extensive experimental data are considered there appears to be no real basis for this
287 generalization, except perhaps as an ill-defined qualitative trend." However, after being so decisive
288 in his judgment Bell relaxes this statement and immediately continues referring indirectly to the
289 correlation he had published in the first edition of his book: "There is some evidence that such a
290 relation holds approximately for a closely related series such as the phenols and alcohols."

291   In 1991, G. Wilse Robinson[33] rediscovered the apparent relation between the equilibrium
292 constant of phenols and alcohols and the isotope effect in their equilibria and extended it to include
293 excited-state photoacids of the phenol family. The correlation he found was over 11 pK_a units in
294 acid strength and looked impressively linear.

295   I have reviewed for this contribution the data appearing in Bell's book and in Robinson's paper
296 and added additional recent data on photoacids that were mainly gathered in the past 10 years
297 (Tables 16.1–16.3).[33–40]

298   Figure 16.1 shows that the general trend observed by Bell and Robinson of decreasing isotope
299 effect with increasing acidity of the acid also appears to exist for very strong photoacids having
300 negative pK_a's. However the spread in the data points is larger than in ground-state acids
301 (Table 16.1) and larger than the one found by Robinson in the smaller set of moderately strong
302 photoacids (Table 16.2). The correlation of the equilibrium constant of the full set of photoacids
303 (including the data appearing in Table 16.3) with their isotope effect is shown in Figure 16.2. The
304 observed spread in the data set of the excited-state equilibrium constant of deuterated photoacids
305 in D_2O may be the result of the considerable experimental error in their determination by several
306 experimental methods of which none is of analytic accuracy.[11,15] In comparison, the equilibrium

TABLE 16.1

pK_a^H and pK_a^D of Ground-State Acids

Compound	pK_a^H	pK_a^D	Reference
2,6-Dinitrophenol	3.58	4.03	35
Bromothymol blue	6.15	6.68	35
3,5-Dinitrophenol	6.70	7.31	35
p-Nitrophenol	7.24	7.80	35
Water	15.74	16.55	36
Phenol	10	10.62	34
4-Phenolsulfonate	8.97	9.52	34
2,4-Dinitrophenol	4.06	4.55	34
2,5-Dinitrophenol	5.19	5.70	34
4-Nitrophenol	7.22	7.77	34
4-Bromophenol	9.35	9.94	34
3-Methoxyphenol	9.62	10.20	34
4-Methoxyphenol	10.24	10.85	34
4-Hydroxyphenyltrimethylammonium chloride	8.34	8.90	34
2-Naphthol	9.47	10.06	34

TABLE 16.2

pK_a^H and pK_a^D of Excited-State Acids Found by Förster Cycle Calculations

Compound	pK_H^*	PK_D^*	Reference
Phenol	4.1	4.6	34
4-Phenolsulfonate	2.3	2.7	34
4-Bromophenol	2.9	3.4	34
3-Methoxyphenol	4.6	5.1	34
4-Methoxyphenol	5.7	6.2	34
4-Hydroxyphenyltrimethylammonium chloride	1.6	2.0	34
2-Naphthol	3.0	3.5	33

TABLE 16.3

pK_a^H and pK_a^D of Excited-State Acids Found by Kinetic Measurements

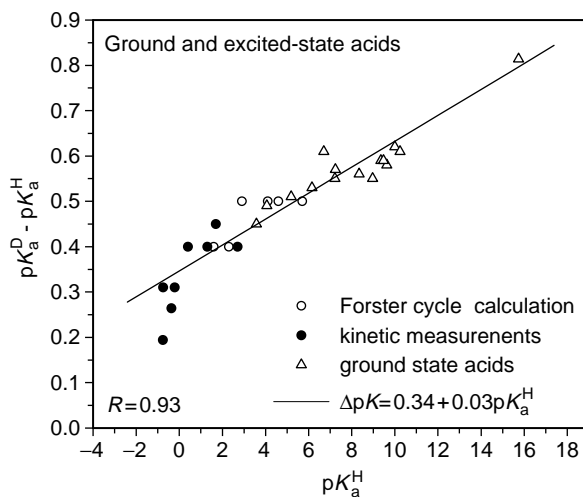| Compound	pK_a^H	pK_a^D	Reference
5-Cyano-2-naphthol	− 0.75	− 0.44	37
6-Cyano-2-naphthol	− 0.37	− 0.11	37
7-Cyano-2-naphthol	− 0.21	0.10	37
8-Cyano-2-naphthol	− 0.76	− 0.57	37
1-Naphthol	0.4	0.81	38
1-Naphthol	− 0.16	0.30	39
2-Naphthol	2.72	3.12	38
1-Naphthol 2-sulfonate	1.58	2.02	38

FIGURE 16.1 Isotope effects on equilibrium dissociation constants in acid dissociation. Open triangles: ground-state acids. Circles: excited-state acids, open symbols: Förster cycle calculation. Full symbols, kinetic calculation. Full line: linear regression $\Delta pK = 0.34 + 0.03pK_a^H$. $R = 0.93$.

constant of deuterated ground-state acids was determined directly by titration in D_2O, a much more reliable method than the methods used for the estimation of excited-state equilibrium constants. These methods have included Förster cycle calculations and analysis of the reversible deuteron-transfer kinetics of photoacids in their excited state.[11,15]

My conclusion is that within the set of phenol-type photoacids that has been investigated so far the linear correlation between the isotope effect in their equilibria and their acidity constant seems to be a valid observation. The observed correlation is similar (but not necessarily identical) to the one observed for ground-state phenols and alcohols. With the aid of Equation 16.11 this apparent correlation makes a useful experimental tool for predicting the KIE in the dissociation of weak photoacids which is discussed below.
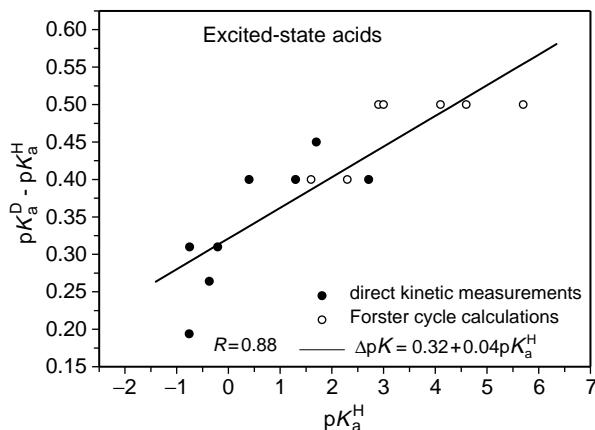

FIGURE 16.2 Isotope effects on equilibrium dissociation constants in excited-state acid dissociation. Open symbols: Förster cycle calculation. Full symbols, kinetic calculation. Full line: linear regression $\Delta pK = 0.32 + 0.04pK_a^H$. $R = 0.88$.

IV. THE KIE IN PHOTOACID DISSOCIATION

The proton-dissociation reaction of strong photoacids is readily measured directly in aqueous solutions. For typical time-resolved dissociation profiles of a strong photoacid (1-naphthol 5-sulfonate)[41] in water and D_2O see Figure 16.3. In this case, the dissociation rate of the strong photoacid is much faster than its fluorescence decay. The situation changes dramatically for photoacids having $pK_a^* > 2$, at which the dissociation rate of the phenol-type photoacids becomes smaller than the decay rate of the excited state which is usually $1-10$ nsec. This limits the observation range of photoacid dissociation to relatively strong photoacids. Until our own recent observation of the proton dissociation reaction of hydroxypyrene in water,[42] there has been no report on the dissociation rate of a phenol-type photoacid having a pK_a^* higher than 3.5. The weakest photoacid for which the KIE has been reported was 2-naphthol, having a pK_a^* of $2.7-2.8$ in H_2O. The available data on the KIE of photoacids were recently reviewed by Formosinho, Arnaut, and Barroso.[43] The KIE was analyzed using the intersecting-state model (ISM) of proton transfer which is an empiric, BEBO-type, reaction model. The fit between the calculated and observed KIE was found to be only qualitative. Not intentionally, the review also gave an opportunity to examine critically the experimental KIE values reported by several research groups. The experimental KIE value given for HPTS (8-hydroxypyrene-1,3,6-trisulfonate), one of the most investigated photoacids, was 4.71 while the calculated value using ISM correlation was 2.68. However, closer examination of the literature reveals considerable variations in the reported value of the KIE of HPTS. Politi et al.[44] reported a value of 5.6 while the Huppert group using much more sophisticated kinetic analysis reported in a variety of experiments extending over a 10-year period KIE values between 2.4 and 4.1,[45] the best value being most probably just in the middle of this range, about 3.2 at 293 K extracted by us from revaluating the kinetic data reported in Ref. 45. The example of HPTS points to the fact that one needs to be critical when analyzing experimentally found KIE values even for strong, well-behaved photoacids such as HPTS.

The dependence of the KIE on the pK_a values of a family of similarly structured acids is predicted by Equation 16.18 to be bell-shaped and symmetric about the center given at $pK_a = 0$. Figure 16.4 shows a schematic drawing of the functional form of the KIE using Equation 16.18. Such a correlation, spanning over a range of 25 pK_a units, was shown by Bell in his book (page 265 in Ref. 1b) for a family of carbon acids. The subject was further discussed in the 1975 symposia of
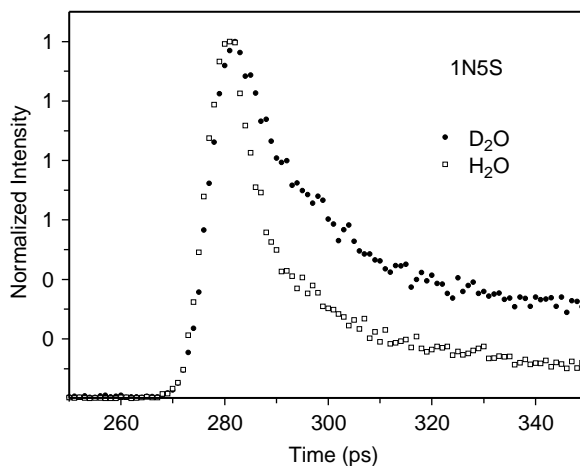

FIGURE 16.3 The isotope effect in the proton dissociation of 1-naphthol-5-sulfonate (1N5S), $T = 20°C$. Upper decay: 1N5S deuteron dissociation in D_2O, $\tau_1 = 31$ psec (93%). Lower decay: 1N5S proton dissociation in water, $\tau_1 = 15$ psec (99%).
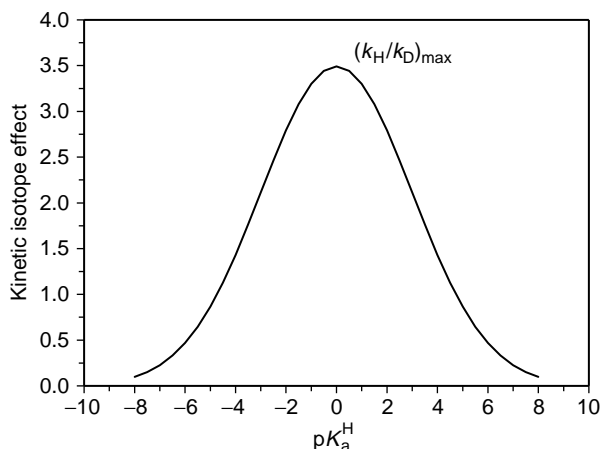
FIGURE 16.4 The functional form of the KIE predicted by the Marcus equation, Equation 16.18.

the Faraday Society by Bell and Kresge[46] with references to Bell's[1b] and Bordwell and Boyle's data.[47] Several less spectacular examples of the bell-shape dependence of the KIE were discussed more recently by Sühnel.[48] Although the situation is far from being clear-cut, the usefulness of Equation 16.18 seems to be established for families of closely-related ground-state proton-transfer reactions. The situation is different for excited-state acids. The limited available data on the KIE of photoacids (Table 16.4) need to be analyzed with caution. One may start such a kinetic analysis by using Equation 16.1, assuming the isotope effect on the intrinsic proton-transfer rate, k_d, has the functional form given by Equation 16.15. Assuming that $k_s \gg k_r$ up to $pK_a^* = 0$ and that $k_r \gg k_s$ from $pK_a^* = 4$ onward (see Figure 16.5 for details) one may arrive by using Equation 16.10 and Equation 16.11 at a projected profile for the observed KIE in photoacid dissociation. Clearly, the measured KIE of photoacids having a pK_a^* below 1 is expected to drop sharply as a function of increased photoacidity (Figure 16.4), while the measured KIE of photoacids with a pK_a^* above 4 is predicted to increase monotonously with the isotope effect in their equilibria (Figure 16.1). The experimental data gathered so far are shown in Figure 16.6 and seem to confirm the first prediction: The isotope effect

TABLE 16.4
The KIE in the Observed Proton-Dissociation Rate of Some Phenol-Like Photoacids

Compound	k_H/k_D	References
1-Naphthol 5-sulfonate	2	41
5-Cyano-1-naphthol	1.6	46
1-Naphthol	3.2	39
1-Hydroxypyrene	3.9	42
2-Naphthol	2.64	38
HPTS (8-hydroxypyrene-1,3,6-tri-solfunate)	3.16	45
5-Cyano-2-naphthol	2.33	37
6-Cyano-2-naphthol	3.33	37
7-Cyano-2-naphthol	3.06	37
8-Cyano-2-naphthol	2.08	37
2-Naphthol-3,6-disulfonate	3.6	40
1-Naphthol-2-sulfonate	2.8	38

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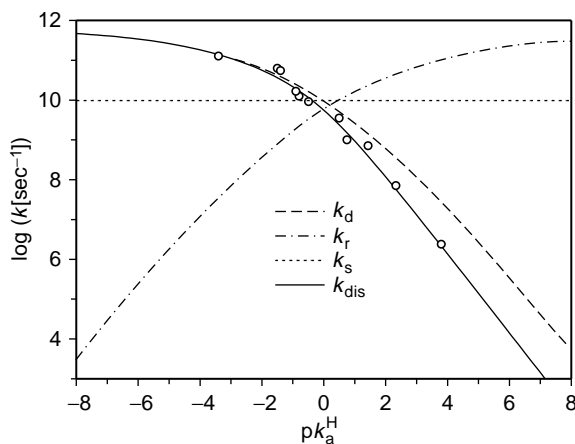513
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FIGURE 16.5 The correlation between the excited-state dissociation rate constants of phenol-type photoacids (k_d) and their corresponding equilibrium constant values using Marcus (BEBO) correlation, Equation 16.15 with $\Delta G_0^{\#} = 2.5$ kcal/mol and $(k_d)_{max} \approx (2 \text{ psec})^{-1}$ (dashed line).[16b] The full line is the overall dissociation rate of the reactions, k_{dis}, (Equation 16.1), as correlated with the experimental data assuming k_d is the intrinsic proton dissociation rate, Equation 16.15. The dashed-dotted line is the calculated proton recombination rate constants, k_r using Equation 16.8. The dotted line is the diffusion controlled rate constant of separation of the products calculated from the contact distance to infinite separation, k_s (Equation 16.4).

of phenol-type photoacids having a pK_a^* value of about zero has been consistently found to be between 3 and 3.6, while stronger photoacids of this family exhibit much smaller KIE's down to about 1.6.[49] At the other limit of the correlation there is only one data point so far, the one for hydroxypyrene (HP, $pK_a^* = 4.1–4.3$). The apparent KIE of HP is the largest found so far for a phenol-type photoacid, KIE (HP) $= 3.8 \pm 0.1$.[42] This value agrees well with the KIE value calculated using Equation 16.11.
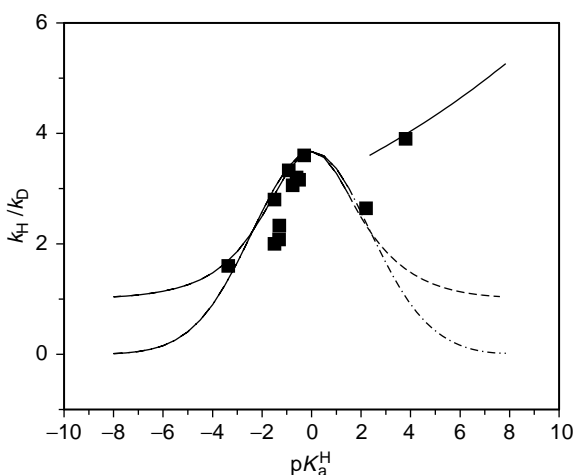

FIGURE 16.6 The KIE of the proton dissociation rate of phenol-type photoacids. Dotted-dashed line: Marcus equation, Equation 16.18. Dashed line: Marcus BEBO expression, Equation 16.19 with $(k_d^H/k_d^D)_{max} = 3.5$. Full line, the predicted isotope effect combining the two-state model, Equation 16.1 with the experimental values of the isotope effect on the acid equilibria ΔpK_a (Figure 16.1 and Equation 16.11).

These findings, although summarizing experimental studies which have been carried out over a period of roughly two decades, are still not fully decisive. However, these findings allow some general observations which are discussed below.

V. CONCLUDING REMARKS ON THE KIE IN PHOTOACID (PHENOL-TYPE) DISSOCIATION

The following conclusions may be drawn from the kinetic data gathered so far on the KIE in photoacid dissociation:

a. The KIE found in strong photoacid dissociation is relatively small and peaks around $pK_a^* = 0$.

b. The KIE tends to decrease rapidly below $pK_a^* = 0$.

c. Judging from the measured KIE values and the isotope effect in the photoacid equilibria, the KIE in the opposite proton-recombination reaction should be close to unity. This conclusion is supported by the kinetic analysis of reversible excited-state proton-transfer reactions.[50] In all reported cases where such a kinetic analysis was undertaken the KIE in the proton-recombination rate was found to be between unity and 1.6.[37–40,45] This important observation about the significant difference in the KIE of k_d and k_r still awaits detailed explanation.

d. Judging from some early measurements on a different class of photoacids, the amine photoacids, the observed KIE in photoacid dissociation may well depend on the chemical nature of the photoacid. In the case of amine photoacids a large isotope effect of about 5 was observed and was found, in most cases studied so far, to be independent of the photoacid pK_a^* in the acidity range, $-7 < pK_a^* < 0$.[51,52] This behavior contrasts with the one found for phenol-type photoacid suggesting that proton dissociation from neutral oxygen acids proceed by a different mechanism than proton dissociation from protonated nitrogen acids which are positively charged prior to dissociation.[16b] Currently, there is no theoretical modeling of these "cationic" proton-transfer reactions. Transient solvation effects and large inner-sphere rearrangements may be the crucial factor in determining the pathway of cationic proton-dissociation reaction of amine acids. These observations on the KIE of amine photoacids draw further attention to Bell's remarks given in Section III of this chapter. As Bell pointed out it has indeed become clear that one should be extremely cautious when trying to generalize processes as sensitive to environment (solvent) and intramolecular structure as the KIE in photoacids proton dissociation. A similar cautionary remark may be made about any theoretical modeling (at any level of theory) that suggests one general mechanistic treatment for the KIE in all proton-transfer reactions. A more realistic view is to assume that there are several possible pathways leading to proton transfer and to the observed KIE in proton-transfer. Generalizations should be limited to specific groups of acids and may become realistic only after extensive data gathering and after critically examining the various experimental findings. It was the aim of this contribution to point out that such a goal is probably possible for the phenol-type family of excited-state acids.

REFERENCES

1 (a) Bell, R. P., *The Proton in Chemistry*, 1st ed., Cornell University Press, Ithaca, NY, Chap. 6, 1959.
   (b) Bell, R. P., *The Proton in Chemistry*, 2nd ed., Cornell University Press, Ithaca, NY, Chap. 6, 1973.
2 Caldin, E. and Gold, V., *Proton-Transfer Reactions*, Chapman & Hall, London, 1975.
3 Hibbert, F., *Adv. Org. Chem.*, 26, 255, 1990.

613   4 Kresge, A. J., *Isotope Effects on Enzyme-Catalyzed Reactions*, Cleland, W. W., O'Leary, M. H., and
614   Northrop, D. B., Eds., University Park Press, Baltimore, MD, p. 37, 1977.
615   5 Kresge, A. J. and Silverman, D. N., *Methods Enzymol.*, 308, 726, 1999.
616   6 Silverman, D. N., *BBA*, 1458, 88, 2000.
617   7 Kohen, A. and Klinman, J. P., *Acc. Chem. Res.*, 31, 397, 1998.
618   8 Förster, Th. Z., *Elecrtochem. Angew. Phys. Chem.*, 42, 531, 1952.
619   9 Weller, A. Z., *Electrochemistry*, 58, 849, 1954. Weller, A. Z., *Phys. Chem. N.F.*, 17, 224, 1958.
620 10 Tolbert, L. M. and Haubrich, J. E., *J. Am. Chem. Soc.*, 116, 10593, 1994.
621 11 Ireland, J. F. and Wyatt, P. A. H., *Adv. Phys. Org. Chem.*, 12, 131, 1976.
622 12 Shisuka, H., *Acc. Chem. Res.*, 18, 41, 1985.
623 13 Arnaut, L. G. and Formosinho, S. J., *J. Photochem. Photobiol.*, A75, 21, 1993.
624 14 Wan, P. and Shukla, D., *Chem. Rev.*, 93, 571, 1993.
625 15 Tolbert, L. M. and Solntsev, K. M., *Acc. Chem. Res.*, 35, 19, 2002.
626 16 (a) Pines, E., UV-Visible spectra and photoacidity of phenols, naphthols and pyrenols, In *The
627   Chemistry of Phenols*, Rappoport, Z., Ed., Wiley, New York, pp. 491–529, 2003; (b) Pines, E. and
628   Pines, D., *Ultrafast Hydrogen Bonding and Proton Transfer Processes in the Condensed Phase*,
629   Elsaesser, T. and Bakker, H. J., Eds., Kluwer, Dordrecht, pp. 155–184, 2002.
630 17 (a) Eigen, M., Kruse, W., and De Maeyer, L., *Prog. React. Kinet.*, 2, 285, 1964; (b) Eigen, M., *Angew.
631   Chem. Int. Ed. Engl.*, 80, 5059, 1958.
632 18 Smoluchowski, M. Z., *Phys. Chem.*, 92, 129, 1917.
633 19 Debye, P., *Trans. Electrochem. Soc.*, 82, 265, 1942.
634 20 Fuoss, R. M., *J. Am. Chem. Soc. Electrochem.*, 54, 531, 1950.
635 21 (a) Brønsted, J. N., *Rev. Trav. Chim.*, 42, 718, 1923; (b) Brønsted, J. N., *Chem. Rev.*, 5, 231, 1928.
636 22 Brønsted, J. N. and Pedersen, K. Z., *Phys. Chem.*, 108, 185, 1924.
637 23 (a) Marcus, R. A., *J. Chem. Phys.*, 24, 966, 1954; (b) Marcus, R. A., *J. Phys. Chem.*, 67, 853, 1963; (c)
638   Marcus, R. A., *J. Ann. Rev. Phys. Chem.*, 15, 155, 1964; (d) Marcus, R. A., *J. Phys. Chem.*, 44, 679,
639   1965; (e) Marcus, R. A., *J. Phys. Chem.*, 72, 891, 1968; (f) Marcus, R. A., *J. Am. Chem. Soc.*, 91, 7244,
640   1969; (g) Marcus, R. A., *Faraday Symp. Chem. Soc.*, 10, 60, 1975; (h) Marcus, R. A., *Faraday Discuss.
641   Chem. Soc.*, 74, 7, 1982.
642 24 Cohen, A. O. and Marcus, R. A., *J. Phys. Chem.*, 72, 4249, 1982.
643 25 Levich, V. G., Dogonadze, R. D., German, E. D., Kuznetsov, A. M., and Kharkats, Yu. I., *Electrochim.
644   Acta*, 15, 353, 1970.
645 26 Kuznetzov, A. M., *Charge Transfer in Physics Chemistry and Biology: Physical mechanisms of
646   Elementary Processes and an Introduction to the Theory*, Gordon and Breach Pubs., Amsterdam, 1995.
647 27 (a) Kornyshev, A. A., Kuznetsov, A. M., Spohr, E., and Ultstrup, J., *J. Phys. Chem.*, B107, 3351–3366,
648   2003; (b) Kuznetsov, A. M., Spohr, E., and Ultstrup, J., *Can. J. Chem.*, 77, 1085, 1999.
649 28 Krishtalik, L. I., *BBA*, 1458, 6–27, 2000.
650 29 (a) Kiefer, P. M. and Hynes, J. T., *J. Phys. Chem.*, A106, 1834, 2002; (b) Kiefer, P. M., and Hynes, J. T.,
651   *J. Phys. Chem.*, A106, 1850, 2002.
652 30 Kiefer, P. M. and Hynes, J. T., *J. Phys. Chem.*, A107, 9022, 2003.
653 31 Westheimer, F. H., *Chem. Rev.*, 61, 265, 1961.
654 32 Melander, L., *Isotope Effects on Reaction Rates*, Ronald Press, New York, 1960.
655 33 Robinson, G. W., *J. Phys. Chem.*, A95, 1038, 1991.
656 34 Wehry, E. L. and Rogers, L. B., *J. Am. Chem. Soc.*, 88, 351, 2003.
657 35 Martin, D. C. and Butler, J., *J. Chem. Soc.*, 1366, 1939.
658 36 Kingerly, R. W. and LaMer, V. K., *J. Am. Chem. Soc.*, 63, 3256, 1941.
659 37 Huppert, D., Tolbert, L. M., and Linares-Samaniego, S., *J. Phys. Chem.*, A101, 4602, 1997.
660 38 Krishnan, R., Lee, J., and Robinson, G. W., *J. Phys. Chem.*, 94, 6365, 1990.
661 39 Pines, E., Tepper, D., Magnes, B.-Z., Pine, D., and Barak, T., *Ber. Buns. fur Phys. Chem.*, 102, 504,
662   1998.
663 40 Masad, A. and Huppert, D., *J. Phys. Chem.*, 96, 7324, 1992.
     Q2 41 Barak, T., Pines, D., and Pines, E., unpublished results.
     Q2 42 Magnes, B.-Z., Barak, T., Pines, D., and Pines, E., unpublished results.
         43 Barroso, S. J., Arnaut, L. G., and Formosinho, S. J., *J. Photchem. Photobiol.*, A154, 13, 2002.
         44 Politi, M. J., Brandt, O., and Fendler, J. H., *J. Phys. Chem.*, 89, 2345, 1985.

664   45 (a) Agmon, N., Pines, E., and Huppert, D. J., *Chem. Phys.*, 88, 5631, 1988; (b) Pines, E., Huppert, D.,
665   and Agmon, N., *Chem. Phys.*, 88, 5620, 1988; (c) Huppert, D., Pines, E., and Agmon, N., *J. Opt. Soc.*
666   *Am.*, B7, 1545, 1990; (d) Pines, E., Huppert, D., and Agmon, N., *J. Phys. Chem.*, 95, 666–674, 1991;
667   (e) Poles, E., Cohen, B., and Huppert, D., *Isr. J. Chem.*, 39, 347, 1999.
668 Q3 46 *Proton Transfer. Simposia of the Faraday Society No. 10*, The Faraday Division, Chemical Society,
669   London, 1975.
670   47 Bordwell, F. G. and Boyle, W. J. Jr., *JACS*, 97, 3447, 1975.
671   48 Sühnel, J., *J. Phys. Org. Chem.*, 3, 62–68, 1990.
672   49 Pines, E., Pines, D., Barak, T., Tolbert, L. M., and Haubrich, J. E., *Ber. Bunsen-Ges. Phys. Chem.*, 102,
673   511, 1998.
674   50 Pines, E. and Huppert, D., *J. Phys. Chem.*, 84, 3576, 1986.
674   51 Shiobara, S., Tajina, S., and Tobita, S., *Chem. Phys. Lett.*, 380, 673, 2003.
675 Q2 52 Barak, T. and Pines, E., unpublished results.

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Author Queries

Q1   Equation 16.22 is cited, but not given. Kindly check.

Q2   Kindly update Refs. [41,42,52].

Q3   Kindly provide complete author details for Ref. [46].