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Proton Transfer of Phenanthridone in the Lowest Excited Singlet State in Moderately Concentrated Mineral Acid Solutions

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A modified form of the steady state treatment for the determination of excited state proton transfer rate constants was successfully applied to a variety of oxygen and nitrogen center aromatic excited acids. These compounds shared the characteristic of being more acidic in the lowest excited singlet state (S_1) than in the ground electronic state (S_0) and of requiring a concentrated mineral acid medium for protonation. This treatment was extended to phenanthridone, the lactam of 6-hydroxyphenanthridine, which is a weak enough base in the ground and the excited states to require moderately concentrated mineral acids for protonation, and becomes a stronger base in the excited state than it is in the ground state. Phenanthridone exists as an α-lactam and is a weaker base in the ground and excited states than the α-lactams derived from 2-hydroxypyridine, 2-hydroxyquinoline, and 1-hydroxyisoquinoline. It is also a much weaker base than the vinylogous γ-lactams. The reasons for this are discussed here.

KEY WORDS: Phenanthridone; proton transfer; excited state; lactam; lactim.

INTRODUCTION

The kinetics of excited state proton transfer was first put on a quantitative basis by Weller, who related the pH dependences of the fluorescence efficiencies of conjugate acids and bases to the competitive kinetics of proton transfer in and photophysical deactivation of the lowest singlet state [1,2]. Weller's equations describing the excited state ionization of those fluorescing molecules which become more acidic in the excited state (proton transfer in S_1 is initiated by directly exciting the conjugate acid) and for which water is the proton acceptor are

$$\frac{\phi}{\phi_0} = \frac{1 + \alpha_b + k_b\tau_0'[H^+]}{1 + k_a\tau_0 + k_b\tau_0'[H^+]} \quad (1)$$

$$\frac{\phi'}{\phi_0'} = \frac{\alpha_b + k_a\tau_0}{1 + k_a\tau_0 + k_b\tau_0'[H^+]} \quad (2)$$

or combining Eqs. [1] and [2]

$$\frac{\phi/\phi_0}{\phi'/\phi_0' - \alpha_B} = \frac{1}{k_a\tau_0} + \frac{k_b\tau_0'[H^+]}{k_a\tau_0}\left(\frac{\phi'/\phi_0'}{\phi'/\phi_0' - \alpha_B}\right) \quad (3)$$

where ϕ and ϕ' are the fluorescence quantum efficiencies of the conjugate acid and base, ϕ_0 and ϕ_0' are the fluorescence quantum efficiencies of the conjugate acid and the conjugate base in the absence of excited state proton transfer, τ_0 and τ_0' are the fluorescence decay times of the conjugate acid and the conjugate base, in the absence of excited state proton transfer (i.e., at very low and very high pH), respectively. k_a and k_b are the rate constants for dissociation of the conjugate acid (pseudo first order) and protonation of the conjugate base (second order) respectively. α_B is the fraction of exciting light directly

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absorbed by the conjugate excited base and $[H^+]$ is the hydrogen ion concentration:

$$\alpha_B = \frac{\varepsilon_B K_a}{\varepsilon_B K_a + \varepsilon[H^+]}$$

where ε_B and ε are, respectively, the molar absorptivities of conjugate base and acid at the nominal wavelength of excitation and K_a is the prototropic dissociation constant of the ground state acid-base reaction.

Equations 1–3 are applicable only to proton transfers that occur in dilute aqueous solution where a_W, the activity of water, is close to unity. Thus, if the rate constants for proton transfer and/or the fluorescence decay times are such that the proton transfer occurs in concentrated mineral acid media (pH $<$ 0) where $a_W < 1$, then Eqs. 1–3 may not be used to quantitate the kinetics of the reaction. Schulman and Vogt [3–5] have used a transition state treatment to modify Eq. 3 to yield

$$\left(\frac{\phi/\phi_0}{\phi'/\phi_0' - \alpha_B}\right) a_W{}^r = \frac{1}{k_a(0)\tau_0} + \frac{k_b(0)\tau_0'}{k_a(0)\tau_0}\, h_0 a_W{}^n \left(\frac{\phi'/\phi_0'}{\phi'/\phi_0' - \alpha_B}\right) \quad (4)$$

where $k_a(0)$ and $k_b(0)$ are analogous to k_a and k_b, except that in concentrated acid k_a and k_b are medium dependent whereas $k_a(0)$ and $k_b(0)$ are not. In Eq. 4, h_0 = antilog$(-H_0)$ where H_0 is the Hammett acidity of the medium. The exponent n represents the number of water molecules which enter into the acid dissociation reaction of the Hammett indicator used to define that portion of the acidity scale in which occurs the excited state proton transfer of the aromatic acid or base of interest. The r represents the number of water molecules required to react with the excited conjugate acid to form the transition state species, which is common to both the excited state protonation and dissociation reactions. Vogt and Schulman have successfully applied Eq. 4 to the determination of rate parameters for molecules of differing charge type and acidity, and have extended the quantitation of excited state proton transfer kinetics in concentrated acid media to approximately 13 M sufuric acid [5]. For those compounds that become more basic in the excited state, proton transfer in S_1 is initiated by exciting the conjugate base and it is preferable to weight the fluorimetric titration data by rearranging Eq. 4 to

$$\left(\frac{\phi'/\phi_0'}{\phi/\phi_0 - \alpha_A}\right) h_0 a_W{}^n = \frac{1}{k_b(0)\tau_0} + \frac{k_a(0)\tau_0}{k_b(0)\tau_0'}\, a_W{}^r \left(\frac{\phi/\phi_0}{\phi/\phi_0 - \alpha_A}\right) \quad (5)$$

where α_A is the fraction of exciting light directly absorbed by the conjugate acid ($\alpha_A + \alpha_B = 1$) and corrects for the direct excitation of the conjugate acid in those cases where the ground and excited state inflection regions overlap.

The present work demonstrates the applicability of Eq. 5 to the excited state proton transfer reaction between phenanthridone (I) (6-hydroxyphenanthridine) and its conjugate acid (II) in aqueous sulfuric acid and shows that phenanthridone is more amide-like than its benzologous aromatic α-lactams, 2-pyridone, 2-quinolone, and 1-isoquinolone.

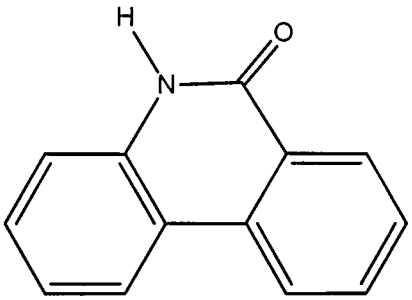
These kinds of studies allow a more complete understanding of the nature of acid-base reactions than is realizable in pure water where one of the reactants (water) does not overtly show its role in the reaction.

EXPERIMENTAL

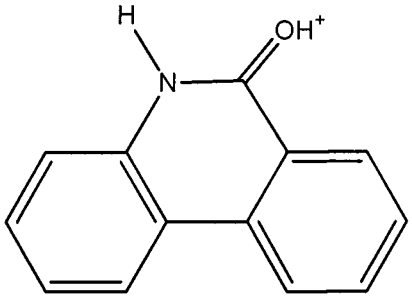
The phenanthridone used was purchased from Pfaltz & Bauer, Inc. (Flushing, NY). It was recrystallized from absolute ethanol. Purity was confirmed by thin layer chromatography on silica gel using 5 vol% MeOH in $CHCl_3$ as the mobile phase. A description of the commercial instrumentation used may be found elsewhere [3]. The fluorescence spectra were corrected for instrumental response by means of a Rhodamine B quantum counter.

We used 1.0×10^{-5} M solutions of phenanthridone for the UV-Vis absorption studies and for the determination of the fluorescence lifetimes. The solutions used for the fluorescence titrations were 1.0×10^{-6} M. Fluorescence was in each case excited at 288 nm, an isosbestic point in the cation-neutral absorption spectrum and monitored quantitatively at 375 nm.

Sulfuric acid solutions were made with analytical grade H_2SO_4 and deionized, distilled water. All acid solutions were checked for spurious emission before use. Values of a_w were taken from [6,7]. It should be noted that $k_a(0)$ and $k_b(0)$ for those compounds that become more basic in the excited state cannot always be determined from a single fluorescence titration [8]. It is a consequence of the variation of water activity during the course of the fluorimetric pH titrations of the very weak bases examined in the present study that allows $k_a(0)$ and $k_b(0)$ to be obtained through application of Eq. 5. However, in dilute solutions where the activity of water does not vary significantly from unity, when the pH of the solution is greater than the ground state pKa of the fluorophore such that $\alpha_A = 0$, the equation corresponding to Eq. 5 for this set of conditions is



(I)



(II)

Scheme 1.

$$\frac{\phi'/\phi_0'}{\phi/\phi_0}[H^+] = \frac{1 + k_a\tau_0}{k_b\tau_0'} \qquad (6)$$

from which $(1 + k_a\tau_0)/k_b\tau_0'$ may be determined, but k_a and k_b cannot be explicitly evaluated. In the latter case it is often necessary to perform a series of fluorimetric pH titrations, each in the presence of a different concentration of a quencher species, and a corresponding series of lifetime measurements in order for k_a and k_b to be extracted from a plot of

$$\frac{\phi'/\phi_0'}{\phi/\phi_0}[H]^+ \tau_0' \text{ vs. } \tau_0$$

The H_0 (Hammett acidity) data [9,10] was used over the applicable concentration range.

The fluorescence lifetimes of the neutral and cationic forms of phenanthridone were measured at pH 4.0 and H_0 −4.6, respectively.

RESULTS AND DISCUSSION

The acidity-dependent quantum yields of fluorescence ϕ of the cationic form of the title compound normalized with respect to ϕ_0, the quantum yield of the isolated cation (relative quantum yields of fluorescence), are shown as a function of Hammett acidity (H_0) in Figure 1. Table I presents the frequencies of the spectral maxima of the absorption and fluorescence bands chosen for use in the Förster cycle approximation of the excited state pKa* [11] as well as the ground state pKa of the protonated phenanthridone, determined spectrophotometrically and the pKa* calculated from the Förster cycle [13]. Fluorescence decay times of the neutral and cationic forms of each compound are also presented in Table I.

According to Eq. (5), a plot of

$$h_0 a_W{}^n \left(\frac{\phi'/\phi_0'}{\phi/\phi_0 - \alpha_A}\right) \text{ vs. } a_W{}^r \left(\frac{\phi/\phi_0}{\phi/\phi_0 - \alpha_A}\right)$$

should be linear and have a slope of $k_a(0)\tau_0/k_b(0)\tau_0'$ and an ordinate intercept of $1/k_b(0)\tau_0'$. Integral values of n and r were selected using the following criteria. (1) The values of n and r chosen resulted in a linear plot of the data. (2) The values of the kinetic parameters $k_a(0)$ and $k_b(0)$ obtained from Eq. 5 upon substitution of n and r should be reasonable. For example, according to Eigen's interpretation of the Brönsted relationship between equilibrium constants and rate parameters [12], the magnitude of the rate constant describing the reaction of the phenanthridone ion and water or between the hydrogen ion and phenanthridone should be no greater than the Debÿe rate constant for a diffusion controlled reaction (5×10^{10} $M^{-1}s^{-1}$). (3) When the assumptions inherent in the

Table I. Spectral Maxima ($\bar{\nu}_a$ for Long Wavelength Absorption, $\bar{\nu}_f$ for Fluorescence, and $\bar{\nu}_{0-0}$ for the Average of $\bar{\nu}_a$ and $\bar{\nu}_f$) and Fluorescence Decay Times (τ_0) of Phenanthridone (B) and Its Conjugate Acid (BH^+)

	BH^+	B	pKa^{BH+}
$\bar{\nu}_a$ ($cm^{-1} \times 10^{-4}$)	2.98 + 0.01	3.11 + 0.01	−2.27 + 0.02
$\bar{\nu}_f$ ($cm^{-1} \times 10^{-4}$)	2.68 + 0.01	2.75 + 0.01	
$\bar{\nu}_{0-0}$ ($cm^{-1} \times 10^{-4}$)	2.83 + 0.01	2.93 + 0.01	
τ_0 (ns)	13.2 + 0.2	5.9 + 0.2	

pKa* = −2.27 −2.10*(28.3–29.3) = −2.27 +2.10 = −0.17.

Also presented is the ground state pK_a determined spectrophotometrically in this work according to the relationship $pK_a = H_0 - \log \frac{[B]}{[BH^+]} - (n - r_g) \log a_w$, where r_g represents the number of water molecules consumed in the ground state acid dissociation of phenanthridone and n is the number of water molecules consumed in the dissociation of the Hammett indicator (s) defining the inflection region of the spectrophotometric titration of phenanthridone. Here $n - r_g = 0$.

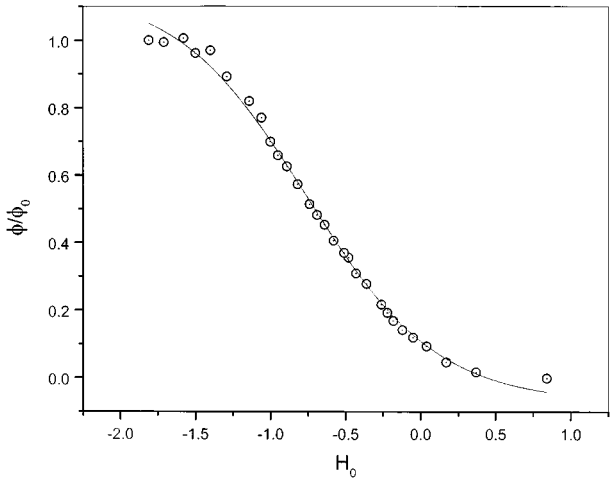


Fig. 1. Plot of ϕ/ϕ_0 of the phenanthridinium cation *vs* H_0 the Hammett acidity function in sulfuric acid.

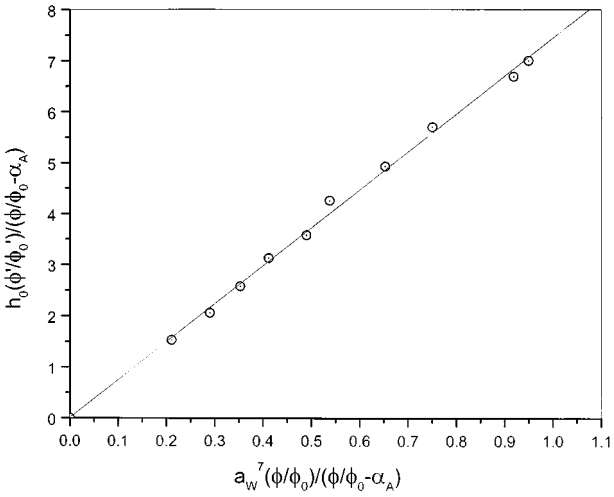


Fig. 2. Plot of $\frac{\phi'/\phi_0'}{\phi'/\phi_0' - \alpha_A} h_0$ vs. $\frac{\phi/\phi_0}{\phi'/\phi_0' - \alpha_A} \alpha_W{}^7$, where a_W is the activity of water.

Förster cycle approximation are valid [5,11], the excited state pK_a* values obtained from the kinetic analysis and the Förster cycle should be comparable if not equivalent.

A number of plots of

$$h_0 a_W{}^n \left(\frac{\phi'/\phi_0'}{\phi/\phi_0 - \alpha_A}\right) \text{ vs. } a_W{}^r \left(\frac{\phi/\phi_0}{\phi/\phi_0 - \alpha_A}\right)$$

for phenanthridone all had $n - r = 7$ and passed through or very close to the origin resulted. All other combinations of n and r resulted in nonlinear plots. This suggests that protonated phenanthridone dissociates so rapidly in the lowest excited singlet state ($k_a\tau_0$ is very large and, hence, $1/k_a\tau_0 \approx 0$) that equilibrium is attained in the lowest excited state so that the slope $k_a(0)\tau_0/k_b(0)\tau_0' = K_a{}^*\tau_0/\tau_0'$ permits determination of K_a* (and hence, Pk_a*) (Fig. 2), but $k_a(0)$ and $k_b(0)$ explicitly remain indeterminate. The value of pKa* so measured was found to be -0.36 ± 0.01, which agreed well with the value of pKa* $= -0.2 \pm 0.3$ calculated from the data in Table I using the Förster cycle [11].

Phenanthridone (6-hydroxylphenanthridine) is similar to other α- or γ-hydroxy *N*-heterocycles in that its uncharged form is tautomerized and electromerized to a lactam (14–18) (or a vinylogous lactam in the γ-hydroxy-*N*-heterocycles). The protonated lactam, in the ground electronic state, is characterized by a dissociation constant (K_a*) of the protonated form, several orders of magnitude higher than expected for the protonated hydroxyaromatic *N*-heterocycles [14]. The degree of lactamization (localization of charge on nitrogen and high bond order of the C−O bond) of hydroxy *N*-heterocycles appears to be a function of whether the compound is an α- or a γ-lactam and the size of the aromatic system. For example, protonated 4-hydroxypyridine is a weaker acid (more phenolic) than protonated 4-hydroxyquinoline, which in turn is weaker acid than protonated 9-hydorxyacridine (acridone) [14]. Protonated 2-hydroxypyridine is a weaker acid than either protonated 2-hydroxyquinoline or protonated 1-hydroxyisoquinoline and all three of these are stronger acids than the corresponding vinylogous lactams [14]. This is understandable if one considers that in the hydroxypyridine, lactamization is energetically less favorable because it destroys the aromatic sextet of the pyridine ring while in the hydroxyquinolines, lactamization allows the preservation of one aromatic sextet. The greater tendency to lactamization in the α-compounds is undoubtedly a proximity effect. Interestingly, 2- and 4-hydroxyquinolines and 1-hydroxyisoquinoline [4], in the lowest excited singlet states, behave somewhat as if they were phenolic zwitterions, rather than as lactams. This is manifested in the lower pKa* values (pKa* is the logarithm of the reciprocal of the dissociation constant in the lowest excited singlet state) of the excited cations relative to those of the corresponding ground state species. This is typical of phenolic and aromatic *N*-heterocyclic behavior [2,4,8]. The hydroxypyridines have lowest excited singlet states that are too short-lived to allow direct measurement of these kinetic and/or thermodynamic parameters. However, protonated acridone has a pKa* that is less acidic (higher) than that of its ground electronic state [8], behavior that is more reminiscent of amides (lactams) than of *N*-heterocyclic phenols [4,5,8]. The tendency for *N*-heterocyclic phenols to act more amide-like in larger fused aromatic systems than in smaller ones is undoubtedly related to the decrease in

aromaticity, per aromatic ring, that occurs as more fused rings are added to an aromatic system. This is also manifested in the increased ease of oxidation as one goes from benzene through the successively higher arenes. As phenanthrene derivatives generally seem to have electronic spectra intermediate between those of the corresponding naphthalene and anthracene derivatives, it might be expected that the hydroxy derivatives, particularly the 6-hydroxy derivative of phenanthridone would demonstrate spectral and chemical properties intermediate between those of the 2- and 4-hydroxyquinolines, on the one hand, and those of 9-hydroxyacridine (acridone), on the other. However, the very low pKa of the ground electronic state of the phenanthridonium cation that becomes more positive in the lowest excited singlet state suggests that the electronic distributions in phenanthridone in both electronic states are much closer to those of "true" arylamides [11–21] than any of the aforementioned lactams. By comparison with the other α-lactams, the lactamization of 6-phenanthridol to phenanthridone requires the disruption of a single aromatic ring while preserving two in the molecule. This probably allows the unsaturation between the 4 and 5 carbon atoms of the phenanthridine ring or between the 6 carbon atom and the nitrogen atom of the phenanthridine ring to demonstrate more olefinic (or iminic) rather than aromatic character. It is then to be expected that the higher *N*-heterocycles (those with more fused aromatic rings) will have even less delocalization energy in the inner rings and their α- and γ-hydroxy derivatives should behave even more like "pure" amides with little or no retained phenolic character in either ground or lowest excited singlet states. These studies suggest that the concept of aromaticity is equally validly applied to the lowest singlet state as to the ground state.

It should be mentioned that an attempt was made to study the acidity of uncharged phenanthridone in the ground and lowest excited singlet states. Although it was possible to spectrophotometrically determine the pK_a of the neutral molecule at 25°C ($pK_{BH} = 12.82 \pm 0.01$), dynamic proton transfer between the photo-excited neutral molecule and anion was not observed. pK_{BH}* was estimated from pK_{BH} and the absorption and fluorescence maxima used in conjunction with the Förster cycle as $pK_{BH}* = 9.9 \pm 0.3$, suggesting that although the carboxamido oxygen accepts negative charge in S_1, the opposite is true for the carboxamide nitrogen atom that becomes less basic (or more acidic) in the lowest excited singlet state.

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