another and their elimination would lead to the corresponding carbazole. At the applied electrolysis potential the carbazole would be oxidized on to its cation radical, so the overall process has an n value of 3.

It is interesting to note that this type of reaction may be fairly general in organic electrooxidations. It has recently been shown that the electrochemical oxidation of tetraarylethylenes leads to the formation of substituted phenanthrenes.[28,29] Since the arylethylene dications are isoelectronic with the arylamine dications, the same mechanism may be operative, i.e., a concerted electrocyclic ring closure (eq 4 and 5).

In summary, then, it would appear that these electrochemical data can do relatively little to further clarify the photolytic mechanism. However, the following points can be forwarded.

(1) The photochemical process cannot go through the amine cation radical; according to the electrochemical studies if the cation radical were implicated some benzidines should be found (unless there is a tremendous solvent effect between polar and nonpolar media).

(2) The photochemical process may go through the amine dication, but again the role of solvent has not been clarified.

(3) The cyclization reaction is general for di- and tri-

arylamines and is not limited to molecules containing roughly neutral or electron-donating substituents as implied in the collected photochemical studies. However, this is not to say that all the molecules converted electrochemically could also be cyclized photochemically with higher energy radiation.

Extensive electrochemical studies in our laboratory and by others have shown that the following conditions are necessary for the electrochemical cyclization reaction.

(1) Only tertiary aromatic amines can be cyclized where at most one alkyl group can be present and all aromatic rings must be substituted in the positions para to the amine nitrogen. In addition, these substituents must not be susceptible to substitution or elimination.

(2) The reactions must be carried out in nonaqueous media since only here can oxidation potentials be applied such that the amines can be oxidized to the dications.

Further studies are now in progress to achieve better yields in the electrochemical oxidations.

(28) J. D. Stuart and W. E. Ohnesorge, *J. Electroanal. Chem.*, 30, App. 11 (1971).
(29) J. D. Stuart and W. E. Ohnesorge, *J. Amer. Chem. Soc.*, 93, 4531 (1971).

Intermediates in Nucleophilic Aromatic Substitution. XI.[1] Spiro Meisenheimer Complex of N,N'-Dimethyl-N-picrylethylenediamine. Partially Rate-Limiting Proton Transfer of Complex Formation. A Temperature-Jump Study

Claude F. Bernasconi*[2] and Constantin L. Gehriger

Contribution from the Thimann Laboratories of the University of California, Santa Cruz, California 95064. Received June 30, 1973

Abstract: N,N'-Dimethyl-N-picrylethylenediamine (AH) is in equilibrium with the spiro Meisenheimer complex (XH and X⁻) formed by intramolecular nucleophilic attack on the 1 position of the picryl moiety (eq 1). The rate of equilibration between the starting material and the two forms of the complex, as determined by the temperature-jump technique, not only shows the expected pH dependence but also increases with buffer concentration. The buffer dependence is curvilinear and reaches a plateau at high concentration. This indicates a change from rate-limiting proton transfer at a low concentration to rate-determining nucleophilic attack at high concentration and is a consequence of the unexpectedly high rate constant of 2×10^5 sec^{-1} for the reversion of the zwitterion (XH) to N,N'-dimethyl-N-picrylethylenediamine. These results suggest that in some cases general base catalysis in nucleophilic aromatic substitutions by amines is a manifestation of slow proton transfer rather than of sluggish leaving group departure.

I n our attempts to deal with the problem of the rates at which intermediates in activated nucleophilic aromatic substitution reactions expel various mobile groups from their sp³ hybridized carbon, we have frequently chosen the reversible interaction of 1,3,5-trinitrobenzene (TNB) with various nucleophiles as a model reaction.[3]

The choice of TNB as a model substrate offers both

(1) (a) Part X: C. F. Bernasconi, R. H. deRossi, and C. L. Gehriger, *J. Org. Chem.*, 38, 2838 (1973). (b) Presented at 12th Pacific Conference on Chemistry and Spectroscopy, San Diego, Calif., Nov 1973.
(2) Alfred P. Sloan Fellow, 1971–1973.

(3) (a) C. F. Bernasconi, *J. Amer. Chem. Soc.*, 92, 129 (1970); (b) C. F. Bernasconi, *ibid.*, 92, 4682 (1970); (c) C. F. Bernasconi and R. G. Bergstrom, *J. Org. Chem.*, 36, 1326 (1971); (d) C. F. Bernasconi and R. G. Bergstrom, *J. Amer. Chem. Soc.*, 95, 3603 (1973).
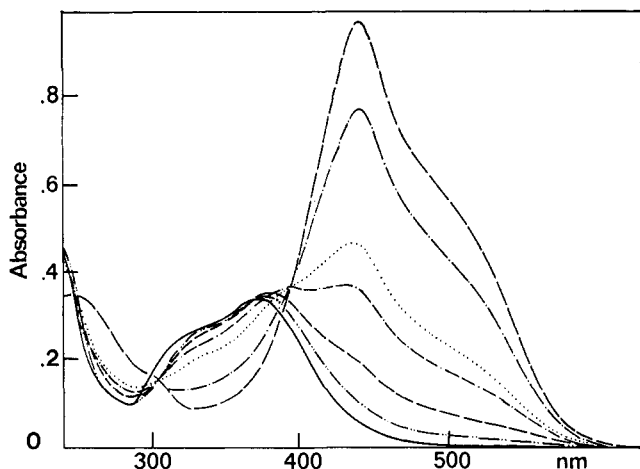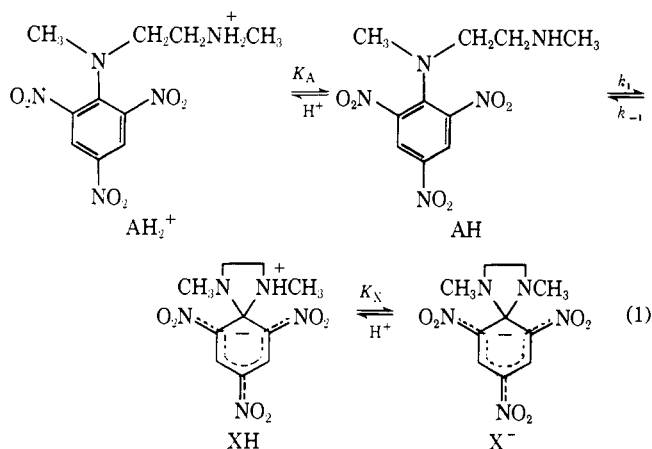
Figure 1. Absorption spectra of AH_2^+ at various pH values in H_2O ($[AH_2^+]_0 = 4 \times 10^{-5}\ M, \mu = 0.5\ M$): (——) pH 6.52; (–···–) pH 8.16; (– – –) pH 8.50; (– –·– –·) pH 8.80; (.....) pH 8.95; (–·–·) pH 9.47; (————) pH 12.14.

Figure 2. Representative plots of τ^{-1} vs. total buffer concentration: (O) pH 8.60, $[AH_2^+]_0 = 4 \times 10^{-4}\ M$; (□) pH 8.00, $[AH_2^+]_0 = 4 \times 10^{-3}\ M$; (△) pH 7.80, $[AH_2^+]_0 = 4 \times 10^{-3}\ M$.

advantages and drawbacks. The absence of a nucleofugic leaving group reduces the complexity of the kinetic system since one has to deal with one equilibrium process only, though possibly coupled with an acid-base equilibrium in the case of amine nucleophiles.

The intramolecular Meisenheimer complex formation of N,N'-dimethyl-N-picrylethylenediamine (AH), eq 1, though still not a "perfect" model reaction, goes

far toward overcoming the disadvantages of TNB since the CH_3–N–CH_2 moiety mimics the electronic and steric effects on k_1, k_{-1}, and K_X exerted by a typical alkoxy leaving group much better than does hydrogen.

The fact that the reaction is intramolecular is a possible drawback; however, this factor is likely to affect k_1 the most whereas the effect on k_{-1} and K_X, the quantities we are principally interested in, should be minor. The choice of an intramolecular system was dictated by the greater stability of the complex and the wish to avoid complications due to competing nucleophilic attack at the 3 position.[4]

In conformity with widely accepted ideas about the detailed mechanism of base catalysis in nucleophilic aromatic substitutions (see Discussion) and some of our own observations regarding the interaction of TNB with aliphatic amines,[3a] we started out by assum-

(4) See, e.g. (a) V. Gold and C. H. Rochester, J. Chem. Soc., 1697, 1727 (1964); (b) M. R. Crampton and V. Gold, Proc. Chem. Soc., London, 298 (1964).
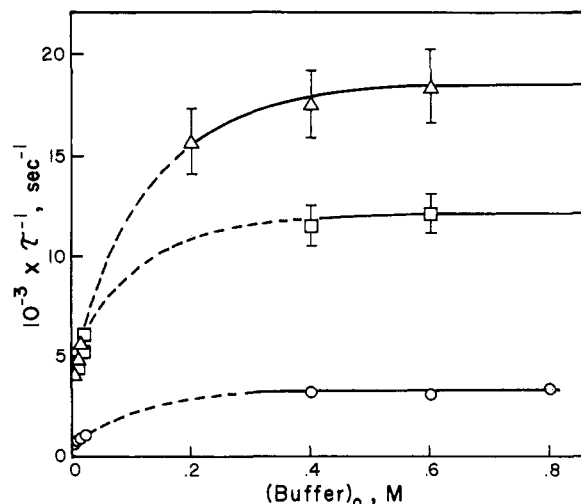
ing that the proton transfer equilibria in eq 1 would under all experimental conditions be rapidly established compared to the steps symbolized with k_1 and k_{-1}. However, we found that below pH ~10 the formation of X^- is significantly general base catalyzed indicating that proton removal from XH is rate limiting.
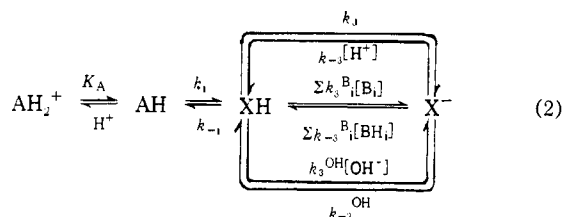
Results

General Features. When base is added to an aqueous solution of N,N'-dimethyl-N-picrylethylenediamine hydrochloride (AH_2^+), there is an immediate formation of a red species with the typical absorption spectrum of a Meisenheimer complex (XH + X^-) as shown in Figure 1. At pH ≥ 12 complex formation appears to be quantitative. The structure of X^- was confirmed by pmr spectroscopy in DMSO solution (see Experimental Section).

Buffered aqueous solutions at varying pH values were subjected to temperature jumps of 3° (end temperature 25°). The ensuing chemical relaxation effect was monitored spectrophotometrically at 546 or 365 nm. We were able to cover a range from pH 7.6 to 10.5. Though measurements at pH <7.6 would have been desirable, the equilibrium concentration of the complex which is limited by the solubility of AH_2^+ becomes too small for monitoring the chemical relaxation.

One single relaxation time was observed under all experimental conditions. It not only depends strongly on the pH as one would expect for a system such as eq 1, but at pH ≤ 9.5 also on the chemical nature and the concentration of the buffer as well as on the substrate concentration. The data are summarized in Tables I and II.

Most revealing is the *curvilinear* dependence of the reciprocal relaxation time on buffer concentration as shown in some representative plots in Figure 2. This is a clear indication that buffer catalysis cannot be due to a concerted process and that we are dealing with a change in the rate-determining step in the reaction mechanism as the buffer concentration is increased.

Mechanism. We shall show that our data are consistent with the mechanism of eq 2; k_3 and k_3^{OH} refer to the rate coefficients for proton abstraction from XH

$$AH_j^+ \underset{H^+}{\overset{K_A}{\rightleftharpoons}} AH \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} XH \overset{k_{-3}[H^+]}{\underset{\Sigma k_3^{B_i}[B_i]}{\rightleftharpoons} \underset{\Sigma k_{-3}^{B_i}[BH_i]}{\overset{k_3^{OH}[OH^-]}{\rightleftharpoons}} } X^- \qquad (2)$$

by the solvent and by hydroxide ion, respectively, whereas the $k_3^{B_i}$'s are the rate coefficients for proton transfer to any general base present, including the buffer base, as well as X^- and AH. Similarly k_{-3} and k_{-3}^{OH} refer to the rate coefficients of protonation of X^- by the hydronium ion and the solvent, respectively, whe eas the $k_{-3}^{B_i}$'s refer to the protonation by general acids such as the buffer acid, AH_2^+, and XH.

Let us assume that $k_{-1} + k_3 + k_3^{OH}[OH^-] + \Sigma k_3^{B_i} \cdot [B_i] \gg k_1$ and $k_{-3}[H^+] + k_{-3}^{OH} + \Sigma k_{-3}^{B_i}[BH_i]$.[5] This assumption which will be justified in the Discussion allows XH to be treated as a steady-state intermediate; the reaction rate in terms of X^- can then be expressed by

$$d[X^-]/dt = k_f[AH] - k_b[X^-] \qquad (3)$$

with

$$k_f = \frac{k_1(k_3 + k_3^{OH}[OH^-] + \Sigma k_3^{B_i}[B_i])}{k_{-1} + k_3 + k_3^{OH}[OH^-] + \Sigma k_3^{B_i}[B_i]} \qquad (4)$$

$$k_b = \frac{k_{-1}(k_{-3}[H^+] + k_{-3}^{OH} + \Sigma k_{-3}^{B_i}[BH_i])}{k_{-1} + k_3 + k_3^{OH}[OH^-] + \Sigma k_3^{B_i}[B_i]} \qquad (5)$$

If we assume further that AH_2^+ and AH are always in rapid equilibrium with respect to the time scale of the processes governed by k_f and k_b (the forward and backward rate constants, respectively), the reciprocal relaxation time derived according to standard procedures[6] is given by

$$\tau^{-1} = k_f \frac{K_A}{K_A + [H^+]} + k_b \qquad (6)$$

Kinetics at High Buffer Concentrations With respect to eq 6 two limiting situations are of particular interest. In the first the deprotonation of XH is much more rapid than its reversion to AH, i.e., $k_3 + k_3^{OH} \cdot [OH^-] + \Sigma k_3^{B_i}[B_i] \gg k_{-1}$. This situation is achieved at high buffer base concentrations (plateaus in Figure 2) and/or at high pH. Equation 6 simplifies to eq 7

$$\tau_{hi}^{-1} = k_1 \frac{K_A}{K_A + [H^+]} + k_{-1}\frac{[H^+]}{K_X} \qquad (7)$$

where "τ_{hi}^{-1}" refers to high base concentration and K_X is the acid dissociation constant of XH.

One can determine the parameters k_{-1}/K_X, k_1, and K_A from the τ_{hi}^{-1} values (equivalent to τ^{-1} at the highest buffer concentration employed; see Table I) as follows. At pH ≤ 8.6 a plot of τ_{hi}^{-1} vs. $[H^+]$ (not shown) is a straight line, indicating that $\tau_{hi}^{-1} \approx k_{-1}[H^+]/K_X$.[7] The slope is $k_{-1}/K_X = 8.4 \times 10^{11} M^{-1} \sec^{-1}$.

(5) We will use the symbol "$\gg$" to indicate "at least tenfold larger" throughout this paper.

(6) M. Eigen and L. DeMaeyer in A. Weissberger, "Technique of Organic Chemistry," Vol. VIII, Part 2, Wiley-Interscience, New York, N. Y., 1963, p 895.

(7) [H⁺] was calculated from the hydrogen ion activity as determined by the pH meter by means of the relation [H⁺] $= a_{H^+}/\gamma_{H^+}$ where γ_{H^+} = 0.74 is the trace activity coefficient in 0.5 M NaCl.[8]

(8) H. S. Harned and R. A. Robinson in "Multicomponent Electrolyte Solutions," Permagon Press, Long Island City, N. Y., 1968, p 50.

Table I. Reciprocal Relaxation Times between pH 7.8 and 10.4 at High Buffer Concentrations[a]

pH[b]	[Buff]$_0$,[c] M	τ^{-1}, sec^{-1}
7.80[d]	0.2	15,700 ± 1600
	0.4	17,500 ± 1700
	0.6	18,400 ± 1800[f]
8.00[d]	0.4	11,500 ± 1000
	0.6	12,100 ± 1000[f]
8.20[d]	0.2	7,020 ± 500
	0.4	7,300 ± 500[f]
8.60[d]	0.4	3,200 ± 200
	0.6	3,100 ± 200
	0.8	3,400 ± 200[f]
9.00[d]	0.15	1,710 ± 85
	0.4	1,840 ± 90
	0.6	1,890 ± 90[f]
9.40[d]	0.4	1,360 ± 70
	0.6	1,290 ± 70
	0.8	1,350 ± 70[f]
9.80[e]	0.2	1,180 ± 60
	0.28	1,210 ± 60
	0.36	1,250 ± 60[f]
10.00[e]	0.2	1,200 ± 80
	0.28	1,220 ± 80
	0.36	1,230 ± 80[f]
10.40[e]	0.12	1,170 ± 80
	0.20	1,200 ± 80
	0.28	1,210 ± 80[f]

[a] At 25 ± 0.5°, μ = 0.5 M maintained by addition of NaCl, substrate concentration $[AH_2^+]_0 = 2 \times 10^{-4}$ (high pH) to 3×10^{-3} M (low pH). [b] pH as measured by glass electrode. [c] [Buff]$_0$ = total buffer concentration. [d] Tris buffer. [e] Carbonate buffer. [f] Equivalent to τ_{hi}^{-1}, see text.

As is apparent from Table I, at pH ≥ 10.0 τ_{hi}^{-1} is no longer dependent on the pH, indicating $K_A \gg [H^+]$ and $k_1 \gg k_{-1}[H^+]/K_X$; thus τ_{hi}^{-1} (pH ≥ 10.0) $= k_1$ = 1200 sec^{-1}.

Rearranging eq 7 to eq 8 allows one to solve for K_A;

$$K_A = \frac{(\tau_{hi}^{-1} - k_{-1}[H^+]/K_X)[H^+]}{k_1 + k_{-1}[H^+]/K_X - \tau_{hi}^{-1}} \qquad (8)$$

at pH 8.6, 9.0, and 9.4 one calculates K_A = 2.87, 2.21, and 1.60×10^{-9} M, average K_A = 2.23×10^{-9} or pK_A = 8.65.[9]

Equilibrium Measurements. Above 500 nm XH and X^- are practically the sole contributors to light absorption (Figure 1). Thus for 1-cm cuvettes Beer's law is given by eq 9. By expressing $[XH]$ and $[X^-]$ in terms

$$OD = \epsilon_{XH}[XH] + \epsilon_{X^-}[X^-] \qquad (9)$$

of the analytical concentration of AH_2^+ and the respective equilibrium constants ($K_1 = k_1/k_{-1}$) one obtains

$$\frac{OD}{[AH_2^+]_0} =$$

$$\frac{\epsilon_{XH}K_AK_1[H^+] + \epsilon_{X^-}K_AK_1K_X}{[H^+]^2 + K_A[H^+] + K_AK_1[H^+] + K_AK_1K_X} \qquad (10)$$

At low pH where there is little XH and X^- in equilibrium with AH_2^+ and AH, the $[H^+]^2$ term in the denominator is dominant and eq 10, after rearrangement, simplifies to eq 11. A plot of the left-hand side of eq

(9) Note that K_A is only calculated at pH values between 8.6 and 9.4 where τ_{hi}^{-1} depends strongly on K_A; at pH <8.6 the numerator, at pH >9.4 the denominator, in eq 8 becomes a small difference between two large numbers.
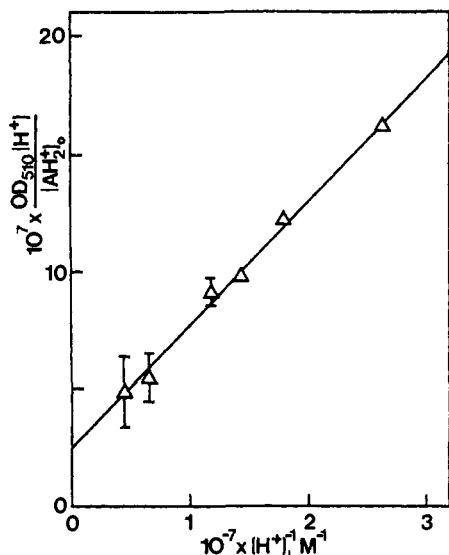
Figure 3. Determination of K_X from absorption measurements; plot according to eq 11.

Figure 4. τ_{no}^{-1} as a function of pH.

$$\frac{OD[H^+]}{[AH_2^+]_0} = \epsilon_{XH}K_AK_1 + \frac{\epsilon_{X^-}K_AK_1K_X}{[H^+]} \tag{11}$$

11 vs. $[H^+]^{-1}$[17] should yield a straight line with intercept $= \epsilon_{XH}K_AK_1$ and slope $\epsilon_{X^-}K_AK_1K_X$.

Such a plot is shown in Figure 3. Due to the rather low OD readings at the lowest pH values the experimental uncertainty is relatively high as indicated in the figure. From the ratio slope/intercept we can estimate K_X since ϵ_{XH} and ϵ_{X^-} are expected to be very similar though probably not identical;[10] one thus obtains K_X

$$\frac{slope}{intercept} = \frac{\epsilon_{X^-}}{\epsilon_{XH}}K_X \approx K_X \quad (\epsilon_{XH} \approx \epsilon_{X^-}) \tag{12}$$

$= 2.29 \times 10^{-7}M$ or $pK_X = 6.64$.

Combining K_X with $k_{-1}/K_X = 8.4 \times 10^{11}\ M^{-1}\ sec^{-1}$ (from slope of τ_{hi}^{-1} vs. $[H^+]$ at low pH) we calculate $k_{-1} = 1.93 \times 10^5\ sec^{-1}$, and further from $k_1 = 1200\ sec^{-1}$ (from τ_{hi}^{-1} at pH ≥ 10.0) one obtains $K_1 = k_1/k_{-1} = 6.21 \times 10^{-3}$. Finally from the intercept of Figure 3 we calculate $K_A = (\text{intercept})/K_1\epsilon_{X^-}^{-11} = 2.82 \times 10^{-9}$ M which is in good agreement with $K_A = 2.23 \times 10^{-9}$ obtained from τ_{hi}^{-1} via eq 8.

Kinetics at Low Buffer Concentrations. The second important limiting situation with regard to eq 6 obtains in the absence of general acids and bases; eq 6 simplifies to eq 13. τ_{no}^{-1} is obtained by extrapolating τ^{-1}

$$\tau_{no}^{-1} = \frac{k_1(k_3 + k_3^{OH}[OH^-])}{k_{-1} + k_3 + k_3^{OH}[OH^-]}\frac{K_A}{K_A + [H^+]} + \frac{k_{-1}(k_{-3}[H^+] + k_{-3}^{OH})}{k_{-1} + k_3 + k_3^{OH}[OH^-]} \tag{13}$$

to zero buffer and zero substrate concentration. This was achieved by measuring the buffer dependence of τ^{-1} at two different substrate concentrations. Extrapolation of the ordinate intercepts of such plots to zero substrate concentration yields τ_{no}^{-1} which is plotted vs. pH in Figure 4.

(10) C. F. Bernasconi, J. Org. Chem., **35**, 1214 (1970).

(11) Assuming ϵ_{XH} 13,400 at 510 nm as measured in a solution of pH 12.
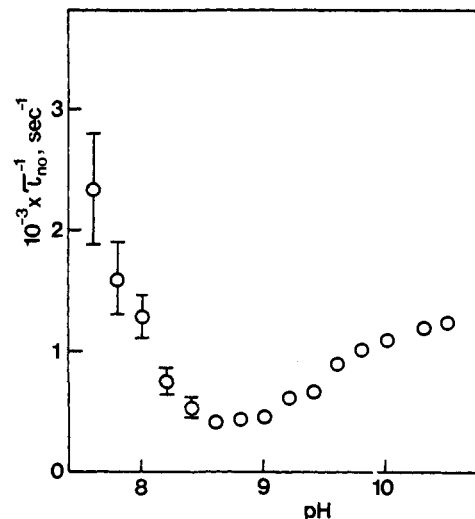
Note that the left-hand side of the plot is dominated by the second term in eq 13, the right-hand side by the first term. The right-hand side approaches a plateau value equal to $k_1 = 1200\ sec^{-1}$ ($k_3^{OH}[HO^-] + k_3 \gg k_{-1}$).

The data of Figure 4, in combination with k_1, k_{-1}, K_A, and K_X, allow us to obtain k_3, k_{-3}, k_3^{OH}, and k_{-3}^{OH} as follows: $5 \times 10^{10}\ M^{-1}\ sec^{-1}$ is an upper limit to be expected for k_{-3}.[12] From this we calculate the upper limit for $k_3 = K_Xk_{-3} \geq 1.05 \times 10^4\ sec^{-1}$. This is much smaller than $k_{-1} = 1.93 \times 10^5\ sec^{-1}$ and consequently eq 13 can be simplified to eq 14.

$$\tau_{no}^{-1} = \frac{k_1(k_3 + k_3^{OH}[OH^-])}{k_{-1} + k_3^{OH}[OH^-]}\frac{K_A}{K_A + [H^+]} + \frac{k_{-1}(k_{-3}[H^+] + k_{-3}^{HO})}{k_{-1} + k_3^{OH}[OH^-]} \tag{14}$$

The reaction of XH with OH⁻ may or may not be diffusion controlled, depending on whether there is significant intramolecular hydrogen bonding to one of the ortho nitro groups or not.[13] Thus, $k_3^{OH} \leq 10^{10}\ M^{-1}\ sec^{-1}$[12] and $k_{-3}^{OH} = k_3^{OH}K_W/K_X \leq 940\ sec^{-1}$.[14] The fact that the plot of τ_{no}^{-1} levels off ($k_3^{OH}[OH^-] \gg k_{-1}$) at a pH as low as 10.4 (Figure 4) ($[OH^-] = 4.9 \times 10^{-4}$[14]) shows that $k_3^{OH} \gg k_{-1}/[OH^-] = 3.7 \times 10^8\ M^{-1}\ sec^{-1}$; thus k_3^{OH} cannot be far below the upper limit of 10^{10} $M^{-1}\ sec^{-1}$. From this we conclude that $k_3^{OH}[OH^-] \gg k_3$ and $k_{-3}^{OH} \gg k_{-3}[H^+]$ at pH ≥ 9.2 which allows a further simplification of eq 14 to eq 15. Substituting

$$\tau_{no}^{-1}(pH \geq 9.2) = \frac{k_1 k_3^{OH}[OH^-]}{k_{-1} + k_3^{OH}[OH^-]}\frac{K_A}{K_A + [H^+]} + \frac{k_{-1}k_{-3}^{OH}}{k_{-1} + k_3^{OH}[OH^-]} \tag{15}$$

$K_Xk_{-3}^{OH}/[H^+]$ for $k_3^{OH}[OH^-]$ and solving for k_{-3}^{OH} affords eq 16. Applying eq 16 to our data at pH 9.2,

(12) M. Eigen, W. Kruse, G. Maass, and L. De Maeyer, Progr. React. Kinet., **2**, 287 (1964).

(13) C. F. Bernasconi, J. Phys. Chem., **75**, 3636 (1971).

(14) $K_W = [H^+][OH^-] = 1.96 \times 10^{-14}$ in 0.5 M NaCl[15] at 25°.

(15) H. S. Harned and B. B. Owen, "The Physical Chemistry of Electrolytic Solutions," Reinhold, New York, N. Y., 1950, p 486.

Table II. Reciprocal Relaxation Times between pH 7.6 and 10.5 at Low Buffer Concentration[a]

pH[b]	$10^{-3}\cdot$ $[AH_2^+]_0$, M	$[Buff]_0^d$ $= 0.005$ M	$[Buff]_0^d$ $= 0.01$ M	$[Buff]_0^d$ $= 0.015$ M	$[Buff]_0^d$ $= 0.02$ M
7.62[e]	1.0	11,450	14,700		17,200
7.60[f]	2.0	3,740	4,440		5,730
	4.0	4,520	5,380	5890	6,590
7.80[f]	1.0	2,800	4,050	4700	
	3.0	4,080	4,650	5640	
8.00[f]	1.0	2,500	3,230	3840	4,720
	4.0		4,480	5250	6,080
8.20[f]	1.0	1,860	2,380	2860	
	3.0	3,080	3,600	4150	
8.40[g]	0.1	620	740	990	1,040
	0.4	774	880	1010	1,140
8.60[g]	0.1	575	675	810	946
	0.4	720	834	952	1,050
8.80[g]	0.1	580	655	790	907
	0.4		820	902	
9.00[g]	0.1	571	650	800	881
	0.4	640	685		
9.20	0.02	643	760	820	
	0.2	815	878	933	999
9.40[g]	0.1	746	780	840	911
	0.4	862	975	1010	1,025
9.60[g]	0.1	880	956	965	1,060
	0.4		957	975	
9.80[g]	0.4		963	995	983
10.00[h]	0.3	1,080	1,175	1275	
10.30[h]	0.4	1,100	1,230	1180	1,170
10.50[h]	0.8	1,220		1260	1,230

[a] At $25 \pm 0.5°$, $\mu = 0.5$ M maintained by addition of NaCl. [b] pH as measured by glass electrode. [c] Estimated uncertainty in τ^{-1}: $\pm10\%$ at pH 7.60–7.80, $\pm8\%$ at pH 8.00–8.40, $\pm5\%$ at pH 8.60–9.80, $\pm8\%$ at pH 9.80–10.30, $\pm10\%$ at pH 10.50. [d] $[Buff]_0 =$ total buffer concentration. [e] Phosphate buffer. [f] Tris buffer. [g] Borate buffer. [h] Carbonate buffer.

$$k_{-3}^{OH} =$$

$$\frac{\tau_{no}^{-1}k_{-1}(K_A + [H^+])[H^+]}{k_1K_AK_X + k_{-1}(K_A + [H^+])[H^+] - \tau_{no}^{-1}K_X(K_A + [H^+])}$$
(16)

9.4, and 9.6[16] yields $k_{-3}^{OH} =$ 379, 392, and 572 sec^{-1}, respectively, average 445 sec^{-1}. Further we obtain $k_3^{OH} = k_{-3}^{OH}K_X/K_W = 5.20 \times 10^9$ M^{-1} sec^{-1}.

With this result we see that at pH ≤ 8.2 the first term in eq 14 becomes negligible and also $k_3^{OH}[OH^-] \ll k_{-1}$. Thus eq 14 reduces to

$$\tau_{no}^{-1} \text{(pH } \leq 8.2) = k_{-3}[H^+] + k_{-3}^{OH}$$
(17)

A plot of τ_{no}^{-1} vs. $[H^+]$ at pH ≤ 8.2 (not shown) affords a straight line with a slope $k_{-3} = 5.9 \times 10^{10}$ M^{-1} sec^{-1} and an intercept $k_{-3}^{OH} = 347$ sec^{-1}. Within experimental error this latter value agrees with the one obtained from eq 16 at pH 9.2–9.6 (445 sec^{-1}). Finally we calculate $k_3 = K_Xk_{-3} = 1.35 \times 10^4$ sec^{-1}.

Table III summarizes the parameters determined from kinetic and equilibrium measurements.

Buffer and Substrate Catalysis. From the data in Table II we can now also calculate k_3^{Bi} and k_{-3}^{Bi} for the various buffer and substrate components catalyzing the proton transfer between XH and X$^-$. At low concentrations the dependence of τ^{-1} on buffer and sub-

(16) At pH >9.6 τ_{no}^{-1} becomes insensitive to k_3^{OH} and k_{-3}^{OH} with the consequence that the denominator in eq 16 becomes a small difference between two large numbers.

Table III. Rate and Equilibrium Constants for the Reactions of Equation 2

k_1, sec^{-1}	$1.20 \pm 0.08 \times 10^3$
k_{-1}, sec^{-1}	$1.93 \pm 0.40 \times 10^5$
$K_1 = k_1/k_{-1}$	$6.21 \pm 1.50 \times 10^{-3}$
K_X, M	$2.29 \pm 0.30 \times 10^{-7}$ [a]
pK_X	6.64 ± 0.05[a]
K_A, M	$2.23 \pm 0.22 \times 10^{-9}$ [b,c]
	$(2.82 \pm 0.43 \times 10^{-9})$[b,d]
pK_A	8.65 ± 0.04[b,c]
	(8.55 ± 0.06)[b,d]
k_{-3}^{OH}, sec^{-1}	$4.45 \pm 0.65 \times 10^2$ [e]
	$(3.47 \pm 0.70 \times 10^2)$ [f]
$k_{-3}^{H_2O} = k_{-3}^{OH}/55.6$, M^{-1} sec^{-1} [j]	8.0 ± 1.2
k_3^{OH}, M^{-1} sec^{-1}	$5.20 \pm 1.04 \times 10^9$ [g]
	$(4.05 \pm 0.81 \times 10^9)$ [h]
k_{-3}, M^{-1} sec^{-1}	$5.90 \pm 2.40 \times 10^{10}$
k_3, sec^{-1}	$1.35 \pm 0.65 \times 10^4$ [i]
$k_3^{H_2O} = k_3/55.6$, M^{-1} sec^{-1} [j]	$2.43 \pm 1.1 \times 10^2$

[a] Defined as $K_X = [H^+][X^-]/[XH]$; the statistically corrected values are $K_X = 4.58 \times 10^{-7}$ or $pK_X = 6.34$; if K_X is defined as $a_H^+[X^-]/[XH]$ the statistically uncorrected (and corrected) values are: $K_X = 1.70 \times 10^{-7}$ (3.40×10^{-7}), $pK_X = 6.77$ (6.47). [b] Defined as $K_A = [H^+][AH]/[AH_2^+]$. [c] If defined as $K_A = a_H^+[AH]/[AH_2^+]$: $K_A = 1.65 \times 10^{-9}$, $pK_A = 8.78$; values obtained from eq 8. [d] From intercept in Figure 3 assuming ε_{XH}^{510} 13,400. [e] From eq 16; statistically corrected value is 2.22×10^2. [f] From intercept of a plot according to eq 17; statistically corrected value is 1.73×10^2. [g] From $K_Xk_{-3}^{OH}/K_W$ with k_{-3}^{OH} from eq 16. [h] From $K_Xk_{-3}^{OH}/K_W$ with k_{-3}^{OH} from eq 17. [i] From $k_{-3}K_X$ with K_X defined as $[H^+][X^-]/[XH]$; statistically corrected value is 2.70×10^4. [j] Converts to same units as the k_3^{Bi} and k_{-3}^{Bi}'s by dividing through molarity of H$_2$O in water; used in Eigen plot of Figure 5; see text.

strate concentration is linear which means that eq 6 can be approximated by eq 18.[17]

$$\tau^{-1} = \frac{k_1(k_3 + k_3^{OH}[OH^-] + \Sigma k_3^{Bi}[B_i])}{k_{-1} + k_3^{OH}[OH^-]} \frac{K_A}{K_A + [H^+]} + \frac{k_{-1}(k_{-3}[H^+] + k_{-3}^{OH} + \Sigma k_{-3}^{Bi}[BH_i])}{k_{-1} + k_3^{OH}[OH^-]}$$
(18)

A plot of τ^{-1} vs. the stoichiometric buffer concentration, $[B]_0$,[18] at any given pH and substrate concentration, or a plot of the intercepts of τ^{-1} vs. buffer concentration plots vs. the stoichiometric substrate concentration has a slope given by eq 19. After substituting

slope =

$$\frac{k_1k_3^B \dfrac{K_AK_B}{(K_A + [H^+])(K_B + [H^+])} + k_{-1}k_{-3}^B \dfrac{[H^+]}{K_B + [H^+]}}{k_{-1} + k_3^{OH}[OH^-]}$$
(19)

$k_3^B K_B/K_X$ for k_{-3}^B and solving for k_3^B one obtains eq 20; K_B for the buffer was determined by potentiometric

$$k_3^B = \frac{slope(k_{-1} + k_3^{OH}[OH^-])(K_B + [H^+])}{\left(k_1\dfrac{K_A}{K_A + [H^+]} + k_{-1}\dfrac{[H^+]}{K_X}\right)K_B}$$
(20)

titration at an ionic strength of 0.5 M.

Calculation of k_3^B according to eq 20 at three to four different pH values for Tris, borate, and the substrate showed some variation with pH. This could be expected since the experimental uncertainties are not insignificant and the various terms in eq 20 have their

(17) k_3 in denominator is omitted since $k_{-1} \gg k_3$.
(18) $[B]_0 = [B] + [BH]$; $[B] = [B]_0K_B/(K_B + [H^+])$, $[BH] = [B]_0\cdot$ $[H^+]/(K_B + [H^+])$ with K_B being the acid dissociation constant of BH.

Table IV. Proton Transfer Rates between XH and B_i

B_i	$pK_{Bi}{}^a$	$k_3{}^{Bi}$, M^{-1} sec^{-1}	$k_{-3}{}^{Bi}$,^{b} M^{-1} sec^{-1}
Phosphate	6.28	$\sim$2 $\times$ 10^7	$\sim$4.6 $\times$ 10^7
Tris	8.06	$\sim$10^7	$\sim$3.8 $\times$ 10^5
AH	8.65	$\sim$10^8	$\sim$10^6
Borate	8.71	$\sim$6 $\times$ 10^6	$\sim$5.1 $\times$ 10^4

^{a} Determined by standard potentiometric titration at $\mu = 0.5$ M.
^{b} $k_{-3}{}^{Bi} = k_3{}^{Bi}K_{Bi}/K_X$.

own uncertainties. Nevertheless, the constants which are summarized in Table IV and which represent averages from data at different pH values certainly give the right order of magnitude.

Discussion

Justification of Our Assumptions. In treating the kinetic data we assumed XH to be a steady-state intermediate, i.e., $k_{-1} + k_3 + k_3{}^{OH}[OH^-] + \Sigma k_3{}^{Bi}[B_i] \gg k_1$, $k_{-3}[H^+] + k_{-3}{}^{OH} + \Sigma k_{-3}{}^{Bi}[BH_i]$. The treatment of data at high buffer concentration or at high pH is, however, not dependent on this assumption. The fact that buffer dependence levels off at high concentrations means that τ^{-1} no longer depends on the *rate* of any proton transfer step and implies that under these conditions the equilibria $AH_2{}^+ \rightleftharpoons AH$ and $XH \rightleftharpoons X^-$ are rapidly established compared to the equilibrium $AH \rightleftharpoons XH$. The slow relaxation time, derived according to standard procedures,[6] in such a situation is given by

$$\tau_{hi}{}^{-1} = k_1\frac{K_A}{K_A + [H^+]} + k_{-1}\frac{[H^+]}{K_X + [H^+]} \quad (21)$$

As long as $[H^+] \ll K_X$ eq 21 and eq 7 are identical which is the case at pH $\geq$7.8. Since all our data on $\tau_{hi}{}^{-1}$ were in fact obtained at pH $\geq$7.8, the parameters k_{-1}/K_X, k_1, and K_A calculated by applying eq 7 must be valid. This analysis demonstrates that $k_{-1} \gg k_1$ and thus one of the steady-state conditions is indeed fulfilled.

The second steady-state assumption, viz., $k_{-1} + k_3 + k_3{}^{OH}[OH^-] + \Sigma k_3{}^{Bi}[B_i] \gg k_{-3}[H^+] + k_{-3}{}^{OH} + \Sigma k_{-3}{}^{Bi} \cdot [BH_i]$, is amply met since at pH $\geq$7.6 the equilibrium favors X$^-$ over XH by $\geq$ tenfold, i.e., $k_3 + k_3{}^{OH}[OH^-] + \Sigma k_3{}^{Bi}[B_i] \geq 10(k_{-3}[H^+] + k_{-3}{}^{OH} + \Sigma k_{-3}{}^{Bi}[BH_i])$.

Finally our assumption that equilibrium $AH_2{}^+ \rightleftharpoons AH$ is rapidly established compared to the equilibrium $AH \rightleftharpoons X^-$, governed by k_f and k_b, not only at high but at low buffer concentrations as well, is validated as follows. One can estimate the reciprocal relaxation time for the equilibration of a normal[19] acid–base pair with a $pK_a = 8.65$ (pK_A) to be $\geq$10^5 sec^{-1} at pH $\geq$7.6;[19] this is indeed much higher than $k_f + k_b$ at any pH studied.

Rates of Proton Transfer. It should be pointed out at the outset that the value for k_{-3} has an estimated uncertainty of $\pm$40%. With this reservation we note that 5.90 $\times$ 10^{10} M^{-1} sec^{-1} is twice as large as the diffusion controlled rate constant of the reaction of $(CH_3)_3N$ with H_3O^+ (3 $\times$ 10^{10} M^{-1} sec^{-1}).[20] The high value is probably due to the fact that k_{-3} refers to the recombination of oppositely charged ions[12,19] although the high ionic strength (0.5 M) is expected to somewhat counteract such a charge effect.
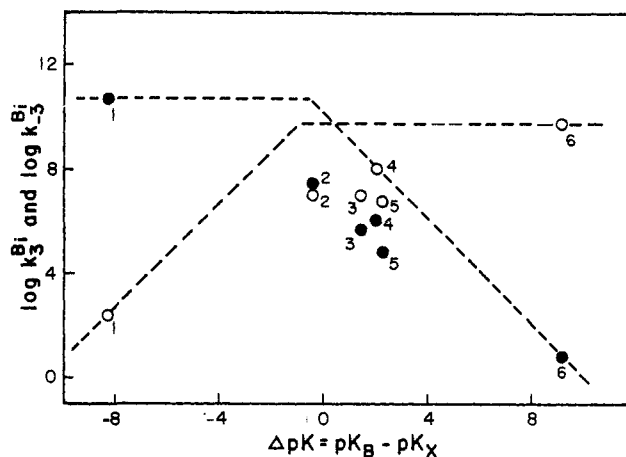
Figure 5. Eigen plot for proton transfer rates: (O) $k_3{}^{Bi}$ (including $k_3{}^{H_2O}$ and $k_3{}^{OH}$); ($\bullet$) $k_{-3}{}^{Bi}$ (including k_{-3} and $k_{-3}{}^{H_2O}$); (1) H_2O; (2) phosphate; (3) Tris; (4) AH; (5) borate; (6) OH$^-$.

The value for $k_3{}^{OH} = 5.2 \times 10^9$ M^{-1} sec^{-1} is close to the value of 7.3 $\times$ 10^9 M^{-1} sec^{-1} for the diffusion controlled reaction of $(CH_3)_2N^+HCH_2COO^-$ with OH^- ($\mu = 0.25$ M)[12] which is of the same charge type as the reaction of XH with OH^-. On the other hand, $k_3{}^{OH}$ is about 25-fold higher than the rate constant reported for the reaction of OH^- with the zwitterionic Meisenheimer complexes formed from TNB with several secondary aliphatic amines.[13] Possible reasons for this difference in rates are under current investigation.

Though the experimental uncertainty in the $k_3{}^{Bi}$ and $k_{-3}{}^{Bi}$ values is high, it is nevertheless instructive to use them, along with $k_3{}^{OH}$, $k_3{}^{H_2O} = k_3/55.6$ and k_{-3}, $k_{-3}{}^{H_2O} = k_{-3}{}^{OH}/55.6$ in drawing an Eigen plot[19] as shown in Figure 5. The plot shows the typical curvature and leveling off at $\Delta pK \gg 0$ ($\Delta pK \ll 0$ respectively) as expected for a simple proton transfer between normal[19] acids and bases and thus further confirms our mechanistic interpretation.

It is, however, noteworthy that the values for Tris and borate lie substantially below the line defined by the other points. In the case of Tris this is probably due to a steric factor whereas the low rate for borate may be related to the fact that proton transfer involves cleavage of a B–O bond;[21] Sayer and Jencks[23] have recently reported similar findings.

Even $k_3{}^B$ for phosphate and the substrate are somewhat lower than expected on the basis on Eigen's[19] data. A possible interpretation could be a weak intramolecular hydrogen bond to one of the o-nitro groups in XH which could lower the proton transfer rate mainly when $\Delta pK \leq 0$, but less when ΔpK is large ($k_3{}^{OH}$).

Implications. In our opinion the two most significant findings of this study are the following: (1) the unexpectedly high value of $k_{-1} = 1.93 \times 10^5$ sec^{-1} with the consequence that under certain conditions deprotonation of XH becomes rate limiting in the formation of X$^-$; (2) the high acidity of XH with a statistically corrected pK_a 2.3 units lower than the pK_a of the amine nucleophile.

(19) M. Eigen, *Angew. Chem., Int. Ed. Engl.*, 3, 1 (1964).
(20) E. Grunwald and E. K. Ralph, *J. Amer. Chem. Soc.*, 89, 2963, 4405 (1967).

(21) (a) J. O. Edwards, G. C. Morrison, V. F. Ross, and J. W. Schultz, *J. Amer. Chem. Soc.*, 77, 266 (1955); (b) R. P. Bell, R. B. Jones, and J. O. Edwards, as quoted in Bell and Evans.[22]
(22) R. P. Bell and P. G. Evans, *Proc. Roy. Soc.*, *Ser. A*, 291, 297 (1966).
(23) J. M. Sayer and W. P. Jencks, *J. Amer. Chem. Soc.*, 95, 5636 (1973).

The high value of k_{-1} is striking because we are dealing with the strongly activated picryl system. It is well known that leaving group departure from Meisenheimer complexes of less activated systems is considerably faster;[24] for example, methoxide ion explusion from the complex between methoxide ion and 2,4-dinitroanisole is about 4×10^4-fold faster than from the analogous complex of 2,4,6-trinitroanisole.[25] We can therefore expect that in typical reactions of secondary amine nucleophiles with the most frequently investigated type of substrate, 1-X-2,4-dinitrobenzene,[26] the reversion of the initially formed zwitterionic intermediate (k_{-1} step) may have rate constants in the order of 10^7–10^8 sec^{-1} or even higher; for less activated substrates such as 1-X-4-nitrobenzene k_{-1} could probably reach a limit of 10^{10} sec^{-1} in some cases.[27]

Let us now consider some implications of our results which may force reassessment of widely accepted notions about base catalyzed nucleophilic aromatic substitutions. In recent years general base catalysis has often been interpreted by a mechanism involving a rapid equilibrium between the initially formed zwitterionic intermediate (2) and its conjugate base (3), followed by a rate-limiting general acid catalyzed leaving group departure (eq 22).[30] This has been called the SB–GA (for specific base–general acid) mechanism.[26a]

$$\text{(22)}$$
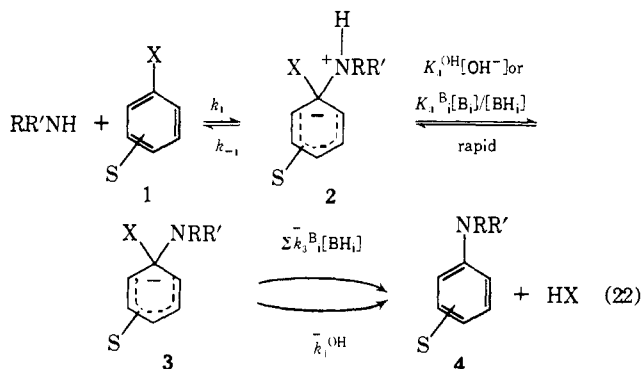
A common way to express the second-order rate coefficient k_A, defined as $(d[4]/dt)/([RR'NH][1])$, is by eq 23[31] where $K_3^{OH} = K_3/K_W$, $K_3^{B}{}_i = K_3/K_{Bi}$ with K_3

(24) For a recent review, see M. J. Strauss, *Chem. Rev.*, **70**, 667 (1970).

(25) C. F. Bernasconi, *J. Amer. Chem. Soc.*, **90**, 4982 (1968).

(26) For recent reviews, see (a) C. F. Bernasconi, *MTP (Med. Tech. Publ. Co.) Int. Rev. Sci.: Org. Chem. Ser. One*, **3**, 33 (1973); (b) F. Pietra, *Quart. Rev., Chem. Soc.*, **23**, 504 (1969); (c) J. Miller, "Aromatic Nucleophilic Substitution," Elsevier, New York, N. Y., 1968; (d) S. D. Ross, *Progr. Phys. Org. Chem.*, **1**, 31 (1963); (e) J. F. Bunnett, *Quart. Rev., Chem. Soc.*, **12**, 1 (1958).

(27) One might argue that the high k_{-1} value of this study to some extent reflects an extra driving force arising from developing resonance stabilization (of the type which renders aniline derivatives weak bases) in the transition state leading from XH to AH. Inasmuch as typical substrates of nucleophilic aromatic substitution reactions are not aniline derivatives such resonance stabilization would be of less importance. However, preliminary results on the same reaction of the Meisenheimer complex of *N*-methyl-β-aminoethyl picryl ether indicate that the effect is not significant.[28] Another possible extra driving force is the release of strain due to eclipsing interactions in the five-membered ring of XH.[29] On the other hand, k_{-1} in the present study relates to an intramolecular complex which probably has the effect of somewhat reducing k_{-1} compared to an intermolecular situation and thus may counteract the effect of eclipsing.

(28) C. F. Bernasconi and C. L. Gehriger, unpublished results.

(29) J. F. Bunnett, personal communication.

(30) See, *e.g.* (a) footnote 27 in J. F. Bunnett and G. T. Davis, *J. Amer. Chem. Soc.*, **82**, 665 (1960); (b) J. F. Bunnett and R. H. Garst, *ibid.*, **87**, 3879 (1965); (c) J. F. Bunnett and C. F. Bernasconi, *ibid.*, **87**, 5209 (1965); (d) A. J. Kirby and W. P. Jencks, *ibid.*, **87**, 3217 (1965).

(31) (a) J. A. Orvik and J. F. Bunnett, *J. Amer. Chem. Soc.*, **92**, 2417 (1970); (b) J. F. Bunnett and C. F. Bernasconi, *J. Org. Chem.*, **35**, 70 (1970).

$$k_A = \frac{k_1(K_3^{OH}\bar{k}_3^{OH}[OH^-] + \Sigma K_3^{B}{}_i\bar{k}_3^{B}{}_i[B_i])}{k_{-1} + K_3^{OH}\bar{k}_3^{OH}[OH^-] + \Sigma K_3^{B}{}_i\bar{k}_3^{B}{}_i[B_i]} \quad (23)$$

and K_{Bi} being the acid dissociation constants of 2 and BH$_i$, respectively.[32]

The results of this study suggest that the assumption of a rapid equilibrium between 2 and 3 may not always be justified. Thus the scheme of eq 24 which takes into

$$\text{(24)}$$
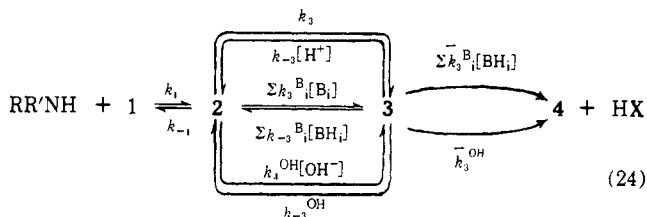
account the *kinetic* importance of the proton transfer steps is more appropriate.

Treating both 2 and 3 as steady-state intermediates but without assuming that 2 and 3 are necessarily in rapid equilibrium, the following expression for k_A is obtained

$$k_A = \frac{k_f(\bar{k}_3^{OH} + \Sigma\bar{k}_3^{B}{}_i[BH_i])}{k_b + \bar{k}_3^{OH} + \Sigma\bar{k}_3^{B}{}_i[BH_i]} \quad (25)$$

where k_f and k_b are defined as in eq 4 and 5.

Two limiting situations with regard to eq 25 are of particular interest. (1) $k_b \gg \bar{k}_3^{OH} + \Sigma\bar{k}_3^{B}{}_i[BH_i]$ which simplifies eq 25 to eq 26. This is identical with

$$k_A = \frac{k_f}{k_b}(\bar{k}_3^{OH} + \Sigma\bar{k}_3^{B}{}_i[BH_i]) =$$

$$\frac{k_1}{k_{-1}}K_3^{OH}\bar{k}_3^{OH}[OH^-] + \Sigma K_3^{B}{}_i\bar{k}_3^{B}{}_i[B_i]) \quad (26)$$

the form taken by eq 23 under the condition $k_{-1} \gg K_3^{OH}\bar{k}_3^{OH}[OH^-] + \Sigma K_3^{B}{}_i\bar{k}_3^{B}{}_i[B_i]$. This means that when leaving group departure is overall rate limiting, it is immaterial whether proton transfer between 2 and 3 is very rapid (eq 23) or whether it is rate limiting with respect to the formation of 3 (eq 22). In other words, the "old views" prevail.

(2) The situation is different when $k_b \ll \bar{k}_3^{OH} +$

(32) Note that in order to account for the noncatalyzed product forming step

$$2 \xrightarrow{k_2} 4$$
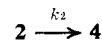
expressions for k_A customarily include an additional term both in the numerator (k_1k_2) and in the denominator (k_2) of eq 23. Since usually $k_{-1} \gg k_2$[26a] this is of no consequence to the denominator. Equation 23 thus accounts for that part of k_A which arises from the base-catalyzed pathways only; under most experimental conditions this is in fact the major part of k_A in many reactions.[26a] We use eq 23 rather than the full equation in order to simplify the discussion to follow. Note also that the "logical pathway"

$$3 \xrightarrow{\bar{k}_3[H^+]} 4$$
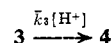
is not included in eq 22 because the sequence

$$2 \xrightleftharpoons{K_3} 3 \xrightarrow{\bar{k}_3[H^+]} 4$$
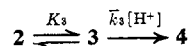
cannot be distinguished from the direct pathway[26a,33]

$$2 \xrightarrow{k_2} 4$$
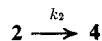
At any rate this sequence has recently been shown to be of very minor importance compared to the k_2 step.[26a,33]

(33) C. F. Bernasconi and R. H. de Rossi, *J. Org. Chem.*, **38**, 500 (1973).

$\Sigma \bar{k}_3{}^{B_i}[BH_i]$; eq 25 becomes eq 27 and general base

$$k_A = k_f = \frac{k_1(k_3 + k_3{}^{OH}[OH^-] + \Sigma k_3{}^{B_i}[B_i])}{k_{-1} + k_3 + k_3{}^{OH}[OH^-] + \Sigma k_3{}^{B_i}[B_i]} \quad (27)$$

catalysis means rate-limiting proton transfer.

Is such a situation likely to occur? To answer this question we must estimate the rates of the processes involved. For simplicity we assume that the concentrations of all B_i's and BH_i's are kept so low that their respective rate terms become negligible, a condition often desired and achieved experimentally (cf. with $\tau_{no}{}^{-1}$).

(1) $\bar{k}_3{}^{OH}$. A recent study of reaction 28 afforded

(28)
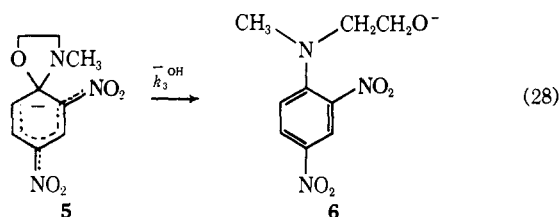
$\bar{k}_3{}^{OH} \approx 10^3 \text{ sec}^{-1}$ in aqueous solution.[33] Since the leaving group in reaction 28 is a rather poor one (pK_a of the conjugate acid of **6** > 14[33]) and due to the intramolecularity, 10^3 sec^{-1} should be regarded as a lower limit for oxygen leaving groups in the 1-X-2,4-dinitrobenzene series; in less activated systems $\bar{k}_3{}^{OH}$ will be correspondingly higher.

(2) k_b. The upper limit for k_b is given by $k_{-3}[H^+] + k_{-3}{}^{OH}$. If we assume $k_{-3} \approx 6 \times 10^{10} M^{-1} \text{ sec}^{-1}$ and $k_3{}^{OH} \approx 5 \times 10^9 M^{-1} \text{ sec}^{-1}$ as determined in this study, this upper limit can be expressed as $6 \times 10^{10}[H^+] + 5 \times 10^{-5}/K_3$.[34]

We need now an estimate for K_3.[34] In our model system the (statistically corrected) pK of the zwitterion (XH) is 2.3 units lower than that of the amine nucleophile ($AH_2{}^+$). In a less activated system (fewer acidifying nitro groups) we expect this pK difference to be somewhat smaller, perhaps ~ 1.6 units with 1-X-2,4-dinitrobenzene substrates. Thus for the reactions in 10% dioxane–90% water of 2,4-dinitrophenyl phenyl ether with piperidine[30c] ($pK \approx 11$), N-methylbenzylamine[35a] ($pK \approx 9.6$), and morpholine[35b] ($pK \approx 8.4$) we estimate $K_3 \approx 4 \times 10^{-10}$, 10^{-8}, and 1.6×10^{-7}, respectively, and $k_b \leq 1.25 \times 10^5$, 5×10^3, and $3.12 \times 10^2 \text{ sec}^{-1}$, respectively.

It thus appears that at least in the morpholine reaction the relation $k_b \ll \bar{k}_3{}^{OH}$ ($\bar{k}_3{}^{OH}$ assumed to be $>10^3 \text{ sec}^{-1}$) and with it eq 27 might indeed hold but probably not in the reactions of the more basic amines.[36]

(34) $k_{-3}{}^{OH} = k_3{}^{OH}/K_3{}^{OH} = k_3{}^{OH}K_w/K_3 = 5 \times 10^{-5}/K_3$, $K_3 = $ acid dissociation constant of **2**.

(35) (a) C. F. Bernasconi, J. Org. Chem., **32**, 2947 (1967); (b) C. F. Bernasconi and P. Schmid, ibid., **32**, 2953 (1967).

Among the reactions of 1-X-2,4-dinitrobenzenes the one of the weakly basic N-methylaniline ($pK \approx 4.9$) with 1-fluoro-2,4-dinitrobenzene[38] appears the most likely to undergo rate-limiting proton transfer, just as Bunnett originally suggested[38] but later rejected[30b] on grounds that deprotonation of **2** should be very rapid.

In still less activated systems, e.g., the reactions of 4-nitrophenyl phosphate with piperidine or dimethylamine,[30d] base catalysis must almost certainly be due to rate-limiting proton transfer.

Experimental Section

Materials. N,N'-Dimethyl-N-picrylethylenediamine hydrochloride was prepared by adding 3.3 g (37 mmol) of N,N'-dimethylethylenediamine (Aldrich) dropwise to a solution of 6.2 g (25 mmol) of picryl chloride in 100 ml of ethanol. A yellow precipitate was formed, presumably the dimer N,N'-dimethyl-N,N'-dipicrylethylenediamine,[39] which was filtered off. The filtrate was acidified with HCl and concentrated to 50 ml in order to precipitate the product which was twice recrystallized from acidic ethanol, mp 208°, yield 25%. Anal. Calcd for $C_{10}H_{14}N_5O_6Cl$: C, 35.75; H, 4.17; N, 20.85. Found: C, 35.85; H, 4.12; N, 20.49. The Meisenheimer complex (X⁻) was prepared in situ by adding one drop of 5 N NaOH to a solution of $AH_2{}^+$ in DMSO-d_6: pmr (DMSO-d_6) δ 2.65 (s, 6, CH_3N), 3.35 (s, 4, CH_2N), 9.04 (s, 2, ring); uv max (H_2O) 442 nm (ϵ 24,200).

Tris[40] (Mallinckrodt), sodium borate (Matheson Reagent Grade), and sodium bicarbonate (Mallinckrodt Analytical Reagent) were used without further purification. HCl and NaOH solutions were prepared from Titrisol.

Rate and Equilibrium Measurements. Kinetic determinations were made on a temperature-jump transient spectrophotometer of Messanlagen Studiengesellschaft, Göttingen, Germany. The solutions, equilibrated at 22°, were subjected to temperature jumps of 3° (end temperature 25°). Relaxation was monitored at 546 or 365 nm; at pH ≤8.80 the relaxtion amplitude[6] was optimal at 546 nm because the equilibrium favors $AH_2{}^+ + AH$ over $XH + X^-$ whereas at pH ≥9.00 the relaxation amplitudes become larger at 365 nm since XH and X⁻ are now the favored species. Each relaxation time reported represents the average of at least three relaxation curves.

The pH was measured at 25 ± 0.1° on a Corning Model 12 pH meter. For the determination of K_X the OD at 510 nm was measured on a Cary 14 spectrophotometer equipped with an expanded slidewire (0–0.2 OD). The residual absorption of $AH_2{}^+$ due to the very high substrate concentration ($3 \times 10^{-3} M$) was compensated by having an acidic solution (pH <5) of $AH_2{}^+$ at the same concentration in the reference cuvette (difference spectrum).

(36) The reaction of 2,4-dinitrophenyl phenyl ether with piperidine in 60% dioxane–40% water shows an isotope effect when the ether oxygen is replaced by ¹⁸O, thus indicating rate-limiting leaving group departure ($k_b > \bar{k}_3{}^{OH}$).[37]

(37) C. R. Hart and A. N. Bourns, Tetrahedron Lett., 2995 (1966)

(38) J. F. Bunnett and J. J. Randall, J. Amer. Chem. Soc., **80**, 6020 (1958).

(39) S. Hünig, private communication, 1970.

(40) Tris(hydroxymethyl)aminoethane.