

Anal. Calcd for $C_{25}H_{42}N_3O_6$: C, 64.63; H, 9.11; N, 9.04. Found: C, 64.75; H, 8.91; N 9.04.

9b (pure minor isomer, 3*R*,4*S*): 540 mg (13.6%); mp 156-157 °C, R_f (D) 0.44; 1H NMR (270 MHz, $CDCl_3$) δ 0.86 (d, 6 H, $J = 6.0$ Hz, 2 CH_3 Iaa), 1.23-1.74 (m, 18 H, includes 1.23, s, C(3) CH_3 , 1.28, s, Boc, 1.40, d, $J = 6.5$ Hz, Ala β - CH_3), 2.31-2.55 (m, 3 H, C(5) H, C(2) H), 3.17 (m, 2 H, $NHCH_2$ Iaa), 3.42 (br dd, 1 H, C(5) H), 3.78 (m, 1 H, C(4) H), 4.37-4.48 (m, 2 H, α -CH, NH), 5.58 (s, 1 H, OH), 6.16 (d, 1 H, $J = 6.5$ Hz, NH), 7.15-7.28 (m, 6 H, NH, aromatic). An additional 665 mg (16.8%) mixture of diastereomers was obtained.

[2H_3]Methyl (4*S*)-Benzyl-5-methyloxazolidone-5-acetate (10). The oxazolidone ring system was prepared by the method reported for the synthesis of a homologue.⁷ Starting with compound **7** (1.408 g, 4 mmol), 536 mg (54.0%) $\geq 95\%$ pure oxazolidone carboxylic acid was obtained as a semisolid.

Esterification of this acid (536 mg, 2.17 mmol) was carried out by reaction with Cs_2CO_3 (1.70 mL of 20% aqueous solution) and CD_3I (0.149 mL, 2.39 mmol) as reported.¹⁵ The ester was then

purified by silica gel chromatography, eluting with 1% methanol in chloroform to give 463 mg (80.8%) of pure **10**.

Acknowledgment. This work was supported by a grant from the National Institutes of Health (No. AM 20100).

Registry No. **2**, 13139-15-6; **3**, 13139-15-6; **4**, 85613-63-4; **5**, 85613-64-5; (3*S*,4*S*)-**6**, 85613-65-6; (3*S*,4*S*)-**7**, 85613-66-7; (3*R*,4*S*)-**7**, 85613-61-2; (3*S*,4*S*)-**7** free acid, 85613-67-8; (3*S*,4*S*)-**8a**, 85613-68-9; (3*R*,4*S*)-**8b**, 85613-69-0; (3*S*,4*S*)-**9a**, 85613-70-3; (3*R*,4*S*)-**9b**, 85613-71-4; **10**, 85613-72-5; Boc-Ala-Iaa, 72155-58-9; H-Ala-Iaa-HCl, 72155-60-3; Boc-Me³Sta-OH, 85613-73-6; ethyl acetate, 141-78-6.

Supplementary Material Available: Full NMR NOE data and spectra for compound **10** (1 page). Ordering information is given on any current masthead page.

(15) Wang, S. S.; Gisin, B. F.; Winter, D. P.; Makofske, R.; Kulesha, I. D.; Tzougraki, C.; Meinenhofer, J. *J. Org. Chem.* 1977, 42, 1286.

(16) Nakanishi, K.; Schooley, D. A.; Koreeda, M.; Muira, I. *J. Am. Chem. Soc.* 1972, 94, 2865.

(17) Jager, Y.; Buss, V. *Liebigs Ann. Chem.* 1980, 101.

General-Base Catalysis in the Reaction of Water with Activated Aromatic Substrates. The Hydrolysis of 3-Methyl-1-picrylimidazolium Ion

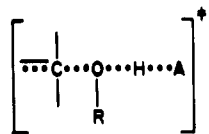
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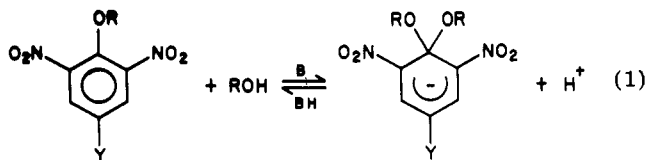
The rate of hydrolysis of 3-methyl-1-picrylimidazolium ion was investigated between pH 1.7 and 9.3 in the presence of several buffer bases at various concentrations. The reaction is strongly catalyzed by oxygen bases. The catalytic constants including water and OH^- are spread over a range of ten powers of ten and show a good Brønsted correlation with $\beta = 0.62$. The kinetic solvent isotopic effect for the water-, acetate-, and OH^- -catalyzed reactions are 2, 0.86, and 0.84, respectively. The mechanism of catalysis is discussed, and it is concluded that it represents concerted addition of water to the aromatic ring.

The reaction of water and alcohols with the carbonyl carbon occurs by a concerted mechanism with an important component of proton transfer in a mobile transition state.¹



This mechanism is supported by a large number of structure-reactivity correlations which provide evidence for a fully concerted mechanism.²

Similar results have been found for the addition-elimination of alcohols to activated aromatic compounds (eq 1)³ where the general base-acid catalysis was interpreted in terms of a concerted mechanism.



(1) Palmer, J. L.; Jencks, W. P. *J. Am. Chem. Soc.* 1980, 102, 6472 and references cited therein.

(2) Jencks, W. P. *Acc. Chem. Res.* 1980, 13, 161.

(3) Bernasconi, C. F.; Gandler, J. R. *J. Am. Chem. Soc.* 1978, 100, 8117.

On the other hand, the reaction of water with activated aromatic substrates is known in less detail mainly because most aromatic nucleophilic substitution reactions have been studied at high pH where the hydroxide ion catalyzed reaction is predominant. It has been suggested that water addition to 4,6-dinitrobenzofuroxan occurs through a concerted mechanism.⁴ Also related are the studies on pseudobase formation from water addition to quaternary nitrogen heterocycles.⁵

We recently reported a study of the hydrolysis of *N*-picrylimidazole (**1**) and interpreted the general buffer catalysis observed along the same lines.⁶ At low pH, the reaction follows the rate law of eq 2, where the term k_{SH}

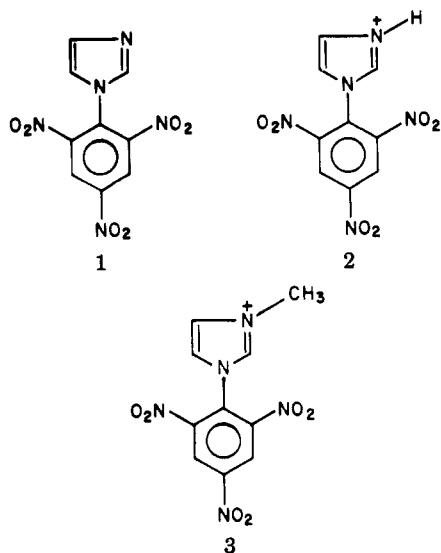
$$v = k_0[S] + k_s[S][B] + k_{SH}[S][BH] \quad (2)$$

was attributed to the reaction of water with *N*-picrylimidazolium cation (**2**), catalyzed by general bases, but this mechanism is kinetically indistinguishable from others that involve **1** and a general-acid catalyst in the transition state. We now report a study of the hydrolysis of 3-methyl-1-picrylimidazolium cation (**3**) which is expected to be a good model for **2**.

(4) Terrier, F.; Millot, F.; Norris, W. P. *J. Am. Chem. Soc.* 1976, 98, 5883.

(5) Bunting, J. W.; Meathrel, W. G. *Can. J. Chem.* 1973, 51, 1965.

(6) de Rossi, R. H.; de Vargas, E. B. *J. Am. Chem. Soc.* 1981, 103, 1533.



Results

The hydrolysis of 3 leads quantitatively to picric acid, and the rate was measured by determining the increase in absorbance at the maximum of the product (358 nm). The experimental pseudo-first-order rate constant k_{obsd} depends on the pH and the total buffer concentration, [Bt]; thus it can be expressed as a sum of two terms corresponding to buffer catalyzed (k_c [Bt]) and uncatalyzed (k_u) pathways (eq 3).

$$k_{\text{obsd}} = k_u + k_c[\text{Bt}] \quad (3)$$

In Figure 1 the observed rate of hydrolysis extrapolated to zero buffer concentration (k_u) is plotted as a function of pH. It can be seen that a smooth pH-rate profile is obtained despite the fact that several types of buffer were used. The rate remains constant from pH 1.40 to 5 and increases at higher pH.

Buffer Catalysis. Buffer catalysis was observed with a variety of buffer species (Table I). The slopes of lines similar to that of Figure 2, k_c , were plotted against the fraction of free base α_B . The left intercept of this plot, which represents catalysis by the acidic portion of the buffer, was undistinguishable from zero with all the buffers used, indicating that only the basic form of the buffer catalyzes the reaction. With the buffer $\text{PO}_4\text{H}_3/\text{PO}_4\text{H}_2^-$, the plot of k_c vs. $\alpha_{\text{PO}_4\text{H}_2^-}$ was not linear, which might indicate that even at this low pH the catalysis by PO_4H_2^- contributes significantly. In this case the observed second-order catalytic rate coefficient for the buffer is given by eq 4 where K_1 and K_2 are the first and second ionization constant for PO_4H_3 .

$$k_c = k_{\text{PO}_4\text{H}_2^-} \frac{K_1[\text{H}^+]}{[\text{H}^+]^2 + K_1[\text{H}^+] + K_1K_2} + \frac{K_1K_2}{k_{\text{PO}_4\text{H}_2^-}([\text{H}^+]^2 + K_2[\text{H}^+] + K_1K_2)} \quad (4)$$

Since we worked at a pH between 1.40 and 2.70, the product K_1K_2 is negligible in the denominator of eq 4, and it can be rearranged to eq 5.

$$k_c \frac{[\text{H}^+] + K_1}{K_1} = k_{\text{PO}_4\text{H}_2^-} + k_{\text{PO}_4\text{H}_2^-} \frac{K_2}{[\text{H}^+]} \quad (5)$$

A plot of the left-hand side of eq 5 vs. $1/[\text{H}^+]$ gives a straight line from which $k_{\text{PO}_4\text{H}_2^-}$ and $k_{\text{PO}_4\text{H}_2^-} K_2$ can be extracted. The value of $k_{\text{PO}_4\text{H}_2^-}$ so obtained is about 15 times higher than the value obtained at high pH for the same

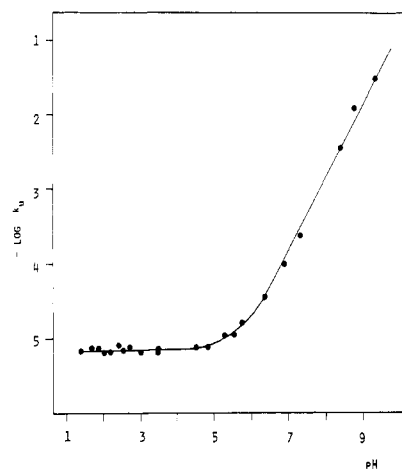


Figure 1. pH dependence of the $\log k_u$ (s^{-1}), the observed pseudo-first-order rate constant extrapolated to zero buffer concentration for the 3-methyl-1-picryl imidazolium ion hydrolysis at 25 °C and $\mu = 1$ M. The solid line was calculated from k_{obsd} (s^{-1}) = $6.69 \times 10^{-6} + 987[\text{OH}^-]$.

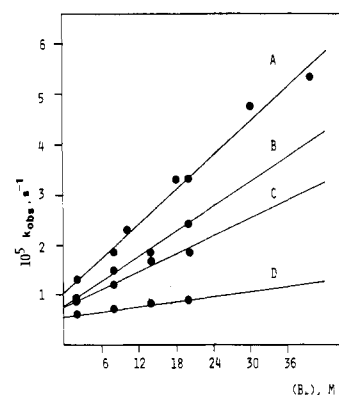


Figure 2. Effect of buffer concentration and pH on k_{obsd} for the hydrolysis of 3-methyl-1-picrylimidazolium ion (acetate buffer, 25 °C, $\mu = 1$ M (NaCl), $[\text{3}]_0 = (5-20) \times 10^{-5}$ M): A, pH 5.50; B, pH 4.80; C, pH 4.51; D, pH 3.40.

Table I. Rate Constants k_B for Base Catalysis of the Hydrolysis of 3-Methyl-1-picrylimidazolium Ion at 25 °C^a

base	pK _{BH}	-log k_B^b	pH range
H ₂ O	-1.61 ^c	6.910 ^d (7.111)	1.4-4.50
PO ₄ H ₂ ⁻	1.72 ^e	4.854 (5.148)	1.40-2.70
ClCH ₂ COO ⁻	2.53	4.955 (5.000)	2.00-3.44
⁻ O ₂ CCH ₂ CH ₂ CO ₂ H	3.85	3.967	3.00-3.40
CH ₃ COO ⁻	4.65	3.886 (3.959)	3.40-5.50
CH ₃ COO ⁻	4.65	3.921 ^f	3.40-5.50
⁻ O ₂ CCH ₂ CH ₂ CO ₂ ⁻	5.00	3.301	5.30-5.70
PO ₄ H ²⁻	6.32	1.298	5.65-7.30
PO ₄ H ²⁻	6.32	1.357 ^f	5.65-7.30
CO ₃ ²⁻	9.78	-0.727	8.75-9.30
OH ⁻	15.58 ^g	-2.994	6.00-9.30

^a Ionic strength 1 M, NaCl as compensating electrolyte unless otherwise noted. ^b k_B in $\text{M}^{-1} \text{s}^{-1}$; values in parentheses correspond to the hydrolysis of 2, taken from ref. 6. ^c $K_{\text{H}_3\text{O}^+} = [\text{H}_2\text{O}]\gamma_{\text{H}^+}$, with $\gamma_{\text{H}^+} = 0.754$ taken from: Harned H.; Owen, B. "The Physical Chemistry of Electrolyte Solutions", 3rd. ed.; Reinhold: New York, 1958; p 748. $[\text{H}_2\text{O}] = 54.4$ was determined by weighing a known volume of a 1 M NaCl solution. ^d $k_B = k_u/54.4$. ^e Fox, J. F.; Jenck, W. P. *J. Am. Chem. Soc.* 1974, 96, 1436. ^f NaNO₃ as compensating electrolyte. ^g $K_{\text{BH}} = K_w\gamma_{\text{H}^+}(\gamma_{\text{H}^+}\gamma_{\text{OH}^-}/a_{\text{H}_2\text{O}})^{-1}[\text{H}_2\text{O}]^{-1}$ with $K_w = 1.007 \times 10^{-14}$, $\gamma_{\text{H}^+} = 0.754$, and $\gamma_{\text{H}^+}\gamma_{\text{OH}^-}/a_{\text{H}_2\text{O}} = 0.535$ from the same reference as in footnote c.

Table II. Effect of Different Compensating Electrolytes on the Rate of Hydrolysis of 3-Methyl-1-picrylimidazolium Ion at 25 °C^a

[buffer], M	10 ² k _{obsd} , s ⁻¹	[buffer], M	10 ² k _{obsd} , s ⁻¹
0.020	0.0935 ^b	0.207	1.07 ^c
0.042	0.159 ^b	0.290	1.90 ^c
0.207	1.11 ^b	0.395	3.10 ^c
0.290	2.10 ^b	0.020	0.121 ^{b,d}
0.395	3.50 ^b	0.020	0.158 ^{b,e}
0.042	0.153 ^c		

^a pH 6.99, ionic strength 1 M, buffer PO₄H₂⁻/PO₄H²⁻.
^b NaCl as compensating electrolyte. ^c NaNO₃ as compensating electrolyte. ^d 0.190 M Na₂SO₄ added. ^e 0.316 M Na₂SO₄ added.

species. Part of the difference may be attributed to errors involved in the determination of $k_{\text{PO}_4\text{H}_2^-}$ under conditions where PO₄H²⁻ is the minor species and also to the influence of salts. Since the increase in the observed rate constants at low pH with weakly basic catalysts is relatively small, it was necessary to work at high buffer concentration, and as the results reported below indicate, the specific salt effect is quite important in this reaction. For this reason we consider that the catalytic rate coefficient for PO₄H₂⁻ is only approximate and that the rate constant reported for PO₄H²⁻ in Table I is the value obtained at high pH with buffer of PO₄H₂⁻/PO₄H²⁻.

The effect of added *N*-methylimidazole was thoroughly investigated because of its importance in regard to the mechanism of catalysis. Only a very small increase in rate was observed (~10–20%) which can hardly be differentiated from a specific salt effect. From the data we can estimate that the second-order rate coefficient for catalysis by *N*-methylimidazole is at most 10⁻³ M⁻¹ s⁻¹.

We have previously found that tertiary amines are poorer catalysts for the hydrolysis of 1 than carboxylate anions of the same p*K*.⁶

Salt Effect. The effect of added salts was examined to discard the possibility that the observed catalysis is due to specific salt effect.⁷ The catalytic coefficients for PO₄H²⁻ and acetate are different when NaNO₃ or NaCl is used as the compensating electrolyte (Table I); however, the observed catalytic coefficients for these bases are high enough to conclude that most of the observed catalysis is effectively due to the base acting as such and not to an activity coefficient effect.

The salt effect is also evident in the reaction catalyzed by PO₄H²⁻ which has some SO₄²⁻ present (Table II). For instance, the presence of 0.3 M of Na₂SO₄ in a solution containing 0.02 M total PO₄H₂⁻/PO₄H²⁻ at pH 6.99 increases the rate by a factor of 1.7, but an increase in buffer concentration from 0.02 to 0.3 M increases the rate by a factor of 20.

In order to minimize the specific salt effect, we always worked under conditions where the compensating electrolyte is at least 10 times the concentration of the buffer salts.⁸

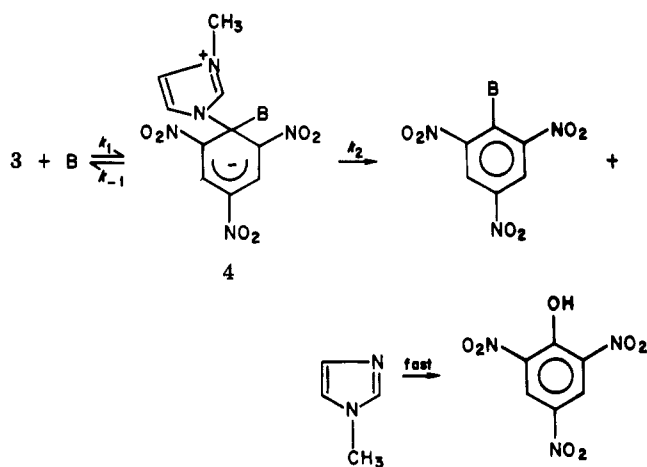
Isotopic Solvent Effect and Activation Parameters. The kinetic solvent isotopic effect (KSIE) was determined for the reactions catalyzed by OH⁻, acetate, and water. The KSIEs were obtained for water and acetate from reactions carried out in D₂O with a AcD/Ac⁻ buffer at pD 4.84, where the OD⁻ reaction contributes very little, as the intercept and slope of a plot of k_{obsd} vs. acetate anion con-

Table III. Activation Parameters and Solvent Isotope Effect ($k_{\text{H}}/k_{\text{D}}$) for the Hydrolysis of 3-Methyl-1-picrylimidazolium Ion Catalyzed by Bases^a

base	ΔH^\ddagger , kcal mol ⁻¹	$-\Delta S^\ddagger$, ^c cal mol ⁻¹ deg ⁻¹	$k_{\text{H}}/k_{\text{D}}$ ^d
H ₂ O	18.4	28.6	2.10 ^e
CH ₃ COO ⁻	19.1	12.2	0.86
OH ⁻	4.0	30.8	0.84 ^f

^a Ionic strength 1 M. ^b Calculated from the relation $\Delta H^\ddagger = E_a - RT$; E_a = Arrhenius activation energy. Estimated error 1 kcal mol⁻¹. ^c Calculated from $k_{\text{B}} = kT/e\Delta S^\ddagger/R_e^{-\Delta H^\ddagger/RT/h}$ at 25 °C. Estimated error 3 eu. ^d At 35 °C. ^e Calculated from the pseudo-first-order rate constant in H₂O and D₂O. ^f This is the ratio of the slopes of plots of k_{obsd} vs. $a_{\text{OL}} = K_{\text{L}}/10^{-\text{pL}}$ with $K_{\text{L}} = 2.089 \times 10^{-14}$ for water and 2.84×10^{-15} for D₂O, taken from: Weast, R. C.; Astle, M. J., Eds. "Handbook of Chemistry and Physics", 61st ed.; CRC Press: Cleveland, OH, 1980-1981; p D-168.

Scheme I



centration. These values were compared with those obtained simultaneously in a similar solution but with water as the solvent. The results are summarized on Table III.

The KSIE for the OH⁻-catalyzed reaction was obtained from the observed pL (L = H or D) dependence of the rate of solutions buffered with borax in water and D₂O (borax does not catalyze the reaction).

The activation parameters for the same three catalytic rate coefficients were obtained from rates measured at three or four temperatures in the range of 15–45 °C and are also shown in Table III.

Discussion

In Table I the rate of hydrolysis of 1-methyl-3-picrylimidazolium ion (3) catalyzed by several bases is compared with the corresponding rate for 1-picrylimidazolium ion (2). There is a fair agreement between the two over a range of three orders of magnitude,⁹ indicating that 3 is a good model for 2 and that both react by the same mechanism.

The fact that the hydroxide ion catalyzed reaction of 3 is 18 times smaller than the calculated rate for the same reaction of 2 confirms our previous conclusion with regard to the mechanism of the neutral hydrolysis of 1.⁶

There are several mechanisms that could account for the kinetic behavior of added bases. In terms of the general

(7) Saloma, P.; Kankaanpera, A.; Lahti, M. *J. Am. Chem. Soc.* 1971, 93, 2084.

(8) Hammett, L. P. "Physical Organic Chemistry", 2nd. ed.; McGraw-Hill: New York, 1970; p 197.

(9) The differences are very likely due to a differential salt effect since the rates of 2 were measured under conditions (high buffer concentration) where the concentration of the compensating electrolyte was not always 10 times higher than the concentration of the buffer salts.

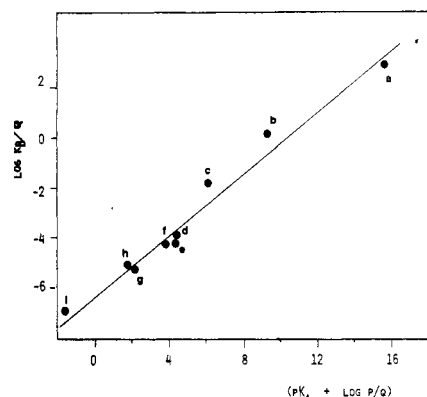


Figure 3. Brønsted plot for the general-base-catalyzed hydrolysis of 3-methyl-1-picrylimidazolium ion: (a) OH^- , (b) CO_3^{2-} , (c) PO_4H^{2-} , (d) $(\text{CH}_2)_2(\text{COO}^-)_2$, (e) CH_3COO^- , (f) $\text{HO}_2\text{C}(\text{CH}_2)_2\text{COO}^-$, (g) $\text{ClCH}_2\text{COO}^-$, (h) PO_4H_2^- , (i) H_2O . P and Q are statistical corrections.

mechanism of nucleophilic aromatic substitution,¹⁰ nucleophilic catalysis by a base may be represented as in Scheme I.

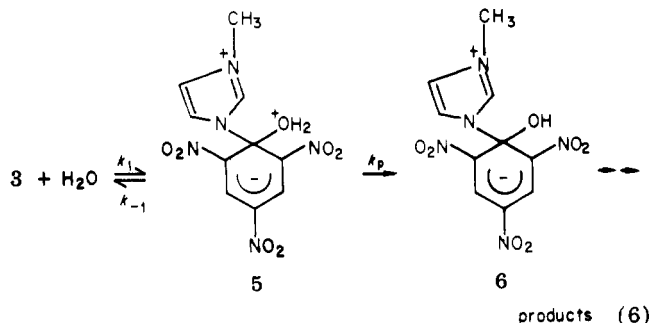
When the base is acetate, picryl acetate should be formed if the mechanism of catalysis is that shown in Scheme I. The observed rate of hydrolysis of 3 in the presence of 0.02 M acetate buffer at pH 3.4 is $6.03 \times 10^{-6} \text{ s}^{-1}$ while the rate of hydrolysis of picryl acetate under the same conditions is $3.94 \times 10^{-3} \text{ s}^{-1}$,¹¹ thus accumulation of picryl acetate is not expected.

The values of k_B in Table I represent k_1 in Scheme I only if the rates of elimination of acetate (k_{-1}) and *N*-methylimidazole (k_2) from the intermediate 4 are such that $k_2 > k_{-1}$.

The rates of leaving group expulsion from aromatic σ complexes have been shown to depend on the p*K* of the leaving group, being about the same for amines and alkoxide ions of similar p*K*.¹²

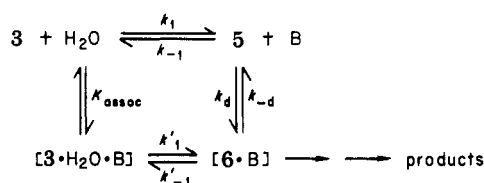
If the same relationship holds for carboxylate ions and amines, a change in rate-determining step would be expected with the bases used, which is inconsistent with the linear Brønsted plot observed (Figure 3).

Another mechanism that could account for the observed base catalysis is the stepwise reaction shown in eq 6, with k_p being rate determining.

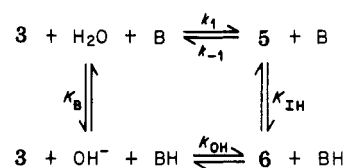


The p*K* of 5 must be on the order of -8 or -9 on the basis of the p*K* of $\text{CH}_3\text{CH}_2\text{O}^+\text{H}_2$ (-1.94)¹³ and the effect of the cyclohexadienide group¹² and the ammonium group.¹⁴

Scheme II

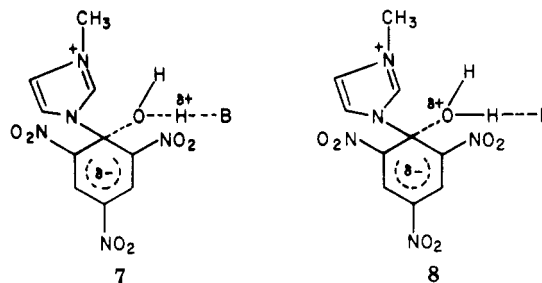


Scheme III



Therefore, proton transfer from 5 to all bases including water should be diffusion controlled, and a β value close to zero (and very little or no catalysis at all)¹⁵ rather than the value observed would be expected.

The stabilization of the transition state by the base may come from a concerted mechanism with C-O bond formation and O-H bond fission occurring at the same time as represented by transition state 7 or by hydrogen



bonding to the base without actual proton transfer, i.e., 8. Transition state 8 corresponds to a preassociation mechanism¹⁶ as described in the lower path of Scheme II.

A preassociation mechanism is expected for a reaction where the lifetime of the intermediate formed in a stepwise mechanism is shorter than the diffusion away of the catalyst, namely, when $k_{-d} < k'_{-1}$ (Scheme II).

In order to estimate the rate of breakdown of intermediate 5, we use the thermodynamic cycle shown in Scheme III from which it follows that $k_{-1} = K_{IH}k_1/K_B K_{OH}$. For acetate as a base, K_B is $4.47 \times 10^{-10} \text{ M}$, $K_{IH} = 10^{13}-10^{14}$, and $K_{OH} = k_{OH}/k_{-OH} = 10^3/10 \text{ (M}^{-1}\text{)}$, which are the rates of the hydroxide ion reaction taken from Table I and the rate of hydroxide expulsion from trinitrocyclohexadienate complexes.¹⁷ The value of k_{-1} so estimated is $10^{15}-10^{16} \text{ s}^{-1}$.

Although this calculation is crude, the value obtained is well above diffusion limit of 10^{10} ,¹⁵ indicating that the lifetime of 5 is, in fact, very short. We conclude that water addition must occur either through an enforced preassociation or a concerted mechanism.²

The small inverse kinetic solvent isotopic effect for the acetate-catalyzed and hydroxide-catalyzed reactions might be taken as an indication that the transition state for these bases is best described by a preassociation mechanism, that they reflect the fact that acetate is more strongly solvated by D_2O than by OH_2 ,^{18a} and that OD^- is a stronger base

(10) Bernasconi, C. F. *MTP Int. Rev. Sci.: Org. Chem., Ser. One* 1973, 3, 33.

(11) Unreported results from this laboratory.

(12) Bernasconi, C. F.; Muller, M. C.; Schmid, P. *J. Org. Chem.* 1979, 44, 3189.

(13) Lee, D. G.; Cameron, R. *J. Am. Chem. Soc.* 1971, 93, 4724.

(14) The p*K* of $\text{MeNH}_2^+\text{-CH}_2\text{OH}$ is 9.98 (Hine, J.; Craiy, J. C., Jr.; Underwood, J. G., II; Via, F. A. *J. Am. Chem. Soc.* 1970, 92, 5194. Hine, J.; Kakesh, F. C. *Ibid.* 1970, 92, 4383) while the p*K* of MeOH is $\sim 14-15$.

(15) Eigen, M. *Angew. Chem., Int. Ed. Engl.* 1964, 3, 1.

(16) Jencks, W. P. *Acc. Chem. Res.* 1976, 9, 425.

(17) Bernasconi, C. F. *J. Am. Chem. Soc.* 1970, 92, 4682.

(18) Loughton, P. M.; Robertson, R. E. In "Solvent-Solvent Interactions"; Coetzee, J. F., Ritchie, C. D., Eds.; Marcel Dekker: New York, 1969: (a) p 482; (b) p 474.

than OH^- .^{18b} However, the lack of a significant isotopic effect must be interpreted with caution because it depends not only on the degree of coupling between bond forming and bond breaking but also on the geometry of the transition state.¹⁹

The large negative entropy of activation is consistent with transition state 7, and similar values have been reported for general-base-catalyzed hydrolysis of a number of activated acyl compounds and in the reactions of water with some quaternary ammonium heterocycles⁵ and with 4,6-dinitrobenzofuroxan.⁴

More over, the Brønsted β is characteristic of concerted reactions,²⁰ and it is the preferred mechanism of many other hydration reactions of electrophilic carbon centers.²¹

It is remarkable that the rate of catalysis by hydroxide ion falls on the same Brønsted line defined by the other bases, suggesting that they all react by the same mechanism. This behavior is similar to that observed in the reaction of other aromatic electrophiles and contrasts with the reactions at carbonyl carbon²² where a positive deviation from the Brønsted line indicates that the direct nucleophilic reaction is the lowest energy reaction pathway for hydroxide ion in water as predicted by the rule formulated by Jencks.²³

We suggest that the general-base-catalyzed reaction of hydroxide ion may compete favorably with direct nucleophilic addition because the unfavorable process of desolvating a hydroxide ion is avoided when the reaction occurs through a water molecule since a hydrogen bond between water and OH^- changes to a fully formed bond as water adds to the electrophile, similar to the Grothaus chain conduction of protons.²⁴ An analogous argument was put forward by Ritchie.²⁵

Experimental Section

Materials. The 3-methyl-1-picrylimidazolium chloride was synthesized from picryl chloride, previously recrystallized from CCl_4 (1 mmol), and *N*-methylimidazole (1 mmol) in 20 mL of anhydrous ether. After 15 min at room temperature in the dark, a light yellow precipitate was separated by centrifugation and washed twice with anhydrous ether, with minimal exposure to the atmosphere.

The precipitate was dried in a desiccator over H_2SO_4 . The product, used without further purification, was an extremely hygroscopic powder, mp 147–150 °C.

The ultraviolet spectrum of the product in acidic water (10^{-2} M HCl) shows a shoulder at 300–310 nm. In basic solution (10^{-2} M NaOH) it hydrolyses within the time required for preparation of the solution, and the spectrum matches one of picric acid.

The purity of the substrate was checked by UV-visible spectra in acidic and basic solutions.

All inorganic salts were reagent grade commercial materials and were used without further purification.

N-Methylimidazole was vacuum distilled and stored over 4-Å molecular sieves in the refrigerator. Water twice distilled in a glass apparatus was used throughout.

Deuterium oxide (99.98%) was a gift from the Atomic Energy Committee (Argentina) and was used as received.

The pK of buffer at a 1 M ionic strength was taken from the literature or determined by potentiometric titration. Standard buffers prepared according to the literature²⁶ were used to calibrate the electrode.

The buffer solutions for the kinetic run were diluted from a stock solution of the required buffer, and the ionic strength was adjusted to 1 M. If the pH of a set of solutions with a constant buffer ratio differed by more than 5×10^{-3} pH units, a drop of diluted acid or base was added to adjust the pH to that of the more concentrated solution. Four or more buffer concentrations were used for each pH.

The UV spectra were recorded on a Beckman 24 spectrophotometer, and the change in optical density during a kinetic run was measured at the maximum absorption of picric acid ($\lambda = 358$ nm).

pH measurement were carried out on a Seybold digital pH meter at the temperature of the experiments.

The pH (pD) of solutions for kinetic solvent isotopic effect determinations were determined on a Corning Model 130 digital pH meter with a Corning combination microelectrode. The pD was obtained by adding 0.4 to the pH meter reading.²⁷

The OH^- and OD^- "concentrations" for this reactions are defined as $[\text{OL}^-] = K_L/10^{-\text{pL}}$ ($L = \text{H}$ or D), where pL is the pH meter reading and K_L is the thermodynamic self-ionization constant of the solvent on a molar scale.²⁸

In all other cases the OH^- concentration is defined as in eq 7,

$$[\text{OH}^-] = \frac{K_w}{\gamma_{\text{H}^+}\gamma_{\text{OH}^-}/\alpha_{\text{H}_2\text{O}}} \frac{\gamma_{\text{H}^+}}{10^{-\text{pH}}} \quad (7)$$

where K_w is thermodynamic value, and γ_{H^+} and the activity coefficient ratio at a 1 M ionic strength (with NaCl) and at the temperature of the experiment were taken from the literature.²⁹

Kinetics. Reactions were initiated by adding 1–2 mg of the solid substrate to a solution containing all other constituents at the required temperature ± 0.02 . Rate constants were computed from plots of $\ln(\text{OD}_\infty - \text{OD}_t)$ vs. time by least-squares analysis of at least 10 data points. Most reactions were followed up to 80–90% conversion, but those having a half-life > 4 h were followed up to 10–50% conversion, and the infinite optical density was obtained by hydrolyzing a small portion of the solution at high pH.

The reactions with a half-life < 5 min were run in the thermostated cell of the spectrophotometer and were initiated by adding 25 μL of a concentrated solution of the substrate to 2 mL of the buffer. The final substrate concentration was $\sim 5 \times 10^{-5}$ M.

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