mixture ($CDCl_3$ + 1 drop $CF_3CO_2H$): δ 1.25 for the 6-Me isomer, δ 1.44 for the 7-Me isomer.

3-Methylindoline, identical with an authentic sample, picrate (mp 146–148 °C), had a retention time of 420 s. 4-Methyl-3,4-dihydro-2,1-benzothiazine 2,2-dioxide had a retention time of 840 s. The quantitative results are summarized in Table I.

**6-Methyl-6,7-dihydro-5*H*-1-pyrindine.** A mixture of 5- and 6-methyloctahydropyrindines (12.0 g)[19] was dehydrogenated with 10% Pd-C in mesitylene. The mixture of 5- and 6-methyl-6,7-dihydro-5*H*-1-pyrindines was obtained (72%): bp 101–104 °C (35 mm). Attempted separation of the isomers on a variety of gas chromatographic columns (OV-17, Versamid 900, diethylene glycol succinate) was unsuccessful. The mixture was converted to the picrates which were fractionally crystallized from 95% ethanol six times to give 6-methyl-6,7-dihydro-5*H*-1-pyrindine picrate: mp 137–137.5 °C (lit.[19] 135–136 °C). Decomposition with $NH_4OH$ gave the free base: NMR ($CDCl_3$ + 1 drop $CF_3CO_2H$) δ 8.46 (d, $J_{2,3}$ = 5.0 Hz, 1 H, $H_2$), 8.26 (d, $J_{3,4}$ = 8.0 Hz, 1 H, $H_4$), 7.70 (dd, $J_{2,3}$ = 5.0 Hz, $J_{3,4}$ = 8.0 Hz, 1 H, $H_3$), 3.6–2.4 (m, 5 H), 1.25 (d, $J$ = 6.1 Hz, 3 H, $CH_3$); if $CF_3CO_2H$ was not added to the $CDCl_3$ solution, $H_2$ gave a broad band at δ 8.2; mass spectrum (70 eV), *m/e* (relative intensity) 133 ($M^{+\cdot}$, 100), 132 (96), 118 (92), 117 (32), 40 (24).

**7-Methyl-6,7-dihydro-5*H*-1-pyrindine.** To a solution of lithium diisopropylamide [from diisopropylamine (0.204 g) in THF (10 mL) at −25 °C and 2 N *n*-butyllithium in hexane (1.05 mL)] was added 6,7-dihydro-5*H*-1-pyrindine (0.24 g) at −25 °C with stirring under an atmosphere of dry $N_2$. After 30 min methyl iodide (0.30 g) in THF (2 mL) was added to the dark red solution at −25 °C. The solution, which became pale yellow, was allowed to come to room temperature and then evaporated. The yellow oil was purified by gas chromatography (single peak) on a 6 ft × $^3/_{16}$ in. 10% OV-17 on Gas Chrom Q column at 160 °C to give 7-methyl-6,7-dihydro-5*H*-1-pyrindine (retention time = 485 s); NMR ($CDCl_3$) δ 8.36 (d, $J_{2,3}$ = 5.4 Hz, 1 H, $H_2$), 7.46 (d, $J_{3,4}$ = 7.9 Hz, 1 H, $H_4$), 6.98 (dd, $J_{2,3}$ = 5.4 Hz, $J_{3,4}$ = 7.9 Hz, 1 H, $H_3$), 3.35–3.00 (m, 1 H, $H_7$) 2.97–2.78 (m, 2 H, $H_5$), 2.78–2.15 (m, 1 H, $H_6$), 1.84–1.54 (m, 1 H, $H_6$), 1.34 (d, $J$ = 7.5 Hz, 3 H, $CH_3$); on addition of 1 drop of $CF_3CO_2H$ the Me doublet moved downfield to δ 1.44; mass spectrum (70 eV), *m/e* (relative intensity) 133 ($M^{+\cdot}$, 58), 132 (53), 118 (100), 117 (33). Anal. Calcd for $C_9H_{11}N$: C, 81.16; H, 8.32. Found: C, 81.20; H, 8.29. Picrate: mp 128.5–129 °C (ethanol). Anal. Calcd for $C_9H_{11}N{\cdot}C_6H_3N_3O_7$: C, 49.73; H, 3.90. Found: C, 49.48; H, 3.88.

**Flash Vacuum Pyrolysis of 1-Phenyl-2-propanesulfonyl Azide at 650 °C.** The products were chromatographed on a column of silica gel to give a mixture of 6- and 7-methyl-6,7-dihydro-5*H*-1-pyrindine (17.3%) (6:7 = 4.3:1), β-methylstyrene (1.4%), and 3-methyl-3,4-dihydro-2,1-benzothiazine 2,2-dioxide (8.4%).

At 400 °C the yields were 10.9% (6:7 = 5.8:1), 1.4%, and 14%, respectively.

**Flash Vacuum Pyrolysis of Substituted 3,4-Dihydro-2,1-benzothiazine 2,2-Dioxides at 650 °C.** These were carried out as described for the parent sultam. In no case was any styrene or dihydropyrindine detected, only indolines and indoles being formed. The results are summarized in Table II.

**Acknowledgment.** We are grateful to the National Science Foundation (Grants GP-33361X2 and CHE-75-09309) for the support of this work. We also thank Dr. Hiteo Iida for a sample of 6-methoxyindoline.

# Buffer Catalysis in the Hydrolysis of Picrylimidazole

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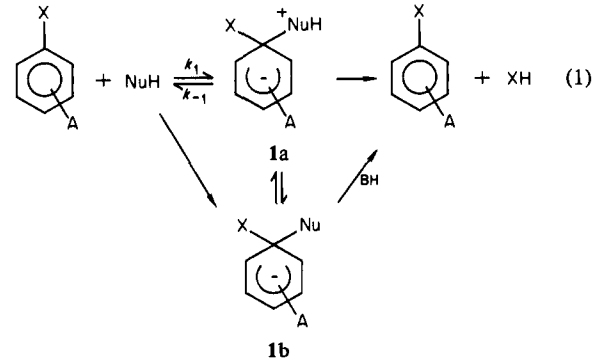
**Abstract:** The kinetics of the hydrolysis of picrylimidazole (S) was studied between pH 0.47 and 10.6 with different buffers at various concentrations. The reaction is strongly catalyzed by carboxylate and phosphate ions. Imidazole also catalyzes the reaction although it is a weak catalyst. At pH below 4 S starts to be protonated and the buffer-independent rate constant increases until it reaches a plateau. Rate constants for the base-catalyzed pathway are separated into the contribution of the rate for the protonated and unprotonated substrates. The Brønsted plots give slopes of 0.5 and 0.58, respectively. The mechanism of catalysis is discussed.

The reaction of amines with activated aromatic substrates was extensively studied during the past decade and important achievements regarding the mechanism of these reactions were obtained.[1] Most of the earliest studies have been done by conventional kinetics from which information about some of the elemental steps was obtained. The application of techniques for the measurement of fast reactions and the study of model reactions were definitive in firmly establishing the detailed mechanism of nucleophilic aromatic substitution and in particular the mechanism of the product-forming steps in reactions with amines as nucleophiles.[2] There are still some questions which remain to be answered, one of them pertains to the mechanism of the first step, namely, the addition of the nucleophile to the aromatic substrate to form the intermediate σ-complex **1a**,**b** (eq 1).




With neutral nucleophiles like amines, alcohols, or water, a proton must be removed at some point in the reaction coordinate to final products. Under conditions where the formation of the intermediate is rate determining, the intermediate **1a** may be so unstable that its formation is avoided through concerted catalysis. The diagonal pathway in eq 1 implies a transition state like **2** and has been suggested for the hydroxide ion catalysis observed in the reactions of amines,[3] under conditions where the formation of the

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

(2) Bernasconi, C. F.; Gehriger, C. L.; de Rossi, R. H. *J. Am. Chem. Soc.* **1976**, *98*, 8451.

(3) (a) Kirby, A. J.; Jencks, W. P. *J. Am. Chem. Soc.* **1965**, *87*, 3209. (b) Kirby, A. J.; Jounas, M. *J. Chem. Soc. B* **1970**, 1187.

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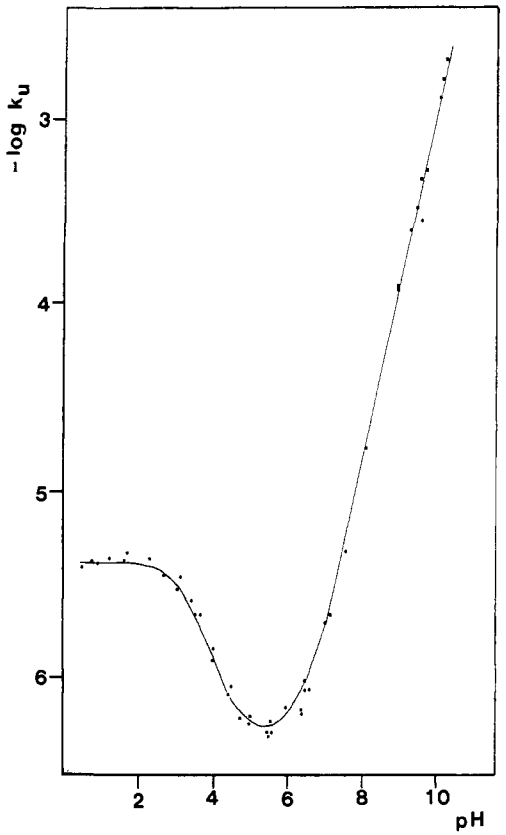
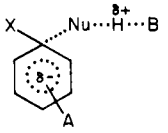



**Figure 1.** pH dependence of log $k_u$ ($s^{-1}$), the observed pseudo-first-order rate constant extrapolated to zero buffer concentration for the picrylimidazole hydrolysis at 25 °C, $\mu$ = 1 M. Solid line calculated from eq 3 with $K_{SH} = 3.67 \times 10^{-4}$ $M^{-1}$, $k_2 = 5.06 \times 10^{-7}$ $s^{-1}$, $k_2^H = 4.3 \times 10^{-6}$ $s^{-1}$, and $k_1 = 14.2$ $M^{-1}$ $s^{-1}$.

$\sigma$ complex is rate determining. More recently a transition state like **2** was suggested for the addition of water to activated aromatic substrates.[4,5]




**2**

Although rate constants for the hydrolysis of a great variety of aromatic substrates are known,[6] in most cases the studies were done at very high pH. Under these conditions the reaction with hydroxide ion is overwhelming and no reaction with water can be detected. In studies carried out under acidic conditions, base catalysis was not attempted.[7]

In this paper we report a study of the hydrolysis of picrylimidazole (S) which is of interest not only because its high reactivity allows the studies to be done at low pH but also it is the first comprehensive study reported with an amine as leaving group as well.

## Results

In spite of the fact that picrylimidazole is a picramide, its spectrum does not have the characteristic maximum in the visible around 400 nm. Besides, it is hydrolyzed faster than picryl chloride, giving picric acid as the only product, under all our reaction conditions, as determined by comparison of the infinity solutions with that of a mock solution of picric acid.

Table I. Experimental Pseudo First-Order Rate Constants for the Hydrolysis of Picrylimidazole in Water at 25 °C[a]

| pH | $10^4 k_u$,[b] $s^{-1}$ | $10^4 k_c$,[c] $M^{-1}$ $s^{-1}$ |
|---|---|---|
| 10.61[d,e] | 53.6 | 323 |
| 10.17[f] | 20.8 | [g] |
| 10.06[h] | 16.8 | [g] |
| 9.99[d] | 13.3 | 242 |
| 9.65[f] | 5.42 | 9.25 |
| 9.51[d] | 4.88 | 143 |
| 9.40[f] | 3.42 | [g] |
| 9.25[i] | 2.50 | [g] |
| 8.94[f] | 1.24 | [g] |
| 8.91[d] | 1.30 | 56.4 |
| 8.05[j] | 0.175 | [g] |
| 7.56[j] | 0.0495 | 0.0735 |
| 7.16[j] | 0.0229 | 0.0608 |
| 7.03[h] | 0.0203 | 1.68 |
| 6.62[j] | 0.00873 | 0.0258 |
| 6.47[h] | 0.00883 | 1.30 |
| 6.39[j] | 0.00690 | 0.0145 |
| 6.39[j,k] | 0.00645 | 0.0151 |
| 6.00[h,l] | 0.0702 | 1.11 |
| 5.90[h] | [g] | 1.09 |
| 5.59[m] | 0.00543 | 0.0131 |
| 5.56[n,o] | 0.00610 | 0.0407 |
| 5.50[n] | 0.00505 | 0.0311 |
| 5.03[n,o] | 0.0102 | 0.0333 |
| 5.01[m] | 0.00583 | 0.0233 |
| 4.73[n,o] | 0.00614 | 0.00489 |
| 4.49[n] | 0.00921 | 0.00590 |
| 4.45[m] | 0.00839 | 0.0417 |
| 4.00[m] | 0.0149 | 0.0438 |
| 3.99[n] | 0.0129 | 0.0795 |
| 3.61[n] | 0.0224 | 0.00767 |
| 3.50[p,q] | 0.0225 | 0.00508 |
| 3.38[n] | 0.0271 | 0.0897 |
| 3.10[n] | 0.0371 | 0.0750 |
| 3.03[p] | 0.0306 | 0.0475 |
| 2.63[p] | 0.0364 | 0.0333 |
| 2.27[p,r] | 0.0446 | 0.0300 |
| 1.67[p,r] | 0.0484 | 0.0125 |
| 1.61[h] | 0.0432 | 0.0617 |
| 1.59[h,r] | 0.0541 | 0.0303 |
| 1.23[s] | 0.0457 | 0.0247 |
| 0.90[h] | 0.0431 | [g] |
| 0.80[s] | 0.0426 | 0.0267 |
| 0.78[s] | 0.0435 | 0.0300 |
| 0.72[s,t] | 0.0432 | 0.0267 |
| 0.72[k,s,t] | 0.0421 | 0.0248 |
| 0.47[s] | 0.0402 | 0.0242 |

[a] Solvent contains 2% dioxane; $\mu$ = 1 M (NaCl as compensating electrolyte unless otherwise quoted). [b] Extrapolated rate constant at zero buffer concentration. [c] $k_c$ is the slope of a plot of $k_{obsd}$ vs. total buffer concentration (0.02–0.20 M unless otherwise indicated). [d] Carbonate buffer. [e] Buffer concentration up to 0.4 M. [f] Borax buffer. [g] Not detected experimentally. [h] Phosphate buffer. [i] Dabco buffer. [j] Imidazole buffer. [k] $NaNO_3$ as compensating electrolyte. [l] Buffer concentration 0.01–0.10 M. [m] Acetate buffer. [n] Phthalate buffer. [o] Buffer concentration up to 0.5 M. [p] Chloroacetate buffer. [q] Buffer concentration up to 0.92 M. [r] Buffer concentration up to 1 M. [s] Trichloroacetate buffer. [t] Buffer concentration up to 0.7 M.

The kinetics of the hydrolysis was studied between pH 10.6 and 0.47 (Table I), and the pH profile is shown in Figure 1. The reaction is catalyzed by the buffers; thus, the rate constants plotted in Figure 1 pertain to the observed rate constants extrapolated to zero buffer concentration.

As can be seen, a smooth pH–rate profile was obtained despite the fact that buffers of varying chemical types were used.

The experimental rate constants, $k_{obsd}$, can be expressed as the sum of the individual pseudo-first-order rate constant for buffer-catalyzed ($k_c$) and -uncatalyzed pathways ($k_u$) (eq 2).

$$k_{obsd} = k_u + k_c(B_t) \quad (2)$$

(4) Terrier, F.; Millot, F.; Norris, W. P. *J. Am. Chem. Soc.* **1976**, *98*, 5883.
(5) de Rossi, R. H.; de Vargas, E. B. *J. Org. Chem.* **1979**, *44*, 4100.
(6) Miller, J. "Aromatic Nucleophilic Substitution"; Elsevier: New York, 1968.
(7) Murto, J. *Acta Chem. Scand.* **1966**, *20*, 310.

Scheme I

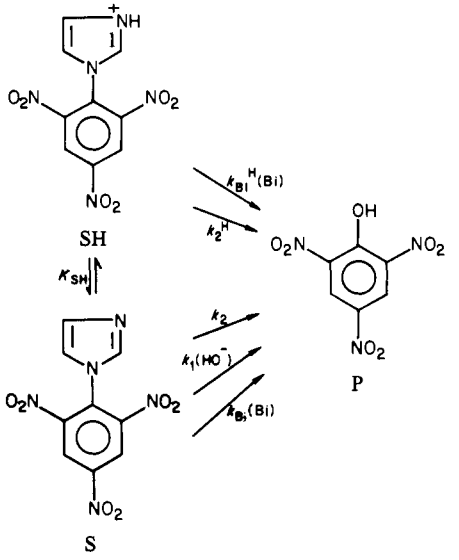



Table II. Rate Constants $k_B$ and $k_B^H$ for Base Catalysis of the Picrylimidazole Hydrolysis in Water at 25 °C[a]

| base | $pK_a$[b] | $k_B$, $M^{-1}$ $s^{-1}$ | $k_B^H$, $M^{-1}$ $s^{-1}$ |
|---|---|---|---|
| $HO^-$ | 15.7[c] | 14.2 | |
| $CO_3^{2-}$ | 9.70[d] | $3.61 \times 10^{-2}$ | |
| $CO_3H^-$ | 7.33[e] | $6.01 \times 10^{-4}$ | |
| imidazole | 7.21[d] | $1.53 \times 10^{-5}$ | |
| $PO_4H^{2-}$ | 6.49[f] | $1.88 \times 10^{-4}$ | |
| $CH_3COO^-$ | 4.65[g] | $1.10 \times 10^{-7}$ | $1.10 \times 10^{-4}$ |
| phthalate | 4.60[e] | $1.69 \times 10^{-6}$ | $1.43 \times 10^{-4}$ |
| biphthalate | 2.58[e] | $7.37 \times 10^{-7}$ | $1.03 \times 10^{-5}$ |
| $ClCH_2COO^-$ | 2.54[e] | | $9.99 \times 10^{-6}$ |
| $PO_4H_2^-$ | 1.72[f] | $8.53 \times 10^{-5}$ | $7.11 \times 10^{-6}$ |
| $Cl_3CCOO^-$ | 0.47[e] | | $3.86 \times 10^{-6}$ |
| $H_2O$ | −1.74[h] | $9.12 \times 10^{-9}$[i] | $7.75 \times 10^{-8}$[j] |

[a] Solvent contains 2% dioxane; $\mu$ = 1 M (NaCl as compensating electrolyte). [b] $pK_a$ at $\mu$ = 1 M unless otherwise quoted. [c] Thermodynamic value: Sayer, J. M.; Jencks, W. P. *J. Am. Chem. Soc.* 1973, *95*, 5637. [d] Jencks, W. P.; Gilchrist, M. *Ibid.* 1968, *90*, 2622. [e] Determined by potentiometric titration, NaCl as compensating electrolyte. [f] Fox, J. F.; Jencks, W. P. *J. Am. Chem. Soc.* 1974, *96*, 1436. [g] Sayer, J. M.; Jencks, W. P. *Ibid.* 1969, *91*, 6353. [h] Hammett, L. P. "Physical Organic Chemistry", 2nd ed.; McGraw-Hill: New York, 1970; p 321. [i] Obtained from $k_2/55.5$. [j] Obtained from $k_2^H/55.5$.

The pH–rate profile of Figure 1 is consistent with a rate equation of the form of eq 3.

$$k_u = [k_2 + k_1(HO^-)]\frac{K_{SH}}{K_{SH} + (H^+)} + k_2^H\frac{(H^+)}{K_{SH} + (H^+)} = \left[k_2 + \frac{k_1K_W}{(H^+)}\right]\frac{K_{SH}}{K_{SH} + (H^+)} + k_2^H\frac{(H^+)}{K_{SH} + (H^+)} \quad (3)$$

Scheme I shows the reactions to which the various rate constants refer; viz, $k_2$ and $k_2^H$ belong to the noncatalyzed water reaction of the unprotonated (S) and protonated (SH) substrates, respectively, while $k_1$ refers to the hydroxide ion catalyzed reaction of S.

The various rate coefficients were determined as follows. We estimate the $pK_{SH}$ as 3.44 ± 0.09[8] from the kinetic data by standard procedures; thus, at pH >5, we have $(H^+) < K_{SH}$ and eq 3 reduces to eq 4 since $k_2^H/k_2$ = 8.5 (Table II).

$$k_u = k_2 + k_1(HO^-) \quad (4)$$

(8) The thermodynamic $pK_{SH}$ is 2.7. In general, the $pK_a$ of amines is higher at ionic strength of 1 M than in pure water. Compare for instance the $pK$ of imidazole which is 6.95 in water[9] and 7.2 at 1 M ionic strength.

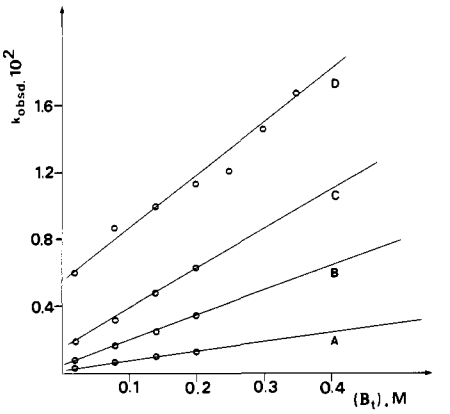



**Figure 2.** Effect of buffer concentration and pH on $k_{obsd}$ for the hydrolysis of picrylimidazole (carbonate buffer, 25 °C, $\mu$ = 1 M): (A) pH 8.91; (B) pH 9.51; (C) pH 9.99; (D) pH 10.61.

In principle $k_2$ and $k_1$ could be estimated from a plot of $k_u$ vs. $(HO^-)$ as the intercept and slope, respectively; however, the fact that $k_2 << k_1$ makes the intercept of such a plot indistinguishable from zero and only the slope can be obtained. In order to get $k_2$, we use only the data at 5 < pH < 7 because in this restricted range of pH the two terms of eq 4 are of similar magnitude.

At pH < 4, $k_1(HO^-)$ is negligible compared with $k_2$ in eq 3; then it can be simplified to eq 5.

$$k_u = k_2\frac{K_{SH}}{K_{SH} + (H^+)} + k_2^H\frac{(H^+)}{K_{SH} + (H^+)} \quad (5)$$

Equation 5 defines the slope of the ascendent part of the pH profile below pH 4. $k_2^H$ was obtained as the mean value of the pH-independent rate constant determined at low pH.

The curve drawn in Figure 1 was calculated by using $pK_{SH}$ = 3.44 and the rate constants given in Table II.

**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. From the left intercept we got the catalytic coefficient for the acidic component of the buffer, whereas the right intercept (fraction of free base equals 1) gives the catalytic coefficient for the basic component of the buffer.

With most of the buffers used we found that the acidic, as well as the basic, species catalyzed the reaction.

According to Scheme I, the rate constant for the catalysis by buffer species, i.e., the slopes of Figure 2 and the like, can be represented by eq 6 where acid catalysis is interpreted as specific acid–general base catalysis. $K_{BiH}$ represents the acid dissociation

$$k_c = k_{Bi}^H\frac{(H^+)}{K_{SH} + (H^+)}\frac{K_{BiH}}{K_{BiH} + (H^+)} + k_{Bi}\frac{K_{SH}}{K_{SH} + (H^+)}\frac{K_{BiH}}{K_{BiH} + (H^+)} \quad (6)$$

constant of the buffer, $k_{Bi}^H$ is the catalytic coefficient for the buffer base reacting with the protonated substrate, and $k_{Bi}$ is the catalytic coefficient for the buffer base reacting with the unprotonated substrate.

At 8.91 < pH < 10.6 we use a carbonate–bicarbonate buffer. Under these conditions $(H^+) < K_{SH}$, and eq 6 simplifies to eq 7.

$$k_c = \left[k_{Bi}^H\frac{(H^+)}{K_{SH}} + k_{Bi}\right]\frac{K_{BiH}}{K_{BiH} + (H^+)} \quad (7)$$

We expect $k_{Bi}^H$ to be not much higher than $k_{Bi}$; considering that $k_2^H$ is $4.3 \times 10^{-6}$ $s^{-1}$ and $k_2 = 5.06 \times 10^{-7}$ $s^{-1}$, thus $k_2^H/k_2$ = 8.5. Then in eq 7 $k_{Bi}^H(H^+)/K_{SH} << k_{Bi}$ and only catalysis by the basic component of the buffer should be expected. The fact that the acidic component of the buffer, bicarbonate, also catalyzes the reaction indicates that it is also acting as a base, which is quite reasonable according to the Brønsted line defined by other bases

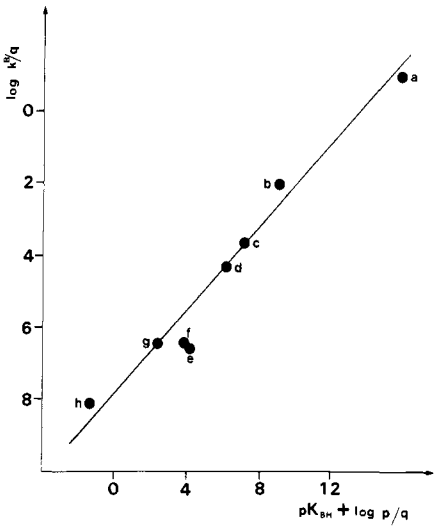



**Figure 3.** Brønsted plot for the general-base-catalyzed hydrolysis of S: (a) hydroxide ion; (b) carbonate; (c) bicarbonate; (d) phosphate dibasic; (e) acetate; (f) phthalate; (g) biphthalate; (h) water.

Table III. Effect of Imidazole Concentration on the Catalyzed and Uncatalyzed Rate of Hydrolysis of Picrylimidazole in Water at 25 °C[a]

| $(Im)_{free}$, M | $10^4 k_{obsd}$, $s^{-1}$ | $(Im)_{free}$, M | $10^4 k_{obsd}$, $s^{-1}$ |
|---|---|---|---|
| | A. Phosphate Buffer[b] | | |
| | 0.404 | 0.0500 | 0.417 |
| 0.0935 | 0.418 | | |
| | B. Carbonate Buffer[c] | | |
| | 9.91 | 0.398 (0.02)[d] | 6.65 (7.80)[e] |
| 0.0969 | 9.61 | 0.398 (0.08)[d] | 15.6 (16.4)[e] |
| 0.194 | 9.19 | 0.398 (0.14)[d] | 20.7 (24.6)[e] |
| 0.291 | 9.31 | 0.398 (0.20)[d] | 32.6 (33.7)[e] |
| 0.388 | 8.54 | | |

[a] Solvent contains 2% dioxane; $\mu$ = 1 M (NaCl as compensating electrolyte). [b] Total buffer concentration = 0.2 M, pH 7.22. [c] Total buffer concentration = 0.2 M unless otherwise indicated, pH 8.20. [d] Values in parentheses indicate the total buffer concentration. [e] Values in parentheses indicate the rate observed in absence of imidazole.

(Figure 3). Similar behavior is observed with $PO_4H_2^-/PO_4H^{2-}$ buffer which we use at 5.90 < pH < 7.03 and gives a plot of $k_c$ vs. free $PO_4H^{2-}$ with a finite intercept. However, if we attribute this value to catalysis by $PO_4H_2^-$ acting as a base, the rate constant obtained is about 3 orders of magnitude above the Brønsted line. We could attribute the catalysis by this species to its kinetic equivalent, namely, catalysis by $PO_4H^{2-}$ acting on the protonated substrate (eq 8).

$$k_{PO_4H_2^-} = k_{PO_4H^{2-}}{}^{H}\frac{K_{PO_4H_2^-}}{K_{SH}} \quad (8)$$

From eq 8 we calculate $k_{PO_4H^{2-}}{}^{H}$ as $8.72 \times 10^{-2}$ $M^{-1}$ $s^{-1}$ which is a value 2 orders of magnitude higher than expected. These reactions are not inhibited by up to 0.1 M imidazole (Table III) (see below).

With an acetic acid–acetate buffer we work at 5.59 > pH > 4. In this pH range we found that $k_c$ increases as the pH is lowered, indicating that there is a significant contribution of the acidic part of the buffer to the total rate. We use a rearranged form of eq 6, viz., eq 9, to calculate $k_{Ac}$ and $k_{Ac}{}^{H}$.[10]

$$k_c[K_{SH} + (H^+)] = k_{Ac}{}^{H}K_{AcH}\frac{(H^+)}{K_{AcH} + (H^+)} + k_{Ac}K_{SH}\frac{K_{AcH}}{K_{AcH} + (H^+)} \quad (9)$$

(9) Bruice, T. C.; Schmir, G. L. *J. Am. Chem. Soc.* **1958**, *80*, 148.

From a plot of the left-hand side of eq 9 vs. the fraction of free base, we obtain $k_{Ac}$ and $k_{Ac}{}^{H}$.

With phthalate buffer we worked at 3 < pH < 5.5. At pH >4.48, the predominant species of the buffer are phthalate monobasic ($BH^-$) and phthalate dibasic ($B^{2-}$); thus, we have to consider the possibility that both species catalyze the hydrolysis of S and SH, and the complete rate expression for the catalysis by the buffer is given by eq 10.

$$\frac{K_{BH} + (H^+)}{K_{BH}}k_c = \left[k_B{}^{H} + k_{BH}{}^{H}\frac{(H^+)}{K_{BH}}\right]\frac{(H^+)}{K_{SH} + (H^+)} + \left[k_{BH}\frac{(H^+)}{K_{BH}} + k_B\right]\frac{K_{SH}}{K_{SH} + (H^+)} \quad (10)$$

Since we work under conditions where $(H^+) \leq K_{BH}$ and $BH^-$ is a weaker base than $B^{2-}$, the catalytic terms of $BH^-$, $k_{BH}$ and $k_{BH}{}^{H}$ can be neglected in eq 10, and $k_B$ and $k_B{}^{H}$ can be obtained by plotting the data according to eq 9.

At pH <4 we have a significant fraction of nondissociated phthalic acid ($BH_2$) whereas the fraction of S changes from 0.78 at pH 3.99 to 0.32 at pH 3.09. The slopes of the buffer plots of $k_{obsd}$ vs. total buffer concentration are then given by eq 11.

$$k_c = \left[k_{BH}\frac{K_{SH}}{K_{SH} + (H^+)} + k_{BH}{}^{H}\frac{(H^+)}{K_{SH} + (H^+)}\right]\frac{K_{BH_2}(H^+)}{D} + \left[k_B\frac{K_{SH}}{K_{SH} + (H^+)} + k_B{}^{H}\frac{(H^+)}{K_{SH} + (H^+)}\right]\frac{K_{BH_2}K_{BH}}{D} \quad (11)$$

$$D = (H^+)^2 + K_{BH_2}(H^+) + K_{BH_2}K_{BH}$$

Since we are working under conditions where $K_{BH} \leq (H^+)$ and $k_B{}^{H}/k_B$ is expected to be >10, the term involving $k_B$ can be neglected in eq 11 and $k_{BH}$ and $k_{BH}{}^{H}$ calculated from a plot of the left-hand side of eq 12 vs. $(H^+)$.

$$\frac{k_cDK_{SH} + (H^+)}{(H^+)} - k_B{}^{H}K_{BH_2}K_{BH} = k_{BH}K_{SH}K_{BH_2} + k_{BH}{}^{H}K_{BH_2}(H^+) \quad (12)$$

With chloroacetic and trichloroacetic acids the total rate increase is low and too difficult to distinguish from a salt effect. However, the fact that the catalysis is independent of the type of salt used (Table I) may indicate that the increase in rate observed represents true buffer catalysis.[12]

We tried to use tertiary amines as catalysts, but they were very ineffective. Imidazole weakly catalyzes the reaction, but it is not nearly as good as other catalysts of similar p$K$. However, the finding that the rate increase is due to base catalysis rather than to a specific salt effect as indicated by the fact that the catalytic coefficient is the same when NaCl or $NaNO_3$ is used as the compensating electrolyte[12] (Table I) is very important in regard to the mechanism (see below).

## Discussion

The curved pH–rate profile shown in Figure 1 indicates that the observed rate of hydrolysis has contributions from acid-catalyzed, neutral, and base-catalyzed reactions, which resembles the hydrolysis of acetylimidazole.[13]

(10) Alternatively $k_{Ac}$ and $k_{Ac}{}^{H}$ could be obtained from a plot according to the following equation

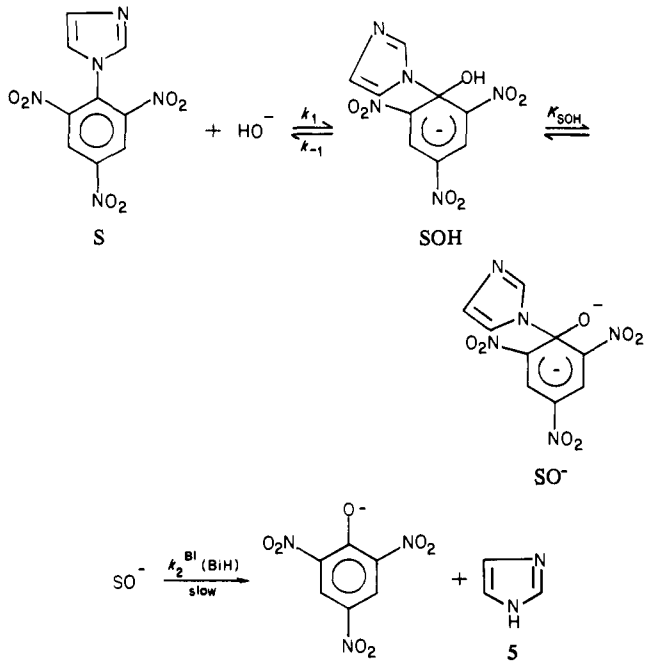
$$k_c\frac{K_{AcH} + (H^+)}{K_{AcH}} = k_{Ac}{}^{H}\frac{(H^+)}{K_{SH} + (H^+)} + k_{Ac}\frac{K_{SH}}{K_{SH} + (H^+)}$$

but the uncertainty in the value of $K_{SH}$ is higher than that in $K_{AcH}$, hence more accurate values are obtained by plotting the data as indicated in eq 10.[11]

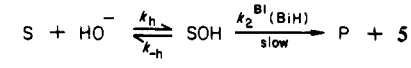
(11) Spiegel, M. R. "Estadistica"; McGraw-Hill: Mexico, 1970; p 226.

(12) Salomaa, P.; Kankaanpera, A.; Lehti, M. *J. Am. Chem. Soc.* **1976**, *93*, 2084.

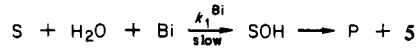
Scheme II



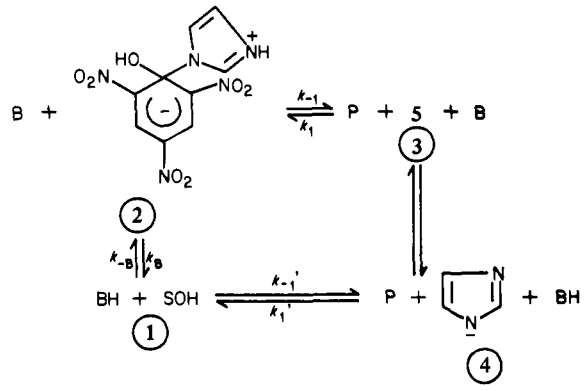
Scheme III



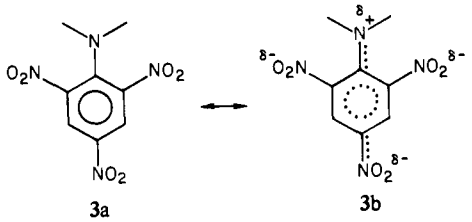
Scheme IV



Scheme V



The hydroxide ion catalyzed rate of hydrolysis of S, $k_1$, is 20 times the rate of hydrolysis of picryl chloride. This result is surprising since ordinary picramides are very unreactive toward nucleophiles.[6] The low reactivity of picramides can be attributed to ground-state stabilization through structures like **3a** ↔ **3b**.



Ground-state stabilization has been shown to contribute strongly in lowering the rate of attack of nucleophiles at the C-1 position of trinitroanisole.[14] Structures like **3b** are expected to destabilize the ground state of picrylimidazole since they disrupt the aromatic system of imidazole. Moreover, molecular models show that steric strain prevents coplanarity between the two rings and structures like **3b** cannot contribute; thus, the high reactivity of picrylimidazole may be attributed only to the inductive effect of the nitrogen of the imidazole ring which has a formal positive charge.[15]

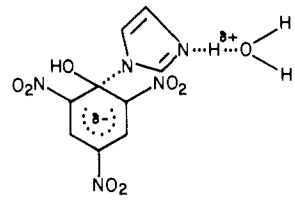
The water reaction can be interpreted in two ways as is shown in eq 13 and 14.

$$S + H_2O \rightarrow P \qquad (13)$$

$$S + H^+ \rightleftharpoons SH + HO^- \xrightarrow{k_2'} P \qquad (14)$$

If the mechanism for the water reaction involves the reaction of hydroxide ion with SH (eq 14), the observed rate constant is $k_2 = k_2'K_W/K_{SH}$; thus we can calculate $k_2' = 1.84 \times 10^4$ $M^{-1}$ $s^{-1}$, which seems to be too high compared with the rate of reaction of hydroxide ion with other picrylic compounds. For example, the reaction of hydroxide ion with picryl fluoride, a very reactive substrate, has a rate constant of 700 $M^{-1}$ $s^{-1}$.[7] Furthermore, preliminary results on the hydrolysis of 3-methyl-1-picrylimidazolium chloride yield a rate constant equal to $1.5 \times 10^3$ $M^{-1}$ $s^{-1}$,[16] 1 order of magnitude smaller than the calculated value of $k_2'$. Hence mechanism 14 is unattractive. On the other hand, for the protonated substrate SH, the only possible mechanism for the uncatalyzed pathway corresponds to the reaction of SH with water. The ratio $k_2^H/k_2 = 8.5$ is reasonable considering the usually small dependence of the rate of nucleophilic aromatic substitution on the leaving group.[17]

A kinetically equivalent mechanism involving **4** as transition state for the rate-determining step of the water reaction can also be rejected as will be discussed below.



**Buffer Catalysis. Picrylimidazole.** There are several mechanisms that could account for the catalysis by the buffer species, and they are shown in Schemes II–IV, where Bi represents any base present in the solution including the solvent.

The observed rate constant for Scheme II is given by eq 15,

$$k_{obsd} = \frac{k_1\sum_i k_2^{Bi}K_{SOH}/K_{BiH}}{k_{-1} + \sum_i k_2^{Bi}(K_{SOH}/K_{BiH})(Bi)}(Bi)(OH^-) \simeq \frac{k_1}{k_{-1}}\sum_i k_2^{Bi}(K_{SOH}/K_{BiH})(Bi)(HO^-) \qquad (15)$$

and it predicts an order higher than 1 for the hydroxide ion, which is contrary to our experimental observations and thus can be easily rejected.

The observed rate constant for Scheme III is given by eq 16.

$$k_{obsd} = \frac{k_h(HO^-)\sum_i k_2^{Bi}(BiH)}{k_{-h} + \sum_i k_2^{Bi}(BiH)} \simeq \frac{k_h}{k_{-h}}\sum_i k_2^{Bi}\frac{K_W}{K_{BiH}}(Bi) \qquad (16)$$

This mechanism implies concerted general-acid catalysis of the amine-leaving group departure from the Meisenheimer complex to be rate determining. The microscopic reverse, i.e., the addition of imidazole to the aromatic ring, should be general base catalyzed

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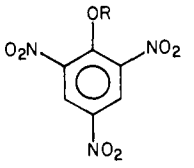
(14) Bernasconi, C. F. *J. Am. Chem. Soc.* **1970**, *92*, 4682.
(15) Adam, W.; Gremison, A. *Tetrahedron* **1966**, *18*, 835.
(16) Unpublished results from this lab.
(17) Bunnett, J. F.; Garbisch, E. W., Jr.; Pruitt, K. M. *J. Am. Chem. Soc.* **1957**, *79*, 385.

under similar reaction conditions.

Three-dimensional reaction coordinate diagrams have been extensively used during the last 10 years to aid the interpretation of reaction mechanisms.[18,19] For the mechanism represented in Scheme III the four corners of this diagram could be represented as in Scheme V.

The standard free energy for the process 1 → 2 when B = $HO^-$ can be calculated as 10.7 kcal/mol from the relationship $\Delta G_{1,2} = -RT \ln K_W/K_{HSOH}$. $pK_{HSOH}$ is expected to be about 1 unit lower than the p$K$ of imidazole.[20] It is somewhat more difficult to estimate $\Delta G_{1,4}$. We can approach a crude estimation as follows. We first assume that the rate of addition of a nucleophile to the one position of compounds of type **6** is about the same when R = H or alkyl. We further assume that the relative rates of the



6

imidazole ion and hydroxide ion are about the same for reactions with 2,4-dinitro compounds and picryl derivatives, namely, 1.9 which is the rate ratio determined for 1-fluoro-2,4-dinitrobenzene.[21] The basic hydrolysis rate of compounds **6** with various R groups is around 1.5 $M^{-1}$ $s^{-1}$,[22,23] we then estimate that $k_1' \approx 3$ $M^{-1}$ $s^{-1}$.

Considering the leaving ability of amines such as piperidine ($pK_a$ = 11.12) and pyrrolidine ($pK_a$ = 11.3) from their trinitrobenzene $\sigma$ complexes, namely, 2.1 × $10^6$ $s^{-1}$ and 1.5 × $10^6$ $s^{-1}$, respectively,[20] we estimate that $k_{-1}'$ should be on the order of $10^4$ $s^{-1}$ ($pK_a$ of imidazole = 14.2) and then $K_{-1}' = k_{-1}'/k_1' \approx 3 \times 10^3$ $M^{-1}$. This is a very large equilibrium constant, and state 4 is thermodynamically favored in Scheme V. The mechanism of Scheme III requires that in Scheme V states 2 and 4 are higher in energy than process 1 and are of similar height;[18] therefore it can be discarded at least for the catalysis by $HO^-$. With less basic catalysts like acetate, for instance, state 2 in Scheme V becomes thermodynamically favored and the mechanism of Scheme III cannot take place.

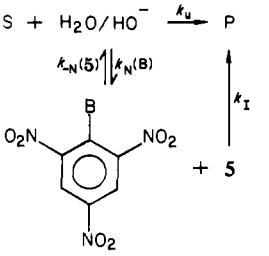
One additional consideration, which disfavors the mechanism of Scheme III, is the fact that it requires that $k_{-h} > k_2^{Bi}(BiH)$; i.e., partially protonated imidazole leaves at a lower rate than the hydroxide ion. The hydroxide ion is more basic than the imidazole anion by 1.7 pK units, and it has been shown that alkoxide ions and amines of about the same p$K$ leave from Meisenheimer complexes at approximately the same rate.[20] Besides, the hydroxide ion is one of the poorest leaving groups known;[14] it leaves more slowly than the methoxide ion even when the latter group is more basic. Then to postulate that partially protonated imidazole will leave more slowly than the hydroxide ion seems totally unrealistic. We then disfavor any mechanism which requires imidazole departure to be rate determining.

A stepwise mechanism as that in eq 17 which involves rate-limiting protonation of SOH would be possible if we could demonstrate that $k_{-Bi}(BiH) < k_{-1}$.

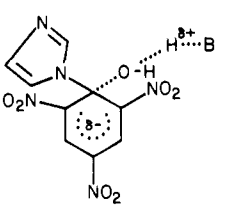
$$S + HO^- \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} SOH \underset{k_{Bi}(Bi)}{\overset{k_{-Bi}(BiH)}{\rightleftharpoons}} HSOH \rightarrow P \qquad (17)$$

When Bi = $CO_3^{2-}$, $K = K_{CO_3H^-}/K_{HSOH} = 3 \times 10^{-4}$, thus $k_{Bi}$ should be on the order of diffusion, $10^{10}$ $M^{-1}$ $s^{-1}$ and then $k_{-Bi} = 3 \times 10^{-4} \times 10^{10} = 3 \times 10^6$ $M^{-1}$ $s^{-1}$. The lowest concentration of $HCO_3^-$ used was 2 × $10^{-3}$ M from which we calculate $k_{-Bi}(BiH)$ ≥ 6 × $10^3$ $s^{-1}$. $k_{-1}$ is not expected to be much higher than 9.8 $s^{-1}$ [14] and then the condition $k_{-Bi}(BiH) < k_{-1}$ is not fulfilled by this buffer. With strongly acidic catalysts the situation is even worse because $k_{-Bi}$ becomes faster and faster until it reaches the value of diffusion.[24] Then a stepwise mechanism as in eq 17 is considered to be very unlikely.

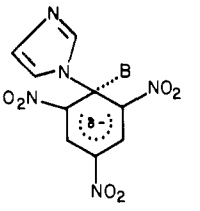
Scheme VI



Since we have discarded all mechanisms where the leaving-group departure is rate determining, buffer catalysis can only be explained by transition state **7** or **8**.



7



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Transition-state **8** corresponds to nucleophilic catalysis and **7** to concerted general-base catalysis of water addition to the aromatic ring (Scheme IV).

If the catalysis is of the nucleophilic type and imidazole is added to the solution, inhibition should be observed (Scheme VI) since in such a case $k_c$ of eq 2 would be given by eq 18. On the other hand, a general-base-catalyzed pathway (Scheme IV) is not expected to be inhibited by imidazole.

$$k_c = \frac{k_N k_1}{k_{-N}(5) + k_1} \qquad (18)$$

With $PO_4H_2^-/PO_4H^{2-}$ buffer acting as the nucleophilic catalyst the intermediate would be 2,4,6-trinitrophenyl phosphate which can go on to products through P–O or C–O bond cleavage ($k_1$ in Scheme VI includes these two pathways).[25]

Inhibition by imidazole would only be possible if attack by imidazole on the aromatic carbon ($k_{-N}$) could compete with $k_1$. At the pH where these rate constants were determined (pH >5.9), if 2,4,6-trinitrophenyl phosphate was formed, it would be in the form of the dianion. Assuming that the increase in rate of hydrolysis ($k_1$) is proportional to the decrease of pH of the leaving group as was reported for 4-nitro- and 2,4-dinitrophenyl phosphate,[26] we estimate $k_1 \approx 1$ $s^{-1}$. On the other hand, $k_{-N}$ is probably on the order of $10^{-2}$ $M^{-1}$ $s^{-1}$.[27] Then the lack of inhibition by imidazole in this particular case is not proof of mechanism. The fact that the catalysis by $PO_4H_2^-$ (or its kinetic equivalent, $PO_4H^{2-}$, reacting with SH) is so much above the Brønsted line defined by the other bases seems to indicate that it reacts by a nucleophilic mechanism.

In the reaction catalyzed by carbonate, there was some inhibition (Table III) at imidazole concentrations higher than 0.1 M which might indicate that the mechanism can be represented

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(19) More O'Ferral, R. A. *J. Chem. Soc. B* **1970**, 274.
(20) Bernasconi, C. F.; Muller, M. C.; Schmid, P. *J. Org. Chem.* **1979**, *44*, 3189.
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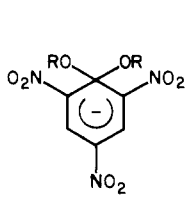
(22) Murto, J.; Murto, M. L. *Acta Chem. Scand.* **1966**, *20*, 297.
(23) Bernasconi, C. F.; Gandler, J. R. *J. Org. Chem.* **1977**, *42*, 3387.
(24) Eigen, M. *Angew. Chem., Int. Ed. Engl.* **1964**, *3*, 1.
(25) Kirby, A. J.; Jencks, W. P. *J. Am. Chem. Soc.* **1965**, *87*, 3209.
(26) Kirby, A. J.; Varvoglis, A. G. *J. Am. Chem. Soc.* **1967**, *89*, 415.
(27) To estimate this rate, we consider that the relative rates of hydrolysis of picryl chloride[5] and 2,4-dinitrofluorobenzene[21] is about 8 and similar ratio of reactivity is expected for imidazole with these two compounds. Moreover, the nucleophilic reactions of aromatic compounds are little dependent on the leaving group.[17]

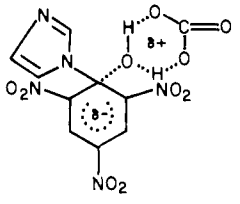
by Scheme VI with B = $CO_3^{2-}$.

However, this depression of rate can totally be accounted for the depression of the uncatalyzed rate of hydrolysis $k_u$. A plot (not shown) of the observed rate constants vs. total buffer concentration gives $k_u = 4.88 \times 10^{-4}$ s$^{-1}$ and $k_c = 1.43 \times 10^{-2}$ M$^{-1}$ s$^{-1}$ when no imidazole is added, whereas it gives $k_u = 3.68 \times 10^{-4}$ s$^{-1}$ and $k_c = 1.38 \times 10^{-2}$ M$^{-1}$ s$^{-1}$ with 0.4 M imidazole. This inhibition can be explained by a complexation of imidazole with the substrate or by an activity coefficient effect. The possibility that $k_{-N} << k_1$ cannot be totally excluded, since we do not know these rates, but, for the acetylimidazolium hydrolysis catalyzed by *N*-methylimidazole inhibition by imidazole takes place at very low imidazole concentrations,[28] which indicates that at least for that reaction imidazole can compete with water as a nucleophile. Besides, in the reaction of 2,4-dinitrophenyl diphenyl carbonate with several tertiary amines, the aminolysis reaction competes with hydrolysis.[29,30]

We think that our data are better explained if the transition state for the reaction catalyzed by bases is represented by **7**. This mechanism has been suggested recently for the addition of water to aromatic substrates.[4,5] Besides, since concerted general-acid catalysis has been observed in the elimination of an alcohol from compounds of type **9**,[31] then the microscopic reverse, i.e., the



**9**



**10**

addition of an alcohol, should be general base catalyzed under comparable reaction conditions and similarly the addition of water should be assisted by general bases. Moreover, the fact that imidazole catalyzes the reaction supports the mechanism proposed because this base cannot react as a nucleophilic catalyst since its reaction with the substrate just leads to regeneration of the starting material.

The Brønsted plot for picrylimidazole is shown in Figure 3 and has a slope of 0.58, typical for a concerted reaction.[32]

The fact that tertiary amines are poorer catalysts than carboxylate ions of the same basicity is puzzling since ammonium ions are much better catalysts for the departure of alkoxide ions from compounds of type **9**.[31] Possible reasons for this discrepancy could be steric repulsion between the imidazole moiety and the bulky tertiary amines on the transition state or the formation of a cyclic transition state **10** which would decrease the requirements of solvation. Such a transition state cannot be formed from compound **9**.

A cyclic transition state involving more than one molecule of solvent was suggested for proton exchange in solutions of benzoic acid in methanol.[33]

**Hydroxide Ion Catalysis.** It is somewhat unexpected that the rate of the hydroxide ion catalyzed reaction falls on the same Brønsted line defined by the other bases because this reaction is a simple proton exchange and does not involve an unfavorable free-energy change. However, a positive deviation from the Brønsted line should be expected if it is reacting by other mechanisms.[34] It is noteworthy that this happens to be the case

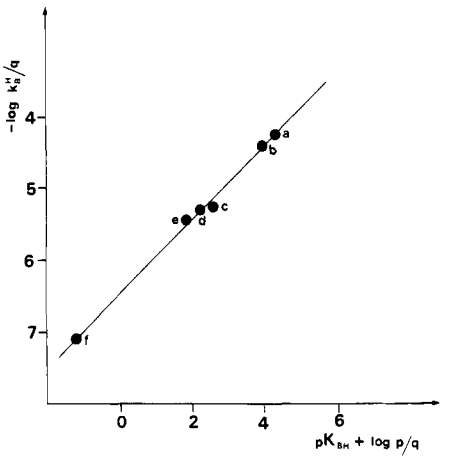



**Figure 4.** Brønsted plot for the general-base-catalyzed hydrolysis of SH: (a) acetate; (b) phthalate; (c) biphthalate; (d) chloroacetate; (e) phosphate monobasic; (f) water.

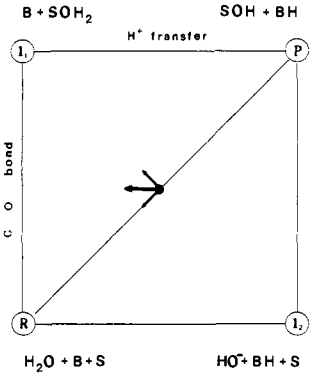



**Figure 5.** Changes in the transition state produced by protonation at the 3-position of the imidazole ring in S.

for all the general-base-catalyzed water reactions with activated aromatic substrates. Ritchie et al.[35] suggested that the proton transfer between water and hydroxide ion places the electrophile in a position originally occupied by the solvation shell of hydroxide ion and this process probably involves an unfavorable free-energy change which is avoided through general-base catalysis.

**Picrylimidazolium Hydrolysis.** The fact that buffer catalysis is observed under conditions where a great fraction of the substrate is protonated represents an evidence against mechanisms which involve imidazole protonation as part of the rate-determining step. The mechanism of catalysis is probably the same as for picrylimidazole although our data do not allow the distinction between nucleophilic and general-base catalysis for this reaction.

The hydrolysis of 3-methyl-1-picrylimidazolium chloride, which is expected to be a good model for picryl imidazolium ion, is catalyzed by acetate and this catalysis is not inhibitable by *N*-methylimidazole,[16] supporting a mechanism involving general-base catalysis. Remarkably, the acetate-catalyzed rate of hydrolysis of the model compound, i.e., $1.13 \times 10^{-4}$ M$^{-1}$ s$^{-1}$, is equal to that of SH.

The Brønsted $\beta$ value obtained from Figure 4 is equal to 0.5 and is within the range expected for concerted general-base catalysis. This value is lower than the $\beta$ value for the hydrolysis of picrylimidazole and this result can be explained considering the More O'Ferrall[19] reaction coordinate diagram of Figure 5.

Protonation of S will stabilize the upper edge $\overline{I_1P}$ relative to the lower $\overline{RI_2}$ which will move the transition state toward S (parallel to the reaction coordinate) and toward $SOH_2$ (perpendicular to the reaction coordinate). Both of these effects tend to reduce the amount of proton transfer; thus $\beta$ decreases as observed.

(28) Oakenfull, J. G.; Jencks, W. P. *J. Am. Chem. Soc.* **1971**, *93*, 178.
(29) Gresser, M. J.; Jencks, W. P. *J. Am. Chem. Soc.* **1977**, *99*, 6963.
(30) It is important to note that in order to inhibit the reaction not only must imidazole compete with water as the nucleophile but also the attack at the aromatic carbon which regenerates the substrate must occur faster than the attack at the carbonyl carbon.
(31) Bernasconi, C. F.; Gandler, J. R. *J. Am. Chem. Soc.* **1978**, *100*, 8117.
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## Experimental Section

**Materials.** Picrylimidazole was synthesized from picryl chloride (1 mmol) and imidazole (2 mmol) in 2 mL of *N,N*-dimethylformamide. After 30 min at room temperature, in the dark, the solution was poured into ice-cooled water, filtered, and washed with water and ethanol. The light yellow crystals decompose at 187-188 °C (lit.[36] 205 °C). The NMR spectrum agrees with that of the literature.[37] Attempts to purify the product by recrystallization or column chromatography resulted in decomposition of picrylimidazole; hence, we use the compound freshly synthesized without further purification. The purity of S was checked by UV-visible spectra under acidic and basic conditions. Under basic conditions (NaOH, 0.01 M) it rapidly hydrolyzes, giving a spectrum coincident with that of picric acid ($\lambda_{max}$ = 358 nm ($\epsilon$ = 15 000 $M^{-1}$ $cm^{-1}$)). In acid solution there was no absorption at $\lambda$ = 358 nm.

Dabco was twice recrystallized from petroleum ether (bp 60-80 °C), mp 155-156 °C. Imidazole was recrystallized from benzene, mp 89-90 °C. Acetic acid was distilled, bp 118 °C. Dioxane was purified by the method of Fieser[38] and was stored over $LiAlH_4$ from which it was distilled as needed. Water twice distilled in a glass apparatus was used throughout.

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

The buffer solutions were prepared from solutions of one of the buffer species of known concentration. After the required amount of acid or base and compensating electrolyte was added, the pH of the solutions were determined. If the pH of the set with constant buffer ratio differ by more than $5 \times 10^{-3}$ pH unit, 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 at each pH.

UV spectra were recorded on a Beckman 24 spectrophotometer, and the change in optical density during a kinetic run was measured on the same instrument at the maximum absorption of picric acid ($\lambda$ = 358 nm). pH measurements were carried out on a Seybold digital pH meter at 25 °C. Standard buffers prepared according to literature[39] were used to calibrate the electrode.

**Kinetic Procedures.** Reactions were initiated by adding the substrate dissolved in dioxane to a solution containing all other constituents. Rate constants were determined by following the appearance of picric acid at 25 °C and $\mu$ = 1 M. All kinetic runs were carried out under pseudo-first-order conditions with substrate concentrations of about $5 \times 10^{-5}$-$4 \times 10^{-4}$ M. Rate constants are accurate to ±3% and were computed from plots of ln ($OD_\infty - OD_t$) vs. time. Most reactions were followed to 80-90% conversion, but for the slowest runs the reaction was followed up to 10-50% conversion. The infinity value for these reactions was determined by hydrolysis of a portion of the substrate solution at high pH. The hydrolyzed substrate was diluted to a concentration suitable for reading the absorption (OD ≈ 0.8) with a buffer of the same pH as the reaction solution. Throughout the paper $(H^+) = 10^{-pH}$.

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# 1,5-Dithia-2,4,6,8-tetrazocine: A Novel Heterocycle of Unusual Properties

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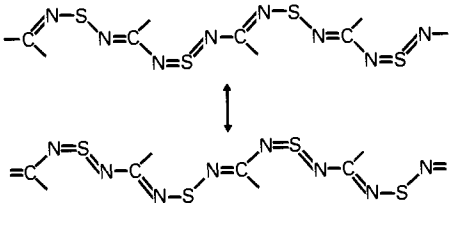
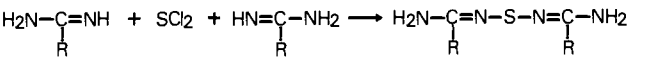
**Abstract:** Four representatives of a novel heterocycle—1,5-dithia-2,4,6,8-tetrazocine—were prepared. In the 3,7-diphenyl (**3**) and 3,7-bis(*p*-methoxyphenyl) (**5**) derivatives, the eight-membered ring was found to be planar by X-ray diffraction analysis; this and the bond lengths as well as the electronic spectra suggest a delocalized, aromatic system of 10 $\pi$ electrons. On the other hand, 3,7-bis(dimethylamino)-1,5-dithia-2,4,6,8-tetrazocine (**4**) is folded along an axis drawn through the sulfur atoms; the short distance between the latter atoms is suggestive of their partial bonding in this compound.

In a research program pursued some time ago at the Woodward Research Institute and directed toward the synthesis of organic conductors, linear polymers of type **1** were considered, among other structures, as promising candidates for conductivity.[2]



**1**

One of the ways considered for the construction of such polymeric compounds involved the intermediates **2**—the expected products from the reaction of two molecules of an amidine and one molecule of sulfur dichloride.



**2**

† Deceased July 8, 1979.

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(2) For a theoretical treatment of the problem of conductivity in this and other types of one- and two-dimensional polymers, see M.-H. Whangbo, R. Hoffmann, and R. B. Woodward, *Proc. R. Soc. London, Ser. A* **366**, 23 (1979).

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