Hydrolyses of 2-Trimethylsilyl N^1-Heterocycles. Involvement of Zwitterions in the Hydrolysis of 2-(Trimethylsilyl)-*N*-methylimidazole, 2-(Trimethylsilyl)pyridine, and 2-((Trimethylsilyl)methyl)-*N*-methylimidazole

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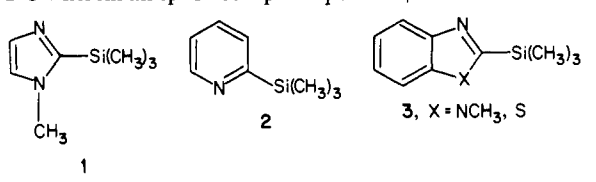
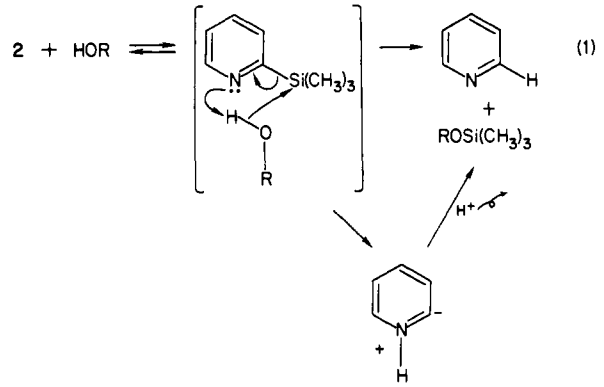
Abstract: The title N-containing aromatic heterocycles (**1**, **2**, and **4**, respectively) with a trimethylsilyl substituent at the 2-position hydrolyze relatively rapidly to form trimethylsilanol and the corresponding 2-H heterocycle. In the case of 2-((trimethylsilyl)methyl)-*N*-methylimidazole (**4**) the heterocyclic product is *N*,2-dimethylimidazole. pH vs. log k_{obsd} profiles establish that the reaction proceeds faster at high than low pH and that above the pK_a of the heterocyclic base the reaction is independent of pH. 2-(Trimethylsilyl)-*N*-methylpyridinium iodide (**6**) and 2-(trimethylsilyl)-*N*,*N*'-dimethylimidazolium iodide (**5**) salts hydrolyze with a first-order dependence on [OH^-] above pH 6 with no evidence for a pH-independent region at high pH. F^- accelerates the reactions of **1** and **2** markedly at low pH but has no effect at pH 12.5. The results indicate that the dominant mechanism for hydrolysis of the title compounds over most of the pH/rate profile involves nucleophilic attack of OH^- on the trimethylsilyl unit of the N-protonated base, the leaving group being the zwitterion (ylide) of the heterocycle. The bimolecular constants for attack of OH^- on protonated **1** and **2** are comparable to those observed for attack of OH^- on **5** and **6**, respectively. Compound **1** below pH 5 is attacked by H_2O as is **4**, but such is not observed for **2**. Only compound **4** suffers nucleophilic attack of OH^- on the deprotonated base. For **1** and **2** at low pH, F^- successfully competes with the intrinsically more nucleophilic OH^- because the latter's concentration is so low, but at high pH the reaction proceeds entirely via OH^- attack.

Rapid solvolytic cleavage of C–Si bonds under mild conditions appears to be a unique process except in heterocyclic systems such as **1–3** wherein an sp^2 N occupies a position β to the C–Si bond[1-3]



1 **2** **3**, X = NCH_3, S

or in cases where the departing C-containing moiety is a moderately good leaving group.[4-6] To account for the greater rate of solvolysis of 2-pyridyltrimethylsilane (**2**) in alkaline (pH >11) over that observed under acidic conditions (pH <4), Webster[3a,b,e] suggested that the process required an unprotonated heterocyclic N to act as an internal general base in assisting the delivery of HOR to the silicon (eq 1). An analogous process was postulated



(1)

by Jutzi et al.[2] for compounds **1–3** with the possible intervention of a zwitterion. Conceptually, the proposed front-side attack[3] on Si was also consistent with the observed retention of configuration in the silinol product of an optically active silypyridine.[3c] However, the latter observation, as pointed out by Seconi and Eaborn,[1] does not necessarily require the mechanism of eq 1. Indeed, the low reported solvent deuterium isotope effect of 1.0–1.2[3e] for the solvolysis of **2** suggests that such an internal delivery is unlikely unless the extent of H^+ transfer in the rate-limiting transition state is highly asymmetric.[6]

During the course of some studies of protecting groups for imidazole,[7] we observed that (*N*-methylimidazol-2-yl)trimethylsilane (**1**) was extremely labile in basic media, the half-life at pH 10 or greater being ~3 s. While such observations for this and related systems has been made before,[2] the general reactivity toward hydrolysis and in our mind never been satisfactorily ex-

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(7) (a) Curtis, N. J.; Brown, R. S. *J. Org. Chem.* **1980**, *45*, 4038–1040. (b) Brown, R. S.; Ulan, J. G., unpublished observations. (c) Repetitive scan UV spectra of **6** (pH 7.7, 0.03 M MOPS) show a dimunitation of a peak at 266 nm ($\epsilon_{T=0}$ = 4.9 × 10^3 M^{-1} cm^{-1}) and an isosbestic point at 256 nm to yield a final spectrum of *N*-methylpyridinium, λ_{max} 258 nm (ϵ_{max} 3.1 × 10^3 M^{-1} cm^{-1}). The initial and final spectra for the hydrolysis of fully protonated **2** (pH 2.0) closely resembles that of **6** (λ_{max}(**2**-H^+) 263 nm (ϵ_{max}(**2**-H^+) 7 × 10^3 M^{-1} cm^{-1}); λ_{max} (protonated pyridine) 254 nm (ϵ_{max} 5 × 10^3 M^{-1} cm^{-1}). Hydrolysis of **5** shows simply a dimunitation of a broad peak at 226 nm.

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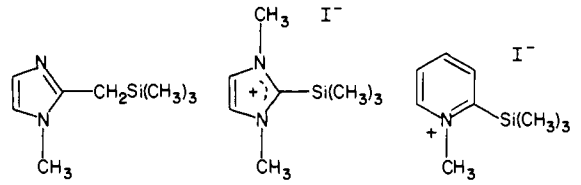
plained. Additionally, heterocycles such as **1–3** appear to exhibit potentially useful synthetic reactions since they add across the C=O unit of nonenolizable ketones, aldehydes, anhydrides, and acyl chlorides to produce alcohols and ketones, respectively.[8] Recently, Moore and Whitesides have demonstrated that **1** and **3** react with a variety of chlorophosphines to form the corresponding heterocyclic phosphine and chlorotrimethylsilane.[9]

In view of the lack of detailed mechanistic information concerning the hydrolysis and peculiarity of the electrophilic additions to such silyl heterocycles, we have undertaken a solvolytic study of **1**, **2**, and **4** as well as that of the *N*-methylimidazolium and -pyridinium analogues **5** and **6**. Also, since the reactivity seems



4 **5** **6**

to be limited to those species having an sp^2 N lone pair adjacent to the C–Si unit, we have determined the UV photoelectron spectra of several substituted imidazoles and pyridines in an effort to ascertain peculiarities in the N lone-pair ionization potentials which might account for the chemical reactivity.

As will be seen, the unique reactivity of each species stems from the intermediacy of a zwitterionic heterocyclic leaving group formed as a result of OH^- attack on the trimethylsilyl unit of the positively charged base. While this possibility was considered by Eaborn and Seconi[1,6] as being in accord with extant data, the following account provides the first clear-cut evidence for such a process.

Experimental Section

Materials. Silanes **1**[2e] and **2**[3e] were prepared in 66% and 63% yield, respectively, according to published procedures.

2-((Trimethylsilyl)methyl)-*N*-methylimidazole (4). To 9.61 g (0.10 mol) of freshly distilled 1,2-dimethylimidazole dissolved in 200 mL of dry THF and cooled to −40 °C under N_2 was added 62.5 mL of 1.6 N *n*-BuLi at such a rate that the flask temperature did not exceed −35 °C. The contents of this flask were added via a cannula to a second flask containing 33 g (0.30 mol) of freshly distilled chlorotrimethylsilane in 100 mL of dry THF cooled to −40 °C. The resultant mixture was allowed to stir and reach room temperature overnight, at which time the solvent and excess $ClSi(CH_3)_3$ were removed by rotary evaporation. To the residue was added 100 mL of H_2O, and the mixture was then extracted with 2 × 150 mL of ether. The combined organic extracts were dried ($MgSO_4$), solvent was removed, and the resulting crude oil (21.2 g) was distilled to yield 10.7 g (63%) of **4** as a mobile clear liquid: bp 58–60 °C/0.6 torr; ^1H NMR ($CDCl_3$) δ 0.87 (s, 9 H, $Si(CH_3)_3$), 2.14 (s, 2 H), 3.54 (s, 3 H), 6.79 (d, 1 H, J = 1 Hz), 6.95 (d, 1 H, J = 1 Hz); IR (film) 3100, 2950, 2880, 1520, 1490, 1450, 1410, 1390, 1280, 1245, 1145, 1080, 850 (br) cm^{-1}; Mass spectrum, m^+/z (intensity) 168 (96.4), 153 (100), 112 (16.2), 96 (15.8), 95 (12.6), 73 (89.5).

2-(Trimethylsilyl)-*N*,*N*′-dimethylimidazolium Iodide (5) and 2-(Trimethylsilyl)-*N*-methylpyridinium Iodide (6). To 10 mL of dry benzene containing 0.01 mol of freshly distilled **1** or **2** was added 7.1 g (0.05 mol) of CH_3I at room temperature, and the stoppered mixture was allowed to stand. In the case of **1**, a white precipitate formed almost immediately, but for **2**, deposition of faint yellow needles required 24 h. The solids were filtered and washed with benzene and then anhydrous ether, with the resultant solids being dried under vacuum for ~7 h. Yield: 1.8 g (61%) of **5**, mp 122 °C dec and 1.25 g (43%) of **6**, mp 129 °C dec. ^1H NMR **5** (0.033 N DCl in D_2O referenced to HOD signal at δ 4.65), δ 0.47 (s, 9 H, $Si(CH_3)$), 3.86 (s, 6 H, NCH_3), 7.36 (s, 2 H); IR (KBr) 3070, 3020, 2990, 1680, 1570, 1480, 1455 (d), 1410, 1390, 1340, 1260, 1250, 1240, 1170, 1100, 850 cm^{-1}. Anal. Calcd for $C_8H_{17}N_2SiI$: C, 32.33; H, 5.72; N, 9.42. Found: C, 32.22; H, 5.49; N, 9.43. ^1H NMR **6** (0.033 N DCl in D_2O referenced to HOD signal at δ 4.65) δ 0.49 (s, 9 H, $Si(CH_3)_3$), 4.43 (s, 3 H, NCH_3), 7.86–8.61 (m, 4 H), 8.82 (d, J = 6 Hz, pyridine 2-H); IR (KBr) 2960, 1600, 1560, 1490, 1440, 1410, 1250, 1140, 1095, 850 cm^{-1}. Anal. Calcd for $C_9H_{10}NSiI$: C, 36.74; H, 5.48; N, 4.76. Found: C, 36.87; H, 5.44; N, 4.81.

(b) Kinetics. Slower kinetic data ($k_{obsd} < 10^{-1}$ s^{-1}) were obtained by observing the rate of change in absorbance of 1 × 10^{-4} M aqueous solutions of **1**, **2**, **4**, **5**, and **6** using a Cary Model 210 UV-visible spectrophotometer previously described.[10] Buffers (0.03 M, ionic strength = 0.03 M $NaClO_4$ except in $HClO_4$ or NaOH cases) were employed to maintain pH; 1–2.6 ($HClO_4$); 3.3–4.3 (formate); 4.7–5.0 (acetate); 5.0–7.0 (MES); 6.0–8.5 (MOPS); 8.3–10.5 (CHES); 9.5–11.5 (CAPS); 11.0–14 (NaOH). Temperature was maintained at 25.0 ± 0.2 °C by means of an external Colora Model NB-ELE water circulating bath and solutions (3.0 mL of buffer in 1-cm quartz cuvettes) were equilibrated in the spectrometer cell holder for 20 min prior to the initiation of a run. Reactions were initiated by injection of 3 μL of an 0.1 M stock solution of silanes **1**, **2**, and **4** in THF, which were prepared freshly each day, into the aqueous buffer. For **5** and **6**, which are insoluble in THF, the 0.1 M stock solutions were made up in 1 × 10^{-3} M HCl and H_2O, respectively. The dimunition in absorbance as a function of time was monitored at 230–240 (**1** and **5**), 275 (**2** and **6**), and 225 nm (**4**). Observed pseudo-first-order rate constants (k_{obsd}) were obtained by fitting the absorbance vs. time data to a standard exponential equation via a nonlinear least-squares treatment.[10] In all cases the reactions were followed to at least 90% completion with the standard deviations in k_{obsd} being 5% or less.

The effects of F^- on the rate of disappearance of **1** and **2** at pH 3.4 and 12.5 were observed by determining the dependence of k_{obsd} on $[F^-]$. Faster kinetic data ($k_{obsd} > 10^{-1}$ s^{-1}) were obtained by using a Durrum–Gibson stopped-flow spectrophotometer interfaced to a microcomputer as previously described[10] and thermostated at 25.0 ± 0.2 °C. Digitally stored absorbance vs. time traces were analyzed by a nonlinear least-squares treatment to give pseudo-first-order rate constants (k_{obsd}), the values being averages of 8–12 separate determinations which are believed accurate to ±5% or better. Into one drive syringe was placed a solution of 2 × 10^{-4} M **1**, **5**, or **6** in 0.01 N $HClO_4$ to retard decomposition. The other drive syringe contained enough $NaClO_4$ to adjust the final ionic strength to 0.03 M after mixing and varying amounts of NaOH and buffer to neutralize the 0.01 N $HClO_4$ and set the final pH to the desired value. pH was measured from the recovered effluent and is assumed to become instantaneously equilibrated upon mixing.

For the pH 12.5 stopped-flow experiments utilizing F^-, the appropriate amount of NaF was placed in the drive syringe containing NaOH solution, and the amount of $NaClO_4$ adjusted accordingly to bring the final value to 0.03 M. However, control experiments established that the k_{obsd} values are insensitive to changes in ionic strength over the range of 0.015–0.10 M.

Activation parameters were obtained by the usual graphical analysis of plots of ln k_{obsd} vs. $1/T$ at five temperatures between 8 and 70 °C, the slopes being evaluated by a linear least-squares treatment.

Solvent isotope effects were monitored at pD 12.5 and 1.6 for **1**, 12.5 for **2** and **4**, and 1.6 for **5**. These regions correspond to the pD (pH) independent regions of k_{obsd} for the various materials. At low pD, two solutions of 0.033 and 0.011 N DCl in D_2O were made by dilution with D_2O (99.7 atom % D) of commercially available 20% DCl in D_2O (Aldrich, 99 atom % D). The high pD solutions were prepared by dilution of a stock 0.12 NaOD (D_2O + metallic Na°). pD values were determined by adding 0.4 unit to the pH reading[11] given by a Radiometer Model GK 2321-C combination electrode.

(c) Photoelectron spectra of various imidazoles were determined with a He I UV-PES machine constructed in our laboratories[12] having an operating resolution of 30–35 meV for the $ArP_{3/2}$–$P_{1/2}$ doublet. Spectra were internally calibrated with argon and CH_3I, which were cointroduced with the sample. Reported values are the averages of 2–3 determinations and are belived accurate to ±0.05 eV.

Results and Discussion

(a) Photoelectron Spectra. Nucleophilic cleavages of $(CH_3)_3Si$–X bonds are generally observed only in those systems where X is a reasonable leaving group[4–6] such as alkoxide,[13]

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Table I. Vertical Ionization Potentials and Assignments for Various Substituted Pyridines and Imidazoles[a,b]

no.	compd	ip_1, eV	ip_2, eV	ip_3, eV	ip_4, eV
1	imidazole[a]	8.89 (π)	10.38 (π,n)	14.03	14.77
2	*N*-methylimidazole[a]	8.69 (π)	9.75 (π,n)	13.12	13.70
3	2-methylimidazole[a]	8.54 (π)	10.11 (π,n)	13.04	13.44
4	1,2-dimethyl-imidazole	8.30 (π)	9.62 (π,n)		
5	*N*-trimethylsilyl-imidazole	8.61 (π)	9.78 (π,n)		
6	*N*-methyl-2-(tri-methylsilyl)-imidazole (**1**)	8.38 (π)	9.50 (n)	9.68 (π)	10.57
7	*N*-methyl-2-((tri-methylsilyl)-methyl)imidazole (**4**)	8.33 (π)	9.69 (π,n)	10.19	
8	pyridine[b]	9.60 (n)	9.75 (π)	10.50 (π)	
9	2-methylpyridine[b]	9.20 (π)	9.50 (n)	10.05 (π)	
10	2-(trimethylsilyl)-pyridine[b] (**2**)	8.90 (n)	9.30 (π)	10.10 (π)	
11	2-(*tert*-butyl)-pyridine	9.22 (n)	9.22 (π)	10.21 (π)	
12	2-((trimethylsilyl)-methyl)pyridine	8.77 (π)	9.16 (n)	10.11 (π)	

[a] Values from ref 20. [b] Values from ref 17.

thiophenoxide,[14] halides,[15] and in some cases benzyl[5g,g] or other carbon-based groups.[6,16] The facile hydrolytic cleavages of trimethylsilyl heterocycles appears to be limited to those systems in which an sp^2 N occupies a position β to the C–Si σ bond. In view of the observed[17-19] hyperconjugative interaction between C–Si σ bonds and adjacent coplanar pairs of electrons, it was of interest to determine whether the PE spectra of **1**, **2**, and **4** would exhibit an unusual destabilization of the sp^2-N_n orbital which might account for the chemical reactivity. For pyridine itself, the assigned ordering is n_N, π, π,[17] with the first two bands being nonresolved. In the case of 2-(trimethylsilyl)pyridine (**2**) C–Si hyperconjugative interaction is observed to selectivity destabilize the N lone pair, which facilitates its ionization relative to the π system[17] and leads to a splitting of the first band. On the other hand, for 2-((trimethylsilyl)methyl)pyridine (entry 12, Table I), extending the C–Si$(CH_3)_3$ bond away from the ring by one CH_2 unit allows a selective hyperconjugative interaction between the C–Si bond and π system which leads to an ordering of π, n_N, π.

Were the analogous situation occurring in the imidazole series, then C–Si–N_n hyperconjugative in **1** should lead to a splitting of the second and third bands from what it is in a suitable model, 1,2-dimethylimidazole, whose assigned ordering is π, π, n_N.[20] Tentatively, when the data for these two are compared (entries 4 and 6, Table I), such is observed, but in both cases the bands are badly overlapped which makes peak position assignment difficult. In any event the n–π splitting in **1** is apparently not large, amounting to ~0.2 eV.

A curious aspect for **4** (entry 7) is that removal of the C–Si-$(CH_3)_3$ unit away from the imidazole by a CH_2 unit produces virtually no change in the observed PE spectrum relative to that of 1,2-dimethylimidazole. The most marked observation within the similarly substituted series (entries 4, 6, and 7) is the general insensitivity of the PE spectra of these imidazoles to the presence and position of the trimethylsilyl group. Hence the PE spectra do not provide interpretable evidence for large differences in the electronic structures which might be related to the hydrolytic reactivity other than a slight destabilization of the n_N orbital which might lead to higher pK_a's for **1** and **2**.

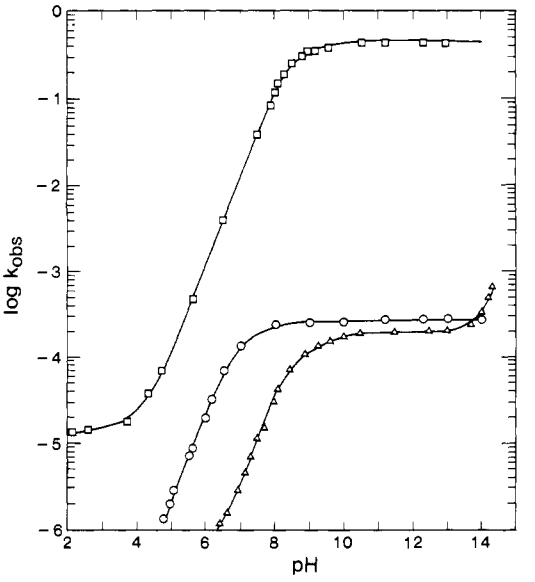


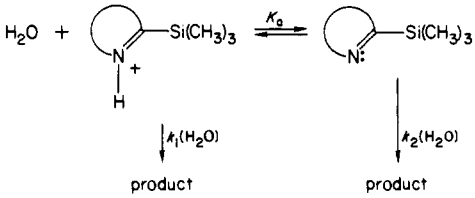
Figure 1. pH vs. log k_{obsd} profiles for the hydrolysis of **1** (□), **2** (○), and **4** (△). Solid lines represent computer fits of the data points to eq 3a (**1**, **2**) and 5 (**4**). T = 25.0 ± 0.2 °C; μ = 0.03 M ($NaClO_4$); buffer strength = 0.03 M except in $HClO_4$ and NaOH media.

Table II. Least-Squares-Generated Constants for the Hydrolysis of **1**, **2**, and **4**[a]

parameter	**1** (eq 3a)	**2** (eq 3a)	**4** (eq 5)
k_1, s^{-1}	1.40×10^{-5}	$<10^{-7}$ [c]	2.78×10^{-7}
k_2, M^{-1} s^{-1}	1.36×10^{5}	2.3×10^{3}	3.81×10^{1}
K_a	3.11×10^{-9}	8.54×10^{-8}	1.94×10^{-9}
pK_a	8.51	7.07	8.71
k_2, s^{-1}	(eq 2a) 0.44[b]	(eq 2a) 2.69×10^{-4} [b]	(eq 4) 1.94×10^{-4} [b]
k_3, M^{-1} s^{-1}			1.88×10^{-4}

[a] Obtained by fitting kinetic data for **1**, **2**, and **4** to indicated expressions, see text. [b] k_2 obtained by fitting data for **1** and **2** to eq 2a, and **4** to eq 4; other parameters are the same as those obtained from eq 3a and 5, respectively. [c] k_1 for hydrolysis of **2** not observable; value represents estimated upper limit.

(b) Hydrolysis of 1, 2, and 4. Given the shapes of the pH vs. log k_{obsd} profiles for **1** and **2** (Figure 1) there appear to be (at least) two kinetically indistinguishable processes which account for the hydrolysis. These are shown in eq 2 and 3 with the derived dependence of k_{obsd} on $[H^+]$ being given in eq 2a and 3a, respectively. Nonlinear least-squares fitting of eq 2a or 3a to the



$$k_{obsd} = \frac{k_1[H^+] + k_2K_a}{K_a + [H^+]} \qquad (2a)$$

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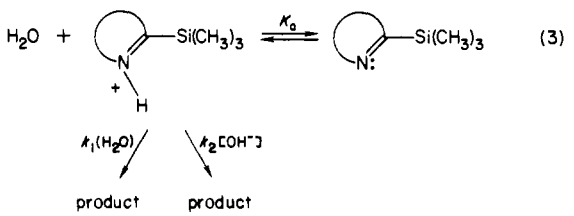


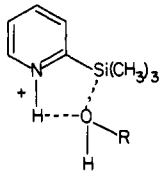
(3)

$$k_{obsd} = \frac{k_1[H^+] + k_2K_w}{K_a + [H^+]} \qquad (3a)$$

$$K_w = [H^+][OH^-] = 10^{-14}$$

observed data for the hydrolysis of **1** and **2** generates the solid curve shown in Figure 1 which can be seen to fit the data well, with the attendent parameters being given in Table II. Conceptually the difference between the mechanisms given in eq 2 and 3 relates to whether nucleophilic attack on the silicon occurs on only one or both of the protonated or neutral forms of the heterocycles, these being connected by the respective acid dissociation constants (K_a). Thus the difference between eq 2a and 3a relates to the interpretation of the second term in the numerator. For eq 2a, k_2 relates to the attack of H_2O on the neutral heterocycle, while in eq 3a, k_2 refers to attack of OH^- on the protonated form. That the latter process should be independent of pH above the pK_a of the heterocycle is best envisioned considering that for each pH unit increase, the [base–H^+] decreases by a factor of 10 while [OH^-] increases by 10-fold, the net effect on k_{obsd} showing no apparent dependence on pH.

The kinetic pK_a of 7.07 for **2** compares favorably with the thermodynamic value of 6.63 obtained by Webster et al.[3d] by half-neutralization techniques. Apparently the 2-trimethylsilyl substituent enhances the basicity of the pyridine N by 1.4–1.8 pK_a units relative to that of pyridine itself (pK_a = 5.21[21]). The same appears true for **1**, the pK_a of which is 1.5 units larger than that reported for *N*-methylimidazole (7.0[21]). Although Webster[3d] attributed the high pK_a for **2** to its facility in forming as associated complex akin to **7**, we feel that a more likely explanation would



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be related to the previously discussed hyperconjugative interaction between the coplanar C–SiR_3 bond and adjacent N lone pair[17] which destablizes the latter and increases the basicity.

For the hydrolysis of **4**, the onset of a second [OH^-]-dependent pathway at high pH (Figure 1) can be best accommodated by either of two kinetically equivalent mechanisms akin to those given in eq 2 or 3 with an additional step whereby OH– nucleophilically captures the unprotonated form (k_3[OH^-]). The derived expressions relating k_{obsd} as a function of [H^+] are given in eq 4 and 5. Nonlinear least-squares fitting of these expressions to the

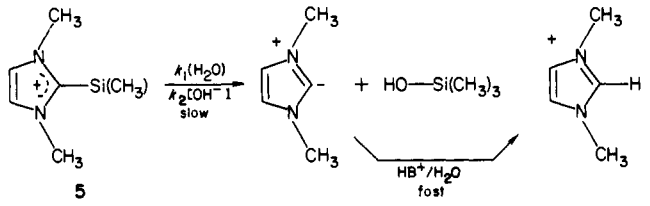
$$k_{obsd}(4) = \frac{k_1[H^+]^2 + k_2[H^+]K_a + k_3K_wK_a}{K_a[H^+] + [H^+]^2} \qquad (4)$$

$$k_{obsd}(4) = \frac{k_1[H^+]^2 + k_2K_w[H^+] + k_3K_wK_a}{K_a[H^+] + [H^+]^2} \qquad (5)$$

observed data for **4** generates the solid curve shown in Figure 1 with the derived values for the constants being given in Table II. Once again the distinction between eq 4 and 5 rests on the interpretation of the meaning of the second term in the numerator.

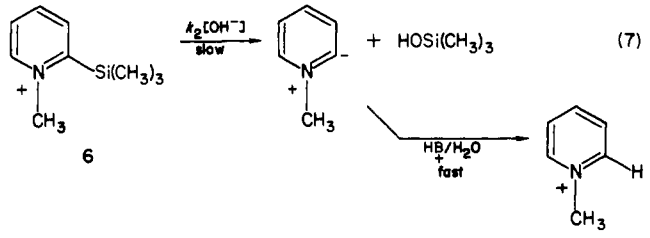
On the basis of the above experiments it is not possible to decide which of the two fundamentally different but kinetically indistinguishable mechanisms is operative. Intuitive bias favors the most nucleophilic species in solution (OH^-) as being responsible for the observed hydrolytic behavior (hence the mechanisms of 3 and 5 for **1**, **2**, and **4**, respectively). Indeed solvolyses of most $(CH_3)_3Si$–X species show first-order dependencies of the reaction rates on [OS^-].[13,22] Consideration of the magnitude of the solvent isotope effects for **1**, **2**, and **4** (Table III) also does not allow one to distinguish between the two mechanisms since the observed k_{H_2O}/k_{D_2O} values at pH 12.5 range from 1.0 in the case of **2** to 2.6 in the case of **1**. Depending upon the system one might argue for or against Webster's general base role for the heterocyclic N.[3] However, as pointed out by Jencks,[23] the interpretation of solvent isotope effects is often not a simple matter.

Distinction between the two mechanisms is most easily made after consideration of the pH vs. log k_{obsd} profiles of the methylated salts **5** and **6** given in Figure 2.[7c] These represent reasonable models for the protonated forms of **1** and **2**, respectively, but undergo no pH-dependent acid dissociation. Also shown in Figure 2 for the sake of comparison are broken lines which represent the hydrolysis profiles for **1** and **2**. At low pH the profile for **5** is remarkably similar to that of **1** in that it shows a pH-independent region which is clearly attributable to attack of H_2O on the trimethylsilyl unit. Above pH 6 the rate of hydrolysis of **5** increases linearly with [OH^-] as far as pH 12, above which the process is too fast to be monitored by stopped-flow techniques (k_{obsd}, pH $\simeq$100 s^{-1}). A similar situation exists for **6** from pH 6 to 12.5. On the basis of the above the only possibly hydrolytic processes which accommodate these data are given in eq 6 and 7, the values of the constants being graphically obtained. The



(6)

k_1 = 8.1 x 10^{-5} s^{-1}
k_2 = (1.3±0.2) x 10^4 $M^{-1}s^{-1}$



(7)

k_2 = (2.0±0.2) x 10^3 M^{-1} s^{-1}

numerical values of k_2, the bimolecular rate constant for OH^- attack, on **5** and **6** compare favorably with the analogous k_2 values (Table II) obtained by fitting the hydrolysis data for **1** and **2**, respectively, to eq 3a. Indeed from Figure 2 the regions of the pH log k_{obsd} profiles of **6** and **2** with slope of +1 lie essentially on top of each other, while the same region for **5** is displaced by $\sim$1 pH unit relative to that for **1**.

Given the above information, the likely mechanism of hydrolysis for **1**, **2**, and related heterocyclic trimethylsilanes is that of eq 3 (eq 5 in the case of **4**) whereby the dominant reaction involves nucleophilic attack of [OH^-] on the protonated heterocycle. In the case of **2** this pertains to the entire observable pH range while for **1** (and **5**) below pH $\simeq$4 the [OH^-] is so low that H_2O acts as the nucleophile. Both **5** and **1** show evidence of catalysis by

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Table III. Activation Parameters and Solvent Deuterium Isotope Effects for Hydrolysis of **1**, **2**, **4**, and **5**

compd	pH[a]	$k_{obsd}^{25°C}$, s^{-1}	E_a, kcal/mol	log A	k_{H_2O}/k_{D_2O}[b]
1	1.6	$(1.49 \pm 0.25) \times 10^{-5}$	21.1 ± 0.1	10.64 ± 0.02	2.3
1	12.5	0.42 ± 0.02	11.5 ± 0.1	8.07 ± 0.02	2.6
2	12.5	$(2.75 \pm 0.02) \times 10^{-4}$	17.0 ± 0.2	8.87 ± 0.02	1.05
4	12.5	$(2.00 \pm 0.02) \times 10^{-4}$	14.7 ± 0.1	7.11 ± 0.01	1.32
5	1.6	$(8.15 \pm 0.01) \times 10^{-5}$	20.8 ± 0.3	11.20 ± 0.03	1.85

[a] pH independent regions. [b] 25.0 °C.

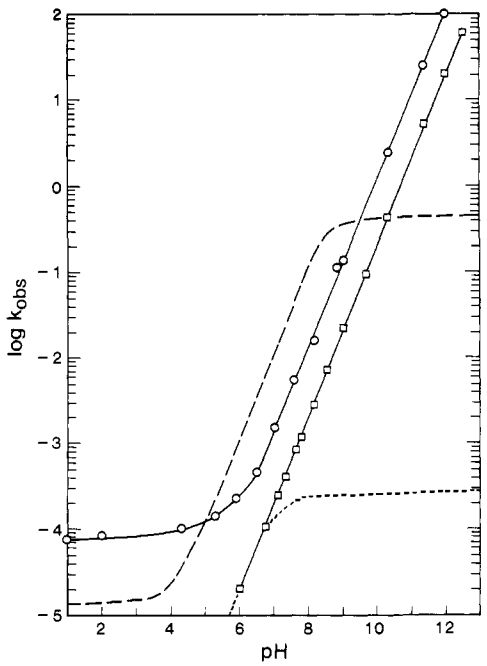


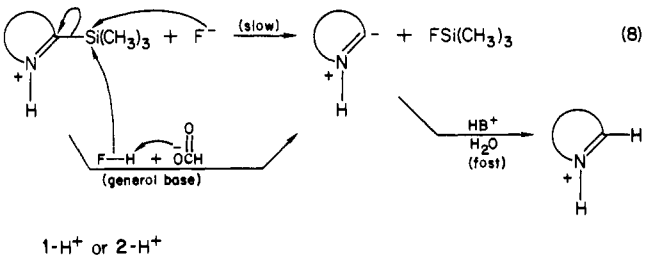
Figure 2. pH vs. log k_{obsd} profiles for the hydrolysis of **5** (○) and **6** (□). Broken lines represent profiles for **1** (--) and **2** (---) transcribed from Figure 1. $T = 25.0 \pm 0.2$ °C; $\mu = 0.03$ M ($NaClO_4$); buffer strength = 0.03 M except in $HClO_4$ or NaOH.

formate and acetate buffers. Judging from the slopes of plots of k_{obsd} vs. buffer this is probably of the general base type since for **1** the slopes of the lines increase with pH (formate, pH 3.7, slope = 9.06×10^{-4} M^{-1} s^{-1}, formate, pH 4.3, slope = 9.6×10^{-4} M^{-1} s^{-1}). For **5** at pH 4.3 (formate) the comparable slope of the plot is 12×10^{-4} M^{-1} s^{-1}.

The buffer catalysis is also consistent with the observed small solvent isotope effects of 2.3 and 1.85 for **1** and **5**, respectively, at pH 1.6 (Table III) which could arise from a general base assistance of the delivery of HO–H by solvent. Furthermore it is clear that the activation energies for **1** and **5** at pH 1.6 are virtually identical, consistent with the attack of a similar weak nucleophile (H_2O) in both cases. It is slightly disturbing that **1** at pH 12.5 exhibits a relatively large solvent isotope effect of 2.6 even though in the simplest scheme no X–H bonds are being broken or formed. Tentatively one might speculate that the effect arises from competing trimolecular pathways in which OH^- assists in the delivery of H–OH to the silicon of the protonated heterocycle or alternatively that in concert with direct OH^- attack, the solvent is undergoing proton transfer to the unprotonated heterocycle. A similar sequence of events might also account for the isotope effects in **4**. Some support of the latter events comes from earlier experiments[7b] which showed buffer catalysis of the hydrolysis of **1** at pH 10–11 in the presence of 0.1–0.3 M CAPS.

Just how much more active toward the (+) heterocycle is OH^- than H_2O can be evaluated by comparing the second-order rate constants for OH^- and H_2O attack on **1** or **5** at low pH (Table II and eq 6, respectively). Toward these two species, OH^- is (5×10^{11})-fold and (8.8×10^9)-fold more active than H_2O, while in the case of **4** it is (7.5×10^9)-fold more active. Unfortunately since the H_2O attack is not observed in the case of **2** or **6**, no direct comparisons can be made with any surity, but from the data of Table II it is reasonable to place a lower limit of 10^{12}-fold greater nucleophilicity for OH^-.

(c) Effect of [F^-]. At pH 3.4 (this value chosen because it represents the pK_a of HF[24]), a linear dependence of the rate of disappearance of both **1** and **2** upon [F^-] is observed, the second-order rate constants being 59 and 3.56×10^{-3} M^{-1} s^{-1}, respectively, these values being obtained from the slopes of plots of k_{obsd} vs. [F^-]. Since both heterocycles are essentially completely protonated at this pH, the reaction likely involves rate-limiting S_N2 attack of F^- on the trimethylsilyl unit of the protonated base as in eq 8. Since halides such as Cl^- and I^- do not accelerate



the reactions it can be assumed that the high Si–F bond strength of 139 kcal/mol[25] provides the driving force for the reaction. Although effective as a nucleophile, F^- is (2.3×10^3)-fold and (6.4×10^5)-fold less potent at attacking the $(CH_3)_3Si$ group of **1** and **2** than is OH^- (Table II). It is of interest that buffer catalysis of the attack of fluoride at pH 3.4 is evident for the decomposition of **1**, since plots of k_{obsd} vs. [formate] in the presence of 3.3×10^{-4}, 1.0×10^{-3}, and 2.0×10^{-3} M added [NaF] give slopes of 4.52×10^{-2}, 1.26×10^{-1}, and 2.32×10^{-1} M^{-1} s^{-1}, respectively. By analogy to the observed buffer catalysis seen for **1** and **5**, this may be envisioned in terms of a general base deprotonation of attacking HF by the anionic form of the buffer as in eq 8.

Finally we consider the lack of observed effect of F^- on the hydrolysis of **1** or **2** at pH 12. Since F^- is 10^3–10^5-fold less effective a nucleophile than OH^-, it is incapable of competing unless its concentration is 10^3–10^5-fold larger than [OH^-]. The only reason that F^- has such a dramatic effect at pH 3.4 is that the [OH^-] is extremely low and OH_2 is a far less effective nucleophile with which F^- can effectively compete.

At first glance it might seem surprising that the hydrolysis mechanisms involve what appears to be a rather unstable intermediate having an anion at the 2-position of the protonated heterocycle. However, it must be recalled that the leaving group in these cases is a zwitterion and formally neutral. In the case of **4**, OH^- attack on the protonated form occurs ~10^5-fold faster than on the unprotonated form (Table II). For the latter event, the *N*-methylimidazol-2-ylmethyl anion is delocalized throughout the π system and therefore somewhat more capable of being a leaving group than the pyridyl or imidazolyl anions which are heavily localized in an sp^2-like orbital at C_2. Indeed the fact that OH^- cannot be induced to attack the unprotonated forms of **1** or **2** indicates that neutralization of charge (to form the ylides) is crucial to converting these heterocycles into reasonable leaving groups.

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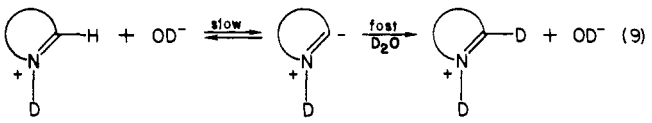
Moreover the process has ample literature precedent in base-catalyzed deuterium exchange of C–H bonds adjacent to a heterocyclic sp^2 N.[26] It is a general observation that pyridines,[27] imidazoles,[28] and benzimidazoles[29] exchange their 2-hydrogens via an OH^--mediated deprotonation of the C_2-H of the N-protonated base as in eq 9.



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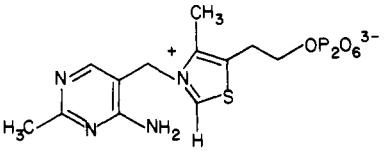
Zoltewicz[26c] has provided evidence that a (+) annular nitrogen atom in an aromatic ring activattes the adjacent sp^2 C–H for deprotonation via ylide formation by a factor of 10^{14}–10^{16}: in the case of imidazole, N-protonation reduces the pK_a of the 2-H to values approaching 17.[27c]

Nature too employs the activating effects of an adjacent (+)-N in facilitating ylide formation within a thiazolium ring in the many reactions catalyzed by the coenzyme thiamine pyrophosphate (**8**).[28]



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Registry No. 1, 35342-89-3; **2**, 13737-04-7; **4**, 91631-71-9; **5**, 91631-72-0; **6**, 91631-73-1; F^-, 16984-48-8; D_2, 7782-39-0; 1,2-dimethylimidazole, 1739-84-0; chlorotrimethylsilane, 75-77-4.

Substituent Effects on the Stability of Three-Electron-Bonded Radicals and Radical Ions from Organic Sulfur Compounds

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Abstract: Optical absorption spectra have been measured for a variety of $(R_2S.\therefore SR_2)^+$, $(RS.\therefore SR)^-$, $R_2S.\therefore Br$ radicals and radical ions by using pulse radiolysis techniques. The results substantiate that the optical properties of such species are a sensitive measure for the strength of their three-electron (2σ, $1\sigma^*$) bonds. A considerable red shift in λ_{max} parallels the inductive power of the substituents R. This is explained by bond weakening due to electron induction into the antibonding σ^* orbital. Linear free energy correlations allow us to quantify the inductive effect and, in addition, to evaluate the influence of steric hindrance on the stability of the three-electron bonds.

The formation of three-electron-bonded radicals which are characterized by a relatively weak bond between two hetero atoms has been well established over recent years in several liquid- and solid-state studies. Typical examples are $(R_2S.\therefore SR_2)^+$,[1-9] $(R_2Se.\therefore SeR_2)^+$,[10] $(>N.\therefore N<)^+$,[11-13] and $(>S.\therefore N<)^+$[14,15] radical cations (the latter two refer to intramolecularly stabilized species), $R_2S.\therefore SR$[16-18] and $R_2S.\therefore X$ (X = halide)[3,19-21] neutral radicals, $(RS.\therefore SR)^-$[16,22] and $(RS.\therefore X)^-$[23] radical anions, and others (although not always assigned to such structures). All these three-electron bonds contain two bonding σ electrons and one antibonding σ^* electron. The bond-weakening effect of the latter,

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