Mechanism of Nucleophilic Substitution Reactions of *o*-Nitrobenzenesulfenamides: Evidence for a Substitution Proceeding through a Sulfuranide Intermediate[1]

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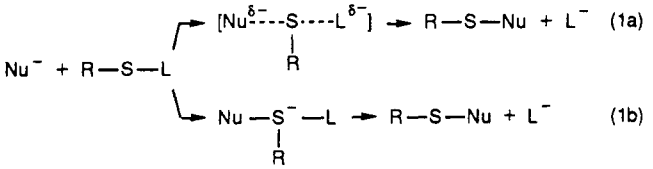
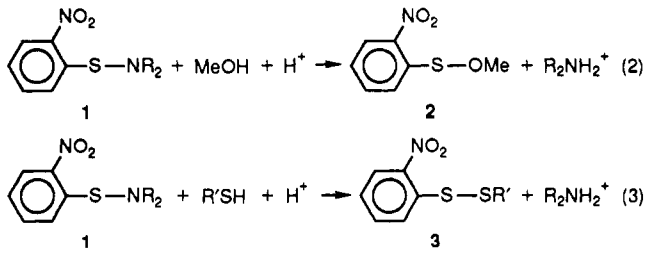
o-Nitrobenzenesulfenamides (1) undergo acid-catalyzed methanolysis (eq 2) in acetonitrile–methanol, affording methyl *o*-nitrobenzenesulfenate (2). They also undergo acid-catalyzed reaction with a thiol, giving an alkyl *o*-nitrophenyl disulfide (eq 3). In MeCN–8.2 M MeOH the thiolysis reaction exhibits a simple first-order dependence on both [R'SH] and [H^+] for acid concentrations in the range 0.001–0.012 M. In marked contrast the rate of the methanolysis reaction is only slightly dependent on [H^+] for the same range of acid concentrations. It is also markedly dependent on [MeOH], being seven times faster in MeCN–8.2 M MeOH than in MeCN–3.0 M MeOH. Measurement of the equilibrium constant for protonation of 1 to 1-H^+ in MeCN–8.2 M MeOH shows that under the conditions used for the kinetic studies of eqs 2 and 3 the extent of protonation of 1 to 1-H^+ is always small (<6%). Consideration of this fact, and the markedly different kinetic behavior of the two substitution reactions, leads to the following interpretation: (a) for the thiolysis the rate-determining step is the reaction of the thiol with 1-H^+; (b) for the methanolysis, on the other hand, the mechanism is that shown in eq 9, where a sulfuranide species 4 is the key intermediate on the reaction coordinate. The lack of dependence of the rate on [H^+] is because acid-catalyzed reversion of 4 to 1-H^+ and methanol (step k_{-6}) is faster than the cleavage of 4 (step k_7) to give 2 and R_2NH. The methanolysis of 1 therefore represents another substitution at dicoordinate sulfur (similar to ref 10) where the kinetics require a mechanism with a hypervalent sulfur species as an intermediate on the reaction coordinate.

Over the years one of the important and still not fully resolved questions regarding nucleophilic substitutions at dicoordinate sulfur is the extent to which such reactions proceed by a mechanism where bond making and bond breaking are synchronous (eq 1a) as opposed to one where bond making precedes bond breaking and a sulfuranide intermediate (Nu–S^-–L) is present on the reaction coordinate (eq 1b).[2]



We recently had occasion, in connection with another investigation, to conduct a study of two acid-catalyzed substitution reactions of *o*-nitrobenzenesulfenamides (*o*-$O_2NC_6H_4SNR_2$, 1): (1) their methanolysis (eq 2) and (2) their reaction with a thiol (eq 3).



The results of this study comprise the present paper. In our opinion the kinetic results for the methanolysis (eq 2) turn out to be compatible only with a two-step mechanism for the substitution where a hypervalent sulfur species is present as an intermediate. The present study therefore provides additional important kinetic evidence that some nucleophilic substitutions at dicoordinate sulfur take place by mechanisms where sulfuranides are intermediates, rather than by one-step mechanisms where bond formation and cleavage are synchronous.

Results

Methanolysis of *o*-Nitrobenzenesulfenamides. When either *N,N*-dimethyl-*o*-nitrobenzenesulfenamide (**1a**, R = Me) or *N*-(*o*-nitrophenylsulfenyl)morpholine [**1b**, R_2 = $(CH_2CH_2)_2O$] was dissolved in acetonitrile–methanol (2:1) and trifluoromethanesulfonic acid (CF_3SO_3H) was added (2 mol/mol of 1), the sulfenamide was found to disappear and be replaced by methyl *o*-nitrobenzenesulfenate (**2**, eq 2). Workup of the reaction solution after 6 h gave **2** as the only sulfur-containing product.

The kinetics of the acid-catalyzed methanolysis of 1 (no reaction takes place in the absence of an acid catalyst) were studied in acetonitrile–methanol mixtures containing from 2.0 to 8.2 M MeOH. The ionic strength was kept constant at 0.2 by the addition of lithium trifluoromethanesulfonate (CF_3SO_3Li). The progress of the methanolysis was followed by monitoring the increase in the absorbance of the solution with time at 415 nm that accompanies the conversion of 1 to 2. In all cases the methanolysis of 1 followed good first-order kinetics. The experimental first-order rate constants (k_1) for the disappearance of 1 under varying conditions are found in Table I.

The following important observations emerge from examination of Table I: (1) although no reaction takes place in the absence of added trifluoromethanesulfonic acid, in MeCN–8.2 M MeOH the rate for **1a** or **1b** is little dependent on [CF_3SO_3H] over the range of acid concentrations (0.003–0.016 M) investigated; (2) in MeCN–3.0 M MeOH, on the other hand, the rate is much more dependent on [CF_3SO_3H], but with plots of k_1 vs [H^+] showing a decreasing dependence on [H^+] at higher [CF_3SO_3H]; (3) the rate is quite sensitive to methanol concentration, being seven times faster in MeCN–8.2 M MeOH than in MeCN–3.0 M MeOH; (4) the rate of methanolysis of **1a** is somewhat faster than that of **1b**.

Reaction of a Thiol with *o*-Nitrobenzenesulfenamides. The reaction of 1-butanethiol (*n*-BuSH) with **1a** and **1b** (eq 3) was studied in acetonitrile–methanol

(1) This research supported by the National Science Foundation, Grant CHE-9000175.

(2) For a review listing the results of many studies pertinent to this question see: Kice, J. L. *Adv. Phys. Org. Chem.* **1980**, *17*, 140–145.

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Table I. Kinetics of the Methanolysis of o-Nitrobenzenesulfenamides at 25 °C in Acetonitrile–Methanol[a]

sulfenamide	[MeOH], M	$[CF_3SO_3H]$ ×10², M	$k_1 \times 10^3$, s^{-1}
1a	8.2	0.53	0.25
		0.80	0.28
		1.1	0.30
		1.6	0.32
	5.2	0.50	0.092
	3.0	0.23	0.020
		0.45	0.032
		0.68	0.043
		0.91	0.050
	2.0	0.50	0.025
1b	8.2	0.40	0.15
		0.80	0.19
		1.2	0.21
		1.6	0.25
	3.0	0.39	0.020
		0.77	0.030
		1.2	0.045
		1.54	0.047

[a] All runs at ionic strength = 0.2 (maintained by addition of CF_3SO_3Li). Initial concentration of 1, 0.0001–0.00024 M.

Table II. Kinetics of the Reaction of 1-Butanethiol with o-Nitrobenzenesulfenamides at 25 °C in Acetonitrile–Methanol[a]

sulfenamide	[MeOH], M	10^3[RSH], M	$10^2[CF_3SO_3H]$, M	$k_1 \times 10^2$, s^{-1}	$k_{RSH} = k_1K_b/[RHS]\alpha$[b]
1a	1.0	1.0	0.30	4.3	26×10^3
			0.20	3.5	
			0.10	1.5	
		3.0	0.10	4.8	22×10^3
		5.0		7.7	21×10^3
	3.0	2.0	0.57	9.3	9.3×10^3
	5.0			4.7	4.3×10^3
		1.0	0.30	1.24	4.2×10^3
	8.2	1.0	0.30	0.62	
			0.60	1.36	
			0.90	1.84	2.2×10^3
			1.2	2.6	
		2.0	0.30	1.2	
		3.0		1.8	2.0×10^3
		4.0		2.4	
		5.0		3.0	
		2.0	1.2	5.3	
		3.0		8.4	
		4.0		10.2	2.3×10^3
		5.0		14	
1b	1.0	1.0	0.35	2.8	
			0.70	4.9	7.9×10^3
			1.05	7.0	
			1.4	9.4	
		1.5	0.70	6.7	
		2.0		9.1	7.4×10^3
		2.5		12.1	
	3.0	2.0	0.60	2.6	2.2×10^3
	5.0			1.0	0.84×10^3
	8.2			0.50	0.42×10^3
	8.2	10.0	0.20	0.63	
			0.40	1.4	0.37×10^3
			0.80	3.0	
			1.2	4.4	
	8.2	15.6	0.10	0.67	
		31		1.3	0.42×10^3
		47		1.9	
		62		2.6	

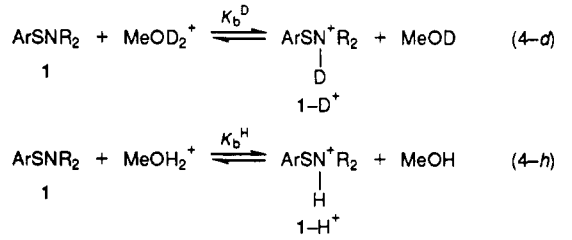
[a] All runs at ionic strength = 0.2 (maintained by addition of CF_3SO_3Li), initial concn of 1, 0.0001–0.00024 M. [b] $\alpha = K_b[H^+]/(K_b[H^+] + 1)$. For runs at constant [RSH]: from slope/[RSH] of a plot of k_1 vs $[CF_3SO_3H]/(K_b[CF_3SO_3H] + 1)$. For runs at constant $[CF_3SO_3H]$: from slope $(K_b[CF_3SO_3H] + 1)/[CF_3SO_3H]$ of a plot of k_1 vs [RSH].

([MeOH] = 1.0–8.2 M) in the presence of CF_3SO_3H as the acid catalyst and with the ionic strength maintained constant at 0.2 by the addition of lithium trifluoromethanesulfonate. The stoichiometry of the reaction was established to be that of eq 3 by observing the change in the ^1H NMR spectrum that took place when CF_3SO_3H (0.15 mmol) and 1-butanethiol (0.05 mmol) were added to a solution of sulfenamide **1b** (0.05 mmol) in 1 mL of CD_3CN–8.2 M CD_3OD. Two minutes after the addition the ^1H NMR had changed from that associated with the sulfenamide to that associated with an equimolar mixture of *n*-butyl *o*-nitrophenyl disulfide (**3**, R = *n*-Bu) and morpholinium ion. No further change occurred upon allowing the solution to stand for an additional 2 h.

The kinetics of the reaction were followed by observing the decrease in the optical density of the reaction solution at 405 nm that accompanies the conversion of **1** to **3**. Under the reaction conditions employed the thiol was present in considerable stoichiometric excess over the sulfenamide, and as a consequence, the disappearance of the sulfenamide followed good first-order kinetics. The experimental first-order rate constants (k_1) for the disappearance of the sulfenamides through reaction with the thiol under the different reaction conditions (varying thiol concentration, varying $[CF_3SO_3H]$, and variation in [MeOH]) are presented in Table II.

Note that the reactions show a first-order dependence of the rate on both thiol concentration and $[CF_3SO_3H]$. This contrasts with the lack of dependence of the rate of methanolysis (see Table I) on $[CF_3SO_3H]$. The rate of reaction of **1** with *n*-BuSH decreases significantly with increasing [MeOH] in the methanol–MeCN mixtures used, being about eight times slower in MeCN–8.2 M MeOH than in MeCN–2.0 M MeOH. This also contrasts sharply with the behavior of the methanolysis of **1**, where the rate *increases* markedly with increasing [MeOH]. Sulfenamide **1a** undergoes thiolysis about five times faster than **1b** (in MeCN–8.2 M MeOH).

Basicity of Sulfenamides in Acetonitrile–Methanol. The equilibrium constant (K_b^D) for protonation of sulfenamide **1a** (eq 4-*d*, Ar = o-$O_2NC_6H_4$, R = Me) in acid



solution in MeCN–CD_3OD could be determined accurately by appropriate measurements of the change in the chemical shift of the $(CH_3)_2N$ protons in **1a** as a function of the trifluoromethanesulfonic acid concentration of the solution. The details of the procedure are given in the Experimental Section.

The equilibrium constant (K_b^H) for the corresponding equilibrium in MeCN–CH_3OH (eq 4-*h*), which would, of course, be expected to be smaller than K_b^D by a factor of around 2,[3] could also be evaluated for either **1a** or **1b** (Ar = o-$O_2NC_6H_4$, $R_2 = (CH_2CH_2)_2O$) by making use of the changes in the ultraviolet spectra of the sulfenamides that accompany their protonation to 1-H^+.

The values for K_b^D and K_b^H for the two sulfenamides in different MeCN–methanol solvent mixtures are tabulated in Table III. From them it is evident that K_b increases with decreasing [MeOH], the increase being particularly marked when the methanol concentration drops below 3.0 M.

Given the lower basicity of morpholine as compared to dimethylamine,[4] it is not surprising that K_b^H for sulfen-

(3) Gold, V.; Grist, S. *J. Chem. Soc. B* **1971**, 1665. Kresge, A. J.; Allred, A. L. *J. Am. Chem. Soc.* **1963**, *85*, 1541.

(4) pK_a of morpholine-H^+, 8.70: Albert, A. *Biochem. J.* **1950**, *47*, 531. pK_a of $Me_2NH_2^+$, 10.8: Somerville, W. C. *J. Phys. Chem.* **1931**, *35*, 2412.

Table III. Basicity of Sulfenamides in Acetonitrile–Methanol

sulfenamide	solvent	K_b^D, M^{-1}	K_b^H, M^{-1}	pK_a of 1-H$^+$	solv. iso. effect, (K_b^D/K_b^H)
1a	MeCN–8.2 M MeOH	6.8	3.6	0.56	1.9
	MeCN–5.0 M MeOH	13.9	8.8	0.94	1.6
	MeCN–3.0 M MeOH		23.7		
	MeCN–1.0 M MeOH		3.6×10^2		
1b	MeCN–8.2 M MeOH		0.33	−0.48	
	MeCN–5.0 M MeOH		0.65	−0.19	
	MeCN–3.0 M MeOH		1.3		
	MeCN–1.0 M MeOH		14.3		

amide **1b** is from 10 to 25 times smaller (depending on [MeOH]) than K_b^H for sulfenamide **1a**.

The K_b values in Table III are expressed in terms of *concentrations*, i.e., $K_b = [1\text{-H}^+]/[\text{H}^+][1]$, rather than *activities*. It seemed likely that the reason for the increase in K_b with decreasing methanol concentration was due to a marked change in the proton activity of a solution containing a given [CF_3SO_3H] as the proportion of the hydroxylic solvent was reduced. This was confirmed by measuring the degree of protonation of two of the weak base amines used to establish the H_o scale[5] in H_2SO_4–H_2O in MeCN–MeOH media of varying methanol content. These measurements confirmed that [amine-H$^+$]/[amine] varied with [H$^+$] and methanol concentration in the same fashion as for the sulfenamides.

The data on K_b^H for **1a** and **1b** in MeCN–8.2 M MeOH indicate that under the conditions used for the kinetic experiments in Tables I and II the extent to which **1** is protonated to 1-H$^+$ at equilibrium is always very small, ranging from as little as 0.03% for **1b** in the presence of 0.001 M CF_3SO_3H to 5.4% for **1a** in the presence of 0.016 M CF_3SO_3H. Because of the increase in K_b^H with decrease in the methanol content of the solvent the extent of protonation of **1a** to **1a**-H$^+$ at equilibrium does become greater than 20% for some of the runs in 1.0 M and 2.0 M MeOH, but even in these cases the extent of protonation is low enough that [1-H$^+$] remains effectively proportional to [H$^+$] throughout. Under none of the reaction conditions in Tables I and II is the extent of protonation at equilibrium high enough that [1-H$^+$] would be even close to independent of [H$^+$].

Discussion

A striking feature of the acid-catalyzed methanolysis of **1** (eq 2) is that, although strong acid is necessary in order for methanolysis of the sulfenamide to take place at a significant rate, the rate in MeCN–8.2 M MeOH is almost *independent* of strong acid concentration for strong acid concentrations of 0.005 M or higher. Measurement of the basicity of **1a** and **1b** (K_b^H) shows that although the sulfenamides are reasonably basic (pK_a of **1a**-H$^+$ = 0.56 in MeCN–8.2 M MeOH; pK_a of **1b**-H$^+$ = −0.48 in the same medium) they are *not* sufficiently basic to be completely protonated to $ArSNH^+R_2$ under the conditions used for study of the kinetics of the methanolysis.

In contrast to the methanolysis, the acid-catalyzed thiolysis of **1** (eq 3) shows a *linear, first-order* dependence of rate on strong acid concentration under the same conditions. This further confirms that the lack of dependence of the rate of eq 2 on [CF_3SO_3H] *cannot* be the result of the sulfenamides being completely protonated to $ArSN^+HR_2$ at quite low [CF_3SO_3H]. Were that the case the acid-catalyzed reaction of **1** with the thiol would also exhibit a rate with little or no dependence on strong acid concentration. Clearly it does not. Nor would this be expected given the values of K_b^H (eq 4-*h*) for **1a** and **1b** and the acid concentrations employed in the kinetic studies.

Mechanism of the Thiolysis Reaction. The kinetics of the reaction of **1** with the thiol are unexceptional and indicative of a mechanism of the general type

$$1 + \text{H}^+ \underset{}{\overset{K_b^H}{\rightleftharpoons}} 1\text{-H}^+ \qquad (4\text{-}h)$$

$$\text{R'SH} + 1\text{-H}^+ \xrightarrow[\text{rd}]{k_5} \text{ArSSR'} + \text{R}_2\text{NH} + \text{H}^+ \qquad (5)$$

where the rate-determining step is the reaction of the thiol with the 1-H$^+$ present in equilibrium with the sulfenamide. With this mechanism the experimental first-order rate constants for thiolysis, k_1 (Table II), will be given by:

$$k_1 = \frac{k_5 K_b^H [\text{H}^+][\text{R'SH}]}{K_b^H[\text{H}^+] + 1}$$

Rate constant $k_{RSH} = k_5 K_b^H$ (also shown in Table II) is thus equal to

$$k_{RSH} = k_5 K_b^H = \frac{k_1(K_b^H[\text{H}^+] + 1)}{[\text{R'SH}][\text{H}^+]}$$

Except for the runs in MeCN–1.0 M MeOH, particularly those with **1a**, $K_b^H[\text{H}^+]$ is small enough compared to 1.0 that $k_{RSH} = k_1/[\text{R'SH}][\text{H}^+]$.

Since we know how K_b^H varies with solvent (Table III), we can evaluate how k_5 ($k_5 = k_{RSH}/K_b^H$) changes with changing methanol content of the solvent. The results for **1b** are as follows [[MeOH], k_5 (M^{-1} s^{-1})]: 8.2 M, 1.2×10^3; 5.0 M, 1.3×10^3; 3.0 M, 1.7×10^3; 1.0 M, 0.54×10^3. For [MeOH] = 3.0–8.2 M k_5 is effectively independent of solvent composition. Below [MeOH] = 3.0 M it decreases somewhat with decreasing [MeOH], being about three times slower in MeCN–1.0 M MeOH than in MeCN–3.0 M MeOH. The results for **1a** are as follows: [[MeOH], k_5 (M^{-1}s^{-1})]: 8.2 M, 0.6×10^3; 5.0 M, 0.5×10^3; 3.0 M, 0.4×10^3; 1.0 M, 0.07×10^3. These data show that the factor responsible for the marked decrease in k_{RSH} with increasing [MeOH] is the decrease in K_b^H that occurs as the methanol content of the solvent is increased. We conclude that the rate of reaction of the thiol (a neutral molecule) with 1-H$^+$ (eq 5) is not subject to any large solvent effect in MeCN–MeOH, and what effect there is actually results in k_5 increasing somewhat with increasing methanol concentration (at least over the range 1.0–3.0 M MeOH).

Table II shows that k_{RSH} for sulfenamide **1a** is 5.2 times larger than that for **1b** in MeCN–8.2 M MeOH. Remembering that $k_{RSH} = k_5 K_b^H$ and that K_b^H for **1a** is 11 times larger than K_b^H for **1b**, this indicates that k_5 is 2.1 times faster for **1b**-H$^+$ than for **1a**-H$^+$. A similar calculation for the data for MeCN–1.0 M MeOH indicates that in that solvent k_5 for **1b**-H$^+$ is about eight times faster than k_5 for **1a**-H$^+$.

We will return to consideration of the detailed sequence of events in eq 5 after considering the implications of the kinetic data in Table I for the mechanism of the methanolysis (eq 2).

Mechanism of the Methanolysis Reaction. How can the acid-catalyzed methanolysis reaction (eq 2) exhibit kinetics where the rate is only slightly dependent on [H$^+$] under conditions where the rate of the acid-catalyzed thiolysis (eq 3) shows a first-order dependence on [H$^+$] and where we know (Table III) that the extent of protonation of **1** to 1-H$^+$ at equilibrium is always modest? One way that it can do so is if the protonation of **1** to give 1-H$^+$ (eq

(5) Paul, M. A.; Long, F. A. *Chem. Rev.* **1957**, *57*, 1.

4) is followed by a sequence of steps (such as eq 6 and 7) where acid can exert a *retarding* effect on the overall rate for the reaction ($k_{-6}[H^+] > k_7$).[6]

$$\mathrm{MeOH} + \mathbf{1}\text{-}H^+ \underset{k_{-6}}{\overset{k_6}{\rightleftharpoons}} \mathbf{4} + H^+ \qquad (6)$$

$$\mathbf{4} \xrightarrow{k_7} \mathrm{ArSOMe} + R_2NH \qquad (7)$$

The mechanism of eqs 4, 6, and 7 gives the following expression for k_1, the experimental first-order rate constant for eq 2 (Table I)

$$k_1 = \frac{k_7 k_6' K_b[H^+]}{(k_{-6}[H^+] + k_7)(K_b[H^+] + 1)} \qquad (8a)$$

(where k_6' includes any dependence of the rate of step k_6 on methanol concentration). Equation 8a can be rearranged to

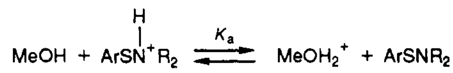
$$\left(\frac{1}{k_1}\right)\left(\frac{K_b[H^+]}{K_b[H^+] + 1}\right) = \left(\frac{k_{-6}}{k_7 k_6'}\right)[H^+] + \left(\frac{1}{k_6'}\right) \qquad (8b)$$

Plots of $(1/k_1)(K_b[H^+]/K_b[H^+] + 1)$ vs $[H^+]$ for the data for the methanolysis of **1a** and **1b** in MeCN–8.2 M MeOH are predicted, and found, to be linear, providing support for the suggested mechanistic interpretation of the methanolysis kinetics. From the slopes and intercepts of these plots k_6' for **1b** in MeCN–8.2 M MeOH is calculated to be 0.26 s^{-1} and (k_{-6}/k_7) to be $3.3 \times 10^2\ M^{-1}$, while for **1a** k_6' is 0.035 s^{-1} and (k_{-6}/k_7) is $3.1 \times 10^2\ M^{-1}$.

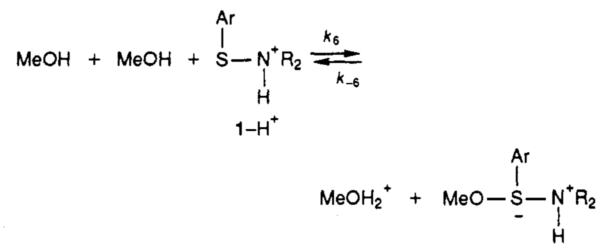
A second notable feature of the kinetics of the methanolysis reaction in MeCN–MeOH is that the rate shows a marked dependence on the methanol content of the solvent, being *seven times faster* for **1a** in MeCN–8.2 M MeOH than in MeCN–3.0 M MeOH, despite the fact that K_b for the sulfenamide is 6.5 times *smaller* in MeCN–8.2 M MeOH than it is in MeCN–3.0 M MeOH (Table III).

Comparison of the slope and intercepts of plots of the data for MeCN–3.0 M MeOH with those for MeCN–8.2 M MeOH allows us to determine to what extent the large increase in $k_7k_6'/(k_{-6}[H^+] + k_7)$ with increase in [MeOH] is due to an increase in k_6', a change in $k_7/k_{-6}[H^+]$, or a combination of both. Plots of the data according to eq 8b for the runs in MeCN–3.0 M MeOH give $k_6' = 0.0048\ s^{-1}$ and $(k_{-6}/k_7) = 56\ M^{-1}$ for **1b**, while for **1a** $k_6' = 0.000\,43\ s^{-1}$ and $(k_{-6}/k_7) = 59\ M^{-1}$. The increase in $k_7k_6'/(k_{-6}[H^+] + k_7)$ therefore results from the fact that k_6' for the sulfenamides is from 54 (**1b**) to 80 (**1a**) times larger in MeCN–8.2 M MeOH than in MeCN–3.0 M MeOH. The fact that (k_{-6}/k_7) for both sulfenamides increases by a factor of about 5.5 with the same change in solvent causes the partitioning of **4** between products (step k_7) and reversion to reactants (step k_{-6}) to be *less* favorable to step k_7 in the medium with the higher methanol content and is responsible for the lesser dependence of the rate on $[H^+]$ in MeCN–8.2 M MeOH.

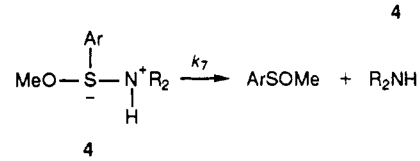
Is there a reasonable explanation for the very large effect of solvent on k_6', particularly since reaction of a thiol with **1**-H^+ is not subject to a significant solvent effect under the same conditions? We believe there is. We know (see Table III) that K_a for **1**-H^+ ($K_a = 1/K_b$), which represents the equilibrium constant for



varies markedly with solvent composition, being from 4.0 (**1b**-H^+) to 6.3 (**1a**-H^+) times *larger* in MeCN–8.2 M MeOH than in MeCN–3.0 M MeOH. This is because the proton *activity* of a solution containing a given $[CF_3SO_3H]$ decreases with increasing percentage of methanol in the MeCN–MeOH solvent. If we write the equilibrium in eq 6 in the fashion shown in eq 9a, then (except for the extra



(9a)



(9b)

molecule of methanol on the left) it also represents an equilibrium much like that involving **1**-H^+ and **1**, and $K_6 = k_6/k_{-6}$ would be expected to vary with solvent composition in the same manner as K_a. If we assume that the increase in k_{-6}/k_7 with increasing methanol content of the solvent is due entirely to an increase in k_{-6} then we might expect that $k_6' = k_6[\mathrm{MeOH}]$ for **1a**-H^+ could be larger in MeCN–8.2 M MeOH than in MeCN–3.0 M MeOH by a factor of

$$\frac{k_6'(8.2\ \mathrm{M})}{k_6'(3.0\ \mathrm{M})} = \frac{(K_6k_{-6})_{8.2\mathrm{M}}}{(K_6k_{-6})_{3.0\mathrm{M}}}\frac{8.2}{3.0} = (6.3 \times 5.3)\frac{8.2}{3.0} = 91$$

A similar calculation for **1b**-H^+ using $K_6(8.2\ \mathrm{M})/K_6(3.0\ \mathrm{M}) = 4.0$ and $k_{-6}(8.2\ \mathrm{M})/k_{-6}(3.0\ \mathrm{M}) = 5.8$ (from the variation in k_{-6}/k_7 with solvent for **1b**-H^+) predicts $k_6'(8.2\ \mathrm{M})/k_6'(3.0\ \mathrm{M})$ for that sulfenamide as 63.

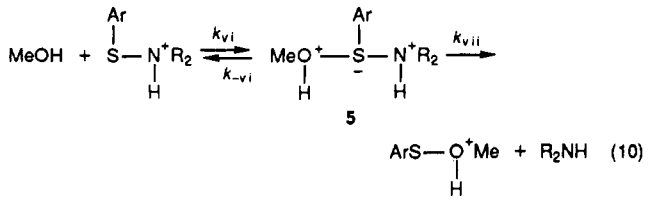
These values are sufficiently close to the observed increases for k_6' (80 for **1a**-H^+, 54 for **1b**-H^+) that we believe the explanation just presented can account very satisfactorily for the large increase in k_6' with increasing methanol content of the solvent.

In step k_6 a molecule of methanol acts as a base to assist the attack of another methanol molecule on **1**-H^+. We believe the reason a proton has to be removed from the attacking methanol molecule in step k_6 is the following. Simple attack of a methanol molecule on **1**-H^+ would lead to **5** (eq 10), rather than **4**. In **5** MeO^+H should be a much better leaving group than R_2N^+H. As a consequence k_{-vi} could easily be so many orders of magnitude larger than k_{vii} that the pathway to products through **5** would not be kinetically viable. On the other hand, if a proton is removed from the attacking methanol molecule synchronous

(6) We also considered, and *were able to rule out* (by the experiments outlined below), that the peculiar dependence of the methanolysis rate on $[CF_3SO_3H]$ might represent a combination of an acid-catalyzed and an uncatalyzed (spontaneous) pathway, *i.e.*, $k_1 = k_{cat}[H^+] + k_o$, with k_o being *much* more dependent on the methanol content of the solvent than k_{cat}. First, we showed that methanolysis of **1** does not occur, even very slowly, in the absence of CF_3SO_3H. This rules out the possibility of any simple uncomplicated spontaneous methanolysis of **1**. It does not, however, rule out a spontaneous methanolysis (eq i) where acid is necessary to convert R_2NH to $R_2NH_2^+$ as soon as it is formed by step k_o because the equilibrium constant ($K_o = k_o/k_{-o}$) for the spontaneous methanolysis is unfavorable.

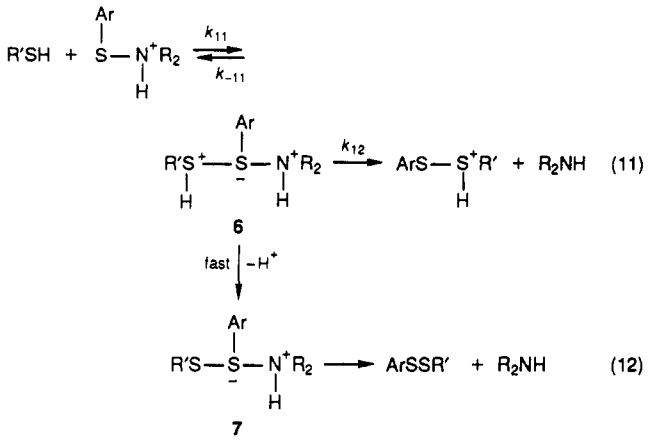
$$\mathrm{MeOH} + \underset{\mathbf{1}}{\mathrm{ArSNR_2}} \underset{k_{-o}}{\overset{k_o}{\rightleftharpoons}} \underset{\mathbf{2}}{\mathrm{ArSOMe}} + R_2NH \xrightarrow[\text{fast}]{H^+} R_2NH_2^+ \qquad (i)$$

However, if this were true, then R_2NH should react *readily* with **2** in neutral solution (step k_{-o}) in MeCN–MeOH to afford **1**. We were able to show by observing such a solution at room temperature over several days that such a reaction does *not* take place at a rapid enough rate.



with its attack on 1-H$^+$, the intermediate formed is **4**, where the MeO$^+$H group of **5** has been replaced by MeO. Loss of the methoxy group is now enough slower that k_7, although smaller under some circumstances than k_{-6}[MeOH$_2^+$], is adequately competitive in rate that the pathway to products via **4** is feasible.[7]

Given this explanation for the observed behavior of the methanolysis reaction, why do we see different behavior for the reaction of the thiol with **1**? Several explanations are possible. One is that the high thiophilicity of R'SH relative to MeOH,[8] and the general stability of sulfur–sulfur bonds, result in k_{-11} for the intermediate **6** formed by attack of the thiol on 1-H$^+$ (eq 11) being much smaller



than k_{-vi} for **5**. As a result k_{12} is either greater than k_{-11} or adequately competitive with it in rate, and the reaction can proceed satisfactorily via the mechanism shown in eq 11. A second possibility is that **6**, as a result of the low proton basicity of a thiol compared to an alcohol,[9] loses a proton to the solvent rapidly, and effectively irreversibly, to give **7**, and **7**, once formed, collapses only to products.

The unusual kinetic behavior of the methanolysis reaction is, as we have seen, satisfactorily explained by the mechanism in eqs 4 and 9 where **4** (a hypervalent sulfur species) is an actual intermediate on the reaction coordinate. On the other hand, it is extremely difficult, if not impossible, to account for the results by any mechanism where formation of the S–O bond and cleavage of the S–N bond are synchronous and there is no actual intermediate on the reaction coordinate. A recent kinetic study of the acid-catalyzed hydrolysis of ethyl benzenesulfenate[10] has also given results that are easily accommodated only by an addition–elimination mechanism where a hypervalent sulfur species is an intermediate on the reaction coordinate. The present study and the work of Okuyama et al.[10] provide important support for the view that addition–elimination mechanisms with sulfuranide intermediates, rather than one-step synchronous bond formation and cleavage, are operative for a number of nucleophilic substitutions at dicoordinate sulfur. Study of the mechanism of the reaction of methoxide ion with a cyclic thiosulfonate ester[11] also led to this conclusion.

Experimental Section

Preparation of Sulfenamides. *o*-Nitrobenzenesulfenyl chloride was prepared[12] from di-*o*-nitrophenyl disulfide. The sulfenamides were synthesized[13] by the addition, at 0 °C, of an ether solution of the sulfenyl chloride to an ether solution containing 2 molar equiv of the amine. The amine hydrochloride that separated was filtered off and washed with a small amount of ether. The filtrate and the washings were evaporated under reduced pressure, and the crude sulfenamide (>90% yield) was recrystallized from 95% ethanol:

***N,N*-Dimethyl-*o*-nitrobenzenesulfenamide (1a):** mp 58–60 °C (lit.[13] mp 62.5–63 °C); ^{1}H NMR ($CDCl_3$) δ 2.92 (s, 6 H), 7.25 (t, 1 H), 7.65 (t, 1 H), 7.92 (d, 1 H), 8.29 (d, 1 H); UV (MeCN) λ_{max} 382 nm (ϵ 4800).

***N*-Morpholino-*o*-nitrobenzenesulfenamide (1b):** mp 90–92 °C (lit.[14] mp 89.5–90.0 °C); ^{1}H NMR ($CDCl_3$) δ 3.1 (m, 4 H), 3.83 (m, 4 H), 7.27 (t, 1 H), 7.66 (t, 1 H), 8.13 (d, 1 H), 8.3 (d, 1 H); UV (MeCN) λ_{max} 382 nm (ϵ 3860).

Methanolysis of Sulfenamides. Products. The *o*-nitrobenzenesulfenamide (**1a** and **1b**), 0.1 mmol, was dissolved in 2 mL of acetonitrile–methanol (2:1), and 0.2 mmol of trifluoromethanesulfonic acid (Aldrich) was added. After 6 h the solvent was evaporated under reduced pressure, the residue was dissolved in methylene chloride and washed with water, and the methylene chloride was dried (Na_2SO_4). Removal of the methylene chloride gave 0.011 g (60%) of methyl *o*-nitrobenzenesulfenate (**2**), mp 48–50 °C (lit.[15] 49–50 °C), with spectral properties identical with a known sample.[15]

Kinetics. An acetonitrile–methanol solution (3 mL) containing the desired concentrations of trifluoromethanesulfonic acid and lithium trifluoromethanesulfonate (Aldrich) was placed in a 1-cm spectrophotometer cell in the thermostated cell compartment of a Beckman DU-50 spectrophotometer, and 10–20 μL of a stock solution of **1a** or **1b** (0.013–0.04 M) in acetonitrile was injected into the cell to initiate the reaction. The progress of the methanolysis was followed by observing the increase in the absorbance (A) of the solution at 415 nm. Plots of log ($A_\infty - A$) vs time were linear in every case. The experimental first-order rate constant for each run was evaluated from the slope of the plot.

Reaction of 1-Butanethiol with Sulfenamides. Products. The 1-butanethiol (Aldrich) employed in this and the kinetic studies was redistilled twice before use. To a stirred solution of *o*-nitrobenzenesulfenamide **1b** (4.0 mmol) in 30 mL of CH_2Cl_2 was added 0.5 mL (5.7 mmol) of CF_3SO_3H and 0.43 mL (4 mmol) of 1-butanethiol. After 15 min the solution was washed with water (30 mL), saturated sodium bicarbonate (30 mL), and water again (30 mL) and then dried over $MgSO_4$. The methylene chloride was removed under reduced pressure to give 0.9 g of a brown, oily liquid which was chromatographed on silica gel (CH_2Cl_2:hexane = 1:4) giving 0.85 g (87%) of *n*-butyl *o*-nitrophenyl disulfide[16] [**3**, R' = *n*-Bu]: ^{1}H NMR ($CDCl_3$) δ 0.90 (t, 3 H), 1.42 (sextet, 2 H), 1.66 (quintet, 2 H), 2.74 (t, 3 H), 7.35 (t, 1 H), 7.69 (t, 1 H), 8.26 (d, 1 H), 8.3 (d, 1 H); ^{13}C NMR ($CDCl_3$) δ 13.51, 21.63, 31.04, 38.17, 125.89, 126.00, 127.28, 133.80, 137.97, 145.73. Anal. Calcd for $C_{10}H_{13}NO_2S_2$: C, 49.37; H, 5.38. Found: C, 49.28; H, 5.50.

(7) It is also possible that formation of **4** in eq 9a actually involves two steps, the first being addition of methanol to 1-H$^+$ to give **5** (as in eq 10) and the second transfer of a proton from **5** to the solvent:

$$MeOH + 5 \rightleftharpoons MeOH_2^+ + 4$$

Regardless of whether **4** is produced in this way, or in a single step (as in eq 9a), the key point is that **4** is an intermediate that partitions to products with reasonable efficiency, whereas **5**, because of MeOH$^+$ being a much better leaving group than R_2NH^+, does not.

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Kinetics. The procedure was generally the same as for the kinetic studies of the methanolysis. The solution initially placed in the spectrophotometer cell contained the thiol, as well as CF_3SO_3H and CF_3SO_3Li. The progress of the reaction was followed by monitoring the decrease in the absorbance of the solution at 405 nm. Plots of log $(A - A_\infty)$ vs time were linear, and the experimental first-order rate constant for a run was obtained from the slope of such a plot.

Measurement of the Basicity of Sulfenamides. ^{1}H NMR Method. For **1a** in CD_3CN–CD_3OD (5.0 or 8.2 M) the chemical shift for the $(CH_3)_2N$ protons occurs at δ 2.91. Addition of trifluoromethanesulfonic acid to such a solution initially causes this signal to move to higher δ, but after enough acid has been added, further addition of acid has no effect on δ, the reason being that once sufficient excess acid has been added **1a** is present entirely as **1a**-D^+, $ArSND^+(CH_3)_2$, and the chemical shift for the $(CH_3)_2N^+$ protons in **1a**-D^+ is not affected by addition of further acid to the medium.

Other studies, where a small amount of $(CH_3)_2ND_2^+$ $CF_3SO_3^-$ was also present in the solution, showed that the chemical shift (δ 2.63) for the $(CH_3)_2ND_2^+$ protons in this solvent was invariant with trifluoromethanesulfonic acid concentration and could therefore serve as a suitable internal standard.

Sulfenamide **1a** (5 mg) was dissolved in 5 mL of CD_3CN–5.0 M (or 8.2 M) CD_3OD. To 0.8 mL of this solution was added a measured amount (10–100 μL) of a 0.7 M stock solution of trifluoromethanesulfonic acid in CD_3CN. Also present in the solution was a sufficient amount of $(CH_3)_2ND_2^+$ $CF_3SO_3^-$ (δ 2.63) to serve as an internal standard. Immediately after injection of the acid the ^{1}H NMR spectrum of the solution was recorded.

The fraction (α) of the sulfenamide present as **1a**-D^+ is related to K_b^D (eq 4-*d*) and $[D^+]$ as follows:

$$K_b^D = \alpha/(1 - \alpha)[D^+]$$

$$(1/\alpha) = 1 + (1/K_b^D[D^+]) \qquad (13a)$$

We define $\Delta\delta$ as the difference between the chemical shift for the Me_2N protons in a particular acid solution and the chemical shift for the same protons in unprotonated **1a** (δ 2.91). Studies[17] of other protonation equilibria of the general type of eq 4-*d* have shown that

$$\alpha \sim \Delta\delta$$

Introducing Y as the constant of proportionality between α and $\Delta\delta$ gives $\alpha = Y(\Delta\delta)$ and

$$1/(\Delta\delta) = (Y/K_b^D)(1/[D^+]) + Y \qquad (13b)$$

Regression analysis of plots of the data according to eq 13b gives Y as the intercept and (Y/K_b^D) as the slope. Such plots ($r \geq 0.996$) gave $Y = 2.1 \pm 0.05$ for the data for **1a** in MeCN containing 5.0 or 8.2 M CD_3OD and slopes leading to the K_b^D values for **1a** shown in Table III.

UV Method. Addition of sufficient CF_3SO_3H to solutions of either **1a** or **1b** in MeCN–MeOH leads to an *immediate* decrease in the absorbance of the solution at 382 nm (λ_{max} for **1**). This change has nothing to do with the absorbance change associated with the methanolysis of **1**; it is due to the fact that the UV absorption spectrum of **1**-H^+ is different from that of **1**. If A_o is the optical density for the particular solution with **1** unprotonated, A_{inf} is the optical density for the same solution with **1** completely protonated, and A is the measured optical density for a particular concentration of added CF_3SO_3H, then

$$\alpha = (A_o - A)/(A_o - A_{inf})$$

For each **1** and methanol concentration, measurements were made using solutions containing $1.0–2.0 \times 10^{-4}$ M **1a** or **1b** and three to five different concentrations of added CF_3SO_3H sufficient to give a conveniently measurable spread of α values. The data for each sulfenamide in a particular MeCN–MeOH solvent mixture were then plotted according to eq 13a. In each case such plots were linear ($r \geq 0.993$) and had an intercept on the $1/\alpha$ axis of 1.0 ± 0.1. Their slope is equal to $1/K_b^H$; those values are tabulated in Table III. The agreement for **1a** between K_b^D (measured by the ^{1}H NMR method) and K_b^H (measured by the UV method), when account is taken of the anticipated solvent isotope effect,[3] provides reassurance that the UV method measures K_b^H accurately.

The degree of protonation of two of the weak base amine indicators (*o*- and *p*-nitroaniline) used to establish the H_o acidity scale[5] in dilute aqueous sulfuric acid was also measured by an analogous UV procedure for MeCN–MeOH mixtures containing from 1.0 to 8.2 M MeOH. The change in K_b with changing [MeOH] in both cases paralleled very closely the changes found for **1a** and **1b** for the change in solvent. The measurements in MeCN–8.2 M MeOH gave the following pK_a's for the two amines: *p*-nitroaniline, +1.70; *o*-nitroaniline, +0.41. These are 0.70 pK units more positive than their pK_a's in dilute aqueous solution, indicating that $a_{H^+}f_B/f_{BH^+}$ is five times *larger* for a given $[H^+]$ in MeCN–8.2 M MeOH than in water.

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Pressure Effects on the Thermal Decomposition of Nitramines, Nitrosamines, and Nitrate Esters

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Solutions of nitramine and nitrate ester explosives and model compounds were thermolyzed at various hydrostatic pressures and their rates of decomposition were measured. The effects of pressure on their rates were used to infer the mechanism of their initial decomposition steps. Most nitramines, including the explosive octahydro-1,3,5,7-tetranitro-1,3,5,7-tetrazocine (HMX), appear to undergo homolysis of the N–NO_2 bond, because their reaction rates decrease with increasing pressure. Exceptions are hexahydro-1,3,5-trinitro-1,3,5-triazine (RDX) and the model compound, 6-nitro-1,2-dinitroso-1,2,3,4-tetrahydroquinoxaline, which react faster with increasing pressure. These two compounds can aromatize by elimination of HNO_2 and HNO, respectively. Secondary nitrate esters shift their major decomposition pathway from homolysis of the O–NO_2 bond to elimination of HNO_3 in the pressure range of 0.4 to 0.8 GPa. The elimination reaction resembles carboxylate ester pyrolysis with E1 character.

Introduction

There is an abundance of kinetic and mechanistic data pertaining to the decomposition of explosives with very little attention to the effect of pressure. The practical application of energetic materials usually involves exposure to intense shock waves or high dynamic pressures, and there is not much basis for theoretical prediction of their behavior under these conditions. Our aim is to increase

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