Denitrosation of Nitrosamines—a Quantitative Study. Reactions of *N*-Methyl-*N*-nitrosoaniline, *N*-Nitrosoproline, Dimethylnitrosamine and *N*-nitrososarcosine

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Rate measurements are reported for the denitrosation of *N*-methyl-*N*-nitrosoaniline (NMNA), *N*-nitrosoproline (NPr), dimethylnitrosamine (DMN), *N*-nitrososarcosine (NS) and *N*-nitrosopyrrolidine (NPy), in acid solution in the presence of nucleophilic catalysts in the following solvent systems: water, ethanol, various aqueous acetic acid solutions (up to 80% acetic acid) and acetonitrile. The reactivity sequence generally is found to be NMNA > NPr = *ca.* NS ≫ DMN = *ca.* NPy. The most reactive solvent system is 80% acetic acid–water containing bromide ion (or thiourea) as a nucleophilic catalyst. The results, together with some earlier work in water, are discussed (*a*) in terms of the changing relative reactivities of the nucleophiles as the polarity of the solvent is changed, and (*b*) in terms of the electron-withdrawing properties of any substituents which can effect both the protonation equilibria and also the rate constant for nucleophilic attack of the protonated form of the nitrosamine.

The carcinogenic properties of nitrosamines are now widely recognised. Even though nitrosamines do not occur naturally, they are readily formed from secondary amines and sources of nitrous acid, particularly nitrite and nitrate ions. Apart from possible *in vivo* formation *e.g.* in the acid environment of the stomach, there is no doubt that nitrosamines are also formed, often in very low concentrations, as by-products of other reactions. The analytical procedures for nitrosamine assay are now quite sophisticated and parts per billion can readily be detected. Clearly there is concern in many areas regarding the presence of even small quantities of nitrosamines in many foods and other products, and a need arises for efficient procedures for nitrosamine removal or destruction. They can be reduced to the corresponding hydrazines with zinc and acetic acid and other reducing agents,[1] and denitrosation to the secondary amine can be effected both thermally and photochemically in some cases.[2] A more useful procedure, which has been used on a laboratory and large scale involves the acid-catalysed denitrosation particularly in the presence of non-basic nucleophiles (Y^-). This is of course the reverse reaction of normal nitrosamine formation; generally the equilibrium favours nitrosamine formation, so for denitrosation to be effective steps must be taken to remove the nitrous acid (or YNO species) formed. Many reagents have been used in this way as 'nitrite traps' (X) including Cu^I and Fe^{II} salts,[3] urea,[4,5] azide ion,[5] sulphamic acid,[5] hydrazine,[5] ascorbic acid,[6] thiols,[7] benzenesulphinic acid[8] and iron pentacarbonyl.[9] The efficiency of a number of X species has been assessed quantitatively.[10]

Mechanistically the most widely studied substrates have been the *N*-alkyl-*N*-nitrosoanilines. In aqueous acid solution denitrosation is believed to occur by nucleophilic attack of Y^- at the nitroso nitrogen atom on the *N*-protonated nitrosamine, as outlined in Scheme 1. So long as the inequality $k_2[X] \gg k_{-1}$ applies then denitrosation is quantitative and effectively irreversible. When the solvent is water (and except at very high concentrations of some Y^- species) attack by Y^- is rate-limiting and the relative reactivity of various Y^- species [(Cl^-, Br^-, SCN^-, N_3^-, I^- and $SC(NH_2)_2$] is similar to that observed for S_N2 reactions at saturated carbon.[11,12] However in ethanol the reaction rate is independent of the nature and concentration of Y^-, and the protonation of the nitrosamine is rate-limiting.[13] A similar situation obtains for reactions in water when the nitrosamine contains powerfully electron-withdrawing substituents such as –CO–[14,15] and $-SO_2-$,[16] and also for

$$PhN(NO)R + H^+ \underset{}{\overset{K}{\rightleftharpoons}} Ph\overset{+}{N}H(NO)R$$

$$Ph\overset{+}{N}H(NO)R + Y^- \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} PhNHR + YNO$$

$$YNO + X \xrightarrow{k_2} \text{Various decomposition products}$$

Scheme 1

PhN(NO)Me when the more powerful nucleophiles SCN^- and $SC(NH_2)_2$ are present at high concentration.[17] In all cases the results have been rationalised in terms of the attack of Y^- competing effectively with deprotonation of the protonated nitrosamine.

Much less quantitative information is available for other nitrosamine types *e.g.* aliphatic dialkyl and heterocyclic nitrosamines, although the relative reactivities of some heterocyclic species (in 0.5 mol dm^{-3} SCN^- at 50 °C) have been established[18] as part of a study relating to the ability of these nitrosamines to transfer the NO group to other amines *via*, in this case, the intermediacy of nitrosyl thiocyanate, ONSCN. The nitrosamine prepared from *N*-acetyltryptophan behaves at high acid concentrations as do nitrosamides, in that protonation is rate-limiting, but at low acidity a different mechanism operates involving a nitroso group rearrangement from N to C-3.[19]

One analytical application of the denitrosation reaction arises in one of the procedures for nitrosoamine determination by the Thermal Energy Analyser technique,[20] in which nitrosamines are treated with 10% HBr solution in acetic acid at a temperature where the nitrosyl bromide formed breaks down to nitric oxide.

The purpose of the present work is to obtain quantitative results for denitrosation for a range of typical nitrosamines and so to establish the most favourable experimental conditions for their rapid and irreversible decomposition. We have chosen to work with dimethylnitrosamine (DMN), nitrososarcosine (NS), nitrosoproline (NPr) and nitrosopyrrolidine (NPy), and have also extended our earlier work with *N*-methyl-*N*-nitrosoaniline (NMNA) to include solvents other than water and ethanol.

Results and Discussion

(a) *Reactions of NMNA.*—Much is already known about the

Table 1 Nucleophile reactivities in the denitrosation of NMNA in acetic acid–water solvents containing sulphuric acid (1.1 mol dm^{-3}) and sodium azide (5×10^{-3} mol dm^{-3})

Solvent	Nucleophile	k_2/dm^3 mol^{-1} s^{-1}	Relative reactivity[b]
10% Acetic acid	Cl^-	Too slow to measure	
	Br^-	6.0×10^{-3}	1.0 (1.0)[c]
	SCN^-	0.56	93 (100)[c]
	$SC(NH_2)_2$	1.48	250 (240)[c]
33% Acetic acid	Cl^-	3.5×10^{-4}	1.0
	Br^-	2.1×10^{-2}	60
	SCN^-	0.68	1940
	$SC(NH_2)_2$	0.98	2800
67% Acetic acid	Cl^-	1.27×10^{-3}	1.0
	Br^-	0.32	250
	SCN^-	0.68	540
	$SC(NH_2)_2$	0.77	610
80% Acetic acid[a]	Cl^-	2.2×10^{-2}	1.0
	Br^-	1.33	60
	SCN^-	0.32	15
	$SC(NH_2)_2$	0.68	31

[a] Sulphuric acid concentration 0.6 mol dm^{-3}. [b] Relative to the least reactive in each group. [c] The figures in parentheses refer to the relative reactivities in water taken from ref. 12.

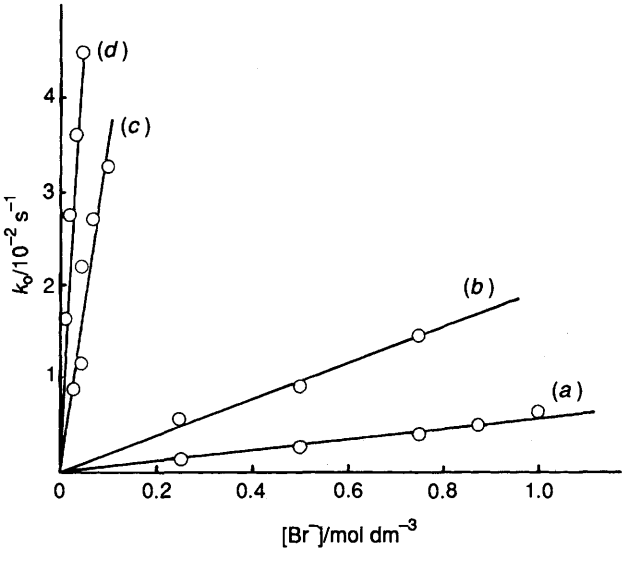


Fig. 1 Effect of [Br^-] on the rate constant for the denitrosation of NMNA in various acetic acid–water solvent mixtures: (*a*) 10%; (*b*) 33%; (*c*) 67%; (*d*) 80% acetic acid

denitrosation reactions of *N*-alkyl (and aryl)-*N*-nitrosanilines in water and ethanol solvents in the presence of acid catalysts. Here we extend our studies to include aqueous acetic acid solvents. Reactions were carried out in 80%, 67%, 33% and 10% acetic acid–water mixtures (v/v) in the presence of sodium azide (5×10^{-3} mol dm^{-3}). Under these conditions denitrosation occurred and the first-order rate constant was unchanged on increasing [NaN_3] further. The reaction mixture contained sulphuric acid (0.6–1.1 mol dm^{-3}) and the kinetic effect of added thiourea and chloride, bromide and thiocyanate ions was investigated. In all cases nucleophilic catalysis was observed and the value of k_2 (defined by, rate = k_2[NMNA][Y^-]) obtained from the slope of the plot of first-order rate constant against [Y^-]. The results for the bromide-ion-catalysed reactions are displayed in Fig. 1. The full results are shown in Table 1, where the final column displays the relative reactivities (relative to the least reactive) of the nucleophiles within each solvent group. Some trends are apparent. The reactivites of both thiourea and thiocyanate ion change very little with solvent composition. This is not surprising in the case of thiourea since solvation is not an important factor. However, at first sight one would expect the reactivity of SCN^- to increase as the solvation ability of the solvent decreases; this is not the case, possibly because the thiocyanate anion becomes progressively more protonated as the solvent moves towards pure acetic acid. The reactivity of both Br^- and Cl^- increases substantially on changing from water to 80% acetic acid as expected given the expected reduction in solvation. The limiting situation, whereby the reactivity of the halide ions is reversed, is however not achieved here, presumably because the solvation effects, although reduced on going to 80% acetic acid, are still a dominant influence. Very recently[21] the reversal of the nucleophilicity trend (compared with water) has been observed for dioxane–water solvent mixtures in the reverse reaction *i.e.* *N*-nitrosation.

The conclusion, from a practical viewpoint for effecting denitrosation of nitrosamines is that reactions become much faster as one changes the solvent from water to 80% acetic acid–water, and the increased reactivity of the bromide ion makes it the 'best' nucleophile of those studied in 80% acetic acid–water solvent, taking over from thiourea by a factor of *ca.* two.

The use of lithium bromide and sulphuric acid in acetic acid–water mixtures should be equivalent to the use of hydrogen bromide in the same solvents. We found that denitrosation of NMNA is rapid in both 60% and 80% acetic acid–water mixtures containing HBr and a nitrite trap. With HBr (1.5 mol dm^{-3}) in 60% acetic acid–water the half-life of the reaction is approximately 7 s at 25 °C. If, as expected, both the acid and bromide ion play a role, then the observed rate constant should depend upon $[HBr]^2$ rather than [HBr]. More correctly the dependence should be upon the product of [Br^-] and some appropriate acidity function. We do not have such an acidity function, but find (see Fig. 2) a much better correlation between the first-order rate constant (k_0) and $[HBr]^2$ than with [HBr] for reactions in 60% acetic acid–water. A similar result has been noted earlier[9] for the reaction of NMNA in hydrochloric acid.

(b) *Reaction of NPy and NPr.*—(i) *In water.* No significant reaction of NPy occurred even in the presence of substantial quantities of the most powerful nucleophile (thiourea) at high acidities and with a nitrite trap present. For NPr however

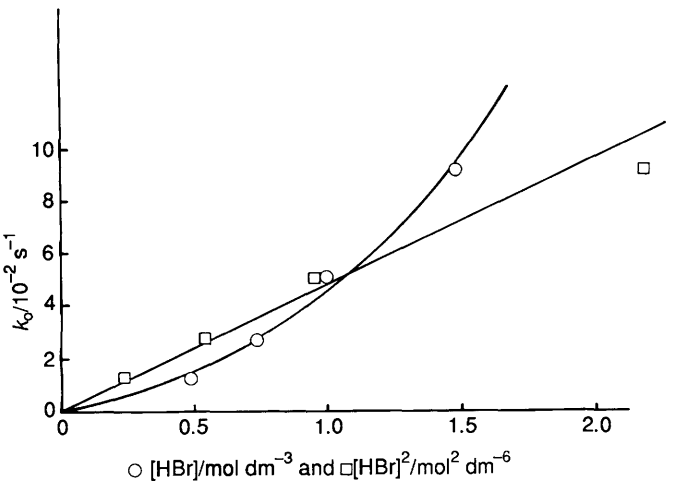


Fig. 2 Plot of k_0 *vs.* [HBr] and [HBr]^2 for the denitrosation of NMNA in 60% acetic acid–water containing HBr

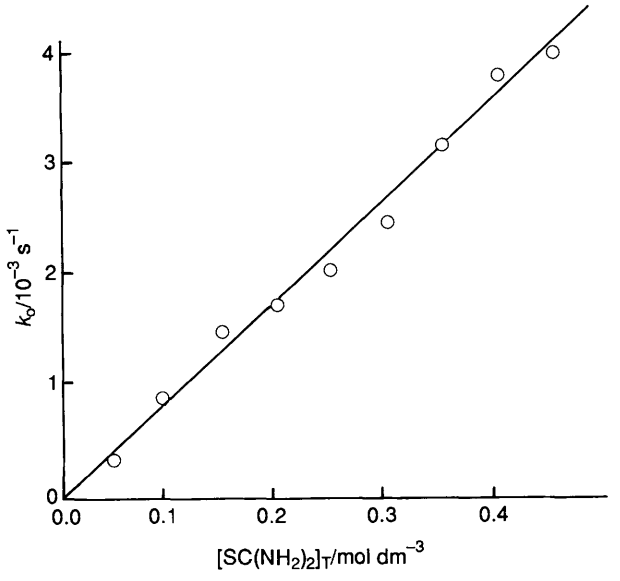


Fig. 3 Dependence of the rate constant for denitrosation of NPr in water upon [SC(NH$_2$)$_2$]$_T$

Table 2 Results for the denitrosation of NPr[a] in aqueous acid solution containing thiourea

[H_2SO_4]/mol dm^{-3}	[Thiourea]/mol dm^{-3}	$k_0/10^{-4}$ s^{-1}
0.80	0.10	0.93
1.60	0.10	2.79
2.40	0.10	6.21
3.20	0.10	9.22
4.00	0.10	11.9
4.80	0.10	13.8
3.20	0.05	3.42
3.20	0.10	8.69
3.20	0.15	11.4
3.20	0.20	16.7
3.20	0.25	20.1
3.20	0.30	24.3
3.20	0.35	31.3
3.20	0.40	38.0
3.20	0.45	40.7

[a] Throughout, [NPr] = 0.01 mol dm^{-3} and [sulphamic acid] = 0.2 mol dm^{-3}.

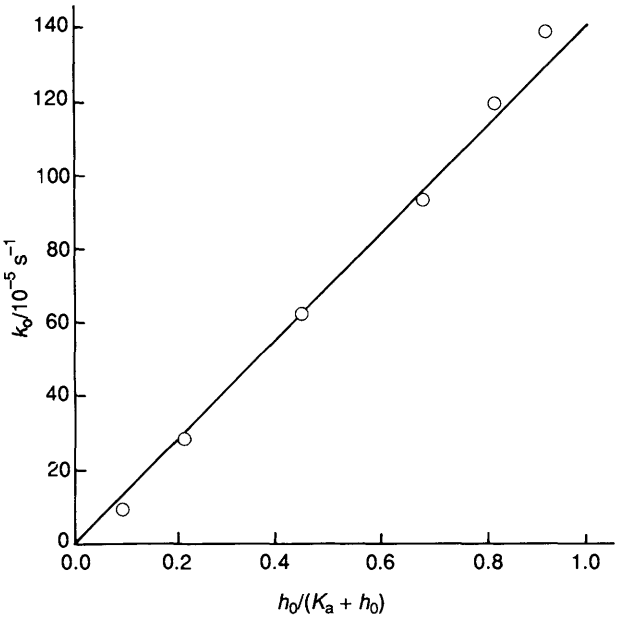


Fig. 4 Verification of eqn. (1) at constant [SC(NH$_2$)$_2$]$_T$

Table 3 Rate constants for the reaction of NPr (5×10^{-3} mol dm^{-3}) in ethanol containing H_2SO_4 (1.0 mol dm^{-3}) and ascorbic acid (0.02 mol dm^{-3})

[Thiourea]	$k_0/10^{-5}$ s^{-1}
0.01	7.90
0.025	8.85
0.04	9.08
0.05	8.26

denitrosation to give proline was achieved in aqueous acid solution in the presence of a nitrite trap, but only when thiourea was present. Reaction was detectable using bromide and thiocyanate at high acidity but was not conveniently measurable. Kinetic measurements were made noting the variation of the first-order rate constant k_0 (*a*) with [H_2SO_4] at constant [thiourea] and (*b*) with [thiourea] at constant [H_2SO_4]. The results are shown in Table 2. The reaction is clearly acid-catalysed and first-order in [thiourea]. Since thiourea is protonated (at the *S*-atom[22]) at these acidities, a correction must be made for the free [thiourea]. We have used as a first approximation the Hammett acidity function h_0 to describe protonation of both the nitrosamine and thiourea, for although other acidity functions may be more appropriate (particularly for thiourea protonation) the differences between these functions up to 4.8 mol dm^{-3} H_2SO_4 are quite small; the detailed form of the acidity dependence is not an issue in this work. We have used the Janssen value[23] (−1.19) for the pK_a for protonated thiourea.

The expected equation for k_0 assuming the sequence in Scheme 1, is given in eqn. (1), where [thiourea]$_T$ is the total

$$k_0 = k_1 K h_0 [\text{thiourea}]_T K_a/(K_a + h_0) \quad (1)$$

stoichiometric concentration of thiourea and K_a the acid dissociation constant for protonated thiourea. We find as expected (see Figs. 3 and 4), good linear relationships between (*a*) k_0 and [thiourea]$_T$ at constant acidity and (*b*) k_0 and $h_0/(K_a + h_0)$ at constant [thiourea]$_T$. The plots yield respectively values of k_1K of 1.0×10^{-3} and 0.9×10^{-3} dm^6 mol^{-2} s^{-1}; the agreement between the two sets of results is satisfactory. For comparison purposes we can note that the corresponding values measured earlier for NMNA[24] and diphenylnitrosamine[12] in aqueous acid solution are 0.52 and

Table 4 Kinetic results for the denitrosation of NPr (5×10^{-3} mol dm^{-3}) in 80% acetic acid–water containing added sodium azide (1×10^{-2} mol dm^{-3})

$[Br^-]$/mol dm^{-3}	[Thiourea]/mol dm^{-3}	$[H_2SO_4]$/mol dm^{-3}	$k_0/10^{-5}$ s^{-1}
0.05	—	1.0	4.5
0.075	—	1.0	6.5
0.10	—	1.0	9.6
0.125	—	1.0	10.8
0.15	—	1.0	11.9
—	0.05	1.0	37
—	0.075	1.0	45
—	0.10	1.0	55
—	0.125	1.0	76
—	0.15	1.0	84
0.05	—	0.5	3.6
0.05	—	1.0	7.5
0.05	—	1.5	11.0
0.05	—	2.0	15.3
—	0.05	0.5	20
—	0.05	1.0	37
—	0.05	1.5	60
—	0.05	2.0	79

Table 5 Kinetic results of the denitrosation of DMN (5×10^{-3} mol dm^{-3}) in 80% acetic acid–water containing sodium azide (1×10^{-2} mol dm^{-3}), thiourea and sulphuric acid

[Thiourea]/mol dm^{-3}	$[H_2SO_4]$/mol dm^{-3}	$k_0/10^{-5}$ s^{-1}
0.05	1.0	1.49
0.075	1.0	2.50
0.10	1.0	3.50
0.15	1.0	4.78
0.20	1.0	6.06
0.10	0.5	1.90
0.10	1.0	4.05
0.10	1.5	5.47
0.10	2.0	7.25

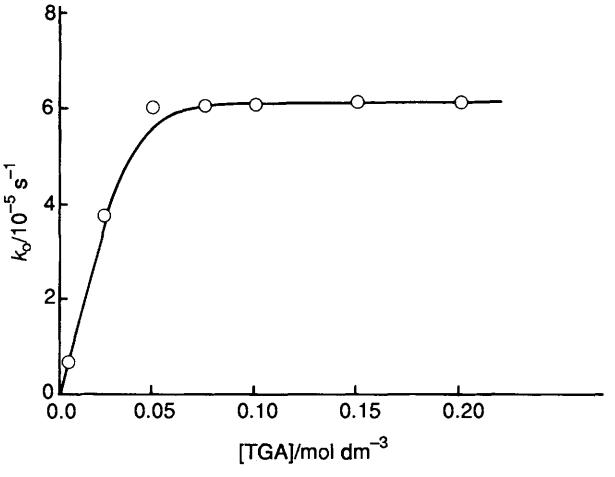


Fig. 5 Effect of [thioglycolic acid] on the rate constant for denitrosation of NPr in acetonitrile

3.0 dm^6 mol^{-2} s^{-1}. Thus the aromatic nitrosamines are many orders of magnitude more reactive to denitrosation than is NPr.

(ii) *In ethanol.* In the absence of an added 'nitrite trap' the denitrosation process was reversible. Ascorbic acid was used to restrict the reverse reaction and an essentially irreversible reaction was achieved when the ascorbic acid concentration was $\geqslant 0.01$ mol dm^{-3}. Good first-order behaviour was obtained, but the reactions were generally quite slow. Acid catalysis (using H_2SO_4) occurs but there is no nucleophilic catalysis (using thiourea). These characteristics suggest that, as for the reaction of NMNA in ethanol,[13] the proton transfer to the nitrosamine is rate limiting. This is supported by the observation of a kinetic isotope effect of k(EtOH)/k(EtOD) = 1.6–1.8 for three measurements made at $[H_2SO_4]$ 0.5, 1.0 and 1.5 mol dm^{-3}. The magnitude of this effect is comparable to earlier measurements on related systems[14,16] and contrasts markedly with the value of *ca.* 0.3 found when the nucleophile attack is rate limiting.[11] A typical set of results, showing the independence upon [thiourea] is shown in Table 3.

(iii) *In 80% acetic acid–water.* Denitrosation of NPr occurred in this solvent containing sulphuric acid and a 'nitrite trap'. Good first-order behaviour was found and both acid and nucleophile catalysis was evident. The dependence upon $[Br^-]$, [thiourea] and $[H_2SO_4]$ is shown in Table 4. In this case, contrasting somewhat with the reactions of NMNA, thiourea is a somewhat more effective catalyst than is bromide ion.

Denitrosation of NPr also occurred in acetonitrile containing thioglycolic acid (which may act as a nucleophile and a trap for any free nitrosating species) and sulphuric acid. Again acid catalysis and a first-order dependence upon [thioglycolic acid] was found at low [thioglycolic acid] indicating a rate-limiting attack by the sulphur nucleophile at the protonated nitrosamine. At high [thioglycolic acid] however, the reaction rate constant k_0 became independent of the nucleophile concentration as shown in Fig. 5. This indicates a change in rate-limiting step from attack of the nucleophile to protonation of the nitrosamine. This effect has previously been noted for some nitrosamines at high concentrations of thiourea and thiocyanate ion for denitrosation of NMMA in water.[17]

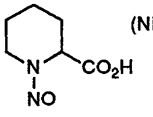
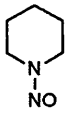
Reactions were not particularly rapid in acetonitrile, *e.g.* the half-life at 0.6 mol dm^{-3} H_2SO_4 using 0.1 mol dm^{-3} thioglycolic acid was *ca.* 1 h at 25 °C, so this solvent system has no particular advantage over those already studied and was not examined further.

(c) *Reactions of DMN and NS.*—These substrates were chosen as representative examples of aliphatic dialkyl nitrosamines. Denitrosation of DMN could not be achieved to any measurable extent in either water or ethanol solvent, even at high acidity and in the presence of quite high concentrations of thiourea or thiocyanate ion. However in 80% acetic acid–water solvent reaction did occur in the presence of thiourea and an acid catalyst, but was much slower than the corresponding reactions of both NMNA and NPr. Details, showing again acid catalysis and thiourea catalysis are in Table 5. *N*-Nitrososarcosine $[CH_3N(NO)CH_2CO_2H]$ is much more reactive than DMN, so much so that its denitrosation can be brought about in aqueous acid solution containing a nitrous acid trap and thiourea as the nucleophilic catalyst. Data for reactions in 0.4–

Table 6 Kinetic results for the denitrosation of NS (2×10^{-3} mol dm^{-3}) in water containing sulphamic acid (2×10^{-2} mol dm^{-3}), thiourea and sulphuric acid

[Thiourea]/mol dm^{-3}	$[H_2SO_4]$/mol dm^{-3}	$k_0/10^{-5}$ s^{-1}
0.20	0.40	13.1
0.20	0.80	20.7
0.20	1.00	22.4
0.20	1.20	29.0
0.20	1.60	43.5
0.20	2.00	65.3
0.01	1.20	1.08
0.02	1.20	2.88
0.03	1.20	3.96
0.04	1.20	4.48
0.05	1.20	6.68

Table 7 Quantitative estimates of denitrosation reactivity in water containing sulphuric acid, thiourea and sodium azide

Nitrosamine	k_1K/dm^6 mol^{-2} s^{-1} [a]
Ph_2NNO (ND)	3.0
PhN(Me)NO (NMNA)	0.52
 (NPr)	1.0×10^{-3}
$HO_2CCH_2N(Me)NO$ (NS)	6.4×10^{-4}
 (NPy)	Too slow to measure
Me_2NNO (DMN)	Too slow to measure

[a] From Scheme 1.

2.0 mol dm^{-3} H_2SO_4 and 0.01–0.05 mol dm^{-3} thiourea are shown in Table 6. If we assume, as we did earlier [section (*b*) (*i*)] for the reactions of NPr under similar conditions, that the Hammett acidity function adequately represents the protonation equilibria of both the nitrosamine and thiourea, then eqn. (1) should be valid for these results for NS. We find, as before, a good linear relationship between k_0 and $h_0/(K_a + h_0)$ from which a value of 7.1×10^{-4} dm^6 mol^{-2} s^{-1} is obtained for the product k_1K. A plot of k_0 *vs.* $[thiourea]_T$ yields another value of 5.7×10^{-4} dm^6 mol^{-2} s^{-1}, which represents a reasonable agreement. These results show that NS and NPr (where the corresponding values of k_1K were found to be 1.0×10^{-3} and 0.9×10^{-3} dm^6 mol^{-2} s^{-1}), undergo denitrosation at much the same rate under comparable conditions.

Conclusions

We find that of the solvent systems studied (water, ethanol, various acetic acid–water mixtures and acetonitrile) denitrosation occurs most readily in 80% acetic acid–water in the presence of either bromide ion or thiourea as nucleophilic catalysts, sulphuric acid and any nitrous acid trap such as added sodium azide. The order of effectiveness of the added nucleophiles is different from that in water probably because of the different solvation effects of the two systems. Reactions in ethanol (for NMNA and NPr) are somewhat different in that the rate-limiting step changes from the reaction of the protonated nitrosamine with the nucleophile, to the formation of the protonated nitrosamine, resulting in a zero-order rate dependence upon the nucleophile concentration.

Of the structural types studied, the aromatic nitrosamines are by far the most reactive. The collected data for the reactions in water containing thiourea are shown in Table 7, where the final column (k_1K) represents the overall reactivity of each substrate. We are unable to separate the effects of structure on (*a*) the protonation of the nitrosamine and (*b*) the reactivity of the protonated nitrosamine with a common nucleophile, since the pK_a values of the protonated nitrosamines have not been determined. It is likely however that the enhanced reactivity of the aromatic nitrosamines is due to the electron-attracting effect of the Ph group(s) which assists nucleophilic attack. In the absence of any such electron-attracting substituents, as in NPy and DMN, no reaction is detectable in the water solvent system even at high acidity. When a β-carboxy group is introduced (NPr and NS) reaction does occur, although not as rapidly as in the case of the aromatic systems. The reactivities of NPr and NS are comparable showing that the heterocyclic system does not play an important part in controlling reactivity.

From the practical point of view it is clear that denitrosation can be effected readily in a number of solvents if the substrates contain electron-withdrawing systems. If not, then reaction can only be accomplished in the most reactive solvent system (of those studied) *i.e.* 80% acetic acid–water containing bromide ion or thiourea.

Experimental

The nitrosamines NMNA, NPr, NPy and NS were prepared by the standard method of nitrosation of the corresponding secondary amines. DMN was a commercial sample. All solvents and reagents used were commercially available and were purified by standard methods. Kinetic measurements were carried out spectrophotometrically at 25 °C noting the disappearance of an absorbance due to the nitrosamine in a conventional recording spectrophotometer. Good first-order behaviour was found throughout and the first-order rate constant k_0 was determined from the integrated first-order rate equation. Values of k_0 were generally reproducible to within better than ±4%.

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References

1 P. A. S. Smith, *The Chemistry of Open-Chain Organic Nitrogen Compounds*, Benjamin, New York, 1966, vol. II, pp. 478–9.
2 Ref. 1 pp. 470–2.
3 E. C. S. Jones and J. Kenner, *J. Chem. Soc.*, 1932, 711.
4 W. G. Macmillan and T. H. Reade, *J. Chem. Soc.*, 1929, 585.
5 D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1975, 655.
6 S. S. Mirvish, L. Wallcave, M. Eagen and P. Shubik, *Science*, 1972, **177**, 65.
7 L. R. Dix and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1984, 109.
8 T. Bryant and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1985, 1083.
9 H. Alper and J. T. Edward, *Can. J. Chem.*, 1970, **48**, 1543.
10 J. Fitzpatrick, T. A. Meyer, M. E. O'Neill and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1984, 927.
11 I. D. Biggs and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1975, 107.
12 J. T. Thompson and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1977, 1932.

13 S. S. Johal, D. L. H. Williams and E. Buncel, *J. Chem. Soc., Perkin Trans. 2*, 1980, 165.
14 C. N. Berry and B. C. Challis, *J. Chem. Soc., Perkin Trans. 2*, 1974, 1638; B. C. Challis and S. P. Jones, *J. Chem. Soc., Perkin Trans. 2*, 1975, 153.
15 J. K. Snyder and L. M. Stock, *J. Org. Chem.*, 1980, **45**, 1990; G. Hallett and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1980, 1372; S. S. Singer and B. B. Cole, *J. Org. Chem.*, 1981, **46**, 3461.
16 D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1976, 1838.
17 S. S. Al-Kaabi, G. Hallett, T. A. Meyer and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1984, 1803.
18 S. S. Singer, G. M. Singer and B. B. Cole, *J. Org. Chem.*, 1980, **45**, 4931.
19 A. Castro, E. Iglesias, J. R. Leis, M. E. Peña, J. V. Tato and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1986, 1165.
20 C. L. Walters, R. J. Hart and S. Perse, *Z. Lebensm. Unters. Forsch.*, 1978, **167**, 315.
21 C. Bravo, P. Herves, E. Iglesias, J. R. Leis and M. E. Peña, *J. Chem. Soc., Perkin Trans. 2*, 1990, 1969.
22 M. J. Janssen, *Spectrochim. Acta*, 1961, **17**, 475; W. Kutzelnigg and R. Mecke, *Spectrochim. Acta*, 1961, **17**, 530; R. J. Gillespie and T. Birchall, *Can. J. Chem.*, 1963, **41**, 2643.
23 M. J. Janssen, *Recl. Trav. Chim. Pays-Bas*, 1962, **81**, 650.
24 G. Hallett and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1980, 624.

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