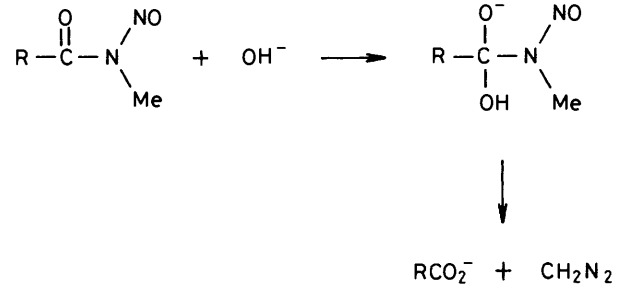
Decomposition of *N*-Methyl-*N*-nitrosotoluene-*p*-sulphonamide in Basic Media: Hydrolysis and Transnitrosation Reactions

Albino Castro, J. Ramón Leis,* and M. Elena Peña
Departamento de Química Física, Facultad de Química, Universidad de Santiago, Santiago de Compostela, Spain

The decomposition of *N*-methyl-*N*-nitrosotoluene-*p*-sulphonamide (MNTS) has been studied in basic and neutral water–alcohol mixtures. In alkaline media and when OH^- was the nucleophile, the known hydrolysis reaction in which OH^- attacks the SO_2 group was observed; this reaction was first order in both OH^- and MNTS. In the presence of ammonia, hydroxylamine, hydrazine, or primary, secondary or tertiary amines, a transnitrosation reaction took place in which the additional nucleophiles attacked the nitrogen atom of the MNTS N=O group; this reaction was first order in both MNTS and free amine. In particular, MNTS proved to be as efficient as some alkyl nitrites for the nitrosation of secondary amines in neutral or alkaline media, in which conventional nitrosating agents do not exist. Similar reaction rates were observed for the more basic tertiary amines (which gave NO_2^- among the final products). Primary amines underwent rather slower reactions, with the exception of hydroxylamine and hydrazine, the nucleophilic nature of which is increased by the α effect. We discuss the relative reactivities of the various amines in terms of their basicity and vertical ionization potentials, and we report the effect of the proportion of alcohol in the medium on the rates of both hydrolysis and transnitrosation reactions.

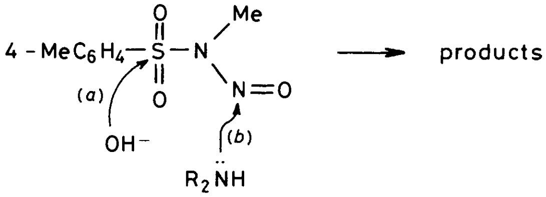
The decomposition of *N*-nitrosamides and related compounds in alkaline media is of interest from both a chemical and biomedical point of view. Traditionally, these reactions have been used to generate diazomethane (Scheme 1); biomedically, the carcinogenic nature of the substrates has been traced to the alkylating capacity of their products (in particular, to the methylation of DNA).



Scheme 1.

Kinetic studies of the above reaction with several nitrosamides and nitrosoureas have revealed significant differences among the substrates used, mainly as regards the kind of catalysis that takes place in the presence of bases, which in some cases act as nucleophiles on the CO group [1] and in others merely favour either the decomposition or the formation of the tetrahedral intermediate by general base catalysis.[2] The reason for these differences is not clear.

A particularly interesting substrate, for which no kinetic information in basic media has hitherto been published, is *N*-methyl-*N*-nitrosotoluene-*p*-sulphonamide (MNTS). The decomposition of MNTS in acid solutions has been studied [3] and its behaviour parallels that of other nitrosamides in that protonation of the substrate is the rate-limiting step. Besides that reaction, MNTS reacts as usual with nucleophiles such as OH^- or EtO^-, which attack the SO_2 group [Scheme 2(*a*)] to afford diazomethane,[4] but unlike nitrosamides and nitrosoureas it also undergoes nucleophilic attack by amines at its N=O group [Scheme 2(*b*)] [5]



Scheme 2.

This latter reaction, which is only mentioned in passing in the literature, implies that MNTS can act as a transnitrosating agent transferring the N=O group to secondary amines to produce *N*-nitrosamines under alkaline conditions in which the usual nitrosating agents, such as NO^+, N_2O_3, and nitrosyl halides, do not exist. No kinetic data have hitherto been published that allow the efficiency and competitivity of this reaction to be assessed and compared with those of agents such as alkyl nitrites [6] or nitroprusside,[7] which are also active in alkaline media. It should be remembered that the ability of a nitroso compound to nitrosate amines under diverse conditions is of importance for estimating the extent to which it may constitute an environmental hazard due to the possibility of its leading to the synthesis of carcinogenic *N*-nitrosamines.

In the work described here we investigate the reactivity of MNTS with the nucleophiles hydroxide, ammonia, hydroxylamine, hydrazine, and primary, secondary, and tertiary amines, and amino acids.

Experimental

MNTS of the greatest commercially available purity was supplied by Merck. Working solutions were prepared directly by dissolving known weights of MNTS in alcohol–water mixtures

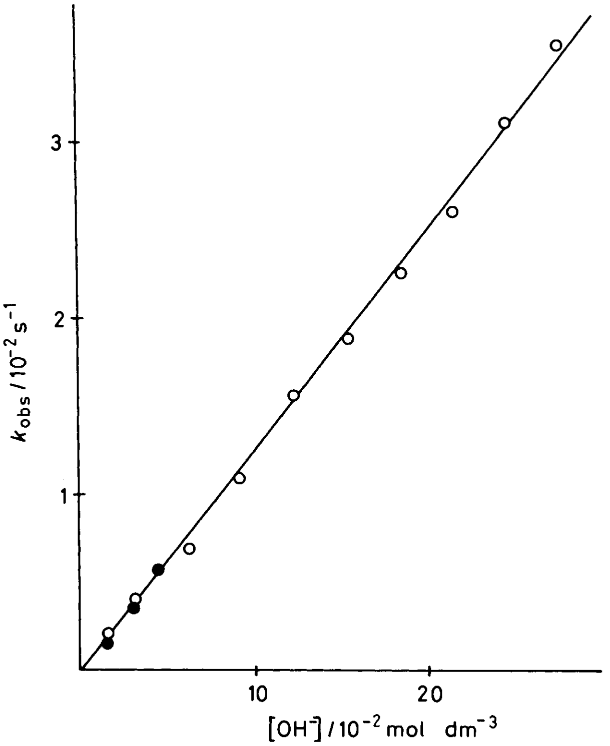


Figure 1. Influence of the concentration of OH^- on the pseudo-first-order rate constant k_0 in the hydrolysis of MNTS at 25 °C. (○) at a constant ionic strength of 0.46 mol dm^{-3} (controlled with NaCl). (●) without addition of NaCl.

Table 1. Influence of the composition of the ethanol–water medium on the rate of hydrolysis of MNTS. $[OH^-] = 0.0268$, $[MNTS]_0 = 3.25 \times 10^{-5}$ mol dm^{-3}.

Ethanol (%) (v/v)	$k_0/10^{-3}$ s^{-1}
8.0	2.64
10.2	2.71
13.4	2.70
16.7	2.96
21.1	3.08
24.3	3.06

(except where stated otherwise, the alcohol was ethanol), and were stored in the dark at low temperature; although stable for months under these conditions, they were frequently freshly prepared.

The reaction of MNTS with ammonia, amines, and amino acids was studied using buffers made up with the amine or amino acid so as to ensure the constancy of pH. The amines, amine salts, or amino acids, purified by distillation or recrystallization, were weighed and used to prepare buffers by addition of the necessary amounts of acid or NaOH to attain the desired pH. Their concentration was always much greater than that of MNTS. A Radiometer pHM 82 pH-meter was used to measure pH.

Reaction kinetics were studied using a Uvikon 820 spectrophotometer at λ 250 nm by recording the change in absorbance due to the disappearance of MNTS. When reagents had appreciable absorbance at λ 250 nm, 260–270 nm wavelength was employed. The initial concentration of MNTS (0.15–1 10^{-4} mol dm^{-3}), was always much smaller than that of the other reagents present.

The kinetics were studied by the integration method except in the case of the reaction between MNTS and NH_3, which was so slow that the initial-rate method was used; in all other cases the absorbance–time data fitted the first-order integrated rate equation perfectly until the reaction was at least 90% complete. In Equation (1)

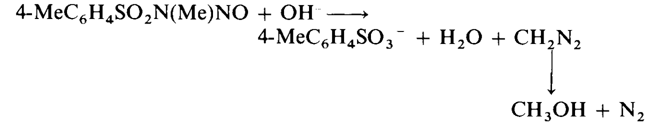
$$\ln(A_t - A_\infty) = \ln(A_0 - A_\infty) - k_0 t \qquad (1)$$

A_0, A_t, and A_∞ are the absorbances at times 0, t, and ∞ respectively and k_0 is the observed pseudo-first-order rate constant. Equation (1) was fitted to the kinetic data using the algorithms of Davies *et al.*[8] All kinetic experiments were performed at 25 °C and the values of pseudo rate constants were reproducible to within 3%.

As indicated below, the identity of the reaction products was confirmed from the characteristics of the u.v. spectrum of the reactions on completion, and in some cases by h.p.l.c. with 1:1 acetonitrile–water as the eluant. The presence of nitrite was detected by Shinn's method.[9]

Results and Discussion

Hydrolysis of MNTS.—In spite of its being frequently employed to synthesise diazomethane (Scheme 3), there has never, to our knowledge, been any kinetic study of the alkaline hydrolysis of MNTS.



$$4\text{-}MeC_6H_4SO_2N(Me)NO + OH^- \longrightarrow 4\text{-}MeC_6H_4SO_3^- + H_2O + \underset{\displaystyle CH_3OH + N_2}{\underset{\downarrow}{CH_2N_2}}$$

Scheme 3.

In an initial series of experiments we studied the influence of OH^- concentration on the observed pseudo-first-order rate constant k_0. In these experiments the concentration of ethanol was 23% (v/v), the initial concentration of MNTS was 6.0×10^{-5} mol dm^{-3}, and the concentration of NaOH was varied between 0.0154 and 0.277 mol dm^{-3}. NaCl was used to keep ionic strength at a constant 0.46 mol dm^{-3}. The results (Figure 1) show the reaction to be first order in OH^-.

$$r = k_1[OH^-][MNTS] \qquad (2)$$

Fitting equation (2) to the experimental data yields a value of 0.127 dm^3 mol^{-1} s^{-1} for k_1. Ionic strength was found to have little effect on the reaction rate, as illustrated in Figure 1 (which includes data that in spite of having been obtained in the absence of NaCl lie exactly on the line calculated from the data obtained at $I = 0.46$ mol dm^{-3}).

In experiments to determine the influence of ethanol concentration on the reaction rate, ethanol was found to produce a very small catalytic effect, the reaction rate increasing by just 15% on raising the ethanol concentration from 8–24% (Table 1). In experiments in which the solvent was 7.1% methanol instead of ethanol, first-order behaviour in OH^- was again observed when the concentration of NaOH was varied (Table 2), but the reaction rate was then approximately twice that in ethanol–water ($k_1 = 0.276$ dm^3 mol^{-1} s^{-1}). Table 2 shows that the catalytic effect of methanol was much greater than that of ethanol. This increase in the rate of hydrolysis as the alcohol content of the medium increases is probably due to the increased amount of RO^- ions, which can also act as nucleophiles on the SO_2 group. In fact, the solvent effect for this reaction is the opposite of that expected, since when the hydrolysis is carried out in dioxane–water a change in the percentage of

Table 2. Influence of [OH$^-$] and the composition of the methanol–water medium on the rate of hydrolysis of MNTS. [MNTS]$_0$ = 8.60 × 10^{-5} mol dm^{-3}.

% Methanol (v/v)	[OH$^-$]/10^{-2} mol dm^{-3}	k_0/10^{-3} s^{-1}
7.1	1.14	2.69
7.1	1.71	4.10
7.1	2.29	5.37
7.1	2.86	7.35
7.1	6.16	16.2
7.1	9.24	25.0
10.0	2.29	6.45
12.1	2.29	7.34
14.3	2.29	8.19
15.7	2.29	8.62
17.9	2.29	9.12
19.3	2.29	9.43

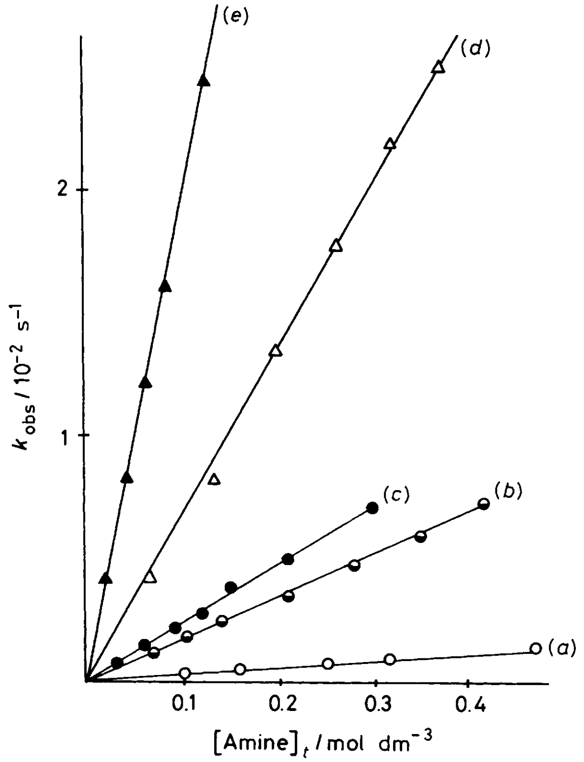


Figure 2. Influence of the concentration of secondary amines on the rate of decomposition of MNTS at 25 °C and ethanol (17%). Buffer ratio = [Free amine]/[Ammonium salt] (*a*) Morpholine, buffer ratio 1. (*b*) Piperazine, buffer ratio 1.4. (*c*) Sarcosine, buffer ratio 0.94. (*d*) Proline, buffer ratio 0.98. (*e*) Dimethylamine, buffer ratio 1.

dioxane from 7.1–40% causes a decrease in the reaction rate by a factor of 2.2.

The Reaction of MNTS *with Amines.—Reaction of* MNTS *with secondary amines.* The ability of MNTS to react with secondary amines to afford nitrosamines was known in the 1950s: Backer and de Boer[10] reported that the decomposition of MNTS in ethanol in the presence of piperidine gave *N*-nitrosopiperidine and not even traces of diazomethane. Takizawa[11] reported that the reaction of MNTS with diphenylamine gave diphenylnitrosamine. The transnitrosating ability of MNTS in neutral or basic conditions is also active in its reaction with cysteine to give *S*-nitrosocysteine, which then gradually decomposes to cystine.[12] The most recent evidence of this ability has been reported by Garcia *et al.*,[5] who found that the reaction between MNTS and diethylamine affords a quantitative yield of *N*-nitrosodiethylamine and drew attention to the contrast between this behaviour and that of ordinary nitrosamides, the CO group of which is the object of nucleophilic attack by amines. Thus, there appears to be no doubt that the reaction between MNTS and amines takes place as shown in Scheme 4.

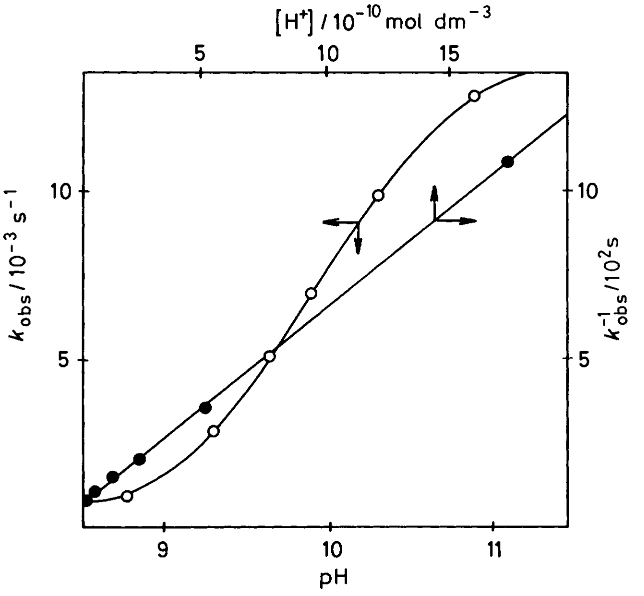


Figure 3. Influence of acidity on the pseudo-first-order rate constant k_0 in the reaction between piperazine and MNTS in ethanol–water (12%). [Piperazine]$_t$ = 0.279 mol dm^{-3}.

$$4\text{-MeC}_6\text{H}_4\text{SO}_2\text{N(Me)NO} + \text{R}_2\text{NH} \longrightarrow 4\text{-MeC}_6\text{H}_4\text{SO}_2\text{N(Me)H} + \text{R}_2\text{NNO}$$

Scheme 4.

For the reaction between MNTS and morpholine under the conditions employed in the kinetic studies described below, we have confirmed the appearance of the products on the right-hand side of Scheme 4 by comparing the h.p.l.c. results with those of authentic samples. The u.v. spectra of the reaction mixtures when reaction was complete were in all cases in keeping with the formation of these products (since nitrosamines have considerable absorbance at 250 nm, the net changes in absorbance recorded in these experiments were smaller than in the hydrolysis reactions discussed above).

The reaction of MNTS with the following seven secondary amines was studied: morpholine (MOR), *N*-methylpiperazine (MPIP), piperazine (PIP), sarcosine (SAR), dimethylamine (DMA), proline (PRO), and piperidine (PIPER). In all cases the influence of the amine concentration on the reaction rate at constant pH was studied, pH being kept constant by the amine/ammonium salt buffer itself. Except where otherwise stated, the solvent contained ethanol (17% v/v). Figure 2 shows some of the results obtained, k_0 always depended linearly on the concentration of buffer.

Figure 3 shows the results of experiments to determine the influence of pH on the rate of reaction with piperazine, which was found to increase with the basicity of the medium and thus confirm that the reaction took place between the free amine R$_2$NH and MNTS (Scheme 5).

Table 3. Values of k_2 for the reaction between MNTS and various secondary amines in ethanol–water (17%).

Amine	pK_a	k_2/dm^3 mol^{-1} s^{-1}
Morpholine	8.49[a]	0.004 96
N-methylpiperazine[b]	9.08[c]	0.007 81
Piperazine	9.73[d]	0.0298
Sarcosine	10.20[e]	0.0487
Proline	10.64[f]	0.129
Dimethylamine	10.80[g]	0.384
Piperidine	11.12[f]	0.160

[a] H. B. Hetzer, R. G. Bates, and R. A. Robinson, *J. Phys. Chem.*, 1966, **70**, 2869. [b] In this case and at concentrations above 0.5 mol dm^{-3} experimental points deviate slightly (downwards) from a straight line; the possibility of a medium effect at this high concentration is not discarded. [c] Ref. 18. [d] Ref. 13. [e] R. M. Izatt, J. J. Christensen, and V. Kothari, *Inorg. Chem.*, 1964, **3**, 1565. [f] Ref. 17. [g] S. Bergstrom and G. Olofsson, *J. Chem. Thermodyn.*, 1977, **9**, 143.

$$R_2NH_2^+ \underset{}{\overset{K_a}{\rightleftharpoons}} R_2NH + H^+$$

$$R_2NH + 4\text{-}MeC_6H_4SO_2N(Me)NO \xrightarrow{k_2} 4\text{-}MeC_6H_4SO_2N(Me)H + R_2NNO$$

Scheme 5.

According to the above mechanism the pseudo rate constant k_0 is given by equation (3), where $[\text{Amine}]_t = [R_2NH] + [R_2NH_2^+]$, so that $1/k_0$ depends linearly on $[H^+]$ [equation (4)].

$$k_0 = k_2K_a[\text{Amine}]_t/(K_a + [H^+]) \quad (3)$$

$$1/k_0 = 1/k_2[\text{Amine}]_t + [H^+]/k_2K_a[\text{Amine}]_t \quad (4)$$

When fitted to the experimental data (Figure 3), this relationship implies that $K_a = 1.2 \times 10^{-10}$ and hence that pK_a 9.9, in agreement with the known value for piperazine.[13] Similar results were obtained with the other amines considered.

Table 3 lists, for each amine, its pK_a and the value of k_2 determined in a series of experiments in which the concentration of amine was varied systematically. There is a fairly good linear relationship of gradient 0.6 between logk_2 and pK_a (Figure 4), showing that the more basic amines produce nitrosamines more readily; the exceptional reactivity of dimethylamine is perhaps due to its smaller steric hindrance.

Table 4 lists the results of a study of the influence of ethanol concentration on the aminolysis of MNTS by piperazine and morpholine. Contrary to what had been observed in the hydrolysis of MNTS, an increase in the percentage of ethanol in the medium causes a decrease in the rate of decomposition of MNTS. This is in keeping with the expected effect of the solvent, since in this case the reagents are non-ionic and react *via* a transition state which features a degree of charge separation.

The reaction of MNTS *with primary amines and related compounds.* The reaction of MNTS with primary amines was studied in a way similar to that described above. In this case, the nitrosamine formed initially is unstable and rapidly decomposes, Scheme 6.

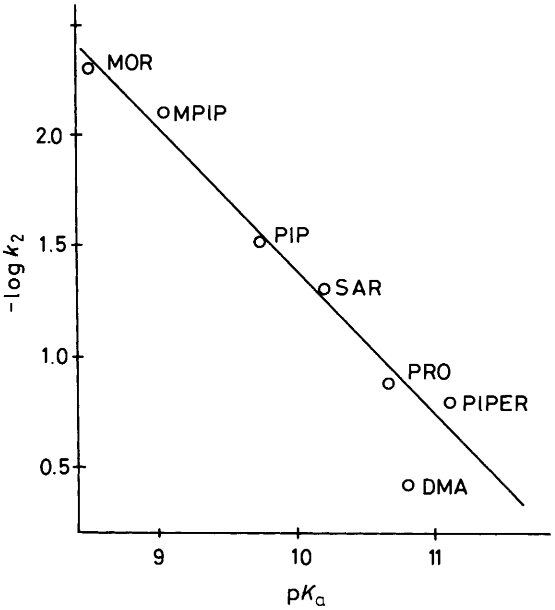


Figure 4. Brønsted plot for the reaction between secondary amines and MNTS.

Table 4. Influence of the composition of the ethanol–water medium on the rate of aminolysis of MNTS. Buffer ratios indicate [free amine]/[ammonium salt].

Ethanol/(%)	k_0/10^{-3} s^{-1}
7.5[a]	1.66
12[a]	1.44
14.1[a]	1.32
18.3[a]	1.13
20.5[a]	1.07
5.3[b]	8.07
8.6[b]	7.61
11.8[b]	7.21
16.2[b]	6.21
20.5[b]	4.85
23.8[b]	4.42
28.1[b]	3.71

[a] Reaction of MNTS with MOR. $[\text{MOR}]_t = 0.471$ mol dm^{-3}. Buffer ratio = 1. [b] Reaction of MNTS with PIP. $[\text{PIP}]_t = 0.279$ mol dm^{-3}. Buffer ratio = 1.4.

$$4\text{-}MeC_6H_4SO_2N(Me)NO + RNH_2 \longrightarrow 4\text{-}MeC_6H_4SO_2N(Me)H + \underset{\substack{\downarrow \\ \text{decomposition products}}}{RNHNO}$$

Scheme 6.

Analysis of reaction products[5] shows that the same is true for the reaction with ammonia, although in this case nitrosation is very slow, which meant our having to determine the approximate rate constant by the initial-rate method. The other compounds used in this study were methylamine, butylamine, glycine, hydroxylamine, and hydrazine. Figure 5 shows the results of experiments typical of those in which the linear dependence of k_0 on $[\text{Amine}]_t$ was confirmed. The non-zero intercept that appears for some reactions corresponds quantitatively with the competitive hydrolysis of MNTS by OH$^-$, so that in view of the results reported above for this latter reaction the empirical rate is given by equation (5).

$$k_0 = k_1[OH^-] + k_2[RNH_2] \quad (5)$$

That it is once more the free amine that is the reactive form was confirmed in a series of experiments carried out at different

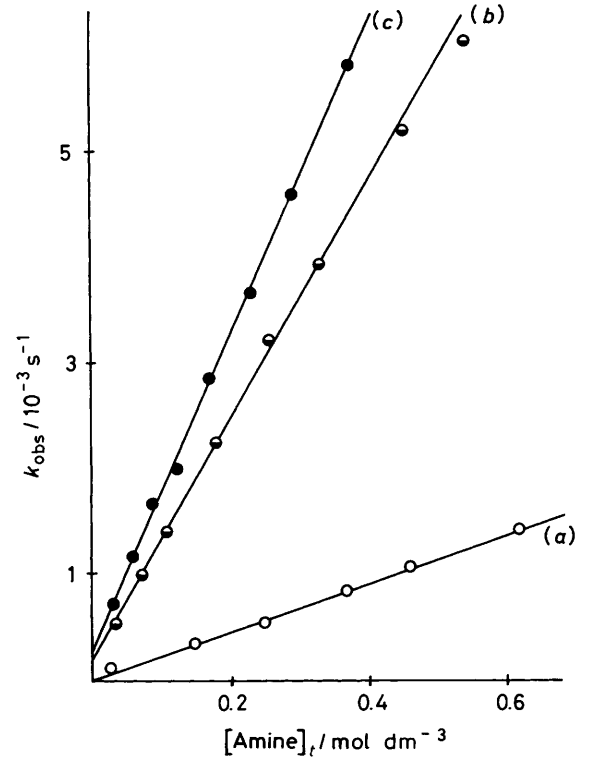


Figure 5. Influence of the concentration of primary amines on the pseudo first-order rate constant k_0 in the decomposition of MNTS at 25 °C and ethanol (17%). Buffer ratio = [free amine]/[ammonium salt]. (*a*) Hydroxylamine, buffer ratio 1.05. (*b*) Methylamine, buffer ratio 1.3. (*c*) Methylamine, buffer ratio 5.

Table 5. Values of k_2 for the reaction between MNTS and various primary amines in ethanol–water (17%).

Amine	pK_a	k_2/dm^3 mol^{-1} s^{-1}
Hydroxylamine	5.95[a]	0.004 36
Hydrazine	7.96[b]	0.0218
Ammonia	9.24[c]	*ca.* 0.000 022
Glycine	9.78[d]	0.000 894
Methylamine	10.64[e]	0.018
Butylamine	10.64[f]	0.005 46

[a] P. Lumme, P. Lahermo, and J. Tummavuori, *Acta Chem. Scand.*, 1965, **19**, 2175. [b] K. Sallavo and P. Lumme, *Suomen Kem., Ser. B*, 1967, **40**, 155. [c] R. G. Bates and G. D. Pinching, *J. Res. Nat. Bur. Stand.*, 1949, **42**, 419. [d] E. J. King, *J. Am. Chem. Soc.*, 1951, **73**, 155. [e] H. S. Harned and B. B. Owen, *J. Am. Chem. Soc.*, 1930, **52**, 5079. [f] R. G. Bates and H. B. Hetzer, *J. Phys. Chem.*, 1961, **65**, 667.

pH. Values of k_2 calculated by first determining the slope of k_0-$[Amine]_t$ plots and then correcting for protonated amine are listed in Table 5.

These data show that for the primary amines studied there is no correlation between $\log k_2$ and pK_a. Comparison with the reactivities of secondary amines shows that ammonia and ordinary primary amines react much more slowly than secondary amines of similar basicity (about 600 times more slowly in the case of ammonia), whereas hydroxylamine and hydrazine are considerably more reactive (40 and 10 times respectively) than the corresponding secondary amine. The greatest reactivity of hydroxylamine and hydrazine is attributed to the α effect.[14]

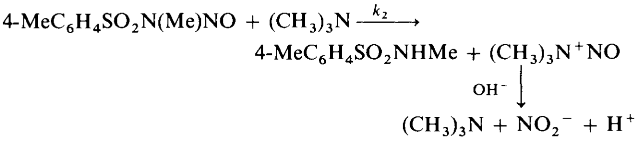
Table 6. Proportion of MNTS NO converted into NO_2^- by reaction with tertiary amines, as measured by Shinn's method.

Amine	$[Amine]_t$/ mol dm^{-3}	[Free amine]/ [Ammonium salt]	NO_2^-/(%)
Trimethylamine	0.0600	1.2	96
Trimethylamine	0.150	1.2	98
Trimethylamine	0.300	1.2	95
N-Methylpiperidine	0.0376	4	82
N-Methylpiperidine	0.0753	4	85
N-Methylpiperidine	0.113	4	83
N-Methylpiperidine	0.188	4	82
N-Methylpiperidine	0.226	4	83
N-Methylpiperidine	0.264	4	87

Table 7. Influence of the acidity on k_0 for reaction of MNTS with trimethylamine. $[Amine]_t$ = 0.0537 mol dm^{-3}. $[MNTS]_0$ = 1.2 × 10^{-4} mol dm^{-3}. Ethanol (13%).

pH	k_0/10^{-3} s^{-1}
9.21	1.51
9.45	2.38
9.80	4.41
9.96	5.60
10.05	6.12

The reaction of MNTS *with tertiary amines.* Though the reactivity of tertiary amines with nitrosating agents is generally relatively small,[15] they considerably increase the rate of disappearance of MNTS. Though in the first step of this reaction the tertiary amine presumably acts, like primary and secondary amines, as a nucleophile, it is not clear what then happens to the R_3N^+NO that is formed (the analogous problem in acid media has been the subject of a number of studies [16]). The simplest of the tertiary amines employed in the present study, trimethylamine (TMA), has a pK_a[17] *ca.* 9.8, so that the basicity of the buffers in which it was used suggests that $(CH_3)_3N^+NO$ might be hydrolysed by OH^-, resulting in a mechanism (Scheme 7) that is analogous to that of the reaction between tertiary amines and alkyl nitrites.[6]



$$\begin{tikzcd}[column sep=small] 4\text{-}MeC_6H_4SO_2N(Me)NO + (CH_3)_3N \arrow[r, "k_2"] & 4\text{-}MeC_6H_4SO_2NHMe + (CH_3)_3N^+NO \arrow[d, "OH^-"'] \\ & (CH_3)_3N + NO_2^- + H^+ \end{tikzcd}$$

Scheme 7.

This hypothesis was verified by analysing the final reaction mixtures using Shinn's method,[9] which confirmed that the conversion of the MNTS nitroso group into nitrite was practically quantitative (Table 6).* In the case of *N*-methylpiperidine (MPIPER), the results of Table 6 show that this was also the major reaction (85%), but that a competing secondary reaction also took place.

Like the primary and secondary amines, both *N*-methylpiperidine and trimethylamine reacted with MNTS at a rate that depended linearly on amine concentration, with values (at 13.3% ethanol) for k_2 of 2.8 × 10^{-2} dm^3 mol^{-1} s^{-1} for the former and 0.24 dm^3 mol^{-1} s^{-1} for the latter. The corresponding value for TMA at ethanol 17% was of 0.19 dm^3 mol^{-1} s^{-1}. The fact that reaction also occurs between MNTS and free amine is confirmed by the results obtained when working with trimethylamine at different pH values (Table 7). The rate constants thus

---

* When reaction mixtures resulting from treatment of MNTS with primary or secondary amines, or OH^- were analysed by the same method, the percentages of NO_2^- detected were not higher than 5%.

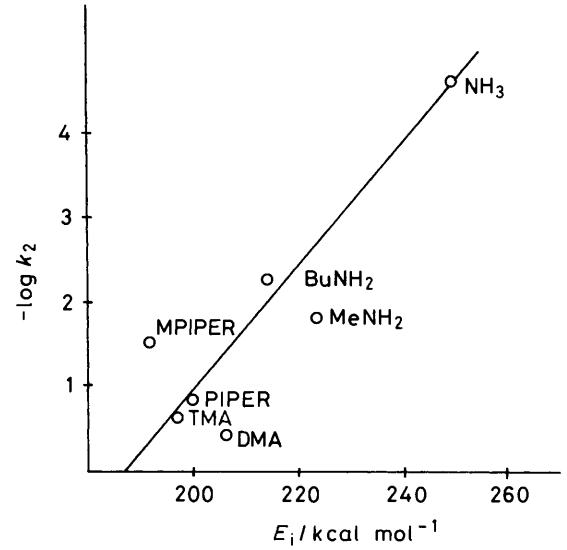


Figure 6. Reactivities of amines with MNTS plotted *vs.* their vertical ionization potentials. (The E_i for $BuNH_2$ is assumed to be the same as that of Bu^iNH_2 and Bu^sNH_2, which are both 214.5 kcal mol^{-1}).*

obtained fit equation (4) well, yielding for pK_a a value of 10.1, in reasonable agreement with the literature value.[17] The fact that trimethylamine is more reactive than *N*-methylpiperidine parallels the similar relationship between the secondary amines dimethylamine and piperidine.

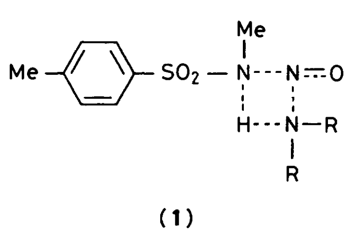
When tertiary amines of pK_a 7–8 were used, k_0 no longer depended linearly on amine concentration, and the formation of NO_2^- was no longer quantitative and depended on the experimental conditions. This may be attributed to the lower concentration of OH^- and the consequent predominance of complex alternative pathways for the decomposition of R_3N^+NO. These latter mechanisms are presumably related to those that likewise complicate the nitrosation of tertiary amines in acid media, and which have not yet been completely elucidated.[16]

Comparison of the reactivities of primary, secondary, and tertiary amines with MNTS. The above results appear to show that amines exert nucleophilic attack on the N=O group of MNTS rather than on its SO_2 group, which contrasts with their behaviour with other nitrosamides. The hydroxide ion, however, reacts efficiently with the SO_2 group but not with the N=O group. The reactivity of secondary amines correlates well with their pK_a, and the tertiary amines investigated have reactivities similar to those of secondary amines of similar pK_a, trimethylamine proving to be remarkably reactive, but primary amines and NH_3 react considerably more slowly (except for those with the α effect), and their basicities and reactivities are uncorrelated.

The above behaviour is remarkably like that observed for reactions between amines and alkyl nitrites (phenethyl nitrite in particular); Oae *et al.*[6] concluded that the reaction was 'orbital-controlled,' a conclusion later confirmed by Casado *et al.*[18] It, therefore, seems likely that a similar cause is responsible for the behaviour observed in the present study; and indeed, considering the significant structural differences among the amines used, there is quite a good correlation between logk_2 and vertical ionization potentials of those for which the latter is known[19] (Figure 6).

Conclusions

MNTS Is able to act as an effective nitrosating agent in basic media by direct transfer of its N=O group to amine N atoms. The rate at which this reaction occurs is similar to that of 2-hydroxyethyl nitrite,[18] the electronegative β-substituent of which makes it relatively reactive among alkyl nitrites, the only homogeneous group of nitrosating agents active in basic media. Finally, the fact that tertiary amines are no less reactive than primary or secondary amines would appear to exclude the possibility of the latter reacting *via* a cyclic transition state (**1**),



(**1**)

the most likely mechanism apparently being stepwise proton transfer.

---

* 1 cal = 4.186 J.

References

1 (*a*) B. C. Challis and S. P. Jones, *J. Chem. Soc., Perkin Trans. 2*, 1979, 703; (*b*) K. Yoshida, K. Yano, and K. Nagamatsu, *J. Chem. Soc., Perkin Trans. 2*, 1985, 437.
2 (*a*) A. Castro, J. R. Leis, M. E. Peña, and J. Vázquez Tato, *J. Chem. Soc., Perkin Trans. 2*, 1986, 117; (*b*) J. K. Snyder and L. M. Stock, *J. Org. Chem.*, 1980, **45**, 1990.
3 D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1976, 691.
4 (*a*) Th. J. de Boer and H. J. Backer, *Recl. Trav. Chim. Pays-Bas*, 1954, **73**, 229; (*b*) M. Pearce, *Helv. Chim. Acta*, 1980, **63**, 887.
5 J. Garcia, J. González, R. Segura, F. Urpi, and J. Vilarrasa, *J. Org. Chem.*, 1984, **49**, 3322.
6 S. Oae, N. Asai, and K. Fujimori, *J. Chem. Soc., Perkin Trans. 2*, 1978, 1124.
7 (*a*) H. Maltz, M. A. Grant, and M. C. Navaroli, *J. Org. Chem.*, 1971, **36**, 363; (*b*) J. Casado, M. A. López Quintela, M. Mosquera, M. F. Rodriguez Prieto, and J. Vázquez Tato, *Ber. Bunsenges. Phys. Chem.*, 1983, **87**, 1208.
8 J. Casado, M. Mosquera, A. Rivas, M. F. Rodriguez Prieto, and J. A. Santaballa, *Comput. Chem.*, 1983, **7**, 209.
9 C. N. Berry and B. C. Challis, *J. Chem. Soc., Perkin Trans. 2*, 1974, 1638.
10 H. J. Backer and Th. J. de Boer, *Proc., Ned. Akad. Wet., Ser. B*, 1951, **54**, 191.
11 T. Takizawa, *J. Pharm. Soc. Jpn.*, 1950, **70**, 490.
12 (*a*) U. Schutz and D. R. McCalla, *Can. J. Chem.*, 1969, **47**, 2021; (*b*) D. L. H. Williams, personal communication.
13 H. B. Hetzer, R. A. Robinson, and R. G. Bates, *J. Phys. Chem.*, 1968, **72**, 2081.
14 R. G. Pearson, 'Advances in Linear Free Energy Relationships,' eds. N. Chapman and J. Shorter, Plenum Press, 1972, ch. 6.
15 D. L. H. Williams, 'Nitrosation,' CUP, Cambridge, 1988, ch. 2, p. 98.
16 (*a*) P. A. S. Smith and R. L. Loeppky, *J. Am. Chem. Soc.*, 1967, **89**, 1147; (*b*) G. E. Hein, *J. Chem. Educ.*, 1963, **40**, 181.
17 'Handbook of Chemistry and Physics,' 60th edn., CRC Press Inc., Boca Raton, 1980.
18 J. Casado, A. Castro, M. A. López Quintela, and F. M. Lorenzo, *Bull. Soc. Chim. Fr.*, 1987, 401.
19 D. H. Aue, H. M. Webb, and M. T. Bowers, *J. Am. Chem. Soc.*, 1976, **98**, 311.

Received 22nd *December* 1988; *Paper* 8/05004J