Nitrosothiosulphate Ion ($S_2O_3NO^-$) as a Nitrosating Species

Tracey Bryant and D. Lyn H. Williams *
Department of Chemistry, University Science Laboratories, South Road, Durham, DH1 3LE
M. Hoshair H. Ali and Geoffrey Stedman *
Department of Chemistry, University College, Swansea, SA2 8PP

Analysis of rate data obtained for the nitrosation of *N*-methylaniline in the presence of thiosulphate ion indicates that the nitrosothiosulphate ion is formed and that it acts as a nitrosating species. The bimolecular rate constant for reaction in water at 25 °C is evaluated as 1.2×10^4 l mol^{-1} s^{-1} from the variation of the observed rate constant with [*N*-methylaniline] and also as 1.2×10^4 l mol^{-1} s^{-1} from the variation of the observed rate constant with [thiosulphate ion]. Results are also reported, for comparison purposes, for the corresponding reactions of nitrosyl bromide and nitrosyl thiocyanate. Nitrosyl bromide (in its reaction with *N*-methylaniline) is more reactive than the nitrosothiosulphate ion by a factor of *ca.* 4×10^5, and nitrosyl thiocyanate is more reactive than the nitrosothiosulphate ion by a factor of *ca.* 1.5×10^4. There is also kinetic evidence of nitrosation by the nitrosothiosulphate ion of methanol, hydrazine, and a thiol. The decomposition of $S_2O_3NO^-$ follows the rate law $-d[S_2O_3NO^-]/dt = k_1[S_2O_3NO^-] + k_2[S_2O_3NO^-]^2$. This is interpreted in terms of two concurrent reaction pathways, the first involving the formation of the radical anion $S_2O_3^{-\bullet}$ in a rate-limiting homolysis (followed by rapid dimerisation) and the second a bimolecular reaction of two $S_2O_3NO^-$ ions with concurrent S–N bond fission and S–S bond formation yielding the tetrathionate ion $S_4O_6^{2-}$.

A number of nitrosating agents (NOX) have been detected kinetically as intermediates by the observation of catalysis of nitrous acid nitrosation reactions in acid solution for a large range of substrates, by the addition of non-basic nucleophiles X^-.[1,2] Species identified in this way include nitrosyl chloride, nitrosyl bromide, nitrosyl iodide, nitrosyl thiocyanate, and the nitrososulphonium ion derived from thiourea. Some of these reagents are known as stable species capable of effecting nitrosation; others have only been detected kinetically by this procedure. The Scheme outlines the generally accepted series of reactions for reaction with a general substrate S. The range of substrates which have been examined is large, and includes amines both aliphatic and aromatic, primary, secondary, and tertiary as well as alcohols, thiols, alkenes, and many other species. The overall efficiency of a catalyst X^- clearly depends on the equilibrium constant K_{NOX} for NOX formation and also the rate constant k for its subsequent reaction. In some cases this is not strictly true since under certain circumstances it is possible to achieve rate-limiting NOX formation, so that the magnitude of k in these cases is not important. However we shall concentrate on the more commonly encountered situation, where the rate-limiting step is the reaction of NOX with S. For the cases thus far examined, it appears that the magnitude of K_{NOX} is the over-riding factor in determining the efficiency of a catalyst X^-. In most cases (nitrosyl iodide is an exception) K_{NOX} has been measured independently; values are given in Table 1, where it can be seen that the values cover a very large range, contrasting with the much smaller change in k for the same species (in reaction with aniline). The net result is that the catalytic efficiency of X^- follows the sequence $Cl^- < Br^- < SCN^- < SC(NH_2)_2$.

$$HNO_2 + X^- + H^+ \overset{K_{NOX}}{\rightleftharpoons} NOX + H_2O$$

$$NOX + S \xrightarrow{k} SNO^+ + X^-$$

$$SNO^+ \xrightarrow{\text{fast}} \text{final product}$$

Scheme.

Table 1. Values of K_{NOX} at 25 °C in water and of k for the reaction of NOX with aniline at 25 °C in water

NOX	K_{NOX}/l^2 mol^{-2}	k/l mol^{-1} s^{-1}
NOCl	1.1×10^{-3} [a]	2.2×10^9 (ref. 7)
NOBr	5.1×10^{-2} [b]	1.7×10^9 (ref. 7)
NOSCN	30 [c]	1.9×10^8 (ref. 8)
$(NH_2)_2CSNO^+$	5 000 [d]	1.3×10^6 (ref. 8)
$S_2O_3NO^-$	1.66×10^7 [e]	—

[a] H. Schmid and E. Hallaba, *Monatsh.*, 1956, **87**, 560. [b] H. Schmid and E. Fouad, *Monatsh.*, 1957, **88**, 631. [c] G. Stedman and P. A. E. Whincup, *J. Chem. Soc.*, 1963, 5796. [d] K. Al-Mallah, P. Collings, and G. Stedman, *J. Chem. Soc., Dalton Trans.*, 1974, 2469. [e] Ref. 3.

Recently Garley and Stedman[3] have investigated mechanistically the reaction between thiosulphate ion and nitrous acid, which has been known for many years.[4] The yellow solution produced is believed to contain the nitrosothiosulphate ion $S_2O_3NO^-$; a number of other species containing the S–NO grouping are similarly coloured yellow in solution, but are often not sufficiently stable to be isolated in pure form.[5] Rate constants were established for the formation of $S_2O_3NO^-$ and its equilibrium constant (K_{NOX}) was estimated as 1.66×10^7 l^2 mol^{-2}. This value is many powers of ten larger than the other K_{NOX} values in Table 1, so that if $S_2O_3NO^-$ has some reactivity as a nitrosating agent, then it is likely that thiosulphate ion catalysis of nitrosation (and diazotisation) should occur, and that this catalysis should be quite marked if the value of K_{NOX} is a major factor. In an earlier comparative study[6] catalysis of morpholine nitrosation by thiourea, thiocyanate ion, and bromide ion was observed in the quantitative ratios 4 200:240:1, the large effect of thiourea highlighting the importance of the size of K_{NOX}. If nitrosation by $S_2O_3NO^-$ were an effective process, then this would represent a rather unusual example of electrophilic nitrosation brought about by a formally *negatively* charged species. This paper presents the results of a study which investigates this possibility.

Experimental

All materials were of the highest purity grade available; *N*-methylaniline was distilled before use. All kinetic measurements were carried out spectrophotometrically at 25 °C in water. Those involving the nitrosation of *N*-methylaniline were followed at 275 nm by noting the appearance of *N*-nitroso-*N*-methylaniline, which was produced quantitatively in all cases. Some of the experiments were carried out in a conventional spectrophotometer; the more rapid reactions were followed by the stopped-flow technique using an instrument interfaced with an Apple II microcomputer. The mean of at least five separate kinetic runs was taken. The decomposition of $S_2O_3NO^-$ and the reactions in the presence of sulphamic acid and hydrazine were followed at 420 nm by noting the disappearance of $S_2O_3NO^-$, which was generated immediately prior to use from equal concentrations of thiosulphate ion and nitrous acid. The products from the reaction of hydrazine were shown to include the thiosulphate ion (by precipitation of the silver salt, which was soluble in excess of thiosulphate), and hydrazoic acid [by its removal in a stream of nitrogen and subsequent reaction with iron(III) to give the characteristic red colour].

Results and Discussion

(a) *Nitrosation of* N-*Methylaniline*.—The kinetics of nitrosation of *N*-methylaniline (NMA) in aqueous acid solution (0.1M) were established by following the appearance of the product *N*-methyl-*N*-nitrosoaniline at 275 nm. Reactions were carried out under first-order conditions with typically [NMA] 1.8—4.8 × 10^{-3}M, [HNO_2] 1 × 10^{-4}M, and [thiosulphate ion] 2 × 10^{-4}M. For comparison, the reaction was also studied kinetically in the presence of the same concentration of bromide ion and also of thiocyanate ion. Throughout, quantitative formation of the product *N*-methyl-*N*-nitrosoaniline was observed, which at the acidity of the experiments (0.1M) reacts further only negligibly slowly by the Fischer–Hepp rearrangement. In all cases good first-order dependence upon total nitrous acid concentration was found and values of the first order rate constant k_0 (defined by $-d[HNO_2]/dt = k_0[HNO_2]$) were obtained. Figure 1 shows the linear dependence of k_0 upon [NMA] for the three reactions with $[Br^-] = [SCN^-] = [S_2O_3^{2-}] = 2 \times 10^{-4}$M. The Scheme predicts that the general form of the expression for k_0 will be given by equation (1), where $[NMA]_T$ is the total concentration of added

$$k_0 = \frac{kK_{NOX}[H^+][X^-][NMA]_TK_a}{(1 + K_{NOX}[H^+][X^-])(K_a + H^+)} \qquad (1)$$

N-methylaniline and K_a is the dissociation constant for its protonated form. However at $[H^+] = 0.1$M, $K_a \ll [H^+]$ so the general expression simplifies for all nucleophilic species to equation (2). Further, for both $X^- = Br^-$ and SCN^- the

$$k_0 = \frac{kK_{NOX}[X^-][NMA]_TK_a}{(1 + K_{NOX}[H^+][X^-])} \qquad (2)$$

inequality $1 \gg K_{NOX}[H^+][X^-]$ applies so that the expression for k_0 is simplified even further to equation (3). This assumes

$$k_0 = kK_{NOX}[X^-][NMA]_TK_a \qquad (3)$$

that the rate-limiting step under these conditions is the reaction of NOX with NMA, and this is confirmed for all three nucleophiles by the pure first-order dependence upon $[NMA]_T$ (Figure 1). This allows the calculation of k for both Br^- and SCN^-, as 5.0 × 10^9 and 3.1 × 10^8 l mol^{-1} s^{-1} respectively. These qualities do not appear to have been reported previously, but the corresponding k values for reaction with aniline are 1.7 × 10^9 and 1.9 × 10^8 (refs. 7 and 8). The values for NMA are thus a little larger than for aniline, as is to be expected, with the NOBr reaction very close to the encounter-controlled limit.

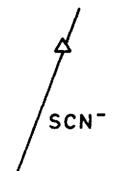


Figure 1. Kinetic substrate dependence for the nitrosation of NMA in the presence of Br^-, $S_2O_3^{2-}$, or SCN^-

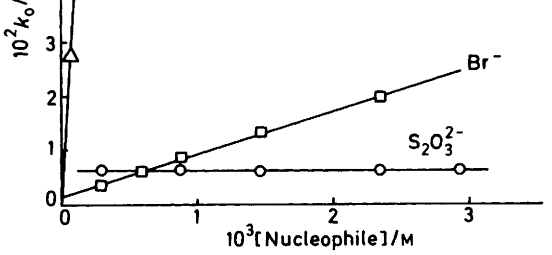


Figure 2. Kinetic dependence upon $S_2O_3^{2-}$, Br^-, or SCN^- in the nitrosation of NMA

When reaction is carried out with added thiosulphate ion however, the situation is somewhat different in that now $1 \ll K_{NOX}[H^+][S_2O_3^{2-}]$, so that the expected form of the

Table 2. Values of k [equations (3) and (4)] for reactions of NMA with NOBr, NOSCN, and $S_2O_3NO^-$

	k/l mol^{-1} s^{-1}	
Electrophile	Method (a)[a]	Method (b)[b]
NOBr	5.0×10^9	3.1×10^9
NOSCN	3.1×10^8	2.5×10^8
$S_2O_3NO^-$	1.2×10^4	1.2×10^4

[a] From the variation of k_0 with [NMA] at constant [H$^+$] and [nucleophile]. [b] From the variation of k_0 with [nucleophile] at constant [H$^+$] and [NMA].

expression for k_0 is equation (4). This enables a value of k to be calculated as 1.2×10^4 l mol^{-1} s^{-1}.

$$k_0 = \frac{k[\text{NMA}]_T K_a}{[\text{H}^+]} \qquad (4)$$

Further experiments were carried out with the same nucleophiles, this time at constant [NMA] and [H$^+$] while varying [Br$^-$], [SCN$^-$], and [$S_2O_3^{2-}$]. The results are given graphically in Figure 2 and show the expected linear dependence of k_0 upon both [Br$^-$] and [SCN$^-$] expected from equation (3) and a complete lack of dependence of k_0 upon [$S_2O_3^{2-}$] as predicted by equation (4). Again the k values were extracted from the kinetic results and are shown in Table 2, together with the earlier results derived from Figure 1. As can be seen, the agreement between the two approaches is excellent for both SCN$^-$ and $S_2O_3^{2-}$ and acceptable for Br$^-$.

So it is clear from these results that $S_2O_3NO^-$ can act as an electrophilic nitrosating agent in its reaction with *N*-methylaniline, and probably more generally. It is however significantly less reactive than NOCl, NOBr, NOSCN, and $(NH_2)_2CS^+NO$, as is perhaps to be expected for an electrophilic reagent which is formally negatively charged. The nitroprusside ion $[Fe(CN)_5NO]^{2-}$ is another example of a negatively charged nitrosating agent which can effect nitrosation of amines (and other substrates) in alkaline solution.[9] Here the mechanism is probably different in that an intermediate complex is formed where the nitroso nitrogen atom is bonded both to the iron and to the amino nitrogen atom.

At lower acidity (1×10^{-2}M) the reaction loses its kinetic dependence upon [NMA] for all three nucleophiles and the observed catalysis by Br$^-$, SCN$^-$, and $S_2O_3^{2-}$ is readily explained in terms of rate-limiting formation of NOBr, NOSCN, and $S_2O_3NO^-$ respectively.

It is not possible in general to use k_0 values to compare the catalytic reactivities of Br$^-$ and SCN$^-$ with $S_2O_3^{2-}$ because the rate law for the first two is different from that of the third. This point is made in Figure 2 where it is clear that thiocyanate ion is always a better catalyst than both bromide ion and thiosulphate ion, whereas when bromide ion and thiosulphate ion are compared, the situation depends on their concentrations. At concentrations less than 6×10^{-4}M Br$^-$ and $S_2O_3^{2-}$ the effect of thiosulphate is greater than that of bromide, whereas above this concentration the reverse is true. This arises from the different fractional conversions of HNO_2 into NOX.

(b) *Decomposition of* $S_2O_3NO^-$.—Solutions of the nitrosothiosulphate ion are relatively unstable, decomposing according to equation (5) to yield the tetrathionate ion and nitric oxide.

$$S_2O_3NO^- \longrightarrow \tfrac{1}{2}S_4O_6^{2-} + NO \qquad (5)$$

This reaction has been studied kinetically, starting with an excess of nitrous acid over the thiosulphate ion, by noting the disappearance of the yellow colour due to the nitroso species at 420 nm. Under these conditions, virtually complete conversion of thiosulphate into nitrosothiosulphate occurs, and there is no evidence of any cloudiness in the solutions due to sulphur precipitation, which normally occurs with free thiosulphate ion in acid solution. Further the extrapolated initial absorbance at 420 nm is directly proportional to the initial thiosulphate ion concentration and is independent of both the acid concentration and the (excess) nitrous acid concentration. All the kinetic runs showed second-order behaviour in the initial stages of the reaction with linear plots of [absorbance]$^{-1}$ *vs.* time, but deviated from second-order behaviour in the later stages in a way which suggested a mixed first- and second-order rate equation [equation (6)]. Rearrangement of equation (6) in the form given in equation (7) enables this to be tested quantitatively.

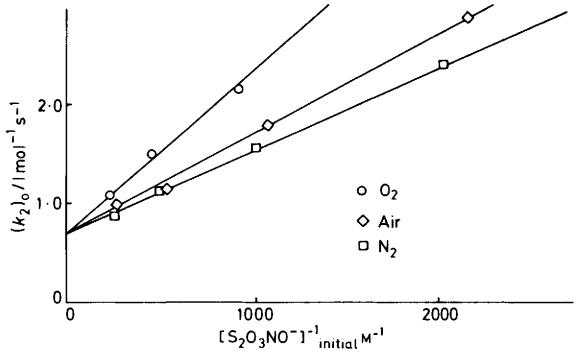


Figure 3. Plot of $(k_2)_0$ *vs.* $[S_2O_3NO^-]^{-1}_{\text{initial}}$ for the decomposition of $S_2O_3NO^-$ in air, oxygen, or nitrogen

$$\text{Rate} = -d[S_2O_3NO^-]/dt = k_1[S_2O_3NO^-] + k_2[S_2O_3NO^-]^2 \qquad (6)$$

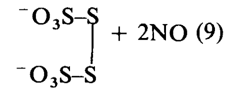
$$\text{Rate}/[S_2O_3NO^-]^2 = k_1/[S_2O_3NO^-] + k_2 \qquad (7)$$

The left hand side of equation (7) represents the observed second-order rate constant $(k_2)_0$ and so a plot of $(k_2)_0$ *vs.* $[S_2O_3NO^-]_0^{-1}$ {where $(k_2)_0$ and $[S_2O_3NO^-]_0$ are the initial values of the second-order rate constant and nitrosothiosulphate ion concentration respectively} for a series of experiments should be linear with a positive slope and intercept. Such plots are shown in Figure 3 for reactions (at constant ionic strength) in air, nitrogen, and oxygen respectively. The different slopes in the three cases show that the pathway for the first-order decomposition is affected by oxygen whereas the fact that the intercept is the same in each case shows that the second-order mechanism is not oxygen-dependent. A similar analysis in terms of initial first-order plots [equation (8)] also gave linear plots

$$\text{Rate}/[S_2O_3NO^-]_0 = (k_1)_0 = k_1 + k_2[S_2O_3NO^-]_0 \qquad (8)$$

for the initial first order rate constant $(k_1)_0$ *vs.* $[S_2O_3NO^-]_{\text{initial}}$ and values of k_1 and k_2 were comparable from the two analyses.

The interpretation of the second-order term in equation (6) is relatively straightforward, the data being consistent with the mechanism outlined in (9), where two $S_2O_3NO^-$ ions react with



$$2S_2O_3NO^- \longrightarrow \left[\begin{array}{l} ^-O_3S{-}S{-}{-}{-}{-}NO \\ \quad\quad\;\, \vdots \\ ^-O_3S{-}S{-}{-}{-}{-}NO \end{array}\right]^{\ddagger} \longrightarrow \begin{array}{l} ^-O_3S{-}S \\ \quad\quad\;\, | \\ ^-O_3S{-}S \end{array} + 2NO \qquad (9)$$

concurrent S–N bond breaking and S–S bond formation. A similar proposal has been made[10] for the mechanism of decomposition of the *S*-nitroso ion derived from thiourea and nitrous acid, where again a disulphide species is formed and nitric oxide eliminated. The interpretation of the first-order kinetic term in equation (6) is not so immediately obvious. A number of mechanisms are possible, but we suggest the rate-limiting formation of the radical anion $S_2O_3^{-\bullet}$ from S_2O_3-NO^- by homolytic fission, followed by rapid dimerisation of the radical anion to give the observed product [see (10)]. The

$$\begin{tikzcd} S_2O_3NO^- \arrow[r, "\text{slow}"] & S_2O_3^{-\bullet} + NO \arrow[d, "S_2O_3^{-\bullet}", "\text{fast}"'] \\ & S_4O_6^{2-} \end{tikzcd} \qquad (10)$$

catalysis by oxygen suggests a radical mechanism and is explicable in terms of a complex formation of the type $[S_2O_3NO{\cdot}O_2]^-$ or by a direct reaction of oxygen with $S_2O_3NO^-$ to give $S_2O_3^{-\bullet}$ and the $NO_3^\bullet$ species, which regenerates nitric oxide and oxygen. It has been suggested[11] that $NO_3^\bullet$ might be a reactive intermediate in the oxidation of nitric oxide by oxygen.

(c) *Nitrosation of Hydrazine, Sulphamic Acid, and Methanol.*—In order to test the generality of application of nitrosothiosulphate ion as a nitrosating agent, we have looked at its reaction with three substrates which are known to be substrates for nitrosation. For both hydrazine and sulphamic acid a different approach from that employed in (a) with *N*-methylaniline was used. The rate of decomposition of the nitrosothiosulphate anion was measured in the presence of increasing concentrations of both hydrazine and sulphamic acid. Equal concentrations of nitrous acid and thiosulphate ion were used to avoid complications due to an excess of either, and reaction was carried out with an excess of either hydrazine or sulphamic acid, as in (b), monitoring the disappearance of the yellow colour due to the nitrosothiosulphate anion. Both substrates catalysed the decomposition of the anion (see Figure 4 for the data for the hydrazine reaction) and the product solutions contained sulphate ion from the sulphamic acid reaction and both hydrazoic acid and thiosulphate ion from the hydrazine reaction, as expected in both cases for a nitrosation reaction.

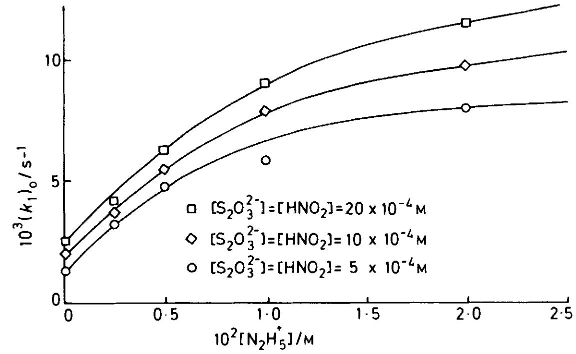


Figure 4. Plot of $(k_1)_0$ *vs.* $[N_2H_5^+]$ for the decomposition of $S_2O_3NO^-$ in the presence of added $[N_2H_5^+]$

The observed rate constant for the hydrazine reaction levels off somewhat with increasing hydrazine concentration. From the limited data available, it is not certain why this occurs, but it is possible that part of the reaction goes *via* free nitrous acid by hydrolysis of the nitrosothiosulphate ion. This is almost certainly the case with sulphamic acid, where the increase in rate constant with [sulphamic acid] is quite small. This is perhaps to be expected for this substrate, which is not subject to attack by other NOX species[12] (probably owing to the reversibility of the reaction), but reacting instead with $H_2NO_2^+$ (or NO^+).

For the reaction with methanol the nitrosothiosulphate anion was first generated from nitrous acid and thiosulphate ion before methanol was added. Reaction occurred rapidly and methyl nitrite is formed, as detected spectrophotometrically.

The experiments described in this paper show clearly that the nitrosothiosulphate ion can act as an electrophilic nitrosating agent, even though it is formally negatively charged, but that is not particularly reactive, being *ca.* 10^4 less reactive than NOSCN in its reaction with *N*-methylaniline. Nevertheless because of the large equilibrium constant for the formation of the nitrosothiosulphate ion from nitrous acid and thiosulphate ion, the catalytic effects obtained by the addition of thiosulphate ion in nitrosation reactions are quite substantial and comparable with that of bromide ion. Such catalysis however is restricted to fairly reactive substrates since the decomposition of the nitrosothiosulphate ion itself can compete with its nitrosating reactions for the less reactive substrates.

References

1 J. H. Ridd, *Quart. Rev.*, 1961, **15**, 425.
2 D. L. H. Williams, *Adv. Phys. Org. Chem.*, 1983, **19**, 381.
3 M. S. Garley and G. Stedman, *J. Inorg. Nucl. Chem.*, 1981, **43**, 2863.
4 S. Akiya, O. Nomoto, and S. Okui, *Science*, 1950, **112**, 463; B. P. Gyani and R. K. Prasad, *J. Ind. Chem. Soc.*, 1965, **42**, 1496.
5 D. L. H. Williams, *Chem. Soc. Rev.*, 1985, **14**, 171.
6 T. A. Meyer and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1981, 361.
7 M. R. Crampton, J. T. Thompson, and D. L. H. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1979, 18.
8 L. R. Dix and D. L. H. Williams, *J. Chem. Res. (S)*, 1984, 96.
9 H. Maltz, M. A. Grant, and M. C. Navaroli, *J. Org. Chem.*, 1971, **36**, 363.
10 P. Collings, M. S. Garley, and G. Stedman, *J. Chem. Soc., Dalton Trans.*, 1981, 331.
11 A. G. Sykes, 'Kinetics of Inorganic Reactions,' Pergamon, Oxford, 1966, p. 83.
12 M. N. Hughes, *J. Chem. Soc. A*, 1967, 902.

Received 28*th May* 1985; *Paper* 5/890