

## Kinetics and Mechanism of the Dehydrohalogenation of 3,5-Dichloro-2,4,6-trimethylbenzohydroxamoyl Chloride and Bromide by Tertiary Amines and Silver Nitrate

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The kinetics of the title reactions in acetonitrile have been studied by using  $\text{Et}_3\text{N}$ ,  $\text{Bu}^n_3\text{N}$ , and *N*-methylmorpholine at 25 °C and  $\text{AgNO}_3$  from 10 to 60 °C. The nitrile oxide produced by the reactions gave 1,3-cycloaddition to the solvent, the rate being determined by independent experiments. The  $\gamma$ -eliminations were found to be second-order processes apart from the reaction of bromide (III) and  $\text{AgNO}_3$  with an order higher than one in silver ion. The leaving group effect  $k_{\text{Br}}/k_{\text{Cl}}$  was found to be *ca.*  $10^3$  in the case of  $\text{AgNO}_3$  and 1.6 for *N*-methylmorpholine. The rate of reaction of chloride (II) with amines followed the trend of their  $\text{p}K_{\text{a}}$  values in water (tentative Brønsted  $\beta$  0.7). For the reaction of (II) with *N*-methylmorpholine,  $k_{\text{H}}/k_{\text{D}}$  was found to be *ca.* 2. The results have been interpreted in terms of concerted eliminations, with transition states having strong anionic character with amines and cationic character with  $\text{AgNO}_3$ . The alternative of a multistep mechanism is presented for the reaction with amines.

THE  $\gamma$ -elimination of hydrogen chloride from hydroxamoyl chlorides with tertiary amines is a general method<sup>1</sup> for the preparation of nitrile oxides, useful compounds for the synthesis of heterocycles. In

<sup>1</sup> Ch. Grundmann and P. Grünanger, 'The Nitrile Oxides', Springer, Berlin, 1971.

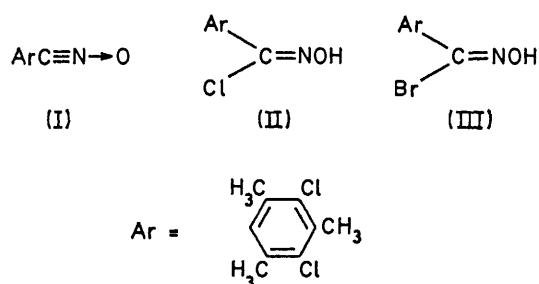
<sup>2</sup> R. H. Wiley and B. J. Wakefield, *J. Org. Chem.*, 1960, **25**, 546.

addition to the ready availability of hydroxamoyl chlorides from oximes and chlorine or nitrosyl chloride,<sup>1-3</sup> the main advantage of this reaction is that it is very fast even at low temperature, *i.e.* at 0 °C or below, a

<sup>3</sup> C. R. Kinney, E. W. Smith, B. L. Woolley, and A. R. Willey, *J. Amer. Chem. Soc.*, 1933, **55**, 3418; H. Rheinboldt, M. Dewald, F. Jansen, and O. Schmitz-Dumont, *Annalen*, 1927, **451**, 161.

condition which is often required to avoid the dimerisation of nitrile oxides to furoxans.<sup>1,4</sup> Aliphatic tertiary amines, particularly Et<sub>3</sub>N, are used for this reaction, whereas aromatic heterocyclic amines give addition products,<sup>5</sup> probably ammonium salts, and secondary and primary amines give amidoximes.<sup>1</sup> Less commonly employed is the dehydrohalogenation of hydroxamoyl chlorides by silver ion<sup>6</sup> although known for a long time.<sup>7</sup> The mechanisms of both these dehydrohalogenations have not yet been investigated nor are quantitative data available. We have thus undertaken a kinetic study of these reactions which belong, among  $\gamma$ -eliminations,<sup>8</sup> to the less common type leading to 1,3-dipolar systems.<sup>9</sup>

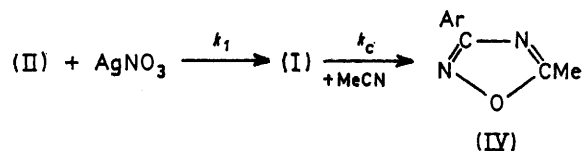
The stability of 3,5-dichloro-2,4,6-trimethylbenzotrile oxide (I) and the difference of its absorption curve in the u.v. region from those of the corresponding hydroxamoyl chloride (II) and bromide (III), which



allows spectroscopic analysis of the mixtures, led us to choose (II) and (III) for kinetic runs.

## RESULTS

*Reaction with Silver Nitrate.*—Under the conditions required for the dehydrohalogenation of (II), consecutive 1,3-cycloaddition of product (I) to the solvent acetonitrile was observed to occur to a considerable extent. A single regioisomer was obtained, which was assigned the structure of a 1,2,4-oxadiazole derivative (IV) on the basis of n.m.r. and mass spectra (see Experimental section). While a detailed discussion on this matter is postponed to a forthcoming paper, it may be stressed that the same structure was found in previous cycloadditions of nitrile oxides to nitriles.<sup>10,11</sup>



After independent kinetic measurements on the cycloaddition (Table 1), the reaction of (II) with excess of silver nitrate was studied, and pseudo-first-order rate coefficients  $k_1$  were evaluated, taking into account the known  $k_c$  values (see Experimental section). In two sets of runs at 60 °C,

coefficients  $k_1$  were found to be independent of the initial concentration of (II) and approximately proportional to that of AgNO<sub>3</sub> (Table 2), according to the rate equation (1).

$$r = k_1 C_{\text{(II)}} = k_2 C_{\text{AgNO}_3} C_{\text{(II)}} \quad (1)$$

Measurements at different temperatures in the range 32–60 °C are reported in Table 3.

TABLE 1

Kinetics of 1,3-cycloaddition of nitrile oxide (I) to solvent acetonitrile

$\theta_c/^\circ\text{C}$	$10^5 k_c/\text{s}^{-1}$
32.1	0.411; 0.413
41.4	1.219; 1.251
50.0	2.79; 3.06
60.0	7.71; 7.59

$$\log A \ 9.78 \pm 0.10; \Delta E_a \ 88.6 \pm 0.6 \text{ kJ mol}^{-1}; \Delta S^\ddagger -66.5 \pm 2.0 \text{ J K}^{-1} \text{ mol}^{-1}.$$

TABLE 2

Reaction of hydroxamoyl chloride (II) with AgNO<sub>3</sub> in acetonitrile at 60.0 °C. Rate coefficients for  $\gamma$ -elimination<sup>a</sup>

$10^4 C_{\text{AgNO}_3}/\text{mol l}^{-1}$	$10^5 C_{\text{(II)}}/\text{mol l}^{-1}$	$10^4 k_1/\text{s}^{-1}$	$k_2^b/\text{l mol}^{-1} \text{ s}^{-1}$
9.78	5.58	5.30	0.54
	5.82	5.36	0.55
3.45	8.17	5.00	0.51
	10.38	4.96	0.51
7.48	~8.6	2.00	0.58
	9.35	2.98	0.53
9.35	7.48	4.08	0.54
	9.35	4.77	0.51

<sup>a</sup> Optimised, using  $10^5 k_c = 7.65 \text{ s}^{-1}$  for the consecutive 1,3-cycloaddition. <sup>b</sup>  $k_2 = k_1/C_{\text{AgNO}_3}$ .

The reaction of (III) with silver nitrate in the range of temperature from 10 to 50 °C was fast enough for 1,3-cycloaddition to be negligible. Rate coefficients  $k_2$  were obtained according to the second-order rate equation, first-order with respect to each reactant. However,

TABLE 3

Reaction of hydroxamoyl chloride (II) with AgNO<sub>3</sub> in acetonitrile. Rate coefficients for  $\gamma$ -elimination<sup>a</sup> at different temperatures [initial concentrations: AgNO<sub>3</sub>, ca.  $9.5 \times 10^{-4}$ ; (II),  $(8 \pm 1) \times 10^{-5} \text{ mol l}^{-1}$ ]

T/°C	$10^4 k_1/\text{s}^{-1}$	$k_2^b/\text{l mol}^{-1} \text{ s}^{-1}$
32.1	0.312; 0.306	0.0318
41.4	0.776; 0.767	0.0805
50.0	1.94; 1.64	0.191
60.0	4.66; 4.77	0.504

$$\log A \ 12.83 \pm 0.27; \Delta E_a \ 83.8 \pm 1.6 \text{ kJ mol}^{-1}; \Delta S^\ddagger -8.1 \pm 5.1 \text{ J K}^{-1} \text{ mol}^{-1} \text{ (from } k_2 \text{ values)}.$$

<sup>a</sup> Optimised, using average  $k_c$  values from Table 1 for the consecutive 1,3-cycloaddition. <sup>b</sup>  $k_2 = (k_1/C_{\text{AgNO}_3})_{\text{av}}$ .

measurements at different initial AgNO<sub>3</sub> concentrations (Table 4) showed that a third-order term is present in the equation;  $k_2$  can be written as in equation (2). The dependence of  $k_2$  values on temperature for two sets of

$$k_2 = k_2' + k_2'' C_{\text{AgNO}_3} \quad (2)$$

<sup>8</sup> W. H. Saunders, jun. and A. F. Cockerill, 'Mechanisms of Elimination Reactions', Wiley, New York, 1973, p. 566.

<sup>9</sup> A. F. Hegarty, M. P. Cashman, and F. L. Scott, *J.C.S. Perkin II*, 1972, 44.

<sup>10</sup> S. Morrocchi, A. Ricca, and L. Velo, *Tetrahedron Letters*, 1967, 331.

<sup>11</sup> G. Leandri and M. Pallotti, *Ann. Chim. (Italy)*, 1957, 47, 376; A. Dondoni and G. Barbaro, *Gazzetta*, 1975, 105, 701.

<sup>4</sup> G. Barbaro, A. Battaglia, and A. Dondoni, *J. Chem. Soc. (B)*, 1970, 588.

<sup>5</sup> A. Quilico, G. Gaudiano, and A. Ricca, *Gazzetta*, 1957, 87, 638.

<sup>6</sup> J. Armand, *Bull. Soc. chim. France*, 1966, 882.

<sup>7</sup> R. C. Elderfield, 'Heterocyclic Compounds', Wiley, New York, 1961, vol. 7, pp. 475–476.

TABLE 4

Reaction of hydroxamoyl bromide (III) (initial concentration  $ca. 7 \times 10^{-5} \text{ mol l}^{-1}$ ) with  $\text{AgNO}_3$  in acetonitrile at 20.0 °C. Rate coefficients for  $\gamma$ -elimination

$10^4 C_{\text{AgNO}_3} / \text{mol l}^{-1}$	$k_2 / \text{l mol}^{-1} \text{s}^{-1}$
2.08	6.52; 7.19
5.15	10.72
8.25	13.0; 15.9
10.32	18.1; 18.1

TABLE 5

Second-order rate coefficients  $a$  for reaction of hydroxamoyl bromide (III) ( $ca. 7 \times 10^{-5} \text{ mol l}^{-1}$ ) with silver nitrate (initial concentration  $C_{\text{AgNO}_3} / \text{mol l}^{-1}$ ), in acetonitrile at various temperatures

$T / ^\circ\text{C}$	$10^4 C_{\text{AgNO}_3} = 2.04 \pm 0.04$ $k_2 / \text{l mol}^{-1} \text{s}^{-1}$	$10^4 C_{\text{AgNO}_3} = 10.3 \pm 0.1$ $k_2 / \text{l mol}^{-1} \text{s}^{-1}$
10.0		$6.42 \pm 0.40$
15.0		$11.77 \pm 0.57$
20.0	$6.85 \pm 0.33$	$18.1 \pm 0.2$
25.0		$28.5 \pm 0.4$
32.1	$22.8 \pm 0.7$	
41.4	$47.9 \pm 0.2$	
50.0	$79.0 \pm 2.5$	
Apparent activation parameters		
$\log A$	$12.39 \pm 0.44$	$13.52 \pm 0.44$
$\Delta E_a / \text{kJ mol}^{-1}$	$64.7 \pm 2.6$	$68.8 \pm 2.4$
$\Delta S^\ddagger / \text{J K}^{-1} \text{mol}^{-1}$	$-16.2 \pm 8.5$	$5.9 \pm 8.4$

<sup>a</sup> Average values from two or more runs.

TABLE 6

Reaction of hydroxamoyl chloride (II) with tertiary amines in acetonitrile at 25.0 °C. Rate coefficients  $a$  for  $\gamma$ -elimination

$10^5 C_{\text{(III)}} / \text{mol l}^{-1}$	$10^4 C_{\text{base}} / \text{mol l}^{-1}$	$k_1 / \text{s}^{-1}$	$10^{-3} k_2 / \text{l mol}^{-1} \text{s}^{-1}$
<i>N</i> -Methylmorpholine ( $\text{p}K_a$ 7.38)			
4.0	19.8	4.48	2.26
8.1 $\pm$ 0.1	19.8	4.98	2.52
	39.6	9.27	2.34
	60.7	15.7	2.59
	79.2	19.0	2.40
	101.0	25.0	2.48
		Mean	2.43
<i>Bu</i> <sub>3</sub> <i>N</i> ( $\text{p}K_a$ 9.93)			
2.9	2.98	63.0	211
	5.96	124	208
4.2	8.06	167	207
8.2 $\pm$ 0.2	8.06	180	223
	12.1	245	202
		Mean	210
<i>Et</i> <sub>3</sub> <i>N</i> ( $\text{p}K_a$ 10.75)			
2.9	2.94	131	446
	5.89	268	454
3.9	7.88	358	454
8.0 $\pm$ 0.2	7.88	362	459
	8.96	411	458
	11.9	501	421
		Mean	449

<sup>a</sup>  $k_1$  values are the average from two or more runs;  $k_2 = k_1 / C_{\text{base}}$ .

runs at different  $C_{\text{AgNO}_3}$  values is given in Table 5. A systematic determination of  $k_2'$  and  $k_2''$  as functions of temperature was not performed. At 20 °C,  $k_2' = 3.91 \text{ l mol}^{-1} \text{ s}^{-1}$ , and  $k_2'' = 1.34 \times 10^4 \text{ l}^2 \text{ mol}^{-2} \text{ s}^{-1}$ .

*Reaction with Tertiary Amines.*—The dehydrohalogenation of (II) and (III) by tertiary amines, a very fast reaction, was followed by the stopped-flow technique. Triethylamine, tri-*n*-butylamine, and *N*-methylmorpholine,

TABLE 7

Reaction of hydroxamoyl bromide (III) (initial concentration  $ca. 7 \times 10^{-5} \text{ mol l}^{-1}$ ) with *N*-methylmorpholine in acetonitrile at 25.0 °C. Rate coefficients  $a$  for  $\gamma$ -elimination

$10^4 C_{\text{base}} / \text{mol l}^{-1}$	$k_1 / \text{s}^{-1}$	$10^{-3} k_2 / \text{l mol}^{-1} \text{s}^{-1}$
37.4	17.0	4.53
39.5	16.8	4.25
60.3	25.7	4.27
77.6	28.0	3.61
78.8	29.5	3.74
100.1	40.9	4.09
102.0	35.2	3.45
	Mean	3.99

<sup>a</sup> See footnote *a* of Table 6.

which have rather different basic strength and give rates which can be followed by the stopped-flow technique, were employed for chloride (II) and *N*-methylmorpholine for bromide (III). The results of experiments carried out at different initial concentrations of the hydroxamoyl halides and of amines are collected in Tables 6 and 7. The reactions, carried out with excess of amine, gave pseudo-first-order rate coefficients  $k_1$  proportional to the initial

TABLE 8

Primary kinetic isotope effect for the  $\gamma$ -elimination of hydroxamoyl chloride (II) [initial concentration  $(8.3 \pm 0.2) \times 10^{-5} \text{ mol l}^{-1}$ ] by *N*-methylmorpholine. Solvent,  $\text{CH}_3\text{CN}$ ; temperature 25.0 °C

$10^4 C_{\text{base}} / \text{mol l}^{-1}$	Substrate	$k_1 / \text{s}^{-1}$	$k_D^a / \text{s}^{-1}$	$k_H / k_D$
23.4	(II)	5.32		< 2.6
	(IIa) <sup>b</sup>	4.73	> 2.0	
50.9	(II)	12.0		< 2.3
	(IIa) <sup>b</sup>	10.8	> 5.3	
74.8	(II)	18.2		< 2.4
	(IIa) <sup>b</sup>	16.3	> 7.7	

<sup>a</sup> Calculated for the fully deuteriated substrate. <sup>b</sup> Isotopic purity (18  $\pm$  1)% D.

concentration of the amine. Second-order coefficients  $k_2$  obtained by an equation similar to (1), were satisfactorily reproducible over the range of concentrations examined.

The reaction of chloride (II) with *N*-methylmorpholine at 25 °C was chosen for a study of the primary kinetic isotope effect. Although a deuteriated hydroxamoyl chloride (IIa) could be prepared only at low isotopic purity (18%), an isotope effect was clearly appreciated, as shown by the  $k_1$  values for (II) and (IIa), the latter being less reactive than the former in three sets of experiments (Table 8). The kinetic isotope effect  $k_H / k_D$ , calculated from the initial deuterium content of (IIa), is  $ca. 2.4$ . However, since the deuterium content of the partially deuteriated substrate increases during the reaction, this value is actually an upper

limit, as indicated in Table 8. A value closer to the true kinetic isotope effect can be computed using the iteratively calculated deuterium content (24%) at half-reaction. Under this approximation,  $k_H/k_D = 1.8 \pm 0.1$ .

## DISCUSSION

Nucleophilic substitutions and eliminations in which silver salts are used as reagents have been described in several cases.<sup>12</sup> For our reactions with silver nitrate, a leaving group effect  $k_{Br}/k_{Cl}$  cannot be determined as a single value, due to the different orders of the reactions of (II) and (III). At 25 °C and  $10^{-3}$  M for  $C^\circ_{AgNO_3}$ , using the experimental  $k_2$  rate coefficient for (III) (Table 5) and the second-order coefficient for (II) calculated by the Arrhenius equation (Table 3), the relative rate of reaction of the bromo- and chloro-compounds,  $k_{Br}/k_{Cl}$ , was evaluated as  $2 \times 10^3$ .

Although lower values are obtained when referring to lower  $C^\circ_{AgNO_3}$  values, the ratio remains sufficiently large in the whole range to prove that the rate-determining step includes C-Hal bond cleavage. This is in agreement with the widely accepted concept of an electrophilic attack by silver ion on halogens in elimination and substitution reactions.<sup>12</sup> The reaction order with respect to the silver salt, which is higher than one in the case of hydroxamoyl bromide (III), indicates in this case the intervention of a second molecule of  $AgNO_3$ . Similar observations have been often reported for reactions with silver salts.<sup>12,13</sup> Among the possible explanations, the one based on salt association seems unacceptable, because of the different behaviour of chloride (II) and bromide (III). Anionic assistance by the second molecule of  $AgNO_3$  is more likely.

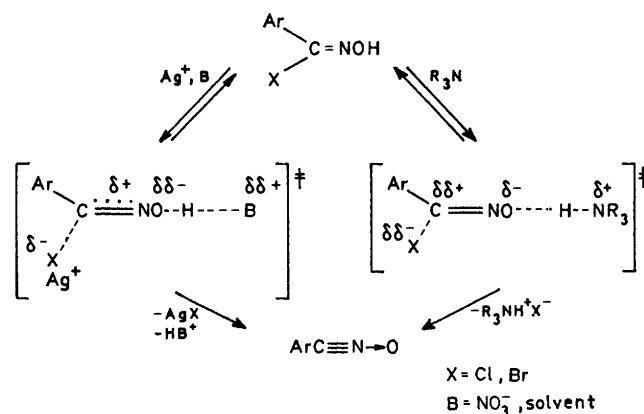
The activation entropy for the  $\gamma$ -elimination of (II) has a value close to zero (Table 3), this indicating a degree of order in the transition state comparable to that of the initial state of the system. On the contrary, the activation entropy for the 1,3-cycloaddition of (I) to MeCN is relatively large and negative (Table 1) as is normally found for these reactions.<sup>14</sup>

Tertiary amines react with chloride (II)  $10^5$ – $10^7$  times faster than  $AgNO_3$  (at 25 °C; Tables 3 and 6). The same comparison cannot be made exactly for bromide (III) owing to the different reaction orders, although this substrate too appears much more reactive with amines than with silver ion. This accounts for the systematic use of tertiary amines when this reaction is employed for synthetic purposes.

The effect of the leaving halogen,  $k_{Br}/k_{Cl}$ , on the reaction with *N*-methylmorpholine at 25 °C is *ca.* 1.6 (Tables 6 and 7). Such a low value proves that the C-Hal bond breaking is scarcely advanced in the rate-controlling step. On the contrary, the cleavage of the O-H bond appears to be an important factor on the basis of two experimental results. First, a substantial

deuterium kinetic isotope effect ( $k_H/k_D$  *ca.* 2) has been measured. Besides, the reactivity of chloride (II) with the three different amines (Table 6) follows the trend of their basicities in water. If a Brønsted correlation is tentatively applied, using  $pK_a$  ( $H_2O$ ) values,  $\beta$  0.7 is obtained. The 'Brønsted  $\beta$ ' increases up to *ca.* 1.0 when the  $pK_a(H_2O)$  of *N*-methylmorpholine is corrected by one unit for the extra basicity shown by amines containing heteroatoms (with respect to monofunctional amines) in acetonitrile.<sup>15</sup>

The results may be interpreted in terms of two concerted mechanisms, markedly different from one another for the degree of breaking of the various bonds at the transition state (Scheme 1).



SCHEME 1

The silver nitrate reaction is envisaged to occur by attack of  $Ag^+$  on halogen, with advanced C-Hal cleavage (large  $k_{Br}/k_{Cl}$ ) accompanied by a small amount of O-H bond cleavage, promoted either by solvent [in the case of chloride (II)] or by both solvent and nitrate anion [in the case of bromide (III)] in agreement with rate equations (1) and (2). Consequently, the transition state is indicated as having a prevailing cationic character.

On the other hand in the reactions with tertiary amines the C-Hal bond appears little extended in the transition state (very small  $k_{Br}/k_{Cl}$ ), while proton transfer to the amine is highly advanced, giving anionic character to this transition state. In fact with the usual approximations<sup>16</sup> a maximum value of the hydrogen isotope effect is expected when hydrogen is half-transferred in the transition state; since in our system the maximum value would fall around 10, the experimental value of *ca.* 2 indicates a proton much more than half-transferred, due account being given to the 'Brønsted  $\beta$ ' value.

It seems difficult, on the basis of the available evidence, to distinguish the proposed mechanisms from multistep processes involving a cationic intermediate in the

<sup>14</sup> P. Beltrame, in 'Comprehensive Chemical Kinetics', eds. C. H. Bamford and C. F. H. Tipper, Elsevier, Amsterdam, 1973, vol. 9, ch. 2.

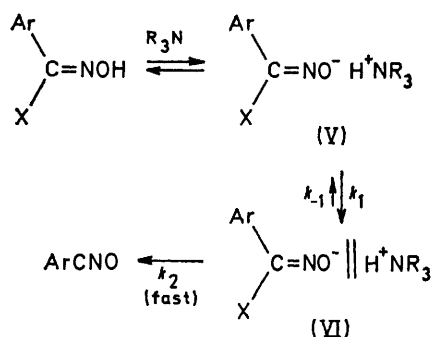
<sup>15</sup> H. K. Hall, jun., *J. Phys. Chem.*, 1956, **60**, 63.

<sup>16</sup> P. J. Smith, in 'Isotopes in Organic Chemistry', eds. E. Buncl and C. C. Lee, Elsevier, Amsterdam, 1976, vol. 2, ch. 6.

<sup>12</sup> Y. Pocker and D. N. Kevill, *J. Amer. Chem. Soc.*, 1965, **87**, 4760, 4771, 4778; Y. Pocker and W. H. Wong, *ibid.*, 1975, **97**, 7097, 7105, and references cited therein.

<sup>13</sup> R. D. Bach and C. L. Willis, *J. Amer. Chem. Soc.*, 1975, **97**, 3844.

AgNO<sub>3</sub> reaction, and anionic intermediates in the case of amines. However for the latter reaction, an indication favouring an ionisation mechanism is given by the large 'Brønsted β' value. Therefore it is worth discussing briefly the E1cb mechanisms for this reaction. A 'pre-equilibrium' process<sup>17</sup> is unlikely, since for the corresponding deuterium isotope effect (which would be an equilibrium isotope effect<sup>16</sup>), a value as low as 1.3 can be calculated from the stretching frequencies<sup>18</sup> of NO-H (*ca.* 3 300 cm<sup>-1</sup>) and <sup>+</sup>N-H (*ca.* 2 950 cm<sup>-1</sup>) bonds. An 'irreversible' E1cb mechanism requires that the intermediate anion ArC(Cl)=NO<sup>-</sup> loses Cl<sup>-</sup> much faster than it is reprotonated. While this behaviour would be unlikely for an intimate ion pair (V) (Scheme 2), it seems conceivable for a solvent-separated ion pair (VI) in a solvent as polar as acetonitrile.



SCHEME 2

No reference has been made to the configuration of hydroxamoyl halides, a problem which is still open.<sup>19</sup> The mechanisms suggested in the Schemes should be little affected by *E*- or *Z*-configurations for the reactant, provided the preferred conformation about N-O bond is *s-trans* as it has been concluded both for the solid state and for solutions.<sup>19</sup>

#### EXPERIMENTAL

Acetonitrile for reactions with amines was carefully purified by distillation through a 2 m Vigreux column over phosphorus pentoxide (twice) and anhydrous K<sub>2</sub>CO<sub>3</sub> and stored in a dark-walled bottle under nitrogen. The required amounts for a set of kinetic experiments were distilled just prior to use. Acetonitrile for reactions with silver nitrate was a Merck Uvasol solvent. Tertiary amines were distilled twice over KOH or CaH<sub>2</sub>. Hexamethyldisiloxane was used as reference for n.m.r. spectra.

3,5-Dichloro-2,4,6-trimethylbenzonitrile *N*-oxide (I) was prepared as described,<sup>20</sup> λ<sub>max</sub>(MeCN) 270 nm (ε = 13 850).

3,5-Dichloro-2,4,6-trimethylbenzohydroxamoyl chloride (II) was prepared by chlorination of mesitaldehyde oxime as previously reported<sup>20</sup> and recrystallized from benzene-light petroleum, m.p. 130–131 °C. The sample employed for reactions with AgNO<sub>3</sub> was further subjected to column

chromatography on silica gel, eluant acetonitrile, m.p. 125–126 °C (Found: C, 45.0; H, 3.7; N, 5.2; Cl, 39.7. C<sub>10</sub>H<sub>10</sub>Cl<sub>2</sub>NO requires C, 45.0; H, 3.75; N, 5.25; Cl, 40.0%) (Cl titrated by AgNO<sub>3</sub>: found 13.3; required 13.3%), δ(CCl<sub>4</sub>) 2.26 (6 H, s), 2.46 (3 H, s), and 7.76 (1 H, s).

[<sup>2</sup>H]3,5-Dichloro-2,4,6-trimethylbenzohydroxamoyl Chloride (IIa).—A solution of mesitaldehyde oxime (4 g) in 2*N*-sodium ethoxide (13 ml) in ethanol was treated with anhydrous ethyl ether at room temperature.<sup>21</sup> The sodium salt of the oxime was filtered, washed thoroughly with ethyl ether, dried at 10<sup>-2</sup> Torr (room temperature), and then treated with D<sub>2</sub>O (12 ml). The deuteriated oxime (ν<sub>max</sub> 2 650 cm<sup>-1</sup>), contaminated with the isotopically normal compound (ν<sub>max</sub> 3 580 cm<sup>-1</sup>) was filtered, washed with D<sub>2</sub>O, and dried *in vacuo*. Chlorination of this material under the usual conditions gave a mixture of (IIa) and (II) whose deuterium content was determined by the difference of the intensity of the 3 600 cm<sup>-1</sup> band from the calculated value obtained by a calibration curve from pure samples of (II).

3,5-Dichloro-2,4,6-trimethylbenzohydroxamoyl bromide (III) was prepared by bubbling dry HBr for 1 h into a solution of nitrile oxide (I) (1.5 g) in anhydrous methylene dichloride (50 ml) at room temperature. The solvent was evaporated under reduced pressure and the product was recrystallized from benzene-light petroleum, m.p. 123–125 °C, ν<sub>max</sub>(CCl<sub>4</sub>-CS<sub>2</sub>) 1 640, 3 320br, and 3 570 cm<sup>-1</sup> (Found: C, 38.0; H, 3.4; N, 4.5. C<sub>10</sub>H<sub>10</sub>BrCl<sub>2</sub>NO requires C, 38.6; H, 3.2; N, 4.5%) (Br titrated by AgNO<sub>3</sub>: found 26.3; required 25.7%).

3-(3,5-Dichloro-2,4,6-trimethylphenyl)-5-methyl-1,2,4-oxadiazole (IV).—Nitrile oxide (I) (2 g) was refluxed for a few hours in MeCN (50 ml). The solvent was evaporated under reduced pressure and crystals were obtained, m.p. 134 °C (from ethanol) (Found: C, 53.1; H, 4.8; N, 10.0. C<sub>12</sub>H<sub>12</sub>Cl<sub>2</sub>N<sub>2</sub>O requires C, 53.1; H, 4.5; N, 10.3%); λ<sub>max</sub>(MeCN) 287 nm (ε 604); δ(CCl<sub>4</sub>) 2.11 (6 H, s), 2.53 (3 H, s), and 2.61 (3 H, s); *m/e* 270–272–274 (*M*<sup>+</sup>), 255–257–259 (*M* - CH<sub>3</sub>), 228–230–232 (*M* - CH<sub>2</sub>CO), and 227–229–231 (*M* - CH<sub>3</sub>CO).

*Kinetics*.—The relatively slow reactions with AgNO<sub>3</sub> and the 1,3-cycloaddition to the solvent were carried out in the thermostatted cells of a u.v. spectrophotometer and followed by variation of the absorbance at 270 nm. At this wavelength, the molar extinction coefficients were ε<sub>(I)</sub> 13 850, ε<sub>(II)</sub> 440, ε<sub>(III)</sub> 800, ε<sub>(IV)</sub> 410 (ε<sub>(II)</sub> and ε<sub>(III)</sub> were determined as quickly as possible using a 'time-drive' attachment at 270 nm, instead of fully scanning the u.v. spectrum).

For the 1,3-cycloaddition of (I) to the solvent, pseudo-first-order rate coefficients *k<sub>c</sub>* were obtained as the slopes of the linear plots of ln(*A<sub>t</sub>* - *A<sub>∞</sub>*) versus time (*A<sub>t</sub>* and *A<sub>∞</sub>* are absorbance at time *t* and infinity, respectively).

For the reactions of (II) with AgNO<sub>3</sub>, the initial concentration of (II) from weighing was corrected since high *A<sub>0</sub>* values evidenced the relatively fast formation of some nitrile oxide (I) after the dissolution of (II) in acetonitrile. *C*<sub>I</sub><sup>o</sup> and *C*<sub>II</sub><sup>o</sup> values were obtained using the known extinction coefficients. Experimental absorbances were compared with calculated values [equation (3)]. *k<sub>c</sub>* Values were known from previous measurements. Step-by-step optimization of *k<sub>1</sub>* was performed by minimising the sum

<sup>20</sup> P. Beltrame, C. Veglio, and M. Simonetta, *J. Chem. Soc. (B)*, 1967, 867.

<sup>21</sup> E. Borello and M. Colombo, *Ricerca sci.*, 1955, 25, 2899.

<sup>17</sup> D. J. McLennan, *Quart. Rev.*, 1967, 21, 490.

<sup>18</sup> 'DMS Working Atlas of Infrared Spectroscopy' Butterworths, London, 1972.

<sup>19</sup> A. Battaglia, A. Dondoni, and O. Exner, *J.C.S. Perkin II*, 1972, 1911; J. P. Declercq, G. Germain, and M. Van Meerssche, *Acta Cryst.*, 1975, B31, 2894; O. Exner, V. Jehlicka, G. Barbaro, and A. Dondoni, *Coll. Czech. Chem. Comm.*, 1977, 42, 833.

of squares of residuals defined by (4) with summation over all experimental determinations. Calculated values of

$$(A_t)_{\text{calc}} = \epsilon_{\text{(IV)}}(C_{\text{(I)}}^{\circ} + C_{\text{(II)}}^{\circ}) + C_{\text{(II)}}^{\circ} \left( \epsilon_{\text{(II)}} + \epsilon_{\text{(I)}} \frac{k_1}{k_c - k_1} - \epsilon_{\text{(IV)}} \frac{k_c}{k_c - k_1} \right) e^{-k_1 t} + (\epsilon_{\text{(I)}} - \epsilon_{\text{(IV)}}) \left( C_{\text{(I)}}^{\circ} - C_{\text{(II)}}^{\circ} \frac{k_1}{k_c - k_1} \right) e^{-k_c t} \quad (3)$$

$$F = \sum_i [(A_t)_{\text{expt},i} - (A_t)_{\text{calc},i}]^2 \quad (4)$$

$C_{\text{(I)}}$  versus time were obtained by equation (5). Experi-

$$(C_{\text{(I)}})_{\text{calc}} = C_{\text{(I)}}^{\circ} e^{-k_c t} + \frac{C_{\text{(II)}}^{\circ} k_1}{k_c - k_1} (e^{-k_1 t} - e^{-k_c t}) \quad (5)$$

mental values of  $C_{\text{(I)}}$  were deduced from experimental  $A_t$  values through relation (6) (valid only for  $\epsilon_{\text{(I)}} \approx \epsilon_{\text{(IV)}}$ ).

$$(C_{\text{(I)}})_{\text{expt}} \approx [A_t - \epsilon_{\text{(II)}}(C_{\text{(II)}}^{\circ} + C_{\text{(I)}}^{\circ})] / (\epsilon_{\text{(I)}} - \epsilon_{\text{(II)}}) \quad (6)$$

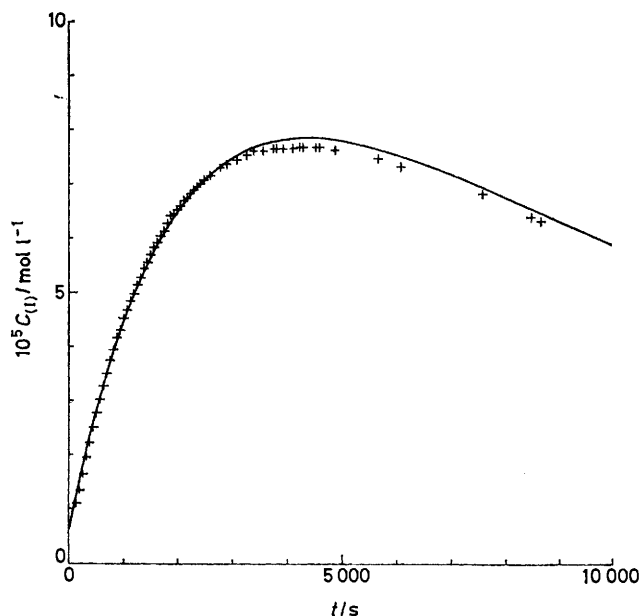
Curves of  $C_{\text{(I)}}$  versus time, obtained from equation (5) using the optimised  $k_1$  values, were in good agreement with the experimental points from equation (6) for all runs. An example is presented in the Figure. The reaction of (III) with  $\text{AgNO}_3$  was interpreted by the second-order rate equation (7) where  $C_{\text{(III)}}^{\circ}$  was corrected for the partial fast

$$k_2 t = \frac{1}{C_{\text{AgNO}_3}^{\circ} - C_{\text{(III)}}^{\circ}} \times \ln \frac{(A_{\infty} - A_0) - \frac{C_{\text{(III)}}^{\circ}}{C_{\text{AgNO}_3}^{\circ}} (A_t - A_0)}{A_{\infty} - A_t} \quad (7)$$

decomposition to (1), as described above for the chloride, and absorbance values  $A_0$ ,  $A_t$ , and  $A_{\infty}$  were measured.

The rate constants of the very fast reactions of (II) and (III) with amines were measured using a Durrum-G type stopped flow apparatus. Stock solutions of the hydrox-amoyl halide and amine in acetonitrile were prepared just prior to their use and thermostatted in a separated bath kept at proper temperature. After the rapid mixing of the solutions in the stopped flow apparatus, the reaction progress was followed by observing on the oscilloscope trace the change of transmittance or absorbance at 270 nm.

Control experiments showed a close superimposition between the  $t_{\infty}$  spectrum and the u.v. spectrum of (I). For a selected concentration of the reactants, several runs were carried out, varying the selected range of time and absorbance scales of the oscilloscope. Pseudo-first-order rate



Calculated curve and experimental points of  $C_{\text{(I)}}$  against time for a typical reaction of (II) with  $\text{AgNO}_3$  ( $C_{\text{(II)}}^{\circ}$ ,  $10.38 \times 10^{-5}$ ;  $C_{\text{AgNO}_3}^{\circ}$ ,  $9.78 \times 10^{-4}$  mol  $\text{l}^{-1}$ ; temperature  $60^{\circ}\text{C}$ )

constants were obtained as the slopes of the linear plots of  $\ln(A_{\infty} - A_t)$  against time (actually the readings on the oscilloscope were quantities proportional to the absorbance values). The experimental error was *ca.*  $\pm 5\%$ .

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