The Kinetics and Mechanism of Aminolysis of Isothiocyanates

Derek P. N. Satchell and Rosemary S. Satchell
King's College London, Strand, London, WC2R 2LS

A study of the kinetics of aminolysis of *p*-nitrophenyl isothiocyanate by n-butylamine, benzylamine, dibenzylamine, *p*-anisidine, *m*-toluidine, *N*-methylaniline, and *p*-chloroaniline in diethyl ether and iso-octane as solvents has shown that many of the reactions involve kinetic terms that are second order in amine. The detailed behaviour reveals that aminolysis occurs *via* an intermediate of amine–isothiocyanate (1:1) stoicheiometry which undergoes subsequent prototropic rearrangement catalysed by a second amine molecule. Added carboxylic acids inhibit aminolysis by strong bases by their removal as inactive amine–acid (1:2) complexes, but with bases with which they form only a negligible amount of complex, acids can catalyse the aminolysis, probably by their effect on the prototropic rearrangement. Added thioureas (products) have negligible catalytic effects in the presence of an excess of amine. Our results and conclusions are similar to findings for the aminolysis of isocyanates in water.

Aromatic and aliphatic isothiocyanates react with a variety of primary and secondary aliphatic or aromatic amines in aqueous and non-aqueous media to form thioureas[1] [*e.g.* equation (1)]. The mechanisms of these reactions have been

$$\mathrm{ArNCS} + \mathrm{RNH_2} \longrightarrow \mathrm{ArNHC(S)NHR} \qquad (1)$$

little studied. For aqueous solutions, kinetic measurements under pseudo-first-order conditions (excess of amine varied over a narrow range) suggest that the additions are first order in both isothiocyanate and amine,[1,2] but only one aromatic amine (aniline) has been examined.[2] For non-aqueous media the reactions are found to be faster in *O*-containing and purely donor solvents than in other types.[2,4] There are reports of inhibition by carboxylic acids,[5] and of possible catalysis by triethylamine.[6] Kinetic studies in non-aqueous solvents have normally used second-order conditions (comparable amounts of isothiocyanate and amine) and the order in amine is not well-established. Examples of first- and second-order dependence on amine have been claimed. Both synchronous (cyclic) and stepwise additions of the amine to the N=C double bond have been proposed. Little of the work with non-aqueous solvents has involved spectroscopic methods of following the reactions, and the additions of aliphatic amines to isothiocyanates have proved too fast for conventional study by sampling techniques.[3,6]

We have now studied the reaction of some aliphatic and aromatic amines with *p*-nitrophenyl isothiocyanate in diethyl ether and iso-octane solutions. We have investigated the reaction order in amine over a wide concentration range and studied the effects of temperature and catalysts.

Experimental

Materials.—*p*-Nitrophenyl isothiocyanate was the Aldrich product. The primary and secondary amines were either fractionally distilled or recrystallised samples. The carboxylic acids were purified as before.[7] Diethyl ether was dried by distillation from sodium. This solvent and iso-octane (Aldrich) were stored over molecular sieves.

Kinetic Methods.—Reactions were studied under pseudo-first-order conditions. Normally an excess of amine was used, and the reaction monitored by observation of the formation of *S*-urea at *ca.* 330 nm with a Perkin-Elmer Lambda 5 spectrophotometer. For the fastest runs a modified Durrum–Gibson stopped-flow instrument was used.[8] Good isosbestic points for loss of isothiocyanate and formation of *S*-urea were obtained. Our concentration and other conditions are in the Tables and Figures. The main temperature used was 13.8 °C but a number of systems were studied at two other temperatures. Normally reaction mixtures were homogeneous throughout a run, but at the lowest amine concentrations of the least reactive amines (when $t_{\frac{1}{2}} > 30$ min) partial precipitation of the *S*-urea sometimes occurred towards the end of a run. (Precipitation problems have been noted in other studies.[3,6]) However, even in these instances, reaction could be followed for at least two half-lives, and k_{obs}, the observed first-order rate constant, was always reproducible to within ±10%. Most reactions were followed to completion. The aromatic amines were too insoluble to permit their study in iso-octane.

Reaction Stoicheiometry.—Preparative-scale reactions in either ether or iso-octane led to high yields of the expected unsymmetrical *S*-ureas, their structures being confirmed by NMR spectroscopy. Artificial product mixtures normally led to UV spectra in good agreement with those obtained at the completion of kinetic runs and suggested that the reactions studied proceed in effectively quantitative yields. There is little evidence of reversibility under our conditions. The pK_a of dibenzylamine was determined by the potentiometric method.

Results and Discussion

Spontaneous Reactions.—Typical plots of k_{obs} against [amine] are shown in Figures 1–4. We find three types of plot: (*i*) a simple first-order dependence on [amine] over a wide concentration range, found for the most basic amines in ether solution (*e.g.* butylamine and benzylamine, Figure 1); (*ii*) plots showing a greater than first-order dependence on [amine] at low amine concentrations, which approaches or becomes a first-order dependence at high amine concentrations, and found for reactive amines in iso-octane, and for amines of intermediate reactivity in ether (*e.g.* *p*-anisidine, *m*-toluidine,

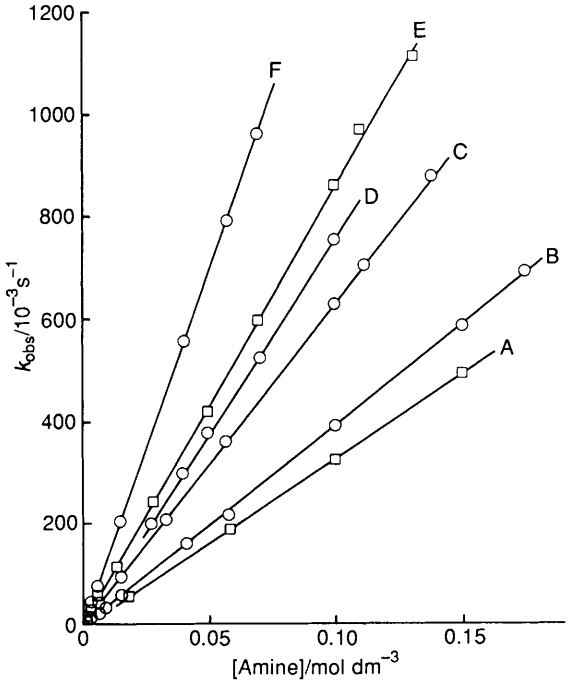


Figure 1. Type (*i*) plots. $[\text{isothiocyanate}]_{\text{initial}} = 3\text{–}5 \times 10^{-5}$ mol dm^{-3} in all Figures; for k_{obs} see the text: A, B, C, benzylamine at 9.8, 14.0, and 24.7 °C; D, E, F, n-butylamine at 9.0, 13.8, and 24.8 °C; solvent, diethyl ether.

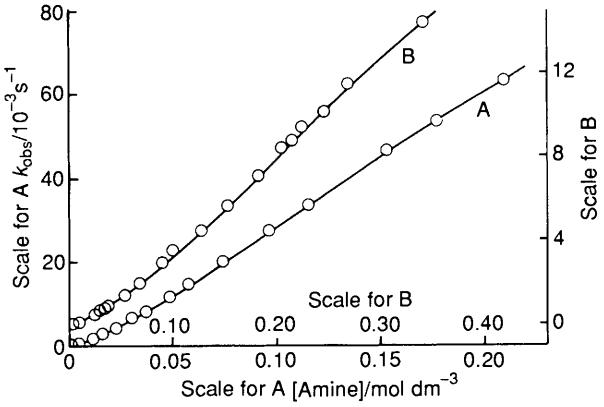


Figure 2. Type (*ii*) plots for diethyl ether.[18] A, *p*-anisidine at 13.8 °C; B, *m*-toluidine at 13.8 °C.

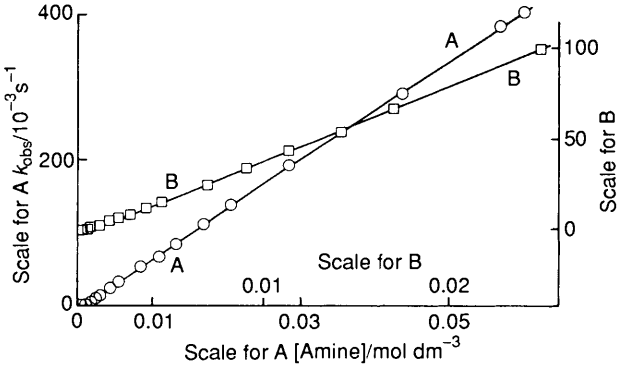


Figure 3. Type (*ii*) plots for iso-octane. A, n-butylamine at 13.8 °C; B, dibenzylamine at 13.8 °C.

butylamine, and dibenzylamine; Figures 2–3); (*iii*) plots showing a greater than first-order dependence over the entire

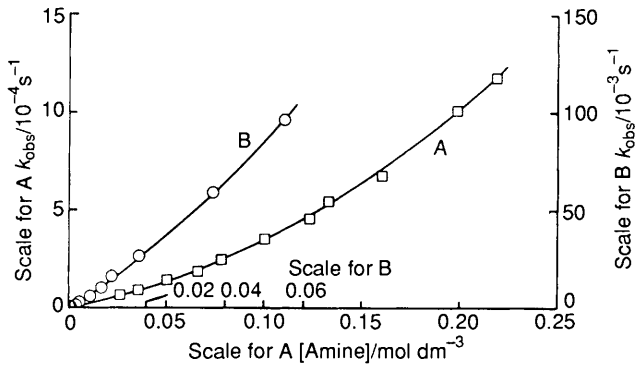
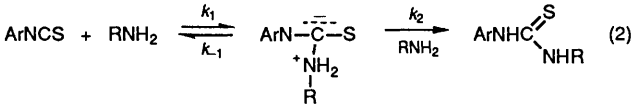


Figure 4. Type (*iii*) plots. A, *p*-chloroaniline in diethyl ether at 13.8 °C; B, benzylamine in iso-octane at 14.0 °C.

amine concentration range used, found for the least reactive amines in ether, and for benzylamine in iso-octane (*e.g.* *p*-chloroaniline, Figure 4). For several systems the amine concentration was varied by a factor of *ca.* 200.

The most significant type of plot is type (*ii*). The curvature is evident at very low amine concentrations (where it is unlikely to represent a medium effect of the amine), and the first-order region (at high amine concentrations) if extrapolated to low amine concentrations, passes (within experimental error) through the origin. Bearing in mind also the simple continuous first-order behaviour shown by some amines [type (*i*) plots], and the continuous curvature shown by others [type (*iii*) plots], it is difficult to explain type (*ii*) plots mechanistically, save by a stepwise mechanism of addition in which a second step, catalysed by the amine, becomes rapid at relatively high amine concentrations [*e.g.*, equation (2)]. With the assumption of a



(2)

low concentration of zwitterionic intermediate and an excess of amine, k_{obs} for equation (2) is given by equation (3).

$$k_{obs} = k_1 k_2 [\text{amine}]^2/(k_{-1} + k_2 [\text{amine}]) \qquad (3)$$

This equation predicts a second-order dependence on [amine] if k_2 [amine] $\ll k_{-1}$, and a first-order dependence if k_2 [amine] $\gg k_{-1}$. Thus, this mechanism allows the second step to be rapid compared with the first step at all concentrations for very reactive amines, and at sufficiently high concentrations for somewhat less reactive amines, but to remain always rate-determining for unreactive amines, in agreement with the kinetic pattern we have observed.

Another mechanism[9] that could, in appropriate circumstances, lead to somewhat similar observed behaviour is that in which a (H-bonded) polymeric form of the addend HX is the reactive species [equations (4) and (5)] and for which polymerisation becomes stoicheiometric at high values of $[\text{HX}]_{\text{stoich}}$. For this scheme $k_{obs} = k[(\text{HX})_x]$. At low values of $[\text{HX}]_{\text{stoich}}$, $[(\text{HX})_x]$ will be proportional to $[\text{HX}]^x_{\text{stoich}}$, but if

$$x\,\text{HX} \rightleftharpoons (\text{HX})_x \qquad \text{Fast} \quad (4)$$

$$\text{ArNCS} + (\text{HX})_x \xrightarrow{k}$$

$$\text{ArHNC(S)X} + (x-1)\,\text{HX} \quad \text{Slow} \quad (5)$$

equation (4) lies to the right at high values of $[\text{HX}]_{\text{stoich}}$, $[(\text{HX})_x]$ would then become proportional to $[\text{HX}]_{\text{stoich}}$. This

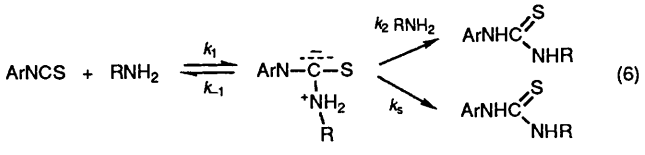
Table 1. Derived rate constants and activation parameters.

Amine	Solvent	Parameter T_p/°C	Derived rate constant[a]		Enthalpy	Entropy
Butylamine (pK_a 10.7)	Diethyl ether	9.0	k_1	7.63	$\Delta H_1^{\ddagger}$ 26 ± 2	$\Delta S_1^{\ddagger}$ −137 ± 5
		13.8		8.66		
		24.8		13.9		
	Iso-octane	13.8	k_1	3.77		
			10^{-3} (k_2/k_{-1})	2.27		
Benzylamine (pK_a 9.4)	Diethyl ether	9.8	k_1	3.32	$\Delta H_1^{\ddagger}$ 29 ± 2	$\Delta S_1^{\ddagger}$ −132 ± 5
		14.0		3.94		
		24.7		6.37		
	Iso-octane	14.0	k_1	1.63		
			10^{-2} (k_2/k_{-1})	7.96		
Dibenzylamine (pK_a 7.6)	Diethyl ether	13.8	k_1	4.17		
			10^{-2} (k_2/k_{-1})	10.3		
	Iso-octane	13.8	k_1	6.93		
			10^{-2} (k_2/k_{-1})	7.39		
p-Anisidine (pK_a 5.3)	Diethyl ether	9.8	$10\,k_1$	2.69	$\Delta H_1^{\ddagger}$ 37 ± 3	$\Delta S_1^{\ddagger}$ −125 ± 6
		13.8		3.12		
		24.6		4.78		
		9.8	k_2/k_{-1}	148		
		13.8		120	$\Delta H_2^{\ddagger} - \Delta H_{-1}^{\ddagger}$ −24 ± 2	$\Delta S_2^{\ddagger} - \Delta S_{-1}^{\ddagger}$ −45 ± 4
		24.6		85.8		
m-Toluidine (pK_a 4.7)	Diethyl ether	9.0	$10^2\,k_1$	3.50	$\Delta H_1^{\ddagger}$ 39 ± 3	$\Delta S_1^{\ddagger}$ −136 ± 5
		17.8		4.75		
		24.7		8.40		
		9.0	k_2/k_{-1}	35.5		
		17.8		27.8	$\Delta H_2^{\ddagger} - \Delta H_{-1}^{\ddagger}$ −32 ± 3	$\Delta S_2^{\ddagger} - \Delta S_{-1}^{\ddagger}$ −69 ± 4
		24.7		18.0		
N-Methylaniline (pK_a 4.8)	Diethyl ether	13.8	10^2 (k_1k_2/k_{-1})	4.40		
			10^2 (k_1k_s/k_{-1})	*ca.* 0.1		
p-Chloroaniline (pK_a 4.0)	Diethyl ether	13.8	10^2 (k_1k_2/k_{-1})	1.54		
			10^2 (k_1k_s/k_{-1})	*ca.* 0.2		

[a] For k_1 (dm^3 mol^{-1} s^{-1}), k_{-1} (s^{-1}), k_2 (dm^3 mol^{-1} s^{-1}), and k_s (s^{-1}), see the text. Units for $\Delta H^{\ddagger}$ kJ mol^{-1}; $\Delta S^{\ddagger}$ J K^{-1} mol^{-1}.

type of mechanism is possible for the addition of alcohols to isocyanates and ketenes in hydrocarbons and similar non-aqueous solvents.[9] It has not been tested for alcoholysis of isothiocyanates, but is not a viable explanation of our present results for two reasons: (*i*) amines are not sufficiently associated[10] to account for the reductions in reaction order observed, and (*ii*) for mechanism (4) and (5) the first-order region of the k_{obs} against $[HX]_{stoich}$ plot does not normally extrapolate through the origin.[9]

The second step of equation (2) is a catalysed prototropic rearrangement. Hydrocarbon solvents are unlikely to be able to participate in such a step, but basic or acidic solvents could in principal do so. Diethyl ether might sensibly be expected to contribute something to this second step and equation (2) can be elaborated to include this possibility, as in equation (6).



(6)

The equation for k_{obs} is now (7). Any contribution from k_s

$$k_{obs} = k_1[\text{amine}]\,(k_2[\text{amine}] + k_s)/(k_{-1} + k_2[\text{amine}] + k_s) \quad (7)$$

would be expected to be most important for the weakest amines.

If, for the strong amines butylamine and benzylamine in ether, we assume that $(k_1[\text{amine}] + k_s) \gg k_{-1}$ for all values of [amine], then equation (7) reduces to $k_{obs} = k_1[\text{amine}]$, and k_1 can be obtained from the slopes of plots like Figure 1. Our values of k_1 at three temperatures are shown in Table 1, with the corresponding activation parameters.

If, for the weakest amines, we assume that $k_{-1} \gg (k_2[\text{amine}] + k_s)$ then equation (7) reduces to (8). For these conditions plots of k_{obs}/[amine] against [amine] will be

$$k_{obs} = \frac{k_1 k_2}{k_{-1}}[\text{amine}]^2 + \frac{k_1 k_s}{k_{-1}}[\text{amine}] \quad (8)$$

rectilinear, and lead to values of k_1k_2/k_{-1} and k_1k_s/k_{-1}. We found that results from our two least reactive amines gave plots of this type (Figure 5). Our derived parameters are in Table 1: they suggest that $k_s < k_2$ [amine] at most values of [amine] for these bases.

For amines of intermediate reactivity in ether, and for the amines studied in iso-octane, the assumption that $k_s \ll k_2$ [amine] leads to a satisfactory fit with experiment. Under such conditions equation (7) reduces to (3), which can be written as equation (9). Plots of [amine]/k_{obs} against [amine]

$$\frac{[\text{amine}]^2}{k_{obs}} = \frac{k_{-1}}{k_1 k_2} + \frac{[\text{amine}]}{k_1} \quad (9)$$

(or of [amine]/k_{obs} against 1/[amine]) are rectilinear (*e.g.*

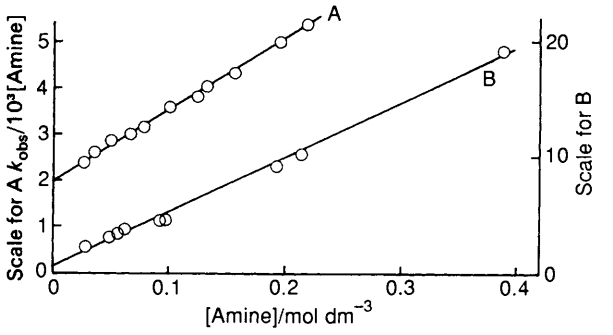


Figure 5. Plots of equation (8). A, *p*-chloroaniline at 13.8 °C; B, *N*-methylaniline at 13.8 °C; solvent diethyl ether.

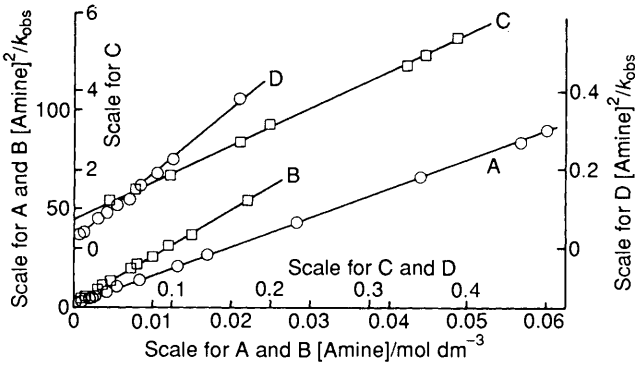


Figure 6. Plots of equation (9). A, Dibenzylamine in iso-octane at 13.8 °C; B, dibenzylamine in diethyl ether at 13.8 °C; C, *m*-toluidine in diethyl ether at 24.7 °C; D, *p*-anisidine in diethyl ether at 24.6 °C.

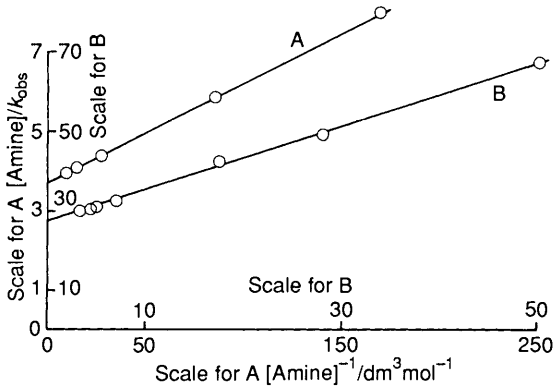


Figure 7. Plots of equation (9). A, *p*-Anisidine in diethyl ether at 9.8 °C; B, *p*-toluidine in diethyl ether at 9.8 °C.

Figures 6 and 7) and provide values of k_1, and k_2/k_{-1}. Our results for dibenzylamine, *p*-anisidine and *m*-toluidine in ether, and for butylamine, benzylamine, and dibenzylamine in iso-octane, are shown in Table 1, together with some activation parameters. For these systems, we mostly find that plots of k_{obs} against amine are of type (*ii*) (*e.g.* Figures 2 and 3). The continuous lines in Figures 2 and 3 were obtained by use of the parameters in Table 1. For systems in which k_s contributes significantly, plots of equation (9), are curved, especially at low amine concentrations.

The two-step mechanism of equation (2) has been proposed before[1,2,4] for aminolysis of isothiocyanates, mostly by analogy with findings and suggestions for the corresponding reaction of isocyanates, and this type of mechanism was favoured by Baker[11] in his early work on isocyanates. For isothiocyanates our results are the first to demand such a mechanism:[1] the one previous report[3] of second-order terms in [amine] was plausibly interpreted as reaction *via* amine dimers in a cyclic transition state. This interpretation now seems unlikely.

For our primary amines in ether at 13.8 and 25.0 °C plots (not shown) of log k_1 against pK_a have slopes of *ca.* 0.32, although the (two) points for the aromatic amines taken alone suggest a slope closer to unity. The point for dibenzylamine lines significantly above the line. For aqueous solutions, where all but one of the measurements refer to aliphatic amines, and only simple first-order behaviour was found, log k_1 *vs.* pK_a slopes of 0.23 and 0.28 were obtained for various amino acids with phenyl isothiocyanate[1] and for primary amines with ethyl isothiocyanate,[2] respectively. A larger slope (*ca.* 1.0) is reported for aromatic amines with phenyl isothiocyanate in benzene,[12] but the rate measurements were probably based on the assumption that the aminolyses were first order in amine concentration, whereas our results suggest that most amines (and especially weakly basic aromatic amines) would exhibit significant second-order kinetic terms in many circumstances in non-hydroxylic solvents. Values of k_1 from previous investigations using non-hydroxylic solvents, and second-order conditions, with the built-in assumption of a first-order dependence on [amine], must be treated with caution.

For butylamine and benzylamine our values for k_1 in iso-octane are more than twofold smaller than in ether (Table 1). Even larger effects in the same direction have been found previously.[4] Dibenzylamine, however, is significantly more reactive in the first (k_1) step in iso-octane than in ether, and therefore even more out of line with the behaviour of the other aliphatic bases than in ether (see above). Abnormally great reactivity (on a pK_a basis) has been found before for aliphatic secondary amines in similar reactions in aqueous solution.[2] The values of k_1 for dibenzylamine at least suggest that steric hinderance is of minor importance in the process reflected by k_1. This conclusion is in line with earlier results for aqueous solutions.[1] However, a similar comparison of the ratio k_2/k_{-1} for dibenzylamine, and especially of the ratio k_1k_2/k_{-1} for *N*-methylaniline (0.044) with that (1.32) for *m*-toluidine (amines of similar pK_a, Table 1) suggests that k_2/k_{-1} is subject to steric requirements. Since k_{-1} reflects the reverse of the apparently sterically unaffected addition, it is presumably k_2 that is sterically sensitive. This seems possible since whatever the exact mechanism of the second step the transition state seems likely to be more crowded than that of the first step.

Aside from the possible steric effects, the values for k_2/k_{-1} (or k_1 k_2/k_{-1}) for different amines vary in a way compatible with qualitative expectations: decreasing base strength might reasonably be expected to decrease k_2 as well as k_1 and to increase k_{-1} so that these rate constant ratios should fall for the weakly basic amines, as is found (Table 1). *p*-Nitrophenyl isothiocyanate is the most reactive isothiocyanate studied so far. For k_1 our $\Delta H^{\ddagger}$ values are small, but increase with decreasing base strength; the $\Delta S^{\ddagger}$ values are compatible with a bimolecular process, and remain largely constant. This rather regular behaviour shows that changes in k_1 mainly reflect changes in $\Delta H^{\ddagger}$, and we find that $\Delta H_1^{\ddagger} = 48.5\text{–}2.1$ pK_a for the primary amines. The effects of base strength on $\Delta H_1^{\ddagger}$ have not been reported previously, but generally similar values of $\Delta H^{\ddagger}$ and $\Delta S^{\ddagger}$ to ours have been found for the aminolysis of less reactive aromatic isothiocyanates by n-butylamine in methanol.[4] That study revealed an irregular dependence of the activation parameters upon the substituent in the isothiocyanate. The values of $\Delta H_2^{\ddagger} - \Delta H_{-1}^{\ddagger}$ (Table 1) show

Table 2. Tests for catalysis in diethyl ether solution.

$[ArNCS]_{initial}$/ 10^{-5} mol dm^{-3}	Amine	$[Amine]_{stoich}$/ 10^{-3} mol dm^{-3}	Tp/°C	Catalyst	[Catalyst]/ mol dm^{-3}	Rate data[a]			
≃3	*m*-$CH_3C_6H_4NH_2$	62	25	$NO_2C_6H_4NHCSNHC_6H_4CH_3$	0	$10^3 k_{obs}$	2.59		
					5.0×10^{-5}		2.59		
					9.6×10^{-5}		2.52		
				$NO_2C_6H_4NHCSNHC_4H_9$	16.8×10^{-5}	$10^3 k_{obs}$	2.59		
≃4	$C_6H_5CH_2NH_2$	6.95	13.8	$EtCO_2H$	0	$10^3 k_{obs}$	26.7	$K^b_{2:1}$	—
					54.8×10^{-3}		17.7		216
					65.5×10^{-3}		15.1		229
					131×10^{-3}		6.37		227
					262×10^{-3}		1.86		220
	m-$CH_3C_6H_4NH_2$	37.4		$EtCO_2H$	0	$10^3 k_{obs}$	0.902	k_3/k_{-1}	—
					0.131		1.17		7
					0.262		1.32		8
					0.502		1.47		7
				$PhCO_2H$	0	$10^3 k_{obs}$	0.902	k_3/k_{-1}	—
					3.48×10^{-2}		1.03		9
					4.45×10^{-2}		1.01		6
					9.17×10^{-2}		1.12		7
					20.8×10^{-2}		1.38		8
				CH_2ClCO_2H	0	$10^3 k_{obs}$	0.902	k_3/k_{-1}	—
					2.76×10^{-2}		0.907		7
					5.5×10^{-2}		0.07		8
					11.0×10^{-2}		1.14		7

For k_{obs} and k_3/k_{-1} see the text. [a] k_{obs}/s^{-1}; $K_{2:1}/dm^6$ mol^{-2}. [b] Free $[PhCH_2NH_2]$ calculated on the assumption that $k_{obs} = 3.94\ [PhCH_2NH_2]$ (see Table 1).

that $\Delta H^{\ddagger}_{-1}$ is the larger. That seems possible if k_2 reflects proton transfer. All in all, our derived rate constants and activation parameters for the spontaneous aminolysis are compatible with previous findings, and consistent with the mechanism proposed.

Catalysis.—With the reactions in diethyl ether solution, we have studied the effects of added carboxylic acids and *S*-ureas (products); these types of compound catalyse the aminolysis of isocyanates in non-hydroxylic solvents.[13,14]

S-ureas. Auto-catalysis by the *O*-urea product is a common feature of the aminolyses of isocyanates.[15] It is often detected when comparable amounts of amine and isocyanate are used. However, under first-order conditions (excess of amine) it tends to be relatively unimportant since the amount of urea (unless deliberately augmented) is always much less than the residual amine concentration, and the latter dominates any second-order (catalytic) terms. For aminolysis of isothiocyanates the effect of the product has been little studied.[1,3] We have tested for catalysis by the addition of *S*-urea at the start of the reaction in amounts comparable to that of the isothiocyanate (higher concentrations were not possible because of the insolubility of the *S*-urea). For aminolysis by *m*-toluidine we find a negligible catalytic effect under these conditions (Table 2). This result, and the good first-order plots obtained for all our systems, suggests that auto-catalytic effects have not influenced our results.

Carboxylic acids. Various carboxylic acids catalyse the reactions of isocyanates and ketenes with anilines in ethers and other solvents,[13] but Siggia found (qualitatively) that acetic acid inhibits the aminolysis of isothiocyanates by aliphatic amines in dioxane.[5] We find an analogous pattern of results. Reaction of *p*-nitrophenyl isothiocyanate with the strong amine benzylamine, is inhibited by added propionic acid. The effects can be quantitatively accounted for by the assumption that the free base is removed as an inactive base–acid (1:2) complex [equation (10)] with $K_{2:1} = 223 \pm 7$ at 13.8 °C (Table 2).

$$PhCH_2NH_2 + 2RCO_2H \overset{K_{2:1}}{\rightleftharpoons} PhCH_2NH_2(RCO_2H)_2 \quad (10)$$

Acid–base complexes of this sort are often detected in solvents of low polarity.[15] For aminolysis by the weaker base *m*-toluidine (which spectroscopic measurements at 290 nm show interacts negligibly or very little with the acids) we find that propionic, benzoic and chloroacetic acids all produce (small) accelerations (Table 2). If one bears in mind the mechanism indicated by the kinetics of the spontaneous reactions [equations (2) and (6)], the accelerations could perhaps be attributed to a medium effect of the polar acids, or to catalysis of the second (prototropic) step of the aminolysis, or to both. (It seems less likely that acids will catalyse the first step since isothiocyanates possess[10,17] relatively weakly basic N-atoms.) If we assume catalysis of the second step, and introduce a term k_3 [acid] (where k_3 is the rate constant for the acid-catalysed prototropy), then equation (3) becomes (11). By use of this equation, and rate constants obtained for the

$$k_{obs} = k_1 \frac{[amine]\,(k_2[amine] + k_3[acid])}{(k_{-1} + k_2[amine] + k_3[acid])} \quad (11)$$

spontaneous reaction (Table 1), values of k_3/k_{-1}, can be calculated. Reasonably self-consistent values are obtained at different values of [acid], and for each acid $k_3/k_{-1} = 7 \pm 1$ dm^3 mol^{-1} (Table 2). This suggests that all three acids are *ca.* four times less effective in the catalysis of the second step than *m*-toluidine itself (k_2/k_{-1} *ca.* 28 dm^3 mol^{-1}). A lack of dependence of catalytic effect on acid strength has been noted previously[16] for aminolysis of phenyl isocyanate in acetonitrile where catalysis by bifunctional reagents was considered to involve cyclic transition states, and the concerted removal and addition of a proton.[13,14] The effectiveness of acids is relatively much greater for isocyanates.

The general pattern of our results is similar to that of results of Williams and Jencks[2] for the aminolysis of cyanic acid in aqueous solutions. They too found the spontaneous additions

to be first-order in amine concentration for strong amines ($pK_a > 6$), and partly second order for weaker amines. For the former, catalysis by buffer acids and bases was absent, but for the latter it could bring the observed rate constant up to a limiting value which (they suggested) represented the rate of the first step of a mechanism like equation (2)

References

1 L. Drobnica, P. Kristiàn, and J. Augustin, in 'The Chemistry of Cyanates and their Thio Derivatives,' ed. S. Patai, Part 2, ch. 22, Wiley, Chichester, 1977.
2 A. Williams and W. P. Jencks, *J. Chem. Soc., Perkin Trans. 2*, 1974, 1753, 1760.
3 W. Vanasshe and G. Hoornaert, *Bull. Soc. Chim. Belg.*, 1971, **80**, 505.
4 P. Kristiàn, G. Suchór, and D. Podhradsky, *Collect. Czech. Chem. Commun.*, 1975, **40**, 2838.
5 J. G. Hanna and S. Siggia, *Anal. Chem.*, 1962, **34**, 547.
6 C. N. R. Rao and R. Venkataraghavan, *Tetrahedron*, 1963, **19**, 1509.
7 S. A. Lammiman and R. S. Satchell, *J. Chem. Soc., Perkin Trans. 2*, 1974, 877.
8 D. P. N. Satchell and T. J. Weil, *J. Chem. Soc., Perkin Trans. 2*, 1980, 1191.
9 D. P. N. Satchell and R. S. Satchell, *J. Chem. Research (S)*, 1988, 190.
10 N. Fuson, M.-L. Josier, R. L. Powell, and E. Utterbank, *J. Chem. Phys.*, 1952, **20**, 145.
11 J. W. Baker and D. N. Bailey, *J. Chem. Soc.*, 1957, 4649.
12 N. I. Yanchuk, USSR Deposited Doc., 1980, SPSTL 942 (*Chem. Abstr.*, **97**, 91336, 91346).
13 D. P. N. Satchell and R. S. Satchell, *Chem. Soc. Rev.*, 1975, **4**, 231.
14 J. M. Briody and D. Narinesingh, *Tetrahedron Lett.*, 1971, 4143.
15 E. J. King in 'Physical Chemistry of Organic Solvent Systems,' eds. A. K. Covington and T. Dickinson, ch. 2, Plenum, London, 1973.
16 D. Hadzi and S. Milicèv, in 'The Chemistry of Cyanates and their Thio Derivatives,' ed. S. Patai, Part 1, ch. 8, Wiley, Chichester, 1977.
17 D. P. N. Satchell and R. S. Satchell, *Z. Naturforsch.*, 1990, in press.
18 A preliminary publication (*Z. Naturforsch.*, 1989, **446**, 1329) contains k_{obs} values for dibenzylamine at 13.8 °C, and for *p*-anisidine at 25 °C in ether solution.

Paper 0/00578I
Received 7*th February* 1990
Accepted 9*th March* 1990