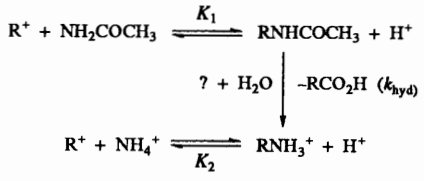
Detropylations of *N*,*N*-dimethyltropylamine and *N*-tropylacetamide

Josefina Palou,[a] Peter M. Robinson[b] and C. Ian F. Watt*[,b]
[a] *Departament de Quimica, Universitat de les Isles Balears, Cra de Valldemossa, 07071 Palma (Balears), Spain*
[b] *Department of Chemistry, University of Manchester, Manchester, UK M13 9PL*

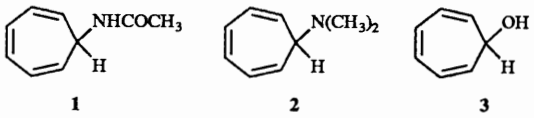
Rates have been measured for the decomposition of *N*-tropylacetamide, **1**, in aqueous acid yielding acetamide and tropylium, and for its formation from tropylium and acetamide in aqueous solution. Reaction of **1** in dilute hydrochloric acid has $k_{H^+} = 2.49 \times 10^{-2}$ dm^3 mol^{-1} s^{-1} at 25 °C, and the equilibrium constant for its formation from tropylium tetrafluoroborate and acetamide is 8.4×10^{-3}. In trichloroacetic and dichloroacetic acid buffers, rates show dependence on buffer concentration. A mechanism for the transformation is proposed, with *N*-protonated **1** as a transition state or high energy intermediate. Decomposition of *N*,*N*-dimethyltropylamine in aqueous acid solution at pH < 3.5 shows good first-order behaviour, $k = 4.58 \times 10^{-1}$ s^{-1} at 25 °C, with no dependence on acid concentration or on ionic strength.

We have recently speculated[1] that carbocations incorporated into crown ethers or related species could *N*-alkylate amides to anchor them within the cavity of the crown and thereby place them in a position where the hydrolytic reactivity of their *N*-acyl bonds might be enhanced. After such cleavage, release of ammonia, as the ammonium ion, from the resulting amine would regenerate the cation and establish a catalytic cycle as shown in Scheme 1.



Scheme 1 Catalysis of amide hydrolysis by functionalised carbocations

As a minimum requirement for operation of such a cycle, equilibria involving alkylation of the amide and dealkylation of the alkylammonium salt would have to be favourable (K_1 and $K_2 \gg 1$ in Scheme 1) and quickly established, and for operation in aqueous medium the carbocations would have to be relatively stable. Tropylium,[2] $C_7H_7^+$, with $pK_{R^+} = 4.76$,[3] stands high in the list of possible carbocation anchors. Its reactions with simple amines and amides to yield the corresponding *N*-tropyl† derivatives were described over 30 years ago by Doering and Knox[4] and the classic studies of Ritchie and Virtannen[5] contain rate measurements for reaction of tropylium with some amines. Nothing in the available data vitiates our speculation and we have already prepared some crowns incorporating tropylium.[1] However, equilibrium measurements for reactions of tropylium with both amines and amides are not available and to underpin our further studies we have made such measurements on the formation and decomposition of *N*-tropylacetamide, **1**, and of a simple *N*-tropylamine, **2**, and present them here. We also compare these results with earlier data on the acid-catalysed decomposition of the tropyl alcohol,[6] **3**, and with the available rate measurements for reactions of the tropylium ion with secondary amines.



1 **2** **3**

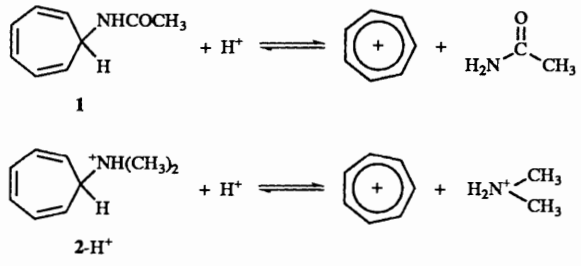
† Tropyl = cyclohepta-2,4,6-trien-1-yl.

Results

N-Tropylacetamide, **1**, was prepared here by a minor modification of the original method. Ideally, the amine studied here should have been the *N*-acyl cleavage product, tropylamine, $C_7H_7NH_2$, but we were unable to prepare this compound. Reaction of tropylium with ammonia yields either di- or tri-tropylamine, and both of these, and presumably tropylamine itself, react with oxidising agents, including the tropylium ion itself, to yield tropenylidenimmonium salts.[7] The tertiary amine, **2**, from reaction of tropylium with dimethylamine, was a more tractable compound and was used in this study of amine detropylation only since pure samples for the rate measurements could be obtained by distillation under reduced pressure.

The UV spectra of aqueous solutions of either **1** or **2** show absorption bands at 256 nm, characteristic of the 7-substituted cycloheptatrienyl residue. On addition of acid this absorption decays with concomitant growth of a band at 276 nm with a shoulder at 281 nm. The changes occur with an isosbestic point at 267 nm and permit monitoring of the regeneration of the tropylium ion from the covalently bound precursors. Rates for detropylation of the amide and amine have been determined in dilute acid solution ([**1** or **2**] < 10^{-4} mol dm^{-3}, pH < 3) where there is less than 2% conversion of the cation to tropyl alcohol.

Simple alkyl amides, including an isomer of **1**, *N*-benzylacetamide,[8] have conjugate acids with $-2 < pK_a < 0$, and are not significantly protonated at the acid strengths used in this work. The reaction stoichiometry for the amide reaction is as shown in Scheme 2. Because of the inductive effect of the cycloheptatrienyl group, the amine, **2**, is expected to be less basic than simple trialkylamines whose conjugate acids have $pK_a \approx 11$. Perrin's additivity scheme,[9] treating the cycloheptatriene as two vinyl substituents at β-carbons, predicts $pK_a = 8.36$ for **2**. Even in the unlikely event that this estimate is high



1

2-H$^+$

Scheme 2 Detropylations of **1** and **2** in aqueous acid

Table 1 First-order rate constants for reaction of *N*-tropylacetamide, **1**, in aqueous hydrochloric acid with ionic strengths maintained at 0.02 mol dm^{-3} by KCl

T/°C	[HCl]/mol dm^{-3}	$k_{obs}/10^{-4}$ s^{-1}
25.00	0.005	1.23
	0.010	2.50
	0.015	3.75
	0.020	5.00
33.00	0.020	12.4
41.00	0.020	26.9

* For rates at [HCl] = 0.02 mol dm^{-3}: $\log k = 11.862 - 4.521 \times 1000/T$. ($r^2 = 0.9999$).

Table 2 First-order rate constants for reaction of *N*-tropylacetamide, **1**, in buffered aqueous acid at 40 °C, with μ = 0.2 mol dm^{-3} (NaCl) and buffer ratios = 1

[Buffer]/ mol dm^{-3}	CCl_3COOH $k_{obs}/10^{-2}$ s^{-1}	$CHCl_2COOH$ $k_{obs}/10^{-3}$ s^{-1}
0.05	0.68	3.90
0.10	1.19	5.27
0.15	1.76	6.18
0.20	2.33	6.55

Table 3 First-order rate constants for reaction of *N*,*N*-dimethyltropylamine, **2**, in aqueous hydrochloric acid. Ionic strengths controlled or adjusted by addition of NaCl

T/°C	[HCl]/mol dm^{3}	μ/mol dm^{-3}	k/s^{-1}
9.23	0.005	0.020	6.72×10^{-2}
	0.010	0.020	6.58×10^{-2}
	0.015	0.020	6.49×10^{-2}
15.11	0.005	0.020	1.44×10^{-1}
	0.010	0.020	1.38×10^{-1}
	0.015	0.020	1.37×10^{-1}
	0.010	0.020	1.30×10^{-1}
	0.010	0.040	1.30×10^{-1}
	0.010	0.060	1.28×10^{-1}
	0.010	0.080	1.34×10^{-1}
	0.010	0.100	1.27×10^{-1}
21.43	0.015	0.020	3.01×10^{-1}
25.00[a]	0.015	0.020	4.58×10^{-1}

[a] Extrapolated from lower temperatures: $\log k = 14.901 - 4.544 \times 1000/T$. ($r^2 = 0.9999$).

by 2 pK_a units, the amine will be fully protonated to give **2**-H^+ under all the acid conditions used in this study. This protonation will be rapid[10] and certainly complete within the time of mixing of acid and amine solutions. The detropylation reaction of **2**-H^+ releases dimethylamine (pK_a = 10.73)[11] and this amine will also be completely converted to its conjugate acid under the reaction conditions and the net reaction occurring for the amine is that also shown in Scheme 1.

For reactions of both amine and amide, the appearance of the cation obeyed good first-order kinetics in all conditions used and the rate data are presented in Tables 1, 2 and 3.

It was also possible to observe equilibrium formation of *N*-tropylacetamide from acetamide and tropylium in aqueous solution. With large excesses of acetamide ([amide] > 0.5 mol dm^{-3}) the spectra of tropylium solutions in dilute acid showed changes consistent with the formation of *N*-tropylacetamide, with the approach to equilibrium showing good first-order behaviour. Rate constants for this approach, and equilibrium constants for formation of **1** under these conditions (*i.e.* the reverse of the reaction in Scheme 1) calculated from initial and final absorbance are also tabulated below in Table 4.

Table 4 Rate constants and equilibrium constants for reaction of tropylium with acetamide in aqueous hydrochloric acid at 30 °C

[Acetamide]/ mol dm^{-3}	[HCl]/mol dm^{-3}	$k_{obs}/10^{-4}$ s^{-1}	$K_{eq}/10^{-3}$
0.736	0.005	3.55	8.05
0.767	0.010	4.46	8.65

Discussion

Formation and cleavage of the amide

For the reaction of **1**, shown in Scheme 1, $K_{eq} = 1.1 \times 10^2$, so that the process is exoergic by 2.8 kcal mol^{-1}.‡ Analysis of the approach to equilibrium by tropylation of acetamide (Table 4) yields estimates for second-order rate constants for forward and reverse reactions of 2.4×10^{-4} and 3.0×10^{-2} dm^3 mol^{-1} s^{-1}, respectively, at 30 °C. The latter is not inconsistent with that determined directly for the detropylation of **1**. For this reaction in hydrochloric acid at 25 °C, the rate data (Table 1) yield the following dependence on acid concentration [eqn. (1)].

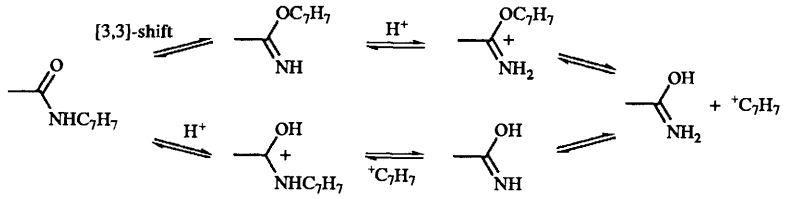
$$k_{obs} = 2.49\ (\pm 0.05) \times 10^{-2}\ [H^+]\ (r^2 = 0.9996) \quad (1)$$

Reaction is thus first-order in acid, with no detectable water-catalysed component. The temperature dependence yields activation parameters, $\Delta H^{\ddagger} = 20.1$ kcal mol^{-1} and $\Delta S^{\ddagger} = -6.25$ eu.§ In buffered aqueous medium at 40 °C, rates show dependence on buffer concentration, with trichloroacetic acid ($pK_a = 0.7$) giving $k_{cat} = 1.1\ (\pm 0.2) \pm 10^{-1}$ dm^3 mol^{-1} s^{-1} and dichloroacetic acid ($pK_a = 1.4$)[11] giving $k_{cat} = 1.8\ (\pm 0.3) \times 10^{-2}$ dm^3 mol^{-1} s^{-1} and the resulting two point Brønsted plot gives $\alpha = 1.12\ (\pm 0.21)$. The uncertainty in the coefficient is large and while Brønsted coefficients greater than one are not unknown,[12] almost all occur in reactions of nitro compounds. We have no structural reason to expect an anomalous exponent here and believe the value is consistent with general acid catalysis of the cleavage with a Brønsted α of just less than one.

A number of mechanisms may be considered for formation and decomposition of *N*-tropylacetamide, but we find it difficult to reconcile the general acid catalysis found in the decomposition of **1** with reaction *via* an *O*-protonated species. Amides show a thermodynamic preference for protonation on oxygen and *O*-protonated **1** might reasonably have $pK_a = -1$, and C–N heterolysis therein could certainly then yield tropylium ion and acetimidic acid (estimated[13] to be 6 kcal mol^{-1} less stable than acetamide) which would tautomerise rapidly (Scheme 3). However, given the likely close balance between the pK_a values of the protonated amide and H_3O^+, the low barriers to proton transfer between oxygen atoms of amide and water[10] and the modest exotherm in the net reaction, the mechanism offers no scope for a coupling of proton transfer with C–N heterolysis which avoids formation of a high energy intermediate in this mechanism.[14] The microscopic reverse of this reaction also contains implausibilities, notably in requiring that formation of **1** involves *N*-alkylation of acetimidic acid.

It is useful to compare reaction of amide **1**, with the acid-catalysed heterolysis of tropyl alcohol, **3**, which was examined by Bunton[6] and by Zuman[15] and their co-workers. This reaction is more exothermic ($\Delta G = -6.5$ kcal mol^{-1}) and the second-order rate constant for hydrolysis of **3** in HCl is 1.46×10^5 dm^3 mol^{-1} s^{-1} (*i.e.* some 1.67×10^7 larger than that for amide cleavage). Notably, the reaction also shows general

‡ 1 cal = 4.184 J.
§ 1 eu = 4.184 J k^{-1} mol^{-1}.



Scheme 3 Possible reactions of **1** *via* *O*-protonated or *O*-alkylated intermediates

acid catalysis with a series of carboxylic acids in water giving a Brønsted $\alpha = 0.8$ and general acid catalysis in this case was rationalised in terms of a concerted protonation and dissociation pathway which avoids a high energy intermediate identified as the protonated alcohol. Bunton estimated (by use of Taft σ^* values) that protonated tropanol has $pK_a = -7$, and tropanol is thus considerably less basic than the amide **1** in forming its *O*-protonated conjugate acid. As the pK_a of the protonated substrate becomes less negative and as the exotherm of its heterolysis also becomes less, so the probability increases that it will be an intermediate rather than a transition state. This is the case with the *O*-protonated amide.

Variations on the *O*-protonation theme, also shown in Scheme 3, might involve an equilibrium between **1** and and its imidate isomer (a retro-Chapman rearrangement),[16] accessible by a [3,3]-sigmatropic shift. These are known to occur in rearrangements of allyl imidates to amides[17] of allylamines, but typically require much higher temperatures (refluxing xylene) than used here. Certainly, this process should be endoergic (by *ca.* 6 kcal mol^{-1}), but cannot be rate limiting. *N*-Protonation and cleavage of the imidate would then yield tropylium and acetamide. In the reverse reaction, formation of **1** would then be initiated by *O*- rather than *N*-alkylation of acetamide, and fit with the known reactivity pattern of amides[18] with reactive alkylating agents, but again the scheme is difficult to reconcile with the observed general acid catalysis since the imidate ester would be considerably more basic than the amide and a concerted *N*-protonation and O–C heterolysis from the imidate again does not avoid any high energy intermediate.

Two possibilities in which neither *O*-protonated nor *O*-alkylated amides lie on the cleavage pathway remain to be considered. Both involve *N*-protonation and differ only in the timing of C···N and H···N bonding changes. The relationship between them is shown in a More O'Ferrall–Jencks diagram (Fig. 1).

Dissociation of amide **1** might yield a tropylium cation–imidate anion pair (top left of Fig. 1). The observed acid

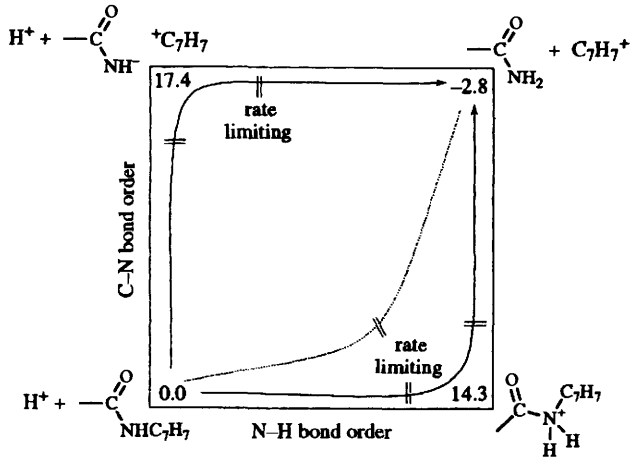


Fig. 1 Reaction pathways for cleavage or formation of **1**. Relative energies are in kcal mol^{-1}.

catalysis would then require that products are formed in a rate limiting acid-induced decomposition of this ion pair. Since acetamide has $pK_a = 15.1$,[11] this ion pair must lie *ca.* 20.6 kcal mol^{-1} above the tropylium and acetamide product or 17.4 kcal mol^{-1} above *N*-tropylacetamide. With the free energy of activation ($\Delta G^{\ddagger}$ at 25 °C) for the detropylation being 21.96 kcal mol^{-1}, this ion-pair is not immediately excluded from the reaction pathway by energy considerations, but since the acetimidate anion is sufficiently basic to be protonated by water and other weak acids, absence of a water-catalysed component militates against its intermediacy. Alternatively, the reaction may take an associative pathway with the *N*-protonated amide (bottom right of Fig. 1) as an intermediate. Perrin and co-workers have shown that N–H hydrogen exchange in secondary amides may occur by an *N*-protonation mechanism[19] and lifetimes of *N*-protonated amides in aqueous media are comparable to those for single bond rotations.[20] Intermediates of the form $CH_3CONH_3^+$ are estimated[21] to have $pK_a < -8$ and, because of the inductive effect of the cycloheptatrienyl substituent, *N*-protonated **1** might be even more acidic, possibly by as much as 2.5 p*K* units, placing this hypothetical intermediate 14.3 kcal mol^{-1} above the amide. Analogies now can be made with alcohol cleavage, although the amide cleavage is *ca.* 3.7 kcal mol^{-1} less exothermic. General acid catalysis could be then associated with rate limiting formation of *N*-protonated **1**, or a concerted pathway (dashed line in Fig. 1), which avoids full formation of either the cation–imidate anion pair, or of the *N*-protonated **1**, but *via* a transition state which more closely resembles *N*-protonated **1**.

In brief, in the amide cleavage, *O*-protonation is not productive compared with the competing *N*-protonation under the reaction conditions.

Formation and cleavage of the amine

The rates of detropylation of the amine salt, **2**-H$^+$, are not dependent on acid concentration nor on ionic strength (Table 3), indicating complete decoupling of C–N heterolysis from the amine protonation. The temperature dependence yields activation parameters, $\Delta H^{\ddagger} = 20.2$ kcal mol^{-1}, $\Delta S^{\ddagger} = 7.7$ eu, and the positive entropy of activation is consistent with a dissociative process.

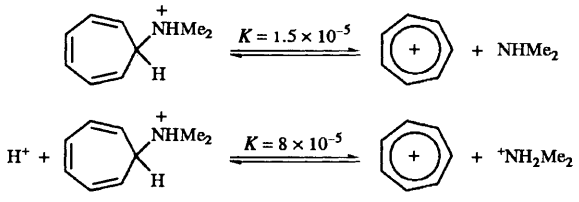
Decompositions of this type are the reverse of the well studied amine–carbocation combination reactions and it has been argued[22] that deviations from Ritchie's constant selectivity relationship in reactions of amines with carbocations arise at least partially from a kinetically significant desolvation of the amine in formation of an amine–carbocation complex. The effect is most evident with reactive electrophiles where amine–cation combination in the complexation step is very fast.[23] In decompositions of the type studied in this work, acid catalysis might then be expected if the reaction pathway involves reversible heterolysis of the alkylammonium ion to an amine–carbocation complex followed by rate limiting solvation or protonation of the amine. Behaviour of this type is probable in decompositions of alkylammonium salts derived from reactive carbocations and has indeed been found by Maskill and co-workers[24] in the solvolytic deamination of 4,4′-dimethoxy-

tritylammonium ions in 80% aqueous methanol, where rates increase with both acid and ionic strength. This alkylammonium salt is 3×10^{-5} less reactive than **2**-H^+ and since 4,4'-dimethoxytrityl cation is 6 pK_{R^+} units[25] less stable than tropylium, the reactivity difference reflects almost all of the stability difference between the two carbocations formed. To account for the absence of acid catalysis in decomposition of **2**-H^+ within the same mechanistic framework, we suggest that with higher stability of the tropylium cation, formation of the amine–cation complex, rather than its decomposition, has become rate limiting.

We have not been able to measure directly the equilibrium constant for formation of **2**, but an estimate can be obtained by combining our rate measurement for decomposition of the salt with available literature data. Although small deviations from the N_+ relationship have been found[23] even for stable cations such as 4,4'-$(Me_2N)_2$trityl ($pK_{R^+} = 7.0$),[26] the absence of any response to acid strength or medium composition in the reaction of **2**-H^+ suggests that its reactions with amines will be well described by the Ritchie equation [eqn. (2)].

$$\log k_{nuc}/k_{H_2O} = N_+ \qquad (2)$$

Bunton *et al.*[27] have estimated that the N_+ value of dimethylamine is 5.8, which is comparable with those of diethylamine and piperidine (5.85 and 6.25).[5] The rate of reaction of the tropylium with water[3] is known ($k_{25} = 0.0474$ $dm^3\ mol^{-1}\ s^{-1}$) and the second-order rate constant for reaction of dimethylamine with the cation can be estimated to be $3 \times 10^4\ dm^3\ mol^{-1}\ s^{-1}$. Taken with our experimental measure of the rate of decomposition of the *N*-tropyldimethylammonium, it is clear that the equilibrium constant for this process (see Scheme 4) is highly unfavourable in the absence of other



Scheme 4 Equilibria in decomposition of **2**-H^+

processes. However, as noted earlier, the net reaction involves protonation of the released dimethylamine ($pK_a = 10.73$), and when this is also taken into account, the process is a favourable one.

Conclusions

From the original viewpoint of testing the feasibility of using tropylium as a covalent anchor for temporary binding of amides in the active sites of amidase mimics, the results are fairly satisfactory. In acidic solution, decomposition of the amine to regenerate the cation is rapid and favourable and within the framework of the reactions of Scheme 1, is unlikely to be a rate limiting process. The binding constant for amide alkylation is not large, but may be sufficient if the hydrolysis of the amide bound within the cavity is sufficiently fast (in Scheme 1, governed by k_{hyd}). Bruice and Marquardt[28] have reported $k_{H^+} = 6.17 \times 10^{-6}\ dm^3\ mol^{-1}\ s^{-1}$ for hydrolysis of butyramide in aqueous acid at 30 °C and taking this as typical for simple primary amide hydrolyses, it is possible to set the requirements for k_{hyd} if cation catalysis is to be observed. For 0.01 mol dm^{-3} acid, an acid-catalysed hydrolysis would have a pseudo-first-order rate constant $k_{obs} = 6.17 \times 10^{-8}\ s^{-1}$.

For the putative catalysed reaction of Scheme 1, the rate is given by eqn. (3). If, for the sake of discussion, equal

$$v = k_{hyd}K_1[R^+][CH_3CONH_2]/[H^+] \qquad (3)$$

concentrations of cation and acid are used, the pseudo-first-order rate constant for amide hydrolysis would be $k_{obs} = k_{hyd}$ $8.4 \times 10^{-3}\ s^{-1}$, so that values of $k_{hyd} > 7.4 \times 10^{-4}\ s^{-1}$ will deliver rate enhancement over simple acid hydrolysis at pH 2. Mechanisms whereby such reactivity might be obtained are under investigation and are expected to involve metal-ion binding of the type known to enhance methanolysis rates of intra-crown phenol esters.[29]

Experimental

IR spectra were recorded on a Perkin-Elmer 1710–FT spectrometer, routinely on thin films deposited on KBr discs. ^1H NMR spectra were run on a Bruker AC 300E spectrometer operating at 300 MHz. Chemical shifts are reported in ppm relative to internal TMS and *J* values are given in Hz. Signal splittings are reported as singlet (s), doublet (d), triplet (t) or complex multiplet (cm).

Preparation of substrates

The amine was prepared by the published method of Doering and Knox[4] and showed analytical properties in accord with its structure and literature values. The amine was a liquid, which readily decomposed on exposure to typical laboratory atmospheres. It was stored under nitrogen and purified immediately before use in kinetic experiments by bulb-to-bulb distillation: bp 65 °C at 2 mmHg; $\delta_H(CDCl_3)$ 1.87 (1 H, t, *J* 5.0), 2.38 (6 H, s), 5.45 (2 H, dd, *J* 5.0 and 9.1), 6.17 (2 H, br d, *J* 9.0) and 6.72 (2 H, t, *J* 2.0).

The published amide preparation was modified as follows. Tropylium tetrafluoroborate (0.85 g) was dissolved in dry acetonitrile (10 cm^3). Acetamide (0.5 g) was added and the mixture warmed to dissolve it. Triethylamine was then added (0.7 g) with some exotherm observed. The solution was then evaporated to an oil on the rotary evaporator and the residue was partitioned between ether and water. The ethereal extracts were dried (K_2CO_3). Evaporation again yielded an oil which was crystallised from a mixture of toluene and ligroin to yield the amide as off-white crystals, mp 103 °C (lit.,103–104 °C); $\delta_H(CDCl_3)$ 1.95 (3 H, s), 4.46 (1 H, q, *J* 5.0), 5.55 (2 H, dd, *J* 5.0 and 9.1), 5.75 (1 H, br s), 6.41 (2 H, br d, *J* 9.0) and 6.72 (2 H, t, *J* 2.0).

Rate measurements for amine

Reactions of **2**-H^+ were monitored at 276 nm using a Pye-Unicam SP8-300 spectrometer fitted with a Hi-Tech stop-flow accessory. Temperatures were controlled with a Haake E3 circulating bath and were constant to ±0.2 °C during a run. Temperatures are accurate to ±0.1 °C. Sodium hydroxide solution, mixed with potassium chloride solution of appropriate concentration, was placed in one syringe of the apparatus and hydrochloric acid in the other, with concentrations arranged to give the desired final concentrations on mixing. A stock solution of the amine in acetonitrile (0.05 mol dm^{-3}) was prepared and *ca.* 20 μl (1 μl = 1 mm^3) was added to the 10 cm^3 reservoir syringe containing the base to give a solution of *ca.* 10^{-4} mol dm^{-3} in the amine. Initial absorbances after mixing were in the range 0.2–0.3 and final absorbances between 0.7 and 0.8. First-order rate constants were extracted from the data by digitising the absorbance data at 276 nm at between 40 and 50 points over four half-lives and non-linear regression of data of an exponential growth curve with initial and final absorbance and rate constant as adjustable parameters using a commercially available fitting routine.[30] Standard deviations of the para-

meters were always less than 1% of the value and calculated and observed infinity values showed excellent agreement. The tabulated rate constants are the means of three separate runs at each temperature or acid condition and agreed to better than 5%.

Amide reactions

For the decompositions of **1**, stoppered cells containing hydrochloric acid or buffer solution were placed in the block of a Zeiss DMR 11 spectrometer and allowed to equilibrate thermally for at least 20 min. To start the reaction, a stock solution of **1** in acetonitrile (*ca.* 20 μl of 0.02 mol dm^{-3}) was added by microsyringe and the cell stirred briefly. The changes in absorbance at 276 nm were monitored and rates obtained as described above. For the formation of **1**, solutions of tropylium tetrafluoroborate (*ca.* 5×10^{-5} mol dm^{-3}) in aqueous hydrochloric acid (3 cm^3 of solutions 0.005 or 0.010 mol dm^{-3}) were placed in a stoppered UV cell in the thermostatted cell block of the UV spectrometer. After 20 min, an accurately weighed amount of solid acetamide (*ca.* 0.120 g) was added, the stopper replaced, the cell shaken to effect solution, before returning to the thermostatted block of the spectrometer. Spectra were then run every 10 or 15 min until no further change took place. To obtain the absorbance for complete conversion of cation to a cycloheptatriene derivative, concentrated sodium hydroxide solution (50 μl of 5 mol dm^{-3} solution) was added and the spectrum re-run. Values of k_{obs} for approach to equilibrium were extracted by fitting the changes in absorbance at 276 nm to an exponential decay as described above. The equilibrium constant for formation of the *N*-tropylacetamide was calculated from the initial (A_o) and final absorbances (A_{eq}) and the absorbance (A_{qn}) after addition of concentrated sodium hydroxide solution (10 μl of *ca.* 10 mol dm^{-3}) to complete quenching of the cation, using the relation $K_{eq} = (A_o - A_{eq})[H^+]/(A_{eq} - A_{qn})$[acetamide].

Acknowledgements

This work was supported by SERC and by the University of Manchester. Dr Josefina Palou acknowledges support from the research fund of Universitat de les Isles Balears. We also thank Dr. Howard Maskill, University of Newcastle upon Tyne, for a useful and lively exchange of data and opinions.

References

1 B. G. Cox, O. S. Mills, N. J. Mooney, P. M. Robinson and C. I. F. Watt, *J. Chem. Soc., Perkin Trans. 2*, 1995, 697.
2 For a review of tropylium chemistry, see K. M. Harmon in *Carbonium Ions*, vol. IV, ch. 29, eds. G. A. Olah and P. v. R. Schleyer, Wiley, London and New York, 1973.
3 C. D. Ritchie and H. Fleischaur, *J. Am. Chem. Soc.*, 1972, **94**, 3841.
4 W. von E. Doering and L. H. Knox, *J. Am. Chem. Soc.*, 1957, **79**, 355.
5 C. D. Ritchie and P. O. Virtannen, *J. Am. Chem. Soc.*, 1973, **95**, 1882.
6 C. A. Bunton, F. Davouzedeh and W. E. Watts, *J. Am. Chem. Soc.*, 1981, **103**, 3855.
7 N. L. Bauld and Y. S. Rim, *J. Am. Chem. Soc.*, 1967, **89**, 6736; H. J. Dauben and D. F. Rhoades, *J. Am. Chem. Soc.*, 1967, **89**, 6746.
8 R. Huisgen and H. Brade, *Chem. Ber.*, 1957, **90**, 1432.
9 D. D. Perrin, B. Dempsey and E. P. Serjeant, pK_a *Prediction for Organic Acids and Bases*, Chapman and Hall, London and New York, 1981.
10 M. Eigen, *Angew. Chem., Int. Ed. Engl.*, 1964, **2**, 1.
11 pK_a Values taken from G. W. Kortum, W. Vogel and K. Andrussov, *Dissociation Constants of Organic Acids*, Butterworths, London, 1961.
12 For references and discussion of anomalous exponents see 'Proton Transfer Reactions', eds. E. Caldin and V. Gold, Chapman and Hall, London, 1975, ch. 7, *The Brønsted Relationship*, by J. Kresge.
13 R. Shaw in *The Chemistry of Amidines and Imidates*, ed. S. Patai, Wiley Interscience, London and New York, 1970, ch. 11.
14 R. A. More O'Ferrall, *J. Chem. Soc. B*, 1970, 240; W. P. Jencks, *Chem. Rev.*, 1972, **72**, 705.
15 P. Zuman, J. Chodkowski and F. Santavy, *Coll. Czech. Chem. Commun.*, 1961, **26**, 380.
16 J. W. Schulenberg and S. Archer, *Org. Reactions*, 1965, **14**, 1.
17 L. Overman, *Org. Syn.*, 1978, **58**, 4.
18 H. Muxfeldt and W. Rojalski, *J. Am. Chem. Soc.*, 1965, **87**, 933.
19 C. L. Perrin and C. P. Lollo, *J. Am. Chem. Soc.*, 1984, **106**, 2749; C. L. Perrin and G. M. L. Arrhenius, *J. Am. Chem. Soc.*, 1982, **104**, 6693.
20 For fuller discussion, see C. L. Perrin, *Acc. Chem. Res.*, 1989, **22**, 269.
21 A. R. Fersht, *J. Am. Chem. Soc.*, 1971, **93**, 3504; J. P. Krug, P. L. A. Popelier and R. F. W. Bader, *J. Chem. Phys.*, 1992, **96**, 7604.
22 W. P. Jencks in *Nucleophilicity*, eds. M. Harris and S. P. McManus, Advances in Chemistry Series, 215, ACS, Washington, 1987, p. 155.
23 R. A. McClelland, V. M. Kanagsabapathy, N. S. Banait and S. Steenken, *J. Am. Chem. Soc.*, 1992, **114**, 1816.
24 C. Bleasdale, B. T. Golding, W. H. Lee, H. Maskill, J. Riseborough and E. Smits, *J. Chem. Soc., Chem. Commun.*, 1994, 93.
25 N. C. Deno, J. Jaruzelski and A. Schreisham, *J. Am. Chem. Soc.*, 1955, **77**, 3044; E. M. Arnett and R. D. Bushwick, *J. Am. Chem. Soc.*, 1964, **86**, 1564.
26 R. J. Goldacre and J. N. Phillips, *J. Chem. Soc.*, 1949, 1724.
27 C. A. Bunton, N. Carrasco, F. Davoudzadeh and W. E. Watts, *J. Chem. Soc., Perkin Trans. 2*, 1980, 1520.
28 T. C. Bruice and F. H. Marquardt, *J. Am. Chem. Soc.*, 1962, **84**, 365.
29 R. Cacciapaglia and L. Mandolini, *Chem. Soc. Revs.*, 1993, 223.
30 Passage F.00, *Graphics and Numerics for the Macintosh*, D. Guenther, Passage Software Inc., 1989.

Paper 5/02841H
Received 3rd *May* 1995
Accepted 20th *June* 1995