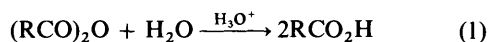


The Kinetics and Mechanism of the Mineral Acid-catalysed Hydrolysis of Carboxylic Acid Anhydrides in Water and in Dioxane–Water Mixtures. Application of the Excess Acidity Function to these Systems

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We report a kinetic study of the mechanism of the mineral acid-catalysed hydrolysis of some benzoic anhydrides, and of pivalic anhydride, in water and in dioxane–water mixtures. Employing a wider range of acid concentrations than used hitherto, we have considered the effects of acid concentration on the rate of hydrolysis *via* the excess acidity approach, and have examined the effects of substituents and of temperature. Contrary to previous suggestions, we find that benzoic and *p*-toluic anhydrides hydrolyse in water predominantly by the A1 mechanism. For these anhydrides, and for the *p*-methoxy and *p*-chloro derivatives, in 60% (v/v) dioxane–water as solvent, a change in mechanism (A2 to A1) is suggested by all the criteria used, but only the A1 mechanism is detectable for mesitoic anhydride. Pivalic anhydride, contrary to previous conclusions, exhibits the A2 and the A1 mechanisms in both water and dioxane–water. The advantage of dioxane–water mixtures over pure water as solvent for detecting mechanisms of acid catalysis, and the value of the excess acidity method for detecting changes in mechanism in the former type of solvent, are both illustrated.

There have been a number of studies^{1,2} of the kinetics and mechanism of the Brønsted acid-catalysed hydrolysis of carboxylic acid anhydrides [eqn. (1)]. Both aliphatic and aromatic



anhydrides have been examined using sometimes purely aqueous solvents, but more often solutions in dioxane–water mixtures or in similar mixed solvents. Results for hydrolyses of a range of aliphatic anhydrides in both water and dioxane–water, using various mineral acids as catalysts, have been interpreted by Bunton^{3–5} in terms of the A2 catalytic mechanism. Both the A1 and A2 mechanisms have, however, been proposed for aromatic anhydrides. The acid-catalysed hydrolysis of benzoic anhydride in both water and dioxane–water mixtures, containing up to 60% (v/v) dioxane has been interpreted by Bunton^{4,5} as involving an A2 mechanism. Nevertheless, a study by Koskikallio⁶ of *p*-substituted benzoic anhydrides in dioxane–water mixtures containing >80% dioxane, and up to *ca.* 0.1 mol dm⁻³ HClO₄ suggests that the mechanism can change from A2 to A1 as the dioxane content of the solvent increases. Koskikallio's approach was developed by Calvaruso^{2,7} who extended it to more aqueous mixtures (including 60% dioxane–water) and who used a range of more concentrated acid solutions (up to *ca.* 5.0 mol dm⁻³ HClO₄) similar to those used by Bunton. Calvaruso, like Koskikallio, concluded that both the A2 and A1 mechanisms can contribute to the hydrolysis of benzoic anhydrides, and that the A1 scheme is favoured by electron-releasing substituents, by a high dioxane content in the solvent, and by concentrated acid. The A2 scheme is favoured by electron-withdrawal.

Koskikallio and Calvaruso's conclusions are based largely upon ΔS^\ddagger values, and on changes in the pattern of substituent effects with changing medium composition. Attempts to use correlations,^{2–5} or the lack of them, with H_0 to diagnose A1 or A2 mechanisms, or to detect changes in mechanism, have proved largely unsuccessful with anhydrides, both for hydrolysis in water and in dioxane–water. This failure suggested to

us⁸ that the use of a wider catalyst concentration range (both higher and lower), and the application of the relatively recently introduced excess acidity (X) approach,⁹ might be valuable. This procedure also facilitates the calculation of the catalytic rate constant for an A1 path as $[\text{H}_3\text{O}^+] \rightarrow 0$. Catalytic constants were based on molarities and quoted as composites of A2 and A1 in earlier work.² Following this approach we have now found that the acid-catalysed hydrolysis of benzoic (and *p*-toluic) anhydride in water involves mainly the A1 mechanism, that pivalic anhydride exhibits both A2 and A1 mechanisms in water and in dioxane–water, and that the use of plots against X are indeed advantageous.

Experimental

Materials.—The anhydrides were either commercial samples (Aldrich) or were prepared from the corresponding acid *via* the sodium salt and acid chloride, using standard procedures.¹⁰ All had physical constants in agreement with literature values. The sources of the other materials were the same as in previous work.¹¹ All the anhydrides hydrolysed in water to give an effectively quantitative yield of the corresponding acid.

Kinetics.—Hydrolyses were followed using UV spectroscopy by observing the fall in the anhydride absorption in the region 240–270 nm. The carboxylic acids all absorb less strongly in this region, and the aromatic acids mostly absorb strongly at slightly shorter wavelengths. Good isosbestic points were always obtained. Initial anhydride concentrations were in the range 5×10^{-5} – 1×10^{-3} mol dm⁻³. The solvent was normally either 60% (v/v) dioxane–water or 1% (v/v) dioxane–water. Hydrolysis in either the presence or absence of added mineral acid led to an accurately first-order loss of anhydride, and the observed first-order rate constant, k_{obs} , was normally reproducible to within $\pm 5\%$. The principal mineral acid catalyst used was perchloric, but some experiments involved sulfuric acid. Dioxane–water mixtures containing mineral acid were made up as previously described.¹¹ Reactions were initiated by adding to the acidic solvent a small volume of a solution of the

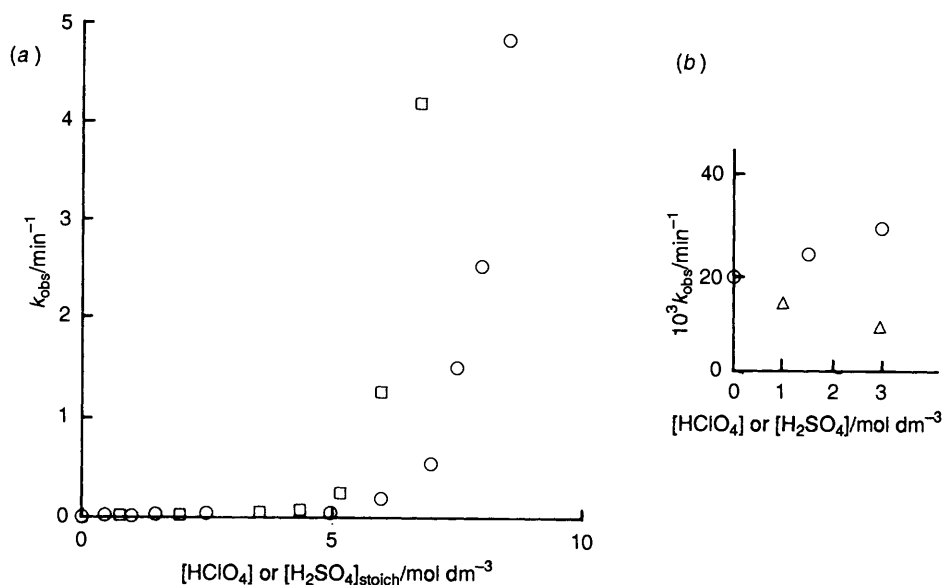


Fig. 1 Hydrolysis of benzoic and *p*-toluic anhydrides in 1% (v/v) dioxane-water at 25 °C: $[\text{anhydride}]_{\text{initial}} \approx 5 \times 10^{-5} \text{ mol dm}^{-3}$. (a) Benzoic anhydride with H_2SO_4 , \circ , and *p*-toluic anhydride with HClO_4 , \square . (b) Benzoic anhydride with H_2SO_4 , \circ , and with HClO_4 , \triangle .

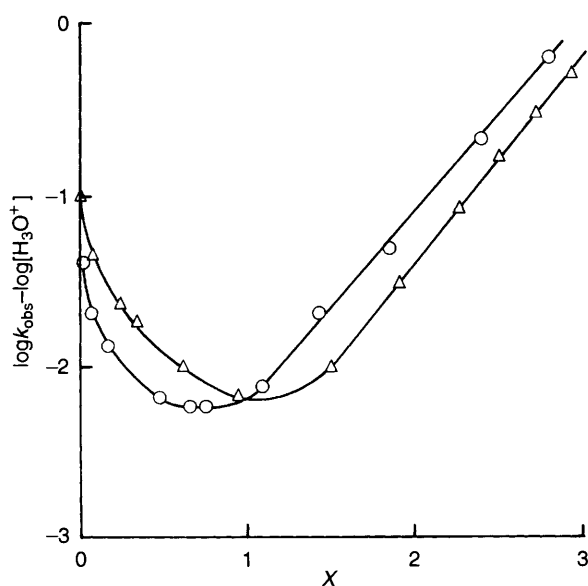


Fig. 2 Plot of $\log k_{\text{obs}} - \log [\text{H}_3\text{O}^+]$ against the excess acidity X in 1% dioxane-water at 25 °C; \circ , *p*-toluic anhydride; \triangle , benzoic anhydride

anhydride in pure dioxane. Determination of k_{obs} were made over a range of acid concentrations, and sometimes at different temperatures. Typical results are given in the Tables and Figs.

Results and Discussion

(i) *Benzoic Anhydride*.—Our results for this anhydride are illustrated in Figs. 1–5 and Tables 1 and 2. Of two previous studies of benzoic anhydride by Bunton^{4,5} the first used a 60% (v/v) dioxane-water solvent. In this solvent at 25 °C acid catalysis is readily detectable when $[\text{H}_3\text{O}^+] = 0.1 \text{ mol dm}^{-3}$, and k_{obs} is found to increase *ca.* 40-fold between 0 and 3.0 mol dm^{-3} perchloric acid.⁴ This effect is much larger than the effect of salts on the spontaneous hydrolysis of anhydrides in this medium,^{2,4,12} an effect which, for perchlorates, reduces rather than increases the rate. Bunton's second study⁵ used water as solvent and found little evidence for acid catalysis when $[\text{H}_3\text{O}^+] \approx 3.0 \text{ mol dm}^{-3}$, but that a significant increase in rate begins at higher concentrations. Bunton interpreted this

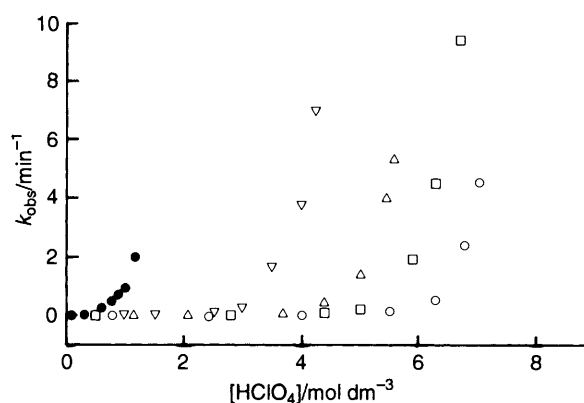


Fig. 3 Hydrolysis of benzoic anhydrides in 60% (v/v) dioxane-water at 25 °C; typical results for k_{obs} : R in (*p*- $\text{RC}_6\text{H}_4\text{CO}_2\text{O}$): Cl, \circ ; H, \square ; Me, \triangle ; MeO, ∇ ; mesitoic anhydride, \bullet ; $[\text{anhydride}]_{\text{initial}} \approx 10^{-3}$ – $10^{-5} \text{ mol dm}^{-3}$

increase in terms of an A2 mechanism. His treatment, however, as he implies, is equivocal, and is unsupported by other criteria of mechanism. We first studied benzoic anhydride using a largely aqueous solvent (1% dioxane), and found, in agreement with Bunton, that low concentrations of mineral acids have little effect on k_{obs} ; an increase in the concentration of perchloric acid from 0 to 3.0 mol dm^{-3} produces an approximately 50% decrease in k_{obs} , whereas with sulfuric acid there is a comparable increase (Fig. 1). Effects of this magnitude can arise from salt effects on the spontaneous hydrolysis,^{12,13} and there is no evidence in the purely aqueous medium for the important catalysis found at low acid concentrations in 60% dioxane-water. However, when the mineral acid concentration is sufficiently increased, a powerful catalysis is detected (*e.g.* Fig. 1). This catalysis is very probably predominantly of the A1 variety, rather than A2, since, in this region, a plot⁹ (Fig. 2) of $\log k_{\text{obs}} - \log [\text{H}_3\text{O}^+]$ against X (the excess acidity¹⁴) is rectilinear with a slope⁹ $m^*m^\ddagger \approx 1.25$, which leads to $m^\ddagger \approx 2$ if m^* for this O-base is taken¹⁴ as 0.5–0.6.† A plot (not shown) of $\log k_{\text{obs}}$ against $-H_0$ when $[\text{H}_2\text{SO}_4]_{\text{stoich}} \approx 5.0 \text{ mol dm}^{-3}$ is

† In the excess acidity approach⁹ the parameter m^\ddagger is related to transition state structure. It is normally found^{9,11} that for A1 mechanisms $m^\ddagger > 1$, and for A2 mechanisms $m^\ddagger \approx 1$.

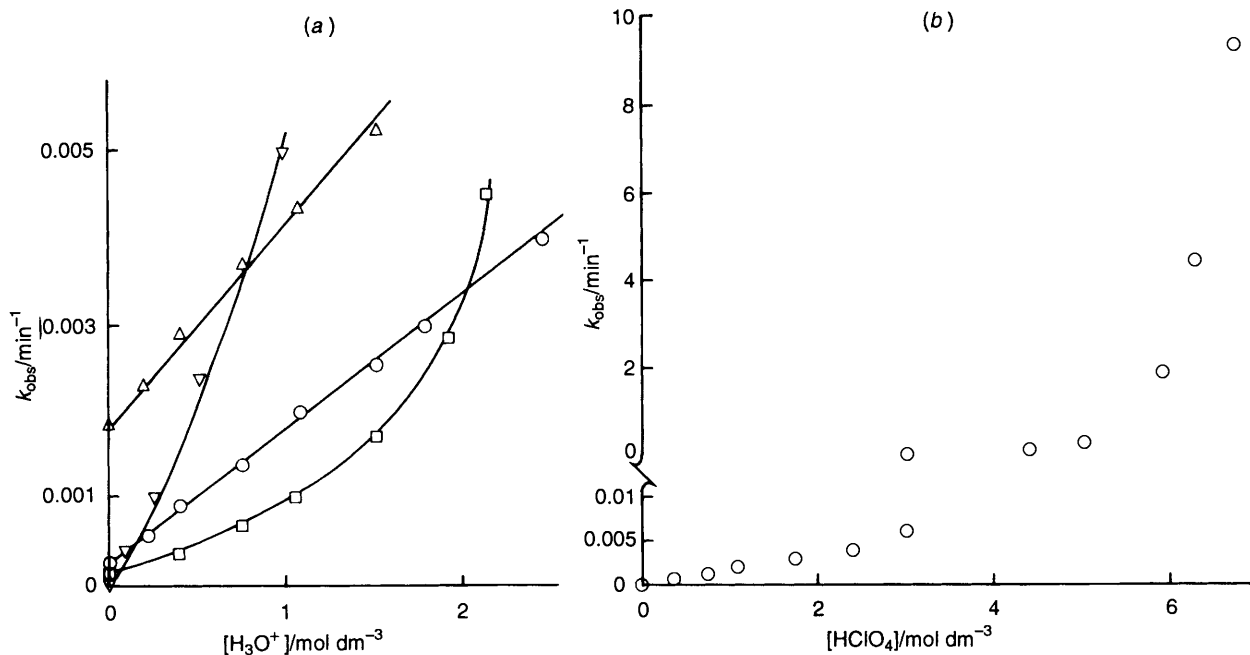


Fig. 4 Behaviour at low acidities in 60% (v/v) dioxane-water at 25 °C. (a) R in $(\text{RC}_6\text{H}_4\text{CO})_2\text{O}$: H, \circ ; Me, \square ; Cl, \triangle ; MeO, ∇ . (b) High and low acidity regions for benzoic anhydride, perchloric acid assumed fully dissociated.

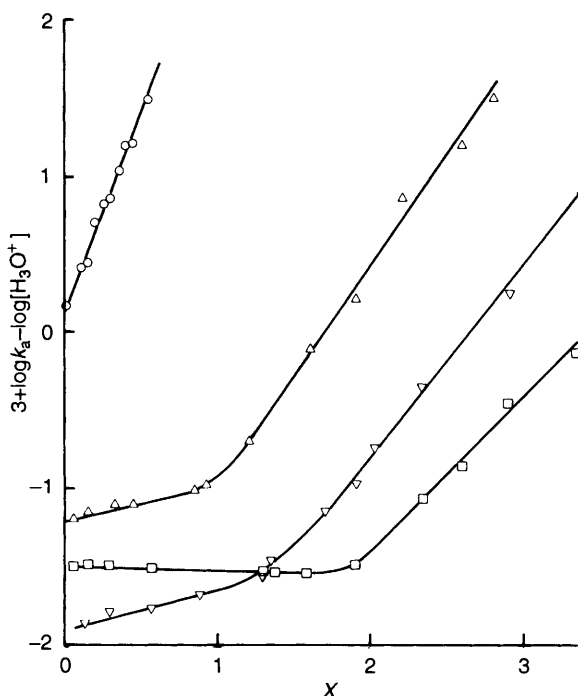


Fig. 5 X-Plots for benzoic anhydrides in 60% (v/v) dioxane-water at 25 °C; benzoic, \square ; *p*-toluic, ∇ ; *p*-methoxybenzoic, \triangle ; mesitoic, \circ ; $k_a = k_{\text{obs}} - k_0$; X values calculated (relative to 60% dioxane-water) from H_0 values¹⁵ by method given in ref. 11 using $\Delta pK_a = 1.36$; k_a in s^{-1}

also a straight line of slope *ca.* 1.0, and the effect of temperature when $[\text{H}_2\text{SO}_4]_{\text{stoich}} = 8.0 \text{ mol dm}^{-3}$ (Table 2) gives $\Delta S^\ddagger = -15 \text{ J K}^{-1} \text{ mol}^{-1}$. These results are all compatible with dominant A1 mechanism of catalysis in relatively concentrated aqueous acid, and this is further supported by the effect of changing the *p*-substituent in the anhydride (see below). Calvaruso's conclusion² that an A1 mechanism contributes to k_{obs} in dioxane-water also suggests that an A1 scheme should be found in water at high enough acid concentrations. We find that any contributions from an A2 scheme are not clearly evident in water.

These results for purely aqueous solutions led us to examine the behaviour reported^{2,4} for the 60% dioxane-water solvent

using a wider range of mineral acid concentrations. Our results using perchloric acid are shown in Figs. 3–5. The findings at low values of $[\text{H}_3\text{O}^+]$ support the previous work. (As noted above perchlorate salts have always been found to lead to a small negative salt effect on the spontaneous hydrolysis so that if anything, the acid catalysis is slightly greater than that observed in the limited region where the spontaneous hydrolysis contributes significantly to k_{obs} .) When $[\text{H}_3\text{O}^+] \gtrsim 2.0 \text{ mol dm}^{-3}$ we find that k_{obs} is largely proportional to $[\text{H}_3\text{O}^+]$ (Fig. 4), and at $[\text{H}_3\text{O}^+] = 1.1 \text{ mol dm}^{-3}$, $\Delta S^\ddagger = -127 \text{ J K}^{-1} \text{ mol}^{-1}$ (Table 2). Faced with similar results Bunton⁴ proposed an A2 mechanism of catalysis, and we concur. This mechanism is also compatible with the effects of substituents² (see also below). However, at higher values of $[\text{H}_3\text{O}^+]$ we find (Fig. 3) a marked increase in k_{obs} of the type found in the purely aqueous medium, and the sharp increase in the acidity dependence occurs at a somewhat lower value of $[\text{H}_3\text{O}^+]$ than in the latter medium. At $[\text{H}_3\text{O}^+] = 6.3 \text{ mol dm}^{-3}$, $\Delta S^\ddagger = 11 \text{ J K}^{-1} \text{ mol}^{-1}$ (Table 2) and an X -plot for 60% dioxane-water (Fig. 5) shows a clear discontinuity at $X \approx 1.9$ ($[\text{H}_3\text{O}^+] \approx 3.0 \text{ mol dm}^{-3}$). In the strongly acidic region this plot has a slope ≈ 1.04 , so that $m^\ddagger \approx 1.7$ – 2.1 , assuming m^* remains the same for dioxane-water.^{14,15} These results all suggest that the A1 mechanism of catalysis, evident in the purely aqueous solvent at high acidity, takes over from the A2 mechanism in 60% dioxane-water as the value of $[\text{H}_3\text{O}^+]$ is increased. This conclusion is compatible with those of Koskikallio and Calvaruso. Table 1 contains the corresponding catalytic rate constants, k_{H^+} and k'_{H^+} . The greater increase in the rate constant (k_0) of the spontaneous hydrolysis of the anhydride between the 60% and the 1% dioxane-water mixtures (Table 1) than in the catalytic rate constant (k_{H^+} , k'_{H^+}), together with the larger change found^{11,15} in the acidity functions (H_0 or X) for a given change in $[\text{H}_3\text{O}^+]$ in dioxane-water mixtures than in water, combine to reduce the relative importance of the catalytic routes in the essentially aqueous solvent, and make (i) the A2 path effectively undetectable for benzoic anhydride in water, and (ii) its A1 route appear at a higher value of $[\text{H}_3\text{O}^+]$ than in 60% dioxane-water. In a 30% (v/v) dioxane-water medium, intermediate behaviour is found (results not shown).

(ii) *p*-Methylbenzoic Anhydride.—This anhydride was studied

Table 1 Spontaneous and derived catalytic rate constants for anhydride hydrolysis at 25 °C

Anhydride	Solvent	$10^5 k_0$ ^a /s ⁻¹	$10^5 k_{H^+}$ ^b /dm ³ mol ⁻¹ s ⁻¹	$10^5 k'_{H^+}$ ^c /dm ³ mol ⁻¹ s ⁻¹
Benzoic	1% D/W	32 ^d	—	0.23
	60% D/W	0.33 ^d	3.0	0.030
<i>p</i> -Methylbenzoic	1% D/W	17	—	0.73
	60% D/W	0.15	1.2	0.10
<i>p</i> -Chlorobenzoic	60% D/W	3.0	4.5	—
<i>p</i> -Methoxybenzoic	60% D/W	0.07	—	0.46
Mesitoic	60% D/W	—	—	76
Pivalic	1% D/W	6.1	40	7.0
	60% D/W	0.50	12.5	1.8

^a Spontaneous rate constant. ^b Second-order (catalytic) rate constant for low acidity route (A2) calculated from $(k_{obs} - k_0)/[H_3O^+]$ as $[H_3O^+] \rightarrow O$. ^c Catalytic rate constant for high acidity route (A1) calculated⁹ from intercept of relevant rectilinear *X*-plot at *X* = 0. ^d Given wrong units in ref. 8.

Table 2 Effects of temperature

Anhydride	Solvent	$[HClO_4]$ /mol dm ⁻³	$[H_2SO_4]$ /mol dm ⁻³	$k_{obs}/10^{-3}$ s ⁻¹	<i>T</i> /°C	ΔH^\ddagger /kJ mol ⁻¹	ΔS^\ddagger /J K ⁻¹ mol ⁻¹	
Benzoic	1% D/W		8.00	13.3	16.5	101 ± 13	-15 ± 4	
				44.0	25.0			
				130	33.0			
	60% D/W	1.12		6.32	0.032	25.0	61 ± 9 ^a	-127 ± 22
					0.067	32.1		
					0.070	35.0		
0.165					45.0			
19.4					16.5			
42.8					21.0			
Mesitoic	60% D/W	0.040		75.2	25.0	116 ± 20	11 ± 4	
				0.041	25.0			
				0.125	33.0			
				0.320	40.0			
				0.321	25.0			
				0.920	37.0			
Pivalic	1% D/W	0.60		1.27	44.0	57 ± 9 ^a	-105 ± 18	
				11.5	13.0			
				23.0	18.5			
				56.1	25.0			
				45.0	14.0			
				138	18.5			
	60% D/W	1.20	4.80		280	25.0	<i>ca.</i> 60 ^b	-106 ^b
					45.0	14.0		
					138	18.5		
					115 ± 20			
					48 ± 12			
					280	25.0		

^a Corrected for spontaneous rate using data in refs. 3 and 4. ^b From ref. 3.

using both 1% and 60% dioxane–water solvents, with perchloric acid as catalyst. The behaviour observed was analogous to that found for benzoic anhydride (Figs. 1–5, Table 1). In the 1% dioxane solvent, there is again little evidence for catalysis at low values of $[H_3O^+]$, but the onset of catalysis occurs at lower values of $[H_3O^+]$ and *X* than for benzoic anhydride, k'_{H^+} being significantly larger and k_0 smaller (Fig. 1, Table 1). The slope of the *X*-plot leads to $m^\ddagger \sim 2$. These facts again suggest a dominant A1 mechanism of catalysis in this solvent. In the dioxane-rich medium at low acidity (≈ 1.0 mol dm⁻³), *p*-toluic anhydride reacts less rapidly, and is less susceptible to acid catalysis (see Fig. 4 and k_0 , k_{H^+} in Table 1) than is benzoic anhydride, but the onset of the rapid acceleration in rate occurs at a lower value of $[H_3O^+]$ than for the latter, and at higher acidities *p*-toluic anhydride is substantially the more reactive (Figs. 3 and 4). Our results up to 4.0 mol dm⁻³ are in good agreement with Calvaruso's.² The *X*-plot (Fig. 5) shows a clear discontinuity, and at high acidity $m^*m^\ddagger \sim 1.1$, giving $m^\ddagger \lesssim 2$. These results support the findings with benzoic anhydride, and the suggested change in mechanism of catalysis in dioxane-rich solvents. Electron release would, on balance, be expected to favour an A1 mechanism (more protonation, more stabilisation of the acylium ion, but somewhat poorer leaving group) but, even while increasing the amount of protonation,

can reduce the rate of an A2 process by dispersing the positive charge on the atom being attacked by water. This dispersion is also the probable reason that the spontaneous rate constant, k_0 , is smaller for *p*-toluic anhydride. Stronger net substituent effects are therefore to be expected on k'_{H^+} and k_0 than on k_{H^+} , for which more balanced effects should operate.

(iii) *p*-Chloro- and *p*-Methoxy-benzoic Anhydrides and Mesitoic Anhydride.—These anhydrides were only studied using a 60% dioxane–water solvent. The *p*-chloro derivative has a behaviour pattern obviously similar to those of the *p*-H and *p*-Me derivatives (Figs. 3 and 4, and Table 1). It is more reactive at low acidities (A2 catalysis) but less reactive at high acidities, and requires a higher value of $[H_3O^+]$ for the onset of the proposed A1 mechanism (Fig. 3). The details for this anhydride are not included in Fig. 5 because the available *X* values do not cover the high acidity region adequately.

p-Methoxy benzoic anhydride, by contrast, is notably more reactive than the *p*-Me derivative at high acidities and the proposed A1 mechanism becomes dominant in relatively dilute acid (Fig. 3). The value of k_{obs} only falls below those for the *p*-H and *p*-Me compounds at comparatively very low values of $[H_3O^+]$ (Fig. 4); this is probably because k_{obs} in the low acidity region, where for the foregoing anhydrides it mainly reflects the

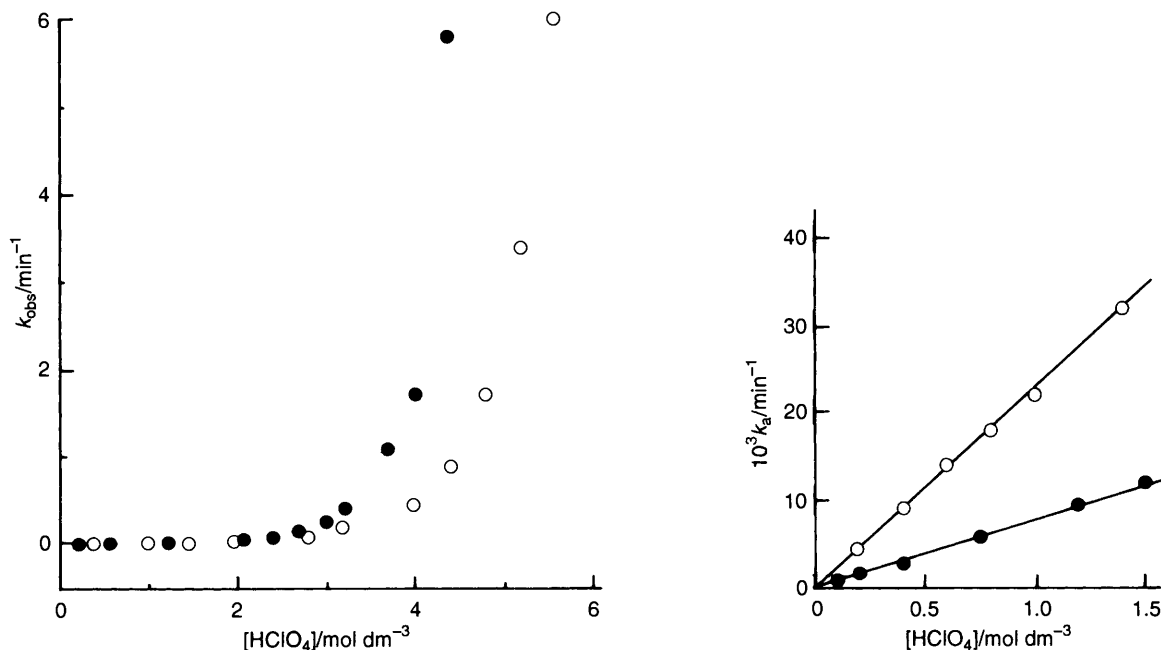


Fig. 6 Catalysed hydrolysis of pivalic anhydride at 25 °C; water, ○; 60% dioxane-water, ●; $k_a = k_{\text{obs}} - k_0$

A2 catalysis, is for the *p*-MeO derivative everywhere affected by the onset of the A1 mechanism. We are not able to obtain a value for k_{H^+} directly, but a Hammett plot for the *p*-Cl, *p*-H, and *p*-Me derivatives indicates that its value as $[\text{H}_3\text{O}^+] \rightarrow 0$ will be less than twice that for the A1 mechanism, k'_{H^+} , and of course the latter mechanism is much more acidity dependent. Support for a change in mechanism with increasing acid concentration is evident in Fig. 5. The slope of the *X*-plot becomes steady when $[\text{H}_3\text{O}^+] \lesssim 1.5$ mol dm^{-3} , it is greater than those for the *p*-H and *p*-Me derivatives, and gives $m^\ddagger \sim 2-3$. This is supportive of a dominant A1 scheme in the high acidity region. A plot² against H_0 up to 3.0 mol dm^{-3} perchloric acid gives no hint of a change in mechanism; the *X*-plot is notably superior. It is also superior for the *p*-H and *p*-Me compounds.

Mesitoic anhydride is so susceptible to the A1 mechanism that it completely dominates even in very dilute acid solutions (Fig. 3), and the *X*-plot for this anhydride shows no evidence of a change in the mechanism of acid catalysis (Fig. 5). This plot has a relatively high slope: $m^*m^\ddagger \approx 2.5$, giving $m^\ddagger \approx 4$. Even at $[\text{H}_3\text{O}^+] = 0.04$ mol dm^{-3} , we find $\Delta S^\ddagger = 55$ J K^{-1} mol⁻¹ (Table 2). With mesitoic anhydride steric crowding probably hinders the A2 scheme, and helps the A1 route, as in hydrolysis of the corresponding esters. Bunton reports⁴ a small number of measurements with mesitoic anhydride, and also suggests an A1 mechanism of catalysis.

Our values for k'_{H^+} (Table 1), and also the details in the Figs, show that the sequence of reactivities of the various benzoic anhydrides in the proposed A1 mechanism is mesitoic \gg *p*-MeO $>$ *p*-Me $>$ *p*-H $>$ *p*-Cl. For the A2 mechanism, the k_{H^+} values give *p*-Me $<$ *p*-H $<$ *p*-Cl. Hammett plots (against σ) using available k_{H^+} and k'_{H^+} values for the mono-substituted derivatives give $\rho \approx -4.5$ for the A1 catalysis, and $\rho \approx 1.3$ for the A2 catalysis. Our k_0 values give $\rho \approx 4.0$ for the spontaneous hydrolysis. Such sequences, and differences in sensitivity to substituent effects, were predicted in (ii) above, and support the (qualitative) conclusions of Koskikallio⁶ and Calvaruso.² For the A1 mechanism a plot against σ^+ is a little better than against σ , and gives $\rho^+ \approx -1.5$.

(iv) *Pivalic Anhydride*.—The acid-catalysed hydrolysis of pivalic anhydride in both aqueous solution and (less extensively) in 60% dioxane-water, was studied by Bunton.³ He compared the behaviour of this anhydride with that of acetic

anhydride, using various mineral acids as catalysts. His acid concentrations were always ≥ 5.0 mol dm^{-3} , and most of the experiments were at acidities < 3.0 mol dm^{-3} . He found, in contrast to the benzoic anhydrides discussed above, that both these aliphatic anhydrides display easily detectable catalysis at low acid concentrations in water as well as in dioxane-water. For the purely aqueous solvent, Bunton found that pivalic anhydride was always significantly less readily hydrolysed (ca. 15–30-fold) than acetic anhydride, that in general plots of $\log k_{\text{obs}}$ against $-H_0$ had irregular slopes < 1.0 , and that $\Delta S^\ddagger \sim -35$ to -110 J K^{-1} mol⁻¹ when $[\text{H}_3\text{O}^+] \gtrsim 3.0$ mol dm^{-3} . A similar situation probably applies for the dioxane-water solvent. Bunton concluded that under these conditions both anhydrides follow an essentially A2 mechanism of catalysis in both solvents. This conclusion is compatible with the structures of the anhydrides, since pivalic anhydride would be expected to be the more sterically hindered towards bimolecular attack by water, quite apart from any effect of the extra electron-release on the charge on the carbonyl carbon atom.

Our findings for the benzoic anhydrides led us to re-investigate the behaviour of pivalic anhydride, again using a wider mineral acid concentration range than Bunton's, and with more emphasis on the behaviour at the higher acid concentrations. We used perchloric acid as the catalyst. Where conditions overlap our observed rate constants are in reasonable agreement with the earlier work, and our results (Fig. 6) for a 1% dioxane-water solvent are plotted against *X* in Fig. 7. This Fig. suggests a change in mechanism. When $X \lesssim 1$ we find $m^\ddagger > 2.0$ which suggests an A1 mechanism of catalysis. Bunton's results (at 0 °C) are also included in Fig. 7 and it can be seen that his last (high concentration) point is out of line with the rest, and indicative of an exceptional increase in the rate of hydrolysis. Careful examination of Bunton's results shows that the evidence for a change in mechanism is becoming detectable in his most acidic solutions. For $[\text{H}_3\text{O}^+] = 5.2$ mol dm^{-3} we find $\Delta S^\ddagger \sim -19$ J K^{-1} mol⁻¹, and for $[\text{H}_3\text{O}^+] = 0.6$ mol dm^{-3} , $\Delta S^\ddagger \sim -105$ J K^{-1} mol⁻¹ (Table 2). This latter value is in satisfactory agreement with Bunton's value for 2.4 mol dm^{-3} perchloric acid solutions, and suggests more solvent involvement in the transition state at low acid concentrations. The results in this region can be interpreted as indicating an A2 mechanism as proposed by Bunton, but the results at high acidity imply an A1 scheme.

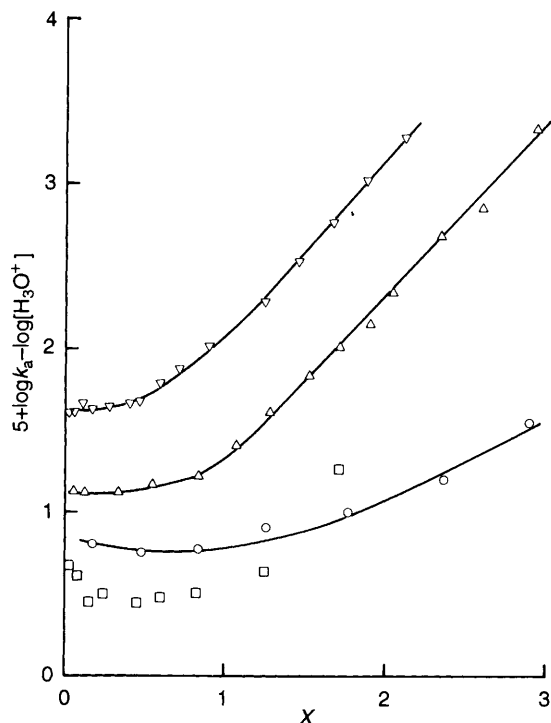


Fig. 7 X-plots for pivalic and camphoric anhydrides. Pivalic: water at 25 °C, ∇ ; water at 0 °C, \square ; 60% dioxane-water at 25 °C, \triangle . Camphoric: water at 55 °C, \circ . $k_a = k_{obs} - k_0$; k_a in s^{-1} .

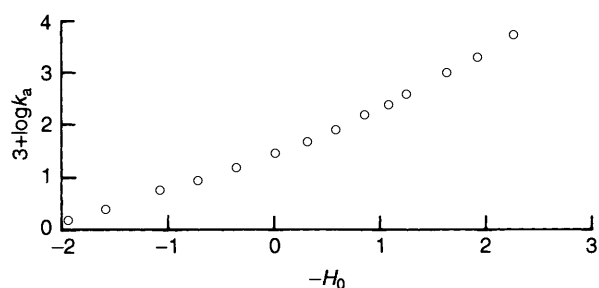


Fig. 8 Plot of $3 + \log k_a$ against H_0 for pivalic anhydride in 60% (v/v) dioxane-water at 25 °C; $k_a (=k_{obs} - k_0)$ in min^{-1}

For the 60% dioxane-water medium our results for pivalic anhydride are also in Figs. 6 and 7. The behaviour pattern is similar to that for water but, as found for the benzoic anhydrides, the change in acidity dependence becomes apparent at a slightly lower value of $[H_3O^+]$ in dioxane-water than in water. A plot against H_0 is again less diagnostic (Fig. 8). At high acidity, the X-plot gives $m^{\ddagger} \sim 1.7$ and at $[H_3O^+] = 4.5 \text{ mol dm}^{-3}$, we find $\Delta S^{\ddagger} = 48 \text{ J K}^{-1} \text{ mol}^{-1}$ (Table 2). Both these results are compatible with an A1 mechanism. In the low acidity region (up to $[H_3O^+] \approx 1.5 \text{ mol dm}^{-3}$) the acid catalysis depends closely on $[H_3O^+]$ and when $[H_3O^+] = 1.2 \text{ mol dm}^{-3}$, Bunton gives $\Delta S^{\ddagger} \sim -106 \text{ J K}^{-1} \text{ mol}^{-1}$; these results are again suggestive of an A2 mechanism at low acid concentrations for this anhydride. For all the hydrolyses studied it is consistently found (Table 2) that ΔH^{\ddagger} for the route identified as an A1 mechanism is about twice as large as that for the A2 route. Larger ΔH^{\ddagger} values are expected for unimolecular mechanisms. We conclude that pivalic anhydride, like most of the benzoic anhydrides in 60% dioxane-water, suffers a change in the mechanism of its acid catalysis as $[H_3O^+]$ is increased. For pivalic anhydride, however, this change in mechanism is also apparent in purely aqueous acid solutions. The reasons that catalysis can readily be detected at low acidities in water for pivalic and acetic anhydrides, whereas it cannot for the benzoic

anhydrides, is that the former anhydrides are more susceptible to the A2 mechanism, and that for them the spontaneous rate of hydrolysis does not increase so much on moving from dioxane-water to water as solvent (Table 1).^{3,4} The A2 mechanism is therefore not swamped by the spontaneous route for the aliphatic anhydrides in water.

Bunton quotes results⁵ for the hydrolysis of camphoric anhydride (a cyclic anhydride, but otherwise rather similar structurally to pivalic) in water at 55 °C over the unusually wide acidity range of 0–7.0 mol dm^{-3} perchloric acid. An X-plot (X at 25 °C) of these results (Fig. 7) is more continuous than that for pivalic anhydride, and its slope does not exceed 0.5, giving $m^{\ddagger} \approx 1$ for camphoric anhydride. This behaviour suggests to us an A2 mechanism of hydrolysis obtains for this cyclic compound even in strongly acidic solutions.

The use of the slope of a plot against the excess acidity as a criterion of mechanism is a relatively recent innovation,⁹ although reasonably well tested for aqueous sulfuric acid solutions.¹⁶ We have recently provided¹¹ a number of applications of the method to aqueous perchloric acid, and to solutions of perchloric acid in 10% dioxane-water. It is accepted that caution is necessary when applying acidity functions to mechanistic analysis, especially where mixed solvents are concerned.^{15,17} However, for the present systems we suggest that the implications of all our results (including those of the X-plots) for both the purely aqueous and the dioxane-water solvents are rather self-consistent and consistent with Calvaruso's conclusions (if not with some of Bunton's). There is, of course, every reason to expect that carboxylic anhydrides will exhibit, in appropriate media, changes in their acid-catalysed mechanism of hydrolysis, just as do^{18,19} the other acylating agents, *viz.* esters, amides, acids, and acyl fluorides. Perhaps, therefore, the excess acidity will prove generally valuable as a mechanistic tool with dioxane-water mixtures. Indeed, the patterns of the X-plots for 60% dioxane-water (Figs. 5 and 7) are reminiscent of those found for esters in aqueous acid.⁹ We have pointed out previously⁸ that, for any hydrolysis, dioxane-rich water solvents are more likely than water to reveal both acid catalysis, and changes in the mechanism of such catalysis. This is because when water contains a non-hydrogen bond donor (dioxane) as a second component there is (i) less hydrogen bonding available to reactants' leaving groups, *etc.*, and (ii) less water available for nucleophilic attack; both factors raise the relative importance of acid-catalysed routes, and these also benefit from the more rapid changes in acidity functions in the mixed medium.

References

- 1 See R. J. E. Talbot, in *Comprehensive Chemical Kinetics*, eds. C. H. Bamford and C. F. H. Tipper, Elsevier, Amsterdam, 1972, vol. 10, ch. 3.
- 2 G. Calvaruso and F. P. Cavasino, *J. Chem. Soc., Perkin Trans. 2*, 1971, 483.
- 3 C. A. Bunton and J. H. Fendler, *J. Org. Chem.*, 1965, **30**, 1365.
- 4 C. A. Bunton and S. G. Perry, *J. Chem. Soc.*, 1960, 3070.
- 5 C. A. Bunton and J. H. Fendler, *J. Org. Chem.*, 1966, **31**, 3764.
- 6 J. Koskikallio, *Acta Chem. Scand.*, 1964, **18**, 2248.
- 7 G. Calvaruso, F. P. Cavasino and E. D. Dio, *J. Chem. Soc., Perkin Trans. 2*, 1976, 993.
- 8 D. P. N. Satchell and W. N. Wassef, *J. Chem. Soc., Perkin Trans. 2*, 1992, 1855.
- 9 R. A. Cox and K. Yates, *Can. J. Chem.*, 1979, **57**, 2944.
- 10 A. I. Vogel, *Practical Organic Chemistry*, 3rd edn., Longmans, London, 1972.
- 11 See M. Ali and D. P. N. Satchell, *J. Chem. Soc., Perkin Trans. 2*, 1993, 917 and earlier papers.
- 12 C. A. Bunton, N. A. Fuller, S. G. Perry and V. J. Shiner, *J. Chem. Soc.*, 1963, 2918.

- 13 C. A. Bunton, J. H. Fendler, N. A. Fuller, S. G. Perry and J. Rocek, *J. Chem. Soc.*, 1963, 5361.
- 14 R. A. Cox and K. Yates, *J. Am. Chem. Soc.*, 1978, **100**, 3861.
- 15 C. H. Rochester, *Acidity Functions*, Academic Press, London, 1970.
- 16 R. A. Cox, *Acc. Chem. Res.*, 1987, **20**, 27.
- 17 D. P. N. Satchell, *J. Chem. Soc.*, 1957, 2878.
- 18 D. P. N. Satchell and R. S. Satchell, in *The Chemistry of Acid Derivatives. Supplement B*, ed. S. Patai, Wiley, Chichester, 1992, vol. 2, ch. 13.
- 19 R. E. Motie, D. P. N. Satchell and W. N. Wassef, *J. Chem. Soc., Perkin Trans. 2*, 1993, 859.

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