1024. *The Hydrolysis of Carboxylic Anhydrides. Part IV.*[1] *Succinic and Tetramethylsuccinic Anhydrides.*

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Hydrolysis of succinic and tetramethylsuccinic anhydride is retarded by aqueous mineral acids (except for dilute sulphuric and hydrochloric acid). It is acid-catalysed in aqueous dioxan. Hydrolysis of succinic anhydride in water is accelerated by bivalent electrolytes, but retarded by univalent electrolytes (except for sodium hydrogen sulphate). In water the feeble catalysis by protons is opposed by a specific negative salt effect, but the rate in aqueous acid is always greater than that in solutions of the alkali-metal salts.

Hydrolysis of open-chain anhydrides is acid-catalysed in water,[2] but to a much smaller extent than that of carboxylic esters, or amides in dilute acid. The early reports on the hydrolysis of cyclic anhydrides in aqueous acid were contradictory,[3] but recently Koskikallio and Errasti[4] have found that mineral acids generally retard hydrolysis, except when in high concentration, although they speed hydrolysis in aqueous dioxan. Retarding effects of acids upon the hydrolysis of succinic and tetramethylsuccinic anhydride are specific, and, in water, acid-catalysis does not overcome the negative salt effects of the ions (Tables 1 and 2.) Strong acids are catalysts in aqueous dioxan (Table 3) and it is convenient to discuss reactions in water and aqueous dioxan separately.

Water.—Most of the experiments were made on succinic anhydride; because of the low solubility of tetramethylsuccinic anhydride in water, we studied its hydrolysis only in neutral solution and in dilute perchloric acid (Table 1).

Table 1.

Hydrolysis in aqueous acid.

				Succinic anhydride.					
In perchloric acid.									
$[H^+]$ (M)	0	0·51	1·00	1·01	2·02	3·00		5·00	10·0
10^5k_1 (sec.$^{-1}$) at 0°	34·7	25·8	21·7	20·3	15·2	11·8	12·5	9·3	30·0
$[H^+]$ (M)	0	0·91	2·07	2·92	4·00				
10^5k_1 (sec.$^{-1}$) at 25·1°	269	218	149	116	120				
In sulphuric acid.									
$[H^+]$ (M)	0·09	0·32	0·44	0·91	1·87	3·03	4·55	7·1	
10^5k_1 (sec.$^{-1}$) at 0°	35·4	35·5	35·4	37·8	29·8	27·0	20·3	11·8	
$[H^+]$ (M)	0·05	0·25	0·50	1·00	1·50				
10^5k_1 (sec.$^{-1}$) at 25·1°	265	252	230	200	177				
In other acids at 0°.									
Acid	HCl					HBr	p-$C_7H_7 \cdot SO_3H$		
$[H^+]$ (M)	0·27	1·01	2·43	3·53	4·00	0·84	0·73		
10^5k_1 (sec.$^{-1}$)	35·7	35·3	30·8	26·5	24·6	30·2	31·8		
				Tetramethylsuccinic anhydride.					
In perchloric acid.									
$[H^+]$ (M)	0	0·116	0·542	1·01					
10^5k_1 (sec.$^{-1}$) at 25·05°	22·7	21·4	19·2	18·5					

Perchloric acid retards hydrolysis, except at high concentration (*ca.* 10M), where it is a major component of the solvent. Hydrobromic and toluene-*p*-sulphonic acid also retard

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[1] Part III, *J.*, 1963, 2918.

[2] Gold and Hilton, *J.*, 1955, 838, 843.

[3] Rivett and Sidgwick, *J.*, 1910, **97**, 1677; Wilsdon and Sidgwick, *J.*, 1913, **103**, 1959; Szabo, *Z. phys. Chem.*, 1926, **122**, 405.

[4] Koskikallio and Errasti, *Suomen Kem.*, 1962, **35***B*, 213.

hydrolysis, but at 0° hydrochloric and sulphuric acid in low concentration are feeble catalysts, although in high concentration they also retard hydrolysis (Table 1 and Fig. 1). Uni-univalent salts of strong acids (except sodium hydrogen sulphate) retard, and sulphates accelerate, hydrolysis. Therefore hydrolyses of cyclic anhydrides are apparently less catalysed by protons than those of open-chain anhydrides,[2, 5] and this catalysis may be less than a specific negative salt effect of the mineral acid.

Acid-catalysis.—The weak acid-catalysis was unexpected because acetic and succinic anhydride have similar reactivities in neutral solution,[1, 6] and might be expected to behave similarly in acid. There are several possible explanations for these differences and we attempted to exclude some of them experimentally. (*a*) Acid-hydrolyses of cyclic are slower than those of open-chain anhydrides because they are multi-step processes in which a readily formed intermediate, *e.g.*, (II) formed by attack of water upon the ion (I), preferentially reverts to reactants, whereas the corresponding intermediate formed from an open-chain anhydride can separate readily into two fragments. [The positions of the proton and the positive charge in (II) are assumed.] If this explanation were correct, we should expect return from ion (II) to be more important in the acid hydrolysis of tetramethylsuccinic than in that of succinic anhydride, because the methyl groups should assist cyclisation,[7] but both anhydrides behave similarly in aqueous acid (Table 1). Also reversion of an intermediate should be important in the spontaneous hydrolysis of cyclic anhydrides, because bond-breaking then involves formation of formally opposite charges. (*b*) In a variant of this hypothesis one supposes that acid-hydrolysis of an anhydride is a multi-stage reaction in which transfer of a proton from one oxygen atom to another is kinetically important (cf. refs. 8 and 9). Addition of water to an ion (I) could give an intermediate, *e.g.*, (III), which must rearrange, *e.g.*, to (IV), with proton transfer, before the

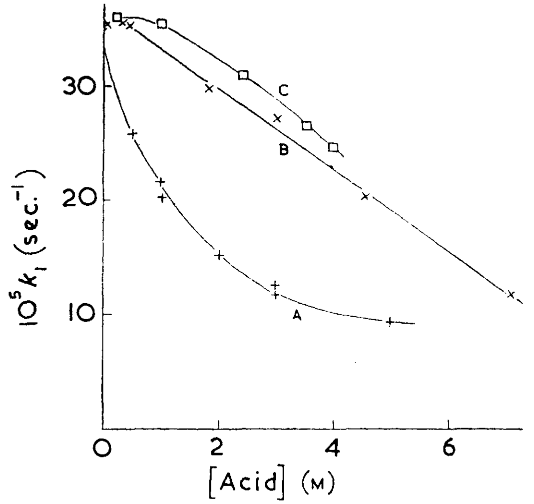
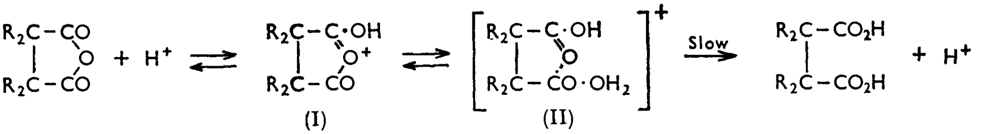


FIG. 1. Hydrolysis of succinic anhydride in aqueous acid at 0°.

(A) $HClO_4$. (B) H_2SO_4. (C) HCl.



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[5] Bunton and Perry, *J.*, 1960, 3070.

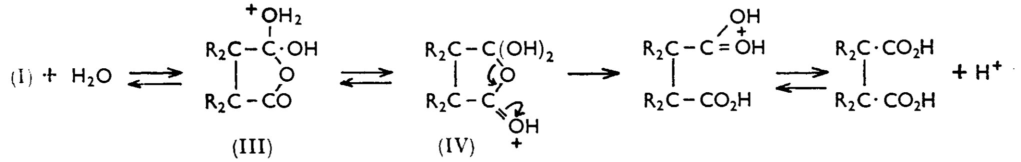
[6] Koskikallio, *Ann. Acad. Sci. Fennicae*, 1954, AII, **57**, 1.

[7] Ingold, Lanfear, and Thorpe, *J.*, 1923, **123**, 3140; Bruice and Pandit, *J. Amer. Chem. Soc.*, 1960, **82**, 5858.

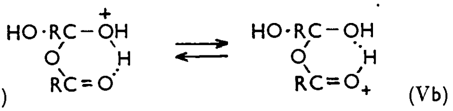
[8] Bunnett, *J. Amer. Chem. Soc.*, 1961, **83**, 4956 and accompanying papers.

[9] Martin, *J. Amer. Chem. Soc.*, 1962, **84**, 4130.

carbon–oxygen bond can break. This proton transfer (III) ⟶ (IV) may require intervention by one or more water molecules,[8,9] and may therefore be inhibited by electrolytes which decrease the activity of water,[10] whereas it could occur intramolecularly during acid-hydrolysis of an open-chain anhydride, by formation of structures such as


(III) (IV)

(Va, b), which cannot be formed readily from cyclic anhydrides. (*c*) In explanations (*a*) and (*b*) one supposes that the transition states for acid hydrolysis of cyclic and open-chain anhydrides are structurally different, whereas they may differ only in the affinity of the


(Va) (Vb)

oxygen atoms towards the proton. Sulphuric acid not only protonates but also solvolyses benzoic anhydride, but phthalic anhydride is unaffected,[11] and therefore it must be less basic than the open-chain anhydrides. Conversely, hydrolysis of phthalic anhydride in aqueous dioxan is not acid-catalysed,[12] but that of benzoic acid is.[5] If the structural factor responsible for the low basicity of the cyclic anhydrides is present also in the transition state, these compounds may be sufficiently insensitive to hydrogen ions for a negative salt effect of the acid to overcome any catalysis.

The structure of the conjugate acid of an anhydride is unknown, but analogy with other carboxyl derivatives [13] suggests that the proton is attached to the carbonyl-oxygen atom, and that the more important canonical forms are (VIa, b):


(VIa) (VIb)

$$\text{(VIa)}\quad R\cdot\overset{\overset{\textstyle OH}{|}}{C}{=}\overset{+}{O}\cdot CO\cdot R \longleftrightarrow R\cdot\overset{\overset{\textstyle \overset{+}{O}H}{\|}}{C}\cdot O\cdot CO\cdot R\quad\text{(VIb)}$$

i.e., protonation of the carbonyl-oxygen atom is assisted by electron-release from an ethereal oxygen atom. Cyclic anhydrides should therefore be less basic than open-chain anhydrides, because the change in the hybridisation of the ethereal oxygen atom from sp^3 in the anhydride to sp^2 in (VIa) will increase the bond angle of this atom and be energetically more costly in a rigid cyclic than in a flexible open-chain anhydride.

Hydrolyses of cyclic phosphates and sulphates afford systems in which cyclisation suppresses electron-release from an ethereal oxygen atom, by inhibiting an increase in its covalency. These reactions are faster than those of the corresponding open-chain compounds, in part because of their instability relative to the open-chain esters.[14] It is suggested that electron-release from the ethereal oxygen atom stabilises the open-chain esters, *e.g.*, (VIIa), but cannot occur in the cyclic esters, *e.g.*, (VIIb).

Suppression of this electron-release destabilises the initial state (VIIb), but not the transition state in which the ring has opened.

In considering a rate, rather than an equilibrium constant, the important factor is the free-energy difference between initial and transition states. Basicity only relates the initial state to its conjugate acid. Therefore, if the inability of the oxygen atom to accept

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[10] Robinson and Stokes, " Electrolyte Solutions," Butterworths, London, 1959, Appendix 8.

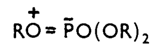
[11] Leisten, *J.*, 1961, 2191.

[12] Perry, Thesis, London, 1958.

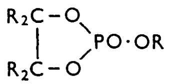
[13] Stewart and Yates, *J. Amer. Chem. Soc.*, 1960, **82**, 4059; Katritsky, *Chem. and Ind.*, 1961, 722.

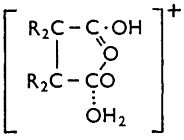
[14] Kumamoto, Cox, and Westheimer, *J. Amer. Chem. Soc.*, 1956, **78**, 4858; Haake and Westheimer, *ibid.*, 1961, **83**, 1102; Kaiser, Panar, and Westheimer, *ibid.*, 1963, **85**, 602.

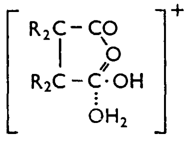
a proton is the reason for the feeble acid-catalysis of the cyclic anhydrides, it follows that the conformations of the transition state and the conjugate acid must be similar, and bond-making must therefore be more important than bond-breaking, *i.e.*, the transition state should be represented as (VIIIa or b).


$\overset{+}{R}O=\overset{-}{P}O(OR)_2$

(VIIa)


(VIIb)


(VIIIa)


(VIIIb)

If bond-breaking had made considerable progress in the transition state there is no reason why the affinity of the oxygen atoms for a proton should be different in the transition states for acid-hydrolysis of cyclic and open-chain anhydrides.

Our experiments eliminate explanation (*a*), but do not differentiate between (*b*) and (*c*). We shall attempt to detect intermediates, *e.g.*, (IV), by the use of oxygen-18.

We have neglected the possibility of an *A*-1 mechanism of hydrolysis because succinic is more reactive than tetramethylsuccinic anhydride in both neutral and acid solution. The methyl groups should hinder nucleophilic attack, but assist heterolysis by releasing electrons.[15]

Salt Effects.—Uni-univalent electrolytes generally retard the spontaneous hydrolysis of both cyclic and open-chain anhydrides (Table 2 and refs. 6 and 16), usually by stabilising the initial and destabilising the transition state.[16] The rates of hydrolysis of succinic anhydride in aqueous acids are in the same sequence as those in neutral solutions of the uni-univalent salts (Table 2 and ref. 6). Sulphates speed the spontaneous hydrolysis. Therefore, neutral salts affect the hydrolysis of succinic and acetic anhydride similarly. Fluoride ions strongly catalyse the hydrolysis of succinic anhydride,[17] probably by forming a reactive acyl fluoride,[16] and $10^5 k_1 = 125$ sec.$^{-1}$ in 0·06M-sodium fluoride at 0°.

TABLE 2.

Hydrolysis of succinic anhydride in aqueous salt solutions, at 0° unless otherwise specified.

Salt	LiCl	NaCl		Me_4NCl	NaBr	$LiClO_4$			
[Salt] (M)	3·0	1·0	4·14	1·0	1·0	1·0	2·0	3·0	4·0
$10^5 k_1$ (sec.$^{-1}$)	12·2	24·9	6·89	34·6	22·7	18·9	7·40	4·21	15·3 *

Salt	$NaClO_4$					$NaHSO_4$		
[Salt] (M)	1·0	2·0	4·0		6·0	0·5	1·0	1·25
$10^5 k_1$ (sec.$^{-1}$)	19·1	10·9	3·84	20·2 *	1·33	36·9	36·5	36·8

Salt	Me_4NHSO_4			Na_2SO_4	$(Me_4N)_2SO_4$	$MgSO_4$	
[Salt] (M)	0·5	1·0	2·0	0·5	0·5	0·5	1·0
$10^5 k_1$ (sec.$^{-1}$)	32·9	30·2	21·6	40·3	53·6	40·7	51·2

Values of *b* (values for acetic anhydride [6,16] in parentheses) are: NaCl, KCl 0·17 (NaCl 0·12); $NaClO_4$ 0·24 (0·24); LiCl 0·15; $LiClO_4$ 0·31; $MgSO_4$ −0·07 (−0·1).

* At 25·1°.

The variation of rate constant with ionic strength fits equation (1) reasonably for

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[15] Ingold, " Structure and Mechanism in Organic Chemistry," Bell, London, 1953, p. 770.

[16] Bunton, Fuller, Perry, and Pitman, *J.*, 1962, 4478.

[17] Fuller, unpublished results.

neutral uni-univalent salts of strong acids in initially neutral solution (except for sodium fluoride):

$$\log k_1/k_0 = b\mu, \tag{1}$$

where k_1 and k_0 are the first-order rate constants in the presence and absence, respectively, of salt, and μ is the ionic strength. The values of b (Table 2) are similar to those observed for the spontaneous hydrolysis of acetic anhydride[16] (in parentheses).

Sodium hydrogen sulphate behaves differently from other uni-univalent salts, in that it speeds hydrolysis. The rate levels off at concentration >M; we do not know whether there is a maximum rate at this concentration, because in higher concentration both sodium sulphate and hydrogen sulphate strongly "salt-out" succinic anhydride at 0°, *i.e.*, destabilise the initial state. The hydrogen sulphate ion is a moderately strong acid[18] ($K_2^{25} = 0{\cdot}01$), and sulphate ions formed by its ionisation could speed hydrolysis. The explanation is incomplete, because tetramethylammonium hydrogen sulphate retards hydrolysis at all concentrations (Table 2) and it does not "salt out" the anhydride as effectively as does the sodium salt, and we could therefore use it in comparatively high concentration. The behaviour of sodium hydrogen sulphate is complicated, and we do not understand it. Its accelerating effect at low concentration is probably caused by destabilisation of the initial state.[16]

Hydrolysis in Acid–Salt Mixtures.—Experiments in acid alone do not tell us whether protons catalyse the hydrolysis of cyclic anhydrides in water, although they do so in aqueous dioxan (Table 3). However, the rate in aqueous acid is always greater than in an equivalent salt solution. If the specific salt effect of the hydronium ion is similar to those of lithium and sodium ions, the rate of the uncatalysed hydrolysis will be approximately constant in mixtures of acid and salt of constant ionic strength.[19] Any increase of rate with increasing acid concentration should then be caused by acid-catalysis. The ionic strength of solutions of hydrochloric or perchloric acid was therefore kept constant by addition of the corresponding lithium or sodium salt (Table 4). Plots of k_1 against

TABLE 3.

Hydrolysis in 60 : 40 (v/v) dioxan–water, with perchloric acid unless specified.

Succinic anhydride.

[H+] (M)	0		0·50		0·51	1·00		1·50	2·00	3·00	5·00	5·48	6·00
$10^5 k_1$ (sec.$^{-1}$) at 0°	1·22	0·28 †	3·60	2·32 *	3·58	4·12	2·89 *	4·60	5·09	5·60	10·2	13·2	21·5

[H+] (M)	0		0·52	1·00	1·08	1·50	2·00		3·00
$10^5 k_1$ (sec.$^{-1}$) at 25·1°	11·0	0·117 ‡	35·1	38·4 †	51·2	61·5	68·4	81·2 *	86·4

Acid	HCl			p-$C_7H_7 \cdot SO_3H$		
[H+] (M)	1·03	2·52	4·02	0·38	0·49	0·95
$10^5 k_1$ (sec.$^{-1}$) at 0°	5·21	11·3	25·9	2·55 §	3·70	5·40

Salt	$LiClO_4$				p-$C_7H_7 \cdot SO_3Na$	
[Salt] (M)	0·43	1·00	2·00	3·00	0·25	0·5
$10^5 k_1$ (sec.$^{-1}$) at 0°	1·18	0·96	0·28	11·7 ¶	1·24	1·01

Tetramethylsuccinic anhydride.

[H+]	0		0·50	1·00	1·08	1·94	2·00	3·11
$10^5 k_1$ (sec.$^{-1}$) at 25·05°	1·18	0·40 ‡	2·10	3·63 †	2·94	5·50	6·52 *	8·66

* M-$LiClO_4$. † 2M-$LiClO_4$. ‡ 3M-$LiClO_4$. § 0·4M-p-$C_7H_7 \cdot SO_3Na$. ¶ At 25·05°.

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[18] Ref. 10, p. 387.

[19] Long, McDevit, and Dunckle, *J. Phys. Chem.*, 1951, **55**, 819; Bunton, Llewellyn, Oldham, and Vernon, *J.*, 1958, 3574.

TABLE 4.

Hydrolysis at constant electrolyte concentration in aqueous acid, at 0° unless specified.

Succinic anhydride.

[HCl] + [NaCl] = 2M.

$[H^+]$ (M)	0	1·00	2·00
10^5k_1 (sec.$^{-1}$)	15·9 *	24·2	32·6 *

[HCl] + [NaCl] = 3M; values for added LiCl in parentheses.

$[H^+]$ (M)	0		1·00		2·00		3·00
10^5k_1 (sec.$^{-1}$)	11·5	(11·9)	17·5	(17·6)	21·6	(21·2)	28·5 *

[HCl] + [NaCl] = 4M.

$[H^+]$ (M)	0	1·00	2·00	3·00	4·00
10^5k_1 (sec.$^{-1}$)	7·0 *	12·2	16·1	19·5	24·6

[HClO$_4$] + [NaClO$_4$] = 3M; values for added LiClO$_4$ in parentheses.

$[H^+]$ (M)	0	1·00		2·00	3.00
10^5k_1 (sec.$^{-1}$)	6·4 *	8·01	(8·49)	(10·2)	11·8

[HClO$_4$] + [NaClO$_4$] = 4M.

$[H^+]$ (M)	0	1·00	1·50	3·03	4·00
10^5k_1 (sec.$^{-1}$)	3·84	5·61	6·16	8·75	10·0

[HClO$_4$] + [NaClO$_4$] = 6M.

$[H^+]$ (M)	0	1·22	2·43	6·00
10^5k_1 (sec.$^{-1}$)	1·33	3·00	4·59	11·0

[HClO$_4$] + [LiClO$_4$] = 4M, at 25·1°; value for added NaClO$_4$ in parentheses.

$[H^+]$ (M)	0		1·00	2·00		3·00	4·00
10^5k_1 (sec.$^{-1}$)	15·3	(20·2)	43·1	73·5	(79·5)	86·3	117, 124

Initial slopes of plots of log k_1 against [HCl] at 0°, from Fig. 2.

μ	1·0	2·0	3·0	4·0
10^5 slope (l. mole^{-1} sec.$^{-1}$)	11·5	8·3	5·5	4·4

Tetramethylsuccinic anhydride, at 25·05°.

[HClO$_4$] + [LiClO$_4$] = M.

$[H^+]$ (M)	0	0·50	1·00
10^5k_1 (sec.$^{-1}$)	13·8	17·0	18·5

* Interpolated.

acid concentration, in molarity, were then approximately linear (Fig. 2). This linearity has no great significance, because it depends on the scale used for concentration, *e.g.*, the plots would not be linear if we had expressed concentration in molality. Lithium and sodium chloride behave similarly, in the presence and absence of acid, but the perchlorates do not. Perchloric acid salts out its lithium salt in water, and it is difficult to obtain solutions of ionic strength greater than 3M, suggesting that acid and salt interact in some way. The slopes of the plots in Fig. 2 are greater for the mixtures of chlorides than for mixtures of perchlorates, and decrease with increasing ionic strength, although the effect is small for the perchlorates. The slope of the plot of k_1 against concentration of perchloric acid at 4M ionic strength is greater at 25° than at 0°, relative to the rate constant in initially neutral solution (from the values in Table 4). This, and the observation that the rate minimum occurs at lower acid concentration at 25° than at 0° (Table 1), show that the acid-hydrolysis has a higher activation energy than the spontaneous hydrolysis. Because the negative salt effect of a proton may differ from that of a lithium or sodium ion we cannot identify the slopes of the plots in Fig. 2 directly with the second-order rate constant for acid-hydrolysis. But we can use the values of these slopes to obtain an approximate value for the rate constant of hydrolysis catalysed by hydrochloric acid, $k_A{}^\circ$, by assuming that the initial slopes give values of the second-order rate constants, k_A, at various ionic strengths, and that equation (2) is followed:

$$\log k_A/k_A^\circ = -a\mu. \qquad (2)$$

The value of k_A° is then given by extrapolating a plot of log k_A against μ to the ordinate (values of log k_A and μ are in Table 4). This plot is only approximately linear, curving at higher ionic strengths, but $10^5 k_A^\circ \approx 16{\cdot}1$ l. mole^{-1} sec.$^{-1}$. Extrapolation of k_A to zero ionic strength removes the uncertainties caused by the different negative salt effects of an acid and its salt.

The second-order rate constant, k_A°, for the acid-hydrolysis of succinic anhydride is smaller than that for acetic anhydride. It is not easy to estimate the value of k_A° for acetic anhydride, because at low acid concentration the spontaneous hydrolysis makes a large contribution and at higher acid concentration the rate increase is proportional to a high, variable power of the acid concentration.[2] The value, in aqueous perchloric acid, at 0° is *ca.* 4×10^{-4} l. mole^{-1} sec.$^{-1}$ (from the spontaneous rate and that in M-acid), or *ca.* 5×10^{-4} l. mole^{-1} sec.$^{-1}$ (from the slope at M-acid of a plot of first-order rate constant against acid concentration). These values are underestimates, because the negative salt effect upon the spontaneous hydrolysis is neglected. The second-order rate constant, calculated at higher acidity, would be much larger.[2] The limited information on the acid hydrolysis of tetramethylsuccinic anhydride in mixtures of perchloric acid and lithium perchlorate (Table 4) suggests that it behaves similarly to succinic anhydride.

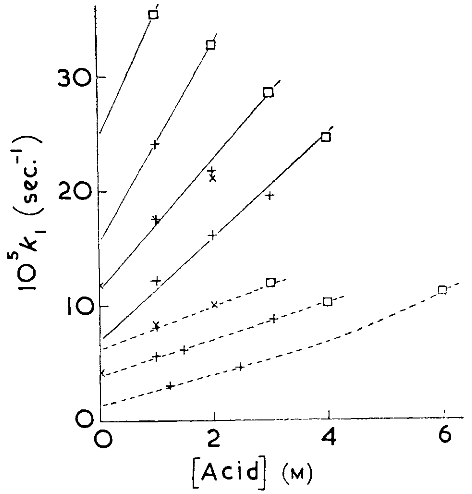

FIG. 2. Hydrolysis of succinic anhydride in aqueous acidic solutions of constant electrolyte concentration.

□ Acid only. + Na salt added. × Li salt added. Broken line, perchlorates. Full line, chlorides.

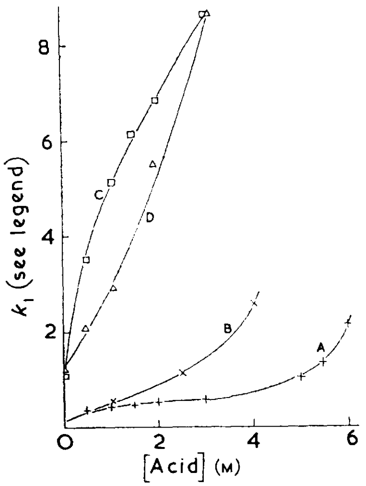

FIG. 3. Hydrolysis in acidic aqueous dioxan.

Succinic anhydride (k_1 in 10^{-4} sec.$^{-1}$): (A) $HClO_4$ at 0°; (B) HCl at 0°; (C) $HClO_4$ at 25·1°.

Tetramethylsuccinic anhydride (k_1 in 10^{-5} sec.$^{-1}$): (D) $HClO_4$ at 25·1°

These effects of acids and salts upon the hydrolysis of succinic anhydride are similar to those observed by Marburg and Jencks for the hydrolysis of the conjugate acid of 1-acetylimidazole (IX).[20] Except for fluorides, uni-univalent salts retard hydrolysis, and equation (1) is followed for the salts, but not for acids in high concentration. The

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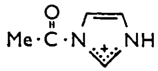
[20] Marburg and Jencks, *J. Amer. Chem. Soc.*, 1962, **84**, 232.

rate in an acid is faster than in its salt solutions (except for perchloric acid), and there may therefore be some slight acid-catalysis involving the diprotonated species (X), although the behaviour of perchloric acid and its salts does not fit this hypothesis, and acids, including sulphuric, retard hydrolysis in all concentrations.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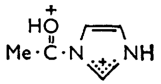
(X)

Rates in Aqueous Dioxan.—Perchloric and hydrochloric acid do not retard the hydrolysis of succinic and tetramethylsuccinic anhydride in dioxan–water (Table 3 and Fig. 3), and the negative salt effects are considerably smaller than in water. Two factors contribute to these differences.

(i) The decrease in the negative salt effect in going from water to aqueous dioxan conflicts with the electrostatic theory, which predicts an increase in salt effect with decreasing dielectric constant.[21] However, the specificity of these salt effects illustrates the shortcomings of simple electrostatic theory. The effect of electrolytes upon the activity coefficients of non-electrolytes in water depends largely upon the changes in the interactions between water and non-electrolyte which reflect the water–electrolyte interactions.[22] These salt effects should, therefore, become less important when an organic diluent is added to the water, because the non-electrolyte will now be appreciably solvated by the organic component of the solvent rather than by the water. Therefore the salt effect of the acid should also be less in the aqueous organic solvent than in water.

(ii) The different behaviour of these cyclic anhydrides in acidic water and aqueous dioxan shows that the spontaneous hydrolysis is more sensitive to the water content of the solvent than is the acid hydrolysis (cf. ref. 4). This is understandable if much of the driving force for the spontaneous hydrolysis comes from hydrogen bonding between water and the reactants, and if protonation provides much of the driving force for acid hydrolysis; and in aqueous dioxan catalysis by hydronium ions may overcome any negative salt effect.

The rate in aqueous dioxan does not vary linearly either with the acid concentration or with the protonating power as measured by Hammett's acidity function,[23] h_0, although the rate is approximately proportional to acid concentration at 25·1° (Fig. 3). As in water, hydrochloric is a better catalyst than perchloric acid. Acid-catalysis of succinic anhydride is much more evident at 25° than at 0° (Fig. 3), suggesting that, as in water (p. 5366), the activation energy is higher for the acid than for the spontaneous hydrolysis. At 0°, in perchloric acid, the initial rate increase with small amounts of acid is followed by a plateau, from 1·5M- to 3M-perchloric acid; the rate then increases sharply (Fig. 3). There is no plateau in hydrochloric acid. Added lithium perchlorate retards the spontaneous hydrolysis, but its effect upon the acid-catalysed hydrolysis is small, particularly at 25°; it decreases the overall rate of hydrolysis in dilute acid and increases it in more concentrated acid (Table 3).

The rates of the spontaneous hydrolysis are sufficiently small, relative to those of the acid-catalysed reactions, for the salt effect of acid to be neglected, and the second-order rate constants (l. mole^{-1} sec.$^{-1}$) for the acid-hydrolyses at 25·1° are: succinic 43×10^{-5}; tetramethylsuccinic, $1·9 \times 10^{-5}$; acetic, 380×10^{-5}. As in water, succinic is less sensitive than acetic anhydride to strong acids.

EXPERIMENTAL

Materials.—The preparation and purification of the anhydrides has been described.[1] The solvent was distilled water, or aqueous dioxan made up by weight so that after addition of acid the solvent composition was 60 vol. of dioxan and 40 vol. of water. The salts were recrystallised

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[21] Ref. 16, p. 362; Frost and Pearson, " Kinetics and Mechanism," Wiley, New York, 1961, Ch, 7.

[22] Long and McDevit, *Chem. Rev.*, 1952, **51**, 119.

[23] Bunton, Ley, Rhind-Tutt, and Vernon, *J.*, 1957, 2327.

specimens, dried over phosphoric oxide *in vacuo* in a drying pistol, except for tetramethylammonium sulphate and the hydrogen sulphates which were prepared in solution by neutralisation of sulphuric acid with the base.

Kinetics.—Reaction was followed by adding aniline to portions of the reaction mixture and titrating the acid produced, or by dilatometry.[1,5] Dilatometry was generally used for the runs in more concentrated acid solution. The concentration of succinic anhydride in water was *ca.* M/20, and of tetramethylsuccinic anhydride *ca.* M/50. The anhydrides were more soluble in aqueous dioxan, and higher concentrations were used. The integrated first-order rate constants, k_1, were estimated graphically.

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