750. *Organic Carbonates. Part VIII.*[1] *The Acid-catalysed Hydrolysis of Ethylene, Trimethylene, and Tetramethylene Carbonate.*

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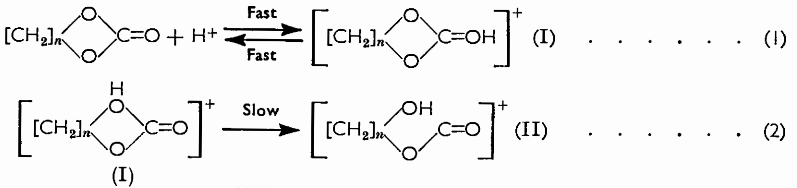
The rates of the acid-catalysed hydrolysis of ethylene, trimethylene, and tetramethylene carbonate have been measured at two temperatures, and the energies and entropies of activation have been calculated. In the hydrolysis of propylene carbonate in >1 M-hydrochloric acid the rates are not proportional to the acidity measured by Hammett's acidity function, h_0. The large negative entropies of activation indicate that all the cyclic carbonates are hydrolysed by the same *A*-2 mechanism. The relative order of reactivities for acid-catalysed hydrolysis of homologous cyclic carbonates at temperatures investigated are 6- $>$ 5- $>$ 7-membered.

In Part V[2] it was shown that hydrolyses of cyclic carbonates in acid solution proceed with carbonyl–oxygen cleavage at both stages of the reaction. The present paper records a kinetic study of the mechanism of this reaction.

In contrast to the alkaline hydrolysis of cyclic carbonates discussed in Part VI,[3] the acid-catalysed reaction is very slow. Possible stages in the reaction are:



$$\underset{\text{(II)}}{[HO\cdot(CH_2)_n\cdot O\cdot CO]^+} \xrightarrow[\text{Fast}]{H_2O} \underset{\text{(III)}}{HO\cdot[CH_2]_n\cdot O\cdot CO_2H} + H^+ \qquad (3)$$

$$\text{(I)} + H_2O \xrightarrow{\text{Slow}} \text{(III)} + H^+ \qquad (4)$$

$$\text{(III)} \xrightarrow{\text{Fast}} \text{Glycol} + CO_2 \qquad (5)$$

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[1] Part VII, Pohoryles, Wislicki, and Sarel, *J. Pharm. Sci.*, 1962, **51**, 348.

[2] Sarel, Levin and Pohoryles, *J.*, 1960, 3079.

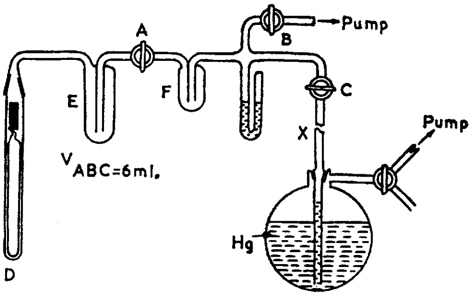
[3] Pohoryles, Levin, and Sarel, *J.*, 1960, 3082.

The purpose of this study was to establish whether the *A*-1 mechanism, steps (1), (2), and (3), or the alternative *A*-2 mechanism, steps (1) and (4), is operative. In contrast to the former, which would involve the formation of a carbonium ion (II), the rate-determining step in the latter, steps (1) and (4), would be concerned with a specific water molecule in the transition state. This difference should be reflected in a difference in the entropies of activation and the h_0 correlation for the two mechanisms, as discussed by Long, Pritchard, and Stafford.[4] We have therefore calculated entropies of activation for the hydrochloric acid-catalysed hydrolysis of ethylene, trimethylene, and tetramethylene carbonate, and the second-order rate-coefficients of propylene carbonate in hydrochloric acid over a range of 0·2—4M. We have also determined the rates of hydrolysis of 2,2-diethyltrimethylene carbonate in water and in deuterium oxide to see whether the reaction proceeds by a pre-equilibrium proton-transfer, step (1).

EXPERIMENTAL

The preparations and source of materials have been described in Part IV.[5] The acids were of analytical grade.

Hydrolysis in Acid Solution.—Solutions for runs were prepared by weighing the desired amount of ester into a 25-ml. volumetric flask to which were added 10 ml. of water. To this, at 0° were added 15 ml. of hydrochloric acid solution. The concentration of ester was 0·0—3 0·04M and that of hydrochloric acid 0·05—3M. Immediately after being shaken, 1-ml. samples of the reaction mixtures were delivered into a series of glass ampules which were then sealed and placed in a thermostat-bath. The extent of reaction was determined from the liberation of carbon dioxide, by withdrawing ampules at intervals and measuring the amount of carbon dioxide as described below.



Apparatus for measuring minute amounts of carbon dioxide. Tube X has internal diameter 6 mm.

Before its estimation, the liberated carbon dioxide was dried and purified by precooling first at −80°, and then at liquid-air temperature. This was performed by the use of an apparatus designed for the measurement of quantities as small as 10^{-4}—10^{-5} mole of carbon dioxide.

Apparatus.—The apparatus is represented in the Figure. All parts are made of Pyrex glass. The tube which is vertically connected to C must not be longer than 60 cm. and may occasionally be used for increasing the standard volume. The standard volume of the apparatus is that volume confined by the stopcocks A, B, and C.

Calibration.—This was performed by measuring the changes in pressure with that of volume of a gas occupying the standard volume of the apparatus by using the Boyle–Mariotte equation.

Rate of Liberation of Carbon Dioxide.—The ampules withdrawn from the constant-temperature bath were immediately cooled and then transferred into tube D of the apparatus along with a magnetised iron rod. After cooling of trap E, at −80°, the whole system was evacuated and then stopcocks A, B, and C were closed. The ampoule was broken with the aid of an

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[4] Long, Pritchard, and Stafford, *J. Amer. Chem. Soc.*, 1957, **79**, 2362.
[5] Sarel, Pohoryles, and Ben-Shoshan, *J. Org. Chem.*, 1959, **24**, 1873.

external magnet and the liberated carbon dioxide was quantitatively collected in the liquid-air trap F after opening of stopcock A. At this stage, the excess of air was pumped out through stopcock B and, after A and B had been closed, the entrapped carbon dioxide was allowed to warm to thermal equilibrium with the surroundings. The amount of carbon dioxide was calculated from the expression $m = MPV/\boldsymbol{R}T$, where m, M, and V are, respectively, the mass, molecular weight, and volume of carbon dioxide, P is the pressure in mm. mercury, T temperature in °K, and $\boldsymbol{R}$ the gas constant (62,360 mm. cm.^{3} mole^{-1} degree^{-1}).

Calculation of Rate Constants.—Since the kinetic form taken by the reaction was $d(CO_2)/dt = k[\text{ester}][H_3O^+]$, the first-order rate-coefficients were calculated by the following expression, derived from the integrated form of the previous one for gaseous systems:

$$k_1 = \frac{2{,}303}{t} \log_{10} \frac{(PV)_\infty}{(PV)_\infty - (PV)_t},$$

where $(PV)_t$ is the pressure–volume product of carbon dioxide liberated after t min., and $(PV)_\infty$ is the volume–product obtained after all the cyclic carbonate has been converted into the corresponding diol. One mole of carbonate being assumed to liberate one mole of carbon dioxide, $(PV)_\infty$ is a measure of the original ester concentration, and $(PV)_\infty - (PV)_t$ gives the quantity of ester remaining after time t. The values of second-order rate coefficients were calculated by dividing the first-order rate-coefficient, k_1, by $[H^+]$, the concentration of the hydrochloric acid at the reaction system. The second-order rate-coefficients, $k_1/[H^+]$, contain the values of the reaction rates of equation (4), k, multiplied by the equilibrium constant, K_{eq}, of step (1). It is plausible to assume that the values of K_{eq} vary little with temperature,[6] so that the temperature coefficients of reaction rates measured are actually those of k. The energy ($\Delta E^\ddagger$) and the entropy ($\Delta S^\ddagger$) of activation were calculated from formula $k = (e\boldsymbol{k}T/\boldsymbol{h}) \exp (\Delta S^\ddagger/\boldsymbol{R}) \exp (-\Delta E^\ddagger/\boldsymbol{R}T)$. The Tables illustrate typical kinetic runs.

TABLE 1.

Acid-hydrolysis of ethylene carbonate with 0·229M-hydrochloric acid at 52°.

Time (min.)	265	340	547	970	1020	1200	1500	∞	Average
$(PV)_t$	48	71	104	160	167	197	234	520	—
10^4K_1 (min.$^{-1}$)	3·66	4·28	4·08	3·78	3·78	3·96	1·98	—	—
$10^3K_1/[H^+]$ (l. mole^{-1} min.$^{-1}$)	1·60	1·87	1·78	1·65	1·65	1·73	1·74	—	1·72

TABLE 2.

Hydrolysis of trimethylene carbonate with 0·0459M-hydrochloric acid at 51·7°.

Time (min.)	10	19	30	50	59	64	88	93	∞
$(PV)_t$	37·5	72	102	172	192	210	270	263	550
10^3k_1 (min.$^{-1}$)	6·99	7·34	6·82	7·42	7·28	7·51	7·66	6·97	—
$10 . k_1/[H^+]$ (l. mole^{-1} min.$^{-1}$)	1·52	1·60	1·48	1·62	1·59	1·64	1·67	1·52	—

Average: $10^3k_1 = 7·25$ min.$^{-1}$; $10^2k_1/[H^+] = 15·8$ l. mole^{-1} min.$^{-1}$.

TABLE 3.

Hydrolysis of tetramethylene carbonate with 2·967M-hydrochloric acid in 1 : 3 dioxan–water at 50·7°.

Time (min.)	87	25	285	345	400	450	480	∞
$(PV)_t$	33	72	84	105	126	129	141	302
10^4k_1 (min.$^{-1}$)	13·25	12·06	11·43	12·35	13·45	12·32	13·05	—
$10^4k_1/[H^+]$ (l. mole^{-1} min.$^{-1}$)	4·46	4·06	3·85	4·16	4·53	4·15	4·40	—

Average: $10^4k_1 = 12·55$ min.$^{-1}$; $10^4k_1/[H^+] = 4·23$ l. mole^{-1} min.$^{-1}$.

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[6] (*a*) Harned and Owen, " The Physical Chemistry of Electrolytic Solutions," 3rd edn., Reinhold Publ. Corp., New York, 1958; (*b*) Ingold, " Structure and Mechanism in Organic Chemistry," Bell, London, 1955.

Table 4.

Rate constants and related parameters for the open-chain and homologous acid carbonates.

Carbonate	Temp.	$10^3k_1/[H^+]$ (l. mole^{-1} min.$^{-1}$)	Relative rate at 52°	$\Delta E\ddagger$ (kcal. mole^{-1})	$\Delta S\ddagger$ e.u.
Dimethyl *	52	0·35	1·0	16·8	−33·4
	71·6	16·4			
Ethylene	52	1·73	5·0	17·9	−26·8
	72·8	9·63			
Trimethylene	32·3	38·5	457	13·5	−31·6
	51·7	158·0			
Tetramethylene †	50·7	0·42	1·5	20·4	−21·7
	69·5	2·49			

* Work done by J. Katzhandler and T. Bino. † Measured in 1 : 3 dioxan–water with 2·96M-hydrochloric acid.

Table 5.

Rate constants and related data for hydrolysis of propylene carbonate in water at 52°.

$[H^+]$ (M)	10^4k_1 (min.$^{-1}$)	$10^4k_1/[H^+]$ (l. mole^{-1} min.$^{-1}$)	H°
0·2	1·07	5·40	+0·67
1·0	4·77	4·77	−0·20
2·0	10·64	5·32	−0·69
3·0	13·31	4·43	−1·05
3·96	16·11	4·07	−1·40

RESULTS AND DISCUSSION

When the logarithms of the rates of acid-hydrolysis of propylene carbonate are plotted against values[7] of Hammett's acidity function, in its logarithmic form, H_0, a curved line is obtained. This result strongly suggests that the reaction is not Hammett-dependent. On the other hand, Table 5 indicates that the reaction is more likely to be concentration-dependent, but that the second-order constants, in the form of $k_1/[H^+]$, fall as the concentration of hydrochloric acid is increased. The decrease is nearly $\frac{1}{4}$-fold over a range 0·2—4M.*

The hydrolysis of 2,2-diethyltrimethylene carbonate (*S*) was found to be 1·26-fold faster in heavy water than in water.† The second-order rate-coefficients, $10^2k_1/[H^+]$, at 32° are: 2·84 l. mole^{-1} min.$^{-1}$ in water and 3·59 l. mole^{-1} min.$^{-1}$ in 100% deuterium oxide. The faster reaction in D_2O follows from the greater concentration of SD^+ ions in D_2O than of SH^+ in H_2O, and from the smaller dissociation constant of D_2O than of H_2O.[10] This provides evidence that the reaction involves a rapid pre-equilibrium proton-transfer stage, as formulated in reaction (1). The position of the proton in the intermediate (I) is unknown. It might be associated either with the carbonyl group or with ring-oxygen atoms.

Considering the latter and the results in Tables 4 and 5 makes it clear that acid-catalysed hydrolysis of all carbonic acid esters, whether cyclic or non-cyclic, occurs by the *A*-2 mechanism, involving steps (1), (4) and (5), and can be labelled $A_{AC}2$. The high negative entropy of activation shown by all carbonates, cyclic and open-chain, accord with this mechanism.[4]

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* In this respect cyclic carbonates show resemblance to γ- and δ-lactones, but they differ from cyclic sulphites.[8]

† A greater value, 1·40, was recorded for the open-chain diethyl carbonate by Wynne-Jones.[9] This value requires checking since it was deduced from data concerning the acid-catalysed hydrolysis of tetraethyl orthocarbonate in D_2O and in H_2O.

[7] Long and Paul, *Chem. Rev.*, 1957, **57**, 1.

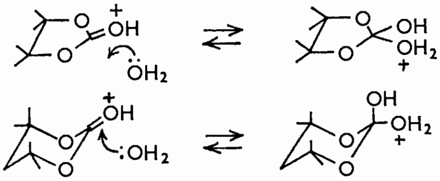
[8] Bunton, de la Mare, and Tillett, *J.*, 1958, 4754.

[9] Wynne-Jones, *Trans. Faraday Soc.*, 1938, **34**, 245.

[10] McIntyre and Long, *J. Amer. Chem. Soc.*, 1954, **76**, 3240; Wiberg, *Chem. Rev.*, 1955, **55**, 713.

The last stage of the reaction, step (5), is indeed very rapid. This is shown by the rates of acid-hydrolysis of alkyl hydrogen carbonates, $RO{\cdot}CO_2H$. The first-order rate-coefficients (min.$^{-1}$) at 0° are:[11] ethyl, 14·0; n-propyl, 10·9; isopropyl, 12·7; n-butyl, 17·3; isobutyl, 16·1; s-butyl, 17·3. By extrapolating the results for ethylene carbonate, the value of 10^{-5} min.$^{-1}$ is obtained for acid-hydrolysis of this compound in water at 0°. This permits the conclusion that the covalency change in step (4) occurs about 10^6 times slower than in step (5), and consequently that step (4) determines the rate of displacement of carbon dioxide.

Structural Effects.—The values for $\Delta E^\ddagger$ and $\Delta S^\ddagger$ (Table 4) illustrate that the 5-fold increase in rate of hydrolysis at 52° on going from the open-chain, dimethyl carbonate, to the cyclic analogue, ethylene carbonate, is due mainly to an increase in the entropy of activation. Whereas the energies of activation for these esters are nearly equal, the difference in entropies of activation amounts to 6·6 entropy units. A much greater influence on rate of reaction can be seen on going from the open-chain to the six-membered cyclic compound, trimethylene carbonate. The latter is hydrolysed 457 times faster than dimethyl carbonate at 52°. This large variation cannot be related to an entropy effect because of the small difference (less than 2 entropy units) in the entropies of activation. However, it seems to be due to an energy effect, which amounts to a decrease in 3·3 kcal./mole. The drop in the energy of activation is even greater on increasing the ring-size from 5 to 6. Nevertheless, trimethylene carbonate is hydrolysed at 52° only 92 times faster than ethylene carbonate. This can be interpreted in terms of a concurrent decrease in the entropy of activation. Thus, the decrease in the energy of activation on going from ethylene to trimethylene carbonate is 4·4 kcal./mole, and the decrease in $\Delta S^\ddagger$ is 4·8 entropy units. It seems likely, therefore, that the greater reactivity of the six- than of the five-membered cyclic carbonate in acid-hydrolysis is due mainly to the energy parameter.



(I) (II)

(III) (IV)

The effect of ring-size on reactivity can be explained in terms of conformational changes on going from the ground to the transition state. In ethylene carbonate the H···OH 1,3-interactions in the intermediate (II), formed in step (4) of the reaction, can account for more internal energy than in the homologous intermediate (IV), formed in the hydrolysis of trimethylene carbonate (III), presumably in the chair conformation.[12] With the aid of Dreiding models it was found that the change from (I) to (II) is accompanied by a shortening of the non-bonded cyclic oxygen distances by 0·02 Å. Thus there should occur stretching of the valency bonds and bending of the valency angles and torsion about the dihedral angles[13] on going from the ground to the transition state of the planar cyclic αβ-carbonate.[14] The change from (III) to (IV) apparently is not accompanied by a similar amount of strain.

$\Delta E^\ddagger$ and $\Delta S^\ddagger$ for the cyclic αδ-carbonate cannot be treated in the above manner since the kinetic study of this homologous ester was not done in comparable reaction conditions. However, the similar reactivity of the seven- and the five-membered ring-compound is worthy of note. The relative order of reactivities for acid-catalysed hydrolysis of lactones

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[11] Faurholt, *Dansk Tidsskr. Farm.*, 1943, **17**, 213; 1945, **19**, 255; 1947, **21**, 243; *Acta Chem. Scand.*, 1952, **6**, 404, 1036.

[12] Sarel and Pohoryles, *J. Amer. Chem. Soc.*, 1958, **80**, 4596.

[13] Westheimer in Newman's "Steric Effects in Organic Chemistry," Wiley, New York, 1956, Chapter 12.

[14] Angell, *Trans. Faraday Soc.*, 1956, **52**, 1178; Cumper and Vogel, *J.*, 1959, 3521.

at 0° are 6- > 7- > 5-membered. 6-Hexanolactone is hydrolysed 23-fold faster than γ-butyrolactone.[15] The entropy and the energy of activation for 6-hexanolactone[15] are $\Delta S^{\ddagger}$ $-20{\cdot}1$ e.u., $\Delta E^{\ddagger}$ $15{\cdot}1$ kcal. mole^{-1}.

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[15] Matuszak and Shechter, Abs. Papers Amer. Chem. Soc. Meeting, New York, Sept. 1957, p. 12-p.