615. *Organic Carbonates. Part VI.*[1] *The Rates of Alkaline Hydrolysis of Substituted Cyclic αβ- and αγ-Carbonates.*

By Leo A. Pohoryles, Iddo Levin, and Shalom Sarel.*

The rates of alkaline hydrolysis of a number of branched cyclic αβ- and αγ-carbonates and analogous lactones in water, at 0°, have been measured. The reactions are of the second order. The αγ- show greater reactivity towards hydroxide ions than do the αβ-carbonates, as do other analogous lactones. Alkyl-substitution in the carbon chain affects the rate of hydrolysis, but 2,2-dialkylpropane-1,3-diol carbonates are more rapidly hydrolysed.

Unlike acid- and base-catalysed hydrolysis of simple esters, which has been studied fairly thoroughly,[2] that of carbonic esters has attracted little attention. Since we were interested in the effect of structure on the polymerisation of cyclic carbonates, the kinetics of their

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[1] Part V, Sarel, Levin, and Pohoryles, preceding paper.

[2] Ingold, " Structure and Mechanism in Organic Chemistry," Methuen, London, 1953.

hydrolysis in acid and alkaline solution were investigated. The present paper reports a kinetic study of the alkaline hydrolysis of a series of branched cyclic αβ- and αγ-carbonates in water, and in 1 : 1 acetone–water at 0°.

Since our main concern was with the effect of substituents on the rates of the ring-opening a technique had to be worked out for measuring the rates for the first stage of the alkaline hydrolysis.

In Part V,[1] the hydrolysis of cyclic carbonates in alkali has been described as a multistage process, involving an intermediate (alkyl carbonate) ion as the first stage. This intermediate is the anion of an acid $RHCO_3$ which would be expected to be similar in strength to carbonic acid in its first dissociation ($K_1 = 3 \times 10^{-7}$). It was therefore assumed that the first stage could be followed by measuring the consumption of alkali, an indicator being used to which carbonic acid behaves as a monobasic acid (pH 8·4). Alkaline hydrolysis of the cyclic αβ- and αγ-carbonates was too rapid even at 0° in very dilute solutions of alkali hydroxides, but it was slow enough in potassium carbonate solutions for convenient measurement. A preliminary account of this study has already been given elsewhere.[3]

EXPERIMENTAL

Starting materials have been described in Part IV.[4] Acetone was purified by the method there described. Doubly distilled water was used in the preparation of all solutions for kinetic runs. Hydrochloric acid was titrated against standard hydroxide. Sodium hydroxide solutions occasionally used in back-titration were prepared by diluting a saturated carbonate-free solution.

Kinetic Studies.—Hydrolyses were carried out in Pyrex glass. Solutions were prepared by weighing the desired amount of ester into a 250 ml. flask to which were added 100 ml. of water. Then, at the reaction temperature, were added 150 ml. of water containing either sodium hydroxide or potassium carbonate of such strength that their final molar concentration was between 1·2—~2 times that of the ester. The extent of reaction was determined from the loss of base during the reaction, by titrating the excess of alkali in aliquot parts at intervals with standard hydrochloric acid by use of a potentiometer at pH 8·2. However, the use of a mixture of Cresol Red and Thymol Blue (end-point, pH 8·2) as indicator yields equally good results. The concentration of esters was 0·025—0·035N and that of potassium carbonate 0·064N.

Calculations.—Since the kinetic form taken by the first-stage of the reaction was

$$dx/dt = k_2[\text{carbonate}][OH^-] \quad . \; . \; . \; . \; . \; . \; . \; . \; . \quad (1)$$

the rate constants were calculated by expression (2) which was derived from the integrated form of the second-order rate equation:[5]

$$k_2 = \frac{2{\cdot}303T_\infty}{t} \times \log \frac{T_t(T_0 - T_\infty)}{T_0(T_t - T_\infty)} \times \frac{v}{N} \quad . \; . \; . \; . \; . \; . \quad (2)$$

where T_0 is the volume of standard acid required to neutralise the hydroxide at the beginning of the reaction, T_t is the volume after t min., T_∞ is the volume equivalent to the excess of alkali remaining when all the cyclic carbonate had been converted into intermediate (III), v is the volume of reaction mixture in ml. removed for each titration, and N is the normality of the acid used. Results of a typical run are shown in Table 3.

When potassium carbonate is the source of hydroxide ions, where the hydrolysis $CO_3^{2-} + H_2O \rightleftharpoons OH^- + HCO_3^-$ occurs, the hydroxide-ion concentration can be derived from the equation

$$(OH^-) = K'[CO_3^{2-}][HCO_3^-] \quad . \; . \; . \; . \; . \; . \; . \; . \; . \quad (3)$$

Thus, equation (1) becomes

$$dx/dt = k_2K'[\text{carbonate}][CO_3^{2-}][HCO_3^-] \quad . \; . \; . \; . \; . \; . \; . \quad (4)$$

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[3] Pohoryles, Levin, and Sarel, *Bull. Res. Council Israel*, 1958, **7A**, 221.

[4] Sarel, Pohoryles, and Ben-Shoshan, *J. Org. Chem.*, 1959, **24**, 2067.

[5] Glasstone, "Textbook of Physical Chemistry," Macmillan, London, 1946, p. 1058.

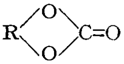
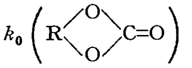
Integration of eqn. (4) by Shields's method [6] gives:

$$k_2 = \frac{2{\cdot}303}{K't} \times \frac{T_0}{T_\infty} \times \log \frac{T_t(T_0 - T_\infty)}{T_0(T_t - T_\infty)} - \log \frac{T_0 - T_\infty}{T_t - T_\infty} = \frac{Z}{K'} \quad . \quad . \quad . \quad (5)$$

where k_2 is the second-order rate constant of hydrolysis by hydroxide ions, T_0 is the volume of standard acid required to change carbonate ions at the beginning of the reaction into bicarbonate, T_t is the volume after t min., and T_∞, that equivalent to the excess of carbonate ions remaining when all the cyclic carbonate has been converted into the open-chain intermediate (III). The quantity K' can be derived from the relation $K' = K_w/K_2$, where K_2 is the second dissociation constant of carbonic acid. For the calculation of K_2, the following values [7] were taken, for 0°: $K_2 = 2{\cdot}36 \times 10^{-11}$; $K_w = 1{\cdot}139 \times 10^{-15}$. Results of a typical run are shown in Table 3.

After this work had been concluded, Kempa and Lee [8] reported a similar derivation for the

TABLE 1. *Alkaline hydrolysis of cyclic carbonates in water at 0°.*

No.	R in 	10^4Z	k_2 (l. mole^{-1} min.$^{-1}$) K_2CO_3	k_2 (l. mole^{-1} min.$^{-1}$) NaOH	k_0  / k_0 (Et_2CO_3)
		Open chain			
1	Et_2	—	—	0·46	1·0
		5-*Membered rings*			
2	$-CH_2{\cdot}CH_2-$	4·92	10·2	—	22·1
3	$-CHMe{\cdot}CH_2-$	2·41	5·0	—	10·8
4	$-CHMe{\cdot}CHMe-$	1·93	4·0	5·0	8·7
5	$-CHMe{\cdot}CMe_2-$	1·0	2·1	2·3	5·0
6	$-CMe_2{\cdot}CMe_2-$	—	—	0·1 [a]	(0·05)
7	$-CHPh{\cdot}CHPh-$	12·0	24·8	—	54
		6-*Membered rings*			
8	$-CH_2{\cdot}CH_2{\cdot}CH_2-$	55	114	—	248
	,,	30 [b]	...		...
9	$-CHMe{\cdot}CH_2{\cdot}CHMe-$	6·7	13·9	—	30
10	${\cdot}CH_2{\cdot}CMe_2{\cdot}CH_2-$	180	364	—	800
	,,	105 [b]	...		...
11	$-CH_2{\cdot}CEt_2{\cdot}CH_2-$	140	239	—	519
	,,	64 [b]	...		...
12	$-CH{\cdot}CEtPh{\cdot}CH_2-$	75 [b]	263 [c]		571
13	$-CH_2{\cdot}CMePr^n{\cdot}CH_2-$	52 [b]	183 [c]		398

[a] Rate in 1 : 1 acetone–water at 30°. [b] Measured in 1 : 1 acetone–water at 0°. [c] Kinetic measurements in 1 : 1 acetone–water of compounds nos. 8, 10, and 11 show that the quantity Z is smaller than in water by a factor 1·7. Accordingly, for comparisons, the rate constant k_2 in water was computed by multiplying the quantity Z, measured in 1 : 1 acetone–water, by the factor 1·7K'.

TABLE 2. *Alkaline hydrolysis of lactones and related esters in water at 0°.*

No.	Substrate	k_2 (l. mole^{-1} min.$^{-1}$) K_2CO_3	k_2 (l. mole^{-1} min.$^{-1}$) NaOH	k_0 (substrate) / k_0 (EtOAc)
1	EtOAc [a]	—	1·2	1·0
2	Et_2CO_3 [b]	—	0·46	0·38
3	$HO{\cdot}CH_2{\cdot}CMe_2{\cdot}CH_2{\cdot}OAc$	—	1·5	1·2
4	$O{\cdot}CH_2{\cdot}CH_2{\cdot}CH_2{\cdot}CO$	6·5	7·56 [c]	5·4
5	$O{\cdot}CHMe{\cdot}CH_2{\cdot}CH_2{\cdot}CO$	2·5	3·5 [d]	2·0
6	$O{\cdot}CH_2{\cdot}CH_2{\cdot}CH_2{\cdot}CH_2{\cdot}CO$	—	428 [c]	(356)

[a] According to Olsson (*Z. phys. Chem.*, 1928, **133**, 233), the rate of hydrolysis of ethyl acetate in water at 0° is 1·2 l. mole^{-1} min.$^{-1}$. [b] By extrapolation of the data of Miller and Case (*J. Amer. Chem. Soc.*, 1935, **57**, 810) the value 0·453 l. mole^{-1} min.$^{-1}$ is obtained for hydrolysis of diethyl carbonate in water at 0°. [c] Data of Matuszak and Shechter (Abs. Papers Amer. Chem. Soc. Meeting, New York, Sept. 1957, p. 12-p) for water–1,2-dimethoxyethane at 0·03°. [d] By extrapolation of the data given by Hückel (" Theoretische Grundlagen der Organischen Chemie," Akademische Verlagsges., Geest und Portig K.-G., Leipzig, 1943, p. 562).

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[6] Shields, *Z. phys. Chem.*, 1893, **12**, 167.

[7] Harned and Owen, " The Physical Chemistry of Electrolytic Solutions," 3rd ed., Reinhold Publ. Corp., New York, 1958.

[8] Kempa and Lee, *J.*, 1959, 1576.

special case in which the initial concentrations of cyclic carbonate and of alkali carbonate were equal.

Results.—Our results are summarised in Tables 1—4 (where k_0 refers to 0°).

TABLE 3. *Hydrolysis of ethylene carbonate in aqueous potassium carbonate at* 0°: Acid = 0·1N-HCl.

Time (min.)	0	30	60	90	120	160	275	300	480	49 hr.
T_t	7·2	6·1	5·8	5·6	5·4	5·1	4·7	4·6	4·2	1·1
k_2 (l. mole^{-1} min.$^{-1}$)	—	10·3	10·3	10·0	9·6	10·9	10·0	10·2	9·6	—

(Average: k_2 = 10·1.)

TABLE 4. *Hydrolysis of tetramethylethylene carbonate in* 1 : 1 *acetone–water at* 32°: Acid = 0·1N. v = 25 ml.

Time (min.)	0	5	45	105	170	230	290	48 hr.
T_b	13·05	12·90	11·90	11·10	10·30	9·95	9·63	8·36
$T_t - T_\infty$	4·71	4·56	3·56	2·76	1·96	1·61	1·29	—
k_2 (l. mole^{-1} min.$^{-1}$)	—	0·125	0·125	0·106	0·113	0·105	0·106	—

(Average: k_2 = 0·113.)

DISCUSSION

From Tables 3 and 4 it can be seen that the first stage in the hydrolysis of cyclic carbonates, whether in alkali hydroxide or in alkali carbonates, proceeds by bimolecular attack of hydroxide ions on the carbonyl carbon atoms, as discussed in Part V. Since cyclic carbonates do not have two axial hydrogen atoms vicinal to the carbonyl group, valuable information could be gained in studying effects of ring-size and substitution upon rates of hydrolysis.

The greater reactivity towards hydroxide ions of the cyclic αγ- than of the αβ-carbonates is qualitatively similar to the behaviour of analogous lactones,[9] but is contrary to the behaviour of analogous sulphites [10] and phosphates.[11] From the quantitative point of view, some significant differences can be noticed between the classical esters of carboxylic and those of carbonic acids. For example, the reactivity toward alkaline hydrolysis increases on going from the open-chain to the five- and six-membered ring carbonates in the order 1 : 22 : 248, as compared with an order of 1 : 6 : 356 in the carboxylic acid series. Further, in the series of five-membered ring esters, ethylene carbonate is hydrolysed at 0° slightly faster than γ-butyrolactone, whereas in the 6-membered ring series δ-valerolactone is hydrolysed some 4-fold faster than trimethylene carbonate.

The kinetic measurements show that accumulation of alkyl groups on the α-carbon atoms of the diol moieties, whether in acid or in alkaline solution, retards hydrolysis of both αβ- and αγ-carbonates. Some pertinent data have been assembled in this laboratory.[12]

For the hydrogen chloride-catalysed hydrolysis of cyclic αβ- and βγ-carbonates in water at 52°, the rates (relative to ethylene carbonate) are: propylene, 0·7; *meso*-1,2,-dimethylethylene, 0·37; trimethylene, 124; 1-methyltrimethylene, 70; 1,3-dimethyltrimethylene 31; 2,2-dimethyltrimethylene, 130; 2,2-diethyltrimethylene, 143.

Unlike substituents, whose effect is relatively small (exception for tetramethylethylene carbonate) changes in ring-size have a large effect on reactivity. This could arise from a difference in the polarities between the two homologous cyclic esters or could be steric in origin. Since it has been shown that changes in ring-size have little effect on the basicities of saturated *N*-heterocycles,[13] it seems reasonable to assume that the basicities of the five- and six-membered cyclic carbonates are similar. Thus, the effect of ring-size upon

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

[10] Bunton, de la Mare, Lennard, Llewellyn, Pearson, Pritchard, and Tillett, *J.*, 1958, 4761.

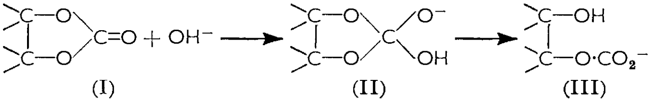
[11] Westheimer, *Chem. Soc. Special Publ.*, 1958, No. 8, p. 1.

[12] Levin, Pohoryles, Sarel, and Usieli, unpublished work; for a preliminary report see *Bull. Res. Council Israel*, 1958, **7**, *A*, 220.

[13] Brown, McDaniel, and Häflinger in " Determination of Organic Structure by Physical Methods," ed. Braude and Nachod, Academic Press, New York, 1955.

reactivity could arise from a difference in ring strains, as suggested by Brown, Brewster, and Shechter.[14]

The effect of substituents in the cyclic αβ-carbonates can reasonably be explained in terms of steric effects. The intermediate (II), formed during the hydrolysis, is much more crowded than either (I) or (III). As a result, changing of trigonal carbonyl in (I) into tetrahedral carbon atom in (II) causes increases in the steric strain of the planar ring [15,16] owing to an increase in the repulsion between cyclic oxygen atoms.



By replacement of hydrogen by methyl in (I) the repulsion between the alkyl groups on the adjacent carbon atoms will be notably large when (II) is obtained. This might deform the over-crowded molecule so as to allow these groups to be further apart and at the same time cause the oxygen atoms of the ring to be brought closer together.[17] The most notable effect in the hydrolysis is the retardation produced by the replacement of all hydrogens in (I) by methyl: tetramethylethylene carbonate is hydrolysed 160—200 times more slowly than trimethylethylene carbonate, which in turn is hydrolysed about 1/5 as fast as ethylene carbonate. With the aid of Dreiding models it was found that the change of a trigonal into a tetrahedral carbonyl carbon atom in tetramethylethylene carbonate is accompanied by a shortening of non-bonded O–O (cyclic) and H–H (methyl) distances by 0·02 Å. This change depicts in essence a " contraction phenomenon " and should increase the internal strain in (II).

The effect of substituents in the six-membered compounds seems to depend upon the position of branching in the diol moiety. Branching at position 1 and/or 3 of the 1,3-diol portion of the cyclic carbonate retards hydrolysis in both acid and alkaline solutions. However, branching at position 2 has different effects on the acid- and the base-catalysed reactions: it does not affect the rate in acid solutions, but has an " accelerating " effect in alkali. It is not known whether 2-alkyl groups exert a true " steric acceleration " effect or whether the lower reactivity of the unsubstituted compounds is to be related to a " steric retardation." It is interesting that similar results were obtained for addition of Grignard reagents to 1,3-cyclic carbonates.[18] Significantly, the carbonate of 2,2-diethylpropane-1,3-diol, showing no tendency to polymerise,[19] gave with Grignard reagents high yields of the expected hydroxy-esters, while poor yields were obtained with trimethylene carbonate under similar conditions. Instead, a substantial amount of polytrimethylene carbonate was obtained. This seems to indicate that the enhanced tendency to polymerisation coupled with lower reactivity towards nucleophiles might be explained in terms of association arising from dipole–dipole interactions between the O–CO–O groups.[15,19,20] These interactions, whether in the solid state or in polar solution, are presumably strong enough to impart definite orientations to the molecules. The effect of branching at position 2 in the 1,3-diol portion of cyclic carbonates upon this behaviour can be explained in terms of steric effects as described above.[19]

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[14] Brown, Brewster, and Shechter, *J. Amer. Chem. Soc.*, 1954, **76**, 467.

[15] Slayton, Simmons, and Goldstein, *J. Chem. Phys.*, 1954, **22**, 1678; Angell, *Trans. Faraday Soc.*, 1956, **52**, 1178.

[16] Otto, *J. Amer. Chem. Soc.*, 1937, **59**, 1590; Bergmann, Fischer, and Pinchas, *Rec. Trav. chim.*, 1952, **71**, 213; Cumper and Vogel, *J.*, 1959, 3521.

[17] Kuhn, *J. Amer. Chem. Soc.*, 1958, **80**, 5950.

[18] Sarel, Pohoryles, and Marcus, unpublished work; *Bull. Res. Council Israel*, 1960, **9**, *A*, 69, contains a preliminary report.

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

[20] Mizushima and Kubo, *Bull. Chem. Soc. Japan*, 1938, **13**, 174; Arbuzov and Shavsha, *Doklady Akad. Nauk S.S.S.R.*, 1949, **68**, 1045; *Chem. Abs.*, 1950, **44**, 886; Hales, Jones, and Kynaston, *J.*, 1957, 622.