[CONTRIBUTION FROM THE RADIATION LABORATORY AND DEPARTMENT OF CHEMISTRY, UNIVERSITY OF CALIFORNIA, BERKELEY]

CARBONATE CLEAVAGE IN THE HYDROLYSIS OF DIETHYL α-NAPHTHYLMALONATE[1, 2]

ARTHUR FRY[3] AND MELVIN CALVIN

Received February 7, 1952

In conjunction with a study of the isotope effect in the decarboxylation of α-naphthylmalonic acid (1), a carbonate cleavage reaction was observed in the basic hydrolysis of the substituted malonic ester.

Such basic carbonate cleavage reactions of substituted malonic esters have been observed by a number of workers (2), but no critical work has been done on the actual point during the hydrolysis at which this cleavage occurs. Wallingford, Homeyer, and Jones (3) were able to prepare substituted malonic esters by an alkoxide-catalyzed Claisen type condensation between diethyl carbonate and the ester of the substituted acetic acid. They showed that the reaction was reversible in many cases and that good yields of the malonic esters could be obtained by using excess diethyl carbonate.

In the present work, the amount of carbonate cleavage has been studied as a function of the amount of hydrolysis of the diethyl α-naphthylmalonate. The data obtained lead to the conclusion that the carbonate cleavage in this case is a reaction of the ion of monoethyl α-naphthylmalonate with base, in competition with the normal hydrolysis reaction.

The kinetic studies on the hydrolysis were carried out as described in the experimental section, and the data obtained are presented in Table I.

On the assumption that the hydrolysis proceeds by the usual second order ester hydrolysis kinetics, values for the specific rate constants for the two ester groups, k_1 and k_2 can be calculated from the integrated rate law. Since k_1 is much greater than k_2, as can be seen by inspection of the data in Table I and of Figure 1, the over-all reaction can be treated approximately as two independent second order reactions with different zero time values. The values of the specific rate constants calculated are given in Table II.

The values of k_1 and k_2 at both temperatures remain quite constant, thus confirming our assumption of second order kinetics. The values of k_1 at the longer times and of k_2 at shorter times are somewhat inaccurate, since in this region our assumption that the hydrolysis of the two ester groups can be treated independently is invalid. The values of k_2 at longer times are probably somewhat inaccurate since the reaction is near completion and we are dealing with small

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[1] The work described in this paper was sponsored by the U. S. Atomic Energy Commission.

[2] This paper was abstracted from the thesis submitted by Arthur Fry to the Graduate Division of the University of California in partial fulfillment of the requirements for the Ph.D. degree, June 1951, and was presented at the Southwest Regional American Chemical Society Meeting, December 1951.

[3] Present address: Department of Chemistry, University of Arkansas, Fayetteville, Arkansas.

TABLE I

KINETIC DATA ON THE HYDROLYSIS OF DIETHYL α-NAPHTHYLMALONATE

TEMP. 25.00°C.			TEMP. 36.95°C.		
Time, hours	mmoles excess base	Amount of ester hydrolyzed,[a] %	Time, hours	mmoles excess base	Amount of ester hydrolyzed[a], %
0	1.500	0	0	2.981	0
0.12	1.308	19.1	0.27	1.976	50.1
.27	1.187	31.2	.32	1.954	51.2
.52	1.060	43.8	.48	1.766	60.6
1.02	0.949	54.9	.70	1.634	67.2
2.00	.834	66.3	1.00	1.474	75.2
3.50	.740	75.7	1.90	1.262	85.7
5.38	.662	83.5	3.03	1.126	92.6
7.67	.613	88.3	4.42	1.049	96.4
15.75	.552	94.4	8.50	0.988	99.5
(∞)	(.496)	(100.0)	(∞)	(.977)	(100.0)

[a] Actually, $\frac{\text{equivalents of base consumed}}{\text{equivalents of ester used}} \times 100$

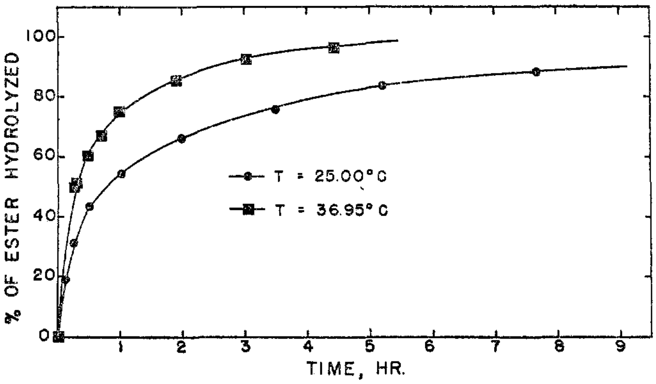


FIG. 1. HYDROLYSIS OF DIETHYL α-NAPHTHYLMALONATE AS A FUNCTION OF TIME

TABLE II

SECOND ORDER RATE CONSTANTS CALCULATED FROM THE INTEGRATED RATE EQUATION AT VARIOUS POINTS

k_1, 25.00°C.		k_2, 25.00 °C.		k_2, 36.95°C.	
t	$k_1 \frac{1}{\text{mole hr.}}$	t	$k_2 \frac{1}{\text{mole hr.}}$	t	$k_2 \frac{1}{\text{mole hr.}}$
0.12	33.2	1.02	2.23	0.48	12.1
.27	31.7	2.00	3.27	.70	11.7
.52	38.3	3.50	3.21	1.00	12.5
		5.38	3.28	1.90	11.3
		7.67	3.10	3.03	11.2
		15.75	2.40	4.42	11.2
				8.50	10.3

differences between large numbers. This method of obtaining data was not deemed sufficiently sensitive to calculate k_1 at the higher temperature.

When the completely hydrolyzed samples were made strongly acid, carbon dioxide was rapidly evolved. This carbon dioxide did not come from the decarboxylation of the malonic acid, as shown by Figure 2. A sample which had been completely hydrolyzed was acidified and the carbon dioxide evolved was collected as a function of time. After carbon dioxide evolution was complete at one temperature, the temperature was increased and the experiment continued.

By examination of this figure it is obvious that the initial carbon dioxide evolution has a different source from that evolved after heating the solution. The final carbon dioxide evolution is undoubtedly from the decarboxylation of the malonic acid, while the initial rapid carbon dioxide evolution arises from carbonate as such.

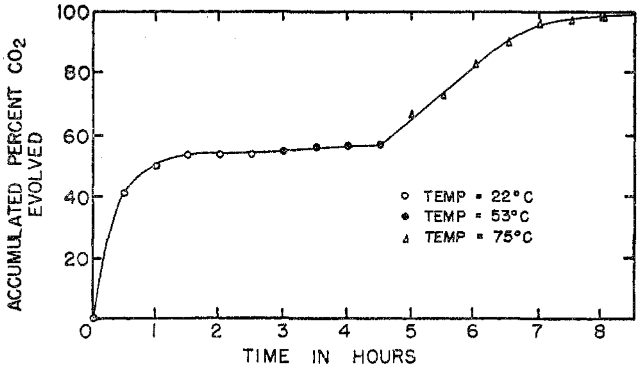


FIG. 2. CARBON DIOXIDE EVOLUTION FROM THE HYDROLYSIS PRODUCT OF DIETHYL α-NAPHTHYLMALONATE.

In order to determine the amount of carbonate obtained as a function of the amount of hydrolysis, the samples in the above 36.95° kinetic run and in an earlier kinetic run at room temperature were acidified in a sweep system and the carbon dioxide evolved in 2½ hours at room temperature was collected. This was sufficient time to insure complete collection of the carbon dioxide as shown by Figure 2. The results of these measurements are given in Table III and are shown graphically in Figure 3. The percentages are calculated on the basis of the hydrolysis of the diester to the diacid, and the decarboxylation of the diacid to the monoacid and carbon dioxide.

From the data presented in Table III it is readily seen that the carbonate *is not* coming from the decarboxylation of the di-ion of the acid, since at both temperatures the amount of carbonate does not increase upon long standing in base after the hydrolysis is complete, even though in each 40–50% of the original ester is present as the di-ion of the malonic acid. Likewise, the carbonate *is not* being formed during the hydrolysis of the first ester group. At 50% hydrolysis at the end of one hour at room temperature (corresponding approximately to complete hydrolysis of the first ester group) no carbonate has been formed, while at the end of the second hour, 7.8% of the carbonate cleavage has occurred while an additional 10.6% hydrolysis has occurred. The higher temperature case is equally clear cut. Finally, the rate of production of carbonate is roughly proportional to the rate of hydrolysis of the second ester group, as is shown by the constancy of the ratio in the final columns at the two temperatures in Table III and by the fit of the data to the straight lines in Figure 3. These

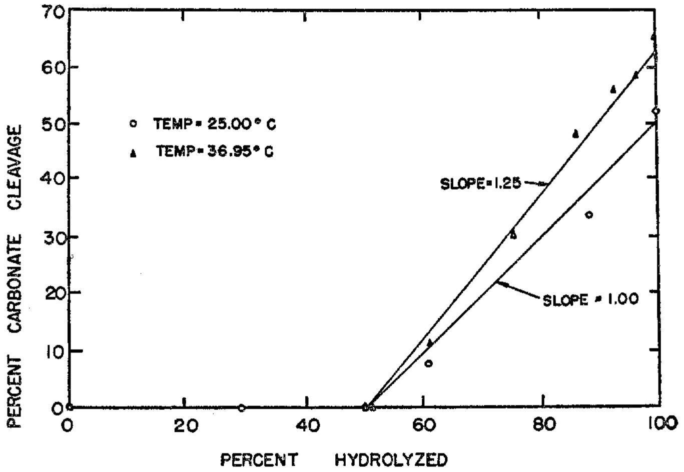


FIG 3. CARBONATE CLEAVAGE AS A FUNCTION OF HYDROLYSIS OF DIETHYL α-NAPHTHYLMALONATE.

TABLE III

AMOUNT OF CARBONATE CLEAVAGE AS A FUNCTION OF THE AMOUNT OF HYDROLYSIS OF DIETHYL α-NAPHTHYLMALONATE

ROOM TEMPERATURE, 25°C.				36.95°C.			
Time, hours	Hydrolysis, %	Carbonate Cleavage,[a] %	$\frac{\% \ CO_3^=}{\% \ Hyd. -50}$	Time, hours	Hydrolysis, %	Carbonate Cleavage,[a] %	$\frac{\% \ CO_2^=}{\% \ Hyd. -50}$
0[b]	0	0	—	0[b]	0	0	—
0.5	29.0	0.0	—	0.27	50.1	0.0	—
1.0	50.0	.0	—	.32	51.2	0.0	—
2.0	60.6	7.8	0.74	.48	60.6	11.3	1.07
17.0	88.0	33.9	1.03	1.00	75.2	30.2	1.20
45.0	102.4	54.5	1.04	1.90	85.7	48.1	1.35
88.5	108.3[c]	52.1	0.90	3.03	92.6	55.8	1.31
				4.42	96.4	58.2	1.25
				8.50	99.5	65.1	1.32
				22.67	99.9	61.8	1.24

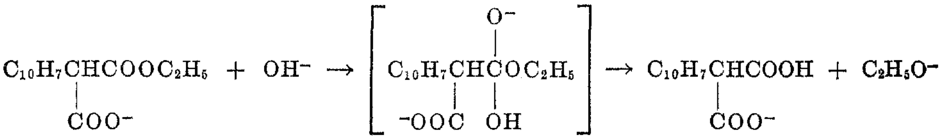
[a] This column is the ratio of $CO_3^=$ formed to moles of ester initially present. [b] Blank run. [c] This value is thought to be high due to an error in titration.

three observations lead to the conclusion that the carbonate is being produced from the ion of the half-ester, either during the hydrolysis or by a path parallel with the hydrolysis with a nearly equal activation free energy.

Combination of the data from Tables II and III enables us to calculate the heat and entropy of activation for the normal hydrolysis and the carbonate cleavage of the monoester ion (4). The values of k_2 given in Table II are actually the sums of the second order rate constants for the hydrolysis reaction and the carbonate cleavage reaction. The slopes of the lines in Figure 2 at the two temperatures are the ratios of these two constants.

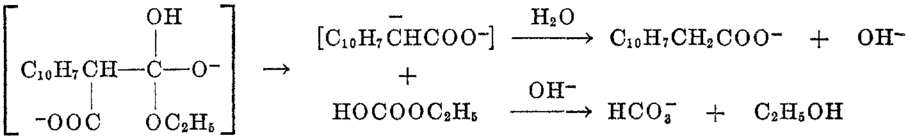
The heat of activation for the hydrolysis reaction is somewhat lower than that of the carbonate cleavage reaction but the hydrolysis reaction is actually slower since the entropy of activation for the hydrolysis reaction is lower than that for the carbonate cleavage reaction. Apparently, the activated complex for the hydrolysis is more restricted than that for carbonate cleavage.

It is interesting to speculate on the mechanism of the reaction of the ion of the half-ester with base. The hydrolysis may be represented by the commonly accepted mechanism (5) of the attack of the hydroxide ion on the carbonyl carbon of the ester, followed by elimination of the ethoxide ion.



$$\underset{\substack{|\\ \mathrm{COO^-}}}{\mathrm{C_{10}H_7CHCOOC_2H_5}} + \mathrm{OH^-} \rightarrow \left[\begin{array}{c} \mathrm{O^-} \\ | \\ \mathrm{C_{10}H_7CHCOC_2H_5} \\ |\quad\; | \\ \mathrm{^-OOC\quad OH}\end{array}\right] \rightarrow \underset{\substack{|\\ \mathrm{COO^-}}}{\mathrm{C_{10}H_7CHCOOH}} + \mathrm{C_2H_5O^-}$$

This same intermediate complex may also decompose by the rupture of the carbon-carbon bond instead of the carbon-oxygen bond, giving the half-ester of carbonic acid and a di-ion of α-naphthylacetic acid.



$$\left[\begin{array}{c} \mathrm{OH} \\ | \\ \mathrm{C_{10}H_7CH{-}C{-}O^-} \\ |\qquad\quad | \\ \mathrm{^-OOC\qquad OC_2H_5}\end{array}\right] \rightarrow \begin{array}{c} [\mathrm{C_{10}H_7\overset{-}{C}HCOO^-}] \xrightarrow{\mathrm{H_2O}} \mathrm{C_{10}H_7CH_2COO^-} + \mathrm{OH^-} \\ + \\ \mathrm{HOCOOC_2H_5} \xrightarrow{\mathrm{OH^-}} \mathrm{HCO_3^-} + \mathrm{C_2H_5OH}\end{array}$$

The carbon-carbon bond in the above complex is undoubtedly weakened by the repulsion between the two negatively charged ends of the molecule, thus facilitating the carbonate cleavage. This weakening effect is not present in the hydrolysis of the diester, and is perhaps the reason why we do not get this cleavage during this first hydrolysis step.

Examination of the above mechanism reveals that it is simply a Claisen ester condensation in reverse, where hydroxyl ion is the base. This type of condensation is known to be reversible (6, 7), and, as mentioned above, Wallingford, Homeyer, and Jones (3) have demonstrated the specific reversibility using diethyl carbonate as one of the esters. Undoubtedly this reverse Claisen condensation takes place in the hydrolysis of all malonic esters, but with most it is so much slower than the normal hydrolysis that it is never observed.

It is well known (6, 7) that esters with only one *alpha* hydrogen give very poor results in the normal Claisen condensation unless a base stronger than ethoxide, such as sodium triphenylmethyl, is used. Since this is the case, we would expect a dialkylacetic ester and ethyl carbonate to give a poor forward Claisen condensation, and, correspondingly, the product, a dialkylmalonic ester, might be expected to give an improved reverse reaction, or carbonate cleavage. In view of this, we would expect α-naphthylalkylmalonic esters to undergo this reverse Claisen condensation especially readily, and it may be that "hydrolysis" of these esters goes almost exclusively by the carbonate cleavage path. Blicke and Feldkamp (8) prepared a series of α-naphthylalkylacetic acids by hydrolysis of the corresponding malonic esters and acidification of the resulting salts. They reported that the "malonic acids" so obtained spontaneously decarboxylated at room temperature. It seems very probable that the source of the carbon dioxide evolved in this case was not the decarboxylation of the malonic acid but rather the reverse Claisen condensation in basic solution.

The question of why α-naphthylmalonic ester (and other arylmalonic esters) undergoes this carbonate cleavage more readily than aliphatic malonic esters is undoubtedly connected with the steric effect of the large bulky aryl groups as well as with the resonance stabilization of the intermediate ion.

Future work is planned on the hydrolysis and carbonate cleavage of pure monoethyl-α-naphthylmalonate, and on the reverse Claisen condensation of other substituted malonic esters.

EXPERIMENTAL

Preparation of diethyl α-naphthylmalonate. Diethyl α-naphthylmalonate was prepared by the condensation of ethyl α-naphthylacetate with diethyl oxalate, followed by decarbonylation of the resulting glyoxalate according to the procedure of Blicke and Feldkamp (8). Recrystallization of the crude ester from alcohol gave material melting at 62–62.5°, reported m.p. 62° (8). Alcohol is a much more satisfactory solvent for recrystallization than the petroleum ether used by Blicke and Feldkamp.

Hydrolysis of diethyl α-naphthylmalonate. The kinetic runs on the hydrolysis of diethyl α-naphthylmalonate were carried out by dissolving weighed samples of 0.502 mole of the ester in 10 cc. of 95% ethanol and adding 1.5 cc. of 1.000 N carbonate-free sodium hydroxide. The alcohol and base solutions were preheated to the bath temperature before being used. An individual sample was prepared for each point since the formation of a precipitate of sodium salts precluded representative sampling during the later stages of the hydrolysis. The experiments were carried out in a constant temperature bath maintained at 25.00 ± 0.05° or 36.95 ± 0.05°C.

After standing for appropriate lengths of time in the constant temperature bath the solutions were washed into excess water and immediately titrated with standard acid using a pH meter to determine the end point. When the samples were poured into the water, the hydrolysis of the diester was stopped for all practical purposes by its precipitation. The time required for titration was short compared to the rate of hydrolysis of the monoester. After titration to pH 7.5, the solutions were quickly transferred to a sweep system, 10 ml. of 1 N hydrochloric acid was added, and the solution was swept for 2½ hours with a stream of carbon dioxide-free air. The carbon dioxide was absorbed from the air stream in a spiral bubbler filled with 1 N sodium hydroxide, and precipitated and weighed as barium carbonate.

Preparation of α-naphthylmalonic acid. The solutions remaining from the above kinetic runs were combined, made alkaline and ether-extracted to remove any unreacted ester. The basic solution was then acidified and ether-extracted again. Most of the ether was removed at room temperature to avoid decarboxylation of the free malonic acid, the solution was filtered, and a large excess of benzene was added. Upon standing, the α-naphthylmalonic acid crystallized as fine white prisms, leaving any half-ester and the α-naphthylacetic acid in solution. After purification by repeated solution in ether and precipitation by benzene, α-naphthylmalonic acid, m.p. 162–165° dec., equivalent weight 115.5, calculated 115.1, was obtained. Ivanov and Pshenichnii (9) reported the melting point as "toward 151° dec."

SUMMARY

A kinetic product study of the basic hydrolysis of diethyl α-naphthylmalonate has shown that the carbonate formed results from the direct fission of the ion of the half ester. The carbonate cleavage reaction is pictured as a type of reverse Claisen condensation.

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