

864. *The Hydrolysis of Amides of Dibasic Acids. Part II.**
The Acid Hydrolysis of Oxamide and Oxamic Acid.

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Rates of hydrolysis of oxamide and oxamic acid in excess of hydrochloric acid have been measured at 79.5° and at 97.5°. In the conversion of oxamide into oxamic acid, the rate of hydrolysis per amide group is very much less than in the second stage of hydrolysis. No fission of the inter-carbon bond, during hydrolysis, could be detected.

EXPERIMENTAL results reported in Part I* showed that when malonamide is hydrolysed in excess of hydrochloric acid at temperatures between 80° and 100°, the value of the first-order rate constant, k_1 (diamide \rightarrow monoamide), is approximately twice that of k_2 (monoamide \rightarrow dibasic acid); *i.e.*, the rate of hydrolysis per amide group is approximately the same for both stages of hydrolysis. Under our experimental conditions, therefore, the damping effect of the intervening methylene group is such that the measured rate of hydrolysis of one amide group is not markedly affected by a change, from amide to carboxyl, in the other functional group. In the amides of oxalic acid the functional groups are directly linked and similar kinetic studies should reveal more clearly the expected difference in electronic effects between amide and carboxyl groups. The acid hydrolyses of oxamide and oxamic acid were therefore followed under experimental conditions comparable with those obtaining in the earlier work. In addition, because the inter-carbon bond length in oxamide is high (Misch and van der Wyk, *Arch. Sci. phys. nat.*, 1938, 20, Suppl. 96), the possibility of hydrolysis being accompanied by fission at this bond was also investigated.

EXPERIMENTAL AND RESULTS

Analytical methods and other experimental details were described fully in Part I. Hydrolyses were carried out at either 79.5° \pm 0.3° or 97.5° \pm 0.5°. Dissolution times of the small quantities of solutes used were again too short to affect the kinetics.

Oxamic acid was prepared by Oelkers's method (*Ber.*, 1889, 22, 1569) and, after several recrystallisations, melted, with decomposition, at 209°.

Oxamic acid reacted extensively with hypobromite but no apparent hydrolysis took place when an aliquot of oxamic acid solution was subjected to the distillation procedure, the buffered solution having a pH of 11.5. This distillation method was therefore used for analysis. A linear $\ln(a-x)-t$ plot was obtained in each hydrolysis run (Fig. 1), and the rate constant obtained therefrom was unaltered by variation in the amide: acid ratio, thus confirming the pseudounimolecular nature of the reaction. Hydrolyses were followed within the range $[\text{HCl}] = 0.0156\text{--}0.75\text{M}$, and the mean values for k_2 at any given concentration of hydrochloric acid were plotted against the molar concentration of acid, giving a linear plot at each temperature (Table). The graphs can be defined by the equations

$$k_2 = 0.0008 + 0.0475 [\text{HCl}] \text{ (at } 79.5^\circ\text{)}; \quad k_2 = 0.007 + 0.184 [\text{HCl}] \text{ (at } 97.5^\circ\text{)}$$

and the approximate activation energy, from these results, is 20.5 kcal. mole⁻¹.

Oxamide was prepared by ammonolysis of ethyl oxalate and, before use, was repeatedly recrystallised from water.

Estimation of ammonia in the presence of oxamide was not easy. The direct titration method was at first disregarded because hypobromite reacted with oxamide. The distillation method, as employed successfully for oxamic acid, was useless because oxamide was hydrolysed to a considerable extent in the relatively strongly alkaline solution; this fact is in accord with the results of Westheimer and Shookoff (*J. Amer. Chem. Soc.*, 1940, 62, 269) who have shown that oxamide is hydrolysed, in alkaline solution, about 100 times faster than oxamic acid. By reducing the pH of the buffered solution to 9.8, however, it was possible to effect 100% recovery of ammonia in 15 minutes, with only 3.5% hydrolysis of the amide. This error, although not insignificant, was by no means large, and allowance could be made for it. A series of hydrolyses were therefore followed at each of the two temperatures. In all cases the $\ln(a-x)-t$ plots were reproducible up to about 50% hydrolysis, consistent values being obtained for $k(=k_1/2)$; see

* Part I, *J.*, 1952, 3264.

Part I). In each instance the logarithmic plot deviated from linearity after the hydrolysis had reached a half-way stage, but consistency was thereafter absent. On occasions the graph broke sharply at this point to give a new straight line from which could be obtained a rate constant somewhat similar to that obtained for oxamic acid hydrolysis (k_2); at other times the deviation from the initial slope was only slight (Fig. 1). The majority of results fell between the two extremes. The inconsistent results obtained for this stage of the hydrolysis are presumably due to the limitations of the analytical method employed. Although relative possible errors will increase with time owing to the decrease in value of the $(a - x)$ term and thus of the titre, the very much lower pH of the distillation buffer (as compared with that used for oxamic acid) further reduces accuracy. With oxamide this analytical method is clearly reaching the limits of application and fails to give reliable information concerning the second stage of oxamide hydrolysis. However, the k_1 values may be combined with values for k_2 , obtained from work on oxamic acid itself, to give a picture of oxamide hydrolysis. Rate constants $k (= k_1/2)$ were again determined at several acid concentrations, at the two specified

FIG. 1.

Hydrolyses in 0.50M-hydrochloric acid at 79.5°.

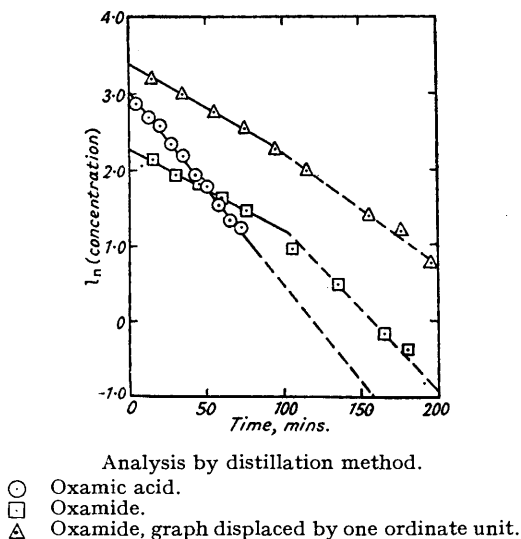
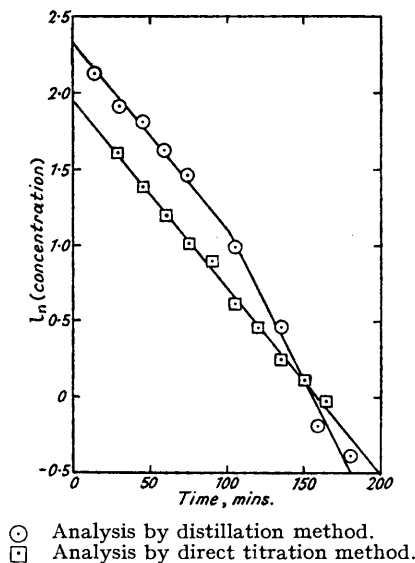


FIG. 2.

Hydrolysis of oxamide in 0.50M-hydrochloric acid at 79.5°.



temperatures. The plot of k against $[HCl]$ at each temperature is a linear one (Table) which passes through the origin and may be described by one of the simple equations:

$$k = 0.0249 [HCl] \text{ (at } 79.5^\circ\text{)}; \quad k = 0.0972 [HCl] \text{ (at } 97.5^\circ\text{)}$$

The approximate energy of activation is 20.5 kcal. mole⁻¹.

The direct titration method, as applied to oxamide, is worthy of further mention because a more detailed examination led to a curious kinetic result. In the reaction between hypobromite and oxamide one mole of hypobromite was found to be consumed per mole of oxamide. The reaction took 3 minutes for completion under our experimental conditions and no further consumption of hypobromite was noted even after some considerable time. This reaction corresponds to an apparent 33% hydrolysis of oxamide. Hydrolysis runs were carried out on oxamide, the direct titration method being used, and the natural logarithm of "apparent concentration of oxamide" was plotted against time. The resulting graph was in each case linear and of slope identical with that of the linear portion of the graph obtained by use of the distillation method. The plot could, however, be projected back to a point on the logarithm axis corresponding to 33% hydrolysis of the oxamide (Fig. 2). This is explicable only in part, and the explanation is only valid for the time during which the oxamic acid intermediate is not present in high concentration. If the extent of the hypobromite-oxamide reaction were constant for all hydrolysate samples during a given experiment, this reaction would show up in each titre as an apparent increase in the extent of hydrolysis; the increase would be

proportional (33%) to the amount of unhydrolysed oxamide. It may readily be shown that the graph of "ln(apparent concentration of oxamide)-time" should have the same slope as that of "ln(actual concentration)-time" and should be displaced to the same extent as that found experimentally. This simple explanation can apply, however, only while the concentration of oxamic acid, which itself reacts with hypobromite, is low. It is true that the early stages of the hydrolysis should be almost free from this complication but calculation of oxamic acid concentrations [by Swain's method (*J. Amer. Chem. Soc.*, 1944, **66**, 1691)] indicates that at 20% hydrolysis the oxamic acid concentration becomes sufficiently large to invalidate the above conclusion. The fact remains, however, that in every case the monolinear form is maintained for the whole of the hydrolysis, and the logarithmic plots lead to rate constants duplicating those obtained for k by the distillation method. Thus the direct titration method could, in practice, be used to extend the number of k values obtained by the more laborious, alternative procedure, and has in fact been so used. The nature of the reaction between oxamide and hypobromite is also not clear, although from the definite reaction proportions, and from Foster's work (*J.*, 1878, **33**, 470; 1879, **35**, 119) it appears possible that initially monobromination occurs. There

10^3 Rate constants (min.^{-1})

Molarity of acid	Oxamide ($k = k_1/2$)		Oxamic acid (k_2)	
	79.5°	97.5°	79.5°	97.5°
	0.0156	—	[1.1]	—
0.0625	—	6.2, 6.7	—	—
0.125	—	13.0, 13.0	—	30.4, 30.6
0.25	6.05, 6.15, 6.18, 6.02, 6.12	23.4, 24.1	12.9, 12.9, 12.6, 13.0	54.0, 53.0
0.35	—	33.8, 33.6	—	70.0, 70.0
0.375	8.3, 10.5, 10.2, 8.8, 9.8, 9.2	—	17.5, 18.0, 18.3	—
0.50	11.5, 12.7, 12.2, 11.9, 12.9, 12.9	48.3, 49.3	25.0, 25.0, 25.5	96, 99
0.75	18.4, 18.4, 18.6, 19.0, 19.2, 18.9	—	39.4, 33.0, 36.0	—

was the possibility that, under our experimental conditions, fission of the inter-carbon bond may take place, with consequent release of oxides of carbon. This was checked during runs on both oxamide and oxamic acid at the lower of the two reference temperatures. Dry air, free from carbon monoxide and carbon dioxide, was drawn through the contents of the reaction vessel throughout the run, the emergent gases being examined for oxides of carbon by standard methods. No positive tests were obtained although the reagents had previously been shown to be sufficiently sensitive to reveal the presence of carbon oxides resulting from any appreciable amount of inter-carbon bond fission.

The theoretical significance of these results is being reserved for discussion in a later paper, when related work on a number of substituted malonamides has been completed.

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