[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY, UNIVERSITY OF SOUTHERN CALIFORNIA]

The Ether Hydrolysis of Some Halogen Substituted Phenoxyacetic Acids

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The rates of ether hydrolysis of phenoxyacetic acid and of 2-fluoro, chloro and bromo, of 4-fluoro, chloro and bromo, of 2,4-difluoro, and dichlorophenoxyacetic acids in concentrated hydriodic acid at 100° and 115° are slow, k being of the order 10^{-5} to 10^{-6} sec.$^{-1}$. The effect of halogen substitution was small. The study was complicated by a competing reduction of the chlorophenols and of the bromophenoxyacetic acids, and bromophenols.

Since not too much is known of the particular structural or functional features of phenoxyacetic acids responsible for biological activity in plants, it was decided to study systematically the effect of halogen substitution in the 2- and 4-positions of the aromatic ring on a particular reaction of this type of compound. The ether cleavage reaction with hydriodic acid was chosen because it involves a site of reaction close to the aromatic nucleus and likely to be influenced by halogen substitution. Data were obtained on the rates of ether hydrolysis of 2-, 4- and 2,4-substituted fluoro-, chloro- and bromophenoxyacetic acids as well as phenoxyacetic acid itself in concentrated hydriodic acid at 100 and 115°. An important side reaction was the dehalogenation of the substituted acids and phenols under these conditions. Since the importance of this side reaction increased in the order F < Cl < Br, the iodophenoxyacetic acids were not studied. The hydrolysis proved to be surprisingly slow, requiring one to two weeks for completion. Concentrated hydrobromic acid, although more stable than hydriodic acid at the temperature used, gave a much slower rate of cleavage. After three days at 80°, only 3% reaction had occurred with concentrated hydrobromic acid and 2,4-dichlorophenoxyacetic acid.

The reaction was followed by diluting samples, making basic, and determining the ultraviolet absorption at the peaks of the salts of the acids and of the corresponding phenols in the region of 270 to 310 mμ.[1] Since this allowed the determination of both reactant and product, a check was provided on the accuracy of the method and on the extent of side reactions. By making the solutions basic before measuring the absorption, the peaks of the acid and phenol were spread about 20 mμ apart, whereas in neutral or acidic solution the peaks of the components were practically superposable. The determination of the phenol concentration by bromide–bromate titration[2] was unsatisfactory. Although the phenols took up bromine quantitatively in the open 2-, 4- and 6-positions in about two minutes, the phenoxyacetic acids also took up bromine in an irregular fashion depending upon the substituents already present, one atom in about two minutes, and part of a second, or a second and part of a third, during the 15 minutes used in the determination.

Experimental

All of the phenoxyacetic acids used were supplied through the courtesy of Dr. R. L. Weintraub except for the 4-bromo- and the 2-fluorophenoxyacetic acids which were prepared here by the usual methods. Of the phenols, all of the fluorophenols were prepared here by standard methods, the rest being commercially available.

The spectrum for each acid and phenol was measured on a solution prepared by dissolving 0.3 to 1.8 millimoles of the material in distilled water which contained 0.2 g. (5 millimoles) of sodium hydroxide and making up to 250 ml. Further dilutions, usually 1 to 10, were made as required in order to bring the absorptions within the range of the instrument, a Beckman model DU. All readings were made against blanks which contained all substances except that being measured. Data for the ranges 255–310 mμ were obtained and those wave lengths chosen which furnished suitable extinction coefficients for the calculation of the concentrations of the components of a given system. Table I lists the wave lengths of maximum absorption and the extinction coefficients.

TABLE I

ABSORPTION MAXIMA OF HALOGEN SUBSTITUTED PHENOXIDES AND PHENOXYACETATES

Substituents	Phenoxides λ_{max}	Phenoxides ϵ_{max}	Phenoxyacetates λ_{max}	Phenoxyacetates ϵ_{max}
Unsubstd.	286.5	2530	269	1390
2-F	283	2680	267	1390
2-Cl	293	3660	273	1810
2-Br	295	3940	273.5	2030
4-F	298	2970	277	2150
4-Cl	298[a]	2400[a]	279	1420
4-Br	298	2300	279	1370
2,4-diF	293	2960	272	1900
2,4-diCl	304	3620	283	2020
2,4-diBr	307.5	3700	284.5	2060

[a] Ref. 1 gives 298 and 2600 for 4-chlorophenoxide.

Concentrations of the starting material and of the phenol formed from it were determined on each 5-ml. sample taken during a run by dilution to a volume of 100 ml. with distilled water after neutralization of the hydriodic acid and conversion of the phenoxyacetic acid and phenol to their salts by the addition of about 1.5 g. of sodium hydroxide. Samples of the diluted mixture were read against a similarly treated blank solution of hydriodic acid. According to Beer's law the optical density, D, at a given wave length of a solution containing two or more components is equal to the sum of the products of the concentration of each substance multiplied by the extinction coefficient, ϵ, of the substance at the wave length.[3] The equations may be solved to find the concentration of each component if absorptions are measured at a number of different wave lengths. In those cases in which reduction of the substituted phenol by hydriodic acid was suspected, a similar set of equations was solved for three components and the optical densities measured at three different wave lengths. In all cases, the blank exhibited a large unexplained absorption beginning at 275 mμ and continuing on down to shorter wave lengths. This absorption persisted whether or not the blank was acidic or basic when measured. It was assumed that this absorption was cancelled out by comparison of the unknown with the blanks.

(1) L. Doub and J. M. Vandenbelt, THIS JOURNAL, **69**, 2714 (1947); **71**, 2414 (1949).

(2) S. Siggia, "Quantitative Organic Analysis via Functional Groups," John Wiley and Sons, Inc., New York, N. Y., 1949, p. 111.

(3) W. West in "Technique of Organic Chemistry," A. Weissberger, ed., Interscience Publishers, Inc., New York, N. Y., Vol. 1, Part II, second ed., 1949, p. 1298.

Another source of error was the iodination of phenols present in the reaction mixture when the aliquots were made alkaline. The hydriodic acid solutions became progressively darker during a run. Since the color disappeared from the aliquots when they were made basic, the color could have been due to the presence of free iodine, at least in part. If so, the error involved must have been small since the total concentration of components usually agreed quite well with the calculated initial concentration of the phenoxyacetic acid used.

Rate Studies.—A phenoxyacetic acid (0.0800 to 0.5000 g., 0.5 to 1.5 mmoles, depending upon the dilution of the samples for measurement of absorption) was weighed into a 100-ml. volumetric flask, and 90 to 95 ml. of hydriodic acid was added. The hydriodic acid was a commercial reagent grade, 47–50% assay, sp. gr. 1.5, and contained 1.5% of hypophosphorous acid as a preservative. Enough of the acid was pooled to ensure an adequate supply of homogeneous solvent for all the runs. The flask was placed in an oil-bath thermostat held at 100 or 115° and allowed to come to temperature, approximately 15–30 minutes. The volume was brought to the mark with additional solvent, already thermostated, the contents mixed by stirring, and the flasks closed with ground glass stoppers lubricated with a heavy silicone grease. A blank solution containing only the solvent was placed in the bath at the same time. Three to six runs usually were made simultaneously. The time of beginning the reaction was chosen as the time at stirring. Samples were withdrawn from time to time and analyzed as described previously. Concentrations so determined fitted a curve for a first-order reaction.

The process of removing a sample with a pipet cooler than the temperature of the bath caused a cooling of the liquid and an expansion of the pipet which was difficult to correct. Instead, no corrections were applied, and the initial concentration was taken as the sum (c) of the concentrations of phenol and phenoxyacetic acid as determined at room temperature from the absorptions rather than the value as calculated from the weight of phenoxyacetic acid used (c_i). In this manner, errors due to temperature and dilution effects were made comparable point to point and run to run, although the concentrations so determined disregarded the difference between the actual solution in the flask at the higher temperature and the contracted volume of the sample taken for dilution. In general c_0 averaged less than c_i by about 1 to 4%. In turn, during a given run c_0 showed an average deviation of 3% or less. Since the extinction coefficients used were considered to be accurate only to 1–2%, this was considered to be excellent precision. Since the expansion of the flask was not corrected for, the slight error as introduced was considered negligible in comparison to the others mentioned.

The first-order plots for phenoxyacetic acid and for the fluorophenoxyacetic acids were good to over 90% completion, about 80 hr. at 115° and 300 hr. at 100°. Although the sums of starting material and product throughout each run at 100° deviated by only small amounts, at 115° the sums dropped toward the end of each run. This was interpreted as loss of the relatively volatile phenols at the higher temperature when the flasks were opened to take samples. No effect on the rate constant was noted.

The first runs on the chlorophenoxyacetic acids displayed a tapering off of the rate constants toward the ends of the runs as well as a lowering of the chlorophenol concentration. Because the chlorophenols gave their ϵ_{max} at longer wave lengths than phenol, and if the chlorophenols were being reduced in the hydriodic acid solutions to phenol, then the calculated concentrations of the remaining chlorophenoxyacetic acid would be in error. In order to check this possibility, weighed amounts of the chlorophenols as well as phenol were dissolved in hydriodic acid under the same conditions which were used for the hydrolyses except that they were thermostated at 115° for 1 day, then for two weeks at 100°. Samples were withdrawn after 6 and 24 hours at 115°, and after 5 days and 2 weeks at 100°. After dilution, the alkaline absorptions were measured against the blank at 272, 280, 287, 293, 298, 304 and 310 mμ and the values obtained were plotted. Phenol remained constant but the three chlorophenols all showed progressive shifts in absorption toward that of phenol. It was possible to calculate that at the end of the experiment, 2-chlorophenol was reduced to the extent of 74% and 4-chlorophenol, 53%. Since the dehalogenation of 2,4-dichlorophenol proceeds stepwise, it was not possible to evaluate the separate steps. By assuming no monochlorophenols present, one could estimate the reduction to be 33% complete.

The hydrolyses were repeated and the samples measured for absorption at three wave lengths corresponding to the ϵ_{max} for chlorophenoxyacetate, chlorophenoxide and phenoxide. In the case of 4 chlorophenoxyacetic acid, the sum of the concentrations of the three components was constant to within 1–2% throughout the runs and was within 1–3% of the initial concentrations as calculated from the weights of starting material used. For the 2-chlorophenoxyacetic acid, the sums dropped markedly after 2 days at 115°, and 4 days at 100°. This could have been caused by the greater volatility of the 2-chlorophenol, b.p. 175.6°, as compared to phenol, b.p. 182°, or to 4-chlorophenol, b.p. 217°. However, in spite of the greater rate of dehalogenation of 2-chlorophenol than 4-chlorophenol, the concentration of phenol in the 2-chlorophenoxyacetic acid runs as calculated was negative (up to 5% of the c_0) during the first half of the reaction. This discrepancy may have been due to some dehalogenation of the 2-chlorophenoxyacetic acid before ether cleavage took place, even though the 4-chloroacid gave no such indication.

The results on 2,4-dichlorophenoxyacetic acid were still unsatisfactory. Over and above the possibility of dehalogenation of the acid, the problem of the stepwise reduction of the dichlorophenol remained. Also, since 2,4-dichlorophenoxyacetate had ϵ_{max} at 283 mμ, while phenoxide had ϵ_{max} at 286.5, and both have broad peaks, the accuracy of resolution of the measured absorptions was poor. In order to get reasonable separation of values of ϵ_{max}, absorptions were measured at four wave lengths. The absorption coefficient for phenoxide, 2,4-dichlorophenoxide and 2,4-dichlorophenoxyacetate at 275.4 mμ were 1620, 735 and 1400; at 283 mμ were 2354, 1235 and 2030; at 300 mμ were 972, 3460 and 183; and at 310 mμ were 116, 3220 and 44. Various combinations of the data for three wave lengths were used to calculate the concentrations. Each combination tried gave negative values for the concentration of phenol. The 275.4, 300 and 310 mμ combination was chosen as best because the phenol concentrations were the smallest negative values found. If only the absorptions at two wave lengths were used and the concentration of dichloroacid and dichlorophenol calculated, it turned out that the concentrations of dichlorophenol by both methods coincided quite well point to point. But the three component calculations for the dichloroacid were higher than the two component calculations by an amount equal to the negative phenol concentrations of the three component calculation. The rate constants calculated from the three component results tended to drift upward during the course of the 100° run, while the constants for the two component results held steady. Also the rate constants from the two component calculations are more consistent with the results for the monochloroacids. Even so, the values given for the rate constants by either calculation should not be considered to be accurate to more than an order of magnitude.

In the first runs on the hydrolysis of 2-bromo-, 4-bromo- and 2,4-dibromophenoxyacetic acids at 100°, it was noticed that after five days the slopes of the first-order plots all became identical to that for phenoxyacetic acid. Also, the optical densities at certain wave lengths dropped below that which would be calculated for either component during the first two to four days and then rose toward the value which would be calculated for phenol. A crude rate study was made of the debromination of 2-bromo-, 4-bromo- and 2,4-dibromophenol as was done for the chlorophenols. The half-lives were of the order of 20, 30 and 360 minutes, respectively. Since these rates were so large in comparison to the rates of hydrolysis, the runs were recalculated on the basis that the absorptions were due only to the presence of the bromoacid and phenol. Although this made a marked difference in the calculated concentrations of the bromoacids, the first-order plots still showed curvature, and the debromination of the bromophenoxyacetic acids must have been proceeding at a rate competitive with the ether hydrolysis. Since the first runs were made by measuring the absorptions at the wave lengths of the ϵ_{max} of the bromophenoxyacetate and bromophenoxide, and since the concentration of phenoxyacetic acid was needed now as well, it became necessary to use the initial concentration of the bromoacid in the calculations. To illustrate, in the case

of the 2-bromophenoxyacetic acid, measurements had been made at 273.5 and 295 mμ. Since both 2-bromophenoxyacetate and phenoxyacetate absorb at less than 1% of the absorption of phenoxide at 295 mμ, the reading at this wave length was used to calculate the concentration of phenol at each point during the run. These concentrations were used to calculate the optical density due to phenol at 273.5 mμ. When subtracted from the observed values the optical densities due to the presence of the two acids were obtained. Likewise, the sum of the concentrations of the two acids at each point was determined as the difference between the initial concentration of the bromoacid for the run and concentration of phenol at each point. For each point a hypothetical density was calculated for each acid at 273.5 mμ for a contration equal to the total concentration of acids. By comparing these optical densities with the observed values which had been corrected for the presence of phenol, the percentage of each acid contributing to the total could be obtained. This method introduced quite a bit of uncertainty, but the first-order plot for the disappearance of the bromoacid now became straight. The 4-bromophenoxyacetic acid run was handled similarly, but no effort was made to untangle the 2,4-dibromophenoxyacetic acid case because of the difficulties of handling such a multicomponent system.

A second set of runs were made with the monobromophenoxyacetic acids in which additional wave lengths were used for the measurement of the optical densities, as was done with the chlorophenoxyacetic acids. The extra wave lengths used allowed measurement of the concentration of the bromophenols in order to whether see there was justification for ignoring them as contributing to the optical densities. To illustrate, the concentration of 4-bromophenol was found to be of the order of 3–5% of the total phenols present at the maximum. Since this difference was barely outside the limit of error, the assumption that no bromophenol was present was considered justified, and the total phenol concentration was thenceforth considered to be only phenol.

Discussion and Results

The rates of hydrolysis of the ether linkage in the phenoxyacetic acids were slow but were of the same magnitude as the value of $(6.6 \pm 1.5) \times 10^6$ found by Tronow and Ladigina[4] for the second-order rate constant for anisole and hydrobromic acid (aqueous) in acetic acid at 20°. Ghaswalla and Donnan[5] obtained first-order constants of about 10^{-5} sec.$^{-1}$ for chloro- and bromoanisoles at 120° with aqueous hydrochloric acid in acetic acid. Hydrobromic acid was about six times faster.

Since the observed rate constants for the hydrolysis of the bromophenoxyacetic acids were for the over-all disappearance by two paths, hydrolysis and reduction, it was necessary to consider the kinetics of the reaction. The scheme employed is shown in equations 1–4.

$$\underset{\text{BA}}{BrC_6H_4COOH} \xrightarrow{k_1} \underset{\text{BP}}{BrC_6H_4OH} + ICH_2COOH \quad (1)$$

$$BA \xrightarrow{k_2} \underset{\text{A}}{C_6H_5OCH_2COOH} \quad (2)$$

$$BP \xrightarrow{k_3} \underset{\text{P}}{C_6H_5OH} \quad (3)$$

$$A \xrightarrow{k_4} P \quad (4)$$

Since the reduction, k_3, was known to be more than 100 times greater than the hydrolysis, k_4, and since k_1, hydrolysis, and k_4 should be about the same magnitude, it was safe to assume $k_3 >> k_1$. Such being the case, for kinetic analysis, the rate constant for the hydrolytic disappearance of BA *via* BP to P could be considered to be k_1, and the constant for the total rate of disappearance of BA became $k_1 + k_2$. Assuming all the steps to be first order in the presence of the very large excess of hydriodic acid

$$-d[BA]/dt = k_1[BA] + k_2[BA] = (k_1 + k_2)[BA] \quad (5)$$

and

$$d[A]/dt = k_2[BA] - k_4[A] \quad (6)$$

BA was obtained as a function of t by integration of (5), which when substituted into (6) gave a differential equation which could be solved to give (7) which defines the concentration of A at any time as a function of the initial concentration of

$$[A] = \frac{k_2[BA_0]}{k_4 - (k_1 + k_2)}(e^{-(k_1+k_2)t} - e^{-k_4t}) \quad (7)$$

BA, of time and of rate constants. Knowing the concentrations of A at various times, and since k_4 was already known from the independent determination of the rate of hydrolysis of phenoxyacetic acid, and using the average value of $(k_1 + k_2)$ as determined from (5), (7) may be rearranged and solved for k_2.

$$k_2 = \frac{[A](k_4 - (k_1 + k_2))}{[BA_0](e^{-(k_1+k_2)t} - e^{-k_4t})} \quad (8)$$

The results as shown in Table II were so erratic that no attempt was made to pursue the study at 115°. It is interesting that the rates of reduction of the bromophenoxyacetic acids exceeded the rates of hydrolysis, especially in the *p*-position.

TABLE II

HYDROLYSIS OF BROMOPHENOXYACETIC ACIDS AT 100°

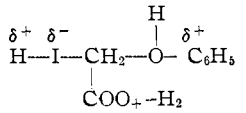
		$k_1 + k_2$, 10^6 sec.$^{-1}$	k_1, 10^6 sec.$^{-1}$	k_2, 10^6 sec.$^{-1}$
2-Bromo-	I	6.3 ± 0.4	2.4 ± 0.8	4.0 ± 0.4
phenoxy-	II	7.5 ± .5	2.5 ± .9	5.0 ± .4
acetic acid	Av.	6.9 ± 1.0	2.4 ± .9	4.5 ± .9
4-Bromo-	I	18.8 ± 0.6	1.8 ± 1.4	17.0 ± 0.8
phenoxy-	II	30.4 ± 2.5	0.8 ± 4.0	29.6 ± 1.5
acetic acid	Av.	24.6 ± 7.4	1.3 ± 3.2	23.3 ± 7.5

Since the chlorophenoxyacetic acids hydrolyzed faster than the competing reduction, and since little or no reduction occurred with the fluorophenoxyacetic acids, the rates of reduction were strongly influenced by the nature of the halogen substituent while the rates of hydrolysis remained relatively constant. The hydrolysis rate constants and the calculated energy factors are shown in Table III.

The range of rate constants was so small that the constancy of $\Delta F^{\ddagger}$ was no surprise. The energies of activation varied within a range of about 3 kcal. and the entropies of activation had a range of 9 e.u. Substitution of a halogen lowered $\Delta H^{\ddagger}$ and $\Delta S^{\ddagger}$ in every instance although the rates varied erratically. If the acid is doubly protonated as it could be in such a medium as was used here, the transition state may be pictured as

(4) B. W. Tronow and L. W. Ladigina, *Ber.*, **62**, 2844 (1929).

(5) R. P. Ghaswalla and F. G. Donnan, *J. Chem. Soc.*, 1341 (1936). See also D. M. Birosel, THIS JOURNAL, **52**, 1944 (1930), and **53**, 1408 (1931), who found *p*-bromoanisole > *o*-bromoanisole > 2,4-dibromoanisole with aqueous hydrobromic acid in acetic acid at steam-bath temperature.



However, such a doubly charged ion would be certain to have some bulky hydriodic acid molecules or iodide ions closely associated with it near the centers of positive charge. This effect could easily provide some steric hindrance to the proper approach of the entering hydriodic acid or iodide ion and result in high $\Delta H^{\ddagger}$ values. Halogen substitution in the ring, specially if in the *o*-position, should be able to help stabilize the proton on the ethereal oxygen by weak hydrogen bonding and lower $\Delta H^{\ddagger}$ somewhat; fluorine should be better than chlorine in this respect and the data agree. In the same way the loss in entropy in going into the transition state for *o*-fluorine should be greater than for *o*-chlorine because of the greater rigidity imparted to the transition state. For halogen in the *p*-position, the partial hydrogen bonding effect would be absent and only the electrical effects would operate.

TABLE III

HYDROLYSIS RATE CONSTANT FACTORS

Substituted phenoxyacetic acid	k^a 115°, 10^6 sec.$^{-1}$	k^b 100°, 10^6 sec.$^{-1}$	E_a, kcal.	log pZ, 100°	$\Delta H^{\ddagger}$, kcal.	$\Delta S^{\ddagger}$, e.u., 100°	$\Delta F^{\ddagger}$, kcal.
None	10.85[c]	2.23	30.4	12.12	29.6	− 5.5	31.7
2-F	9.14	2.23	27.1	10.20	26.3	−14.3	31.6
4-F	10.87	2.36	29.3	11.53	28.6	− 8.2	31.7
2,4-diF	10.14	2.42	27.5	10.48	26.7	−13.0	31.6
2-Cl	8.90	2.01	28.5	11.02	27.8	−10.6	31.7
4-Cl	7.61	1.74	28.3	10.82	27.6	−11.5	31.9
2,4-diCl	6.68[d]	1.61[e]	27.3	10.19	26.6	−14.3	31.9

[a] Average deviations ≈ 0.6. [b] Average deviations ≈ 0.08. [c] Average of 10.82 and 10.89. [d] Average of 7.63 and 5.72. [e] Average of 1.32 and 1.90.

A Hammett plot of the data *vs.* σ-values gives a scatter of points. However, as discussed by Jaffé,[6] the slope of the line from a 4-halophenoxyacetic acid to phenoxyacetic acid should be the same as that from a 2,4-dihalophenoxyacetic acid to a 2-halophenoxyacetic acid. For F, at 115° the values are 0.013 and 0.74; at 100°, 0.40 and 0.58. For Cl, at 115°, the values are −0.68 and −0.55; at 100°, −0.48 and −0.43. Except for the case of F at 115°, the agreement is fair.

Although the data do not fit a Hammett plot, a plot of $\Delta H^{\ddagger}$ *vs.* $\Delta S^{\ddagger}$ gives a straight line. A least squares fit gave a slope of 125° and an intercept of 32 kcal. If the slope is taken as the "isokinetic temperature"[7] at which the Hammett relationship fails, the temperatures selected for the study of the reaction were not the best possible choice. However, it is of interest that the linear relationship between $\Delta H^{\ddagger}$ and $\Delta S^{\ddagger}$ holds for the *o*-substituted acids as well as for the non-*o*-substituted materials.

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(6) H. H. Jaffé, *Chem. Revs.*, **53**, 191 (1953).
(7) J. E. Leffler, *J. Org. Chem.*, **20**, 1202 (1955).

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The Hydrazinolysis of Methyl Acetate

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The reaction of methyl acetate and hydrazine in benzene solution in the presence of methanol has been studied. The reaction was first order in both hydrazine and ester when the molar ratio of ester to hydrazine was 1:2, but became first order in ester and second order in hydrazine at ratios of 1:1 and 2:1.

In previous work the catalytic effects of hydroxylated solvents on the rates of hydrazinolysis of ethyl acetate were determined.[1] The energies and entropies of activation for these reactions were calculated also. In that work the ratio of ester to hydrazine was 1 mole to 2 moles in every case and the reactions followed second-order kinetics, being first order with respect to each of the reactants. In the present study of the hydrazinolysis of methyl acetate, it was decided to follow the reactions dilatometrically. As long as the ratio of ester to hydrazine was maintained at $1M$:$2M$, the reactions were second-order. When the ratio was changed to $2M$:$1M$, the reaction order changed and the dilatometric method became inapplicable. As a result, a chemical titration method was utilized. In this way it was found that the reactions became third order when the ratio of ester to amine was $1M$:$1M$ and $2M$:$1M$. In these cases the reactions were first order with respect to the ester and second order with respect to the hydrazine.

Solvents of the type *n*-hexane and cyclohexane would have been desirable for these reactions because of the non-solvating nature of these solvents. The insolubility of hydrazine in these solvents rendered them useless, however, and benzene was finally chosen as solvent. Even with benzene it was found that a minimum of 5 M methanol was necessary to achieve homogeneity. While this imposed certain limits on the system, it had certain advantages also. For example, it made it possible to observe the effect of methanol on the reaction rates. The use of methanol rather than some other alcohol avoided any complication of the reaction kinetics which might otherwise have occurred as the result of ester interchange. Molar ratios, ester to hydrazine, of 1:2, 1:1 and 2:1 were used and the methanol concentration was varied from 5 M to 19.1–21.4 M (pure methanol).

(1) R. A. Ferren, J. G. Miller and A. R. Day, THIS JOURNAL, **79**, 70 (1957).