Pseudo-First-order Alkaline Hydrolysis of Diethyl Tartrate: A Baseline Study for a Polymer Matrix Used in Controlled-Release Delivery Systems

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Abstract ☐ The hydrolysis kinetics of a bifunctional group compound, diethyl tartrate, was studied as a function of temperature and pH in the alkaline region. A pH-stat was used to maintain constant pH conditions in the alkaline region. This allowed the studies to be carried out at low ionic strengths and without the use of buffers. The results indicate that the hydrolysis for both steps followed specific base catalysis. The ratio of the two rate constants was 13.31, which was attributed to a strong charge effect in the second step. The results also show that the use of an overall average rate constant may not be acceptable for multifunctional group compounds.

The hydrolysis kinetics of a bifunctional group compound, diethyl tartrate, was studied at different pH and temperature conditions. This work was initiated as a result of difficulties encountered in the degradation studies of a tartrate polyester matrix which is being used in our laboratories to develop drug delivery systems. Semilog plots of the hydrolysis product, (H^+), as a function of time were not linear, indicating that an overall rate constant determined from experimental data would be a function of time and therefore not useful in describing the time course of hydrolysis (see Figure 1). This suggested that the individual microrate constants were not equal and that neighboring ester groups and/or their products were influencing the rate of reaction. Since the nearest neighboring group exerts the maximum influence, it was reasoned that the simplest system to study the effect of neighboring ester or acidic groups with minimal experimental and analytical difficulties would be the diester compound.

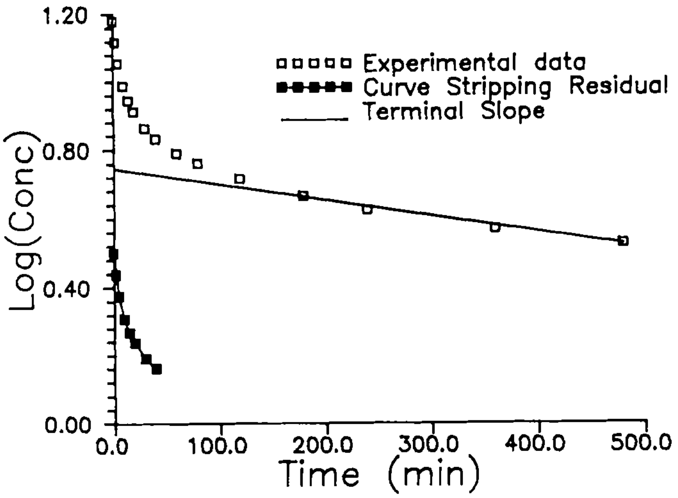

Figure 1—Alkaline hydrolysis of tartrate polyester at 45 °C and pH 9 using a pH-stat to maintain constant pH. The concentration of H^+ remaining to be produced (or equivalent number of ester bonds to be hydrolyzed) is plotted as a function of time.

A search of the literature showed that previous studies of diester hydrolysis were conducted under conditions of high ionic strength, buffer concentrations, extreme acidic pH conditions (usually pH <1), and mixed solvents.[1-11] Because of the limitations of the above studies and the need for data as a function of pH in the alkaline region which were not available in the literature, it was decided to conduct the necessary diester work in the alkaline region as baseline studies for the experimental design and analysis of the polyester development program being contemplated. A pH-stat technique was adopted to eliminate the need for high buffer concentrations to maintain a constant pH. The elimination of buffer requirements also would permit the ionic strength of the reaction media to be kept minimal.

The purpose of this paper is to evaluate a pH-stat technique to study diester hydrolysis under alkaline conditions, to determine accurately the value of the two individual rate constants, and to report the effect of pH and temperature on the two rate constants calculated from the hydrolysis data.

Experimental Section

Materials—Diethyl tartrate was obtained from Eastman Kodak, Rochester, New York, and was tested for purity by titrating to a neutral pH with standard NaOH.

pH-stat—A pH-stat (Radiometer A/S, Copenhagen, Denmark) equipped with Auto-burette (type ABU 1b), titrator (type ABU 1c), and titrigraph (type SBR 2c) was used for all experiments in the hydrolysis studies.

Reaction Vessel—The hydrolysis studies were carried out in a jacketed beaker (Figure 2) which was equipped with an inlet and an outlet for water circulation. Water maintained at constant temperature was circulated through the jacketed beaker from a water bath. The top lid had provisions for a thermometer and pH electrode, an

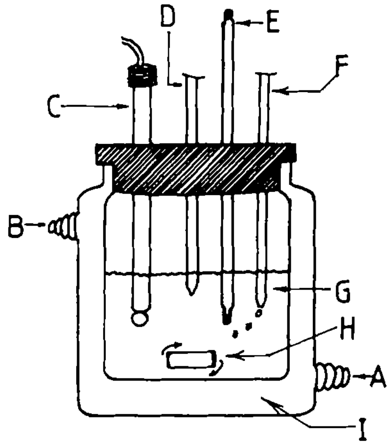

Figure 2—The reaction vessel used for hydrolysis studies: (**A**) water inlet; (**B**) water outlet; (**C**) electrode; (**D**) autoburette tip; (**E**) thermometer; (**F**) nitrogen inlet; (**G**) reaction media; (**H**) magnetic stirrer; (**I**) water circulating at constant temperature.

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inlet for the burette to titrate the hydrolyzed diester to the required pH, and an inlet for nitrogen gas to remove any carbon dioxide absorbed from the atmosphere, as absorbed carbon dioxide will change the pH significantly at high pHs.

Procedure for Kinetic Experiments—Two-hundred milliliters of water was poured into the reaction vessel and constantly stirred with a magnetic stirrer. The thermometer, electrode, nitrogen inlet, and autoburette tip were set in place. The initial pH was adjusted by titrating water to the required pH. The system was allowed to equilibrate for 30 min before starting the experiment.

Diethyl tartrate is a viscous liquid that cannot be poured through a pipette. For this reason, the diester was introduced in the reaction vessel by using a syringe. The syringe was weighed before and after adding the diester to the reaction vessel to obtain by difference the weight of the diester added to the reaction system.

After initiating the experiment, the volume of NaOH used by the pH-stat to reach and maintain the reaction mixture at a fixed pH was automatically recorded as a function of time. For analysis, the data were read off the pH-stat recordings.

The initial concentration, $\mathrm{Di_o}$, of the diester was obtained from the weight of the diester introduced in the reaction system. As a check, the initial concentration of the diester was also obtained at the end of the reaction by determining the total amount of carboxylic acid initially present by refluxing the diester overnight with an excess of NaOH and then back titrating the mixture to a neutral pH using standard HCl.

Theoretical Section

Background and Experimental Limitations—Diester hydrolysis is a two step process. Under pseudo-first-order conditions, the reaction is expressed by the following:

$$\underset{\text{(Di)}}{\text{diester}} \xrightarrow{k_1} \underset{\text{(Mo)}}{\text{Monoester}} \xrightarrow{k_2} \underset{\text{(Ac)}}{\text{Carboxylic Acid}} \qquad (1)$$

where k_1 and k_2 are observed first- and second-step rate constants. The rate equations for this reaction are:

$$\mathrm{d(Di)}/\mathrm{d}t = -k_1 t \qquad (2)$$

$$\mathrm{d(Mo)}/\mathrm{d}t = k_1\mathrm{(Di)} - k_2\mathrm{(Mo)} \qquad (3)$$

$$\mathrm{d(Ac)}/\mathrm{d}t = k_2\mathrm{(Mo)} \qquad (4)$$

These rate equations can be integrated to obtain the concentration of each species as a function of time. For $\mathrm{(Di)} = \mathrm{(Di_o)}$ at $t = 0$, the integrated equations are given below:

$$\mathrm{(Di)} = \mathrm{(Di_o)}\mathrm{e}^{-k_1 t} \qquad (5)$$

$$\mathrm{(Mo)} = \frac{\mathrm{(Di_o)}k_1}{k_2 - k_1}\left(\mathrm{e}^{-k_1 t} - \mathrm{e}^{-k_2 t}\right) \qquad (6)$$

$$\mathrm{(Ac)} = \mathrm{(Di_o)}\left[1 + \frac{1}{k_1 - k_2}\left(k_2\mathrm{e}^{-k_1 t} - k_1\mathrm{e}^{-k_2 t}\right)\right] \qquad (7)$$

The concentrations of (Di), (Mo), and (Ac) can be easily calculated if the rate constants, k_1 and k_2 are known. But, conversely, calculating the rate constants from the experimental data has been a difficult problem. The data obtained from most experimental studies measure the sum of the products appearing in the first and second step. For example, in a study of the hydrolysis kinetics of a diester under alkaline conditions, the concentration of hydrogen ions produced as a function of time includes the hydrogen ions produced in both the first and second step of the reaction.

As the reaction proceeds at a constant pH, an equivalent amount of standard NaOH is required to neutralize the hydrogen ions produced to maintain pseudo-first-order conditions. The amount of NaOH used (F) as a function of time can be expressed by the following equation:

$$F = \mathrm{(Mo)} + 2\mathrm{(Ac)} \qquad (8)$$

Substituting the values of (Mo) and (Ac) from eqs 5 and 6 yields the following:

$$F = \mathrm{(Di_o)}\left[2 - \frac{k_1 - 2k_2}{k_1 - k_2}\cdot\mathrm{e}^{-k_1 t} - \frac{k_1}{k_1 - k_2}\cdot\mathrm{e}^{-k_2 t}\right] \qquad (9)$$

Rearranging the above equation will give:

$$2\mathrm{(Di_o)} - F =$$

$$(Di_o)\left[\frac{k_1 - 2k_2}{k_1 - k_2}\cdot\mathrm{e}^{-k_1 t} + \frac{k_1}{k_1 - k_2}\cdot\mathrm{e}^{-k_2 t}\right] \qquad (10)$$

The term $[2\mathrm{(Di_o)} - F]$ is the amount of equivalent ester units or hydrogen ions remaining to be produced for the complete hydrolysis of the diester. The methods to calculate the rate constants from eq 9 or eq 10 have been suggested by various authors[12,13] and were recently evaluated by Kalonia and Simonelli.[14,15] Their analysis showed that accurate constants from eq 10 can only be obtained in a limited range of $k_2{:}k_1$ ratios.

Acid- or base-catalyzed diester hydrolysis can be studied as a pseudo-first-order reaction by maintaining a constant pH. This is generally accomplished by using a large concentration of one of the reacting species (e.g., hydrogen ions), so that the concentration of this species does not change significantly during the reaction as compared with the relative changes in the diester concentration. This method restricts studies to the extreme conditions of acidic or alkaline pH. The use of relatively high concentrations of acid or base also results in large contributions to the overall ionic strength of the reaction mixture. Under such conditions, the ionic strength can significantly influence the rate constants. Therefore, this type of experimental technique does not permit one to investigate diester hydrolysis under different pH conditions and at low ionic strengths (<0.01).

In many cases where acid or base is produced during the reaction, a constant pH is maintained by using buffers, and the progress of the reaction can be followed by measuring one of the physical properties such as the UV absorption as a function of time. This task of maintaining a constant reaction pH is complicated for diester hydrolysis by the fact that tartaric acid and its esters do not show any characteristic peak in the UV range. Therefore, the progress of the reaction is monitored by quantitating the amount of acid produced during the reaction. If a buffer is used to maintain the reaction pH, the amount of acid produced during the reaction cannot be measured. Additionally, buffer catalysis and ionic strengths can have significant contributions. It is not uncommon to see reports where investigators use buffers concentration which have >0.2 M ionic strength.[16,17]

Desired experimental conditions were maintained by using a pH-stat which is a device that continuously monitors the reaction pH. As soon as any acid due to diester hydrolysis is produced, the pH-stat will titrate the reaction media with an equivalent amount of standard base. The amount of base used is recorded on a chart recorder. Use of a pH-stat has one major advantage over other methods for the study of hydrolysis reactions; that is, that there is little or no contribution due to the ionic strength. For most cases, the ionic strength effects will be insignificant, except when the hydrogen or hydroxyl ions contribute significantly to the overall ionic strength of the reaction mixture.

Results and Discussion

The pH-stat technique was utilized for all studies and found most suitable for diethyl tartrate hydrolysis when the reaction was studied over a time range of 30 min to 6 h. Reactions that took <30 min or >6 h exhibited errors in the data due to various reasons which are described in the next section. For diester hydrolysis studies under alkaline conditions, the use of the pH-stat proved to be an excellent technique to obtain hydrolysis kinetic data.

Limitations to the Use of the pH-Stat Technique—Despite its usefulness to study diester hydrolysis at constant pH conditions and negligible ionic strengths, some limitations were encountered with the pH-stat technique during this investigation. The first problem encountered with the pH-stat is unique to multistep reactions (two steps in this case), especially, when the first step is much faster than the later steps. For faster reactions that required 3–4 min to complete the first step, the rate of titration always seemed to lag behind the rate of hydrolysis. This type of problem can generally be corrected by changing the size of the burette or the concentration of the titrant. When this is done, however, one observes over-titration for the later, much slower, part of the reaction.

The second limitation was encountered for slower reactions, especially, when it required >6 h to complete the study. Erratic titration behavior was noticed after ~6 h of continuous use, even when the reaction vessel contained only water which should not exhibit any change in pH. These errors could be due to one or more of the following reasons: long exposure of the electrode to the reaction mixture, absorption of carbon dioxide, instrumental heating, or electronic changes in the circuitry of the pH-stat. Some of these errors can be corrected with instruments having state of the art electronic design. This is the reason that studies at lower pHs were not conducted (see Figure 3) as reactions were too slow.

At pHs >9, significant absorption of carbon dioxide was evident from blank titrations (only water present in the reaction mixture). Therefore, an additional step was required to avoid this problem. The reaction vessel used was sealed and a positive pressure was maintained by bubbling nitrogen into the reaction mixture throughout the course of the titration. Reproducible results were obtained when nitrogen was bubbled in the reaction vessel.

As the diester hydrolysis proceeded, acid produced was titrated with standard NaOH. Continuous addition of NaOH caused an increase in the total volume of the reaction mixture. The change in volume contributes to the changes in the concentration of reactants as a function of time. The concentration, therefore, must be corrected to avoid error in the calculation of the rate constants due to dilution during titration. The following equation was used to correct the concentration of NaOH as a function of time:

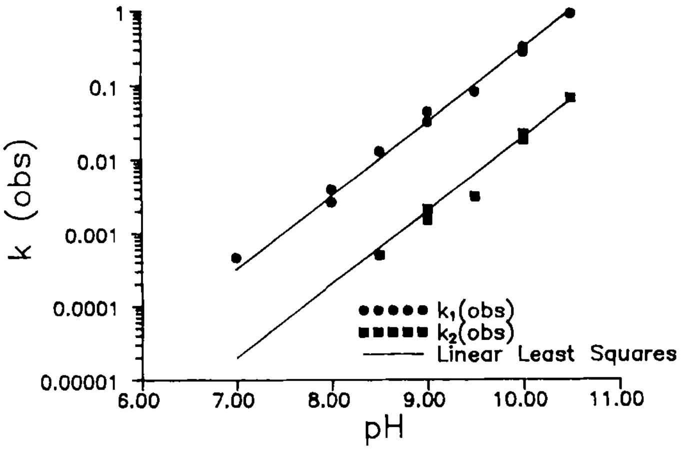

Figure 3—The pH-rate profile for diethyl tartrate hydrolysis at 55 °C using pH-stat analysis. Lines were drawn as linear least squares, forcing a slope of 1, using eq 14.

$$C_i = \frac{V_i \cdot S \cdot 1000}{V_o + V_a + V_i} \tag{11}$$

where C_i is the concentration (mmol/L) of NaOH added; S is the strength of NaOH (mmol/mL); V_i is the volume of NaOH (mL) used at time t; V_o is the initial volume (mL) of water used in the reaction vessel; and V_a is the volume (mL) of NaOH used to adjust the initial pH. The concentration of hydrogen ions produced as a function of time is equal to the concentration of NaOH used which can be calculated from the above equation.

Diester hydrolysis data obtained at pH 9 and 55 °C are shown in Table I. Column 3 of this table shows the concentration of hydrogen ions when the volume correction is used. Column 4 shows the uncorrected concentration. The readings at 200 min show a 5.8% difference between the corrected and uncorrected concentrations. This data, if not corrected for the dilution effect, can introduce a significant error in the calculated value of the rate constants, particularly at later times.

Data Analysis—The corrected hydrolysis data were plotted on a semilograithmic scale to identify characteristic shapes. In an earlier paper,[15] a comprehensive evaluation of pseudo-first-order consecutive reactions identified five different cases (Case I–V) based on the $k_2{:}k_1$ ratios. These cases could be identified from the shape of $\ln(2Di_o - F)$ versus time plots. It was established that only Case I ($k_1 >> k_2$ so that k_2 can be neglected compared with k_1) and Case II ($k_1 > k_2$ and k_2 cannot be neglected compared with k_1) data would yield correct rate constants from the coproduct data. Other cases, Case III ($k_1 = 2k_2$), Case IV ($k_1 << k_2$), and Case V ($k_1 < 2k_2$), do not yield unique solutions.

Table I—Concentration of Carboxylic Acid Produced during Diester Hydrolysis at pH 9.0 and 55 °C

Time, min	NaOH Used, mL	Acid Concentration, mmol/L		Concentration Remaining (2A − F)[c]
		Corrected[a]	Uncorrected[b]	
0.00	0.00	0.000	0.000	6.5347
2.00	0.78	0.340	0.342	6.1944
4.00	1.50	0.652	0.657	5.8826
6.00	2.04	0.884	0.894	5.6503
8.00	2.58	1.116	1.130	5.4191
10.00	3.12	1.346	1.367	5.1892
20.00	4.98	2.128	2.181	4.4065
30.00	6.24	2.650	2.733	3.8843
40.00	7.08	2.995	3.101	3.5397
50.00	7.62	3.215	3.338	3.3197
60.00	8.13	3.422	3.561	3.1129
70.00	8.52	3.579	3.732	2.9555
80.00	8.82	3.700	3.863	2.8348
90.00	9.09	3.808	3.981	2.7264
100.00	9.36	3.916	4.100	2.6184
120.00	9.84	4.108	4.310	2.4269
140.00	10.32	4.298	4.520	2.2364
160.00	10.74	4.464	4.704	2.0704
180.00	11.22	4.653	4.914	1.8814
200.00	11.67	4.830	5.111	1.7051

[a] Corrected concentration of hydrogen ions as a function of time calculated from amount of NaOH (mL) used. [b] Uncorrected concentration of hydrogen ions produced as a function of time. [c] Concentration of carboxylic acid remaining to be produced for complete hydrolysis.

It was found that the log $(2Di_o - F)$ versus time plots indicate that Case I applies for the alkaline hydrolysis data as they exhibited Case I shapes. Therefore, Case I analysis was used for the calculation of rate constants. The initial estimates for the rate constants were obtained by KSTRIP, an interactive curve-stripping program that can be used on microcomputers (KSTRIP is available from the author). The estimated values were then used in TKSolver,[18] an equation solving program, to obtain accurate rate constants. For reactions below pH 8.0, sufficient data were not available for the second step because the reaction was too slow. In such cases, k_1 was calculated from the initial data independent of k_2. Since k_1 is much greater than k_2, initial data (as much as 40% depending on the $k_2{:}k_1$ ratio) will have negligible contribution from the very slow second step. Therefore, the initial part of the reaction can be treated as a one-step process for which eqs 2 and 5 will apply. If the reaction indeed follows this analysis, a plot of ln (H^+ remaining to be produced) for the first step versus time should be linear. As the $k_2{:}k_1$ ratio increases, the terminal portion of the plot will start curving, indicating that there is significant contribution from the second step. In such cases, only the linear segment should be used to calculate the first rate constant, k_1.

Effect of pH on the Rate Constants—Hydrolysis studies were conducted at different pH values at 55 °C. The pH–log (rate) profiles are shown in Figure 3. It can be seen that the profiles for both rate constants, k_1 and k_2, are linear, with slopes close to 1. This indicates a first-order hydroxyl ion catalysis in this region. The pseudo-first-order behavior in the high pH region can be explained by analyzing the specific acid-base catalysis given by the following equation:

$$k_{obs} = k_o + k_{OH^-}(OH^-) + k_H + (H^+) \tag{12}$$

where k_{obs} is the experimental rate constant, k_o is the spontaneous first-order rate constant, and k_{H^+} and k_{OH} are the specific acid and base second-order rate constants, respectively.

Under alkaline conditions, the concentration of hydrogen ions will be much smaller than hydroxyl ions. The slope of the pH profile was 1, indicating that the OH^--catalyzed reaction is dominant and, therefore, that the contributions from the terms $(k_{H^+})(H^+)$ and k_o can be assumed to be negligible compared with the contribution from the term $(k_{OH^-})(OH^-)$. The above equation would thus be reduced to:

$$k_{obs} = k_{OH}(OH^-) \tag{13}$$

or

$$\log(k_{obs}) = \log(K_w \cdot k_{OH^-}) + \mathrm{pH} \tag{14}$$

When the above condition is true, the plot of log (k_{obs}) versus pH becomes linear with a slope equal to 1. Examination of Figure 3 shows that eq 14 is applicable, indicating that the hydrolytic reaction follows specific base hydrolysis in the alkaline pH region.

The specific base rate constants, $k_{1(OH)}$ and $k_{2(OH)}$, can be calculated for the first and second steps from eq 13 by plotting k_{obs} versus OH^-. The plots will be linear with slopes equal to $k_{1(OH)}$ and $k_{2(OH)}$, respectively. The calculated values of the specific rate constants were as follows: $k_{1(OH)} = 392.01$/min and $k_{2(OH)} = 29.38$. The two specific rate constants exhibit a ratio of 13.31. In the absence of any extraneous effect, the two rate constants should exhibit a ratio of 2 based on the statistical factor (i.e., the diester has two sites available for hydrolysis and the monoester has only one site). These data clearly show that the determination of an overall average rate constant for multifunctional kinetic studies can lead to highly erroneous results.

The difference in two specific rate constants can be explained on the basis of the commonly accepted mechanism for base-catalyzed ester hydrolysis,[19] according to which the rate-determining step is the addition of hydroxide ion to the carbonyl group of the ester. The diester molecule involved in the first step is a neutral molecule under various pH conditions, whereas the monoester molecule involved in the second step can exist both in the ionized and unionized forms. In the alkaline region above pH 8, the monoester exists in >99.99% ionized form. The ionized carboxyl group carries a negative charge which exerts a field effect. This field effect repels the hydroxyl ions and slows the attack on the carboxyl carbon of the ester group in the monoester molecule.

Effect of Temperature on the Observed Rate Constants—Hydrolysis studies were conducted at various temperatures. The energy of activation for the reaction was calculated from the slope of the Arrhenius plot. A plot of ln (k_{obs}) versus $1/T$ is shown in Figure 4. The energy of activation, ΔH_{act}, for the first step was 17.6 kcal/mol and for the second step 18.9 kcal/mol. These energies are within the range characteristic for most esters hydrolyzing via the A_{AC^2} mechanism. The activation energy for the second step was found to be slightly higher than the first step ($\Delta H_{act(1)} - \Delta H_{act(2)} = 1.3$ kcal/mol). Although this appears to be a small difference, linear least squares analysis indicated that this difference was statistically significant. In addition, if one attempts to force parallel lines through k_1 and k_2 data, the fits of the curve are sufficiently poor to support rejection of an hypothesis of no difference in the respective enthalpies. This difference, if real, can be rationalized if one assumes that the energy in the transition state is influenced by the negative charge present on the monoester molecule. Since the attacking hydroxyl ion also has a negative charge, the molecule in the transition state is highly energized. The thermodynamic data of activation (i.e., enthalpies, $\Delta H^\ddagger$, and entropies, $\Delta S^\ddagger$) were also calculated. The following equation was used to calculate $\Delta H^\ddagger$ and $\Delta S^\ddagger$:

$$k = \frac{RT}{h} \cdot e^{\frac{\Delta S^\ddagger}{R}} \cdot e^{\frac{-\Delta H^\ddagger}{RT}} \tag{15}$$

The enthalpies of activations were $\Delta H_1^\ddagger = 17.07$ kcal/mol and $\Delta H_2^\ddagger = 18.35$ kcal/mol, and entropies were $\Delta S_1^\ddagger = -245$ cal/deg · mol and $\Delta S_2^\ddagger = -273$ cal/deg · mol. The highly neg-

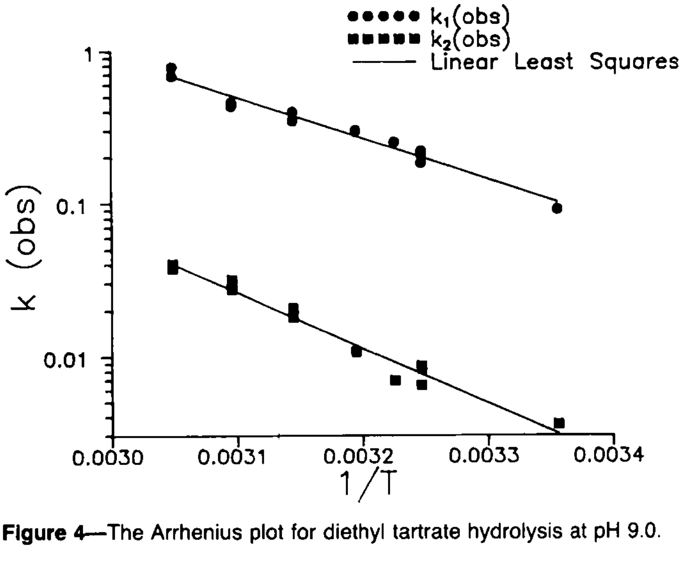

Figure 4—The Arrhenius plot for diethyl tartrate hydrolysis at pH 9.0.

ative values of $\Delta S^{\ddagger}$ are in concordance with the A_{AC^2} mechanism.

Conclusions

The pH-stat technique is easy and effective for the study of diester hydrolysis. Its main advantages are that the reactions can be studied without the use of buffers and data can be readily obtained. The first step rate constant for the base-catalyzed diethyl tartrate hydrolysis is much faster than the second step rate constant which clearly showed that the use of overall average rate constants of multifunctional groups reactions are not acceptable. The difference in the two rate constants was attributed to the charge effect exhibited by the ionized monoester molecule.

References and Notes

1. Maugh, Tom, II; Bruice, Thomas C. *J Am. Chem. Soc.* 1971, *93(13)*, 3237–3248.
2. Radhakrishnamurti, P. S.; Patro, P. C. *J. Indian Chem. Soc.* 1969, *46(10)*, 903–908.
3. Radhakrishnamurti, P. S.; Patro, P. C. *J. Indian Chem. Soc.* 1971, *48(9)*, 812–816.
4. Khalil, Fayez Y.; Assad, Adel N. *Z. Naturforsch.* 1982, *37a*, 390–394.
5. Khalil, Fayez Y.; Assad, Adel N. *Z. Naturforsch.* 1984, *39a*, 95–101.
6. Khalil, Fayez Y.; Assad, Adel N. *Z. Naturforsch.* 1985, *40a*, 84–91.
7. Rao, G. Venkoba; Venkatasubramanian, N. *Ind. J. Chem.* 1972, *10*, 178–181.
8. Ingold, C. K. *J. Chem. Soc.* 1931, 2170–2179.
9. Homan, J. D. H. *Rec. Trav. Chim.* 1944, *63*, 181–188.
10. Newberger, Mark R.; Kadkec, Robert H. *AIChE J.* 1973, *19(6)*, 1272–1275.
11. Kucherov, V. F.; Yanovskaya, L. A. In *The Chemistry of Carboxylic Acid and Esters*; Patai, Saul, Ed.; Interscience: New York, 1969; p 193.
12. Swain, C. G. *J. Am. Chem. Soc.* 1944, *66*, 1696.
13. Pearson, Ralph G.; Moore, John W. *Kinetics and Mechanism;* Wiley: New York, 1981; p 300.
14. Kalonia, D. S.; Simonelli, A. P. *J. Pharm. Sci.* 1988, *77(12)*, 1055–1061.
15. Kalonia, D. S.; Simonelli, A. P. *J. Pharm. Sci.* 1989, *78(1)*, 78–84.
16. Nelson, N. M.; Bundgaard, Hans *J. Pharm. Sci.* 1988, *77(4)*, 285–298.
17. Nelson, N. M.; Bundgaard, Hans *Int. J. Pharm.* 1987, *39*, 75–85.
18. TKSolver plus; Universal Technical Systems; Rockford IL, 1987.
19. Kirby, A. J. In *Comprehensive Chemical Kinetics*; Bamford, C. H.; Tipper, C. H., Eds.; Elsevier: New York, 1972; p 162.