drochloride solution necessary to give a 0.500 M solution was transferred to a 250-ml. volumetric flask; the solution also was made 0.500 M with respect to pyridine; final addition of methanol was made at the reaction temperature. The ketone solution (4.00 × 10⁻²M) and the thiosemicarbazide hydrochloride solution (2.00 × 10⁻²M) were made up in the buffered methanol. Equal volumes of the two solutions were mixed for the reaction. Samples of 1 ml. were removed at intervals and diluted 100 times with ethanol-water (1:1), and the optical density was read against a similarly diluted sample of the original buffer solution. Runs were made at 24.97 and 38.90 ± 0.02° (34.83° for the semicarbazone). The reading of the solution on a Beckman pH meter was 3.49 ± 0.02 throughout the reaction. The rate constants were reproducible within 1–2%; activation energies are believed to be accurate to ±0.4 kcal./mole.

[Contribution from Cobb Chemical Laboratory, University of Virginia, Charlottesville, Va.]

The Mechanism of Solvolysis of Nitrostyrenes

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The kinetics of hydrolysis of 3,4-methylenedioxy-β-nitrostyrene, to piperonal and nitromethane, are characteristic of two consecutive, pseudo-first-order reactions over the pH range −0.8 to 6. The first is reversible and shows general base catalysis, with the rate constant a non-linear function of the acetate ion concentration. The rate of the second is pH dependent. A mechanism is proposed consistent with these observations.

The base-catalyzed cleavage of 3-methoxy-4-hydroxy-β-nitrostyrene in strongly alkaline solution was shown by Stewart[1] to proceed with attack of a hydroxyl ion via the colorless intermediate nitroalcohol. The fact that this cleavage will take place slowly in acid solution[2a] proves, however, that a slightly different mechanism must be possible. In the course of some preliminary work on nitrostyrene formation in acetate buffers (our previous work dealt only with amine catalysis[2]) we had occasion to study the hydrolysis in acid solution and can report new kinetic features.

Experimental

The nitrostyrene used in this work was 3,4-methylenedioxy-β-nitrostyrene (piperonylidenenitromethane), which was prepared by condensation of piperonal and nitromethane.[3] Buffer solutions and hydrochloric acid solutions were prepared from reagent-grade chemicals and distilled water. The ionic strength, here equal to the sodium ion concentration, was kept constant at 0.1 M by adding sodium chloride, except for a series of runs at 0.3 M.

The kinetic runs were started by adding 1 ml. of a freshly prepared 0.002 M methanol solution of nitrostyrene to the aqueous buffer components and adding water to make 100 ml. The flasks were placed in the thermostat and samples analyzed spectrophotometrically without further dilution, at 372 mμ. The optical density at this wave length decreased as the nitrostyrene hydrolyzed, and at first no peak at 312 mμ appeared to indicate the presence of piperonal unless the pH was as high as 6. Piperonal gradually was produced at lower pH, however, but only after a considerable quantity of an intermediate had formed.

In very acidic solution (3 and 6 M HCl) the λmax of the nitrostyrene shifted to 385 mμ.

Results

The hydrolysis of nitrostyrene showed complex kinetics, though all the reactions involved were pseudo first order because of the high dilution of the substrate in comparison with the buffer components. Figure 1 shows typical logarithmic plots of concentration vs. time. At very low pH, equilibrium is reached when about 60% of the nitro-

(1) R. Stewart, J. Am. Chem. Soc., **74**, 4531 (1952); see also a study of nitrochalcones by E. A. Walker and J. R. Young, J. Chem. Soc., 2045 (1957).

(2) (a) T. I. Crowell and F. A. Ramirez, J. Am. Chem. Soc., **73**, 2268 (1951); (b) T. I. Crowell and D. W. Peck, ibid., **75**, 1075 (1953).

(3) N. A. Lange and W. E. Hambourger, ibid., **53**, 3865 (1931).
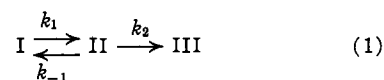
Table I

RATE CONSTANTS (SEC.⁻¹) FOR NITROSTYRENE HYDROLYSIS

[HAc], M	HCl, M	[Ac⁻], M	pH^a	$10^6 k_1$	$10^6 k_{-1}$	$10^6 k_2$	k_1/k_{-1}
0.005		0.05	5.6	3.22	3.86	4.36	0.83
.003		.03	5.6	4.25	6.37	3.72	.67
.001		.01	5.6	1.92		c	
.1		.1	4.6	6.34	8.94	0.44	.71
.05		.05	4.6	5.08	7.78	.42	.65
.03		.03	4.6	4.03	6.42	.47	.63
.01		.01	4.6	2.16	3.47	.43	.62
.3		.3	4.5^b	6.94			
.2		.2	4.5^b	6.94			
.12		.12	4.5^b	6.83			
.07		.07	4.5^b	6.04			
.02		.02	4.5^b	3.84			
.01		.01	4.5^b	2.03			
.008		.008	4.5^b	1.69			
.006		.006	4.5^b	1.39			
.004		.004	4.5^b	0.91			
.002		.002	4.5^b	.58			
.25		.025	3.6	3.16	4.83	0.075	0.65
.05		.005	3.6	1.24	2.13	.103	.58
	0.0012		2.9	0.181	0.42	.114	.44
	0.012		1.9	.186	.42	.114	.44
	0.12		0.9	.181			.53
	3.0		−0.5	.091		c	
	6.0		−0.8	.086		$k_2 > k_1$	
	0.0012		2.9^d	1.92			
	.012		1.9^{d,e}	0.15			
	.006		2.2^{d,e}	.12			

^{a} Using 2.8 × 10⁻⁵ for K_{HAc} in 0.1 M NaCl (M. Kilpatrick and R. D. Eanes, J. Am. Chem. Soc., **75**, 586 (1953)). ^{b} Ionic strength 0.3, K_{HAc} = 3.2 × 10⁻⁵. ^{c} k_2 to close to k_1 for accurate determination. ^{d} Temperature 45°. ^{e} Solvent 99% D₂O.

styrene has reacted, and a plot of log $(x_e - x)$ is linear, indicating opposing first-order reactions.

At a slightly higher pH, a slow reaction becomes apparent after the first equilibrium is established. This reaction is so rapid at pH 7 that the first step is rate controlling and simple first-order kinetics are obtained. These observations may be represented by the scheme

$$I \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} II \overset{k_2}{\longrightarrow} III \qquad (1)$$

which is confirmed by the spectral changes described in the Experimental section.

Fig. 1.—Logarithmic plots of nitrostyrene concentration: curve A, pH 0.9; B, 3.6; C, 4.6; D, 5.6; E, 7.7.

Fig. 2.—Variation of k_1 with acetate ion concentration: ○, pH 3.6; ●, pH 4.5; ○, pH 4.6; ◐, pH 5.6.

The value of k_1 usually was calculated by assuming a value of 0.6 for k_1/k_{-1}, plotting as for a reversible first-order reaction and measuring the initial slope. By the approximation of McDaniel and Smoot,[4] k_2 then was evaluated and finally k_{-1}, this last value being an improvement on the tentative assumption for k_1/k_{-1}. These constants are given in Table I for all runs.

The first step shows general base catalysis: k_1 depends only on the acetate ion concentration between the pH values 3.6 and 5.6. In the absence of any base except the solvent, for example in 0.1 M or 0.01 M hydrochloric acid solutions, k_1 attains a minimum value of 0.183 × 10⁻⁶ sec.⁻¹. The relationship between k_1 and acetate ion concentration, given by the points in Fig. 2, is not a linear one. The constancy of k_1/k_{-1} at different pH, within the limits of experimental error, is an indication that this first step is not an ionization nor an addition of hydroxyl ion.

The second step, on the other hand, proceeds at a rate which depends not on the acetate ion but on the pH. The values of k_2 given in Table I seem proportional to hydroxyl ion concentration at higher pH, but also increase in very acidic solutions.

At 45° in 10⁻³M hydrochloric acid, k_1 is 1.92 × 10⁻⁶ sec.⁻¹, yielding an apparent activation energy of 22.6 kcal./mole. In 99% deuterium oxide, the corresponding value at 45° is 0.13 × 10⁻⁶ sec.⁻¹. The ratio k_H/k_D is 14 ± 2.

Methanolysis.—Many experiments were carried out in methanol. The methanolysis shows the general features of the hydrolysis and is only slightly accelerated by 1 to 5% water. The reverse reaction, condensation of piperonal and nitromethane, clearly shows the formation of an

(4) D. H. McDaniel and C. R. Smoot, *J. Phys. Chem.*, **60**, 966 (1956).

intermediate, presumably the nitroalcohol, in acetate buffers in methanol. Only a small fraction of the stoichiometric concentration of nitrostyrene was formed; whether this was due to unfavorable equilibrium with the nitroalcohol or to a competing methanolysis, possibly forming a methyl ether, was not determined.

Discussion

We explain the above results in terms of the mechanism shown in the accompanying diagram. The

$$ArCH{=}CHNO_2 \text{ (I)} + H_2O \underset{k_{-4}}{\overset{k_4}{\rightleftharpoons}} \underset{\underset{B}{\overset{+}{\underset{|}{OH_2^+}}}}{ArCH\bar{C}HNO_2}$$

$$\underset{\underset{OH}{|}}{ArCHCH_2NO_2} \text{ (II)} \underset{k_{-5}}{\overset{k_5}{\rightleftharpoons}} \underset{\underset{OH}{|}}{ArCH\bar{C}HNO_2} \text{ (IIa)}$$

$$\updownarrow K_2 \qquad\qquad \overset{+}{HB^+}$$

$$\underset{\underset{O^-}{|}}{ArCHCH_2NO_2} \overset{k_6}{\longrightarrow} ArCHO + \bar{C}H_2NO_2 \rightleftharpoons CH_3NO_2$$

$$\text{IIb} \qquad\qquad \text{III}$$

first stage of eq. 1 is separated into three steps: an initial hydration which proceeds only to a slight extent ($k_4 \ll k_{-4}$); removal of a proton by a base, in these experiments water or acetate ion, so that $k_5 = \Sigma k_b[B]$; and protonation at the carbon atom to form the nitroalcohol II. As shown by Zollinger for the azo coupling reaction,[5] such a system would

(5) (a) H. Zollinger, *Helv. Chim. Acta*, **38**, 1597, 1617, 1623 (1955); (b) R. P. Bell, "The Proton in Chemistry," Cornell University Press, Ithaca, N. Y., 1959, p. 139.

follow the rate equation

$$\frac{1}{k_1} = \frac{1}{k_4} + \frac{k_{-4}}{k_4 \Sigma k_b [B]} \qquad (2)$$

Accordingly, the values for k_1 were analyzed to give equation 3 which is plotted as the curve in Fig. 2.

$$\frac{1}{k_1} = \frac{1}{8.18 \times 10^{-6}} + \frac{1}{1.7 \times 10^{-7} + 2.78 \times 10^{-4} [OAc^-]} \qquad (3)$$

In the denominator of the second term appear the ratios $k_4 k_b / k_{-4}$ for the bases water and acetate ion. These rate constants cannot be separated; it is certain, however, that k_4/k_{-4} is very small and k_5 large.

At low base concentration, proton removal is rate controlling and the rate linear in base concentration. As the amount of base becomes large, hydration becomes rate controlling and k_1 approaches the limiting value k_4. The isotope effect on k_1 in the absence of acetate would be determined by the nucleophilic strength of D_2O in the hydration step, its basicity in the second step, the usual effect on the breaking of a carbon–hydrogen bond, and finally the effect of changing the medium from

H_2O to D_2O. The observed value of k_H/k_D seems consistent with this mechanism.

The rate constant k_2 for the second stage of the reaction would be $K_2 k_6 / [H^+]$, which seems to account for k_2 above pH 4. It is quite reasonable that an acid-catalyzed dehydration would become significant below pH 4.

Preliminary studies of more basic solutions (pH 11 to 13) showed a hydroxyl-ion catalyzed fading of the nitrostyrene color. This reaction differs from the hydrolysis already described in that the nitrostyrene is regenerated if the solution is reacidified immediately. If an acidic partially hydrolyzed solution of nitrostyrene is treated with base for a minute or two, acidification restores only the concentration of nitrostyrene recently present. Longer exposure to base yields piperonal. These experiments indicate a hydroxyl ion adduct other than the form IIb if our formulation of the intermediate II is correct. One possibility would be $I \underset{H^+}{\overset{OH^-}{\rightleftarrows}} IIa$. These acid–base reactions deserve further study.

The Arrhenius Parameters of the Deuterium Isotope Rate Effect in a Base-promoted Elimination Reaction[1]: Evidence for Proton Tunnelling

By V. J. SHINER, JR.,[2] AND MORRIS L. SMITH[3]

The rates of reaction of 1-bromo-2-phenylpropane and the 2-deuterio analog with sodium ethoxide in ethanol have been determined from 5 to 55°. The reaction product was shown to be quantitatively α-methylstyrene with no competing substitution reaction. By reference to the infrared spectra taken with a lithium fluoride prism instrument the tertiary carbon–hydrogen and carbon–deuterium vibrational stretching frequencies of these compounds were assigned as 2904 and 2160 cm.$^{-1}$, respectively. The approximately doubly-degenerate bending frequencies of these bonds were tentatively assigned as 1299 and 934 cm.$^{-1}$, respectively. The experimentally determined Arrhenius equations are: $\log k_H = 11.033 - (20,655/2.303 RT)$; $\log k_D = 11.443 - (22,406/2.303RT)$; $\log (k_H/k_D) = -0.437 + (1,790/2.303RT)$. An upper limit for the isotope activation energy effect, neglecting tunnelling, was estimated using the observed ground state vibrational frequencies and found to be significantly less than that obtained experimentally. The frequency factors are abnormal in that the one for the deuterium compound is almost three times that for the hydrogen analog. These apparent discrepancies between the experimental and theoretical activation parameters are rationalized on the basis of non-classical barrier penetration by the proton. The proton tunnelling through the potential barrier effectively increases the activation energy difference for the two isotopes. The general utility of the deuterium isotope rate effect in mechanistic studies is not greatly altered by the occurrence of tunnelling.

The hydrogen isotope rate effect has been developed into a fairly general and powerful tool for the investigation of reaction mechanism.[4] Application of the absolute rate theory[5] allows the theoretical calculation of the magnitude of the isotope rate effect[6,7] providing that the configurations

and vibrational frequencies of the initial and transition states are known. Since the frequencies of the initial state and the isotope rate effect are subject to experimental observation while the transition state frequencies are not, the theory is generally used to infer something about the latter from the first two. In addition, where independent information on the structure of the transition state is available, observation of the isotope rate effect provides a test of the absolute rate theory.

A much more exacting test of the theory is provided by a comparison of the observed and predicted isotope Arrhenius parameters. Little definitive evidence of this type is available in spite of the wide interest in deuterium isotope effects. One obvious difficulty is the experimental one of measuring rates

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(3) Abstracted in part from a thesis submitted by Morris L. Smith to the Graduate School of Indiana University in partial fulfillment of the requirements for the Ph.D. degree.

(4) For a general review see K. B. Wiberg, Chem. Revs., 55, 713 (1955).

(5) S. Glasstone, K. J. Laidler and H. Eyring, "The Theory of Rate Processes," McGraw–Hill Book Co., Inc., New York, N. Y., 1941, pp. 153–201.

(6) J. Bigeleisen, J. Chem. Phys., 17, 675 (1949).

(7) L. Melander, Arkiv. Kemi., 2, 211 (1950).