In addition there was obtained 15.0 g. of unchanged alcohol and 9.5 g. of a high-boiling material of b. p. 179–185°. This latter fraction was a mixture of the acids formed during the oxidation and their esters with the alcohols present in the starting material. It had an 18% acid and an 82% ester content based on its acid value and saponification value.

An oxidation of 88 g. (1 mole) of alcohol with $[\alpha]_D$ −3.55° corresponding to 59.2% (−)2-methylbutanol-1 gave 35 g. of aldehyde with rotation $[\alpha]_D$ +17.73°, 14.5 g. of unchanged alcohol and 24 g. of high boiling residue. The yield of aldehyde was 49% of the theoretical.

Similarly, a third oxidation of 88 g. (1 mole) of amyl alcohol with $[\alpha]_D$ −1.79° corresponding to 29.8% (−)2-methylbutanol-1 yielded 30 g. of aldehyde with $[\alpha]_D$ +8.72°, 11 g. of unchanged alcohol and 21 g. of high boiling material. The aldehyde was obtained in 40% yield. The rotations of the three aldehydes *versus* the rotations of the alcohols from which they were prepared is shown graphically in Fig. 1. Since the aldehydes were susceptible to oxidation by atmospheric oxygen, they were stored in a nitrogen atmosphere in glass-stoppered flasks. Ehrlich,[1] by oxidation of an alcohol with $[\alpha]_D$ −5.47°, obtained only a 15% yield of aldehyde with $[\alpha]_D$ +23.6°. The extrapolated value for pure (+)2-methylbutanal-1 from Fig. 1 is $[\alpha]_D$ +31.2° as compared with the 24% lower value of $[\alpha]_D$ +23.6° obtained by Ehrlich. No experimental details are given for comparison in the work of Levene and Kuna.[2]

The unrecrystallized 2,4-dinitrophenylhydrazones of the three aldehydes obtained had $[\alpha]_D$ +30.3° (c = 4.89, acetone, 1-dm. tube), m. p. 125.5–126.5° from the aldehyde with $[\alpha]_D$ +28.50°; $[\alpha]_D$ +17.5° (c = 4.84, acetone, 1-dm. tube), m. p. 111.5–113.5° from the aldehyde with $[\alpha]_D$ +17.5°; $[\alpha]_D$ +8.06° (c = 4.96, acetone, 1-dm. tube), m. p. 114–115.5°, from the aldehyde with $[\alpha]_D$ +8.72°. For the determination of the constants of the 2,4-dinitrophenylhydrazone of pure (+)2-methylbutanal-1, the 2,4-dinitrophenylhydrazone ($[\alpha]_D$ +30.3°) from the aldehyde of rotation $[\alpha]_D$ +28.50° was recrystallized three times, once from 80% ethanol and twice from 90% ethanol when a melting point of 132.5–133° was obtained. The rotation of this derivative was $[\alpha]_D$ +32.1° (c = 4.99, acetone, 1-dm. tube); the value obtained by extrapolation from the curve of $[\alpha]_D$ of the derivative *versus* $[\alpha]_D$ of the aldehyde was $[\alpha]_D$ +32.2°.

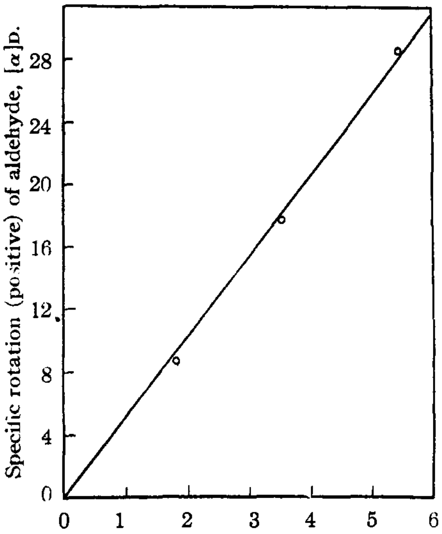


Fig. 1.

Anal. Calcd. for $C_{11}H_{14}N_4O_4$: C, 49.57; H, 5.30; N, 21.04. Found: C, 49.35; H, 5.32; N, 21.50.

Summary

The optically active aldehyde, (+)2-methylbutanal-1, has been prepared by oxidation of active amyl alcohol and the optical rotation of the pure aldehyde and the properties of its 2,4-dinitrophenylhydrazone determined.

PRINCETON, NEW JERSEY RECEIVED MARCH 20, 1945

[CONTRIBUTION FROM THE FRICK CHEMICAL LABORATORY OF PRINCETON UNIVERSITY]

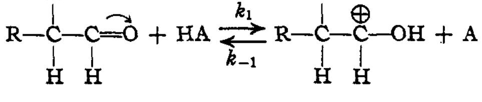
Kinetics and Mechanism of the Acid Catalyzed Racemization of (+)2-Methylbutanal-1

BY ELMER J. BADIN AND EUGENE PACSU

In probing the mechanism of hydrogenation and dehydrogenation of aldehydes and alcohols,[1] it was found necessary to prepare the optically active aldehyde, (+)2-methylbutanal-1. This aldehyde was found to racemize readily on a nickel catalyst surface. A study of the acid catalyzed racemization of this aldehyde was undertaken since no data on the rate of racemization of an aldehyde have previously been reported. The study has permitted further information to be obtained in regard to the extent and nature of solvation in acid catalyzed reactions in addition to information regarding the racemization reaction.

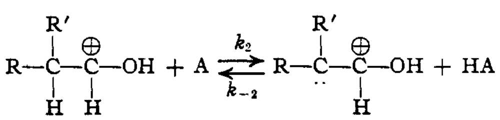
The rate in alkaline medium was found to be much greater than in acid solution; work in this investigation has dealt mainly with the rate of acid catalyzed racemization in aqueous dioxane solvents. Results indicated that the general mechanism applying for acid catalysis was best represented by the following steps where HA represents a molecule of acid and A its conjugate base.

(1) Addition of a proton to the carbonyl oxygen



$$\mathrm{R{-}\underset{H}{\overset{R'}{C}}{-}\underset{H}{C}{=}O + HA \underset{k_{-1}}{\overset{k_1}{\rightleftarrows}} R{-}\underset{H}{\overset{R'}{C}}{-}\underset{H}{\overset{\oplus}{C}}{-}OH + A}$$

(2) Removal of a proton from carbon atom 2



$$\mathrm{R{-}\underset{H}{\overset{R'}{C}}{-}\underset{H}{\overset{\oplus}{C}}{-}OH + A \underset{k_{-2}}{\overset{k_2}{\rightleftarrows}} R{-}\underset{\cdot\cdot}{\overset{R'}{C}}{-}\underset{H}{\overset{\oplus}{C}}{-}OH + HA}$$

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[1] Badin and Pacsu, THIS JOURNAL, **66**, 1963 (1944).

Racemization probably occurs in step (2) in which the carbanion carbon tends to "flatten out," causing an increasing rate of racemization with increasing temperature. Either (1) or (2) is rate determining and, if it may be assumed that the addition of a proton to the carbonyl oxygen is fast, that is, that k_1 is greater than k_2, then step (2) would be rate determining. This mechanism is essentially that postulated by Schwab, Taylor and Spence[2] and by Hammett[3] involving addition of a proton to the carbonyl oxygen and subsequent elimination of a proton on carbon atom 2, but differs in that the "enol" has been written in polarized form.

Rates of racemization were determined in dioxane solutions containing various amounts of water. Rate constants were obtained using $\log[\alpha_0]_D - \log[\alpha]_D = 0.4343kt$ where $[\alpha]_D$ is the specific rotation. A plot of log $[\alpha]_D$ *versus* time in seconds gave values of k in reciprocal seconds. The results are shown in Table I. Values of $\Delta H^\ddagger$, the heat of activation for the reaction, were calculated and from these values corresponding changes in entropy, $\Delta S^\ddagger$, during the formation of the activated complex, were obtained from the equation for absolute reaction rates.[4]

TABLE I

RATES OF RACEMIZATION OF (+)2-METHYLBUTANAL-1 IN WATER–DIOXANE MIXTURES CONTAINING HYDROGEN CHLORIDE

Expt.	Temp., °C.	Mole fraction, dioxane	$[OH_3^+]$, N	$k \times 10^5$, (sec.$^{-1}$)	$k' \times 10^5$ (sec.$^{-1}$) (mole liter^{-1})$^{-1}$
1	25.0	0.274	0.0692	0.330	4.77
2	25.0	.274	.140	0.640	4.58
3	35.3	.274	.140	1.63	11.6
4	25.0	.274	.603	5.88	9.75
5	25.0	.500	.0695	1.11	16.0
6	36.3	.500	.0695	3.54	50.9
7	25.0	.649	.140	3.77	26.9
8	36.0	.649	.140	18.6	133
9	25.0	.960	.139	54.2	390
10	36.1	.960	.139	64.0	460
11	25.0	.984	.148	15.9	107
12	25.0	1.00	.148	4.67	31.6

$$k = \frac{k_0T}{h}\, e^{-\Delta H^\ddagger/RT}\, e^{\Delta S^\ddagger/R} \qquad (1)$$

In this equation k is the specific reaction rate constant, k_0 is the Boltzmann constant, T the absolute temperature, h Planck's constant, and $\Delta S^\ddagger$ and $\Delta H^\ddagger$ are the entropy and heat of activation, respectively.

The values of $\Delta S^\ddagger$ obtained in the case of aqueous dioxane solutions indicated that a highly solvated activated complex existed. Variations in rate constant, heat of activation, and entropy of activation with change of dioxane–water concentration have been obtained. Values of $\Delta H^\ddagger$ and $\Delta S^\ddagger$ are summarized in Table II. Increasing concentration of dioxane indicated a change in the amount of solvation in the activated state as compared with the reactant molecule; a maximum in the rate–solvent curve occurred when the composition of the solvent was varied.

TABLE II

THERMAL AND ENTROPY DATA FOR THE ACID CATALYZED RACEMIZATION OF (+)2-METHYLBUTANAL-1 IN WATER–DIOXANE MIXTURES

Mole fraction dioxane	$[OH_3^+]$, N	$\Delta H^\ddagger$, kcal./mole	$\Delta S^\ddagger$, entropy units
0.274	0.140	16.6	−22.6
.500	.0695	19.0	−12.2
.649	.140	26.6	+14.4
.960	.139	2.7	−60.4

Experimental

Materials.—The preparation of the aldehydes used in these experiments is described in the preceding article.[5] Rate measurements were carried out on an aldehyde with $[\alpha]_D$ +28.50° (homogeneous), b. p. 90–92°, $d^{25}{}_4$ 0.8055, and containing 90.3% 2-methylbutanal-1 and 9.7% 3-methylbutanal-1. Another aldehyde with $[\alpha]_D$ +8.72° and containing 29.8% 2-methylbutanal-1 was used in some of the experiments. The aldehydes were susceptible to air oxidation and were stored in glass-stoppered flasks flushed with nitrogen; an amount of acid amounting to less than 1% was present in the aldehyde used.

Experiments were carried out mainly in aqueous dioxane solution; the dioxane used was acid free and was purified according to Fieser[6] and fractionated from sodium.

Method of Rate Measurement.—Reactions were carried out by adding a weighed quantity of the aldehyde at the required temperature to the reaction solvent contained in a 50-ml. glass-stoppered Erlenmeyer flask in a water-bath thermostatically maintained at the required temperature ±0.05°. Timing of the reaction with a stop watch was coincident with the addition of the aldehyde. After mixing, the homogeneous sample was transferred to a metal-jacketed one-decimeter polarimeter tube through which water from the constant temperature bath was circulated by means of a comparatively high speed centrifugal pump. The polarimeter tube was stoppered with a tinfoil covered stopper to prevent access of air and evaporation of the solution during the experiment. Rotations were taken at various times using the D sodium line of wave length 5893 Å. Results on homogeneous samples were within ±0.01°.

Reaction in Aqueous Ethanol Solution.—Racemizations were first carried out in ethanol–water solutions. An experiment was carried out in aqueous ethanol (containing 0.365 mole fraction ethanol and 0.635 mole fraction water) 0.710 N with respect to hydrochloric acid. A 0.7220-g. sample of aldehyde, $[\alpha]_D$ +28.50°, was dissolved in 9.90 ml. of the reaction solvent. The reaction at 25.6° was characterized by a sharp drop in rotation followed by a more gradual decrease in slope. Results were

Time, min.	α_D	$[\alpha]_D$
2.5	+1.06°	15.86°
12	+0.84°	12.57°
51	+0.78°	11.67°
2456	+0.68°	10.17°

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(2) Schwab, Taylor and Spence, "Catalysis," D. Van Nostrand Co., Inc., New York, N. Y., 1937, p. 124.

(3) L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, p. 232.

(4) Wynne-Jones and Eyring, *J. Chem. Phys.*, **3**, 492 (1935).

(5) Badin and Pacsu, THIS JOURNAL, **67**, 1352 (1945).

(6) L. F. Fieser, "Experiments in Organic Chemistry," 2nd ed., D. C. Heath and Co., New York, N. Y., 1941, p. 369.

A graph of log $[\alpha]_D$ *versus* time permitted the data to be represented by a curve consisting of two straight lines. The rate constant for the initial portion of the curve was estimated as 142×10^{-5} sec.$^{-1}$; for the second portion of the curve k was 0.098×10^{-5} sec.$^{-1}$. This indicated compound formation between alcohol and aldehyde.

Reactions in Aqueous Dioxane Solution.—Since experiments in ethanol solution indicated compound formation, dioxane was used for the subsequent experiments. In one experiment 10.00 ml. of aqueous dioxane solution, 0.140 N with respect to hydrogen chloride and containing 0.274 mole fraction dioxane and 0.726 mole fraction water, and 0.645 g. aldehyde with rotation $[\alpha]_D$ +28.50° were used. The temperature was maintained at 25.0 ± 0.05°. The results of this experiment are given in Table III. The specific rotation has been calculated from the formula

$$[\alpha]_D = \alpha_D\ [10.00 + 0.800]/0.645$$

TABLE III

RATE OF RACEMIZATION, 0.140 N HYDROGEN CHLORIDE, 0.274 MOLE FRACTION DIOXANE

Reaction time, min.	Observed rotation (α_D)	Specific rotation ($[\alpha]_D$)
0	.....	......
3	+1.77°	+29.63°
5	+1.74°	+29.13°
14.5	+1.72°	+28.79°
88	+1.67°	+27.96°
278	+1.53°	+25.61°
1357	+0.96°	+16.07°
2872	+ .52°	+ 8.70°
3183	+ .46°	+ 7.70°
4421	+ .31°	+ 5.19°
5721	+ .20°	+ 3.35°

These results are shown in Fig. 1 together with those of other experiments carried out at various acid concentrations.

Anhydrous hydrogen chloride in dioxane was obtained by passing dry hydrogen chloride gas into anhydrous dioxane. A portion of the hydrogen chloride–dioxane solution was then diluted to the required acid concentration with dioxane. The solution containing 0.984 mole fraction dioxane was prepared by adding 0.0807 g. of water to 27.376 g. of anhydrous hydrogen chloride–dioxane 0.148 N with respect to hydrogen chloride. Other solutions were prepared by adding aqueous hydrochloric acid to anhydrous dioxane.

Discussion and Results

When ethanol was used as the solvent, the initial sharp drop in rotation indicated compound formation of some type which was attributed to hemiacetal formation when ethanol and the aldehyde were mixed. Adkins and Broderick,[7] have shown by density and refractive index measurements that hemiacetal formation occurs simply by mixing an aldehyde and an alcohol. A plot of log $-d[\alpha]_D/dt$ *versus* log $[[\alpha]_D - [\alpha_0]_D]$ where $[\alpha]_D$ is the specific rotation and $[\alpha_0]_D$ is the specific rotation at that point where the second portion of the curve begins, had a slope equal to one. This indicated that the first portion was first order; the kinetics and mechanism in ethanol solution were not investigated further.

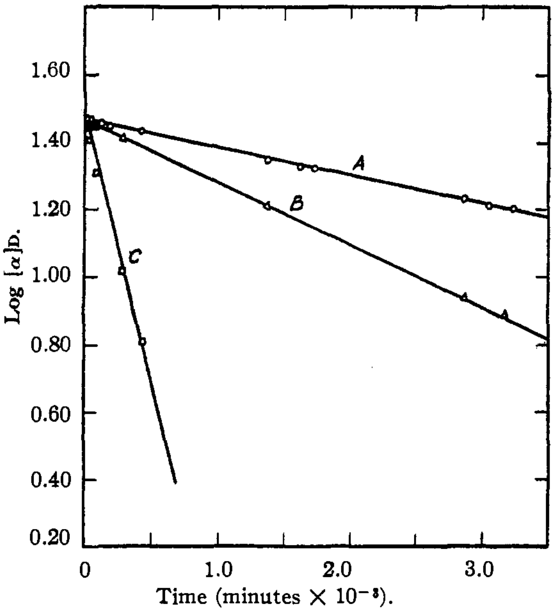

Fig. 1.—Variation of rate with acid concentration: Curve A, $[OH_3^+]$ 0.0692; curve B, $[OH_3^+]$ 0.140; curve C, $[OH_3^+]$ 0.603; temperature 25.0°; 0.274 mole fraction of dioxane.

The rate constants for the racemization in dioxane solution have been expressed as $k = k'[OH_3^+]$. Use of the Christiansen formulation[8] leads to the expression

$$-\frac{1}{[\alpha]}\frac{d[\alpha]}{dt} = \frac{k_1k_2[HA]}{k_{-1} + k_2} \tag{2}$$

where $[\alpha]$ is the specific rotation. Based on the integrated form of (2), a plot of log $[\alpha]_D$ *versus* time in seconds gave values of k in reciprocal seconds. Equation (2) implies proportionality between the specific rate constant for the reaction and the acid concentration. Actually a plot of log k for the reaction *versus* log $[OH_3^+]$ was not linear. Braude,[9] similarly, has found that in the aniontropic acid catalyzed rearrangement of propenylethynylcarbinol, the logarithm of the specific rate constant was not proportional to the logarithm of the acid concentration. A consideration of the equilibria involved in acid catalyzed reactions would undoubtedly lead to a result differing from (2) and accounting for deviation from linearity of the log k *versus* log $[OH_3^+]$ curve.

Dioxane was chosen as the reaction solvent because of its inertness to the aldehyde molecule. In the place of $[OH_3^+]$, however, it might be assumed that the dioxanonium ion would exert a catalytic influence. In the absence of water, a stronger base than dioxane, the basicity of dioxane should permit formation of, and catalysis by, the dioxanonium chloride. When water is present, a competition between oxonium ion formation and dioxanonium ion formation results. It would be expected that the stronger of the two bases

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(7) Adkins and Broderick, THIS JOURNAL, **50**, 499 (1928).

(8) Christiansen, *Z. physik. Chem.*, **28B**, 303 (1935); Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, p. 107.

(9) Braude, *J. Chem. Soc.*, 443 (1944).

would unite with a proton and thus be the better catalyst. The basicity of dioxane has been placed by Gordy[10] in a scale of basicities rated on the magnitude of the shifts which various solvents produced in the OD fundamental band of CH_3OD. Gordy and Martin[8] have further shown from infra-red absorption of hydrogen chloride in solution that hydrogen chloride is not completely ionized in dioxane but that the binding force between the hydrogen and the chlorine has been weakened perceptibly by association with the solvent. Bartlett and Dauben[11] have pointed out that in a system containing dioxane and hydrogen chloride a great deal of ion association into clumps or clusters of the type suggested by Fuoss and Kraus[12] might occur. Therefore, dioxane may be taken as a considerably weaker base than water although the rate in anhydrous dioxane was appreciable as shown by Expt. 12.

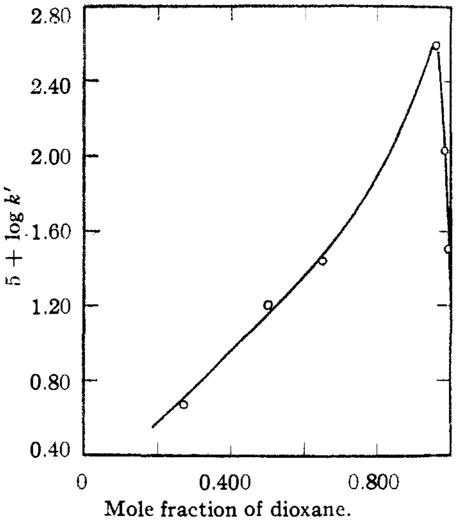

Fig. 2.

Values of the rate constant, k', have been expressed in sec.$^{-1}$ mole^{-1} liter but it should be emphasized that the concentrations rather than the activities of the acid have been used since incomplete data are available for activity coefficients of hydrogen chloride in water–dioxane mixtures. By increasing the $[OH_3^+]$ an increase in rate resulted; by further variation of acidity from *p*H 1 to *p*H 14 it would be expected that a catalytic catenary would result. That is, the rate would be extremely fast at both high basicity and high acidity but that at some intermediate value a minimum would result if the rate constant were plotted *versus* *p*H. The variation of rate with OH_3^+ concentration is given in Fig. 1. When log k was plotted *versus* log $[OH_3^+]$, the points deviated from a straight line. Expressing the absolute rate equation exactly for reactions in solution

$$k = \frac{k_0T}{h}\frac{\gamma_A\gamma_B}{\gamma^{\ddagger}}e^{-\Delta H^{\ddagger}/RT}e^{\Delta S^{\ddagger}/R} \qquad (3)$$

where $\gamma_A\gamma_B$ is the product of the activity coefficients of the reactants and $\gamma^{\ddagger}$ is the activity coefficient of the activated complex, it is obvious that k should involve the activity of the hydrogen chloride.

Variation of the amount of dioxane and water in the reaction solvent led to a curve, as shown in Fig. 2, with a maximum in the rate at approximately 0.96 mole fraction dioxane. At dioxane concentrations higher than this the rate rapidly decreased. This maximum can be interpreted as due primarily to changes in solvation of the activated complex resulting from the varying activities of the two components, water and dioxane, as the solvent was changed. Further it may be concluded that, since a slight amount of water increased the rate greatly, the oxonium ion is a far better catalyst than the dioxanonium ion.

Values of heats of activation, $\Delta H^{\ddagger}$, and entropies of activation, $\Delta S^{\ddagger}$, are tabulated in Table II. These results show that at the maximum rate corresponding to the peak in Fig. 2, $\Delta H^{\ddagger}$ is 2.7 kilocalories per mole and $\Delta S^{\ddagger}$ is -60.4 entropy units. When more water is present, $\Delta S^{\ddagger}$ becomes more positive and then decreases again. Since activities of hydrogen chloride were not considered in determining the rate constants, the apparent values of $\Delta S^{\ddagger}$ obtained will differ from the true values of $\Delta S^{\ddagger}$ by a term involving the activities

$$\Delta S^{\ddagger}_{\text{apparent}} = \Delta S^{\ddagger} + R \ln (\gamma_A\gamma_B/\gamma^{\ddagger}) \qquad (4)$$

The values of $\Delta S^{\ddagger}$ are indirect measures of the amount of solvation, large negative values indicating a great amount of "freezing" or association of solvent molecules with aldehyde molecules during reaction.

Summary

The kinetics and mechanism of the racemization of (+)2-methylbutanal-1 in water–dioxane solutions have been studied at various $[OH_3^+]$ concentrations and solvent compositions. A maximum in the rate curve was obtained when the solvent composition was varied throughout the entire range. Values of heats of activation, $\Delta H^{\ddagger}$, and entropies of activation, $\Delta S^{\ddagger}$, for the reaction were determined. The maximum in rate and changes in the values of $\Delta S^{\ddagger}$ have been attributed to changes in solvation between the reactant state and the activated state.

PRINCETON, NEW JERSEY RECEIVED MARCH 20, 1945

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[10] Gordy, *J. Chem. Phys.*, **7**, 93 (1939); Gordy and Martin, *ibid.*, **7**, 99 (1939).

[11] Bartlett and Dauben, THIS JOURNAL, **62**, 1339 (1940).

[12] Fuoss and Kraus, *ibid.*, **55**, 2387 (1933).