Further experimental data, over a wider temperature range, if feasible, or in other solvents would seem desirable. It does not appear, from the standard error given in Table V, that the failure of eq. 19 and 20 is due to lack of precision in measurements of ΔH^* and ΔS^*. The largest standard errors in ΔH^* and ΔS^* were found for reaction 23. However, in all regression analyses involving reaction 23, no significant improvement was obtained by omitting the data for 23. Thus, it is indicated that the values of ΔH^* and ΔS^* for reaction 23 are fairly reliable.

Finally, it may be of interest that the ranges in values of the activation parameters given in Table VI show generally that structural changes in this reaction series result in larger changes in ΔF^* than in ΔH^* and $T\Delta S^*$.

Experimental

Materials.—The materials used in this study have been described previously,[5,6] except for 3,5-dinitrobenzoic acid which was obtained from a commercial source and was recrystallized from toluene until the melting point agreed with previously reported values (204–205°).

Rate Measurements.—A Beckman Model DU spectrophotometer was used to follow the reactions. The spectrophotometer was equipped with thermospacers and the cell compartment was thermostatted at the desired temperature by circulating water from a constant-temperature bath through the thermospacers. The reaction temperatures were maintained within ± 0.01° of the values given in Table I.

Due to low solubility in toluene, all reactions of 3,5-dinitro- and 3-nitrobenzoic acids were run in 10-cm. cells and all reactions of 3-bromo- and 4-methoxybenzoic acids were run in 5-cm. cells. All other reactions were run in 1-cm. cells. The reactant solutions were prepared and mixed as described previously[5,6] and then an aliquot of the mixture was transferred to the spectrophotometer cell which had been placed in the thermostatted cell compartment 2 hr. previously. The cell containing the reacting mixture was left in the thermostatted cell compartment until the reaction had proceeded through about 3 half-lives.

Otherwise, the rate measurements were made as described previously.[5,6]

Three to six rate determinations were made on each reaction at each temperature. The precision of the results is comparable to that reported previously.

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Hydration, Hydrate Acidity, and Aldolization of Isobutyraldehyde[1a]

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The equilibrium constants for the hydration and the aldolization of isobutyraldehyde in aqueous solution have been determined by n.m.r. and ultraviolet measurements. The acidity of isobutyraldehyde hydrate and the kinetics of the dealdolization of isobutyraldol have also been studied. From the kinetic study it appears that the combination of an isobutyraldehyde carbanion with an isobutyraldehyde molecule is comparable in rate to its protonation, so that the rate-controlling step in dealdolization may change with a change in the nature of the basic catalyst used.

In relation to a study of the α-hydrogen exchange of isobutyraldehyde it became desirable to investigate possible complications due to aldolization and hydration of the aldehyde, and the acidity of the aldehyde hydrate. The aldol condensation of acetaldehyde, which has been studied by a number of workers[2–5] has been found to be sufficiently irreversible that, even in relatively dilute aqueous solutions, the kinetics may be studied without allowance for reversibility (in the early part of the reaction, at least). The equilibrium constant for the aldolization of acetone at 25°, which climbs from about 0.023 M^{-1} in 4% water–96% acetone–diacetone alcohol to 0.037 M^{-1} in 80% water–20% acetone–diacetone alcohol,[6] is so unfavorable that the pure organic equilibrium mixture contains only about 13% diacetone alcohol and in dilute aqueous solution the dealdolization reaction can be run essentially to completion.[6–8]

Usherwood showed that the aldol condensation of isobutyraldehyde is reversible and reported that in the organic layer of the two-phase potassium carbonate catalyzed reaction about 5.5, 33.2, and 90.1% isobutyraldehyde is present at 13, 60, and 100°, respectively.[9] No measurements were made in aqueous solution, however, and the interpretation of the results obtained is complicated by the subsequent demonstration that the principal higher boiling constituent of the organic phase is not isobutyraldol but 2,6-diisopropyl-5,5-dimethyl-4-*m*-dioxanol, the product of the addition of a molecule of isobutyraldehyde to one of aldol.[10–12]

The extent of hydration of aldehydes and ketones in aqueous solution, ranging from more than 99% for chloral and formaldehyde to an amount too small to detect for acetone, has been measured or estimated by

(1) (a) Abstracted in part from the Ph.D. thesis of James G. Houston, 1965. (b) National Defense Education Act Fellow, 1960–1963.

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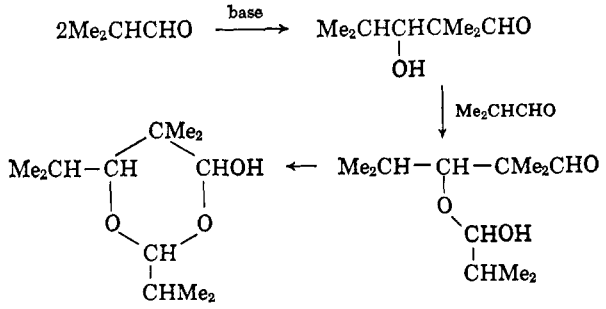

several investigators.[13-18] The acidity of several aldehyde and ketone hydrates has been determined[19,20] and, in the case of acetaldehyde, the kinetics of the base-catalyzed aldolization have been found to be complicated by the acidity of the aldehyde hydrate,[5] which neutralizes part of the basic catalyst and simultaneously decreases the fraction of the aldehyde present in the reactive "free" aldehyde form.

Results

As Lombardi and Sogo pointed out, the n.m.r. spectra of aqueous solutions of isobutyraldehyde contain peaks due to isobutyraldehyde hydrate as well as those due to the free aldehyde.[18] The integrated intensities of the peaks due to the methyl groups of these two species showed that 38 ± 1% of the aldehyde is present as the hydrate at 25° and 30 ± 1% at 35°. These figures correspond to K_h values of 0.0110 and 0.0077 M^{-1}, respectively, where $K_h = [Me_2CHCH(OH)_2]/[Me_2CHCHO]\,[H_2O]$ and to an enthalpy of hydration of −6.5 kcal./mole.

Treatment of 0.1–0.5 molar aqueous solutions of isobutyraldehyde with dilute alkali resulted in the formation of an organic layer that consisted largely of diisopropyldimethyldioxanol with smaller amounts of isobutyraldol and isobutyraldehyde. The solubility of the dioxanol in water is too low for its presence in the aqueous phase to be detected by n.m.r. measurements. The concentrations of isobutyraldehyde and isobutyraldol could be determined by n.m.r. and ultraviolet measurements. Several sets of measurements on the aqueous phase from the reaction at 35° gave aldehyde and aldol concentrations from which equilibrium constants for aldolization can be calculated. The isobutyraldehyde concentration in this expression refers only to the free aldehyde, not its hydrate. Values of K_a did not change significantly when the reaction time was more than doubled. Essentially the same K_a value was obtained when equilibrium was approached from the other side, starting with isobutyraldol.

$$K_a = \frac{[Me_2CHCHOHCMe_2CHO]}{[Me_2CHCHO]^2} = 1.05 \pm 0.1\ M^{-1}$$

Quantitative n.m.r. measurements were also found useful for studying the reaction kinetics. Since isobutyraldehyde is the predominant species present at equilibrium in homogeneous aqueous solution, it is much more convenient to study the dealdolization of isobutyraldol than the aldolization of isobutyraldehyde. Satisfactory first-order rate constants for the dealdolization reaction were not obtained when sodium hydroxide was used as the catalyst. The rate constants tended to decrease as the reaction proceeded; in very dilute base traces of oxidation destroyed part of the catalyst and in stronger base the reaction was too fast to study conveniently. The reaction rate was therefore measured at 35° in the presence of 1,4-diazabicyclo[2.2.2]octane ($C_6H_{12}N_2$) and trimethylamine buffers. In order to interpret the results, the ionization constant of the first amine was determined in water at 35° and that of the latter amine was interpolated from the measurements of Everett and Wynne-Jones at higher and lower temperatures.[21] Within a given run the reaction was found to fit the first-order rate equation. First-order rate constants (k_{fo}) were determined in three different runs with the results shown in Table I. Attribution of all of the catalysis to the hydroxide ions present leads to the calculation of the second-order rate constants (k_{so}) listed. These values are believed to be in satisfactory agreement when it is considered that an error of 0.05 in the relative pK values of the two bases used would put the third value between the first two.

TABLE I

KINETICS OF THE DEALDOLIZATION OF ISOBUTYRALDOL IN THE PRESENCE OF TERTIARY AMINE BUFFERS

[B], M	[BH$^+$], M	10^5[OH$^-$], M	$10^5 k_{fo}$, sec.$^{-1}$	k_{so}, M^{-1} sec.$^{-1}$
0.0390[a]	0.00734	5.1	15.5	3.0
0.0228[a]	0.0159	1.7	4.63	2.7
0.0086[b]	0.0070	12	41.0	3.4

[a] 1,4-Diazabicyclo[2.2.2]octane buffer. [b] Trimethylamine buffer.

At the concentrations used, the dealdolization reaction should proceed past 90% completion. No attempt was made to correct for reversibility since it should produce no more uncertainty in the first-order rate constants than that already present (probably about 10%) due to various experimental errors.

The acidity of isobutyraldehyde hydrate was determined by ultraviolet measurements on aqueous solutions of isobutyraldehyde in the presence of various concentrations of sodium hydroxide. As the hydrate is used up by the reaction shown below, additional free

$$Me_2CHCH(OH)_2 + OH^- \overset{K}{\rightleftharpoons} Me_2CHCHOHO^- + H_2O$$

aldehyde is hydrated so as to keep the free aldehyde–hydrate ratio constant. From the resultant decrease in absorbance a K value of 93 was obtained at 25°.

$$K = \frac{[Me_2CHCHOHO^-]\,[H_2O]}{[Me_2CHCH(OH)_2]\,[OH^-]} = 93$$

Complications due to aldolization should be minor. At the concentrations used, less than 10% of the aldehyde would be present as the aldol at equilibrium (according to the equilibrium constant determined at 35°). Furthermore, the extinction coefficient of the aldol, which is formed from two molecules of aldehyde, is almost twice the apparent extinction coefficient of

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(21) D. H. Everett and W. F. K. Wynne-Jones, *Proc. Roy. Soc.* (London), **A177**, 499 (1941).

isobutyraldehyde (*i.e.*, the value calculated by treating the aldehyde concentration as the sum of the hydrate and free aldehyde concentrations), so that about the same contribution to the absorbance is made by the aldol as by the equivalent amount of isobutyraldehyde in its equilibrium free aldehyde–hydrate mixture.

Discussion

The extent of hydration of isobutyraldehyde has been studied by ultraviolet measurements,[13-17] but in such cases it is necessary to estimate the extinction coefficient of the free aldehyde by measuring it in a solvent where addition to the carbonyl group cannot occur and assuming the same value for aqueous solution, or by choosing that value required to give the closest approach to a straight line in a plot of log K_h *vs.* $1/T$, or by some other means. The n.m.r. method,[18] in which both the aldehyde and hydrate are measured directly, eliminates some of the uncertainties of the ultraviolet method. The enthalpy of hydration (−6.5 kcal./mole) that is calculated from the equilibrium constants for hydration that we have determined at 25 and 35° is within the combined experimental uncertainties of that (−7.3 kcal./mole) reported recently by Gruen and McTigue.[16] Our individual K_h values are some 40% higher than those of Gruen and McTigue. Herold did not give the temperature at which his measurements were made but Gauditz, in a continuation of his work, made measurements at about 18°. Our value of K_h extrapolated to 18° is about 3% higher than that of Herold.

Although the equilibrium constant for the aldolization of acetaldehyde does not appear to have been determined quantitatively, the available qualitative observations show that the constant is much larger than what we have found for the aldolization of isobutyraldehyde. Although most reagents add to acetaldehyde to a somewhat greater extent at equilibrium than they do to isobutyraldehyde (K_h for acetaldehyde[18] is about 2.5 times as large as that for isobutyraldehyde), the much smaller equilibrium constant for the aldolization of isobutyraldehyde seems attributable largely to the steric interactions that are produced when *two* isobutyraldehyde molecules are brought together. The fact that the equilibrium constant for aldolization of isobutyraldehyde is larger than that for acetone must be due to the greater resistance of the carbonyl group of a ketone (due largely to stabilization of the carbonyl group by the second alkyl group) to addition by any reagent, *e.g.*, water.[15]

The dealdolization of isobutyraldol would be expected to proceed by the following mechanism, analogous to that for the dealdolization of diacetone alcohol[8] and the reverse of that for the aldolization of acetaldehyde.[2-5] It is assumed that the first step is at equilibrium (*i.e.*, $k_{-1} >> k_2$) so that there is no need to consider other versions of the first step in which the hydroxide ion is replaced by the other bases in the solution. Studies of the kinetics of the base-catalyzed deuterium exchange of isobutyraldehyde[22] show that under the conditions of the present kinetic experiments on dealdolization, the attack of water on isobutyraldehyde contributes negligibly to the total rate of carb-

(22) J. G. Houston, unpublished work from this laboratory.

$$Me_2CH\underset{|}{\overset{OH}{C}}HCMe_2CHO + OH^- \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} Me_2CH\underset{|}{\overset{O^-}{C}}HCMe_2CHO + H_2O$$

$$Me_2CH\underset{|}{\overset{O^-}{C}}HCMe_2CHO \underset{k_{-2}}{\overset{k_2}{\rightleftharpoons}} Me_2CHCHO + Me_2\bar{C}CHO$$

$$Me_2\bar{C}CHO + BH^+ \underset{k_{BH}}{\overset{k_{BH}}{\rightleftharpoons}} Me_2CHCHO + B$$

$$Me_2\bar{C}CHO + H_2O \underset{k_{OH}}{\overset{k_{H_2O}}{\rightleftharpoons}} Me_2CHCHO + OH^-$$

$$Me_2\bar{C}CHO + H_3O^+ \underset{k_W}{\overset{k_{H^+}}{\rightleftharpoons}} Me_2CHCHO + H_2O$$

anion formation from isobutyraldehyde. It follows from the principle of microscopic reversibility that under these conditions a negligible fraction of the carbanions derived from isobutyraldehyde are protonated by hydronium ions. Hence the steps governed by k_{H^+} and k_W may be neglected. Neglecting reversibility and making the steady-state approximation for the intermediate alkoxide ion and carbanion, the rate of disappearance of isobutyraldol may be expressed as eq. 1, where HAAH represents the aldol, HA is iso-

$$v_d = \frac{k_1k_2(k_{BH}[BH^+] + k_{H_2O})[OH^-][HAAH]}{k_{-1}(k_{-2}[HA] + k_{BH}[BH^+] + k_{H_2O})} \quad (1)$$

butyraldehyde, and BH^+ the acidic constituent of the buffer, and where reactions involving the solvent water have been treated as first order. The fact that the reaction gives satisfactory first-order rate constants during individual runs in spite of the increasing concentration of isobutyraldehyde suggests that the term $k_{BH}[BH^+] + k_{H_2O}$ is considerably larger than $k_{-2}[HA]$, that is, that protonation is the usual fate of isobutyraldehyde carbanions under the conditions in question. In such a case eq. 1 reduces to eq. 2.

$$v_d = \frac{k_1k_2}{k_{-1}}[OH^-][HAAH] \quad (2)$$

Just how negligible the term $k_{-2}[HA]$ is may be estimated as follows. The equilibrium constant for aldolization may be expressed as eq. 3 where K_B is

$$K_a = \frac{k_{-1}k_{-2}k_{OH}^H}{k_1k_2k_{H_2O}} = \frac{k_{-1}k_{-2}k_B^H}{k_1k_2k_{HB}K_B} \quad (3)$$

the ionization constant of B.

Although k_{OH}^H, the rate constant for the removal by a hydroxide ion of a proton from free isobutyraldehyde, is not available, the corresponding k_{OH}^D value ($4.7 \times 10^{-2}\ M^{-1}$ sec.$^{-1}$), relating to isobutyraldehyde-2-*d*, has been determined.[22] Since K_a and k_1k_2/k_{-1}, the second-order rate constant for dealdolization, are known, k_{H_2O}/k_{-2} may be expressed as eq. 4. Similarly,

$$\frac{k_{H_2O}}{k_{-2}} = \frac{(k_{OH}^H/k_{OH}^D)k_{OH}^D}{(k_1k_2/k_{-1})K_a} = 1.5 \times 10^{-2}(k_{OH}^H/k_{OH}^D)M \quad (4)$$

from the values of k_B^D (9.3×10^{-3} and $2.3 \times 10^{-2}\ M^{-1}$ sec.$^{-1}$, respectively)[22] and pK_B (5.15 and 4.10, respectively) for diazabicyclooctane and trimethylamine, values of 420 (k_B^H/k_B^D) and 91 (k_B^H/k_B^D) may be calculated for k_{HB}/k_{-2}. These results show that in the dealdolizations we have studied in the presence of buffers the intermediate carbanion is pro-

tonated almost every time it is formed ($k_{BH}[BH^+] + k_{H_2O} >> k_{-2}[HA]$). Even with the least favorable assumption (that $k_B{}^H/k_B{}^D$ is unity) and in the least favorable case (that of the trimethylamine buffer), the intermediate carbanions are protonated more than 90% of the time even at 80% reaction. In the absence of a buffer, however, the situation is different. According to the results in eq. 4 the dealdolization of 0.0529 M isobutyraldol in the presence of a sodium hydroxide catalyst would yield intermediate carbanions that, at 80% reaction, might combine with aldehyde as often as they are protonated if the kinetic isotopic effect ($k_{OH}{}^H/k_{OH}{}^D$) is small enough (~4). This is probably one reason why we observed falling first-order rate constants in our sodium hydroxide catalyzed dealdolizations.

The K value obtained for the acidity of isobutyraldehyde hydrate corresponds to an ionization constant of 1.7×10^{-14}. This is slightly smaller than the ionization constants of acetaldehyde hydrate (2.7×10^{-14}) and formaldehyde hydrate (5.4×10^{-14}).[19]

Experimental

N.m.r. Spectra.—The n.m.r. spectra[23] of isobutyraldehyde and its aqueous solutions were found to resemble those reported by Lombardi and Sogo[18] and by Ranft.[24] In aqueous solution the n.m.r. spectrum of the free aldehyde consists of doublets due to the aldehyde and methyl hydrogens (J values of 1.5 and 6.6 c.p.s., respectively) at τ 0.43 and 8.91 and a multiplet due to the α-hydrogen at 7.44. This spectrum is accompanied by that of the aldehyde hydrate, consisting of the methyl doublet (J = 6.2 c.p.s.) at τ 9.07, and the α-hydrogen multiplet at about 8.3. In heavy water another peak appears at about τ 5.2; this is believed to be a doublet due to the dihydroxymethyl group, with the upfield portion obscured by the protium oxide impurity in the heavy water. At 25° the doublet at τ 8.91 was found to be 1.66 ± 0.1 times as large as that at 9.07 and at 35° it was 2.36 ± 0.1 times as large.

Isobutyraldol, prepared from 2,6-diisopropyl-5,5-dimethyl-4-m-dioxanol[10] by a method based on that of Hagemeyer and coworkers,[12] had its best resolved n.m.r. spectrum in carbon tetrachloride solution, where the two γ-methyl groups (part of an isopropyl group attached to an asymmetric carbon atom) gave doublets (J values of 6.7 and 7.0 c.p.s.) at τ 9.10 and 9.07; the α-methyl groups, a singlet at 8.93; the γ-hydrogen, a broad multiplet centered around 8.20; the β-hydrogen, a doublet (J = 4 c.p.s.) at 6.52; and the aldehydic hydrogen, a singlet at 0.50. In aqueous solution the doublets due to the γ-methyl groups were essentially superposed; no absorption other than that attributable to the aldol and the solvent was observed.

Equilibrium in the Aldolization of Isobutyraldehyde.—In a typical experiment 50 ml. of 0.422 M aqueous isobutyraldehyde that was 0.0037 M in isobutyric acid was added to 3 ml. of 0.3345 M sodium hydroxide solution. Within a few minutes an organic layer separated and, after 1 hr. at 35°, a 15-ml. sample was removed and 0.18 ml. of 1.45 M acetic acid was added to neutralize it. This sample was centrifuged until it consisted of two clear layers. To 5 ml. of the aqueous layer was added an equal volume of water. The resultant solution had an absorbance of 1.052 at 2840 Å. (measured at 35°).[25] A second 5-ml. portion of the aqueous layer was extracted twice with 0.75-ml. portions of chloroform for the n.m.r. measurements. The n.m.r. peak (measured at 35°) at τ 8.85 (the downfield part of the methyl doublet of isobutyraldehyde) had an area 9.14 times that of the 8.93 peak (due to the α-methyl groups of isobutyraldol). Separate experiments showed that the extraction procedure removes the isobutyraldehyde and isobutyraldol from the water almost quantitatively. It was also shown that at 35° the extinction coefficient of isobutyraldol at 2840 Å. is 25.2 M^{-1} cm.$^{-1}$ and that of free isobutyraldehyde is 22.1 M^{-1} cm.$^{-1}$ (calculated from an apparent extinction coefficient of 15.5 M^{-1} cm.$^{-1}$ and the fact that 30% of the aldehyde is present as the hydrate under these conditions). Therefore, if y represents the concentration of isobutyraldol and x the total concentration of isobutyraldehyde (free aldehyde plus hydrate), the following relation (eq. 5) for the absorbance holds. The appropriate

$$(0.70)(22.1)x + 25.2y = 1.052 \quad (5)$$

relationship for the relative areas of the n.m.r. peaks is given by eq. 6 since 54% of the area of the methyl doublet of isobutyraldehyde is in the downfield part and none of the isobutyraldehyde is hydrated in chloroform. Simultaneous solution of eq. 5 and 6 followed by calculations based on the equilibrium constant for hydration show that the undiluted aqueous phase at 35° was 0.0870 M in free isobutyraldehyde, 0.0372 M in isobutyraldehyde hydrate, and 0.00734 M in isobutyraldol. These figures give a K_a of 0.97 M^{-1}. A number of runs carried out for different times and using different concentrations of dilute alkali, some starting with isobutyraldehyde and some with isobutyraldol, gave a K_a value of 1.05 ± 0.1 M^{-1}.

$$0.54x = 9.14y \quad (6)$$

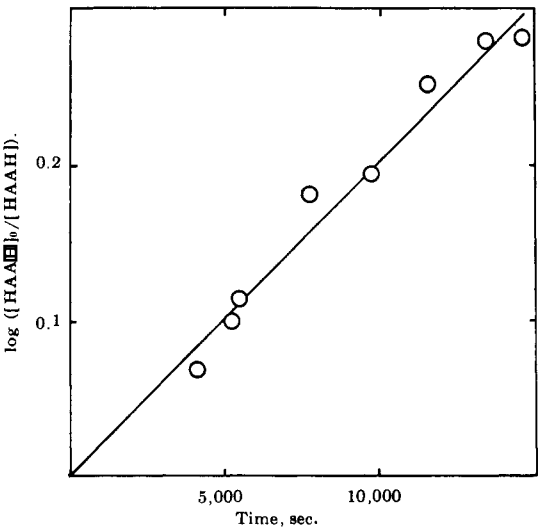

Figure 1.—Kinetic plot for the dealdolization of isobutyraldol in the presence of a 1,4-diazabicyclo[2.2.2]octane buffer.

pK Values of 1,4-Diazabicyclo[2.2.2]octane and Trimethylamine at 35°.—Weighed samples of diazabicyclooctane were dissolved in 25 ml. of water and titrated at 35 ± 0.5° using a Beckman research pH meter, Model 101900. The pK_a of the conjugate acid of the amine was taken as equal to the pH of the half-neutralized solution (the hydrogen ion concentrations were negligible compared to the concentrations of amine and its conjugate acid). Values of 8.577, 8.603, and 8.622 were observed at ionic strengths of 0.0063, 0.0090, and 0.0120 M, respectively. A plot of pK_a *vs.* ionic strength, extrapolated to zero ionic strength, gave a pK_a value of 8.530. Combination with the ion-product constant of water at 35°[26] gives a pK_b value of 5.150 for the amine at zero ionic strength.

To determine the pK values of trimethylamine at 35° the pK_a values for trimethylammonium ions at 10, 20, 30, 40, and 50°[21] were plotted against temperature at each of the ionic strengths 0.05, 0.10, 0.15, and 0.20 M (ionic strengths due to potassium chloride). The ionization constant of trimethylamine at a given ionic strength may then be determined from the pK_a value from this plot and the ion-product constant of water at the same ionic strength.[26] The pK_b for trimethylamine at 35° and zero ionic strength is 4.096.

Kinetics of the Dealdolization of Isobutyraldol.—Preliminary studies showed that in the presence of fairly strong sodium hydroxide the dealdolization of isobutyraldol is too fast to follow

(23) All n.m.r. spectra were determined using a Varian A-60 instrument. Chemical shifts are τ-values referred to tetramethylsilane for those spectra run in organic solvents and to sodium 3-(trimethylsilyl)-1-propanesulfonate for those spectra run in water.

(24) J. Ranft, *Ann. Physik*, **10**, 1 (1962).

(25) All ultraviolet spectral measurements were made using a 1.00-cm. cell in a Carey spectrophotometer, Model 14.

(26) H. S. Harned and B. B. Owen, "The Physical Chemistry of Electrolytic Solutions," 3rd Ed., Reinhold Publishing Corp., New York, N. Y., 1958, pp. 638, 752.

conveniently, and in the presence of dilute sodium hydroxide the observed first-order rate constants fall as the reaction progresses. Acidimetric titration showed that the base concentration decreases during the reaction, presumably due to the autoxidation of small amounts of aldehyde to acid. The reaction was therefore studied in buffer solutions. In a typical run the reaction was started by adding 0.25 ml. of 0.0716 *M* 1,4-diazabicyclo-[2.2.2]octane–0.0445 *M* 1,4-diazabicyclo[2.2.2]octane perchlorate and 0.50 ml. of a 0.0794 *M* isobutyraldol solution, that was 0.0016 *M* in acid, by syringe to an n.m.r. tube under nitrogen. The tube was kept at 35 ± 0.1° and at recorded times the areas of the τ 8.85 peak of isobutyraldehyde and the 8.93 peak of isobutyraldol were measured. Using a relationship of the form of eq. 6 and a second relationship based on the concentrations of isobutyraldol known to be present in the original reaction solution the concentrations of aldol and isobutyraldehyde may be calculated. The results of this calculation are shown in Table II

TABLE II

BASE-CATALYZED DEALDOLIZATION OF ISOBUTYRALDOL IN WATER AT 35°[a]

Time, sec.	$A_{8.85}$[b]	$A_{8.93}$[c]	[Aldol]
0			0.0529
4,110	1.752	13.44	0.0451
5,200	3.40	17.64	0.0422
5,460	3.29	13.97	0.0407
7,740	4.25	10.92	0.0349
9,720	4.99	11.61	0.0338
11,520	6.14	10.34	0.0296
13,440	6.51	9.50	0.0278
14,610	7.95	11.50	0.0277

[a] $[C_6H_{12}N_2]$ = 0.0228 *M*, $[C_6H_{12}N_2H^+]$ = 0.0159 *M*. [b] Area of the τ 8.85 peak. [c] Area of the τ 8.93 peak.

and a plot of log ($[HAAH]_0/[HAAH]$) *vs.* time is shown in Figure 1.

The hydroxide ion concentrations in the various runs were calculated from the buffer ratio and the ionization constant of the buffer base at the appropriate ionic strength (calculated by the method described explicitly for trimethylamine in the preceding section).

The Acidity of Isobutyraldehyde Hydrate.—Two solutions of isobutyraldehyde were prepared simultaneously, with the total concentration of aldehyde (hydrate plus free aldehyde) being 0.107 *M* in each. One solution was 0.0964 *M* in sodium hydroxide and the other contained no added base. The absorbances (measured immediately) of the two solutions at 2850 Å. were 1.385 and 1.465, respectively, at 25°. From these observations the following *K* value may be calculated.

$$K = \frac{[Me_2CHCHOHO^-][H_2O]}{[Me_2CHCH(OH)_2][OH^-]} = 93$$

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Further Studies of Mechanisms of Chlorinolysis of Sulfur–Carbon Bonds. The Mechanism of Abnormal Chlorinolysis and Desulfonylation of Sulfonyl Chlorides. III[1]

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The incidence of the abnormal mechanism of chlorinolysis of sulfur–carbon bonds is apparently restricted to sulfides possessing the structural feature RS–(C=C)$_n$–C=N–. The factors involved in the competition of normal and abnormal reactions are examined in this context. The possibility that the abnormal course proceeds by a nucleophilic addition–elimination mechanism of the type encountered usually in aromatic and heteroaromatic systems has been evaluated. The effect of medium composition on the relative occurrence of the competing reaction mechanisms in particular cases has been explored. The inference that might be drawn from these results is that N-chlorination may create special facilitation for a nucleophilic displacement mechanism in the abnormal reaction. This has been examined by tracer experiments and discredited. Further consideration of the product distribution in experiments with 4-alkylthio-7-chloroquinoline substrates and the factors determining the stabilities of related sulfonyl chlorides suggest that an SNi mechanism can account for all the known characteristics of both the abnormal chlorinolysis reaction and the catalyzed decomposition of the sulfonyl chlorides.

An earlier article from these laboratories[2] has discussed the chlorinolysis of 4-benzylthio-7-chloroquinoline (I) and some of its derivatives. The principal cleavages products were reconciled with the intervention of a carbonium ion. However, 14% of the products arose from an alternative course which most certainly was not of carbonium ion character. It has been suggested[3,4] that benzylsulfonyl chloride (III) was formed by an SN2 displacement of the sulfenyl chloride from the 4-position by chloride ion. These available, competitive reaction paths are summarized in Scheme I.

The exclusive preference of chlorinolysis to occur *via* the normal SN1 path in most substrates is well documented.[2,5] However, in certain cases the displacement reaction path (SN2?), which we shall hereafter call *abnormal*, does manifest itself as a competing path.

(1) (a) Part of the data discussed here is taken from the Ph.D. dissertation of R. W. Body submitted in partial fulfillment of the requirements at the University of Delaware, June 1963. (b) Part of this paper has been presented before the American Association for the Advancement of Sciences Symposium on Recent Advances in Organic Chemistry, Philadelphia, Pa., Dec. 1962.

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