TABLE IV

EFFECT OF TEMPERATURE

Ionic strength = 0.034; initial H_2O_2 concn. = initial Na_2-SO_3 concn. = 1.12 × 10^{-2} M

T, °C.	pH	k_2, 1. mole^{-1} sec.$^{-1}$
0.4	9.95	0.0196
15.0	10.18	.0829
25.0	10.08	.196
35.0	9.95	.464

shown in the third column of Table III. Some effect is to be expected in the absence of EDTA salts, owing to loss of sulfite through reaction with the oxygen of the air and through loss of H_2O_2 through decomposition. However, the importance of these side reactions relative to the peroxide–sulfite reaction was minimized by choosing conditions leading to rapid peroxide–sulfite reaction. The experiment was run to see if there was any important, unanticipated effect of EDTA salts on the rate of the reaction. On comparing the second and third columns of Table III, it is obvious that there was no such effect.

Effect of Metal Ions.—The chemicals used in the experiments probably contained traces of ions that usually show catalytic activity. To determine whether or not metal-ion catalysis might have played an important part in the experiments, a run was made in an unbuffered solution to which sufficient Fe^{+++} and Cu^{++} were added to double

the maximum expected level of these ions in the kinetic experiments. EDTA salts were also present and Na_2SO_4 was used to give an ionic strength of 1.0. The pH varied from 8 to 6.6 during the course of the reaction. For comparison, an identical run was made without the added metal ions. The polarographic method was used. The added metal ions had no detectable effect on the rate of disappearance of H_2O_2 from the reaction mixture. It is unlikely that metal-ion catalysis was of any importance in the kinetic runs.

Effect of Light.—No attempt was made to protect the solutions during reaction from the illumination in the room, since it was believed that the reactants absorbed a negligible amount of radiation from this source. However, as a precaution, one reaction in pH 10 carbonate buffer was carried out in the beam of a Kodak Signet 500 slide projector placed 16 inches from the reaction vessel. There was no detectable effect of this high-intensity illumination on the reaction rate.

Errors.—For pH values above about 8 and for an initial H_2O_2 concentration of 0.01 M, the probable errors in k_2 for several representative initial sulfite concentrations are: B_0 = 0.04 M, 1.5%; B_0 = 0.01 M, 2.0%; B_0 = 0.0055 M, 3.2%. These errors were calculated using assumed values for the probable errors in the various quantities used in calculating k_2.

ROCHESTER 4, N. Y.

Deuterium Isotope Effects in the Bromine Oxidation of Ethanol and of Acetaldehyde[1]

BY LOUIS KAPLAN

The rate of oxidation of ethanol-1,1-d_2 by bromine in aqueous solution has been measured over a range of alcohol concentrations from one to 41% by weight, and over a pH range from two to ten. In unbuffered acid solution the ratio of the rates of oxidation of ordinary to that of dideutero alcohol is close to four, proving that a methylene C–H bond is broken in the rate-determining step of the reaction. From the isotope effect and the reaction kinetics it is concluded that ethyl hypobromite plays no role in the oxidation of ethanol, neither as the precursor of acetaldehyde nor as the active oxidant. Under the same conditions acetaldehyde is oxidized at a rate four times that of acetaldehyde-1-d, in accord with the suggestion that the aldehyde is oxidized, *via* its hydrate, by a mechanism very similar to that for ethanol. At pH 5.5 and 10.3 the isotope effects for the alcohol oxidation are 2.4 and 3.0, respectively. The significance of these lower values cannot be assessed in the absence of detailed knowledge of the reaction products and kinetics under these conditions.

Introduction

In an earlier investigation[2] of the oxidation of ethanol-1-t by bromine it was found that the rate of removal of a tritium atom from the methylene carbon was about 0.15 that for a protium atom. From this isotope effect and the previously investigated[3] reaction kinetics it was concluded that ethyl hypobromite is not a precursor of acetaldehyde in the oxidation. The possibility remained that ethyl hypobromite or an ethanol–bromine complex might be formed in the rate-determining step of the reaction and oxidize a different molecule of ethanol in a subsequent fast step. It was pointed out that such

a mechanism could account for the isotope effect observed in the competitive system studied with tritium, but should result in little or no difference between the separately measured reaction rate constant for ethanol from that for ethanol-1,1-d_2. The present paper presents the results of a comparison of the reaction rates of these two compounds in dilute unbuffered solutions comparable to those used in the tritium studies, as well as at higher alcohol concentrations and higher pH. An estimate has also been made of the deuterium isotope effect in the concurrent oxidation of acetaldehyde.

Experimental

Reagents.—Ethanol was the U. S. Industrial Chemicals Co. absolute grade. It was found to contain less than five parts per million of aldehyde.

Ethanol-1,1-d_2 was prepared by a procedure similar to that used for the tritium-labeled compound.[2] To a sus-

(1) Work performed under the auspices of the U. S. Atomic Energy Commission. Presented at the 132nd Meeting, American Chemical Society, New York, N. Y., September, 1957.

(2) L. Kaplan, THIS JOURNAL, 76, 4645 (1954).

(3) L. Farkas, B. Perlmutter and O. Schächter, *ibid.*, 71, 2829 (1949).

pension of 4.54 g. (0.108 mole) of lithium aluminum deuteride (Metal Hydrides, Inc.) in 150 ml. of absolute ethyl ether, which had been stirred for two hours in a nitrogen atmosphere, was added a solution of 30.8 g. (0.205 mole) of benzyl acetate in 30 ml. of ether over a period of 40 minutes. The mixture was refluxed for two hours, 3 g. of acetophenone was added to destroy excess hydride, and the mixture was again refluxed for one-half hour. The ether was removed *in vacuo*, 10 ml. of water was added and the mixture was well stirred. The volatile product was collected in a Dry Ice trap by pumping with an oil pump while heating the reaction mixture to a temperature of 50°. The distillate was fractionated through a spiral-wound column and the alcohol–water azeotrope was refractionated through a Nester micro spinning-band column.[4] The yield of azeotrope was 6.27 g. (58% based on LiAlD₄), containing 96.0% of alcohol. The deuterium content of the alcohol corresponded to 1.94 atoms per molecule. The alcohol was subsequently found to contain 0.11 mole % of aldehyde, necessitating a correction of some of the measured reaction rates.

Other chemicals used were reagent grade commercial products.

Analytical Procedures.—Bromine was determined iodometrically: an aliquot was added to a mixture of 15 ml. of H₂O, 1 ml. of 5% KI and 1 ml. of 1 N HCl and the liberated iodine was titrated amperometrically[5] with 0.002 N Na₂S₂O₃.

Acetaldehyde was determined colorimetrically as the dinitrophenylhydrazone, following the procedure of Böhme and Winkler.[6] Optical densities were measured at 435 mμ on a Beckman model DU spectrophotometer. The procedure was standardized against a freshly distilled sample of acetaldehyde. Measurements were reproducible to about one per cent. of the acetaldehyde content if the size of the aliquot was adjusted to give a density between 0.5 and 1.0 in a 5 mm. cell.

Measurement of Reaction Rates.—All rates were measured at 24.8°. Alcohol solutions were prepared by diluting weighed aliquots with water or buffer solutions in 25- or 50-ml. volumetric flasks, the stoppers of which were lubricated with silicone to avoid evaporation of bromine. Liquid bromine was added from a micropipet to the thermostated solution which was then agitated rapidly to dissolve the bromine and returned to the thermostat. At appropriate intervals aliquots were run into acidified KI solution and titrated as described above. The micropipets were calibrated by discharging their content of bromine into water and analyzing an aliquot.
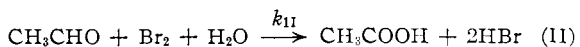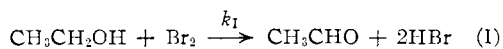
Two buffer solutions were used: 0.2 M potassium acid phthalate with NaOH for pH 5.5 and 0.2 M boric acid with NaOH for pH 10.3. The initial pH of the buffer and that of the reaction mixture after reduction of the bromine were measured with a pH meter.

Rate runs with 41% alcohol were carried out on a 1-ml. scale in order to conserve deuteroethanol. These runs were made in 0.1 M KBr solution in order to achieve a convenient reaction rate and to facilitate the calculation of reaction rate constants. The entire sample was analyzed for bromine after approximately half the bromine had been consumed.

The final concentration of acetaldehyde was estimated, for some of the rate runs, by analyzing an aliquot for acetaldehyde after the bromine color had disappeared.

Results

The oxidation of ethanol by bromine takes place in two stages[3,7]

$$CH_3CH_2OH + Br_2 \xrightarrow{k_I} CH_3CHO + 2HBr \quad (I)$$

$$CH_3CHO + Br_2 + H_2O \xrightarrow{k_{II}} CH_3COOH + 2HBr \quad (II)$$

In dilute, unbuffered, aqueous solution (1 to 4% alcohol), the ratio k_{II}/k_I is 202.[7b] The rate of reaction I is proportional to the concentration of alcohol and of free bromine and is independent of acid concentration. The accumulation of Br⁻ during

(4) R. G. Nester, *Anal. Chem.*, **28**, 278 (1956).

(5) D. P. Evans, *Analyst*, **72**, 99 (1947).

(6) H. Böhme and O. Winkler, *Z. anal. Chem.*, **142**, 1 (1954).

(7) (a) S. Bugarszky, *Z. physik. Chem.*, **38**, 561 (1901); (b) **71**, 705 (1910).

the course of oxidation slows down the reaction by tying up bromine as tribromide according to the equilibrium

$$Br_2 + Br^- \rightleftharpoons Br_3^- \quad (III)$$
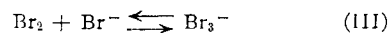
The initial rate of bromine consumption is that of reaction I, but as the reaction proceeds, the rate increases because of the build-up of the more rapidly oxidized acetaldehyde; under appropriate conditions the rate approaches twice that of reaction I.

The kinetic equations for the oxidation of ethanol in dilute solutions have been solved by Bugarszky.[7b] The computations are, however, extremely tedious. They are, moreover, considerably influenced by the value assigned to the tribromide equilibrium constant. For the present purpose, namely, the comparison of the rates of oxidation of CH_3CH_2OH and CH_3CD_2OH, the explicit calculation of k_I is unnecessary. It is found experimentally that if the alcohol is in considerable excess over bromine, a plot of the logarithm of the *total* bromine concentration *vs.* time gives a nearly straight line. If, in parallel experiments with the two isotopic alcohols, the initial concentrations of bromine are identical, the extent of Br_3^- formation will be the same for a given fraction of bromine reduced. If, in addition, the ratio k_{II}/k_I is the same for each of the isotopic alcohols, the effect of acetaldehyde oxidation will be the same for the two experiments. Under these conditions, the ratio of the rate constant, $k_I{}^H$, for oxidation of the normal alcohol, to that, $k_I{}^D$, for the deuteroalcohol, is equal to the ratio of the slopes of the experimental rate curves. Since, as will be shown below, the value of $k_{II}{}^H/k_I{}^H$ is quite close to that of $k_{II}{}^D/k_I{}^D$, this method of treating the data gives a valid measure of $k_I{}^H/k_I{}^D$.

The results of oxidations carried out on the two isotopic alcohols at three concentrations are plotted in Fig. 1. The ratios of the slopes (hydrogen/deuterium) for parallel runs are 4.27, 4.51 and 4.75 for runs II, III and IV, respectively. To obtain the corresponding isotopic rate ratios, k_H/k_D, it is necessary to correct for the small differences in initial alcohol concentration between parallel runs by multiplying the ratio of the slopes by the ratio of the molarity of the deuteroalcohol to that of the normal alcohol. The resulting values of k_H/k_D are 4.33, 4.37 and 4.15; mean value 4.28. The reasonable agreement among these values indicates that no serious error is introduced by the fact that in run D-III the initial Br_2 concentration is somewhat higher than in run H-III, while in run D-IV it is somewhat lower than in run H-IV.

Some of the conclusions arrived at in the previous paper[2] are based on the kinetic results of Farkas, *et al.*,[3] in 41% (by weight) alcohol, although the tritium isotope effect measurements[2] were perforce carried out in much more dilute solutions. An experiment was therefore performed to establish whether the deuterium isotope effect in 41% alcohol was significantly different from that already determined for 1 to 4% solutions. The oxidations were carried out at very low bromine concentrations in the presence of a large excess (0.1 M) of KBr. Under these conditions the effect of the additional bromide, as well as of the acetaldehyde, formed in
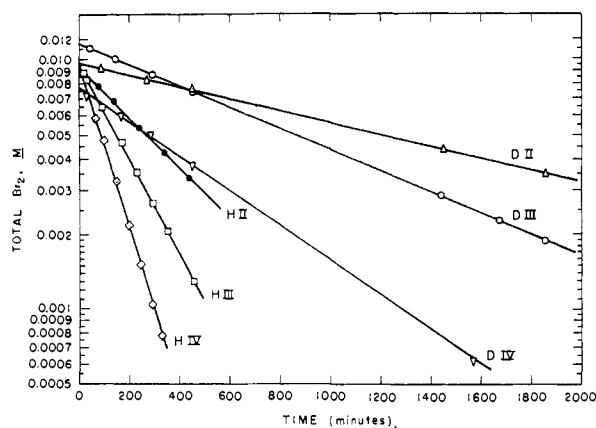
Fig. 1.—Oxidation of isotopic ethanols by bromine at 24.8°: D-II, D-III, D-IV: CH_3CD_2OH, 0.206 M, 0.408 M, 0.735 M; H-II, H-III, H-IV: CH_3CH_2OH, 0.203 M, 0.421 M, 0.843 M.

the oxidation is negligible and the rate of the reaction is given by the expression[3]

$$-\frac{dx}{dt} = kK\frac{x}{(Br^-)} \qquad (IV)$$

where x represents the *total* titratable bromine, k is the rate constant for reaction I, and K is the dissociation constant[3] (0.0106 mole l.$^{-1}$) of tribromide in 41% alcohol. In order to conserve deuteroethanol, the reactions were carried out on a 1-ml. scale, and the residual bromine was analyzed after about one half-time had elapsed. The value of k was calculated from the integrated form of equation IV

$$k = \frac{2.3(Br^-)\log x_0/x}{K(t - t_0)} \qquad (V)$$

At the high alcohol and low bromine concentration used in these experiments, the acetaldehyde content of the deuteroethanol has an appreciable effect on the observed reaction rate. The true value of k_I may be calculated from the measured rate constant k by means of the expression

$$k_1 = k - k_{II}(\text{aldehyde}) \qquad (VI)$$

using $k_{II}{}^D = 0.74$ obtained by combining Farkas' value[3] of 2.9 for $k_{II}{}^H$ with the ratio $k_{II}{}^H/k_{II}{}^D = 3.9$ determined as described below. For the aldehyde concentration in equation VI, the initial concentration, 0.0089 M was averaged with the value 0.0079 M calculated to be present when the reaction was stopped. The results are shown in Table I. The value of 0.097 min.$^{-1}$ found for $k_I{}^H$ in 8.37 M ethanol is in reasonable agreement with Farkas' value,[3]

TABLE I
OXIDATION OF ISOTOPIC ETHANOLS AT 24.8°

	CH_3CH_2OH	CH_3CD_2OH
Alcohol, M	8.37	8.10
(KBr), M	0.100	0.100
$(Br_2)_0$, M	.00730	.00730
$(Br_2)_t$, M	.00394	.00358
$t - t_0$, min.	60.0	238.0
(Aldehyde)$_0$, M	0	0.0089
k, min.$^{-1}$	0.097	.0282
k_I, min.$^{-1}$	0.097	.0220
$\frac{k_I{}^H}{k_I{}^D} \times \frac{M_D}{M_H}$		4.3

0.0942 min.$^{-1}$, in 8.28 M solution. The observed ratio of $k_I{}^H$ to $k_I{}^D$ was multiplied by the ratio of alcohol concentrations to give a value of $k_I{}^H/k_I{}^D = 4.3$.

The reaction rates for the oxidations at pH 5.5 and 10.2 in 0.17 M alcohol are shown in Fig. 2. The ratios, k_H/k_D, estimated from the slopes, are 2.4 at pH 5.5 and 3.0 at pH 10.3.
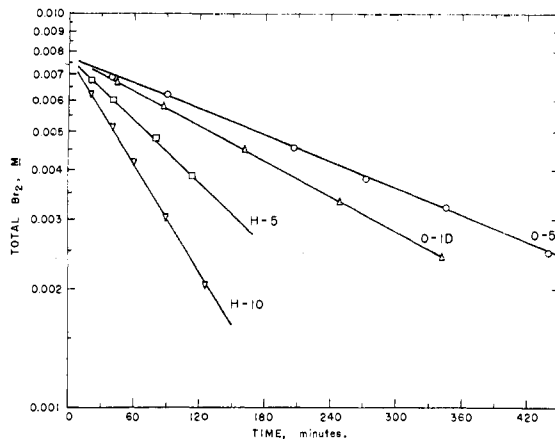

Fig. 2.—Oxidation of isotopic ethanols by bromine in buffered solutions at 24.8°: D-5, D-10: 0.165 M CH_3CD_2-OH, pH 5.5, 10.3; H-5, H-10: 0.168 M CH_3CH_2OH, pH 5.5, 10.3.

The ratio of the rate constant for reaction II to that for reaction I can be determined,[3] under appropriate conditions, by measuring the steady-state concentration of acetaldehyde during the oxidation of ethanol. If this ratio is measured for each of the isotopic alcohols, then, since the isotope effect, $k_I{}^H/k_I{}^D$, for reaction I is known, it is possible to calculate the isotope effect, $k_{II}{}^H/k_{II}{}^D$, for reaction II from the relationship

$$k_{II}{}^H/k_{II}{}^D = \frac{k_{II}{}^H/k_I{}^H}{k_{II}{}^D/k_I{}^D} \times k_I{}^H/k_I{}^D \qquad (VII)$$

As demonstrated by Farkas, the final concentration of acetaldehyde after complete consumption of bromine is practically equal to the steady-state concentration provided that the initial ratio of bromine to alcohol concentration is greater than eight times the ratio k_I/k_{II}. The value of k_{II}/k_I is then equal to the ratio of the final concentration of alcohol to that of aldehyde. The results of a number of such

TABLE II
STEADY-STATE CONCENTRATION OF ACETALDEHYDE DURING OXIDATION OF ISOTOPIC ETHANOLS

Alcohol	(Et-OH)$_0$, M	(Et-OH)$_{final}$, M	$(Br_2)_0$, M	(CH$_3$-CHO), $M \times 10^3$	k_{II}/k_I
CH_3CH_2OH	0.165	0.160	0.0094	0.891	180
	.202	.196	.0113	1.10	178
	.203	.198	.0094	1.09	182
	.418	.404	.0283	2.19	184
	.833	.805	.0566	4.34	185
				Mean	182
CH_3CD_2OH	0.205	0.200	0.0094	0.99	202
	.408	.402	.0113	2.04	197
	.407	.393	.0283	1.95	201
	.812	.784	.0566	3.89	202
				Mean	201

experiments in unbuffered, dilute alcohol solutions are shown in Table II. From the values $k_{II}{}^H/k_I{}^H = 182$, $k_{II}{}^D/k_I{}^D = 201$ and $k_I{}^H/k_I{}^D = 4.3$, the value of $k_{II}{}^H/k_{II}{}^D$ is calculated to be 3.9.

Discussion

The deuterium isotope effect for the oxidation of ethanol by bromine in dilute, unbuffered solution may be compared directly with the previously determined tritium isotope effect.[2] Making the very rough approximation that the isotope effects arise from differences in the zero point energies associated with the stretching vibrations involving the isotopic hydrogen atoms in the normal and transition states of the reaction, and that the vibration frequencies for protium, deuterium and tritium are in the ratio $1:\sqrt{2}:\sqrt{3}$, it can be calculated that a value of k_H/k_D of 4.3 at 24.8° corresponds to a value of k_H/k_T of 7.5 at 37.5°. The latter value agrees, within the rather large uncertainties in both the experimental measurements and the over-simplified calculation, with the experimental value[2] of 6.7. This agreement between the competitively measured k_H/k_T and the kinetically measured k_H/k_D is in contrast to the case of the chromic acid oxidation of 2-propanol[8] in which reaction the processes studied by the competitive and non-competitive methods were not identical.

The observation that k_H/k_D for the oxidation of ethanol in 41% solution is identical with that in the 1 to 4% solutions removes any doubt of the validity of combining the isotope effect measurements in the dilute solutions with Farkas' kinetic results[3] in the concentrated solutions. The fact that the isotopic rate ratio, k_H/k_D, is as large as 4.3 requires the conclusion that a methylene C–H bond is broken in the rate-determining step of the reaction. This rate-determining step cannot, therefore, be the alcoholysis of bromine to ethyl hypobromite.

$$CH_3CH_2OH + Br_2 \longrightarrow CH_3CH_2OBr + HBr \quad (VIII)$$
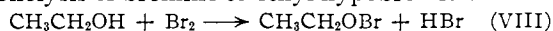
Neither can equation VIII represent a rapid equilibrium preceding the rate-determining step, since such an equilibrium would require an inverse de-

(8) L. Kaplan, This Journal, 77, 5469 (1955).

pendence of the reaction rate on HBr concentration, contrary to the findings of Farkas, et al.[3] It must be concluded, therefore, that ethyl hypobromite plays no role in the oxidation of ethanol, either as the precursor of acetaldehyde or as the oxidant of a second ethanol molecule. The lack of intervention by ethyl hypobromite in this oxidation is of particular interest since a number of oxidations of alcohols by inorganic oxidants in polar solutions have been shown or postulated to involve esters as intermediates.[9] The bromine oxidation, in contrast, appears to involve direct attack by molecular bromine, or a species with which it is in rapid equilibrium, at the C–H linkage of the alcohol, possibly abstracting hydrogen as hydride ion.[2,3] The similarity of the isotope effect $(k_H/k_D = 3.9)$ in the acetaldehyde oxidation to that in the alcohol oxidation is in accord with the idea[3] that these reactions have similar mechanisms, the aldehyde reacting in the form of its hydrate or hemiacetal.

The occurrence of appreciable isotope effects in the oxidation of ethanol in solutions buffered at pH 5.5 and 10.2 indicate that under these conditions, also, the rate-determining steps involve rupture of a carbon–hydrogen bond. The reaction rates in solutions above pH 3, in contrast to those in unbuffered solutions at lower pH, are, however, markedly pH-dependent.[3,10] In the absence of a detailed analysis of the reaction kinetics, it is impossible to draw any conclusions with respect to possible intermediates in the oxidation under conditions of high pH.

Acknowledgment.—The author is indebted to Mrs. Judith Jero for technical assistance. He also expresses his appreciation to Dr. B. Perlmutter-Hayman for making available to him some unpublished results on the oxidation of ethanol solutions at high pH.

(9) L. S. Levitt and E. R. Malinowski, ibid., 77, 4517 (1955), suggest that such oxidations always go through the intermediate stage of ester formation.

(10) B. Perlmutter-Hayman, private communication.

LEMONT, ILLINOIS

Effect of Temperature, Pressure, Acidity and Solvent on an Aquo Ion Exchange Reaction

BY H. R. HUNT AND H. TAUBE

$\Delta V^\ddagger$, $\Delta S^\ddagger$ and $\Delta H^\ddagger$ for the exchange of water between $(NH_3)_5CoOH_2{}^{+++}$ and solvent have been measured as $+1.2 \pm 0.2$ ml., $+6.7 \pm 1$ e.u. and 26.6 ± 0.3 kcal. mole^{-1}. The values of $\Delta V^\ddagger$ and $\Delta S^\ddagger$ observed, and the added observation that $\Delta V^\ddagger$ is independent of pressure, appear to rule out an extreme or compact S_N2 mechanism. The observations require that the Co(III)–OH$_2$ bond in the activated complex be stretched to a critical value; however, they provide no basis for deciding whether a true intermediate is formed on decomposition of the activated complex or whether some use is made of the entering water molecule in the activation process. The rate of water exchange is almost independent of the concentration of water when CH$_3$OH is principal component of the solvent mixture. The exchange rate for $(NH_3)_5CoOH^{++}$ is much less (by a factor of at least 40) than it is for $(NH_3)_5CoOH_2{}^{+++}$.

Many substitution reactions of metal complex ions in water proceed by intermediate aquation. The exchange of water with an aquo ion, therefore, has special significance to the problem of the mech-

anism of the substitution reactions.[1] Because rate law methods fail to define the composition of the activated complex with respect to solvent, the

(1) J. P. Hunt and H. Taube, J. Chem. Phys., 19, 602 (1951).