2. With isatin alpha-chlorides, either alpha or beta-condensations were obtained, depending upon the experimental conditions.

3. As dyes, these products proved to be of little value.

NEW YORK, N. Y. RECEIVED JUNE 15, 1935

[CONTRIBUTION FROM THE CHEMISTRY LABORATORY OF THE UNIVERSITY OF MICHIGAN]

The Exchange Reaction between Acetone and Deuterium Oxide. Kinetics and Equilibrium

BY J. O. HALFORD, LEIGH C. ANDERSON, JOHN R. BATES AND R. D. SWISHER

The direct introduction of deuterium atoms into acetone by means of the reaction between acetone and dilute deuterium oxide, using an alkaline catalyst, has been reported independently by several investigators.[1] In the original communication from this Laboratory, the exchange in the presence of potassium carbonate was described as rapid, and the equilibrium constant of the reaction

$$CH_3COCH_3 + DOH = CH_3COCH_2D + HOH \quad (1)$$

was estimated to be approximately two.

This paper presents the results of a study of the kinetics at 35 and 40°, and of the equilibrium at temperatures from 35 to 80°. The reaction is actually very slow at room temperature and only moderately rapid at the boiling point of the acetone–water solution.

Experimental

The deuterium content of the acetone was calculated from the decrease in density of the water, measured by the "plummet" method to within 20 parts per million. The quantitative separation of acetone and water was the principal experimental problem. For this purpose a 90-cm. vacuum-jacketed fractionating column was found to be satisfactory, as demonstrated by the ready recovery of water of normal density from an acetone–water solution. In the determinations at the boiling point of the solution, in which potassium carbonate was present during the entire separation, no difficulty was encountered in obtaining water free from organic impurities. When the reaction was studied at lower temperatures, it was necessary to neutralize the catalyst with phosphoric acid before distilling. The water fraction obtained from the acid solution contained traces of an organic impurity, presumably diacetone alcohol, which interfered with the density determination. The impurity was removed as acetone by redistilling the water from potassium carbonate. The water fraction constituted 60–70% of the amount originally introduced into the reaction, the remainder being divided about equally between rejected initial and final fractions, thus ensuring that the density measurement would be representative.

The acetone was obtained by fractional distillation of a good grade of c. p. acetone, using the 90-cm. column. A middle fraction boiling well within 0.1° and constituting 50–70% of the total was retained. It was decided that purity of the acetone could best be tested by mixing it with a deuterium oxide solution, without the catalyst, and then examining the water for decrease of deuterium content after the separation by distillation. The deuterium content of a sample of water, on treatment with twice its volume of acetone, was decreased from 4.17 to 4.09%. If the effect is due entirely to dilution by water in the acetone, its concentration is calculated to be 1.0% by volume or 3.15 mole per cent. Reaction without the catalyst, as indicated by the velocity data, would be expected to be negligible. In calculating the concentrations at equilibrium a correction for the water content of the acetone has been made.

Recovery of Deuterium Oxide by Combustion.—A sample of 12.05% deuterium oxide was decreased to 5.19% by the exchange reaction with twice its volume of acetone. The acetone fraction was dried with powdered fused potassium hydroxide, carefully redistilled, and the middle fraction burned in a small lamp with an asbestos wick. The gases from the combustion were passed through a tube containing hot cupric oxide, to produce water containing 4.60% deuterium oxide. This is to be compared with 4.64% calculated from the decrease in deuterium content of the water.

Kinetics

The kinetics of reaction 1, which is found to be pseudo-unimolecular, are consistent with any one of several postulated mechanisms, including the mechanism which assumes an intermediate enolization of the acetone. It is therefore not possible to say definitely that enolization is responsible for the reaction, or to obtain from the temperature coefficient the energy necessary to convert the keto to the enol form. However, this mechanism is by no means excluded.

Each experiment was performed with a solution made up of 20 cc. of 2.55% deuterium oxide, 40 cc. of acetone, and a weighed amount of anhydrous potassium carbonate. After the solution

---

[1] Halford, Anderson and Bates, THIS JOURNAL, **56**, 491 (1934); Bonhoeffer and Klar, *Naturwissenschaften*, **22**, 45 (1934); Schwarz and Steiner, *Z. physik. Chem.*, **B25**, 153 (1934); Klar, *ibid.*, **B26**, 335 (1934).

had been submerged in a thermostat for the required length of time, the catalyst was neutralized with phosphoric acid, and the analysis was carried out in the manner already described. Figure 1 shows the concentration of deuterium oxide in the water as a function of the time. Curve I shows the results obtained at 35° with 0.08 g. of potassium carbonate, curve II with the same amount of catalyst at 40°. Curve III was obtained at 35° with 0.008 g. of the catalyst. It can be seen that the curves are asymptotic to a DOH concentration of 0.0157, expressed as the mole fraction in the acetone–water solution. This leads to a value of the apparent equilibrium constant of reaction 1, $K(\Sigma) = 2.52$. This constant is defined by the relation

$$K(\Sigma) = (\Sigma)/[(1 - \Sigma)R] \qquad (2)$$

where $\Sigma = (\text{acetone-}d) + 2(\text{acetone-}d_2) + 3(\text{acetone-}d_3) + \ldots \qquad (3)$

and R is the ratio (DOH)/(HOH) at equilibrium. The concentrations of the acetones are expressed as fractions of the total acetone. The fact that there is no detectable change of this constant with the temperature indicates that the heat of reaction is approximately zero.

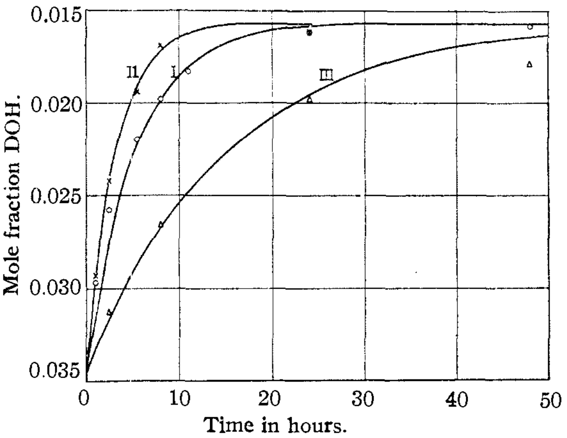

Fig. 1.

Since the mechanism cannot be determined it serves the purpose to consider the over-all reactions

$$CH_3COCH_3 + DOH \underset{k_2}{\overset{k_1}{\rightleftharpoons}} CH_3COCH_2D + HOH$$

An approximate solution for the rate constants k_1 and k_2 may now be made. The concentrations of acetone and ordinary water are considered as constants, and the concentration of acetone-d is set equal to the decrease in DOH concentration. It is then possible to derive equation 4

$$k_{ps.} = k_1(CH_3COCH_3) + k_2(HOH) = \frac{1}{t}\ln\frac{(DOH)_0 - (DOH)_e}{(DOH) - (DOH)_e} \qquad (4)$$

which expresses the pseudo-unimolecular rate constant $k_{ps.}$ in terms of the constants k_1 and k_2 and also in terms of the time t and the initial and final concentrations $(DOH)_0$ and $(DOH)_e$. The combination of equation 4 with the relation $k_1/k_2 = K(\Sigma)$ permits the calculation of k_1 and k_2, since $k_{ps.}$ and $K(\Sigma)$ are expressed in terms of measured quantities. Table I shows the results from the data represented by the three curves of Fig. 1, and from two additional single measurements at 35° with higher concentrations of the catalyst. The first row shows the temperature, the second the weight of the catalyst, the third the over-all rate constant $k_{ps.}$, and the fourth and fifth the rate constants, respectively, of the forward and reverse reactions, in reciprocal hours. The average difference between the experimental points and calculations based upon the velocity constants is about 3%.

Table I

Kinetics

$(HOH) = 0.678$, $(CH_3COCH_3) = 0.322$

	I	II	III		
Temp.	35	40	35	35	35
K_2CO_3, g.	0.08	0.08	0.008	0.024	0.24
$k_{ps.}$	.190	.311	.0668	.167	.233
k_1	.323	.527	.113	.282	.393
k_2	.128	.209	.0448	.112	.156

The rate can be seen to be extremely sensitive to temperature. A calculation using the two values of k_1 from curves I and II yields an activation energy of 17,900 cal. The rate is also dependent upon the potassium carbonate concentration up to a certain point, after which additional amounts have little effect. It is futile to speculate upon the significance of this with our present knowledge of the behavior of electrolytes in mixed solvents.

Equilibrium

The equilibrium data are summarized in Table II. In addition to the results obtained in connection with the reaction velocities, several determinations at higher temperatures are included. The slight temperature coefficient of the equilibrium constant made it possible to carry out relatively rapid measurements (arbitrarily marked 80°) in which the acetone and water were separated by direct distillation from the solution containing the catalyst. The first column shows the temperature, the second the time in hours, the third the mole fraction of deuterium oxide in the water initially introduced, and the fourth its mole fraction in the water separated at equilibrium. The fifth column shows the ratio of total moles of water to total moles of acetone, and the sixth the ratio R as defined above, calculated with the aid of the equilibrium constant of the reaction

$$H_2O + D_2O = 2DOH; K = 3.26 \qquad (5)$$

as given by Topley and Eyring.[2] In the seventh column is shown the acetone fraction Σ, defined by Equation 3, and in the eighth the apparent equilibrium constant $K(\Sigma)$.

TABLE II

EQUILIBRIUM

Temp., °C.	Time, hrs.	D_2O, initial	D_2O, final	$n(H_2O)/n(\text{acetone})$	DOH/HOH	Σ	$K(\Sigma)$
80	5	0.1201	0.0517	2.105	0.1080	0.2789	3.58
80	3	.0952	.0414	2.105	.0893	.2153	3.07
80	5	.0399	.0189	2.105	.0384	.0860	2.45
80[a]	12	. . .	.00919	2.539	.01852	.0400	2.25
60	96	.0241	.01136	2.105	.02293	.0524	2.41
40,35	∞	.0255	.01167	2.105	.02356	.0561	2.52

[a] Reversed reaction, using the acetone obtained from the previous experiment.

It may be seen from the last column that in order to obtain the true equilibrium constant of reaction 1, correction must be made for the higher deuterioacetones. This may be done expeditiously in the following manner.

When $K(\Sigma)$ is plotted as a function of R, the intercept at the axis $R = 0$ gives K_1, the equilibrium constant of reaction 1. The curve so obtained is tangent at $R = 0$ to a line of slope $2K_2$, which lies below the curve, as shown by the following equation

$$K(\Sigma) = K_1 + 2K_2R + 3K_3R^2 + \ldots \qquad (6)$$
$$dK(\Sigma)/dR = 2K_2 + 6K_3R + \ldots \qquad (7)$$

K_2 and K_3 are the equilibrium constants of the reactions

$$\text{acetone} + 2DOH = \text{acetone-}d_2 + 2HOH, \text{ and}$$
$$\text{acetone} + 3DOH = \text{acetone-}d_3 + 3HOH$$

Equations 6 and 7 are approximate only to the slight extent introduced by using the quantity $(1-\Sigma)$ to represent the concentration of ordinary acetone. It is of assistance to plot at the same time Σ/R as a function of R, obtaining the same slope and intercept at $R = 0$ because of the approach of $(1-\Sigma)$ toward unity as R approaches zero. This latter curve is more nearly horizontal and lies below the tangent line of slope $2K_2$.

The application of this analysis to the first four measurements recorded in Table II yields the value $K_1 = 2.1$, which should be more accurate than the corresponding values of $K(\Sigma)$, which are reproducible to within ±0.1 unit. Applying $K_1 = 2.1$ to the data leads to the value $K_2 = 4$. No value of K_1 different by any considerable amount from the one selected gives a reasonable interpretation of the data, but K_2 is subject to a relatively large error, especially since it is sensitive to any trend which might have been introduced by the experimental procedure.

For comparison with the experimental results, the entropy of each of the reactions may be calculated by means of the equation proposed by Halford.[3] If free rotation of the substituted methyl group is assumed, the entropy of Reaction 1 is calculated to be $-R \ln 1.92$, which indicates a small negative heat of reaction consistent with the slightly higher values of the equilibrium constant at the lower temperatures. If it is further assumed that the heat of any reaction of distribution of deuterium atoms among the various substituted acetones is zero, it may be predicted that $K_2 = 0.42 \times K_1^2 = 1.85$, about half the experimental result. Unfortunately, the discrepancy is large enough to make any definite statement concerning the rotation of the methyl groups unwise at present. If, on the other hand, it is assumed in calculating the entropy that the substituted methyl group is vibrating with a low saturated frequency, rather than rotating, the discrepancy between prediction and measurement is still greater. The experimental result lies between the limits imposed by free rotation and completely restricted rotation, but is in better agreement with free rotation.

Summary

The introduction of the first deuterium atom into acetone by dilute deuterium oxide is a pseudo-unimolecular reaction, with a high temperature coefficient, approximately zero heat of reaction, and a limiting equilibrium constant of 2.1. The extent of formation of the various acetones of higher deuterium content is somewhat greater than might be predicted.

ANN ARBOR, MICHIGAN RECEIVED JUNE 15, 1935

---

(2) Topley and Eyring, *J. Chem. Phys.*, **2**, 217 (1934).

(3) Halford, *ibid.*, **2**, 694 (1934).