[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY OF THE UNIVERSITY OF PENNSYLVANIA]

Solvent Effects and Ester Interchange in Basic Hydrolysis of Esters

BY GEORGE A. GALLAGHER, JOHN G. MILLER AND ALLAN R. DAY

RECEIVED MARCH 1, 1957

The saponification of ethyl acetate and methyl acetate has been measured at 30° in dioxane–water mixtures containing additional solvents found earlier to facilitate ammonolysis and concurrent hydrolysis of these esters. The effect of these additional solvents in the highly basic solutions is different from that found for the ammoniacal solutions. Ester interchange was observed with methanol and ethylene glycol as additives. The alcohols as well as glycol and some lactamides decrease the velocity of the hydrolysis, probably by solvating the hydroxyl ion through hydrogen bonding, thereby decreasing its reactivity. The effect of acetone, which increases the rate of saponification, is probably due to solvation of the activated complex.

In studies of the effect of hydroxylated substances and especially substances containing the groupings O–C–C–O or O–C–C–N upon the rate of ammonolysis of methyl acetate, it has been observed that these substances also accelerated the concurrent hydrolysis of the ester.[1] To investigate separately the effect of these and other solvents upon the hydrolysis reaction alone, the present study deals with the saponification of esters in dioxane–water systems to which these solvents were added in partial replacement of the dioxane. Thus, the reference standard reaction system was 20 molar water in dioxane and at the start of reaction contained 0.048 *M* NaOH and 0.040 *M* ester. In the other systems, the added substances merely replaced some of the dioxane, the standard amounts of water, NaOH and ester remaining constant.

Experimental

Reagents.—Dioxane, methyl acetate, methyl alcohol, ethylene glycol, lactamide, N-methyllactamide and N,N-dimethyllactamide were purified as described previously.[1] The pinacol was Eastman Kodak white label material and was used without further purification; m.p. 42.5–43°.

Ethyl acetate was dried over anhydrous potassium carbonate and fractionally distilled through a 35 cm. packed column; b.p. 77.0°, $n^{25}D$ 1.3704. Acetone was purified by refluxing with addition of small quantities of solid potassium permanganate until a pink color persisted. It was then fractionally distilled; b.p. 56.1°, $n^{25}D$ 1.3567. Ethyl alcohol was dried by the method of Lund and Bjerrum[2] and was then distilled; b.p. 78.4°, $n^{25}D$ 1.3599.

Procedure.—A stock solution of 0.1 *N* NaOH in dioxane was made by weighing the calculated amount of carbonate-free 0.2 *N* NaOH into a 1-liter volumetric flask, followed by weighed addition of sufficient water to give a final water concentration of 40 molar. The flask was filled close to the mark with dioxane, thermostated for two hours at 30° and then filled to the mark with dioxane. This stock solution was stored in a paraffin-coated flask under nitrogen. New solutions were made every three weeks. Determinations of the normality of the solution at the beginning and end of the three-week period showed a variation of only one part per thousand.

Runs were made in the following typical way. A freshly-prepared dioxane solution of known added solvent concentration was added volumetrically in desired amount to a reaction flask, followed by the requisite amount of stock solution, also measured volumetrically. After mixing, the flask was thermostated for at least two hours and then brought to volume with dioxane. Reproducibility of NaOH concentration at this stage was shown to be within one part per thousand.

Next, the ester was weighed to the nearest 0.1 mg., in a 1-ml. glass-stoppered weighing bottle and added to the thermostated reaction flask. The empty weighing bottle was also dropped in to ensure complete transfer of ester. By this procedure, shaking could be started within one to two seconds after addition of ester. After shaking for 25 seconds, the flask was replaced in the thermostat. Ten-ml. aliquots were removed periodically and added to 25 ml. of chilled 0.025 *N* HCl to stop reaction. Back-titration was run with 0.025 *N* NaOH.

The thermostat was kept within ±0.01 of 30° in all runs. Frequent reference determinations (no solvent catalyst added) were run.

Results

TABLE I

THE EFFECT OF VARIOUS ADDED SUBSTANCES ON THE RATE OF HYDROLYSIS OF ETHYL ACETATE AT 30° IN 0.048 *M* NaOH IN DIOXANE CONTAINING 20 *M* H_2O[a]

Added substance	Concn., *M*	Rate constant[b] ($k \times 10^2$)	Mean error $\times 10^2$
	..	6.98[c]	0.06
Water[d]	0.12	6.98	.02
	.20	7.12	.03
	.40	7.19	.17
	2.00	7.93	.08
	4.00	8.12	.10
	8.00	8.93	.13
Ethanol	0.40	6.33	.05
	2.00	6.01	.01
	6.00	4.24	.03
	8.00	3.63	.01
Acetone	0.40	7.21	.05
	2.00	7.36	.08
	6.00	7.56	.06
Pinacol	0.20	6.97	.04
	1.00	6.80	.14
Lactamide	0.12	4.49	.03
N-Methyllactamide	0.12	4.54	.04
	.20	3.51	.07
	.40	2.39	.02
	.80	1.49	.01
N,N-Dimethyllactamide	.12	6.54	.11

[a] Initial ester concentration 0.04 *M* throughout. [b] The rate constant in the usual bimolecular rate expression, $dx/dt = k(a - x)(b - x)$, expressed in liter mole^{-1} sec. Except where otherwise noted, each *k* is the mean of the separate values obtained from six samples taken at 120 second intervals in a single run. [c] This is the grand mean for eight separate runs. [d] The concentrations listed for H_2O are, of course, the concentrations above 20 *M*.

For the saponification of ethyl acetate, Table I lists the results showing the effects of acetone, pinacol, and a group of substances found earlier[1] to facilitate the ammonolysis and concurrent hydrolysis of esters. The values for the water–dioxane systems at 2, 4 and 8 *M* additional water may be compared with the results of other workers.

---

[1] T. A. Koch, J. G. Miller and A. R. Day, THIS JOURNAL, **75**, 953 (1953); G. R. Wolf, J. G. Miller and A. R. Day, *ibid.*, **78**, 4372 (1956).

[2] H. Lund and J. Bjerrum, *Ber.*, **64B**, 210 (1931).

The 2 and 8 M solutions correspond nearly to 60 and 50 volume % dioxane, concentrations at which Nair and Anantakrishnan[3] found k to be 7.44 × 10^{-2} and 8.58 × 10^{-2} l. mole^{-1} sec.$^{-1}$ at 30° for ethyl acetate. The 4 M system is about 57.5% dioxane by weight, and at exactly 57.9 weight % dioxane, Tommila and co-workers[4] found the value 8.16 × 10^{-2} l. mole^{-1} sec.$^{-1}$ at 30° for this ester. In the concentration range in which we have worked these and other workers have observed the same increase in rate with increase in amount of water.

The other results are not directly comparable to any other values in the literature for the reason that no one else has reported measurements of the hydrolysis of ethyl acetate in these ternary solvent mixtures.

Table II shows the behavior of the hydrolysis of ethyl acetate in water–methanol–dioxane and water–ethylene glycol–dioxane systems at 30°. Closely related to these data are the results shown in Table III for the hydrolysis of methyl acetate. While the k values were constant throughout any run for methyl acetate, they show large drifts upward for ethyl acetate in the presence of methanol and ethylene glycol. That these increases are due to ester interchange is apparent from the dependence of the change upon the amount of methanol or ethylene glycol present and of the near equality of the velocity constants of methyl acetate hydrolysis at 4 and 8 M methanol with the specific rates for ethyl acetate near the end of the runs in the presence of 4 and 8 M methanol. For ethyl acetate and NaOH concentrations nearly the same as ours, Tommila and co-workers[4] reported that a similar increase in k took place during reaction in methanol–water and ethylene glycol–water mixtures, this increase being ascribed to transesterification. These workers noted that when the solutions contained 70 weight % or more of methanol, k was almost constant, which they took to indicate that the conversion of ethyl acetate to methyl acetate is very rapid at that concentration. Our highest methanol concentration was only 27 weight %.

TABLE II

THE HYDROLYSIS OF ETHYL ACETATE AT 30° IN 0.048 M NaOH IN DIOXANE CONTAINING 20 M H_2O AND METHANOL OR ETHYLENE GLYCOL[a]

Additional substance	Concn., M	$k \times 10^2$ for sample taken at time (in sec.) shown: 120	240	360	480	600	720
Methanol	0.20	7.40	7.45[b]	7.74	7.71	7.81	7.74
	0.40	7.54	..	8.13	8.13	8.12	8.40
	2.00	6.80	7.51	8.15	8.49	8.91	9.18
	4.00	5.19	5.95	6.43	6.58	7.05	7.34
	4.00	5.17	5.83	6.35	6.67	7.44	7.10
	8.00	3.26	3.41	3.70	3.81	3.93	4.05
Ethylene glycol	0.20	7.35	7.69	7.92	7.93	8.17	8.22
	1.00	5.92	6.65	7.31	7.56	7.95	7.83
	2.00	3.85	4.62	5.15	5.58	5.88	6.10

[a] Initial ester concentration 0.04 M throughout. [b] At 245 seconds.

TABLE III

THE EFFECT OF ADDITION OF WATER OR METHANOL ON THE RATE OF HYDROLYSIS OF METHYL ACETATE AT 30° IN 0.048 M NaOH IN DIOXANE CONTAINING 20 M H_2O[a]

Added substance	Concn., M	$k \times 10^2$	Mean error × 10^2
Water	..	16.28	0.19
	0.40	16.18	.43
	2.00	16.92	.23
	6.00	16.99	.38
	8.00	17.22	.41
Methanol	..	16.27	.16
	0.40	15.25	.32
	2.00	10.94	.17
	4.00	7.39	.15
	6.00	5.50	.02
	8.00	4.22	.05

[a] The procedures, initial concentration of ester, and units of k were the same as for ethyl acetate, except that samples were taken at 90 instead of 120 second intervals.

Extrapolation of the ethyl acetate rates to zero time, *i.e.*, to a condition of no transesterification, shows that the saponification of that ester is hindered by both methanol and ethylene glycol, the magnitude of the effect of methanol being nearly the same as that of ethanol. The results also agree with the finding of Tommila[4] that in aqueous systems the solvent effects of ethylene glycol and methanol were closely similar in depressing the alkaline hydrolysis of ethyl acetate. Figures 1 and 2 show the extrapolation of the ethylene glycol runs and the methanol runs at higher methanol concentrations. At the lowest concentrations

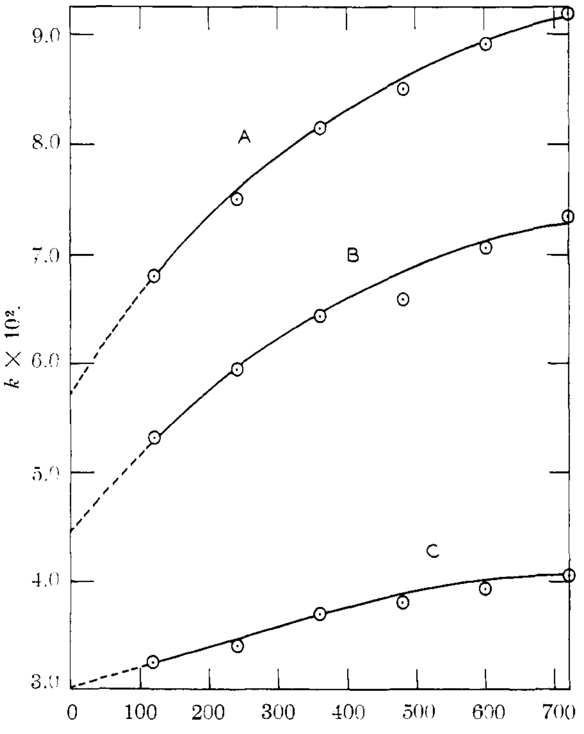


Fig. 1.—The extrapolation to zero time of the velocity constants for the alkaline hydrolysis of ethyl acetate in methanol–water–dioxane mixtures at 30°: (A) 2 M methanol; (B) 4 M methanol; (C) 8 M methanol.

---

[3] P. M. Nair and S. V. Anantakrishnan, *Proc. Indian Acad. Sci.*, **32A**, 187 (1952).

[4] E. Tommila, A. Koivisto, J. P. Lyrra, K. Antell and S. Heimo, *Ann. Acad. Sci. Fennicae.* [A] II, **47**, 3 (1952).

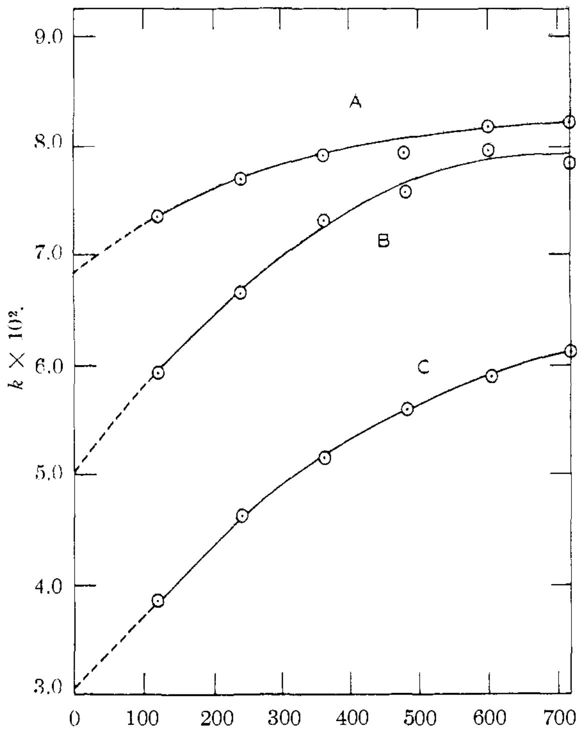


Fig. 2.—The extrapolation to zero time of the velocity constants for the alkaline hydrolysis of ethyl acetate in ethylene glycol–water–dioxane mixtures at 30°: (A) 0.2 *M* glycol; (B) 1 *M* glycol; (C) 2 *M* glycol.

the extrapolations of the methanol data were uncertain.

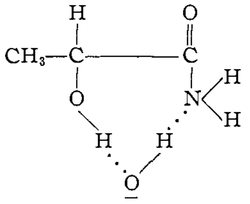
Discussion

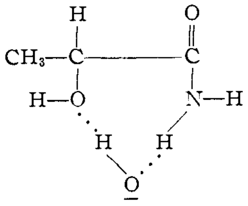
The normal alkaline hydrolysis of an ester[5] such as studied here is a reaction between an uncharged ester molecule and an hydroxyl ion. Due to the low solubility of esters in water, the reaction is often carried out by adding organic solvents to the water. The extensive recent work by Tommila and co-workers[4,6] has reviewed the many studies in this field and stresses the role of solvation of both the reactants and the transition state.

In measurements of the solvent effect in concurrent ammonolysis and hydrolysis of esters,[1] there were observed strikingly parallel effects upon the competing reactions by solvents containing structures suitable for production of hydrogen-bonded complexes with ammonia and water. The action of these solvents was looked upon as catalytic in the sense that by attachment of the solvent to the attacking species in these pre-equilibria, a more reactive particle was produced, in contrast to the usual stabilization caused by solvation. Since water was present in large amount in those systems, it appeared improbable that the action of solvent was through complete removal of a proton from ammonia to form free amide ions as the principal attacking species in ammonolysis. It also appeared unlikely that the effect of the added solvent was exerted through transesterification, for the rate constants did not show strong drifts during progress of reaction. Indeed, direct tests showed that no appreciable ester interchange in the simple ammonolysis of esters is probable in ethylene glycol–water mixtures.[7] In anhydrous systems containing alcohols, the reaction of free amide ions and the role of ester interchange may be more important.[8–10]

It is now apparent that the simulated hydroxyl ions pictured for hydrolysis in the ammoniacal reaction systems are of small importance in competition with the hydroxyl ions in the more basic media. Indeed, the addition of alcohol at nearly constant water concentration causes a decrease in reactivity of the hydroxyl ions in the highly basic media. Since Tommila[4,6] has shown that the effect due to complete transfer of a proton from alcohol to the ion probably is not appreciable until alcohol concentrations much higher than those used here are reached, the decrease in rate may be due to solvation of the type R–O–H. . .$\bar{\mathrm{O}}$H.

The lactamides exerted a strong effect in comparison with methanol or ethanol. Here, cyclic hydrogen bonded complexes of the types shown here for lactamide





might be operative. For N,N-dimethyllactamide, which showed relatively little activity, the second type would not exist and the first type might not form due to the bulk of the methyl groups attached to the nitrogen atom.

Pinacol would be expected to have small effect because of the steric hindrance and electron release effects of the methyl groups attached to the carbon atoms holding the hydroxyl groups. Kuhn[11] has suggested that in the presence of any given hydrogen acceptor, the hydrogen bond formed by various alcohols would be stronger the more acidic the alcohol. We note also that little ester exchange would be expected for this tertiary alcohol, as confirmed by the constancy of the velocity constants throughout the runs with pinacol.

Finally, we may inspect the effect of acetone. In binary mixtures with water, methanol and ethylene glycol always decrease the rate of basic hydrolysis of esters. On the other hand, addition of ethanol and other aliphatic alcohols, as well as acetone and dioxane, increases the reaction velocity at low temperatures and up to a mole fraction of

---

(5) C. K. Ingold, "Structure and Mechanism in Organic Chemistry," Cornell University Press, Ithaca, N. Y., 1953, p. 752 ff.

(6) E. Tommila, *Suomen Kemistilehti*, **25B**, No. 9, 37 (1952); E. Tommila and S. Hietala, *Acta Chem. Scand.*, **8**, 257 (1954); E. Tommila and S. Maltamo, *Suomen Kemistilehti*, **28B**, No. 3, 73, 118 (1955).

(7) M. Gordon, J. G. Miller and A. R. Day, THIS JOURNAL, **71**, 1250 (1949).

(8) W. H. Watanabe and L. R. DeFonso, *ibid.*, **78**, 4542 (1956).

(9) R. Baltzly, I. M. Berger and A. A. Rothstein, *ibid.*, **72**, 4149 (1950).

(10) F. H. Wetzel, J. G. Miller and A. R. Day, *ibid.*, **75**, 1150 (1953).

(11) L. P. Kuhn, *ibid.*, **74**, 2492 (1952).

less than 0.1 for the organic solvent. Tommila[4,6] has ascribed the increase in rate to the existence of both hydrophilic and hydrophobic groups in the activated complex. Because of the specific attractions of these groups, solvation of the complex should be greater in some mixtures of water and organic solvents than in water alone. The effects found for dioxane and acetone in binary mixtures with water were remarkable for their near equality,[4] and our results show that in the ternary solvent system dioxane–acetone–water these effects augment each other up to a mole fraction of acetone higher than 0.2.

PHILADELPHIA 4, PENNA.

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY, UNIVERSITY OF CALIFORNIA, LOS ANGELES]

The Photochemical Decomposition of Cyclohexanone, Cyclopentanone and Cyclobutanone[1]

BY F. E. BLACET AND ARNOLD MILLER

RECEIVED MARCH 2, 1957

Quantum yields of the photodecomposition products of these cycloketones have been determined in the ranges 3130–2537 Å. and 100–300°. Cyclohexanone principal products were carbon monoxide, cyclopentane and pentene-1; minor products were ethylene and propylene. At 2537 and 2654 Å. quantum yields were temperature independent. At 3130 Å. the yields of carbon monoxide, cyclopentane and pentene-1 increased with temperature. Cyclopentanone products were carbon monoxide, cyclobutane and ethylene. The yields were temperature dependent at 3130 Å., but independent at 2654 and 2537. Cyclobutanone products were carbon monoxide, cyclopropane, ethylene, propylene and presumably ketene. Yields were temperature independent at 3130 and 2654 Å. All the ketones gave some evidence of polymer formation and a polymerization quantum yield is given based on material balance. A general mechanism of photodecomposition has been proposed involving two primary chemical reactions, which accounts moderately well for all experimental results.

The photochemistry of cyclic ketones was initially investigated by Norrish and co-workers[2,3] and later by Benson and Kistiakowsky,[4] utilizing full radiation from mercury arcs. In this paper are given quantum yields obtained under controlled conditions of wave length and temperature for the vapor phase photolyses of cyclohexanone, cyclopentanone and cyclobutanone. Dunn and Kutschke[5] have published a few results obtained with cyclohexanone since the information given herein was reported originally.[1]

Experimental

Photolyses were performed in a 20-cm. cylindrical fused silica cell of 3 cm. diameter which could be maintained at desired temperatures. Monochromatic radiant energy obtained by the use of a quartz monochromator was measured with the aid of a phototube which had been calibrated at each wave length involved with a thermopile-galvanometer system. Photolysis products were identified and quantitatively analyzed by means of both mass spectrometric and microgas analytical techniques.

The cyclohexanone and cyclopentanone were prepared by drying and fractionating standard Eastman White Label material. Both the boiling points and refractive indices of the final products checked published values. Cyclobutanone was synthesized by the method of Roberts and Sauer.[6] The boiling point was 98.2° at 755 mm. and $n^{25}\text{D}$ was 1.4191. These values are to be compared to 98–100° and $n^{25}\text{D}$ 1.4189 reported by Roberts.

Absorption spectra of the three compounds determined in heptane solution, and determined also in the vapor phase for cyclohexanone and cyclopentanone, were in agreement with those reported previously.[4] The electronic absorption region of interest extends approximately from 2400 to 3400 Å. Cyclobutanone and cyclopentanone have several prominent absorption bands in this region, while the cyclohexanone spectrum has only a faint indication of structure.

Quantum yields obtained in the study of all three of the ketones are given in Table I. All values are averages based on duplicate runs at each temperature, except for one experiment shown for cyclobutanone in which water vapor was added to the reaction system. Mean deviations of quantum yields were from 4 to 7%. Exposure times averaged between three and four hours and at the several wave lengths absorbed intensity, I_a, varied between 1×10^{-8} and 1×10^{-9} Ein./l./sec. Where temperature ranges are indicated, quantum yields were found to be constant for runs

TABLE I

QUANTUM YIELDS OF PRODUCTS IN THE PHOTOLYSIS OF CYCLOHEXANONE, CYCLOPENTANONE AND CYCLOBUTANONE

Cyclohexanone

Wave length, Å. Temp., °C.	3130 125	3130 300	2654 100–300	2537 100–300
$\Phi_{\text{carbon monoxide}}$	0.22	0.91	0.78	0.81
Φ_{ethylene}	.02	.02	.02	.03
$\Phi_{\text{propylene}}$	.03	.03	.02	.02
$\Phi_{\text{cyclopentane}}$	.06	.31	.28	.21
$\Phi_{\text{pentene-1}}$	.12	.59	.45	.32
$\Phi_{\text{polymerization}}$	.01	−.02	.03	.25

Cyclopentanone

Wave length, Å. Temp., °C.	3130 125	3130 225	2654 100–300	2537 100–300
$\Phi_{\text{carbon monoxide}}$	0.33	0.42	0.61	0.74
Φ_{ethylene}	.15	.34	.25	.21
$\Phi_{\text{cyclobutane}}$	.26	.17	.27	.21
$\Phi_{\text{polymerization}}$	.01	.10	.21	.43

Cyclobutanone

Wave length, Å. Temp., °C.	3130 100–300	3130[a] 100	2654 100–300
$\Phi_{\text{carbon monoxide}}$	0.35	0.40	0.53
Φ_{ethylene}	.51	.54	.53
$\Phi_{\text{cyclopropane}}$	.13	.14	.17
$\Phi_{\text{propylene}}$	.014	.014	.12
$\Phi_{\text{polymerization}}$	.21	.25	.24

[a] 70 mm. H_2O vapor and 136 mm. cyclobutanone in reaction cell.

---

[1] Presented before the Division of Physical and Inorganic Chemistry at the Chicago meeting of the American Chemical Society, September, 1953.

[2] M. Saltmarsh and R. G. W. Norrish, *J. Chem. Soc.*, 455 (1935).

[3] C. H. Bamford and R. G. W. Norrish, *ibid.*, 1421 (1938).

[4] S. W. Benson and G. B. Kistiakowsky, THIS JOURNAL, **64**, 80 (1942).

[5] J. R. Dunn and K. O. Kutschke, *Can. J. Chem.*, **32**, 725 (1954).

[6] J. D. Roberts and C. W. Sauer, THIS JOURNAL, **71**, 3925 (1949).