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α,α-Dimethylacetoacetic Acid. Hydrolysis of the Ethyl Ester. Ketonic Decomposition. Reaction with Iodine and Bromine. Dissociation Constant

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The kinetics of the ketonic decomposition of acetoacetic acid

$$CH_3COCH_2COOH \longrightarrow CH_3COCH_3 + CO_2$$

have been investigated several times. Special importance attaches to the work of Widmark[1] and Ljunggren.[2]

The undissociated acid decomposes with considerable velocity, while the ion decomposes very slowly. The reaction is catalyzed by primary amines.

From the work on acetoacetic acid it was impossible to decide whether the keto or the enol form is active. Pedersen[3] found that α,α-dimethylacetoacetic acid, which cannot be rearranged into an enol form, undergoes the same kind of decomposition

$$CH_3COC(CH_3)_2COOH \longrightarrow CH_3COCH(CH_3)_2 + CO_2$$

It was therefore concluded that the keto form is unstable. The investigation of this acid has been continued in order to make a further study of the mechanism of ketonic decomposition, and especially of the amine catalysis. The catalysis is still being investigated and the results will be published later. This paper gives various other results of the study.

Purification of α,α-Dimethylacetoacetic Ethyl Ester.—A preparation from Kahlbaum was shaken for eight minutes with an equal volume of 25% potassium hydroxide.[4] It was washed with water, dried over anhydrous sodium sulfate and distilled twice in an all-glass apparatus fitted out with a Widmer column. At the last distillation three fractions were collected, the second containing two-thirds of the distillate.

Samples of the three fractions were left at room temperature with a little more sodium hydroxide than is necessary for hydrolysis. The next day, when the hydrolysis was complete, phenolphthalein was added, and the excess of sodium hydroxide was titrated with hydrochloric acid. The sodium hydroxide used in the hydrolysis agreed with the calculated amount within an accuracy of 0.2%. Now more hydrochloric acid was added, the total amount being more than equivalent to the sodium hydroxide originally used. The flasks were left for two days at 25°. After this time the ketonic decomposition is complete. The carbon dioxide was removed from the solutions by means of a current of pure carbon dioxide-free air, after which the solutions were titrated with sodium hydroxide. If the preparation is pure only neutral decomposition products of the ester are left, that is, the total amounts of acid and base used are equivalent. Actually, a little more base was used, the discrepancy corresponding to 0.3–0.5% of the ester. Thus the preparations contain a small amount of an ester of a stable acid.

The second fraction was used for most of the work. A few experiments were carried out with the two other fractions and an old preparation from the previous investigation.[3]

Rate of Hydrolysis of Dimethylacetoacetic Ester.—The reaction was started by pouring a carbon dioxide-free aqueous solution of the ester into a silica flask containing an excess of carbonate-free sodium hydroxide solution. At the start both solutions were at the temperature of the experiment, 0 or 24.95°. At suitable time intervals samples were taken with a pipet. In order to retard the reaction they were run into ice water to which had been added phenolphthalein and some 0.02 *N* hydrochloric acid, not quite sufficient to neutralize the sodium hydroxide. The solutions were titrated immediately with hydrochloric acid. The reaction was followed until at least four-fifths of the ester was hydrolyzed. The results always agree well with the bimolecular law. The velocity constants of four experiments are given in Table I. The second and third columns give the initial stoichiometric concentrations of sodium hydroxide and ester. The formulas given under the table summarize the results. The critical increment is 10.6×10^3 cal./mole.

It is of interest to compare this reaction with the alkaline hydrolysis of acetoacetic ethyl ester, studied by Goldschmidt and Oslan.[5] The latter is a first order reaction, the rate being independent of the excess of alkali. The investigators give the following explanation. Acetoacetic ethyl ester is a weak acid (dissociation constant[5] 2×10^{-11}). In the alkaline solution the equilibrium

$$EH + OH^- \rightleftarrows E^- + H_2O \qquad (1)$$

where EH is the ester, is practically completely

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[1] Widmark, *Acta Med. Scand.*, **53**, 393 (1920); *Scand. Arch. Physiol.*, **42**, 43 (1922).

[2] Ljunggren, "Katalytisk Kolsyreavspjälkning ur Ketokarbonsyror," Diss., Lund, 1925.

[3] Pedersen, THIS JOURNAL, **51**, 2098 (1929).

[4] Auwers, *Ber.*, **46**, 505 (1913).

[5] Goldschmidt and Oslan, *Ber.*, **33**, 1140 (1900).

TABLE I

RATE OF HYDROLYSIS OF DIMETHYLACETOACETIC ETHYL ESTER

$k^* = 0.4343\ k$, where k is the bimolecular constant

T., °C.	$(\mathrm{NaOH})_0$	$(\mathrm{Ester})_0$	k^* (liter/mole min.)
24.95	0.03664	0.03182	0.253
	.04024	.02974	.255
		Mean	.254
0.0	.05847	.04347	.0489
	.06800	.04066	.0500
		Mean	.0494

$\log k^* = 7.189 - 2.320 \times 10^3 \times T^{-1}$ (liter/mole min.)

or $k = 5.93 \times 10^5 e^{-\frac{10.6 \times 10^3}{RT}}$ (liter/mole sec.)

displaced in favor of the ester ion. The hydrolysis takes place between EH and OH^-, while E^- does not react, but $(\mathrm{EH}) \times (\mathrm{OH}^-)$ is proportional to (E^-), that is, to the total ester concentration. If K is the mass law constant of equilibrium 1, k_{uni} and k_{bi} the apparent and real velocity constants, we have

$$k_{\mathrm{uni}}\ (\mathrm{E}^-) = k_{\mathrm{bi}}\ (\mathrm{EH})(\mathrm{OH}^-), \text{ or}$$
$$k_{\mathrm{uni}} = k_{\mathrm{bi}}/K$$

from which we obtain

$$Q_{\mathrm{uni}} = Q_{\mathrm{bi}} - Q$$

where Q_{uni} is the apparent critical increment, Q_{bi} is the actual critical increment for the bimolecular reaction, and Q is the heat absorbed when reaction 1 takes place from left to right. By calorimetry Pedersen[6] has found that $Q = -7.28 \times 10^3$ cal./mole at 20°. If we use for Q_{bi} the value found above for the dimethyl substituted ester 10.6×10^3 cal./mole, we calculate $Q_{\mathrm{uni}} = 17.9 \times 10^3$ cal./mole. In the absence of experimental data for acetoacetic ethyl ester, we may compare this value with the critical increment for the methyl ester $Q_{\mathrm{uni}} = 17.7 \times 10^3$ cal./mole calculated from Ljunggren's[2] measurements at 25 and 37°. The good agreement is a new evidence in favor of the explanation given by Goldschmidt and Oslan.

Rate of Ketonic Decomposition in Hydrochloric Acid

The ethyl ester was weighed out, dissolved in a suitable excess of sodium hydroxide and left at 25 or 35° until hydrolysis was complete. Just before the experiment a suitable amount of hydrochloric acid and, in most cases, sodium chloride solution were added. The initial concentration of dimethylacetoacetic acid was always about 0.02 M.

The decomposition was followed by measuring the carbon dioxide pressure above the solution by means of an apparatus described in an earlier paper.[7] The concentration of undecomposed dimethylacetoacetic acid at the time, t, is proportional to the difference, P, between the final pressure and the pressure at the time, t.

Dimethylacetoacetic acid is partly dissociated in the solutions of hydrochloric acid used. We may call the degree of dissociation α. The ion decomposes so slowly that we may disregard it here. Consequently the unimolecular velocity constant calculated from the experiments is $k = k_0(1-\alpha)$, where k_0 is the velocity constant for the undissociated acid. Owing to the decomposition of the acid the hydrogen-ion concentration decreases a little during the reaction. However, it is only in the experiments with the smallest concentration of hydrochloric acid that this change is of any importance. We may write

$$\mathrm{d} \log_e P/\mathrm{d}t = k_0(1 - \alpha_\infty) + k_0(\alpha_\infty - \alpha)$$

where α_∞ is the final value of α. By integration we find

$$\log P_0 - \log P = k_0^*(1 - \alpha_\infty)t + \Delta$$

where

$$\Delta = k_0^* \int_0^t (\alpha_\infty - \alpha)\mathrm{d}t$$

The small correction Δ is easily found by rough graphical integration, where α is calculated from the known composition of the solution at different times during the reaction, the dissociation constant of the acid and a rough value of k_0^*. We plot $\log P + \Delta$ against t. The numerical value of the slope is $k^* = k_0^*\ (1-\alpha_\infty)$.

The experiments, which were carried out at 24.97 and 34.93°, always agree well with the unimolecular law. Fraction 2 of the ester was used. In two cases parallel experiments with all three fractions were carried out. The results are given in Tables II and III. There is a distinct decrease in k^* as the hydrochloric acid concentration decreases. In the last column is given $k_0^* = k^*/(1 - \alpha_\infty)$. α_∞ has been calculated from the concentration of hydrochloric acid and the dissociation constant K determined electrometrically in the last part of this paper. The constancy of k_0^* is usually quite good within a series of experiments at constant total salt concentration.

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[6] Pedersen, *Kgl. Danske Videnskab. Selskab, Math. fys. Medd.*, **12**, No. 2 (1932).

[7] Pedersen, THIS JOURNAL, **53**, 18 (1931).

TABLE II

KETONIC DECOMPOSITION OF DIMETHYLACETOACETIC ACID IN MIXTURES OF HYDROCHLORIC ACID AND SODIUM CHLORIDE AT 24.97°. $k^* = 0.4343\ k$ (min.$^{-1}$)

	(HCl)	$k^* \times 10^3$	$k_0^* \times 10^3$
(HCl) + (NaCl)	0.0272	1.790	1.818
= 0.050	.0272	1.798[a]	1.826
$K = 4.2 \times 10^{-4}$	.0272	1.790[b]	1.818
	.0270	1.789	1.817
	.0169	1.768	1.812
	.0104	1.743	1.814
	.0053	1.672	1.805
		Mean	1.816
(HCl) + (NaCl)	0.0768	1.801	1.811
= 0.100	.0751	1.797	1.808
$K = 4.65 \times 10^{-4}$	.0517	1.793	1.809
	.0248	1.781	1.814
	.0140	1.763	1.822
	.0102	1.742	1.821
	.0056	1.695	1.835
		Mean	1.817
(HCl) + (NaCl)	0.1772	1.813	1.818
= 0.200	.1772	1.813[a]	1.818
$K = 5.1 \times 10^{-4}$	.1772	1.807[b]	1.812
	.0562	1.806	1.822
	.0167	1.786	1.841
	.0049	1.680	1.851
		Mean	1.827
(HCl) + (NaCl)			
= 0.500	0.478	1.833	1.835
$K = 5.7 \times 10^{-4}$			
(HCl) + (NaCl)			
= 1.1	1.06	1.886	1.886
(HCl) + (NaCl)	2.06	1.924	1.924
= 2.1	0.077	1.903	1.914
$K = 4.5 \times 10^{-4}$	.0072	1.812	1.925
		Mean	1.921

[a] Ester, fract. 3. [b] Ester, fract. 1.

TABLE III

KETONIC DECOMPOSITION OF DIMETHYLACETOACETIC ACID IN MIXTURES OF HYDROCHLORIC ACID AND SODIUM CHLORIDE AT 34.93°. $k^* = 0.4343\ k$ (min.$^{-1}$)

	(HCl)	$k^* \times 10^3$	$k_0^* \times 10^3$
(HCl) + (NaCl)	0.0776	6.30	6.33
= 0.100	.0526	6.29	6.34
$K = 4.2 \times 10^{-4}$	.0273	6.30	6.40
	.0272	6.30	6.40
	.0132	6.15	6.35
	.0081	6.05	6.36
	.0065	5.96	6.34
	.0051	5.90	6.39
		Mean	6.36
(HCl) + (NaCl)	0.1768	6.34	6.36
= 0.200	.0776	6.34	6.38
$K = 4.8 \times 10^{-4}$	.0273	6.30	6.41
	.0081	6.03	6.39
		Mean	6.38

The salt effect on k_0^* is only measurable in rather concentrated solution. The following may be considered the best values in dilute solution

at 24.97° $k_0^* = 0.001816$ (min.$^{-1}$), and
at 34.93° $k_0^* = 0.00636$ (min.$^{-1}$)

The results may be expressed by the formula

$$\log k_0^* = 14.094 - 5.018 \times 10^3 \times T^{-1} \text{ (min.}^{-1}\text{), or} \quad (2)$$

$$k_0 = 4.76 \times 10^{12} e^{-\frac{23.0 \times 10^3}{RT}} \text{ (sec.}^{-1}\text{)} \quad (3)$$

From the last expression we see that the critical increment is 23.0×10^3 cal./mole. If we extrapolate to 18.0° by means of formula 2, we find $k_0^* = 7.18 \times 10^{-4}$. In the earlier paper[3] we found by experiment at this temperature $k_0^* = 7.23 \times 10^{-4}$.

From Ljunggren's[2] measurements of the ketonic decomposition of unsubstituted acetoacetic acid at 25 and 37° we derive the formulas

$$\log k_0^* = 14.352 - 5.282 \times 10^3 \times T^{-1} \text{ (min.}^{-1}\text{), or} \quad (4)$$

$$k_0 = 8.63 \times 10^{12} e^{-\frac{24.2 \times 10^3}{RT}} \text{ (sec.}^{-1}\text{)} \quad (5)$$

At 25° the dimethyl substituted acid decomposes 4.2 times as fast as the unsubstituted acid. As seen from a comparison of the formulas 3 and 5, substitution of methyl groups for the α-hydrogen atoms has the effect of diminishing both the critical increment and the factor before the exponential function. At ordinary temperatures the acceleration due to the former outweighs the retardation due to the latter.

Rate of Ketonic Decomposition in Alkaline Solution.—The very slow ketonic decomposition of the dimethylacetoacetate ion was studied in barium hydroxide solution at 24.97 and 34.93°. Dimethylacetoacetic ethyl ester was weighed out and dissolved in a graduated flask in 28–62% more barium hydroxide solution than sufficient for hydrolysis. After a few hours at the temperature of the experiment hydrolysis was complete. The solution, containing barium dimethylacetoacetate together with the excess of barium hydroxide, was now poured into a glass-stoppered bottle covered inside with a layer of paraffin wax. It was again placed in the thermostat. The following total reaction takes place

$$CH_3COC(CH_3)_2COO^- + OH^- + Ba^{++} \longrightarrow CH_3COCH(CH_3)_2 + BaCO_3$$

One equivalent of strong base is used per mole of dimethylacetoacetate ion. Barium carbonate forms a precipitate. At suitable intervals samples were taken from the clear solution and titrated with 0.02 N hydrochloric acid. The reaction follows the unimolecular law with no detectable discrepancy. The numerical data are given in Table IV. The experiments show that the hydroxyl-ion concentration has no detectable influence. Aniline, which under special circumstances is a very powerful catalyst, has no influence in sufficiently alkaline solution. This is seen from experiment 5 where 0.1 *M* aniline has been added. We consider the mean of k_1 from the two first experiments as the best value at 24.97° because the experiments 3–5 are less accurate.

TABLE IV

KETONIC DECOMPOSITION OF THE DIMETHYLACETOACETATE ION

k_1^* velocity constant (dec. log, min.$^{-1}$). Initial Concentration of Dimethylacetoacetate Ion in all Experiments c = 0.1

Expt. no.	Ester	(OH$^-$) at t = 0	$k_1^* \times 10^6$	Length of experiment (min.) $\times 10^{-3}$
		24.97°		
1	Fract. 1	0.0460	4.00	35
2	Old prep.	.0374	3.98	17
3	Old prep.	.0283	4.0	12
4	Old prep.	.0621	4.1	12
5	Old prep.	.0283[a]	4.2	12
	Means of nos. 1 and 2		3.99	
		34.93°		
6	Fract. 1	0.0409	17.42	13
7	Fract. 3	.0434	17.50	13
		Mean	17.5	

[a] The solution contained aniline, 0.1 *M*.

It is well known that acetoacetic acid in strongly alkaline solution undergoes the so-called acid decomposition

$$CH_3COCH_2COOH + H_2O \longrightarrow 2CH_3COOH$$

In order to show that a corresponding acid decomposition of dimethylacetoacetic acid does not take place to any appreciable extent in the above experiments, the following experiments were carried out.

Solutions containing 0.1 *M* ethyl ester and about 0.15 *M* sodium hydroxide were hydrolyzed and poured into glass-stoppered bottles covered inside with paraffin wax. They were placed in a thermostat at 25 and 35° for ketonic decomposition in alkaline solution. At suitable intervals samples were taken. They were added to a little more hydrochloric acid than was equivalent to the original concentration of sodium hydroxide, and then left for two days at 25° or for one day at 35° for complete ketonic decomposition. Finally, they were titrated with sodium hydroxide. If we have a pure ketonic decomposition, both in alkaline and acid solution, the total amount of sodium hydroxide used will be equivalent to the hydrochloric acid. This was always found to be the case, even when the alkaline decomposition had been continued for 3×10^4 minutes at 25° or for 10^4 minutes at 35°. Only a small and constant deviation was found, corresponding to the slight content of an ester of a stable acid in the preparation.

From the experiments we derive the following formulas for the velocity constant k_1 of the ketonic decomposition of the dimethylacetoacetate ion

$$\log k_1^* = 14.452 - 5.917 \times 10^3 \times T^{-1} \text{ (min.}^{-1}) \quad (6)$$

or

$$k_1 = 1.09 \times 10^{13} e^{-\frac{27.1 \times 10^3}{RT}} \text{ (sec.}^{-1}) \quad (7)$$

By comparison with formulas 2 and 3 we see that the slowness of the decomposition of the ion is due to the large critical increment, 27.1×10^3 cal./mole for the ion compared with 23.0×10^3 cal./mole for the acid.

Ljunggren[2] has studied the ketonic decomposition of the acetoacetate ion at 25 and 37°. His results lead to the expressions

$$\log k_1^* = 11.322 - 4.880 \times 10^3 \times T^{-1} \text{ (min.}^{-1}) \quad (8)$$

or

$$k_1 = 8.05 \times 10^9 e^{-\frac{22.3 \times 10^3}{RT}} \text{ (sec.}^{-1}) \quad (9)$$

The large difference between the expressions for the dimethylacetoacetate ion and the acetoacetate ion is rather surprising.

Rate of Reaction with Iodine and Bromine.—The experiments were carried out in a thermostat at 24.97°. The ester was hydrolyzed with an excess of sodium hydroxide. Equal volumes of the hydrolysate were run into a series of bottles all containing the same concentration of hydrochloric acid and potassium iodide, but varying concentrations of free iodine. The time t from the mixing of the solutions until the disappearance of the yellow color was measured. We denote by c and x the initial concentrations of dimethylacetoacetic acid and iodine in moles/liter. In the figure $\log c/(c - x)$ has been plotted against t for two experiments (circles). The composition of the solutions is given under the figure. The straight lines indicate the rate of carbon dioxide cleavage in a solution of the same concentration. It is seen that the points fall exactly on the lines. This shows that the rate of reaction with iodine is completely determined by the rate of ketonic decomposition.

The results of similar experiments with bromine instead of iodine are also given in the figure. In one experiment (open squares) the hydrolysate was added to hydrochloric acid containing varying concentrations of bromine. In another experiment (solid squares) the procedure was slightly altered in order to avoid evaporation of bromine. Here the hydrolysate was run into a mixture of hydrochloric acid and potassium bromide. Immediately afterward bromine was liberated by dropping in a micro crucible containing varying amounts of potassium bromate. As seen from the graph the rate of bromination is undoubtedly almost completely determined by the ketonic decomposition. However, the agreement is not quite so good as for the iodination.

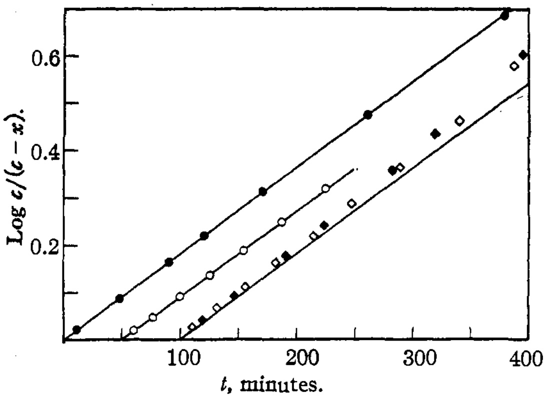

Fig. 1.—Reaction of dimethylacetoacetic acid with iodine (circles) and bromine (squares). Solid circles: HCl 0.16 *M*, NaCl 0.04 *M*, KI 0.10 *M*. Open circles: HCl 0.087 *M*, NaCl 0.013 *M*, KI 0.025 *M*. Solid squares: HCl 0.16 *M*, NaCl 0.04 *M*, KBr 0.10 *M*. Open squares: HCl 0.08 *M*, NaCl 0.02 *M*. Straight lines, carbon dioxide cleavage.

It is most unlikely that dimethylacetoacetic acid reacts directly with halogen, and the ketone which is the final product of the ketonic decomposition reacts only very slowly with halogen. The reacting molecule must be an intermediate product, undoubtedly an enol form. We may conclude that the carbon dioxide cleavage of β-keto carboxylic acids leads directly to the enol form of the reaction product. This reacts instantaneously with free halogen. In the absence of halogen it is rapidly transformed into the corresponding ketone.

In an earlier paper[8] the carbon dioxide cleavage of the α-nitrocarboxylic acids was studied. From experiments on the bromination of dimethylnitroacetic acid it was concluded that the decomposition leads directly to the aci-form of the nitrohydrocarbon. This is rapidly transformed into the true nitrohydrocarbon. There is thus an analogy between the mechanisms of the cleavages of the α-nitro and β-keto carboxylic acids. This analogy was already pointed out in the earlier paper where a possible mechanism was discussed.

In addition to the β-keto and α-nitro carboxylic acids a third group, malonic acid and its substitution products, split off carbon dioxide according to the same mechanism. This has been shown by Jytte Muus[9] for dibromomalonic acid.

$$CBr_2(COOH)_2 \longrightarrow CBr_2HCOOH + CO_2$$

If bromine is present, it is used up at nearly the same rate as carbon dioxide is split off. From this result it is concluded that the decomposition leads directly to an unstable "enol" form of dibromoacetic acid $CBr_2{=}C(OH)_2$, which reacts instantaneously with bromine, forming tribromoacetic acid.

The Dissociation Constant of Dimethylacetoacetic Acid.—For the study of the ketonic decomposition, and especially the amine catalysis, it was necessary to know the strength of dimethylacetoacetic acid in salt solutions. Solutions of the acid (denoted by HB) and the sodium salt (NaB) of nearly equal concentrations, were measured with the hydrogen and quinhydrone electrodes against the solution: 0.01000 *M* hydrochloric acid + 0.0900 *M* sodium chloride. 3.5 *M* potassium chloride was used as bridge solution. The buffer solutions were prepared by hydrolyzing the ester at 25° with an excess of sodium hydroxide, and, just before the measurement, adding hydrochloric acid and sodium chloride.

Owing to the unavoidable ketonic decomposition of the acid, (HB) decreases during the measurement. From the equation

$$\frac{F \log e}{RT}\frac{\mathrm{d}E}{\mathrm{d}t} = -\frac{\mathrm{d}\log(\mathrm{H}^+)}{\mathrm{d}t} = -\frac{\mathrm{d}\log(\mathrm{HB})}{\mathrm{d}t} = k_0{}^* \quad (10)$$

where k_0 is the unimolecular velocity constant, we see that the potential difference E will rise with constant rate. By linear extrapolation we may find E to the time of the addition of hydrochloric acid. However, other complications also occur.

In the measurements with the quinhydrone electrode E was found to increase linearly with the time, but the rate of increase was much greater than would be expected from the ketonic decomposition. Quinhydrone reacts with the buffer. An analogous reaction was found in earlier meas-

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(8) Pedersen, *J. Phys. Chem.*, **38**, 559 (1934).

(9) Jytte Muus, *ibid.*, **39**, 343 (1935); and unpublished paper.

urements in acetoacetate buffers.[10] E for the original buffer may be found by linear extrapolation, but not without making the determination less reliable.

The hydrogen electrode vessels were fitted out with a porous glass filter (Schott und Gen., G.2) through which the hydrogen was passed into the solution, thus producing a large number of tiny bubbles. Owing to the great surface of the hydrogen, we get a quick saturation, and therefore a quick adjustment of the electrode. However, a new difficulty arose. The potential measured was dependent on how quickly the hydrogen passed through the solution. Probably a reaction between the keto acid and the hydrogen takes place on the platinized electrode. When the solution is not stirred violently enough with the current of hydrogen the reaction products may accumulate in the neighborhood of the electrode and cause a change of E. The hydrogen was therefore passed through the solution as quickly as possible without risking undue foaming. The fact that hydrogen and quinhydrone electrode measurements usually agree well is an evidence of the reliability of the results.

According to measurements of Bjerrum and Unmack[11] the activity coefficient of the hydrogen ion f_{H^+} in sodium chloride solution is given by the formulas

$$\text{at } 25° \quad -\log f_{\mathrm{H}^+} = 0.161\sqrt[3]{s} - 0.178s - 0.003 \qquad (11)$$
$$\text{at } 35° \quad -\log f_{\mathrm{H}^+} = 0.146\sqrt[3]{s} - 0.161s - 0.005 \qquad (12)$$

where s is the total salt concentration (equal to the ionic strength). The formulas hold when $0.001 < s < 1.5$. Although the formulas are only found to hold for sodium chloride solutions containing a little hydrochloric acid, we have used them both for the solution 0.01000 M hydrochloric acid + 0.0900 M sodium chloride and for the buffer solutions. For the former we get (at 25°) $-\log f_{\mathrm{H}^+} = 0.054$ and (at 35°) $-\log f_{\mathrm{H}^+} = 0.047$. We may therefore calculate $-\log a_{\mathrm{H}^+}$, where a_{H^+} is the hydrogen-ion activity, for the buffer solutions by means of the formulas

$$(\text{at } 25°) \quad -\log a_{\mathrm{H}^+} = 2.054 + E/0.05915$$
$$\text{and } (\text{at } 35°) \quad -\log a_{\mathrm{H}^+} = 2.047 + E/0.06113$$

In the quinhydrone measurements we further add 0.049 s, an approximate correction for the salt error of the electrode.[12] From $-\log a_{\mathrm{H}^+}$ thus computed, and $-\log f_{\mathrm{H}^+}$ found from the formulas 11 and 12 we calculate the hydrogen-ion concentrations of the solutions.

We denote by $(\mathrm{HB})_{stoich.}$ and $(\mathrm{NaB})_{stoich.}$ the stoichiometric concentrations of the buffer constituents. Their real concentrations are

$$(\mathrm{HB}) = (\mathrm{HB})_{\mathrm{stoich.}} - (\mathrm{H}^+) \text{ and}$$
$$(\mathrm{B}^-) = (\mathrm{NaB})_{\mathrm{stoich.}} + (\mathrm{H}^+)$$

We now calculate the dissociation constants

$$K_{\mathrm{ac}} = a_{\mathrm{H}^+}(\mathrm{B}^-)/(\mathrm{HB}) \text{ and}$$
$$K = (\mathrm{H}^+)(\mathrm{B}^-)/(\mathrm{HB})$$

TABLE V

THE DISSOCIATION CONSTANT OF DIMETHYLACETOACETIC ACID IN SOLUTIONS CONTAINING SODIUM CHLORIDE AT 25°

$(\mathrm{HB})_{\mathrm{stoich.}}$ and $(\mathrm{NaB})_{\mathrm{stoich.}}$ about 0.0350

Electrode	s	$-\log K_{\mathrm{ac}}$	$\Delta_{\mathrm{ac}} \times 10^3$	$-\log K$	$\Delta \times 10^3$
H	0.0755	3.420	−1	3.368	−1
H	.0755	3.420	−1	3.368	−1
Q	.0755	3.420	−1	3.368	−1
Q	.0755	3.413	+6	3.361	+6
H	.125	3.387	+1	3.332	+1
Q	.125	3.384	+4	3.329	+4
H	.199	3.353	+3	3.297	+4
H	.199	3.360	−4	3.304	−3
Q	.199	3.349	+7	3.293	+8
H	.292	3.327	−1	3.275	+1
Q	.292	3.324	+2	3.272	+4
H	.385	3.305	−1	3.359	+1
Q	.385	3.297	+7	3.251	+9

The numerical results of a series of measurements are given in Table V. Here H and Q denote hydrogen and quinhydrone electrode, respectively. The solutions contain so much sodium chloride that the total salt concentration s is that given in the second column of the table. The third column gives $-\log K_{\mathrm{ac}}$. The results of this and three other series of measurements may be expressed by the four formulas, where $b \equiv (\mathrm{HB}) + (\mathrm{B}^-)$

$$(25°,\ b = 0.025) \quad -\log K_{\mathrm{ac}} = 3.522 - 0.504\sqrt{s} + 0.243s$$
$$(25°,\ b = 0.070) \quad -\log K_{\mathrm{ac}} = 3.542 - 0.504\sqrt{s} + 0.193s$$
$$(25°,\ b = 0.125) \quad -\log K_{\mathrm{ac}} = 3.567 - 0.504\sqrt{s} + 0.189s$$
$$(35°,\ b = 0.070) \quad -\log K_{\mathrm{ac}} = 3.582 - 0.513\sqrt{s} + 0.205s$$

The coefficients of the square root terms have been calculated from the Debye–Hückel theory. The fourth column of Table V gives Δ_{ac}, the difference between $-\log K_{\mathrm{ac}}$ calculated from the formulas and found by the measurement. In the three other series of measurements, which were also carried out both with the hydrogen and quinhydrone electrode, the agreement is equally good.

By means of the formulas we may extrapolate to sodium chloride-free solutions. The results at

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(10) Pedersen, *J. Phys. Chem.*, **38**, 993 (1934).

(11) Bjerrum and Unmack, *Kgl. Danske Videnskab. Selskab, Math. fys. Medd.*, **9**, No. 1 (1929).

(12) Sörensen, Sörensen and Linderström-Lang, *Compt. rend. trav. lab. Carlsberg*, **14**, No. 14 (1921).

25° are given in the second column of Table VI, and may be expressed by the formula

$$-\log K_{ac} = 3.511 - 0.504\sqrt{s} + 1.09s \qquad (13)$$

TABLE VI

$s = b/2$	$-\log K_{ac} + 0.504\sqrt{s}$	
	Expt., by extrap.	Form 13
0.0125	3.525	3.525
.0350	3.549	3.549
.0625	3.579	3.579

The third column of Table VI contains $-\log K_{ac} + 0.504\sqrt{s}$ calculated from this formula. By means of formula 13 we may extrapolate to infinite dilution. We thus find the activity dissociation constant $K^0 = 3.08 \times 10^{-4}$ at 25°. We may compare this value with that found in an earlier work[3] from the ketonic decomposition of dimethylacetoacetic acid in glycolate buffer solutions at 18°. Here it was found that the ratio of the K^0 values for dimethylacetoacetic acid and glycolic acid is 2.04. If we use the value $K^0 = 1.48 \times 10^{-4}$ for glycolic acid[10] we find for dimethyl acetoacetic acid at 18° $K^0 = 3.02 \times 10^{-4}$, in good agreement with the value found electrometrically.

Column 5 of Table V gives $-\log K$ calculated from the experiments. It is convenient to have a single formula which with sufficient accuracy expresses all the results at a given temperature. We find the formulas

$$(\text{at } 25°)\ -\log K = 3.555 + 0.40b - 0.55\sqrt[3]{s} + 0.20s$$

and

$$(\text{at } 35°)\ -\log K = 3.599 + 0.40b - 0.55\sqrt[3]{s} + 0.20s$$

In the last column of Table V is given Δ, the difference between $-\log K$ calculated from the formula and found by experiment. In the three other series of experiments the agreement is equally good.

The values of E used for the calculation of the dissociation constants have always been found from the E values measured at different times by linear extrapolation to the time when the solutions were mixed. In the hydrogen electrode measurements at 25° the rate of change of E was in average $dE/dt = 0.110 \times 10^{-3}$ volts/min. According to formula 10 this rate corresponds to a ketonic decomposition with the velocity constant $k_0^* = 0.00186$. In the kinetic experiments (Table II) was found $k_0 = 0.00182$. At 35° the average value was $dE/dt = 0.387 \times 10^{-3}$ volt/min., corresponding to $k_0^* = 0.00633$. In the kinetic experiments (Table III) was found $k_0^* = 0.00636$. In the measurements with the quinhydrone electrode the rate of change of E was greater and varied with the concentration of the buffer constituents. For the solutions measured at 25° it varied from 0.14×10^{-3} to 0.28×10^{-3} volt/min., while at 35° it was about 0.73×10^{-3} volt/min.

I wish to express my thanks to the head of the laboratory, Professor Niels Bjerrum, for his kind interest in my work.

Summary

The paper contains miscellaneous studies of α,α-dimethylacetoacetic acid.

1. The rate of hydrolysis of its ethyl ester at 0 and 25°.
2. The rate of ketonic decomposition of the acid in mixtures of hydrochloric acid and sodium chloride at 25 and 35°.
3. The rate of ketonic decomposition of the ion at 25 and 35°.
4. The rate of reaction with iodine and bromine. (From the experiments it is concluded that the ketonic decomposition of β-keto carboxylic acids leads directly to the enol form of the reaction product.)
5. The dissociation constant, measured by means of the hydrogen and quinhydrone electrodes in solutions containing sodium chloride at 25 and 35°.

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