[CONTRIBUTION FROM THE NOYES CHEMICAL LABORATORY, UNIVERSITY OF ILLINOIS]

The Existence of a 1,2-Addition, 1,2-Elimination Mechanism for Decarboxylation[1]

BY ELIAS J. COREY AND GIDEON FRAENKEL

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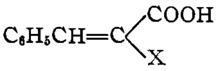
The rate equation for the decarboxylation of benzylidenemalonic acid derivatives (I) in pyridine is $dCO_2/dt = k'[HA][C_5H_5NH^+] + k''[HA]^2$ where [HA] is the concentration of un-ionized acid. The ratio of k'/k'' is at least 350. The order of reactivity of *p*-substituted acids is $p\text{-}NO_2 > p\text{-}H > p\text{-}CH_3 > p\text{-}OCH_3$. The presence of two ortho substituents strongly inhibits decarboxylation. Together the experimental data indicate the operation of a mechanism involving the concerted addition of a proton (from $C_5H_5NH^+$ or HA) and a solvent molecule to the α,β-double bond of I to give an intermediate which undergoes decarboxylative elimination. Catalysis of the decarboxylation in pyridine by thioacetic acid, which would be expected on the basis of the above interpretation, has been found.

It has been demonstrated in previous work[2] that alkylidenemalonic acid derivatives undergo decarboxylation by a three-step process involving isomerization to a β,γ-unsaturated malonic acid derivative, anionic decarboxylation and protonation of the resulting mesomeric anion. The present paper deals with a study of the decarboxylation processes which are operative in the case of arylidenemalonic acid derivatives. It has been observed frequently in the past that these acids decarboxylate readily upon heating. From the outset it seemed likely that such decarboxylation proceeds *via* an intermediate which does not possess the original α,β-ethylenic linkage. One might anticipate, on the basis of structural considerations, that direct decarboxylation should be quite slow, because of the presence of the α,β-double bond.

Results

Ethyl hydrogen benzylidenemalonate (IA) and benzylidenecyanoacetic acid (IB) decarboxylate readily in pyridine near the boiling point. The acid ester IA, which could be obtained only as a mixture of isomers (m.p. 85–90°), yields pure *trans*-ethyl cinnamate. The cyano acid IB, which was obtained as a single isomer (C_6H_5, CN *cis*) yields a mixture consisting of 65% *trans*- and 35% *cis*-cinnamonitrile. In each case the product obtained represents the equilibrium mixture of *cis* and *trans* isomers.[3] It appears likely that equilibration of the isomeric decarboxylation products takes place under the conditions required for reaction.[3]



$$C_6H_5CH{=}C\begin{matrix}COOH\\X\end{matrix}$$

IA, X = $COOC_2H_5$
IB, X = CN
IC, X = COOH

The kinetics of the decarboxylation of IA and IB were determined in pyridine solution at 111.0° by measuring the rate of evolution of carbon dioxide. Interpretation of the kinetics was based upon the previous finding[4] that acids such as IA and IB are quite weak acids in pyridine with dissociation constants of the order of 10^{-8} and ionization constants (for ionization to ion-pairs) of about 10^{-4} at 25°. Unimolecular decarboxylation of the un-ionized acid (HA) would be indicated if a plot of $\ln V_\infty/(V_\infty - V_t)$ *vs.* t yielded a straight line (V_∞ is the total volume of carbon dioxide evolved in the reaction and V_t is the volume of carbon dioxide at time, t). Unimolecular anionic decarboxylation, *i.e.*, decarboxylation *via* the carboxylate ion (A^-), would be indicated if the same plot were a curve which is markedly concave up.[2] The plots of $\ln V_\infty/V_\infty - V_t)$ *vs.* t for both IA and IB are *downward* bending curves and suggest the possibility that decarboxylation actually occurs by a second-order process.

Three equations which would correspond to second order kinetics are

$$\left.\begin{array}{lll}(1) & HA + HA & \rightarrow\\(2) & HA + A^- & \rightarrow\\(3) & HA + C_5H_5NH^+ & \rightarrow\end{array}\right\}\text{products}$$

The rate equation for the first process is

$$\frac{d[HA]}{[HA]^2} + \frac{K^{1/2}}{2}\frac{d[HA]}{[HA]^{5/2}} = -k\,dt \qquad (a)$$

where K is the *dissociation* constant of the acid and k is the second-order rate constant. Integration of this equation gives

$$\frac{1}{[HA]} + \frac{K^{1/2}}{3[HA]^{3/2}} = kt + c \qquad (b)$$

If K is of the order of 10^{-6} or 10^{-8}, the second term is very small. For [HA] values ranging from 0.2 to 0.01 a plot of $1/[HA]$ *vs.* t should yield a straight line, if process (1) is occurring.

The rate expression for reactions of type (2) or (3) is

$$\frac{d[HA]}{[HA]^{3/2}} - \frac{K^{1/2}d[HA]}{2[HA]^2} = -kK^{1/2}dt \qquad (c)$$

and the integrated form is

$$\frac{2}{[HA]^{1/2}} + \frac{K^{1/2}}{2[HA]} = kK^{1/2}t + c \qquad (d)$$

For reactions of type (2) or (3) if K is small ($\sim 10^{-6}$ or less) a plot of $1/[HA]^{1/2}$ *vs.* t should yield a straight line.

When a plot of $1/\sqrt{HA}$ *vs.* t is made for IA and IB good straight lines are obtained (Fig. 1). Plots of $1/[HA]$ *vs.* t, however, afford lines of marked *upward curvature*.

The data for the acid ester IA require special comment since the yield of carbon dioxide obtained in several runs was only *ca.* 85% of the theoretical. Titration with standard base of a sample of IA which had been heated in pyridine until decarboxylation was complete indicated the presence of *ca.* 15% of the amount of acid originally present.

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[1] Presented before the Division of Organic Chemistry at the 122nd Meeting of the American Chemical Society in Atlantic City, N. J., September, 1952, and taken in part from the B.S. thesis of G. Fraenkel.

[2] E. J. Corey, THIS JOURNAL, **74**, 5897 (1952).

[3] G. B. Kistiakowsky and W. R. Smith, *ibid.*, **58**, 2428 (1936).

[4] E. J. Corey, *ibid.*, **75**, 1163 (1952).

A reasonable explanation for these facts is that IA undergoes transesterification to yield ethyl benzylidenemalonate and benzylidenemalonic acid. The latter subsequently undergoes decarboxylation to cinnamic acid which is fairly stable in hot pyridine. While it may be only fortuitous that a straight line was obtained in the $1/\sqrt{HA}$ plot for IA, this evidence none the less indicates that the decarboxylation follows second-order kinetics. The competing reaction, transesterification, is probably second order and as a consequence it seems reasonable that the occurrence of this reaction to the extent of *ca.* 15% of total should not alter significantly the shape of the $1/\sqrt{HA}$ *vs.* *t* plot for reactions such as (2) and (3). It should be noted that the rate of decarboxylation of benzylidenemalonic acid is quite fast compared to the rate of transesterification (Fig. 1).

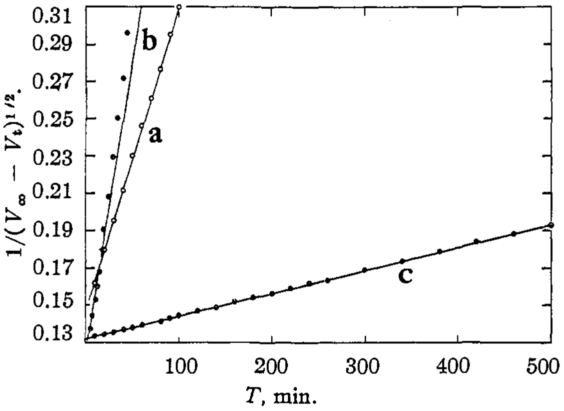


Fig. 1.—Plots of $1/\sqrt{HA}$ *vs.* *t* for decarboxylations in pyridine at 111.0°: a, IA (C_0 = 0.16 *M*); c, IB (C_0 = 0.192 *M*); b, IC (C_0 = 0.20 *M*).

Decarboxylation of IA and IB by a HA, A^- reaction (type 2) is excluded as a possibility by the fact that the rate of decarboxylation of one equivalent of these substances is decreased strongly by the presence of 0.5 equivalent or less of N-ethylpiperidine, a strong base. As the relative amount of N-ethylpiperidine is increased the rate of reaction is decreased still further. In the presence of added pyridinium nitrate as a source of pyridinium ion the decarboxylation of IA and that of IB are clearly first order (Fig. 2) and in each case the rate of reaction is increased. Both reactions give essentially quantitative yields of carbon dioxide which indicates that side reactions of IA such as transesterification have been rendered of negligible importance.

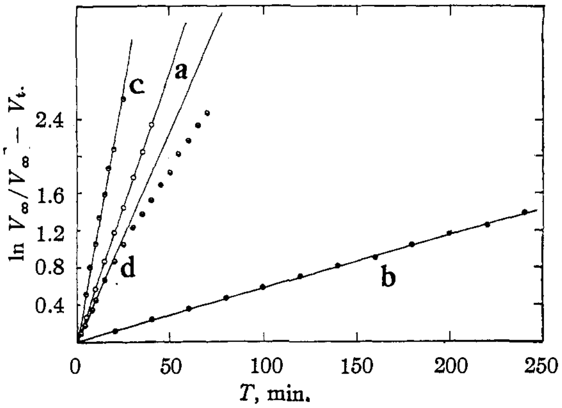
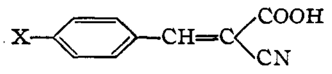


Fig. 2.—First-order plots for decarboxylations in pyridine: a, IA (C_0 = 0.149 *M*) plus 0.113 *M* pyridinium nitrate at 111.0°; b, IB (C_0 = 0.20 *M*) plus 0.20 *M* pyridinium nitrate at 111.0°; c, IC (C_0 = 0.20 *M*) plus 0.20 *M* pyridinium nitrate at 99.0°; d, IC (C_0 = 0.20 *M*) at 111.0°.

Various *p*-substituted benzylidenecyanoacetic acids II, X = NO_2, CH_3, OCH_3 exhibit



II

the same behavior as the unsubstituted acid; in pyridine plus added pyridinium nitrate first-order kinetics are observed and in pyridine alone there is a linear relationship between $1/\sqrt{HA}$ and *t*. The order of reactivity of the series of benzylidenecyanoacetic acids both with and without added pyridinium nitrate was found to be *p*-NO_2 > *p*-H > *p*-CH_2 > *p*-OCH_3 (Table I, Fig. 3).

TABLE I

RATE CONSTANTS FOR THE DECARBOXYLATION OF BENZYLIDENECYANOACETIC ACIDS IN PYRIDINE AT 111.0°

Acid	Rate constant, k_2[a] l./mole min. × 10^3
p-H	2.1
p-NO_2	7.3
p-CH_3	1.7
p-OCH_3	0.66

[a] Calculated from the expression $k = b\sqrt{V_\infty}/\sqrt{C_0}$, where *b* is the slope of the line obtained by plotting $[1/(V_\infty - V_t)]^{1/2}$ *vs.* *t* and C_0 is the initial concentration of the acid in moles/l. Half lives calculated for these reactions from the slope of the experimental plot ($t_{1/2} = 0.414/b\sqrt{V_\infty}$) agreed with the observed values to within 4%.

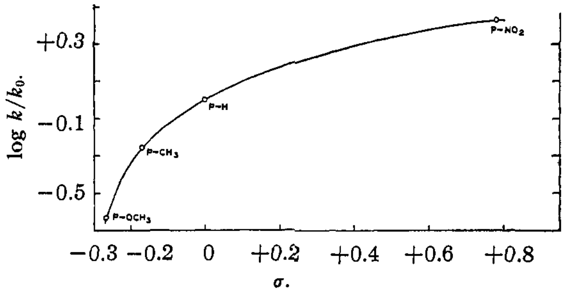
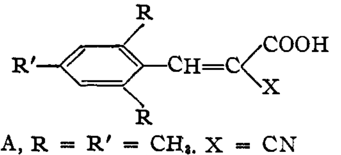


Fig. 3.—The relationship between log k/k_0 and Hammett's σ for the decarboxylation of *p*-substituted benzylidenecyanoacetic acids (C_0 = 0.20 *M*) plus pyridinium nitrate (0.20 *M*) in pyridine at 111.0°.

A significant ortho effect is evident in the decarboxylation of diortho-substituted acids of type III. The presence of two ortho substituents in IIIA, B, C results in strong inhibition of decarboxylation as is evidenced by the fact that the rate



IIIA, R = R′ = CH_3, X = CN
IIIB, R = Cl, R′ = H, X = CN
IIIC, R = Cl, R′ = H, X = $COOC_2H_5$

of reaction in pyridine at 111° is too slow to measure.[2]

The decarboxylation of benzylidenemalonic acid (IC) was also studied. In pyridine plus pyridinium nitrate the reaction exhibits good first-order kinetics. In pyridine alone, however, the plot of $1/\sqrt{\text{HA}}$ *vs. t* yields an upward bending curve, while the plot of $\ln (V_\infty/(V_\infty - V_t))$ *vs. t* gives a downward bending curve. The first-order plot showed somewhat more curvature than the second-order plot. Upward curvature for the $1/\sqrt{\text{HA}}$ *vs. t* plot would be expected for an HA, $C_5H_5NH^+$ reaction (type 3), since the product of decarboxylation is itself an acid and a source of pyridinium ion. In addition an HA, $C_5H_5NH^+$ reaction requires downward curvature of the first-order plot, since the starting diacid should furnish at least twice as much pyridinium ion as an equimolar concentration of cinnamic acid. Thus, it appears likely that the behavior of the diacid IC upon decarboxylation parallels that of IA and IB.

Thioacetic acid catalyzes greatly the decarboxylation of IA and IB in pyridine. The half-lives for decarboxylation of IA and IB in 20% thioacetic acid–80% pyridine at 111.0° are reduced to *ca.* one minute and three minutes, respectively, from values of 70 and 400 minutes in pure pyridine. Use of 20% acetic acid–80% pyridine at 111.0° as the decarboxylation medium in place of pyridine produces a relatively small change in rate of reaction.

Discussion

Significance of the Kinetic Data.—The kinetic data clearly indicate that the decarboxylation of IA and IB involves reaction between a molecule of un-ionized acid and a pyridinium ion in or previous to the rate-determining step, *i.e.*, an HA, $C_5H_5NH^+$ reaction (3). That a pyridinium ion is implicated in reaction is demonstrated in two ways: (1) the kinetics of decarboxylation in pyridine alone and in pyridine plus N-ethylpiperidine are consistent only with an HA, $C_5H_5NH^+$ reaction and (2) in the presence of added pyridinium nitrate, which serves to maintain the pyridinium ion concentration at a constant value,[5] the reaction shows strict first-order kinetics with no observable deviation.

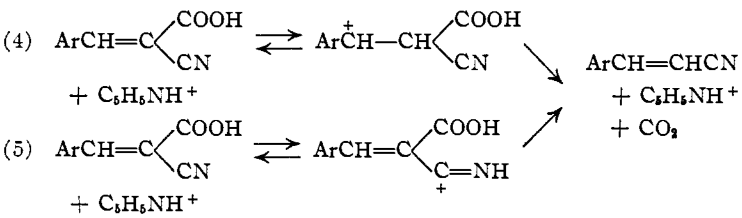
Unfortunately, since the difference in the integrated rate equations for the HA, $C_5H_5NH^+$ reaction (3) and the HA, HA reaction (1) is not a very great one, the present data on the decarboxylation of IA and IB in pyridine alone do not exclude the possibility that the latter process may occur to a considerable extent during decarboxylations in pure pyridine. Even if the rate constant for the HA, $C_5H_5NH^+$ reaction were a hundred times that for the HA, HA reaction a considerable amount of the second reaction would take place because of the relatively low concentration of pyridinium ions.

It is clear, however, that in the case of decarboxylation in the presence of added pyridinium nitrate the HA, HA reaction is of negligible importance, since strict first-order kinetics are observed. If the assumption is made that 3% of the total decarboxylation in the presence of pyridinium nitrate takes place by an HA, HA reaction (this value probably represents an upper limit) then the ratio k'/k'' in the rate equation (e)

$$dCO_2/dt = k'[HA][C_5H_5NH^+] + k''[HA]^2 \qquad (e)$$

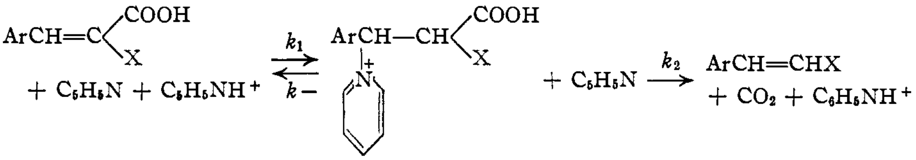
is at least 350. A value of 10^{-4} was assumed for the dissociation constant of pyridinium nitrate.[5]

The Effect of Para Substituents.—The relative rates of decarboxylation of *p*-substituted benzylidenecyanoacetic acids, I, (order of reactivity, $p\text{-NO}_2 > p\text{-H} > p\text{-CH}_3 > p\text{-OCH}_3$) provide information which is useful in determining the nature of the decarboxylation process. On the basis of the observed effect of *p*-substituents upon the rate of reaction two of the most obvious mechanisms by which HA and $C_5H_5NH^+$ can combine may be excluded.[6] They are



Both of these mechanisms require that the reactivity sequence of *p*-substituted acids be $p\text{-OCH}_3 > p\text{-CH}_3 > p\text{-H} > p\text{-NO}_2$ which is opposite to that actually observed. Furthermore these processes provide no reasonable explanation for the ortho effect which appears to operate in the case of ortho-disubstituted acids.

An interpretation of all of the above data which presents a reasonable and consistent picture of the decarboxylation process may be arrived at by assuming that a solvent molecule is involved in the reaction of HA with pyridinium ion. The termolecular combination is assumed to proceed in a concerted fashion to give the intermediate IV. Decarboxylative elimination[7] of IV would then lead to the products. This interpretation is in accord with the kinetic data regardless of the relative magnitude of the rate constants k_1, k_{-1} and k_2, as long as a steady-state concentration of IV is estab-



IV

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[5] This value is much less than that which corresponds to the amount of pyridinium nitrate added. The dissociation constant of pyridinium nitrate in dilute solutions has been found to be $0.5 \cdot (10)^{-4}$ at 25°, D. S. Burgess and C. A. Kraus, THIS JOURNAL, **70**, 706 (1948).

[6] These are analogous to the mechanisms proposed by W. S. Johnson and W. E. Heinz, THIS JOURNAL, **71**, 2913 (1949), for the acid-catalyzed decarboxylation of certain cinnamic acids.

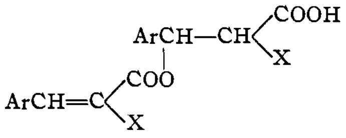
[7] *Cf.* ref. 3; *cf.* E. Grovenstein, Jr., and D. E. Lee, Abstracts of the 122nd Meeting of the American Chemical Society. Further evidence for the occurrence of decarboxylative elimination has been obtained by D. Y. Curtin and B. J. Luberoff, S. J. Cristol and E. R. Trumbull (unpublished work).

lished rapidly. The manifestation of ortho effect in the decarboxylation is also consistent with this mechanism.

The observed effect of para substituents also fits the above representation of the decarboxylation process. The substitution of *p*-OCH_3 for *p*-H or *p*-NO_2 would decrease strongly the value of k_1 and would increase the values of k_{-1} and k_2. The increase in k_{-1} would be considerably greater than that in k_2.[8] If the first step is fast and reversible and fast in comparison with the second step, substitution of *p*-OCH_3 for *p*-H should decrease to the over-all rate of reaction. If the first step is reversible and slow compared to the second step replacement of *p*-H by *p*-OCH_3 should again have an inhibitory effect upon decarboxylation. Thus, either assumption leads to agreement between the expected and the experimentally observed sequence of reactivity.

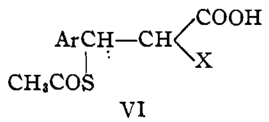
To the extent that the second term in the rate equation (e) is important, the proton required for the first step of the above process might be furnished by H^+A^- ion-pairs or HA instead of dissociated pyridinium ion. In the latter event the decarboxylation would be general acid-catalyzed instead of specific pyridinium-ion catalyzed.

It is also possible that the HA, HA reaction could afford (by a synchronous addition of HA to the α,β-double bond of a second HA molecule) the intermediate V which would subsequently undergo decarboxylative elimination.



V

Catalysis of the Reaction.—The strong catalytic influence of thioacetic acid provides further evidence for the operation of a concerted 1,2-addition, 1,2-elimination mechanism for the decarboxylation of benzylidenemalonic acid derivatives. Since thioacetic acid is highly nucleophilic, it would be expected to perform the function of pyridine in the uncatalyzed reaction with great efficiency and, hence, increase the rate of decarboxylation. It seems quite likely that the intermediate involved in the catalyzed reaction is VI



$$\underset{\displaystyle CH_3COS}{ArCH}\text{—}CH\begin{cases}COOH\\X\end{cases}$$

VI

Quantitative Effect of Para Substituents upon Rate.—An interesting feature of the observed effect of para substituents upon rate of decarboxylation is the absence of the linear relationship between log k/k_0 and Hammett's substituent constants (σ) which is predicted by the Hammett equation, log $k/k_0 = \rho\sigma$.[9] Shown in Fig. 3 is the log k/k_0 *vs.* σ plot for decarboxylations of *p*-substituted benzylidenecyanoacetic acids which were carried out in the presence of added pyridinium ion to ensure the operation of a single mechanism. The log k/k_0 *vs.* σ plot for the reaction in pyridine alone has the same shape. It should be noted that the experimental points fall on a smooth curve which is concave downward and that the degree of curvature increases sharply as more negative values of σ are approached. Apparently ρ is a function of σ and is more sensitive to negative than to positive values of σ.

The interdependence of ρ and σ for a number of reactions has been pointed out previously.[10,11] Important deviations from the behavior predicted by the Hammett equation appear to be associated with reactions of two types: (1) those in which the degree of resonance stabilization in the transition state is strongly dependent upon the substituent present[10] and (2) reactions in which there is enhanced resonance in the ground state between certain substituents and the reacting group. Examples of reactions of the first type are aromatic substitution,[12] hydrolysis of benzyl halides[12] and the pinacol rearrangement.[11] Transformations of *p*-nitrophenol or *p*-nitroaniline, on the other hand, represent reactions of the second type as is evidenced by the need for a special σ for the nitro group in these compounds.

Divergence from the Hammett relationship in the decarboxylation of *p*-substituted benzylidenecyanoacetic acids is apparently due to enhanced resonance interaction between certain substituents (OCH_3, CH_3) and the groups affected directly by the reaction (CN, $COOC_2H_5$, COOH).

Experimental[13]

Reagents.—Pyridine was reagent grade, stored over barium oxide.

Pyridinium nitrate was made from reagent pyridine and concentrated nitric acid and was purified by recrystallization from *t*-butyl alcohol.

N-Ethylpiperidine was the Eastman Kodak Co. product for penicillin G determination.

Thioacetic acid, which was kindly furnished by Dr. C. S. Marvel, was purified by distillation, b.p. 92.8–93.0°.

Ethyl hydrogen benzylidenemalonate was prepared essentially by the method of Reinicke[14] and purified by several recrystallizations from ligroin, m.p. 85–90°.

Benzylidenecyanoacetic acid, m.p. 180°,[15] *p*-methoxybenzylidenecyanoacetic acid, m.p. 229–230°,[16] and *p*-methylbenzylidenecyanoacetic acid, m.p. 209–210°,[16] were prepared by the method of Lapworth and McRae.

p-Nitrobenzylidenecyanoacetic acid was prepared by heating a solution of 5.0 g. (0.031 mole) of *p*-nitrobenzalde-

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[8] The basis for this argument is that conjugation of *p*-methoxyphenyl with COOH and or X results in much more resonance stabilization than does conjugation of phenyl or *p*-nitrophenyl. In addition there will be more resonance stabilization when *p*-methoxyphenyl is conjugated with X and COOH than with X alone. Since the transition state for each of the three reactions considered has some double bond character (intermediate in degree between starting material and product), the effect of substituents upon the rate constants, k_1, k_{-1} and k_2 for these reactions follows immediately.

[9] L. P. Hammett, THIS JOURNAL, **59**, 96 (1937); "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, Chapter VII.

[10] C. G. Swain and W. P. Langsdorf, Jr., THIS JOURNAL, **73**, 2813 (1951).

[11] W. E. McEwen and N. B. Mehta, *ibid.*, **74**, 526 (1952).

[12] The data for aromatic substitution given in ref. 10, p. 199, fit a curve more satisfactorily than a straight line. The same is true of the data for hydrolysis of benzyl halides [G. M. Bennett and B. Jones, *J. Chem. Soc.*, 1815 (1935)].

[13] We are indebted to Miss Emily Davis and Mrs. Katherine Pih for the microanalyses.

[14] G. Reinicke, *Ann.*, **341**, 80 (1905).

[15] A. Lapworth and J. A. McRae, *J. Chem. Soc.*, **121**, 1699 (1922).

[16] E. Fiquet, *Ann. chim.*, [6] **29**, 472 (1893).

hyde, 5.0 g. (0.059 mole) of cyanoacetic acid and 0.1 g. of piperidine acetate in 75 ml. of acetonitrile at reflux for four hours. The reaction was carried out in a flask surmounted by a Soxhlet extractor charged with barium oxide. The acetonitrile was removed by distillation under reduced pressure and the residue was extracted with sodium bicarbonate solution. Acidification and recrystallization of the resulting solid from ethanol–water afforded 2.9 g. (38.7%) of pure acid, m.p. 208–209° (lit.[16] 208°).

Mesitylidenecyanoacetic acid was prepared by heating a mixture of 4.45 g. (0.30 mole) of mesitaldehyde, 5.1 g. (0.60 mole) of cyanoacetic acid, 1.5 ml. of piperidine and 15 ml. of dry pyridine at 70°. The reaction mixture was poured into dilute hydrochloric acid and the solid which precipitated was purified by extraction with sodium bicarbonate, acidification and recrystallization of the material so obtained from ethanol. After two recrystallizations the product had m.p. 193.2–195.5°. A third recrystallization failed to change the m.p.

Anal. Calcd. for $C_{13}H_{13}NO_2$: C, 72.54; H, 6.09; N, 6.51. Found: C, 72.30; H, 6.24; N, 6.56.

Benzylidenemalonic acid, m.p. 195–196, was prepared by the method of Claisen and Crismer.[17]

Reaction Products.—Ethyl hydrogen benzylidenemalonate (3.0 g.) was heated in pyridine solution under kinetic conditions until decarboxylation was complete and the product isolated by pouring the reaction mixture into cold dilute hydrochloric acid, extracting with ether and distilling the liquid obtained in the extract through a semi-micro column. There was obtained 1.85 g. (77%) of ethyl cinnamate, b.p. 98–99° (0.5 mm.), $n^{20}D$ 1.5597. It is apparent from the refractive index that the product obtained was the pure *trans* isomer (lit.[18] $n^{20}D$ 1.5598 for ethyl *trans*-cinnamate, $n^{20}D$ 1.545 for ethyl *cis*-cinnamate).

By the same procedure there was obtained from 5.0 g. of benzylidenecyanoacetic acid 3.35 g. (89.9%) of a mixture of 65% *trans*- and 35% *cis*-cinnamonitrile, b.p. 86–89° (0.7 mm.), $n^{20}D$ 1.5968 (lit.[3] $n^{20}D$ 1.6032 for *trans*-cinnamonitrile and $n^{20}D$ 1.5849 for *cis*-cinnamonitrile).

The same products of decarboxylation were isolated from runs in which thioacetic acid was used as catalyst. In the case of the cyano acid IB the yield of nitrile obtained was low (45%) due to reaction of the nitrile once formed with thioacetic acid.

Kinetic Data.—The apparatus and procedure used for measuring rates of decarboxylation have been described previously.[2] The only modification made in the present work was the agitation of the reaction mixture by vigorous and constant *vibration* rather than by shaking.

URBANA, ILLINOIS

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[17] L. Claisen and L. Crismer, *Ann.*, **218**, 131 (1883).

[18] K. v. Auwers and E. Schmellenkamp, *Ber.*, **54**, 631 (1921).

[CONTRIBUTION FROM THE NOYES CHEMICAL LABORATORY, UNIVERSITY OF ILLINOIS]

The Ionization and Dissociation of Certain Weakly Acidic Substances in Pyridine

BY ELIAS J. COREY

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A study has been made of the behavior of certain phenols and carboxylic acids in pyridine. 2,6-Dinitro-3,4-xylenol (I) and 2,5-dinitrophenol (II) are partially ionized in pyridine solution in contrast to the extremes, *p*-nitrophenol (not appreciably ionized) and 2,4,6-trinitrophenol (essentially completely ionized). Solutions of I and II contain as ionic species both dissociated ions and pyridinium-phenoxide ion-pairs. Carboxylic acids have been shown to be only slightly ionized and dissociated in pyridine.

The nature and relative amounts of the species which result when weakly acidic compounds—phenols, carboxylic acids, etc., are dissolved in pyridine have not been determined previously despite the frequency with which reactions involving solutions of acids in pyridine are carried out.[1]

In connection with an investigation of decarboxylation reactions in pyridine[2] we have determined in an approximate way the fate of several acidic substances in this solvent.

Results and Discussion

From the data in Table I it is apparent that *p*-nitrophenol is not appreciably ionized in pyridine and that 2,4,6-trinitrophenol is completely ionized. The dinitrophenols are intermediate and the acidities of these substances in pyridine parallel qualitatively the pK_a values in water.

The behavior of 2,6-dinitro-3,4-xylenol (I) and that of 2,5-dinitrophenol (II) were investigated in detail since I and II are moderately ionized in pyridine and are potential indicators for the study of colorless acids.

Spectra of I and II.—Solutions of (I) in pyridine exhibit bands at 360 and 430 mμ (Fig. 1). The spectrum of I in absolute ethanol is characterized by a single band at 347 mμ (log ε 3.6) which corresponds to the 360 mμ band observed in pyridine. In ethanol containing a small amount of the strong base N-ethylpiperidine absorption occurs at 425 mμ (log ε 3.79) only, and there is no band at 347 mμ. The absorption manifested in pyridine of I at 360 mμ is, therefore, attributable to the phenol (the shift $\lambda_{max}^{ethanol}$ 347 mμ to $\lambda_{max}^{pyridine}$ 360 mμ probably being the result of hydrogen bonding of the phenol with pyridine) and the absorption at 430 mμ to the phenoxide ion.

TABLE I

Compound	pK_a (25°)	Solutions in pyridine, color	Solutions in pyridine + N-ethyl piperidine Color	Change in color from pure pyridine
p-Nitrophenol	9.19	Almost colorless	Yellow	++
2,5-Dinitrophenol	5.15	Pale yellow	Yellow-orange	+++
2,6-Dinitro-3,4-xylenol	4.92[a]	Yellow	Deep yellow	+++
2,4-Dinitrophenol	4.00[b]	Deep yellow	Deep yellow	+
2,4,6-Trinitrophenol	..	Deep yellow	Deep yellow	None

[a] In 50% ethanol. *pK* 4.19 in 50% ethanol.

The spectrum of II in pyridine is analogous to that of I. Phenol absorption occurs at 360 mμ as in the case of I and phenoxide absorption is manifested at 465 mμ (Fig. 2).

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[1] In a brief communication L. E. Krohn and V. K. LaMer, THIS JOURNAL, **53**, 3563 (1931), have reported a study of the change in light absorption of indicators (*e.g.*, thymol blue) in pyridine with varying concentrations of trichloroacetic acid and diethylamine.

[2] E. J. Corey, *ibid.*, **75**, 1163 (1953).