Purification of Substances to be Reduced.—Commercial C. P. products were obtained and further purified. Thiophene-free benzene, toluene, ethylbenzene, xylene, mesitylene, cymene and diphenylmethane were refluxed for three hours over sodium and then distilled. Only the constant boiling fraction was used for the reduction. Triphenylmethane, α,α-diphenylethane, dibenzyl, phenylacetic acid and β-phenylpropionic acid were crystallized alternately from alcohol and glacial acetic acid. After several recrystallizations the compounds were used for reduction.

Method of Isolation of Products.—After filtration and evaporation of the solvent, the reduced compounds, with the exception of the two acids, were treated several times with a mixture of 10 cc. of concd. nitric acid and 25 cc. of sulfuric acid at 0–5° for five to ten minutes. After each treatment the compound was washed with water. When the acid mixture no longer turned yellow, the compound was considered to be pure and was washed, dried and distilled.

The cyclohexylacetic acid and the β-cyclohexylpropionic acid were merely distilled from the acetic acid in order to purify them.

Summary

It has been found that platinum-oxide platinum black is an excellent catalyst in the reduction with hydrogen of various types of aromatic hydrocarbons in glacial acetic acid as a solvent.

URBANA, ILLINOIS

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

THE EFFECT OF THE CYANIDE GROUP ON THE BASICITY OF ALIPHATIC AMINES AS DETERMINED IN WATER AND ALCOHOL SOLUTIONS

BY T. D. STEWART AND VIVIEN COOK

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Stewart and Aston[1] studied the influence of alkoxyl substitution upon the basicity of tertiary aliphatic amines, using derivatives of the type $ROCH_2N(C_2H_5)_2$. They found that methoxy, ethoxy and *iso*butoxy substitution lowered the base strength and gave to the basic dissociation constants of the amino ethers the values 3.6×10^{-9}, 1.8×10^{-8} and 4×10^{-7}, respectively. The comparisons were made colorimetrically in alcohol, using methylaniline and benzylamine as standards. By assuming that the relative base strengths in water and alcohol solution would be the same, the above figures were assigned as for an aqueous solution. The work of Goldschmidt and Mathiesen[2] indicates that in general amines are stronger bases in alcohol than in water, but with notable exceptions, so that comparisons of basicity in alcohol cannot be considered to offer a similar comparison in water. The above values, then, must be considered not as absolute but as relative to themselves and to methylaniline in alcohol solution only. It is, nevertheless, apparent that alkoxyl sub-

[1] Stewart and Aston, THIS JOURNAL, **48**, 1642 (1926).

[2] Goldschmidt and Mathiesen, *Z. physik. Chem.*, **119**, 439 (1926).

stitution upon the α-carbon of an amine causes a marked decrease in the tendency of the nitrogen atom to attach a proton, and particularly interesting to observe that the effects of different alkoxyl groups are considerably different in magnitude.

The present paper deals with the effect upon the base strength of substituting a cyanide group for a hydrogen atom upon a carbon in the α-position to the nitrogen in methyldiethylamine, triethylamine, benzyldiethylamine and *iso*propyldiethylamine. The comparisons are made both in alcohol and water solutions in the first and third cases, but only in alcohol in the other two because of the rapid decomposition of these bases in aqueous acid.

Experimental Results

The electrometric titrations of α-diethylamino-acetonitrile and of α-diethylaminophenylacetonitrile were carried out in water as described in the Experimental Part. The results indicate basic dissociation constants,

TABLE I

THE ELECTROMETRIC TITRATION OF SOLUTIONS OF α-DIETHYLAMINO-ACETONITRILE IN EXCESS OF 0.1 *N* HYDROCHLORIC ACID WITH 0.1 *N* SODIUM HYDROXIDE

Part A

Expt.	Sample, equiv.	Total acid, cc. of 0.1 *N*	Temp., °C.	Concn. of amino salt at the true end-point, equiv./L.
1	0.003976	45.49	22.8	0.07777
2	.003597	45.56	22.6	.06405
3	.004135	45.13	22.5	.0837

Part B

Expt.	Percentage titration of salt by NaOH	Voltage, from plot	$H^+ \times 10^{-3}$, equiv./L.	$K_B \times 10^{10}$
1	0	0.481	3.353	.66
	25	.535	.399	.75
	50	.561	.1438	.69
2	0	.489	2.401	1.07
	25	.545	.2647	1.13
	50	.573	.0886	1.13
3	0	.488	2.462	1.31
	25	.550	.2218	1.35
	50	.575	.0810	1.23
				Av. 1.03

TABLE II

THE ELECTROMETRIC TITRATION OF SOLUTIONS OF α-DIETHYLAMINOPHENYLACETONITRILE IN EXCESS OF 0.1 *N* HYDROCHLORIC ACID WITH 0.1 *N* SODIUM HYDROXIDE

Part A

Expt.	Sample, equiv.	Total acid, cc. of 0.1 *N*	Elapsed time before titration, min.	Weak base found, equiv.	Temp., °C.
1	0.00275	36.17	35	0.00240	22.5
2	.002785	36.03	25	.00238	22.8
3	.002875	36.05	120	.001225	22.5

TABLE II (*Concluded*)

Part B

Approximate Solubility of Free Base, 0.00176 Equiv./L.

Expt.	0.1 *N* NaOH, cc.	Total vol. cc.	Salt present (approx.), equiv./l.	Voltage	(H^+), equiv./l.	Cation, equiv./l.	$K_B \times 10^{11}$
1	9.58	45.75	0.0496	0.4477	0.01234	0.0373	1.72
1	16.05	52.22	.0328	.4559	.00895	.0239	1.52
1	25.38	61.55	.0126	.4812	.00331	.00932	1.60
2	8.96	44.99	.0515	.4491	.01175	.0398	1.92
2	15.98	52.01	.0308	.4618	.00714	.0237	1.89
2	24.09	60.12	.0132	.4811	.00335	.0099	1.68
						Av.	1.72

K_B, of 1.0×10^{-10} and 1.7×10^{-11}, respectively, and are summarized in Tables I and II and Fig. 1. The method of calculation is described in the later section.

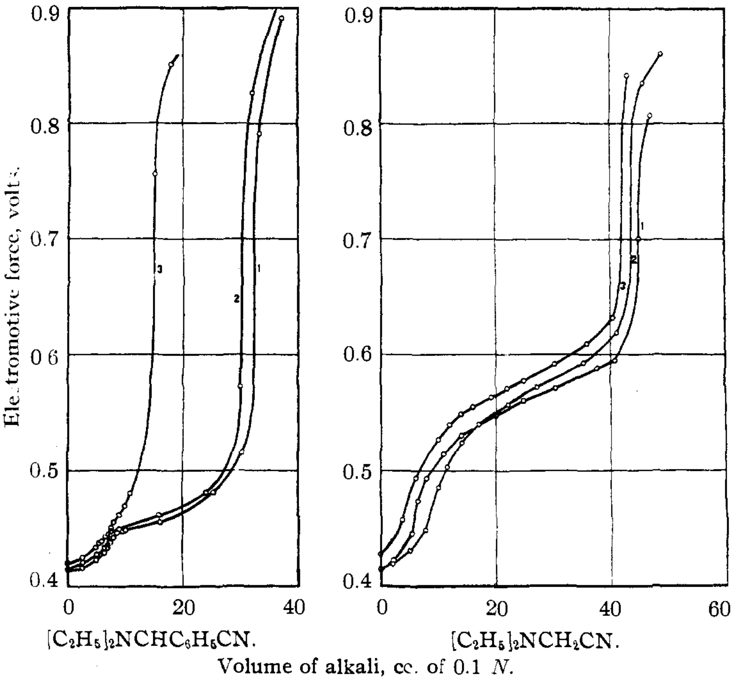

Fig. 1.—Titration curves of salts of α-aminonitriles in water.

In each of the three titrations of α-diethylamino-acetonitrile (Table I, Part B) consistent values of K_B are obtained, but the three do not agree well among themselves, perhaps due to constant small errors in the electrode condition. Still greater difficulty was experienced in attempting to get reproducible results when the base was titrated directly with acid, probably due to poisoning of the electrodes in the alkaline solution. The numerical average given may be considered accurate within 30%.

Since the α-diethylaminophenylacetonitrile is sparingly soluble in water and upon titration of its acid solution started to precipitate before the excess acid was neutralized, the inflection point of the titration curve did not signify the point of equivalency of acid and base present. Moreover, the substance decomposes more or less rapidly in dilute acid, presumably according to the equation

$$C_6H_5CHCNN(C_2H_5)_2 + H_2O \longrightarrow C_6H_5CHOHCN + (C_2H_5)_2NH$$

The net result is the transformation of a weak base to a stronger base, so that the concentration of the salt present (Table II, Part B, Col. 4) was corrected for this reaction (see Experimental Part). Curve III, Fig. 1, for this substance was for an acid solution which had stood for two hours before titration. The weak base had almost disappeared.

Table III summarizes the results of the colorimetric comparison in alcohol at room temperature. An alcoholic hydrogen chloride solution of known strength, containing the indicator in question, was added to a weighed sample of the pure base. Details of the color comparison are given later and only the average of several determinations is given in Table III. The ratios of the various basic dissociation constants are taken from the expression

$$K = \frac{(\text{Cation})}{(\text{Base})}(\text{OH}^-) = \frac{(\text{Salt})}{(\text{Base})}(\text{OH}^-)$$

wherein it is assumed that the acid is completely dissociated and the concentration of hydrogen ion is negligible in comparison to the concentration of cation, that is, all of the acid is neutralized. At equal hydrogen ion (or hydroxyl ion) concentration, as shown by the indicator, the ratios of the constants are the ratios of the respective salt/base ratios.

TABLE III

THE PERCENTAGE TITRATIONS OF SOME α-DIETHYLAMINO NITRILES, *m*-CHLOROANILINE AND ANILINE WITH ALCOHOLIC HYDROGEN CHLORIDE AT CORRESPONDING INDICATOR COLORS, AND THE RELATIVE STRENGTHS OF THE RESPECTIVE BASES IN ABSOLUTE ALCOHOL SOLUTIONS

Part A

	Methyl red and propyl red				Methyl red	
m-Chloro aniline (K_C)	α-Diethylamino-propionitrile (K_P)		α-Diethylamino-acetonitrile (K_A)		α-Diethylamino-phenylacetonitrile (K_D)	
%	%	K_P/K_C	%	K_A/K_C	%	K_C/K_D
40	58	2.07	56	1.91	11	5.4
50	66	1.94	64	1.78	13	6.7
60	72	1.72	72	1.72	16	7.8
70	80	1.72	80	1.72	18	10.6
80	87.5	1.75	87.5	1.75	25	12.0
90	..	..	..	..	41	12.9
		Av. 1.8		Av. 1.8		Av. 9.2

Table III (*Concluded*)

	Part B		Part C		
	Propyl red		Methyl red and thymol blue		
m-Chloro-aniline (K_C) %	α-Diethylamino-phenylacetonitrile (K_D) %	K_C/K_D	α-Diethylamino-acetonitrile (K_A) %	Aniline (K_N) %	K_N/K_A
50	14.5	5.9	40	92	17.3
60	17.0	7.3	50	95	19
70	21.5	8.5	60	97.2	..
80	31	8.9			
		Av. 7.7			Av. 18

The trend in the ratio of the constants in the case of the phenylated derivative may be due to experimental error, although there is no suggestion in the actual data of an uncertainty of this magnitude.

The fact that the trend is less marked in the experiment using Propyl Red than it is in those with Methyl Red, suggests some specific action of the base upon the indicator, or it may involve the assumptions used in the comparison of the constants.[3] It is doubtful whether the numerical average should be taken rather than an extrapolation toward higher fractional neutralization. Since the averages agree in the two series better than the extrapolations, they are taken as more nearly approaching the correct value and a final value of 8.5 assumed for the ratio of the constants of *m*-chloro-aniline to α-diethylaminophenylacetonitrile in alcohol. It is believed that this is correct within 20%.

Table IV gives a résumé of the results. The values for aniline and *m*-chloro-aniline in water are taken from Landolt-Börnstein's "Physikalische-Tabellen." K_B is the basic dissociation constant in water, K_A, the acidic dissociation constant of the cation, obtained for water solution from the relation $K_A = K_W/K_B$, where $K_W = 1.0 \times 10^{-14}$. K_A in alcohol solution is taken in the case of aniline, methylaniline and *m*-chloro-aniline from the work of Goldschmidt and Mathiesen,[2] and in the other cases calculated from these and from the above determined ratios.[4]

[3] Goldschmidt and Mathiesen, ref. 2, give the dissociation constant of *m*-chloro-anilonium ion ($ClC_6H_4NH_3^+$) in alcohol as 6.1×10^{-5}, and of anilonium ion ($C_6H_5NH_3^+$) as 2×10^{-6}. A solution of aniline hydrochloride which is 0.05 *M* would then contain about 3×10^{-4} *N* hydrogen ion, since hydrogen chloride is a strong acid, even in alcohol. A solution of chloro-aniline and its hydrochloride, in equivalent amounts, would contain about 6×10^{-5} *N* hydrogen ion, and at the same point our data show aniline to be about 98% titrated, whereas calculated from the above figures it should be 97%. We are, then, in substantial agreement and are apparently justified in the assumption of our simplified formula.

[4] In addition to the results of Table IV, we have made measurements on a base prepared from acetone, and which is presumably α-diethylamino-*iso*butyronitrile. It had the correct equivalent weight and is a new preparation so far as we know. The yield was small and the preparation has not been duplicated, hence we have not detailed the results at this time. In alcohol solution the substance titrated precisely as does aniline. If this is true, it means that whereas the introduction of one methyl

Table IV

Basic and Acidic Dissociation Constants of Some Amine Bases and Salts in Water and in Alcohol

	Water solution			Alcohol solution	
Base	K_B	K_A	Ratios of the basic constants	K_A	Ratios of the constants
α-Diethylaminophenylacetonitrile	1.7×10^{-11}	5.9×10^{-4}	...	5.2×10^{-4}	....
m-Chloro-aniline	3.4×10^{-11}	2.9×10^{-4}	2	6.1×10^{-5}	8.5
α-Diethylamino-acetonitrile	1.0×10^{-10}	1.0×10^{-4}	2.9	3.4×10^{-5}	1.7
α-Diethylaminopropionitrile	....	....	...	3.4×10^{-5}	1.0
Aniline	4.6×10^{-10}	2.2×10^{-5}	4.6	2×10^{-6}	17
Methyl aniline	....	....	...	1.3×10^{-5}	0.154
α-Diethylamino-methyl methyl ether	....	....	...	9.2×10^{-7}	14.1
α-Diethylamino-methyl ethyl ether	....	....	...	1.84×10^{-7}	5.0
α-Diethylamino-methyl *iso*butyl ether	....	....	...	8.27×10^{-9}	22.2
Methyldiethylamine	2.7×10^{-4}	3.7×10^{-11}	...	....	....

Discussion of Results

We may record in Table V for reference[5] the dissociation constants in water of the parent compounds of the cyano and alkoxyl amines and the data for a similar series of substitutions in acetic acid.

Table V

Basic and Acidic Dissociation Constants in Water

Bases	K_B	Acids	K_A
Methyldiethylamine	2.7×10^{-4}	Acetic	1.8×10^{-5}
Diethylbenzylamine	3.6×10^{-5}	Phenylacetic	5.3×10^{-5}
Triethylamine	6.4×10^{-4}	Cyanoacetic	3.7×10^{-3}
		Methoxyacetic	3.3×10^{-4}
		Ethoxyacetic	2.3×10^{-4}

From a comparison of Tables IV and V, using either the ratios of the dissociation constants or the logarithms of such ratios, it is seen that the cyano and alkoxyl groups substituted in amines reduce the basicity to a much greater extent than they increase the strength of acetic acid when substituted upon the α-carbon atom. The two systems are not strictly comparable, since that of the amines is a 1,3-substitution from the affected atom (nitrogen) whereas in acids the substitution is 1,4. The effect of substituting a methyl or phenyl group in a cyano-amine is the same as for the same substitution in the simple amine. The logarithm of the dissociation constant of methyldiethylamine (in water) is decreased

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group into the acetonitrile derivative causes no change in the base strength, a second methyl group produces a 17–18 fold increase in base strength. Mr. A. J. Alstrom, to whom we are indebted for this provisional result, is repeating the preparation.

[5] Landolt-Börnstein, "Physikalische-Chemische Tabellen," Berlin, 1923.

about 0.85 by phenyl substitution in the α-position, and about 6.4 by cyano substitution. Similarly substitution of hydrogen by phenyl in diethylamino-acetonitrile decreases the logarithm of the dissociation constant by 0.77. The increase of the logarithm of the dissociation constant of acetic acid is 0.47 for phenyl substitution, 2.32 for cyano substitution and 1.27 for methoxy substitution. Methoxy substitution in the amines likewise produces a smaller effect than cyano substitution, at least in alcohol solution. From the data now available, it would not be safe to assume that the effect of change of solvent from water to alcohol would produce in the amino nitriles and in the amino ethers the same change or even a similar change in the basic dissociation constants of the respective bases. It is, however, probable that the two classes of amines closely maintain in water the same relative basicities as in alcohol.

It so happens that the bases studied here maintain the same order in base strength in alcohol and in water, although the ratios of the dissociation constants differ widely in the two solvents. The series of bases studied by Goldschmidt and Mathiesen[2] differed even in order. The reasons for this will not be discussed fully at this time. It is evident, however, that in the expression

$$K_B = \frac{(\text{Cation})}{(\text{Base})}(\text{OH}^-)$$

the parentheses denote the activities of the respective substances. Since normally it is the concentration and not the activity that is measured, and since in the present work solutions of equal hydrogen ion activity are compared, we may substitute for the above activities the product of the concentration by a factor γ and write for two different bases in a given solvent

$$\frac{K'_B}{K''_B} = \frac{\dfrac{\gamma_1 C'_{\text{Cation}}}{\gamma_2 C'_{\text{Base}}}}{\dfrac{\gamma_3 C''_{\text{Cation}}}{\gamma_4 C''_{\text{Base}}}}$$

A similar expression could be written for a second solvent. It is then seen that the ratio of the constants, K'_B and K''_B, as calculated *only* from the concentrations, can be the same in two solvents only in case the *ratio* of the four γ factors remains constant with change of solvent. The γ factors involve deviations from the laws of perfect solutions and may, in addition, not be constant with respect to changes in concentration. It is not surprising that numerical constancy of the above ratios is not found. The question is further complicated by the assumption of infinitely dilute solutions in each solvent as the standard state of a given molecular species in that solvent. There would be a difference in the energy contents of the hypothetical molal solution of a given solute in two different solvents which should be included in a strict comparison of the basic dissociation constants of a substance in those solvents.

Finally, if a base is composed of two or more tautomers, the net base strength is determined in part by the relative amounts of the tautomers present. A change of solvent may change these relative proportions and in a manner unrelated to the relative base strengths of the individual tautomers. Then the comparison of a base which may tautomerize with one which may not, in different solvents, might lead to results which would be difficult of interpretation in terms of their respective structures.

Experimental Part

Preparation of Materials.—The following amino nitriles were prepared using the method of Knoevenagel and Mercklin.[6]

α-Diethylamino-acetonitrile from formaldehyde, sodium bisulfite, diethylamine and potassium cyanide; colorless oil, soluble in water; b. p. 35–40° (6 mm.) and 62–63° (14 mm.); yield, 85% based on diethylamine.

α-Diethylaminopropionitrile from acetaldehyde and other reagents as above; colorless oil, soluble in water; rapidly hydrolyzes in aqueous acid; b. p. 47–49° (7 mm.) and 67–68° (17 mm.); yield, 65% based on diethylamine.

α-Diethylaminophenylacetonitrile from benzaldehyde and other reagents as above; slightly soluble in water, soluble in dilute acid, usually light yellow or greenish yellow after distillation; b. p. 112° (7 mm.) and 130–131° (11 mm.); yield, 83% based on diethylamine. The base decomposes more or less rapidly in acid solution, most rapidly in dilute acid, but even then the reaction requires some hours. It may be dissolved in acid and be reprecipitated with base unchanged. The equivalent weight may be accurately determined by dissolving a weighed sample in a known excess of 0.1 *N* acid and back-titrating immediately with 0.1 *N* base, using the difference between the precipitation end-point and the phenolphthalein end-point as a measure of the base present. The latter end-point should coincide with that of an equal amount of acid to which no base has been added.

α-Diethylamino-*iso*butyronitrile, from acetone and the other reagents as above. The yield was very poor and the preparation could not be duplicated after several trials; b. p. 75–77° (23 mm.); equivalent weight to methyl orange titration, 140 g.; colorless oil, soluble in water, rapidly hydrolyzed in aqueous acid.

Aniline.—A redistilled sample of a commercial preparation was used. It was colorless and remained so, in the dark, for several days.

***m*-Chloro-aniline.**—A redistilled sample of Kahlbaum's preparation was used. It distilled within a range of one degree and was colorless.

Alcoholic Hydrogen Chloride was prepared by generating hydrogen chloride from pure concentrated hydrochloric and sulfuric acids, drying the gas by passage through sulfuric acid and collecting it in cooled commercial absolute alcohol of good grade. It was then diluted to the desired normality with the same alcohol and the indicator added to the stock solution in the following concentrations, respectively: methyl red, 0.0083 g. per liter; propyl red, 0.012 g. per liter; thymol blue (acid), 10 cc. of a saturated solution in alcohol, per liter. The normality of the alcoholic acid was determined by diluting a measured volume of the solution with water and titrating with standard alkali, using the contained indicator. One series of experiments was made using 0.1 *N* acid and several using about 0.05 *N* acid.

Solubility of α-Diethylaminophenylacetonitrile in Water.—This was determined at 20° as follows. A 0.4580-g. sample (0.002349 equiv.) was dissolved in 30.0 cc. of 0.1 *N*

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[6] Knoevenagel and Mercklin, *Ber.*, **37**, 4087 (1904).

acid and 2.00cc. portions of this solution were added to known volumes of water. Excess akali in known volume was then added to the mixture and the presence or absence of a turbidity noted. In one experiment, a final volume of 97.0 cc. allowed no turbidity; in another a final volume of 91 cc. allowed a very slight turbidity which cleared on the addition of 1.0 cc. of water. Since a high degree of accuracy was not desired, this volume of 92.0 cc. was taken as accurate and represents a solubility of 0.00176 ± 2% equivalents per liter. At 30° the solubility is about 0.00185 ± 2% equivalents per liter.

Electrometric Titration Procedure.—The side arm of an ordinary decinormal potassium chloride calomel electrode was dipped directly into the solution prepared for the titration. In this was immersed an apron type foil platinum electrode. No attempt was made to exclude liquid junction potentials because of other uncertainties involving poisoning of the electrode which could be of much greater magnitude. A Leeds and Northrup type K potentiometer, a calibrated standard cell and electrolytic hydrogen generator completed the equipment. The hydrogen electrode was first brought in equilibrium with a known amount of 0.1 *N* standard acid and then an amount of organic base added which was not quite equivalent to this acid. The solution was then titrated with 0.1 *N* alkali as usual.

The titration curves are given in Fig. 1 and the data for the calculations given in Tables I and II (see previous section). The hydrogen ion concentration was calculated from the equation

$$\log \frac{1}{(\mathrm{H}^+)} = \frac{E - 0.3358}{0.00019844T}$$

In the case of α-diethylamino-acetonitrile, the constant of hydrolysis of the salt and the basic dissociation constant of the base were calculated from the expressions

$$K_{\mathrm{H}} = \frac{(\mathrm{H}^+)^2}{(\text{concn. of salt}) - (\mathrm{H}^+)};\ K_B = \frac{K_W}{(K_{\mathrm{H}})}$$

The concentration of salt was determined from the total volume (volume of acid plus volume of alkali) at the true end-point, and the amount of base used. The other calculations were based on the simplified expression

$$K_B = \frac{(\text{Salt})}{(\text{Base})}\frac{K_W}{(\mathrm{H}^+)}$$

In the case of the phenylacetonitrile derivative the last equation given was used. The concentration of the base was constant and equal to that in a saturated solution; that of the salt (cation) was calculated from: (*a*) total volume of the solution and (*b*) total equivalents of salt present (base used minus base decomposed minus salt neutralized and hydrolyzed). The amount of base decomposed was given by the difference between the acid started with and acid found, as shown by the second inflection of the titration curve. Most of the reaction of decomposition took place before the back-titration started and used up acid by the formation of a strong base, presumably diethylamine.

Colorimetric Determination of Relative Base Strengths in Alcohol.—A wedge type colorimeter was used. The two wedges were fitted with the stock alcoholic acid–indicator solution, and one of them was then made alkaline with a few drops of diethylamine. They were each mounted separately on movable frames provided with a scale reading from 0 to 100. The comparator cup was of such dimensions as to give the same depth of solution as that provided by the wedges when their combined scale readings were 100. Careful calibration of the colorimeter was not necessary because it was used in a purely empirical way. A daylight lamp furnished illumination, reflected from a frosted glass background. Manipulation of the wedges permitted variation of the color shades and intensity, and the recorded scale readings allowed reproduction at will of any previously used color standard. In practice it was found that the alkaline wedge could be shifted appreciably without sensible alterations of the color shade. Readings of the acid wedge scale only were used.

The organic bases to be compared were weighed into a small flask provided with glass tubing attachments similar to those of a wash bottle. The stock acid–indicator solution was added in desired quantity from a buret the tip of which passed through the stopper of the flask. At any given time some of the solution could be forced over into the comparator cup, the color comparison made, withdrawn back into the flask, more acid added and the operation repeated until the possible range of color comparison was covered. Small amounts of water caused no appreciable change in the color of the solutions, and in the system employed there was no exposure of the solution to the air except through calcium chloride filled tubes, and except that the comparator cup had a small opening for the insertion of a glass tube through which the liquid was introduced and withdrawn. The acid wedge reading was plotted for each base against its percentage titration and smooth curves were drawn through the points. A given wedge reading then represented a corresponding hydrogen ion activity for each of the various solutions. A summary of the data is given in Table III. By the above method, using the same stock solution for all the bases, the comparisons are made at equal cation, hydrogen ion and anion concentrations (assuming complete ionization of acid and salts). The ratio of the dissociation constants is then very closely the inverse ratio of the concentrations of the free bases under these conditions. Moreover, the total ionic strength of the solution changes but very little during the whole course of the titration.

Summary

The basic dissociation constants of α-diethylamino-acetonitrile and of α-diethylaminophenylacetonitrile in water are 1.0×10^{-10} and 1.7×10^{-11}, respectively. In alcohol solution, the acidic dissociation constants of their salts (chlorides) are 3.4×10^{-5} and 5.2×10^{-4}, respectively, while those of α-diethylaminopropionitrile and of α-diethyl*iso*butyronitrile in alcohol are 3.4×10^{-5} and 2×10^{-6}, respectively. The base strengths of some

amino ethers in alcohol have been recalculated. Some of the factors which determine the apparent change in relative base strength with change of solvent are discussed.

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

CYCLOPROPYLMETHYL ALKYL ACETIC ACIDS AND THEIR BACTERICIDAL ACTION TOWARD B. LEPRAE. XIII[1]

BY J. A. ARVIN[2] AND ROGER ADAMS

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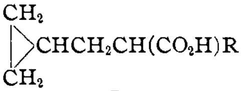
The various synthetic acids which have been studied for their bactericidal action have so far contained either a five or six membered ring. The present investigation has covered the synthesis of a series of acids containing a three-membered ring, namely, the cyclopropylmethyl alkyl acetic acids, (I) where the R varies from *n*-amyl to *n*-myristyl. These acids are of interest because they lead to the possibility of determining whether the significant factor for the bactericidal action in the many acids that have been studied is the character of the ring or is the molecular weight of the whole molecule, or merely that part of the molecule outside of the ring structure.



$$\begin{matrix} CH_2 \\ | \rangle CHCH_2CH(CO_2H)R \\ CH_2 \end{matrix}$$

I

In Table I are given the bacteriological results.

TABLE I

CYCLOPROPYLMETHYL ALKYL ACETIC ACIDS, $C_3H_5CH_2CH(CO_2H)R$

	Dilutions of sodium salt in thousands																		
R =	10	20	30	40	50	60	70	80	90	100	111	125	133	143	153	167	176	185	200
n-C_5H_{11}	+	+	+	+	+	+	+	+	+	+									
n-C_6H_{13}	+	+	+	+	+	+	+	+	+	+									
n-C_7H_{15}	−	+	+	+	+	+	+	+	+	+									
n-C_8H_{17}	−	−			+			+		+									
n-C_9H_{19}	−	−			−			−		+									
n-$C_{10}H_{21}$										−	−	−	−	−	+	+	+	+	+
n-$C_{11}H_{23}$										−	−	−	−	−	±	±	−	±	+
n-$C_{12}H_{25}$										−	−	±	−	+	+	+	+	+	+
n-$C_{14}H_{29}$										+	+	+	+	+	+	+	+	+	+

It is noticeable that practically no bactericidal effect appears until the R group is *n*-octyl or larger. The two most effective acids were those containing the decyl and the undecyl groups. It appears, then, that in this as in other series, at least sixteen carbon atoms in the molecule must

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[1] Paper XII in this series, THIS JOURNAL, **50**, 1790 (1928).

[2] This communication is an abstract of a portion of a thesis submitted by J. A. Arvin in partial fulfilment of the requirements for the degree of Doctor of Philosophy in Chemistry at the University of Illinois.