glycolic acids is already low, and near the value of D_i, the internal dielectric constant. Computations of D_E for methanol and ethanol show that, for the case of the dipoles in chloroacetic and glycolic acids, using alcohol instead of water as solvent will not appreciably increase the electrostatic effect.

While it is not advisable to extrapolate the data for the aliphatic dicarboxylic acids[14] obtained in water–alcohol mixtures to pure alcohol as solvent, it is obvious that the ratio of the first and second dissociation constants for malonic and succinic acids, while greater in methyl and ethyl alcohols than in water, is certainly not increased by as much as the simple theory predicts, but by a smaller amount, a fact at least qualitatively in agreement with the prediction of the present theory for these cases. Quantitative agreement, in the case of non-aqueous solutions, must await further experimentation.

The errors inherent in the present method of estimating the electrostatic effect, other than those given here, have been discussed in the previous publications. The reasons for omitting

consideration of Ingold's attempt to take electrostriction and electrical saturation into account[16] have been presented elsewhere.[2] The molecular volumes necessary for the calculations have been estimated from Traube's rule.[17]

The authors wish to thank Professor J. G. Kirkwood for his advice and assistance in preparing this paper.

Summary

The new mathematical formulation, by Kirkwood and Westheimer, of the electrostatic effect of a substituent, has been shown to account satisfactorily for the ratio of the ionization constant of a dibasic acid, for the ratio of the ionization constant of a dipole substituted acid to that of the unsubstituted acid, for the ratio of the ionization constant of an amino acid with that of a salt of its ester. The theory also accounts for the effect of alkyl groups and, at least in the few cases which can now be examined, for the effect of solvent, upon the ionization constant ratio.

(16) Ingold, J. Chem. Soc., 2179 (1931).
(17) Traube, Saml. chem. chem.-tech. Vortr., 4, 255 (1899).

The Hydrolysis of Nitriles with Acids

BY VERNON K. KRIEBLE AND CLARENCE I. NOLL

In previous communications from this Laboratory attempts have been made to correlate reaction rates in hydrolytic reactions catalyzed by concentrated solutions of strong mineral acids with physical properties of these acid solutions.[1] One of the most interesting of these reactions was the hydrolysis of hydrogen cyanide with hydrochloric acid, in which case the velocity constants varied as the square of the mean ion activity. It also was shown that in this particular reaction hydrobromic acid and especially sulfuric acid are much inferior to hydrochloric acid as catalysts for concentrations greater than 3 molal. The investigation has been continued covering the hydrolysis of substituted hydrogen cyanide or organic nitriles using the same mineral acids as catalysts.

Preparation of Material and Method of Procedure

Acetonitrile.—The Eastman Kodak Company preparation was fractionated through a Widmer distilling column.

The fraction used distilled over completely in the temperature range 82.0–82.5°.

Propionitrile.—The Eastman product was treated in the same way, using a fraction boiling over between 96.6–96.8°.

β-Hydroxypropionitrile.—The Eastman preparation was fractionated at 21.5 mm. pressure. The fraction used came over at 122–123°.

Cyanoacetic Acid.—The preparation used was manufactured by C. A. F. Kahlbaum. It was purified by heating it in chloroform and then shaking the solution until cold. In this way large white crystals separated. In a sealed tube they melted 69–70°.

α-Hydroxypropionitrile.—We were unable to get a pure sample of this nitrile. An attempt was made to distil a Kahlbaum preparation under reduced pressure but it decomposed. The best sample we could get was obtained by taking the middle fraction in a vacuum distillation with the aid of solid carbon dioxide–ether bath. This fraction was free of inorganic compounds which would act as catalysts in the hydrolytic studies. The nitrile, however, still had organic impurities as complete hydrolysis only gave 91.2% of the theoretical ammonia.

Acids.—The C. P. grades of acid supplied by Baker and Adamson were used. The hydrobromic acid was distilled

(1) (a) THIS JOURNAL, 51, 3368 (1929); (b) 55, 2326 (1933); (c) 57, 15 (1935); (d) 57, 19 (1935); (e) 60, 2976 (1938).

three times from red phosphorus in an apparatus having ground glass connections.

Procedure.—The concentrated acids were analyzed and then diluted to the exact molality given in the tables. The acid was cooled to 0° and the nitrile added. With acetonitrile and propionitrile one mole per liter of nitrile was added except where the acid concentration was 2 molal or less then 0.5 mole was used. In case of the other nitriles 0.5 mole was used in all of the experiments. The solution was quickly and thoroughly mixed and then 10 to 12 Pyrex tubes were filled and tightly stoppered with rubber stoppers leaving an air bubble for expansion of the liquid when the tubes were warmed. The stoppers were wired in to prevent them from blowing out during an experiment. The tubes were placed in a wire basket, stoppers pointing downward, and suspended in a thermostat kept at 65 ± 0.1°. From time to time tubes were removed and immediately suspended in an ice-bath. It was assumed that it took as long for the tubes to warm from 0 to 65° as it did to cool them to 0°.

Method of Following the Hydrolysis.—Nitriles hydrolyze to amides first and then to the acid and ammonia. Our own investigations as well as that of Taylor[2] showed that the second reaction is many times as fast as the first so that by estimating the amount of ammonia present in the hydrolytic solution it is possible to calculate the percentage of nitrile hydrolyzed. The ammonia was estimated by Folin's aspiration method. The velocity constants were calculated on the basis of a monomolecular reaction with respect to the nitrile, time being taken in hours. Each constant recorded in Tables I and II is the average of at least five experimental rate determinations. All determinations made before 10% and after 80% of the nitrile was hydrolyzed were excluded from the average. In only a few cases did experimental values vary more than 5% from the recorded velocity constant.

Experimental Results

TABLE I
HYDROCHLORIC ACID AS CATALYST

Molality HCl	$K \times 10^3$	$(K \times 10^3)/a^\pm$	$(K \times 10^3)/(a^\pm)^2$
		Acetonitrile	
2	0.60	0.29	0.145
4	2.07	.28	.038
6	7.7	.38	.018
8	38.7	.77	.015
10	107	1.01	.010
		Propionitrile	
4	3.6	0.49	0.066
6	14.4	.71	.035
8	78.4	1.55	.031
14.9	140.4	0.241	.0004
		β-Hydroxypropionitrile	
4	1.85	0.25	0.034
6	6.6	.32	.016
8	34.7	.69	.014

(2) Taylor, *J. Chem. Soc.*, 2741 (1930).

	α-Hydroxypropionitrile		
4	28	3.8	0.52
6	259	12.7	.62
8	1022	20.3	.40
	Cyanoacetic acid		
2	0.28	0.14	0.067
4	1.2	.16	.022
6	8.3	.41	.020
8	62.1	1.23	.024

TABLE II
SULFURIC ACID AS CATALYST

Acetonitrile		β-Hydroxypropionitrile	
2	0.85	4	2.7
4	2.65	6	6.3
6	6.04	α-Hydroxypropionitrile	
8	10.60		
Propionitrile		4	6.1
		6	17.5
4	4.6	Cyanoacetic acid	
6	9.1	4	1.26
		6	2.66
		8	5.66

Discussion

It is apparent that the velocity constants vary widely if a comparison is made for one acid concentration with the different nitriles, for the same nitrile and different concentrations of acid or for the same nitrile, same acid concentrations but different acids. The important question, however, is can this wide variety of velocity constants be correlated to other properties of the acid or nitrile so that predictions can be made in regard to the hydrolysis of other organic substances, particularly nitriles.

One cannot help being impressed by the large increase in the rate of hydrolysis as the concentration of hydrochloric acid is stepped up. This is similar to the hydrolysis of hydrogen cyanide, in which case the velocity constant divided by the square of the mean ion activity was a constant. This indicates that either the undissociated molecule or both ions acting simultaneously are the sole catalyst. This relationship, last column Table I, does not hold as well for nitriles as it did for hydrogen cyanide. It, however, holds best at the higher concentrations, 4 to 8 molal, where a slight change in concentration has the greatest effect on the velocity. It is for this reason that we believe that the principal catalyst is either the undissociated molecule or both ions just as in the case of hydrogen cyanide.

It is quite obvious from Table II that sulfuric acid is not only a poorer catalyst than hydro-

chloric acid but also that a change in concentration of the acid has a much smaller effect especially at the higher concentrations. This is in agreement with the hydrolysis of hydrogen cyanide.

During recent years Hammett and co-workers[3] have developed a new method of treating the acidity of solutions in terms of an acidity function which is a measure of the hydrogen ion concentration as determined by basic indicators. They have been quite successful in correlating velocity constants for a widely divergent series of chemical reactions with this function by applying the equation $H_0 + \log K =$ constant where H_0 is the acidity function. We have pointed out before[1b] that this relationship does not hold for the hydrolysis of hydrogen cyanide where sulfuric acid is the catalyst, nor does it hold for hydrochloric acid. Table III shows that this simple relationship between velocity constants and acidity function in nitrile hydrolysis does not exist. It shows that the velocity constants for hydrochloric acid go up faster than the acidity function while for sulfuric acid the reverse is the case.
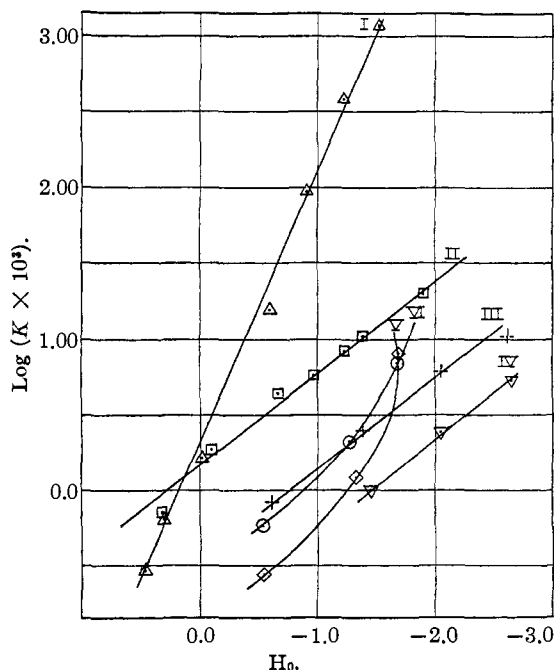

Fig. 1.—Curves indicate the relationship of the acidity function to the velocity constant for nitrile hydrolysis in hydrochloric acid and sulfuric acid: I, HCN in HCl; II, HCN in H_2SO_4; III, CH_3CN in H_2SO_4; IV, $CNCH_2COOH$ in H_2SO_4; V, CH_3CN in HCl; VI, $CNCH_2COOH$ in HCl.

(3) (a) Hammett and Deyrup, This Journal, 54, 2721 (1932); (b) Hammett and Paul, ibid., 56, 827 (1934); (c) Hammett, Chem. Rev., 16, 67 (1935).

TABLE III

Molality	$H_0 + \log K \times 10^3$ HCl	$H_0 + \log K \times 10^3$ H_2SO_4
	CH_3CH_2CN	
4	−0.65	−0.79
6	− .59	−1.09
8	− .39	
	$CH_3CHOHCN$	
4	+0.25	−0.67
6	+ .66	− .81
8	+ .72	
	$COOHCH_2CN$	
2	−0.10	
4	− .12	−1.35
6	+ .17	−1.63
8	+ .50	−1.90

The acidity function for an 8 molal hydrochloric acid solution was an extrapolated value based on the data given by Hammett.[3c]

In those reactions where the relationship, $H_0 + \log K =$ constant, holds, one obtains a straight line with a slope of −1 when $\log K$ is plotted against H_0. This relationship as we have said before does not hold in hydrogen cyanide or nitrile hydrolysis. It is, however, worth while to point out that in both of these hydrolyses, a straight line is obtained, Fig. 1, when sulfuric acid is the catalyst but the slope is not −1. In the case of hydrochloric acid a straight line is obtained only where hydrogen cyanide is the substrate.

We have only meager data on hydrobromic acid and then only on acetonitrile. It is not as good a catalyst for this nitrile as is hydrochloric acid. This can be seen easily by comparing the velocity constants in Table I with those in Table IV. The difference between acids in nitrile hydrolysis is not as great as it is in hydrogen cyanide hydrolysis. Table IV shows that varying the concentration of hydrobromic acid has a much greater effect on hydrogen cyanide than on acetonitrile hydrolysis. By plotting these rates it is apparent

TABLE IV

HBr Molal	$K \times 10^3$ CH_3CN	$K \times 10^3$[a] HCN
1.03		0.94
2.00	0.59	
2.12		2.80
3.25		8.50
4.00	1.89	
4.45		47
5.74		292
6.00	4.1	
7.13		1640

[a] This Journal, 55, 2327 (1933).

that a 6 molal hydrobromic acid solution hydrolyzes hydrogen cyanide 75 times faster than acetonitrile.

One of the most interesting results brought out by these experiments is the profound influence that the R in RCN has on nitrile hydrolysis.

TABLE V

	4 Molal H_2SO_4	6 Molal H_2SO_4	4 Molal HCl	6 Molal HCl	8 Molal HCl
CH_3CN	2.6	6.	2.1	7.7	39
CH_3CH_2CN	4.6	9.1	3.6	14.4	78
$COOHCH_2CN$	1.2	2.6	1.2	8.3	62
$\dfrac{CH_3}{OH}{>}CHCN$	6.1	17.5	28	259	1022
CH_2OHCH_2CN	2.7	6.3	1.9	6.6	35

The rates are tabulated in Table V. If one leaves out of consideration for the moment the hydrolysis of cyanoacetic acid with 6 and 8 molal hydrochloric acid, it is seen that the R has the same influence irrespective of the acid or the concentration of the acid. This shows that the polarity of R is just as important in nitrile hydrolysis as it is in the hydrolysis, reduction or inverse substitution of alkyl halides.

If acetonitrile is taken as a reference standard then the substitution of a methyl, a mild basylous radical for a hydrogen atom in the methyl increases the velocity very markedly and the further substitution of a hydroxyl, another basylous radical, increases the rate still more. If, however, an acylous radical like the carboxyl group is substituted the rate is cut in half. The order of decreasing activity of the radicals attached to the CN group investigated is as follows

$CH_3CHOH > CH_3CH_2 > CH_3, CH_2OHCH_2 > COOHCH_2$

Since, according to Kharasch, ethyl is less electronegative than methyl, this series of radicals is in the order of increasing electronegativity. Shoesmith with co-workers and Olivier found the same relationship for alkyl halide hydrolysis in neutral or acid solutions while Berger and Olivier found the opposite effect for acyl halide hydrolysis.[4]

(4) Shoesmith and Slater, *J. Chem. Soc.*, **125**, 2278 (1924); Shoesmith and Slater, *ibid.*, 214 (1926); Shoesmith and Connor, *ibid.*, 1769 (1927); Olivier, *Rec. trav. chim.*, **41**, 646 (1922); **42**, 516 (1923); Berger and Olivier, *ibid.*, **46**, 517 (1927).

The reason that 6 and 8 molal hydrochloric acid solutions acting on cyanoacetic acid do not fall in line is no doubt due to secondary reaction between the carboxyl group and the hydrochloric acid.

From the practical point of view it is worth while to point out that a 10 molal hydrochloric acid solution is probably the most favorable concentration for nitrile hydrolysis. The rates for propionitrile in Table I show that a 14.9 molal solution is the highest, but from the rates for acetonitrile it is observed that the 10 molal solution is 3 times as effective as the 8 molal. If this increase is applied to the rate for 8 molal on propionitrile it is found to be nearly double that for the 14.9 molal. This indicates a maximum rate for nitrile hydrolysis which has already been observed for amides. It is quite possible that at these high hydrochloric acid concentrations the hydrolysis of the nitrile to the acid is held up by a much slower hydrolysis of the intermediate amide.
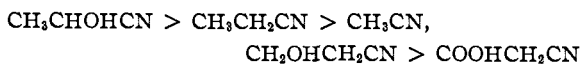
Summary

1. The rates of hydrolysis of five nitriles, using hydrochloric acid and sulfuric acid as catalysts, have been studied. The results show that for hydrochloric acid the rate of hydrolysis increases approximately as the square of the mean ion activity increases.

2. Sulfuric acid is a poorer catalyst than hydrochloric acid, especially at the higher concentrations of the acid.

3. There is no relationship between the acidity of the solutions, as determined by basic indicators, and the velocity of hydrolysis.

4. The order of hydrolysis for the nitriles studied is as follows

$CH_3CHOHCN > CH_3CH_2CN > CH_3CN,$

$CH_2OHCH_2CN > COOHCH_2CN$