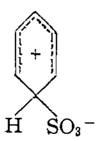
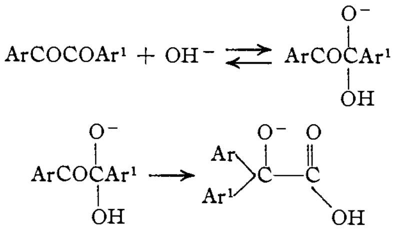
that an isotope effect was observed in aromatic sulfonation. This is accounted for readily if the intermediate V is formed. Since the intermediate in this case would be neutral it might have unusual


V

stability and the removal of a proton from it might involve considerable weakening of the C–H bond before the transition state is reached. Again it is possible that the intermediate is a fiction since the same description of the transition state could be arrived at without reference to it.

A last example will be discussed briefly since it illustrates the useful application of the postulate in a concrete manner. The benzilic acid rearrangement of an unsymmetrical benzil can be formulated as a stepwise process in the manner


$$ArCOCOAr^1 + OH^- \rightleftarrows ArCO\overset{O^-}{\underset{OH}{C}}Ar^1 \quad (3)$$

(4)

Alternatively, the Ar group could migrate from $ArC(O^-)(OH)COAr^1$. Relative migratory aptitudes of two groups would be determined by influence of substituents on the two equilibrium constants and the two velocity constants. Furthermore, the effects in the two steps should tend to compensate since substituents which promote the formation of the tetrahedral structure at the adjacent carbonyl group should also tend to slow down migration of an aryl group from that position. The prediction of the over-all effect can, however, be made unambiguously. Since the adduct ion can be considered as an unstable intermediate the structure of the transition state in reaction 4 will be closely related to it. Therefore, the dominating influence on the over-all reaction rate will be that predicted for reaction 3 alone. It has recently been found, by the use of isotopic carbon to distinguish between the carbonyl groups, that the direction of rearrangement of unsymmetrical benzils is in accord with this prediction.[28]

Acknowledgment.—The author is much indebted to friends, too numerous to mention, who have read this manuscript and made many important suggestions as to its composition. Particular mention should be made of the contribution which has been made in this way by Professor John D. Roberts who suggested among other things the application of the postulate to the benzilic acid rearrangement.

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[28] J. D. Roberts, D. R. Smith and C. C. Lee, THIS JOURNAL, **73**, 618 (1951).

[CONTRIBUTION FROM THE CHEMICAL LABORATORY OF IOWA STATE COLLEGE]

The Dissociation of Sterically Hindered Acids

BY GEORGE S. HAMMOND AND DONALD H. HOGLE

RECEIVED APRIL 7, 1954

The ionization constants of a series of highly hindered aliphatic acids have been measured and it was found that the accumulation of bulky groups in the vicinity of the carboxyl function has a pronounced acid weakening influence. This is interpreted as indicating that the dominant effect is the shielding of the carboxylate ions from solvation. The commonly encountered view that acid strengthening effects observed in some hindered systems is due to the compression of the acidic proton is criticized. Two factors, steric hindrance to solvation and steric inhibition of resonance, are sufficient to give a qualitative account of steric effects on the strength of proton acids.

Because a number of sterically hindered acids were made available to us[1] we have measured the ionization constants in order to assess the effect of steric hindrance in unconjugated systems. The measurements were carried out by potentiometric titration in "50 per cent." methanol–water with a glass electrode and a calomel reference electrode. The solvent was chosen as representing the best available compromise based upon the slight solubility of the acids in water and the non-ideal behavior of electrolytes in non-aqueous solvents. It is anticipated that the unknown correction for the difference in the junction potential from that involved in the use of the same electrodes in water is small and constant throughout the series and that the data in Table I represent a reasonable approximation to the *relative* thermodynamic dissociation constants. For comparison data previously reported[2] for the acidity constants of other aliphatic acids in a similar medium are summarized in Table II. While the values for benzoic acid do not check well between the two series of measurements this difference may be due to a medium effect as the earlier works did not specify the way in which their solvent was prepared. Another possible source of the discrepancy between the two series of measurements may be found in the fact that different electrode systems, involving different variations in the junction potentials, were used in the two investigations. It is evident, however, that the usual

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[1] The acids were made available for the purpose of this study by Prof. M. S. Newman. Most of the compounds were originally a part of the chemical collection of the late F. C. Whitmore. The exact history of the samples is unknown, but it is inferred that they were prepared by the methods described in the various references given in the Experimental section.

[2] W. L. Bright and H. T. Briscoe, *J. Phys. Chem.*, **37**, 787 (1933).

variations[3] of acidity with structure are found with the unhindered acids and that the steric effects are an order of magnitude larger than any of the others.

TABLE I

IONIZATION CONSTANTS OF HINDERED ACIDS IN 50% BY VOLUME METHANOL–WATER AT 40°

Acid	Apparent pK_a		
	1/4	1/2	3/4
Benzoic (measured at 23°)		5 62	
Acetic	5.57	5.55	5.55
Methylneopentylacetic	6.04	6.05	6.07
Methyl-*t*-butylacetic	6.25	6.25	6.24
Ethyl-*t*-butylacetic	6.31	6.32	6.31
Diisopropylacetic	6.40	6.40	6.39
Triethylacetic	6.44	6.44	6.44
Dimethylneopentylacetic	6.49	6.51	6.51
Dineopentylacetic	6.51	6.54	6.62
Dimethyl-*t*-butylacetic	6.74	6.72	6.71
Methyl-*t*-butylneopentylacetic	6.88	6.97	7.02

TABLE II

ACID STRENGTHS IN 50% METHANOL–WATER AT 25°[2]

Acid	pK_a	Acid	pK_a
Benzoic	5.15	Isobutyric	6.23
Propionic	5.84	Isovaleric	5.89
n-Butyric	5.88	*n*-Caproic	5.91

Steric hindrance decreases the ionization constants by substantial amounts with the maximum effect being reached with methyl-*t*-butylneopentylacetic acid which is 1/25 as strong as acetic acid. We attribute the result to steric hindrance to solvation of the carboxylate anions which apparently overshadows any effect due to the compression of the acidic proton of the acids.

A question arises as to the most appropriate way of describing the solvation effect. On the one hand the introduction of bulky groups lowers the effective dielectric constant in the vicinity of the negative carboxyl group which should increase its classical electrostatic free energy. At the same time the availability of the carboxyl oxygens for hydrogen bonding to solvent molecules is decreased. Even if these effects are separable in principle they certainly cannot be differentiated on the basis of our data.

Steric effects on the acidity of both carboxylic acids and ammonium ions have been discussed frequently in the earlier literature. It has been known for many years that *o*-substituted benzoic acids are regularly much stronger than their *para* isomers and that the *cis* isomers of α,β-unsaturated acids are usually stronger than their *trans* isomers. The effect has been attributed to the steric requirements of the acidic protons.[2,4,5] The results with hindered aliphatic acids render such an explanation rather untenable. It appears much more probable that the acid-strengthening effect is due to the twisting of the carboxyl groups away from the plane of the rest of the unsaturated systems as suggested by Ingold.[6] Recently a discussion of the possible interplay of steric effects on solvation and on resonance interaction in hindered benzoic acids has been presented.[7] There are, of course, a variety of special effects, such as hydrogen bonding, which can come into play when polar substituents are located close to the acidic function.[8]

Also of interest in this connection is the analysis by Brown and Cahn[9] of *ortho* effects on the basicity of anilines. In addition to the base-strengthening effect due to steric hindrance of resonance in the free base there is some other factor which is base weakening since there is not a monotonic response to the accumulation of steric effects. We believe that steric hindrance to the solvation of the anilinium ions should be of more importance than the increase in the internal steric requirements of the functional group on addition of a proton. Pearson and Williams[10] have reached the same conclusion on the basis of the observation that the relative basicities of hindered amines are strongly solvent dependent. This problem is of considerable importance since the neglect of solvation effects will lead to overestimation of the steric strain in neutral "homomorphs" of the hindered anilinium ions.[9]

Another observation bears upon this question. It has been shown[11] that di-*o*-*t*-alkyl phenols are much weaker acids than other alkyl phenols. As with the aliphatic acids it is found that the bulky groups stabilize the acid form relative to the anion, presumably by exclusion of solvent from the site of the negative charge.[12] Since the steric situation is rather similar to that in hindered aromatic amines the implication is rather strong that bulky groups decrease the stability of ionic species irrespective of the presence or absence of the acidic protons.

Experimental

Acetic acid was purified by refluxing over potassium permanganate followed by fractional distillation through a glass helix packed column from acetic anhydride.

Methylneopentylacetic acid was fractionated in a glass center rod column at *ca.* 1 mm. pressure, b.p. 86° (uncorrected), n^{20}_D 1.4231; literature[13] b.p. 217° (730 mm.); n^{20}_D 1.4233; neutral equivalent, calcd. 144, found 145.

Dimethylneopentylacetic acid was used without further purification, m.p. 45°; literature[14] m.p. 45°; neutral equivalent, calcd. 156, found 156.

Methyl-*t*-butylneopentylacetic acid was purified by extracting a methanolic sodium hydroxide solution with pentane, acidification of the methanol–water solution with concentrated HCl and extraction of the acid solution with pentane. After several repetitions of this procedure, the m.p. rose from 102–116° to 116–126° and the neutral equivalent decreased from 270 to 250 (calcd. 200). The small size of the sample prevented any further purification. The literature[13,15] gives m.p. 130°. Since a small amount of neutral material could be extracted from the sodium salt with pentane it was inferred that the acid had undergone partial decarboxylation during the many years which had elapsed since the sample was originally prepared.

Dineopentylacetic acid was used without further purifica-

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[3] J. F. J. Dippy, *Chem. Revs.*, **25**, 151 (1939).

[4] B. Flurscheim, *J. Chem. Soc.*, 725 (1909).

[5] J. F. J. Dippy, *et al.*, *ibid.*, 1421 (1937).

[6] C. K. Ingold, "Structure and Mechanism in Organic Chemistry," Cornell University Press, Ithaca, N. Y., 1953, pp. 743–750.

[7] H. L. Goering, T. Rubin and M. S. Newman, THIS JOURNAL, **76**, 787 (1954).

[8] G. E. K. Branch and M. Calvin, "The Theory of Organic Chemistry," Prentice-Hall, Inc., New York, N. Y., 1941, p. 259.

[9] H. C. Brown and A. Cahn, THIS JOURNAL, **72**, 2939 (1950).

[10] R. G. Pearson and F. V. Williams, *ibid.*, **76**, 258 (1954).

[11] G. H. Stillson, D. W. Sawyer and C. K. Hunt, *ibid.*, **67**, 303 (1945).

[12] P. D. Bartlett, *J. Chem. Education*, **30**, 22 (1953).

[13] F. C. Whitmore, *et al.*, THIS JOURNAL, **63**, 643, 2028 (1941).

[14] F. C. Whitmore, *et al.*, *ibid.*, **63**, 2035, 3237 (1941).

[15] F. C. Whitmore, *et al.*, *ibid.*, **56**, 1128 (1934).

tion, m.p. 52–53°; literature[16] m.p. 53–55°; neutral equivalent, calcd. 132, found 130.

Dimethyl-*t*-butylacetic acid was used without further purification, m.p. 200°; literature[17] m.p. 196–197°; neutral equivalent, calcd. 144, found 146.

Ethyl-*t*-butylacetic acid was used without further purification, m.p. 79.5–80.5°; literature[18] 76–77°; neutral equivalent, calcd. 144, found 146.

Triethylacetic acid was used without further purification, m.p. 39°; literature[19] m.p. 39°; neutral equivalent, calcd. 144, found 142.

Diisopropylacetic acid was used without further purification, b.p. 82–83° at 2 mm., $n^{24.4}_{\mathrm{D}}$ 1.4243; neutral equivalent calcd. 144.2, found 144.7.

Ionization Constants.—The apparent ionization constants were determined in 50 volume per cent. methanol-water solution at 40° by the method of one-quarter, one-half and three-quarter points in the titration curve.

A Beckman *p*H meter, model G, was used with calomel and glass electrodes and was standardized against buffer solutions at *p*H 4 and 7.

The measurements were obtained by dissolving approximately 0.75 meq. of the acid in 100 ml. of the aqueous methanol and titrating with 0.0957 *N* methanolic sodium hydroxide.

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[16] J. G. Aston, J. T. Clarke, K. A. Burgess and R. B. Greenburg, THIS JOURNAL, **64**, 300 (1942).

[17] F. C. Whitmore, *et al.*, *ibid.*, **63**, 1626 (1942).

[18] K. L. Loening, A. B. Garrett and M. S. Newman, *ibid.*, **74**, 3929 (1952).

[19] F. C. Whitmore, *et al.*, *ibid.*, **64**, 2964 (1942).

[A CONTRIBUTION FROM THE CHEMICAL LABORATORY OF IOWA STATE COLLEGE]

The Displacement of Halogen from 1-X-2,4-Dinitrobenzenes

BY GEORGE S. HAMMOND AND LAWRENCE R. PARKS

RECEIVED JUNE 4, 1954

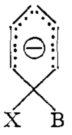
The rates of the reaction of N-methylaniline with the 1-X-2,4-dinitrobenzenes in which X is F, Cl and Br have been measured in nitrobenzene and in 99.8% ethanol as solvents. In both media the order of reactivity, at the temperatures studied, is Br > Cl > F. Preliminary estimates show that variations in the activation entropies make significant contributions to the relative reactivities. It is inferred that bond breaking must have progressed to a significant extent in the transition states.

Considerable attention has recently been directed toward the study of the nucleophilic displacement reactions of activated aryl halides.[1-9] The review of Bunnett and Zahler[1] established the frame of reference for the discussion of the mechanism of such reactions. Further discussions have been presented by Berliner and co-workers[2] and by Chapman and his co-workers.[4-7] The latter have stated the status of the problem succinctly.[6,7]

One of the striking features of the reaction is the regularity with which fluoride appears as a more reactive leaving ion than the other halide ions. This observation stands in sharp contrast to the behavior of the halides in displacements at saturated carbon atoms,[10] and the phenomenon has been the focal point of the work reported at this time.

Results and Discussion

The high reactivity of activated aryl fluorides as compared with other, similarly activated, aryl halides may be taken as an implication that bond-breaking has not made nearly as much progress in the transition states of the rate-controlling steps as is inevitably accomplished in the transition states involved in nucleophilic aliphatic displacements. This view is in substantial agreement with the conclusions concerning the nature of such transition states.[4-7] While there is much uncertainty as to exact timing of the reaction it is generally conceded that the transition states are usually near configuration I in which the carbon atom undergoing substitution is nearly tetrahedral. It is understandable that, if the free energy of formation of such a structure determines the rate of the reaction, the relative reactivity of the halogens might be largely determined by the C–X bond moments rather than by the other factors, such as atom polarizability,[11] which come into prominence in reactions in which the rate-controlling steps involve stretching the old bonds well beyond normal covalent distances. If this view is correct it should be possible to accomplish the inversion of the "normal" reactivity order by changing the reactions in such a way as to require more bond breaking in the rate limiting transition states.



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That such an inversion has been accomplished is borne out by the data in Table I. With N-methylaniline as the nucleophile, the relative rates of halide displacement from 1-X-2,4-dinitrobenzenes are Br > Cl > F in both nitrobenzene and ethanol

[1] J. F. Bunnett and R. E. Zahler, *Chem. Revs.*, **49**, 273 (1951).

[2] E. Berliner, M. J. Quinn and P. J. Edgerton, THIS JOURNAL, **72**, 5305 (1950); E. Berliner and L. C. Monack, *ibid.*, **74**, 1574 (1952).

[3] O. L. Brady and F. R. Cropper, *J. Chem. Soc.*, 507 (1950)

[4] N. B. Chapman, R. E. Parker and Soanes, *Chem. and Ind.*, 148 (1951).

[5] N. B. Chapman and R. E. Parker, *J. Chem. Soc.*, 3301 (1951).

[6] R. R. Bishop, E. A. S. Cavell and N. B. Chapman, *ibid.*, 437 (1952).

[7] E. A. S. Cavell and N. B. Chapman, *ibid.*, 3392 (1953).

[8] J. Miller, *ibid.*, 3550 (1952); A. L. Beckwith, J. Miller, and G. D. Leahy, *ibid.*, 3552 (1950); J. Miller, THIS JOURNAL, **76**, 448 (1954).

[9] C. K. Ingold, "Structure and Mechanism in Organic Chemistry," Cornell University Press, Ithaca, N. Y., 1953, Chapt. 15.

[10] C. K. Ingold, ref. 9, p. 338.

[11] J. O. Edwards, THIS JOURNAL, **76**, 1540 (1954), mentions the fact that the second term in his double basicity scale is apparently of much greater importance in determining relative nucleophilic reactivity in displacements at unsaturated carbon atoms than in saturated displacements. This leads to conclusions similar to those developed herein except that the application is to the extent to which the new bond has developed covalent character in the transition state.