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result from the presence of methyl radicals which can remove a hydrogen atom from many organic molecules, forming methane and a more complicated free radical which may then react further. The great difference in the effect of methyl radicals on various organic substances, some of them quite similar, is of considerable interest. Azomethane itself is presumably very little attacked if at all. It is impossible to say whether methane reacts with methyls or not as the products would be the same as the reactants. Ethane and propane are presumably only slightly attacked by methyls, but isobutane, acetaldehyde[17] and propionaldehyde[18] are so readily attacked at the temperature at which azomethane decomposes that the side reaction becomes a main reaction. It is, of course, to be remarked that relatively small differences in the activation energies involved would account for the observed effects.

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[17] Allen and Sickman, J. Am. Chem. Soc. **56**, 2031 (1934).
[18] O. K. Rice and Sickman, J. Am. Chem. Soc. **57**, 1384 (1935).

The kinetics of the polymerization of ethylene induced by methyl radicals from azomethane has already been described in outline.[18] It should be remarked that it is the over-all rate of change of pressure with time which is proportional to the square root of the azomethane pressure and the three-halves power of the ethylene pressure. When an attempt is made to correct for the increase of pressure produced by the decomposition of azomethane this simple relation seems to be upset.

Azomethane also causes propylene to polymerize. The reaction is, we believe, similar to the ethylene polymerization. Acetylene also polymerizes under the influence of methyl radicals from azomethane, the rate of pressure change being proportional to the square root of the azomethane pressure and to the first power of the acetylene pressure, a rather more usual result. These various reactions will be described in detail later.

Erratum. In Part II an error was made in the calculations for Run 217. For this run $\log P_i$ should be 0.295 and $\log (K \times 10^5)$ should be 0.691.

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(Received June 1, 1936)

The effect of a substituent upon the heat of ionization of an organic acid is far from identical with its effect upon the free energy of ionization. In the case of benzoic acid derivatives the entropy of ionization varies with structure in a way which can be predicted from the temperature dependence of the dielectric constant of the solvent, and which is closely correlated with the effect of a change in the dielectric constant of the solvent upon the relative strengths of substituted and unsubstituted acids. The behavior of aliphatic acids is more complicated, probably because of their less rigid structure. Similar considerations apply to the effect of changes in structure of reactant upon other equilibria and upon reaction rates.

THE effect of a change in structure of reactant upon the equilibrium or rate of an organic chemical reaction has been interpreted in a variety of ways.[1] Thus the shift in the ionization constant of formic acid produced by the substitution of another group for the nonionizable hydrogen has been attributed to an increase or decrease in the electrical work involved in separating the ions of the acid which results either from the dipole field of the substituent or from internal electron displacements of various sorts produced by the substitution. These interpretations have in common an identification

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[1] See for instance: Waters, *Physical Aspects of Organic Chemistry* (New York, 1936); or Ingold, Chem. Rev. **15**, 225 (1934).

TABLE I. *Reference acid formic acid.*

Acid RCOOH	K	ΔF	ΔH	ΔS	$-.005\Delta F$	ΔpK
Acetic	0 0990	+1372	−99	−4 94	−6 9	+0.13
Propionic	0 0754	+1533	−155	−5.67	−7.7	+0.40
Butyric	0.0855	+1460	−678	−7.17	−7.3	+0 37
Chloracetic	7.78	−1216	−1157	+0 20	+6.1	+0.27
Glycollic	0.833	+109	+223	+0.38	−0.55	—
Lactic	0.775	+147	−14	−0 64	−0.74	—

TABLE II. *Reference acid benzoic acid.*

Acid XC₆H₄COOH	K	ΔF	ΔH	ΔS	$-.005\Delta F$	ΔpK
m-NO₂ benzoic	5.24	−980	+350	+4 5	+4 9	−0.35
m-I benzoic	2.00	−410	+90	+1 7	+2.1	−0.22*
m-CH₃ benzoic	0.85	+100	+120	0.0	−0.5	+0 04
p-OCH₃ benzoic	0.50	+410	+110	−1.0	−2.1	+0.12
p-CH₃ benzoic	0.68	+230	+210	0.0	−1.2	+0.04
o-NO₂ benzoic	96.4	−2690	−2960	−0.9	+13.5	+0 27
o-OH benzoic	15.6	−1630	+670	+7.7	+8.2	−0.27
o-I benzoic	19.8	−1770	−2310	−1.8	+8.9	—
o-Cl benzoic	19 3	−1750	−2290	−1.8	+8 8	+0 28
o-CH₃ benzoic	1.93	−390	−1230	−2.8	+2 0	+0 28

* This is the value for m Cl-benzoic. The value for m-I benzoic should be nearly identical.

of the effect of the substituent upon the electrical work with its effect upon the free energy of the reaction, an identification whose theoretical correctness is not beyond question. It would be justified (except for the effect of possible differences in zero-point energy) were the effect of the substituent upon the free energy equal to its effect upon the heat of the reaction, as has been not infrequently supposed to be the case. Actually however, as the data in Tables I and II demonstrate, no such equality exists.

Table I contains under the heading K the equilibrium constants at 25°C for the series of reactions

$$RCOOH + HCOO^- \rightleftharpoons RCOO^- + HCOOH. \quad (1)$$

The quantity K is of course the ratio of the ionization constant of the acid named to that of the reference acid formic acid. Under the headings ΔF, ΔH, and ΔS are given the changes in free energy, in heat content, and in entropy for the same series of reactions, all at 25°C. These quantities measure the effect of the substitution of the group R for hydrogen upon the free energy, heat content, and entropy of ionization. The data in this table derive from the recent very accurate work of Harned and co-workers and of Nims and co-workers,[2] and their carefully fitted empirical equations for the temperature variation of the ionization constant were used in calculating the heats of ionization and entropies given in the table.

Table II contains similar information for the series of reactions

$$XC_6H_4COOH + C_6H_5COO^-$$
$$\rightleftharpoons XC_6H_4COO^- + C_6H_5COOH \quad (2)$$

and the quantities given represent therefore the

effect of a substituent in the benzene ring on the ionization of benzoic acid. The table is based upon the work of Schaller,[3] which, while it is certainly less precise than Harned's work, does represent measurements by a single careful investigator under constant conditions and in the same apparatus at approximately 10° intervals from room temperature to the boiling point. Under these conditions many of the errors which might affect the results as absolute measures of ionization constant or of its temperature coefficient vanish when, as in the present case, only the ratios of ionization constants are considered. Over the range and within the precision of these measurements the plot of log K against $1/T$ is linear, and the values of ΔH in the table were obtained graphically from such plots. The same relationship is nearly but not quite valid within the precision of Harned's work. It follows that the difference in the heats of ionization of two carboxylic acids is nearly independent of temperature, although the heat of ionization of a single acid varies rapidly with temperature.

These results are at first sight extremely discouraging for any attempt at theoretical interpretation. The entropy changes are large, equivalent to one or even two orders of magnitude in the ionization constant, and the heats of ionization show no relation whatsoever to the free energies. An "electronic theory" of the effect of substituents based upon the heats of reaction would be in striking contradiction to one based upon ionization constants, most notably in the case of m-nitro benzoic acid. Such large entropy changes might arise from the internal motions of the acids and the corresponding anions, for the

² Harned and Embree, J. Am. Chem. Soc. **56**, 1042 (1934); Harned and Ehlers, ibid. **54**, 1350 (1932); **55**, 652, 2379 (1933); Harned and Sutherland, ibid. **56**, 2039 (1934); Wright, ibid. **56**, 314 (1934); Nims, ibid. **58**, 987 (1936); Nims and Smith, J. Biol. Chem. **113**, 145 (1936).

³ Schaller, Zeits. f. physik. Chemie **25**, 497 (1898).

recent data of Kassel[4] shows that the reaction
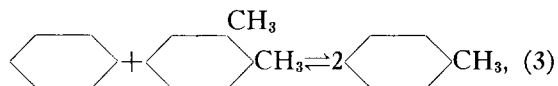
which has the same sort of symmetry as the reactions of Eqs. (1) and (2), is accompanied by an entropy change of 5.9 units in the gas state. There are however indications that, in the case of the rigid structures of the benzoic acid derivatives, the entropy change depends chiefly upon the solvent and can be accounted for in terms of the temperature dependence of its dielectric constant.

The hypothesis that the change in ionization constant produced by a substituent results from the effect of its dipole field upon the potential energy of the hydrogen ion has been suggested several times.[5] the equation which gives it expression is

$$\Delta F = e\Delta(\mu \cos \theta)/Dr^2, \qquad (4)$$

where ΔF has the same significance as in Tables I and II, $\Delta(\mu \cos \theta)$ is the difference between the components in the direction of the carboxyl group of the dipole moments of the $X-C$ and $H-C$ groups, r is the distance from the dipole to the carboxyl group, and D is the dielectric constant. It is not however a satisfactory equation. In the first place the quantity $\Delta(\mu \cos \theta)$ should be just equal to the total moment of the substance C_6H_5X in the case of a *para* substituted benzoic acid derivative in which the moment of the substituent is directed toward or away from the center of the benzene ring. But substitution of the moments of various benzene derivatives into Eq. (3) along with a value of 80 for D and a reasonable figure (6A) for r leads to values of ΔF for the corresponding *para* derivative of benzoic acid about ten times too large. Furthermore as Dippy and Watson[6] have shown the relation between dipole moment of C_6H_5X and free energy of ionization of XC_6H_4COOH is not a linear one, nor indeed even a unique correlation. This result is not surprising for a number of reasons, of which the most fundamental is the fact that the identification of an electrical

potential energy which involves a temperature dependent dielectric constant with a free energy change can at best give an approximate result. That it may be a useful approximation and even an accurate one for large distances appears from the success of the Debye-Hückel theory of electrolytes, in which the same identification is made.

Schwarzenbach[7] has attempted to account for dielectric saturation and for the transmission of the field through the solute molecule by a modified equation in which the actual dielectric constant possessed by the solvent in bulk is corrected and the corrected value is used in Eq. (4). He has verified some of the corollaries of the modified equation in a most ingenious fashion by comparing the effects of substitution upon the ionization of phenols and of thiophenols.

It is possible however to account for the entropies of ionization as well as for the effects observed by Schwarzenbach by a different modification

$$\Delta F = A/Dr^2. \qquad (5)$$

The quantity A is characteristic of a given substituent and position, is independent of temperature, and its magnitude is very approximately proportional to, but several times larger than, the value of $\Delta(\mu \cos \theta)$. The equation is of course a first-order empirical correction of the inadequate Eq. (4), which is to be justified by its results. It differs from Schwarzenbach's equation by putting the burden of the correction on the $\Delta(\mu \cos \theta)$ factor rather than on D.

Differentiating Eq. (5) with respect to temperature

$$\frac{\delta \Delta F}{\delta T} = -\frac{A}{D^2 r^2}\frac{\delta D}{\delta T} = -\Delta F \frac{\delta \ln D}{\delta T}$$

and

$$\Delta S = \Delta F \frac{\delta \ln D}{\delta T}. \qquad (6)$$

Since the quantity $\delta \ln D/\delta T$ has the value -0.005 for water this treatment predicts an entropy change equal to the free energy change times -0.005. The sixth column of Tables I and II shows a fairly satisfactory verification for the *meta* and *para* derivatives of benzoic acid, but a

[4] Kassel, J. Chem. Phys. **4**, 276 (1936).
[5] For a review of the literature see Wooten and Hammett, J. Am. Chem. Soc. **57**, 2289 (1935).
[6] Dippy and Watson, J. Chem. Soc. 436 (1936).

[7] Schwarzenbach and Egli, Helv. Chim. Acta **17**, 1184 (1934); Schwarzenbach and Epprecht, ibid. **19**, 493 (1936).

complete failure for the *ortho* derivatives of benzoic acid and for the aliphatic acids.

If now we suppose that a substituent may have a further effect upon the work of ionization which is independent of medium and temperature we obtain

$$\Delta F = A/Dr^2 + C, \tag{7}$$

from which we get

$$\Delta S = \frac{A}{r^2}\left(\frac{1}{D^2}\frac{\delta D}{\delta T}\right) \tag{8}$$

for the temperature effect and

$$\Delta F_1 - \Delta F_2 = \frac{A}{r^2}\left(\frac{1}{D_1} - \frac{1}{D_2}\right) \tag{9}$$

for the effect of a transfer from a medium of dielectric constant D_2 to one of constant D_1. Consequently

$$\frac{\Delta S}{\Delta F_1 - \Delta F_2} = \frac{D_2}{(D_2 - D_1)D_1}\frac{\delta D_1}{\delta T}, \tag{10}$$

and the ratio of the entropy change in one solvent to the difference in the free energy changes in two solvents of different dielectric constants should be independent of the substituent X.

The only extensive data for a test of this relation are contained in the work of Wooten and Hammett[8] on the relative ionization constants of various acids in the solvent butyl alcohol in the presence of a considerable concentration of neutral salt. The effect of the salt is to alter all ionization constants in much the same way that an increase in dielectric constant would. This follows from the fact that the clustering of alkali ions around the acid anions is determined by the same field which determines the potential energy of the hydrogen ion, and that the effect of the field upon the alkali ions must produce a change in free energy of ionization which is opposite in sign to that produced by the effect upon the hydrogen ions.

In general therefore it is to be expected that the linear relation between ΔF and $1/D$ discovered by Wynne-Jones,[9] which is entirely in agreement with Eq. (7), will be more valid the

lower the ion association. That is to say, it should be most satisfactory in solvents of relatively high dielectric constant and the ionization constant should in any case be the limiting value for zero ionic strength.

Because of the effect of the salt we cannot test Eq. (10) with respect to the numerical value of the right-hand side, since we cannot expect to use the actual value of the dielectric constant of butyl alcohol. We may nevertheless predict that the ratio of ΔS to the ΔpK of Wooten and Hammett, which equals $-(\Delta F_1 - \Delta F_2)/2.3RT$ will be independent of the substituent, negative in sign, and greater in magnitude than the value 1.9 calculated for $D_1 = 80$ and $D_2 = 17.4$.

As a comparison of columns 5 and 7 of Tables I and II show this prediction is fairly well verified (probably within the precision of Schaller's work) for *ortho* as well as for *meta* and *para* substituted benzoic acid derivatives, with the exception of salicylic acid, but is entirely invalid for the aliphatic acids. The obvious interpretation of this result is that the simple assumptions used which include a neglect of the internal motions of the reactant molecules are valid only for the rigid structures of the benzoic acid derivatives. The considerable deviation in the case of salicylic acid may well depend upon some specific interaction between the adjacent hydroxyl and carboxyl groups, perhaps of the nature of a hydrogen bond.[10]

The validity of Eq. (5) for the *meta* and *para* derivatives shows that the quantity C in Eq. (7) must be zero or nearly so for these substitutions, whereas it is certainly not zero for *ortho* substitutions. This contrast between the effects of close and distant substitutions is an extension to benzoic acid derivatives of a conclusion which Schwarzenbach and Egli[7] have previously reached from a study of phenols and thiophenols. There also it was found that *ortho* substituents exert large and specific effects which cannot possibly be attributed to dipole fields.

When we consider other reactions than the ionization of acids we find that Eq. (5) is insufficient even for *meta* and *para* substituents. This follows both from the fact that substituents

[8] Wooten and Hammett, J. Am. Chem. Soc. **57**, 2289 (1935).
[9] Wynne-Jones, Proc. Roy. Soc. **A140**, 440 (1933).

[10] See Branch and Yabroff, J. Am. Chem. Soc. **56**, 2568 (1934).

do have an effect upon reactions in which neither reactants nor products are electrically charged and also from the fact that the effect of a substituent upon one reaction may be many times greater than it is upon another. For instance the free energy of the process

$$XC_6H_4NH_3^+ + C_6H_5NH_2$$
$$\rightleftarrows XC_6H_4NH_2 + C_6H_5NH_3^+ \quad (11)$$
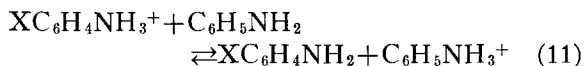
is three times as great for the same substituent X, provided this occupies the *para* or *meta* position, as it is for the reaction of Eq. (2).[11]

Since the distance between substituent and dissociable hydrogen does not differ enough in the two cases to account for this effect, the linear relation between the free energies of the two series of reactions requires an equation of the following form for the free energy:

$$\Delta F = \frac{A}{r^2}\left(\frac{1}{D} + B\right). \quad (12)$$

The quantity B is of the nature of a function of the polarizabilities of the linkages being made and broken in the reaction. It is independent of temperature and medium, and its magnitude must be of the order of twice that of the term $1/D$ in the case of the reaction of Eq. (11). It follows that while entropy and medium effect should have the same absolute value for reaction (11) as for reaction (2), they will represent considerably smaller fractions of the free energy in the former case.

Data are not available for a test of this prediction for the ionization of aniline derivatives, but it can be tested by studies of reaction rates and their temperature variation. It is now apparent from the work of Scheffer and

Brandsma,[12] of La Mer,[13] and of Eyring[14] that a reaction rate problem is essentially similar to an equilibrium one. In the rate problem the quantity $-RT \ln k_1/k_2$, where k_1 and k_2 are rate constants for the reaction of substituted and of unsubstituted reactant is equivalent to our ΔF; the difference in the Arrhenius activation energies for the two reactions is equivalent to our ΔH; and the difference in the entropies of activation to our ΔS. In the recently studied case of the alkaline hydrolysis of benzoic esters,[15] the quantity $-RT \ln k_1/k_2$ must include the electrical work involved in transferring a hydroxyl ion from the activated complex of hydroxyl ion with substituted ester to the complex of hydroxyl ion and unsubstituted ester. This electrical work must be very nearly the same as the electrical work involved in reaction (2), which as we have seen is practically identical with the free energy change of that reaction. Since, however, the quantity $-RT \ln k_1/k_2$ for the ester reaction is 2.5 times the ΔF for the ionization reaction (2),[16] the quantity B of Eq. (12) must equal 1.5 times $1/D$ for the hydrolysis reaction. It is therefore easy to see why, as Ingold and Nathan[15] have observed, the changes in entropy of activation are small for this series of reactions compared with the changes in free energy of activation, or that, what amounts to the same thing, the changes in the factor P of the English workers will be small compared with the changes in k. By contrast the reactions of alkyl halides studied by Winkler and Hinshelwood[17] present a situation much like that of the ionization of the aliphatic acids. Changes in the entropy of activation or in the P factor are relatively large, are unpredictable, and probably depend upon the internal motions of the reactants and of the activated complex as well as upon the solvent molecules.

[11] This is one of a large number of simple linear relations between the effects of substituents upon various reactions of benzene derivatives. Attention to some of these has previously been called by myself, Chem. Rev. **17**, 125 (1935), and by Burkhardt, Ford, and Singleton, J. Chem. Soc. 17 (1936). This subject will be discussed in detail in another paper.

[12] Scheffer and Brandsma, Rec. trav. chim. **45**, 522 (1926); Brandsma, ibid. **47**, 94 (1928); **48**, 1205 (1929).
[13] La Mer, J. Chem. Phys. **1**, 289 (1933).
[14] Eyring, J. Chem. Phys. **3**, 107 (1935).
[15] Ingold and Nathan, J. Chem. Soc. 222 (1936).
[16] Hammett, Chem. Rev. **17**, 125 (1935).
[17] Winkler and Hinshelwood, J. Chem. Soc. 1147 (1935).