water was always run side-by-side with each series of salt solutions.

The saturated solutions were analyzed as follows: 1-naphthoic acid, by acid-base titration, using phenolphthalein as indicator. Potassium acid phthalate obtained from the U. S. Bureau of Standards served as primary acidimetric standard.

Naphthalene, by measurement of optical density at 3110 Å. with a Beckman model DU ultraviolet spectrophotometer, after 1:20 dilution of the saturated solutions with 50% dioxane–water. Using 1-cm. cells and a constant slit width of 0.7 mm. the optical densities were in the neighborhood of 0.5 and could be measured with a precision of better than 0.4%. Actual concentrations were obtained from the optical densities by interpolation on a large scale calibration curve, which was constructed using solutions of known concentration and identical instrument settings. (Under our experimental conditions there was some slight but significant deviation from Beer's law.)

The optical density of each undiluted salt solution was measured using 50% dioxane as a blank. No appreciable absorption was found with any of the salts in the region around 3110 Å., except with sodium 2-naphthalenesulfonate. This salt absorbed so strongly that it was not feasible to determine the naphthalene concentration in its presence.

Owing to the rapid change of solubility with solvent composition (Table IV), great care was exercised to minimize changes in composition due to volatilization during preparation and handling of the solutions. (The vapor contains about 83% dioxane.) We believe, on the basis of the good precision of a large number of "controls," that we have succeeded in repressing this source of error. The over-all accuracy of S_3^0/S_3 is believed to be better than 0.5%.

Density Measurements.—Density measurements were made by William B. Brogden, Jr., using a set of bicapillary pycnometers made to specification by the Daylight Vacuum Products Co., Chicago, Ill. The pycnometers were made in the form of flat-bottomed bulbs with a pair of precision capillaries about 10 cm. long attached to the upper surface of the bulb. These capillary tubes were graduated along the middle 6 cm. of their length, and terminated at the top with a pair of standard taper joints. The pycnometers were filled by gravity flow from an auxiliary reservoir designed in such a way that the solutions could be out-gassed prior to filling. Out-gassing was done in the usual way by repeated freezing, evacuation and thawing under vacuum. The volume per unit length of each capillary was obtained by measuring the length of a weighed drop of mercury, and the bulb volume of each pycnometer was then determined by filling with water and weighing. In each series of measurements, the density of 50.00 wt. % dioxane–water was measured as a "control" along with that of several salt solutions. From the reproducibility of a number of "controls" we estimate the precision of our results as 0.003%.

[Contribution from the Chemistry Department of Florida State University, Tallahassee, Fla.]

Salt-induced Medium Effects. II. Kinetic Salt Effects in the Solvolysis of Neophyl *p*-Toluenesulfonate and in the Racemization of L-(+)-*threo*-3-Phenyl-2-butyl *p*-Toluenesulfonate in 50 Wt. % Dioxane–Water[1a]

By E. F. J. Duynstee, Ernest Grunwald[1b] and Martin L. Kaplan

Received April 14, 1960

In order to resolve the problem whether kinetic salt-effects in the ionization step of SN1 reactions are colligative, rate constants, k_1, were measured for the solvolysis of neophyl *p*-toluenesulfonate and for the racemization of L-(+)-*threo*-3-phenyl-2-butyl *p*-toluenesulfonate in 50 wt. % dioxane–water in the presence of the following salts: NaOH, NaCl, LiCl, NaBr, $NaNO_3$, NaI, $NaClO_4$, 2-$C_{10}H_7SO_3Na$, $(CH_3)_4NCl$, $(C_6H_5)_4PCl$. The salt effects were highly specific; values of the parameter, S, in the empirical equation: $\log (k_1/k_1^0) = Sm_4$ (where m_4 is the salt molality) ranged from *ca.* +0.5 for $NaClO_4$ to *ca.* −0.6 for NaOH. Comparison of k_1 with rate constants for direct second-order reactions with anions indicated that such reactions could not account for the observed specificity of salt-effects on k_1. For the solvolysis of neophyl *p*-toluenesulfonate, such parallel reactions should be insignificant; for *threo*-3-phenyl-2-butyl *p*-toluenesulfonate, newly measured rate constants for the substitution reaction of this compound and of ethyl *p*-toluenesulfonate with lithium bromide in anhydrous acetone indicated that direct reaction with the anion might possibly make a small contribution to the observed salt effect on the racemization rate when the salt is NaOH, NaBr and NaI. According to Grunwald and Butler's semi-empirical treatment of salt-effects on the activity coefficients of non-electrolytes,[2] if the variation of S for different salts were due to long-range electrostatic and salt-induced medium effects, one would expect eq. 12, where Z_1 is the mole-fraction of water in the solvent, F_{4m}^0 is the standard partial molal free energy of the salt, and A and B are adjustable constants, to apply. This equation was found to fit the effect of all simple inorganic salts on the solvolysis of neophyl *p*-toluenesulfonate, and of all inorganic salts except NaBr and NaI on the racemization of *threo*-3-phenyl-2-butyl *p*-toluenesulfonate. Significant deviations from this equation were observed for all organic salts, suggesting that short-range interaction between organic ions and organic non-electrolytes is significant.

In the preceding study of salt effects on the activity coefficients of non-electrolytes,[2] we found that electrostatic models alone could not accommodate the facts, but that a combination of electrostatic effects and "salt-induced medium effects" could account for the observations involving simple inorganic salts. We now use the equations developed in that connection to resolve one of the classical problems in the kinetics of SN1 substitution of organic compounds (RX): whether or not the effect of salts at low concentrations on the rate constant, k_1, for the initial ionization of RX is a colligative function of the ionic strength.

According to the transition-state theory, the effect of salts on k_1 is formally given by eq. 1

$$k_1 = k_1^0 \gamma_{RX}/\gamma_{\ddagger} \qquad (1)$$

where k_1^0 denotes the rate constant at infinite dilution, and γ_{RX} and $\gamma_{\ddagger}$ denote the molal activity coefficients of RX and of the transition state complex with respect to their infinitely dilute reference states. In a highly influential paper published in 1940, Bateman, Church, Hughes, Ingold and Taher[3] developed a theory according to which most of the

---

[1] (a) Work supported by Office of Naval Research. Reproduction in whole or in part is permitted for any purpose of the United States government. (b) Alfred P. Sloan fellow, 1959.

[2] E. Grunwald and A. F. Butler, This Journal, 82, 5647 (1960).

[3] L. C. Bateman, M. G. Church, E. D. Hughes, C. K. Ingold and N. A. Taher, *J. Chem. Soc.*, 979 (1940).

interactions leading to deviations from ideal behavior in the presence of salts are so similar for RX and for the transition state complex that their effect on the *ratio*, $\gamma_{RX}/\gamma_{\ddagger}$, may be neglected. The only non-negligible interaction, according to this theory, is that of the highly polar transition state complex with its ionic atmosphere. The effect of this interaction on $\gamma_{\ddagger}$, and hence on k_1, was calculated, the final expression being eq. 2.

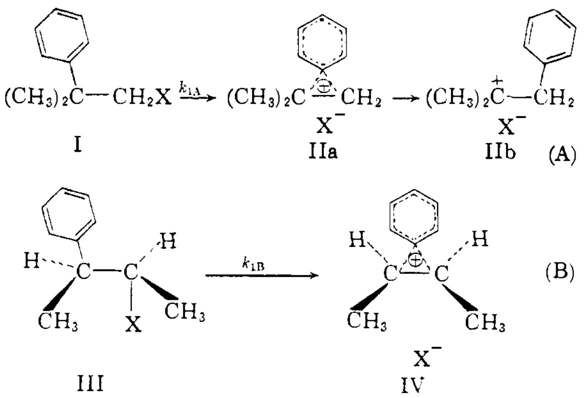
$$\log (k_1/k_1^0) = 0.912 \times 10^{16}\, \sigma\mu/\epsilon^2 T^2 \qquad (2)$$

In this equation, μ is the ionic strength, ϵ the dielectric constant and $\sigma = Z^2 d$ is a measure of the charge separation in the transition state complex, the latter being regarded as a dipole with charges $\pm Ze$ separated by a distance, d.

According to this theory, deviations of the experimental "rate constants" from eq. 2 must be explained in terms of active participation of the salt in the reaction. However, Lucas and Hammett,[4] studying the solvolysis of *t*-butyl nitrate in dioxane–water mixtures, shortly thereafter proposed that non-colligative salt-induced medium effects might also be important. These authors found that for a number of salts the kinetic effect closely paralleled the effect of the salt on the water activity of the solution. For example, when 0.1 *M* NaOH was added, the water activity decreased, and so did the rate, while addition of 0.1 *M* $NaClO_4$ resulted in an increase both of the water activity and of the rate. Moreover, in the absence of salt an increase in the water activity (caused by adding water to the solvent) similarly resulted in an acceleration.[4]

Unfortunately, it is difficult to prove that the measured rate constant for the solvolysis of *t*-butyl nitrate is equal to that for the initial ionization. Thus Benfey, Hughes and Ingold[5] suggested that the high specificity of salt effects observed in this reaction[4] be explained on the basis of direct participation of certain salts *subsequent* to the ionization step, and that eq. 2 is still correct for the ionization itself. Since then, a number of kinetic investigations, including some of the highest precision, of SN1 reactions have provided evidence both for[6] and against[7,8] this view.

Reactions Chosen for Study.—In the two decades since the start of this controversy, a number of reactions have been studied for which it is quite certain that the measured rate constant is equal to that for the ionization of RX. We now report kinetic salt effects in 50 wt. % dioxane–water for two such reactions: (A) the solvolysis of neophyl *p*-toluenesulfonate,[7,10a] which we have followed by measuring either the production of acid or the disappearance of RX; (B) the racemization of L-(+)-*threo*-3-phenyl-2-butyl *p*-toluenesulfonate.[9,10] Probable mechanisms for the ionization step in these processes are



The evidence concerning the mechanism of reaction A has been summarized by Fainberg and Winstein.[7] Concerning reaction B, the racemization is due mainly to the ionization of III to the optically inactive intimate ion pair IV, the process being assisted by participation of the adjacent phenyl group.[9,10] A careful analysis of the products obtained in glacial acetic acid has shown, however, that the formation of IV is accompanied by a small percentage of SN1 solvolysis without phenyl participation, and by some solvolysis with participation of the adjacent hydrogen.[11] Most of the products obtained in these side-reactions are optically inactive,[9b,11] their total optical activity being less than 1% of that of the starting material III.[10a] Moreover, salt effects on these side reactions should parallel those on the major reaction, the conversion of III to IV. Following the example of previous workers,[10] we shall therefore neglect the existence of these side reactions.

Some of the anions used in this study, such as hydroxide and iodide, have considerable nucleophilic character,[12] and we were concerned lest there be a contribution to the measured rate constants by reactions proceeding through SN2 or E2 mechanisms. We shall return to this problem in a later section, but indicate here that such reactions are apparently negligible in our systems, except possibly for the reaction of sodium bromide and sodium iodide with *threo*-3-phenyl-2-butyl *p*-toluenesulfonate.

Results

Reaction rates were followed for at least two half-lives and proceeded according to first-order rate equations. The standard errors of the rate constants were within ±2%, except for a few of the racemization runs, indicated in Table III, for which the errors may have been as large as ±5%. Data for typical runs are shown in the Experimental Part.

The sensitivity of the rate constants to change in the solvent composition is shown by the data in Table I. The magnitude of the medium effect is similar to that observed in other solvolysis reactions of the SN1 type, the slope *m* in the *mY* equa-

---

[4] G. R. Lucas and L. P. Hammett, THIS JOURNAL, **64**, 1928 (1942).

[5] O. T. Benfey, E. D. Hughes and C. K. Ingold, *J. Chem. Soc.*, 2488, 2494 (1952).

[6] For example: B. Bensley and G. Kohnstam, *ibid.*, 3408 (1955).

[7] A. H. Fainberg and S. Winstein, THIS JOURNAL, **78**, 2763 (1956).

[8] F. Spieth and A. R. Olson, *ibid.*, **77**, 1412 (1955).

[9] (a) D. J. Cram, *ibid.*, **71**, 3863 (1949); (b) **74**, 2129 (1952).

[10] (a) S. Winstein, B. K. Morse, E. Grunwald, K. C. Schreiber and J. Corse, *ibid.*, **74**, 1113 (1952); (b) S. Winstein and K. C. Schreiber, *ibid.*, **74**, 2165 (1952); (c) A. H. Fainberg and S. Winstein, *ibid.*, **78**, 2780 (1956); (d) S. Winstein and G. C. Robinson, *ibid.*, **80**, 169 (1958).

[11] D. J. Cram, *ibid.*, **74**, 2137 (1952).

[12] C. G. Swain and C. B. Scott, *ibid.*, **75**, 141 (1953).

tion[13] being 0.678 for the solvolysis (reaction A), and 0.869 for the racemization (reaction B).

The production of acid accompanying the racemization of III was measured also. The rate followed first-order kinetics throughout the range of observation (two half-lives), and the rate constant in 50 wt. % dioxane–water was 0.614 times that for the racemization (data in Table I). The greater value for the rate of racemization is consistent with previous kinetic data in other solvents[10] and supports our assumption of a similar reaction mechanism.

TABLE I

EFFECT OF SOLVENT COMPOSITION AND TEMPERATURE ON THE FIRST-ORDER RATE CONSTANT FOR THE SOLVOLYSIS OF NEOPHYL *p*-TOLUENESULFONATE AND THE RACEMIZATION OF L-(+)-*threo*-3-PHENYL-2-BUTYL *p*-TOLUENESULFONATE IN DIOXANE–WATER MIXTURES

Water in solvent				
Wt. %	Mole fraction (Z_1)	Temp., °C.	$10^6 k_1^0$, sec.$^{-1}$	$\frac{\delta \ln k_1^0}{\delta Z_1}$
(A) Solvolysis of neophyl *p*-toluenesulfonate				
44.74	0.798	50.00	4.11	16.7
50.00	.830	50.00	7.01	16.5
54.77	.855	50.00	10.6	
50.00	.830	25.00	0.267	..[a]
50.00	.830	84.0	269	..
			Av.	16.6
(B) Racemization of L-(+)-*threo*-3-phenyl-2-butyl *p*-toluenesulfonate				
45.01	0.8001	25.00	2.70	20.2
50.00	.8302	25.00	4.95	23.4
52.28	.8426	25.00	6.62	
			Weighted av.	21.1
50.00	0.8302	25.00	(3.04)[b]	..

[a] Arrhenius activation energy, 24.1 kcal. [b] Solvolysis rate constant, measured by following the production of acid.

The data for the kinetic salt effects are listed in Tables II and III. The rate measurements for each of the salt solutions were done side-by-side with the measurement of a "blank" in the same solvent without salt, and k_1/k_1^0 is the ratio of the rate constant for the salt solution to that of the accompanying "blank." The tables show the high specificity of the effect of individual salts, which ranges all the way from considerable acceleration to considerable retardation.

Plots of log (k_1/k_1^0) *vs.* the salt molality, m_4, were linear within experimental error for both reactions. The data were therefore fitted to equations of the form

$$\log (k_1/k_1^0) = Sm_4 \qquad (3)$$

where S is a parameter characteristic of the salt and the reaction. Values of S were computed by the method of least squares and are included in Tables II and III. It must be remarked, however, that the experimental error of log (k_1/k_1^0), about ± 0.008 unit, is fairly large relative to the absolute values of this quantity. Thus, if there had been moderate curvature in the plots of log (k_1/k_1^0) *vs.* m_4, we could not have detected it.

TABLE II

KINETIC SALT EFFECTS IN THE SOLVOLYSIS OF NEOPHYL *p*-TOLUENESULFONATE IN 50.00 WT. % DIOXANE–WATER AT 50.00°

Salt	m_4	k_1/k_1^0	S	dF_{4m}^0/dZ_1 kcal.
NaOH[a]	0.092	0.863	−0.68	−20.4
	.205	.727		
NaCl[b]	.108	.976	− .15	−16.6
	.206	.923		
LiCl[b]	.101	1.004	+ .02	−14.3
	.189	1.007		
NaBr[b]	.103	1.048	+ .13	−14.1
	.203	1.053		
$NaNO_3$[b]	.100	1.065	+ .19	−12.6
	.198	1.082		
NaI[a]	.097	1.062	+ .26	−11.3
	.199	1.123		
$NaClO_4$[b]	.098	1.157	+ .59	− 6.8
	.196	1.300		
2-$C_{10}H_7SO_3Na$[b]	.095	1.026	+ .20	− 7.3
	.191	1.103		

[a] Measured by following unreacted RX. [b] Measured by following production of acid.

TABLE III

KINETIC SALT EFFECTS IN THE RACEMIZATION OF L-(+)-*threo*-3-PHENYL-2-BUTYL *p*-TOLUENESULFONATE IN 50.00 WT. % DIOXANE–WATER AT 25.00°

Salt	m_4	k_1/k_1^0	S	dF_{4m}^0/dZ_1, kcal.
NaOH	0.193	0.798	−0.51	−20.4
NaCl	.097	.980[a]	− .19	−16.6
	.193	.907		
$NaNO_3$	.195	1.075	+ .16	−12.6
NaBr	.102	1.090[a]	+ .26	−14.1
	.202	1.112		
$NaClO_4$	.097	1.062[a]	+ .42	− 6.8
	.190	1.223		
NaI	.094	1.238[a]	+ .77	−11.3
	.194	1.376		
2-$C_{10}H_7SO_3Na$	.201	1.036	+ .08	− 7.3
$(CH_3)_4NCl$	.196	1.029	+ .06	−16.3
$(C_6H_5)_4PCl$	.204	0.914	− .19	− 2.6

[a] The standard error for this value is greater than ±2% and may be as large as ±5%.

Tables II and III also list experimental values[14] of dF_{4m}^0/dZ_1, where F_{4m}^0 is the standard partial molal free energy of the salt (on the *molal* concentration scale), and Z_1 is the mole-fraction of water in the solvent.

It remains to discuss to what extent the added salts participate in the over-all reactions. If participation involves reaction of the anions of the salt with IIa, IIb or IV, it will have no effect on the measured rate constants. Reaction of IIa or IIb with an anion, Y^-, will lead to the tertiary structure, $(CH_3)_2 \cdot CY \cdot CH_2 \cdot C_6H_5$, which solvolyzes so rapidly that the production of acid is not appreciably delayed by the formation of this intermediate.[7] Reaction of the optically inactive IV will, of course, have no effect on the rate of race-

---

(13) (a) E. Grunwald and S. Winstein, THIS JOURNAL, **70**, 846 (1948); (b) A. H. Fainberg and S. Winstein, *ibid.*, **78**, 2770 (1956).

(14) E. Grunwald, G. Baughman and G. Kohnstam, *ibid.*, **82**, Nov. 29 (1960).

mization. That such reaction is not negligible, however, is suggested by the data in Table IV. Reaction mixtures containing salt, and 0.01 molal in III, were heated at 80° for several hours, during which time the solvolysis of the original tosylate was virtually complete. Although the final optical rotations did not differ significantly from zero, the acid titers did differ significantly from those of the blanks. Presumably the production of acid has been retarded here by the formation of reactive intermediates of the general formula, RY.

TABLE IV

SOLVOLYSIS OF L-(+)-*threo*-3-PHENYL-2-BUTYL *p*-TOLUENESULFONATE AT 80.0°; INITIAL RX, 0.01 *m*

Time, hr.	Salt, 0.2 *m*	Acid titer, % of RX	α[a]
4.5	..	99.8	−0.001
7	..	98.5	.....
9	..	98.3	+ .004
4.5	NaCl	93.8	+ .003
7	$(CH_3)_4NCl$	92.9	+ .002
7	$(C_6H_5)_4PCl$	89.2	− .007
9	$NaNO_3$	94.7	+ .004
7	$2\text{-}C_{10}H_7SO_3Na$	95.5	+ .007
9	NaOH	..	+ .001

[a] Optical rotation. Initial values of α are about +0.25°.

TABLE V

SECOND-ORDER RATE CONSTANTS FOR REACTION OF LITHIUM BROMIDE WITH VARIOUS SUBSTRATES IN ANHYDROUS ACETONE

Substrate	Reaction	Temp., °C.	10^4k_2, sec.$^{-1}$ M^{-1}
CH_3CH_2OTs	Substitution[a]	77.8	738
		25.0	5.37[b]
$C_6H_5CH_3CHCHCH_3OTs$	Substitution	77.8	5.1[c]
(*threo*)	Elimination	77.8	15.2[c]

[a] The amount of elimination occurring under these conditions is undetectably small. [b] Arrhenius activation energy is 19.34 kcal. [c] Based on a single run with nearly constant LiBr concentration; computed on the assumption that reaction is first order in LiBr and second order over-all.

In addition, there might be direct reaction between Y^- and RX by SN2 or (in the case of III) E2 mechanisms. Several observations suggest that this is not important: (i) the over-all reactions proceed with good first-order kinetics up to several half-lives even in the presence of NaI and $2\text{-}C_{10}H_7SO_3Na$, where the product of an SN2 reaction would solvolyze at an appreciable rate. (ii) The rate in the presence of 0.2 *m* NaOH is slower than that in the presence of any other salt at 0.2 *m* concentration, even though hydroxide ion is one of the most effective agents for promoting SN2 or E2 reaction. (iii) In the racemization of III, the final optical rotations are virtually zero. SN2 reactions of the L-(+)-*threo*-*p*-toluenesulfonate should lead to optically active products with the L-*erythro* configuration.[15] Previous data suggest that their rotations are not negligibly small.[9b]

In analyzing this problem further, we made use of available rate constants for the SN2 reaction of ethyl *p*-toluenesulfonate with Cl^-, Br^-, I^- and OH^- in 60% dioxane–water.[16] This solvent is sufficiently like ours for the present purpose. We define a factor, *f*, by (4)

$$k_2^{RX} = f \times k_2^{EtOTs} \quad (4)$$

where k_2 is in each instance the rate constant for the direct reaction with Y^-. We further specify that such reaction is negligible in relation to the experimental error of k_1 if

$$k_2^{RX} \times (Y^-) = f \times k_2^{EtOTs} \times (Y^-) < 0.01\, k_1 \quad (5)$$

Upon using the known values of k_1, k_2^{EtOTs} and 0.2 *m* for (Y^-), we find that the maximum value of *f* for which the inequality (5) is still correct is $^1/_{1000}$ for the solvolysis of neophyl *p*-toluenesulfonate, and $^1/_{115}$ for the racemization of III. In other words, if the actual k_2^{RX} values are smaller than k_2^{EtOTs} by at least these factors, their contribution to the experimental rate constants may be neglected. Data for the SN1 reactivity of neophyl compounds are apparently not available, but data for the sterically similar neopentyl halides have been obtained in anhydrous acetone, and the factors, *f*, relative to the corresponding ethyl halides are close to $^1/_{5000}$.[17] This is sufficiently smaller than the upper limit of $^1/_{1000}$ obtained above to leave a comfortable margin of safety.

In order to estimate *f* for *threo*-3-phenyl-2-butyl *p*-toluenesulfonate, we have measured rate constants for the reaction of lithium bromide with III and with ethyl *p*-toluenesulfonate in anhydrous acetone. The results are shown in Table V. Unfortunately, the formation of RBr from III was accompanied by a parallel reaction, the elimination of *p*-toluenesulfonic acid from III. We believe that the latter proceeds by an E1 mechanism, the initial ionization being promoted by LiBr. For this reason the formation of RBr may also proceed by an ionization mechanism rather than by SN2, at least in part, so that the experimental rate constant is an upper limit to the SN2 reactivity. At any rate, the calculated value of *f* is $^1/_{145}$. This is just barely less than the upper limit of $^1/_{115}$ estimated above. Allowing for possible solvent effects on *f* between acetone and 50% dioxane–water, we conclude that there might be (but probably isn't) a contribution to the rate constant of racemization of III of several per cent. due to direct reactions with I^- and OH^-, the most reactive of the anions, and a smaller but possibly still significant percentage with Br^-.

Discussion

We now apply the semi-theoretical equations obtained in the preceding paper[2] in order to interpret the observed pattern of kinetic salt effects. The symbols we employ are consistent with those defined in the preceding paper.[2] We shall assume that γ_{RX} and $\gamma_\ddagger$ in eq. 1 are given as functions of salt molality by equations of the type 6

$$\log \gamma_3 = k_{34}m_4 \quad (6)$$

where k_{34} is the appropriate salt–non-electrolyte interaction constant. It was found[2] that, if short-range interaction between salt and non-electrolyte can be neglected, the sum total of electrostatic and salt-induced medium effects is given by eq. 7

---

[15] Cram's system IIA.[9b] For assignment of configuration, see D. J. Cram, THIS JOURNAL, **74**, 2149 (1952).

[16] H. R. McCleary and L. P. Hammett, *ibid.*, **63**, 2254 (1941).

[17] P. B. D. de la Mare, L. Fowden, E. D. Hughes, C. K. Ingold and J. D. H. Mackie, *J. Chem. Soc.*, 3200 (1955).

$$k_{34} = A_3 + B_3 \frac{\mathrm{d}F_{3\mathrm{m}}{}^0}{\mathrm{d}Z_1} \frac{\mathrm{d}F_{4\mathrm{m}}{}^0}{\mathrm{d}Z_1} \qquad (7)$$

where A_3 and B_3 are constants characteristic of the solvent and the non-electrolyte. Equation 7 was found to fit the data for simple inorganic salts, but not for salts with large organic ions such as $(CH_3)_4NCl$ or $2\text{-}C_{10}H_7SO_3Na$.[2] This was taken to indicate that specific short-range interaction is not negligible for salts with large organic ions.

Upon applying eq. 7 to RX and its transition-state complex, we obtain from 1 and 6

$$\log(k_1/k_1{}^\circ) = \left[A_{\mathrm{RX}} - A_{\ddagger} + \frac{\mathrm{d}F_{4\mathrm{m}}{}^0}{\mathrm{d}Z_1}\left(B_{\mathrm{RX}}\frac{\mathrm{d}F_{\mathrm{RX}}{}^0}{\mathrm{d}Z_1} - B_{\ddagger}\frac{\mathrm{d}F_{\ddagger}{}^0}{\mathrm{d}Z_1}\right)\right] m_4 \qquad (8)$$

On comparing eq. 8 with the phenomenological eq. 3, we find that the term in brackets is equal to the experimental slope, S. Thus eq. 8 predicts a linear relationship between S and $\mathrm{d}F_{4\mathrm{m}}{}^0/\mathrm{d}Z_1$.

Before testing this prediction, however, we shall transform eq. 8 into a more convenient form. We define

$$A \equiv A_{\mathrm{RX}} - A_{\ddagger} \qquad (9)$$

$$B \equiv \left(B_{\mathrm{RX}}\frac{\mathrm{d}F_{\mathrm{RX}}{}^0}{\mathrm{d}Z_1} - B_{\ddagger}\frac{\mathrm{d}F_{\ddagger}{}^0}{\mathrm{d}Z_1}\right)\Big/\left(\frac{\mathrm{d}F_{\mathrm{RX}}{}^0}{\mathrm{d}Z_1} - \frac{\mathrm{d}F_{\ddagger}{}^0}{\mathrm{d}Z_1}\right) \qquad (10)$$

Furthermore

$$\frac{\mathrm{d}F_{\mathrm{RX}}{}^0}{\mathrm{d}Z_1} - \frac{\mathrm{d}F_{\ddagger}{}^0}{\mathrm{d}Z_1} = -\frac{\mathrm{d}(\Delta F^0)^{\ddagger}}{\mathrm{d}Z_1} = RT\frac{\mathrm{d}\ln k_1{}^0}{\mathrm{d}Z_1} \qquad (11)$$

Thus we obtain

$$S = A + BRT\frac{\mathrm{d}\ln k_1{}^0}{\mathrm{d}Z_1}\frac{\mathrm{d}F_{4\mathrm{m}}{}^0}{\mathrm{d}Z_1} \qquad (12)$$

where A and B are parameters characteristic of the reactant, transition state and the solvent.

The observed salt effects on the solvolysis of neophyl *p*-toluenesulfonate are beautifully consistent with the pattern observed for the activity coefficients of non-electrolytes.[2] Experimental values of S for simple inorganic salts are plotted *vs.* $\mathrm{d}F_{4\mathrm{m}}{}^\circ/\mathrm{d}Z_1$ in Fig. 1A. The hoped-for linear relationship is indeed observed, within experimental error. On the other hand, the point for the organic salt, $2\text{-}C_{10}H_7SO_3Na$ (not shown in the figure), lies far off the line. Since it is clear from independent evidence[7] (see preceding section) that k_1 measures the rate of ionization of neophyl *p*-toluenesulfonate (I → II), we conclude that (i) contrary to simple electrostatic models, salt effects on the rate constant for the ionization of RX are highly specific; (ii) in favorable cases (*e.g.*, certain simple inorganic salts) the specificity can be ascribed largely to the operation of salt-induced medium effects; (iii) in other cases (*e.g.*, salts with large organic ions) specific short-range interaction is important also.

The data for the racemization of 3-phenyl-2-butyl *p*-toluenesulfonate are less easy to interpret. Values of S are plotted *vs.* $\mathrm{d}F_{4\mathrm{m}}{}^0/\mathrm{d}Z_1$ in Fig. 1B. Two different straight lines can reasonably be drawn: the solid line through the points for NaOH, NaCl, $NaNO_3$ and $NaClO_4$, and the dotted line through the points for NaCl, NaBr and NaI. We believe that one of these lines represents the simple case in which specific short-range interaction is negligible. The question is, which one? In view of the possibility that direct reaction of RX with Y^- may make a slight contribution to the rate constant for racemization, and because such reaction would probably be more important for NaI than for any of the other salts, we have chosen the solid line as representing the simple case. On this basis, specific short-range interaction (which may but need not be such as to lead to SN2 reaction) causes a small positive deviation for NaBr and a somewhat larger positive deviation for NaI. However, regardless of whether we choose the solid line or the dotted line to represent the simple case, the data for the salts with large organic ions deviate widely. These observations again suggest that specific short-range interaction between large organic ions and organic non-electrolytes is important and, in view of the consistency of such observations, that it is a general phenomenon.

On the basis of the solid lines drawn in Fig. 1, the parameters A and B in eq. 12 have the following values for 50 wt. % dioxane–water: solvolysis of neophyl *p*-toluenesulfonate, 50°: $A = 1.30$, $B = 0.0085$ racemization of *threo*-3-phenyl-2-butyl *p*-toluenesulfonate, 25°: $A = 0.94$, $B = 0.0055$. The values of B are somewhat larger than, but of the same order of magnitude as, the analogous parameters, B_3, in the expression for k_{34} for naphthalene and 1-naphthoic acid.[2] The intercepts, A, are about twice the magnitude to be expected for the long-range colligative electrostatic salt effect as calculated from eq. 2. Using the plausible values of 0.5 for Z and 2.8×10^{-8} cm. for d, the value of σ for the transition state complex in our reactions is found to be 0.7×10^{-8}. On substituting this value and numerical data for 50 wt. % dioxane–water (ϵ 35.85 at 25°, 31.11 at 50°) in eq. 2, the expected contribution to S from these interactions is a constant term, independent of the salt, amounting to 0.63 for the solvolysis and 0.56 for the racemization.

In conclusion, our data lend strong support to previous indications[7,8] that eq. 2 is not adequate for describing kinetic salt effects on the ionization step in solvolysis reactions. Our data show not only that these salt effects are *not* colligative, but also that they need not even have the same sign for different salts. A change in the sign of the salt effect can result from salt-induced medium effects (for example, $NaClO_4$ *vs.* NaOH), or from specific short-range interaction (for example, $(CH_3)_4NCl$ *vs.* $(C_6H_5)_4PCl$ in the racemization reaction).

Experimental Part

Neophyl *p*-Toluenesulfonate.—Neophyl alcohol was prepared by the procedure of Whitmore, Weisberger and Shabica[18]; b.p. 130–132° (30 mm.), $n^{20}\mathrm{D}$ 1.5260. To 1 g. of this alcohol in 2 ml. of anhydrous pyridine was added 1.27 g. of *p*-toluenesulfonyl chloride, and the mixture allowed to stand overnight at room temperature. The reaction mixture was then poured into 30 ml. of cold 2*N* H_2SO_4, and the solid mass which separated was extracted with a 2:1 mixture of ether and petroleum ether. After washing and drying of the organic layer, the solvent was removed at room temperature, and the residue of crude crystalline neophyl *p*-toluenesulfonate was recrystallized twice from benzene–petroleum ether (1:5). The pure product (m.p. 75–76°) was stored in a refrigerator until

---

[18] F. C. Whitmore, C. A. Weisberger and A. C. Shabica, THIS JOURNAL, 65, 1469 (1943).

used. The equivalent weight, as determined by titration for acid after solvolysis, was 310 (calcd. 304).

L-(+)-*threo*-**3-Phenyl-2-butyl *p*-Toluenesulfonate.**—The starting material was α-phenylpropionaldehyde, practical grade, obtained from Brothers Chemical Co. The commercial product was vacuum distilled, b.p. 65–66° (2.5 mm.), n^{21}D 1.5198—a large residue of viscous material being discarded.

The purified aldehyde was converted to 3-phenyl-2-butanol by the following method: 102 g. (1.07 moles) of methyl bromide was bubbled into a mixture of 30 g. of magnesium and 160 ml. of dry ether in a 500-ml. 3-neck round-bottom flask, equipped with stirrer and condenser, and the Grignard reagent was prepared. A solution of 105 g. of α-phenylpropionaldehyde (0.80 mole) in 140 ml. of ether was then added dropwise to the Grignard solution, and the addition product was decomposed by pouring the ether solution into a mixture consisting of 80 g. of ammonium chloride, 370 ml. of ice-water and 500 ml. of ether. The ether layer was separated, the water layer was extracted twice with *n*-pentane–ether, and the combined organic layers were dried over anhydrous potassium carbonate. After removal of the solvent, vacuum distillation yielded 92 g. (76%) of the racemates of *threo*- and *erythro*-3-phenyl-2-butanol (b.p. 84–88° (2.5 mm.), n^{21}D 1.5200).

L-(+)-*threo*-3-Phenyl-2-butyl *p*-toluenesulfonate was prepared from the mixture of isomeric 3-phenyl-2-butanols according to the method of Cram.[9a] In Cram's original paper,[9a] the L-(+)-*threo* isomer was designated as isomer IA, but the configuration of this isomer has since been established.[15]

Salts.—The preparation or purification of **sodium chloride, sodium perchlorate, tetramethylammonium chloride** and **sodium 2-naphthalenesulfonate** followed previous practice.[3]

Tetraphenylphosphonium chloride was prepared according to the method of Willard, Perkins and Blicke.[19] The salt is hygroscopic and was dried in a vacuum oven at 80° just before use; equivalent weight by Volhard titration for chloride: 374.2 (obsd.), 374.8 (calcd.). **Sodium nitrate** (reagent grade) was recrystallized twice from water and dried *in vacuo*. **Sodium bromide** (reagent grade) was recrystallized from water, and the water of crystallization removed by drying the hydrate for several hours at 130°. **Sodium iodide** (reagent grade) proved to be alkaline and was therefore recrystallized from water while nitrogen was bubbled through the solution. The neutral product was dried for 9 hours in a vacuum oven at 100°. **Lithium chloride** (reagent grade) was recrystallized from water and dried for several hours at 130° just before use. **Sodium hydroxide** pellets (reagent grade) were used to make up a 50% aqueous stock solution.

Solvents.—Purified **dioxane** and doubly distilled **water** were freshly prepared when required, as described previously.[3] Reagent grade **acetone** was dried by means of Linde "molecular sieves," type 4A, $^1/_{16}$ inch pellets, and distilled. The Karl Fischer titer for water in the final solvent was zero within experimental error.

Preparation and Analysis of Solutions.—Solutions were prepared from freshly purified chemicals by standard analytical procedures with the customary attention to detail required in good analytical work. To titrate for acid, aliquot portions of the reaction mixtures were diluted with twenty volumes of neutral acetone and titrated with *ca.* 0.01 *N* standard aqueous sodium hydroxide, either potentiometrically or—more usually—with lacmoid as an end-point indicator. Titrations for bromide (to obtain the rate constants for substitution listed in Table V) were done potentiometrically with standard silver nitrate.

To follow the solvolysis of 0.005 *M* neophyl *p*-toluenesulfonate in the presence of 0.1 *M* or 0.2 *M* NaOH or NaI, it was necessary to determine the amount of unreacted ester as a function of time. The following procedure was devised: 10.00 ml. of the reaction mixture was pipetted into 30 ml. of water in an erlenmeyer flask. The solution immediately became cloudy, indicating that the ester had "oiled out." The oil was induced to crystallize by freezing the mixture solid and allowing it to stand overnight in the freezing compartment of a refrigerator. Thawing the next day resulted in a clear liquid phase and a nicely crystalline solid phase. The crystals were transferred quantitatively onto a sintered glass funnel, washed thoroughly with doubly distilled water, dissolved in 5 ml. of pure dioxane, and the solution carried by suction into an ampoule. Next, 5 ml. of water was allowed to pass through the sintered glass funnel and into the ampoule. The ampoule was then sealed and heated at 83° for 8–10 hours, during which time the solvolysis of the unreacted ester was virtually complete. The amount of

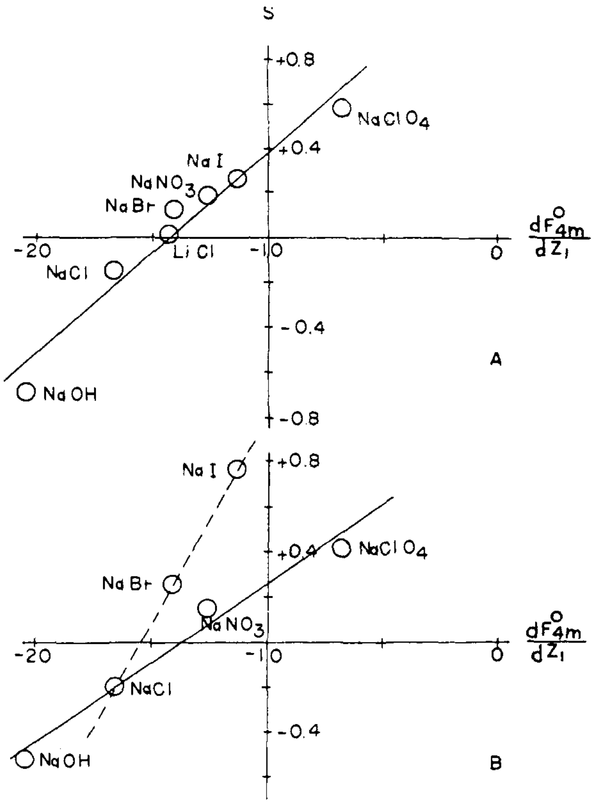


Fig. 1.—Plot of *S vs.* dF^0_{4m}/dZ_1: (A) solvolysis of neophyl *p*-toluenesulfonate in 50% dioxane–water, 50°; (B) racemization of L-(+)-*threo*-3-phenyl-2-butyl *p*-toluenesulfonate in 50% dioxane–water, 25°.

TABLE VI

SAMPLE KINETIC DATA

10^{-3} *t*, sec.	Titer or rotation	$10^6 k_1$, sec.$^{-1}$
Solvolysis of neophyl *p*-toluenesulfonate with 0.2 *M* LiCl		
0	0	..
62.58	2.70	6.96
84.36	3.48	7.19
145.38	4.92	7.10
172.80	5.37	7.02
227.28	6.06	6.93
253.97	6.32	6.91
Infinite	7.64	..
Solvolysis of neophyl *p*-toluenesulfonate with 0.1 *M* NaI		
0	7.72	..
86.4	3.81	8.17
172.8	1.84	8.30
259.2	0.87	8.43
345.6	0.43	8.36
Racemization of L(+)-*threo*-3-phenyl-2-butyl *p*-toluenesulfonate with 0.2 *M* NaCl		
0	0.222	..
75.0	.156	4.70
165.3	.098	4.95
256.5	.069	4.56
342.9	.046	4.59

---

(19) H. H. Willard, L. R. Perkins and F. F. Blicke, THIS JOURNAL, **70**, 737 (1948).

unreacted ester was then determined by titration of the acid liberated during its solvolysis. The method was found to give quantitative results in several control experiments. Values of $k_1°$ obtained by this method were highly precise and showed no drifts during more than three half-lives, but inexplicably were 10% higher than values of $k_1°$ obtained by following the progress of the solvolysis by direct titration for acid. Although this discrepancy has disturbed us, we believe that the error, if any, will cancel out in the ratio $k_1/k_1°$. Typical kinetic data are shown in Table VI.

Measurement of Optical Rotations.—Rotations were measured with a Bellingham and Stanley polarimeter at the wave length of the sodium D line (589 mμ), or at the 546 mμ mercury line, the latter giving a 20% larger rotation.

Due to the limited solubility, the concentration of III could not be made higher than 0.01 *M*, and in a 4-dm. tube the initial rotations were of the order of 0.25°. Before each measurement, the zero point of the polarimeter was checked by taking a set of 8 readings without the polarimeter tube. Next, a set of 8 readings was taken with the polarimeter tube in place; to eliminate errors from polarization of the end plates, the tube was then reversed and another set of 8 readings was taken. The difference between the mean of the 16-tube readings and the mean of the 8 air readings was taken as the rotation of the solution.

The standard deviation in each set of 8 readings was about 0.003°, and the difference between the means of the 2 sets of tube readings was of the order of 0.005°.

Rate Measurements.—The half-lives of all reactions were sufficiently long so that accurate timing presented no problem. The solvolysis of neophyl *p*-toluenesulfonate was followed by a standard sealed-ampoule technique, the ampoules having been flushed with pure nitrogen prior to filling. The racemization of L-(+)-*threo*-3-phenyl-2-butyl *p*-toluenesulfonate was measured as follows: Reaction mixtures were prepared in 50-ml. volumetric flasks, and the volumetric flasks were then immersed in a water thermostat maintained at 25.00°. Optical rotations were measured immediately after attainment of temperature equilibrium, and then once a day for 4 days thereafter. The measurements of the optical rotations required approximately 30 minutes each, and were carried out in a 4-dm. polarimeter tube thermostated at 25.00°. The mid-points of the periods of measurement were recorded as the effective time. After each measurement the reaction mixture was returned to the volumetric flask and replaced in the thermostat.

Least-squares Treatment of the Data.—For a first-order solvolysis or racemization, $\ln|(y_\infty - y_t)| = a - k_1t$, where y denotes the acid titer or the optical rotation, and a and k_1 are adjustable constants. Since the error in $|(y_\infty - y_t)|$ is constant throughout the kinetic run, the error in $\ln|(y_\infty - y_t)|$ is inversely proportional to $|(y_\infty - y_t)|$, and the statistical weight of the corresponding point may be equated to $|(y_\infty - y_t)|$. Thus

$$\Sigma[(y_\infty - y_t)^2 \ln|(y_\infty - y_t)|] = a\,\Sigma(y_\infty - y_t)^2 - k_1\Sigma[(y_\infty - y_t)^2 t]$$

$$\Sigma[(y_\infty - y_t)^2\, t \ln|(y_\infty - y_t)|] = a\,\Sigma(y_\infty - y_t)^2 t - k_1\,\Sigma(y_\infty - y_t)^2 t^2$$

[CONTRIBUTION NO. 1066 FROM THE DEPARTMENT OF CHEMISTRY, UNIVERSITY OF PITTSBURGH, PITTSBURGH 13, PENNA.]

Stereoelectronic Effects on Organic Bases. II.[1] Base Strengths of the Phenolic Ethers[2]

BY EDWARD McCOLLIN ARNETT AND CHING YONG WU

RECEIVED MARCH 23, 1960

The pK_a's of the oxonium ions of phenol and the phenyl ethers of the following alkyl groups have been studied: methyl, ethyl, *n*-propyl, isopropyl, *n*-butyl, isobutyl, *sec*-butyl, *n*-amyl and also 5,5-dimethylhomochroman. Several others were too insoluble or too unstable to investigate. Decomposition of the ethers by the aqueous sulfuric acid medium was controlled by working at 0° and a special treatment of the data was developed to deal with sulfonation and solvent effects. The technique gives highly precise pK_a's by the Hammett method, but discrepancies in the H_0 scale at strong acidities are noted. The results show that these phenolic compounds are protonated on the oxygen rather than the ring and are in general more sensitive to stereoelectronic effects than are their nitrogen analogs.

Much of the theoretical foundation of modern organic chemistry is based upon observations of the effects of changing molecular structure upon acid-base equilibria[3,4] of carboxylic acids, phenols and amines, all of which may be readily investigated in the *p*H range of acidity. Also a few types of weak bases[5-9] have been studied by means of the Hammett H_0 indicator method which serves, in effect, to extend the *p*H scale to the solutions of high acidity required for the protonation of weak bases. There remains to be answered the question of how analogous compounds of nitrogen, oxygen, phosphorus and sulfur compare in their abilities to accept a proton under similar conditions. Such comparisons should yield fundamental knowledge about the relative importance of electronic and steric effects as they operate on the central atom and be of considerable use in predicting the point of attack by protons in polyfunctional molecules and the best conditions for separating weak bases by extraction. A good point of entry to this field should be a comparison of alcohols and ethers with primary and secondary amines. Although the literature contains numerous attempts to measure basicities of simple aliphatic alcohols and ethers, most of the methods involve techniques that make it risky to refer the results to the standard state of infinite dilution in aqueous solution on which the pK_a scale is based. Comparison of the results shows disagreement not only between the magnitudes of these base strengths but often even differences in order. The difficulty is mainly an analytical one since the simple aliphatic compounds do not have the spectral absorption characteristics that allow them to be treated as Hammett indicators.

We have approached this problem through the

---

(1) Part I. E. M. Arnett and C. Y. Wu, *Chemistry & Industry*, 1488 (1959).

(2) We are grateful to the Research Corporation and the National Science Foundation for supporting this work.

(3) H. C. Brown, D. H. McDaniel and O. Häfliger, "Determination of Organic Structures by Physical Methods," edited by E. A. Braude and F. C. Nachod, Academic Press, Inc., New York, N. Y., 1955, Chapter 14.

(4) H. K. Hall, Jr., THIS JOURNAL, **79**, 5441 (1957).

(5) M. A. Paul and F. A. Long, *Chem. Revs.*, **57**, 1 (1957).

(6) N. C. Deno, J. J. Jaruzelski and A. Schriesheim, THIS JOURNAL, **77**, 3044 (1955).

(7) R. Stewart and K. Yates, *ibid.*, **80**, 6355 (1958).

(8) K. Yates and R. Stewart, *Can. J. Chem.*, **37**, 664 (1959).

(9) K. Yates, Thesis, University of British Columbia, 1959.