ORGANIC AND BIOLOGICAL CHEMISTRY

[CONTRIBUTION FROM THE VENABLE CHEMICAL LABORATORY, UNIVERSITY OF NORTH CAROLINA]

Effects of *o*-Methyl and *o*-Bromine on Rates of Nucleophilic Displacement at Aromatic Carbon. Evidence for the Operation of London Forces[1]

BY J. D. REINHEIMER[2] AND J. F. BUNNETT[3]

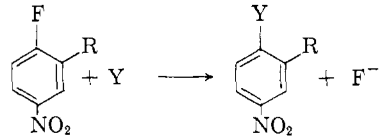
RECEIVED JULY 9, 1958

Rates of reactions of five nucleophilic reagents with 1-fluoro-4-nitrobenzene and its 2-bromo and 2-methyl derivatives have been determined. The data reveal the operation of both steric and London forces interactions with the methyl and bromine *ortho* substituents. London forces are especially important in reactions with thiophenoxide ion; they are dramatically illustrated by the net accelerating influence of the *o*-methyl group in reaction with this reagent.

A new factor affecting rates of bimolecular nucleophilic displacements recently was recognized.[4] This factor manifests itself in a tendency of nucleophilic reagents of high polarizability, such as iodide ion and thiophenoxide ion, to be especially reactive with substrates having large halogen atoms at or near the site of substitution. The physical principle underlying this new factor is believed to be the operation, in the transition state, of London forces[5] between regions of high polarizability in the reagent and in the substrate.

The new factor was recognized through comparison of suitable rate ratios. In several reaction series the thiophenoxide-methoxide rate ratio ($k_{\text{PhS}^-}/k_{\text{OMe}^-}$) was seen to increase with increase in size or weight of the halogen displaced. Similar trends in the thiophenoxide-methoxide ratio were noted with increase in size of halogen at the site of substitution *but not displaced in the rate-determining step*. Indeed, evidence was presented[4] that, in SN2 displacements in benzyl chlorides, even the introduction of large halogen substituents on the benzene ring can especially favor reaction with a nucleophile of high polarizability (iodide ion as compared to hydroxide ion). This led to the expectation that *o*-substituents of high polarizability would in general tend, after allowance had been made for their electronic and steric effects, to accelerate reactions with nucleophiles of high polarizability.

The present research was a probe of this expectation. In line with our interests in nucleophilic substitution at aromatic carbon,[6,7] we chose to investigate the effects of *o*-bromine and *o*-methyl, as compared to *o*-hydrogen, on the rates of displacement of nuclear fluorine by reagents of varying polarizability. Bromine and methyl were chosen as *o*-substituents for study because they differ considerably in polarizability and yet are similar in size[8] so that direct steric hindrance effects are about the same for both groups. That bromine and methyl differ in direction of electronic effect was of interest but not of concern inasmuch as allowance for electronic effects is made by considering rate ratios. The substrates chosen for investigation were 1-fluoro-4-nitrobenzene, 1-fluoro-2-bromo-4-nitrobenzene and 1-fluoro-2-methyl-4-nitrobenzene. Equation 1 is a general representation of the reactions studied (details of balancing electronic charge and proton distribution are for the moment ignored).



(1)

The reagents studied were hydroxide ion, methoxide ion, ammonia, piperidine and thiophenoxide ion; these are arranged in approximate order of increasing effective polarizability as expected from the previous study.[4] Methanol was the solvent for all reactions except those with hydroxide ion which were run in 50% aqueous dioxane. Several analytical methods were used to follow the progress of reactions, as explained in the Experimental section. In general rates could be determined with high precision.

Experimentally measured rate coefficients and calculated rate coefficients at standard temperatures are listed in Table I. The order of reactivity of the several reagents varies somewhat according to the substrate, but it is clear that of this set of five reagents the methoxide and thiophenoxide ions are "fast," that hydroxide ion and piperidine are of intermediate activity, and that ammonia is "slow." Energies and entropies of activation are listed in Table II; it can be seen that most differences in rate derive from changes in both activation parameters.

Certain predictions of the absolute substituent effects to be expected for *o*-bromine and *o*-methyl can be made from conventional theory. Electronically, for these reactions which are assisted by electron withdrawal from the site of substitution, bromine should accelerate and methyl should decelerate reaction. Sterically, each group should hinder reaction. The actual substituent effects (Table I) are in all cases but one in the direction predicted from electronic considerations. The exception is that *o*-methyl *accelerates* reaction with thiophenox-

---

[1] This research was supported in part by a grant from The Petroleum Research Fund administered by the American Chemical Society. Grateful acknowledgment is hereby made to the donors of this fund.

[2] On Research Leave from the College of Wooster, 1957–1958.

[3] Brown University, Providence, R. I.

[4] J. F. Bunnett, THIS JOURNAL, **79**, 5969 (1957).

[5] O. K. Rice, "Electronic Structure and Chemical Binding," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, p. 354; J. R. Partington, "An Advanced Treatise on Physical Chemistry," Vol. I, Longmans, Green and Co., London, 1949, p. 741.

[6] J. D. Reinheimer, W. F. Kieffer, S. W. Frey, J. C. Cochran and E. W. Barr, THIS JOURNAL, **80**, 164 (1958).

[7] J. F. Bunnett, *Quart. Revs.*, **12**, 1 (1958).

[8] R. W. Taft, Jr., in M. S. Newman, "Steric Effects in Organic Chemistry," John Wiley and Sons, Inc., New York, N. Y., 1956, p. 598.

TABLE I

SECOND-ORDER RATE COEFFICIENTS (IN L. MOLE^{-1} SEC.$^{-1}$ × 10^5) FOR REACTIONS OF VARIOUS NUCLEOPHILIC REAGENTS WITH 2-SUBSTITUTED-4-NITROFLUOROBENZENES[a]

	2-Substituent					
	Hydrogen		Methyl		Bromine	
Reagent	Temp., °C.	k	Temp., °C.	k	Temp., °C.	k
OH^-	46.10	4.37 ± 0.00	46.15	1.40 ± 0.02	46.15	106 ± 3.5
	86.3	178 ± 0	86.3	50.8 ± 2.8	68.3	633 ± 17
	0	*0.0163*	*0*	*0.00612*	*0*	*1.01*
	100	*525*	*100*	*145*	*100*	*5720*
OCH_3^-	46.05	162	46.05	44.0 ± 0.6	46.05	4750 ± 50
			96.8	4680 ± 60		
	0	*0.579*[b]	*0*	*0.142*	0.0	46.9 ± 0.6
	100	*20300*[b]	*100*	*6020*	*100*	*249000*
NH_3	*46.05*	*2.44×10^{-3}*	*46.05*	*1.00×10^{-3}*	*46.05*	*0.158*
	97.2	0.209 ± 0.005	97.2	0.0670 ± 0.0003	86.2	4.14 ± 0.01
	129.0	1.88 ± 0.05	129.0	0.533 ± 0.015	129.4	67.1 ± 1.4
	0	*1.06×10^{-5}*	*0*	*5.93×10^{-6}*	*0*	*1.15×10^{-8}*
	100	*0.258*	*100*	*0.0814*	*100*	*10.8*
$C_5H_{10}NH$	46.05	16.0 ± 0.3	46.05	0.358	46.05	119 ± 1
	68.0	65.8 ± 0.2	96.8	5.80 ± 0.02	68.0	363 ± 5
	96.8	305 ± 3	128.4	21.2 ± 0.4		
	0	*0.428*	*0*	*0.0125*	*0*	*6.04*
	100	*358*	*100*	*6.36*	*100*	*1460*
$C_6H_5S^-$	46.25	153 ± 3	46.25	159 ± 1	46.22	17500 ± 300
			97.2	7710 ± 60	24.60	3050 ± 50
	0	*0.778*[c]	*0*	*1.34*	*0*	*299*
	100	*13700*[c]	*100*	*9250*	*100*	*561000*

[a] Data in italics are calculated from data at other temperatures. [b] Calculated from data of G. P. Briner, J. Miller, M. Liveris and P. G. Lutz, ref. 30. [c] Calculated from data of C. W. L. Bevan and J. Hirst, ref. 26.

TABLE II

ACTIVATION PARAMETERS FOR REACTIONS OF VARIOUS NUCLEOPHILIC REAGENTS WITH 2-SUBSTITUTED-4-NITROFLUOROBENZENES

	2-Substituent					
	Hydrogen		Methyl		Bromine	
Reagent	$\Delta H^\ddagger$, kcal./mole	$\Delta S^\ddagger$, cal./deg.	$\Delta H^\ddagger$, kcal./mole	$\Delta S^\ddagger$, cal./deg.	$\Delta H^\ddagger$, kcal./mole	$\Delta S^\ddagger$, cal./deg.
OH^-	20.4	−11.3	19.8	−19.0	16.9	−19.4
OCH_3^-	20.6[a]	− 7.0[a]	20.9	− 8.4	16.7	−12.3
NH_3	19.8	−31.4	18.7	−36.8	17.9	−29.1
$C_5H_{10}NH$	13.0	−35.0	12.0	−46.0	10.4	−39.6
$C_6H_5S^-$	19.2[b]	−11.6[b]	17.3	−17.7	14.6	−16.3

[a] Ref. 30. [b] Ref. 26.

ide ion at 46.24°. On both electronic and steric grounds methyl ought to retard this reaction. This is, as we shall discuss further, an example of the predominance of London forces over conventional steric and electronic effects.

In general it is necessary to compare rate ratios in order to see the operation of London forces. In Table III, reagent:hydroxide ion rate ratios at 0° are displayed. Hydroxide ion is chosen as the basis for comparison because of its low polarizability and small size. Such ratios are taken to make allowance for substituent electronic effects; the ratio k_Y/k_{OH^-} reflects differences in inherent nucleophilic reactivity of reagents and differences in local interaction (steric and London forces) of *o*-substituents with reagents, but not differences in electronic effects of the *o*-substituents.[9] Allowance is made for differences in inherent nucleophilic reactivity of the reagents by adjusting such ratios to the scale of 1.0 for *p*-fluoronitrobenzene for each reagent. The adjusted ratios, which represent only differences between the reagent under consideration and hydroxide ion in respect to local interactions with *o*-substituents,[9] are listed in Table IV. These adjusted ratios are presented at three standard temperatures;

---

(9) These conclusions may be justified with reference to a modification of the Hammett equation[10] resembling one used by Taft.[8]

$$\log (k/k_0) = \rho\sigma + g$$

The term $\rho\sigma$ in this equation represents electronic effects and the term g represents the total of all steric and London forces interactions. One may then write

$$\log k_Y - \log k_{o,Y} = \rho_Y\sigma + g_Y$$

$$\log k_{OH^-} - \log k_{o,OH^-} = \rho_{OH^-}\sigma + g_{OH^-}$$

Subtracting

$$\log(k_Y/k_{OH^-}) = \sigma(\rho_Y - \rho_{OH^-}) + \log(k_{o,Y}/k_{o,OH^-}) + (g_Y - g_{OH^-})$$

Since ρ-values in aromatic nucleophilic substitutions with various reagents are all of about the same magnitude,[11] one may equate ρ_Y and ρ_{OH^-}. This causes the electronic term to drop out, and one obtains

$$k_Y/k_{OH^-} = (k_{o,Y}/k_{o,OH^-})e^{(g_Y - g_{OH^-})}$$

Values of k_Y/k_{OH^-} are given in Table III; from this equation one can see that they represent both differences in inherent nucleophilic reactivity of Y and OH^-, as represented by $k_{o,Y}/k_{o,OH^-}$, and differences in steric and London forces interactions as represented by $e^{(g_Y - g_{OH^-})}$. This equation is transformed easily into

$$\left(\frac{k_Y}{k_{OH^-}}\right) \Big/ \left(\frac{k_{o,Y}}{k_{o,OH^-}}\right) = e^{(g_Y - g_{OH^-})}$$

Values of the left-hand term of this equation are set forth in Table IV. It is thus shown that these values represent only differences in steric and London forces interactions (as well as differences in any other local interactions of type as yet unrecognized).

(10) L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, p. 186; H. H. Jaffé, *Chem. Revs.*, **53**, 191 (1953).

(11) J. F. Bunnett and R. F. Snipes, THIS JOURNAL, **77**, 5422 (1955); J. F. Bunnett, H. Moe and D. Knutson, *ibid.*, **76**, 3936 (1954); E. Berliner and L. C. Monack, *ibid.*, **74**, 1574 (1952); J. Miller, *Aust. J. Chem.*, **9**, 61 (1956).

one can thereby judge temperature dependence. Adjusted ratios for the *o*-nitro group at 0°, calculated from data in the literature, are presented also for all reagents but ammonia.

TABLE III

RATE RATIOS (k_Y/k_{OH^-}) AT 0°

Reagent	H	2-Substituent CH_3	Br	NO_2[a]
OCH_3^-	36	23	47	18
NH_3	6.5×10^{-4}	9.7×10^{-4}	11×10^{-4}	..
$C_5H_{10}NH$	26	2.2	6.0	16
$C_6H_5S^-$	48	208	298	1090

[a] Source of NO_2 values explained in Experimental section.

TABLE IV

ADJUSTED RATE RATIOS: $\left(\frac{k_Y}{k_{OH^-}}\right)_R \Big/ \left(\frac{k_Y}{k_{OH^-}}\right)_H$

Reagent	Temp., °C.	H	CH_3	Br	NO_2
OCH_3^-	0	1.0	0.65	1.3	0.50
	46.05	1.0	0.86	1.2	
	100	1.0	1.0	1.1	
NH_3	0	1.0	1.5	1.7	
	46.05	1.0	1.3	2.7	
	100	1.0	1.1	3.9	
$C_5H_{10}NH$	0	1.0	0.084	0.23	0.62
	46.05	1.0	.070	.30	
	100	1.0	.063	.37	
$C_6H_5S^-$	0	1.0	4.3	6.2	23
	46.05	1.0	3.2	4.6	
	100	1.0	2.5	3.8	

The adjusted rate ratios in Table IV reveal the operation of both steric and London forces effects. In the case of ammonia, a small reagent similar to hydroxide ion in steric requirement but larger in polarizability,[12] the adjusted rate ratio is increased by *o*-methyl and increased to a greater extent by *o*-bromine. That is, ammonia is relatively more effective when methyl or bromine is present in the *o*-position. This increase in effectiveness is ascribed to London forces operating between the reagent of higher polarizability (ammonia as compared to hydroxide ion) and *o*-substituents of higher polarizability (methyl or bromine as compared to hydrogen). That the increase is greater for bromine than for methyl is reasonable because bromine has higher polarizability.[13]

The data concerning piperidine indicate the strong incursion of adverse steric factors. This was expected from work of Capon and Chapman[14] who showed that the effective bulk of piperidine is considerably greater than that of methoxide ion or aniline in reaction with 6-methyl-2,4-dinitrochlorobenzene. All three *o*-substituents cause a net decrease in the rate ratio for piperidine. Thus, owing to steric repulsions of *o*-substituents with a large reagent, piperidine is relatively less effective with any of the *o*-substituted substrates. However one can still discern the play of London forces. The adjusted rate ratio for *o*-bromine is nearly three times greater than for *o*-methyl; since these groups are nearly the same size, the increase is assigned to London forces which operate more strongly with the bromine atom because of its greater polarizability. There is a further increase with *o*-nitro in spite of the fact that nitro is bigger than bromine or methyl[8]; this is interpreted as evidence that the nitro group has even higher effective polarizability.[15]

The data for methoxide ion in Table IV appear to indicate a close balance of steric and London forces factors: steric factors (having to do with the greater size of methoxide than of hydroxide ion) appear to dominate with *o*-methyl, to be overcome by London forces with *o*-bromine which is no larger than methyl but has greater polarizability, and to regain dominance with *o*-nitro which is considerably larger than the other substituents. This analysis is weakened, however, by the considerable temperature dependence of the adjusted rate ratios: the differences have practically vanished at 100°. An inversion in the methyl:bromine position occurs on going to still higher temperatures.

It was expected that thiophenoxide ion, because of its high polarizability, would be most responsive to London forces. The data of Table IV fully support this expectation. Although this reagent is assuredly larger than hydroxide ion, no adverse steric effect is discernible. The adjusted rate ratios change as expected from consideration of London forces: the greater the polarizability of the *o*-substituent, the greater the increase in the adjusted rate ratio. Again the nitro group shows very high effective polarizability.[15]

The London forces interaction of thiophenoxide ion with *o*-methyl is dramatically illustrated (Table I) by the absolute accelerating effect of the *o*-methyl group at 46°. In this case London forces overcome both the unfavorable steric effect and the unfavorable electronic effect of the methyl group.[16]

The above analysis shows that a satisfying interpretation of the observed rate ratios can be made with reference to London forces in addition to conventional steric and electronic effects. Let us now see how one might attempt to rationalize the results without invoking London forces. First, let us continue with the assumption[9] that adequate allowance for substituent electronic effects is made by considering k_Y/k_{OH^-} ratios. The adjusted rate ratios in Table IV then must be explained with reference only to steric effects. The three principal *o*-substituents would have to be ranked in order of increasing size, either $H < CH_3 < Br$ or $H < Br < CH_3$. If the former order is assumed and if it is further assumed that ammonia and thiophenoxide ion have less effective bulk than hydroxide ion (this

---

[12] The refractions (R_D) of hydroxide ion and ammonia are, respectively, 5.15 and 5.9; "Landolt-Bornstein," 6 Aufl., I Band, 1 Teil, p. 402, and 3 Teil, p. 514.

[13] The refraction constants (R_D) of hydrogen, methyl and bromine are, respectively, 1.10, 5.72 and 8.86 cc./mole; J. R. Partington, "An Advanced Treatise on Physical Chemistry," Vol. 4, Longmans, Green and Co., London, 1953, p. 46.

[14] B. Capon and N. B. Chapman, *J. Chem. Soc.*, 600 (1957).

[15] At first glance this conclusion appears to be at variance with the facts; R_D for the aromatic nitro group is 7.30 compared to 8.86 for bromine. It is, however, necessary to consider the polarizability of the nitro group *in the transition state* in which it carries increased negative charge with resultant increase in polarizability.

[16] E. L. Eliel and R. S. Ro, *Tetrahedron*, **2**, 353 (1958), have recorded another instance in which the introduction of methyl groups brings about an absolute acceleration of nucleophilic displacement induced by thiophenoxide ion. These authors found that the rate of displacement, to form cycloalkyl phenyl sulfides, is about twice as great with 4,4-dimethylcyclohexyl *p*-toluenesulfonate as with unsubstituted cyclohexyl *p*-toluenesulfonate.

assumption is dubious for ammonia and very improbable for thiophenoxide), the adjusted rate ratios for these two reagents can be explained. However it would be necessary to assume the order of size, $CH_3 > Br > H$, for which there is no precedent,[17] together with the reasonable postulate that piperidine is larger than hydroxide, in order to rationalize the ratios for piperidine. This approach to rationalization of the results thus involves inconsistent and unreasonable assumptions.

In a further attempt to interpret the data of Table IV without reference to London forces, let us accept the established order of size, $H < CH_3 \leqslant Br$, and postulate that reactions with all the other reagents are more susceptible to substituent electronic effects than are reactions with hydroxide ion. The latter would mean greater Hammett ρ-values with the other reagents than with hydroxide ion, a possibility which cannot be excluded on the basis of evidence now available. This would result in ratios with *o*-bromine which were raised and ratios with *o*-methyl which were depressed with respect to some intermediate level determined by steric factors. If one assumes the order of effective reagent size, $C_5H_{10}NH > OH^- \sim OCH_3^- > NH_3 >> C_6H_5S^-$, the adjusted rate ratios of Table IV can be successfully rationalized. A major difficulty with this interpretation is that it requires some strange assumptions about the relative sizes of reagents; the assumption that thiophenoxide ion is effectively much smaller than hydroxide is especially unpalatable.

Furthermore, neither of these rationalizations *sans* London forces can interpret the absolute acceleration of reaction with thiophenoxide ion by the *o*-methyl group. On both steric and electronic counts, methyl should decelerate this reaction. Because the interpretation with reference to London forces gives a good account of this observation as well as all the other data of Table IV, without requiring any strange assumptions, we consider it to be correct. The important part of London forces in determining the rates of these reactions is demonstrated.

Experimental

Materials.—***p*-Fluoronitrobenzene** was prepared by the Finger reaction.[18] The product was purified by distillation through a 15-cm. Widmer column; b.p. 74.0–74.4° (7–8 mm.). **1-Fluoro-2-bromo-4-nitrobenzene** was prepared as described by Bunnett and Rauhut.[19]

1-Fluoro-2-methyl-4-nitrobenzene.—Attempts to prepare this compound by nitration of *o*-fluorotoluene according to Schiemann[20] were not fruitful. 2-Amino-5-nitrotoluene (tech., 95.4% pure)[21a] was converted to 2-chloro-5-nitrotoluene by the procedure of Goldschmidt and Hoenig in[21b] 78% yield. One hundred grams of the 2-chloro-5-nitrotoluene was dissolved in 300 cc. of dimethyl sulfoxide. Anhydrous potassium fluoride (80 g.) was added and the mixture was heated at 185–190° for 12–16 hours with vigorous stirring. The product was isolated by steam distillation and then vacuum distillation through a Widmer column. The first fraction of distillate solidified in the condenser, then an intermediate liquid fraction was obtained, and finally a higher boiling solid fraction. These observations indicated good separation of the starting chloro compound (m.p. 44°) and the product fluoro compound (m.p. 41.5°). The first solid fraction, b.p. 95.5–97.0° (10 mm.), m.p. 37–40°, was the crude product, 38.7 g., 43% yield. The melting point after two recrystallizations from ethanol was 40–41° (lit.[19] 41.5°).

Reagent grade *methanol* was dried with use of magnesium metal.[22] **Dioxane** was purified by refluxing with hydrochloric acid and then refluxing over sodium metal for several days.[23] The pure dioxane was distilled through a Widmer column, stored over sodium metal and redistilled from sodium immediately before each use. **Piperidine** was refluxed over sodium metal and redistilled; b.p. 105–105.5°. It was stored in a desiccator over sodium hydroxide. **Thiophenol**, Eastman Kodak Co. White Label, was distilled and a middle cut was used. It was stored in a desiccator over calcium chloride.

Kinetic Runs with Sodium Methoxide.—The procedure previously described[6] was used. Runs at 97° were, however, carried out with aliquots of the reaction solution in sealed ampoules. In all runs the concentration of sodium methoxide was equal to that of the substrate and about 0.03 *M*. Good linear second-order plots, good infinity values and good reproducibility between runs were obtained.

Kinetics Runs with Sodium Hydroxide.—Reactions were run in polyethylene bottles to avoid attack of hydroxide ion on glass. Each reaction solution was prepared as follows: A weighed sample of the substrate was placed in a 100-cc. volumetric flask and a small amount of dioxane was added to dissolve the sample. Then 50 cc. of a standard solution of sodium hydroxide in water was added by pipet, and the solution was diluted to the mark with pure dioxane and shaken thoroughly. This solution was transferred to a polyethylene bottle which had been flushed with nitrogen, and this bottle was placed in the thermostat. Samples were removed by pipet and the time of release into a hydrochloric acid quenching solution was noted. The extent of reaction was determined by bromination of the phenol formed.[24] The substrates were not brominated under the conditions used but dioxane gave a blank which increased with time of bromination. It was necessary to brominate each sample for a definite period (1 or 2 minutes) and to correct each titration value by subtracting the corresponding blank. It was shown that the auto-oxidation of dioxane during two days at 46° formed no peroxides detectable by bromine titration; interference from this source was thus negligible for fast runs. However, runs which lasted three days or more gave plots of $1/(a - x)$ *vs.* t with a characteristic downward curvature for the last two or three points. The rate coefficients for such runs were derived from the first five or six points which were in a good straight line. Inclusion of the last three points would have decreased the rate coefficient by about 2.5% in a typical run.

Kinetic Runs with Sodium Thiophenoxide.—The faster runs were performed in volumetric flasks with use of the sampling procedure of Bunnett and Merritt[25] while the slower runs were performed in sealed tubes as described by Bevan and Hirst.[26] The only modifications were the use of a sintered glass filter stick to bubble the reaction solutions with nitrogen and the use of an aqueous rather than a methanolic quenching solution. The nitrogen was passed through a sulfuric acid tower and then through a methanol tower before being used to bubble the solutions. The extent of reaction was determined by iodimetric titration of thiophenol in the quenched solutions.[27]

Kinetics Runs with Piperidine.—The spectrophotometric procedure of an earlier investigation[28] was used with the

---

[17] R. L. Shriner, R. Adams and C. S. Marvel, in H. Gilman, "Organic Chemistry," 2nd ed., John Wiley and Sons, Inc., New York, N. Y., Vol. I, pp. 362, 368.

[18] G. C. Finger and C. W. Kruse, THIS JOURNAL, **78**, 6036 (1956).

[19] J. F. Bunnett and M. M. Rauhut, *J. Org. Chem.*, **21**, 937 (1956).

[20] G. Schiemann, *Ber.*, **62**, 1799 (1929).

[21] (a) Our thanks to E. I. du Pont de Nemours and Co. for the gift of this amine; (b) H. Goldschmidt and M. Hoenig, *Ber.*, **20**, 199 (1887).

[22] L. F. Fieser, "Experiments in Organic Chemistry," 2nd ed., D. C. Heath and Co., New York, N. Y., 1941, p. 360.

[23] A. Weissberger, "Technique of Organic Chemistry," 2nd ed., Interscience Publishers, Inc., New York, N. Y., 1955, Vol. VIII, p. 372.

[24] S. Siggia, "Quantitative Analysis *via* Functional Groups," 2nd ed., John Wiley and Sons, Inc., New York, N. Y., 1954, p. 162.

[25] J. F. Bunnett and W. D. Merritt, Jr., THIS JOURNAL, **79**, 5967 (1957).

[26] C. W. L. Bevan and J. Hirst, *J. Chem. Soc.*, 254 (1956).

[27] Reference 24, p. 131.

[28] J. F. Bunnett, E. W. Garbisch, Jr., and K. M. Pruitt, THIS JOURNAL, **79**, 385 (1957).

following modifications: (a) Solutions of the reactants for a run were bubbled with nitrogen, in separate flasks, for one minute at room temperature; these flasks then were placed in the thermostat and their contents were mixed at zero time by means of a bent tube with two standard taper male joints. (b) Runs at higher temperatures were performed in sealed tubes; the reaction solutions were made up at room temperature and were bubbled with nitrogen before aliquots were dispensed into nitrogen-flushed tubes. Photometric measurements were made at 400 mμ in runs with *p*-fluoronitrobenzene and at 380 mμ in runs with the other two substrates. Runs carried out without added piperidine hydrochloride had infinity absorbance values less than theoretical, but from runs with piperidine hydrochloride present the infinity absorbance was equal to that expected on the assumption of a quantitative reaction. That piperidine hydrochloride has this beneficial effect indicates that, in its absence, there is a methanolysis complication.[28] Rate coefficients *for reaction with piperidine* were the same in the presence as in the absence of piperidine hydrochloride. In all runs, concentrations were about as follows: substrate, 0.001 *M;* piperidine, 0.1 *M;* piperidine hydrochloride (if present), 0.06 *M*. The pseudo-first-order rate coefficients obtained were converted to second-order coefficients by dividing by the piperidine concentration.

Kinetic Runs with Ammonia.—The spectrophotometric procedure[28] was used and runs were set up so as to provide pseudo-first-order kinetics. Concentrations were about as follows: substrate, 1.0×10^{-3} *M;* ammonia, 1.0 *M*. All runs were performed in sealed ampoules. Ammonium chloride was present in the reaction solutions to suppress methanolysis. The ammonia stock solutions (about 2 *M*) were saturated or nearly saturated in ammonium chloride. The concentration of ammonium chloride in one stock solution of ammonia was higher than in another, but duplicate runs based on the two stock solutions gave the same rate coefficients. Absorbance was measured at 380 mμ in all runs. It was shown that Beer's law is obeyed by solutions of the various products. Most runs were so slow that it was impractical to wait ro obtain an "infinity" sample; theoretical infinity absorbances therefore were used. In runs with 1-fluoro-2-bromo-4-nitrobenzene at 86.2°, an experimental infinity absorbance of 1.030 was obtained; the theoretical infinity absorbance was 1.045. The same reaction at 129.4°, however, gave an infinity absorbance of 1.190 which is higher than expected. This reveals an unknown complication. Since the rate coefficients are based on data from the initial stages of reaction, during which the extent of the side reaction should be small, it is regarded as a minor perturbation.

General Kinetics Procedures.—Thermostats were constant to +0.02° at 46° and ±0.1° at higher temperatures. Temperatures were taken with thermometers certified by the National Bureau of Standards. The rate coefficients in Table I are average values from two or more supposedly identical runs (except where no deviation is noted). The average deviation of individual values from the mean is listed. Enthalpies and entropies of activation were calculated from standard expressions[29]; entropies were obtained with respect to rate coefficients at 46.05°.

Comparison of Rate Coefficients with Determinations by Other Workers.—For reaction of *p*-fluoronitrobenzene with sodium thiophenoxide at 46.25°, we found $k = 1.53 \times 10^{-3}$ l. mole^{-1} sec.$^{-1}$ while 1.48×10^{-3} is interpolated from data of Bevan and Hirst.[26] For reaction of this substrate with sodium methoxide at 46.05°, we found $k = 1.62 \times 10^{-3}$ l. mole^{-1} sec.$^{-1}$; the identical value is interpolated from data of Miller, *et al.*[30] For reaction of *p*-fluoronitrobenzene with piperidine in *methanol* at 96.8°, we found $k = 3.05 \times 10^{-3}$ l. mole^{-1} sec.$^{-1}$; for the same reaction in *ethanol* the value 3.18×10^{-3} is extrapolated from data of Chapman, *et al.*[31]

Rate Ratios in Table III for the *o*-Nitro Group.—Rate coefficients for reactions with methoxide ion, thiophenoxide ion and piperidine were listed by Bunnett.[4] Bunnett and Randall[32] have found, for reaction of 2,4-dinitrofluorobenzene with sodium hydroxide at 0° in 60% dioxane, $k =$ 0.114 l. mole^{-1} sec.$^{-1}$. Bunnett and Pruitt[33] reported for reaction of 2,4-dinitrochlorobenzene with sodium hydroxide at 45.95° in 50% dioxane and in 60% dioxane, respectively, $k = 0.377$ and 0.453 l. mole^{-1} min.$^{-1}$. These figures are used to establish a tentative ratio of hydroxide ion rates in 50% *vs.* 60% dioxane, and one calculates $k = 0.095$ l. mole^{-1} sec.$^{-1}$ for reaction of 2,4-dinitrofluorobenzene with sodium hydroxide in 50% dioxane at 0°. The various rate ratios are based on this figure.

Isolation of Reaction Products.—These various products were isolated from the relevant reactions: *p*-nitrophenol, m.p. 112.5–114° (lit.[34] 114°); 2-methyl-4-nitrophenol, m.p. 95–95.5° (lit.[34] 96°); 2-bromo-4-nitrophenol, m.p. 113–114.5° (lit.[34] 114°); 2-methyl-4-nitroanisole, m.p. 64.5–65° (lit.[34] 64°); 2-bromo-4-nitroanisole, m.p. 103.5–105° (lit.[34] 106°); 2-methyl-4-nitrophenyl phenyl sulfide, m.p. 82.5–84.5° (lit.[35] 82°); and 2-bromo-4-nitrophenyl phenyl sulfide, m.p. 112–114°. The last is a new compound; it was recrystallized from methanol.

Anal. Calcd. for $C_{12}H_8BrNO_2S$: C, 46.46; H, 2.60. Found (Micro-Tech. Labs.): C, 45.95, 46.25; H, 2.72, 2.67.

---

(29) S. Glasstone, K. J. Laidler and H. Eyring, "The Theory of Rate Processes," McGraw-Hill Book Co., Inc., New York, N. Y., 1941, p. 199.

(30) G. P. Briner, J. Miller, M. Liveris and P. G. Lutz, *J. Chem. Soc.*, 1265 (1954).

(31) N. B. Chapman, R. E. Parker and P. W. Soanes, *ibid.*, 2109 (1954).

(32) J. F. Bunnett and J. J. Randall, THIS JOURNAL, **80**, 6020 (1958).

(33) J. F. Bunnett and K. M. Pruitt, *J. Elisha Mitchell Sci. Soc.*, **73**, 297 (1957).

(34) I. Heilbron and H. M. Bunbury, "Dictionary of Organic Compounds," Oxford Press, London, Eng., 1953.

(35) J. Toussaint, *Bull. Soc. Roy. Sci., Liege*, **12**, 533 (1943).

CHAPEL HILL, N. C.