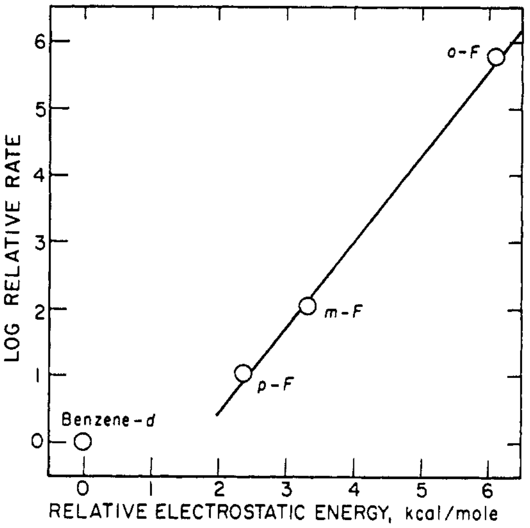

Figure 4. Results of a Kirkwood–Westheimer inductive field effect calculation applied to fluorophenyl anions and compared to kinetic acidities of fluorobenzenes.

k_{rel} for these positions *vs.* the calculated electrostatic energies, as in Figure 4, is adequately linear. Benzene itself does not fall on this line presumably because optimum parameters were not chosen; nevertheless these results do show that ordinary inductive field effects can account for the positional effects of a substituent in stabilizing a phenyl anion.

The kinetic acidity of the *meta* position of benzotrifluoride relative to benzene reported in Table II, 4.6 × 10^2, is also considerably lower than that obtained by Hall, Piccolini, and Roberts,[5,15] 10^4. Benzotrifluoride is also unstable toward LiCHA and probably also to KNH_2; possibly this decomposition, which was not considered in the earlier work, is responsible for the discrepancy. The rate enhancement caused by a *m*-trifluoromethyl group is about five times that of *m*-fluorine. The dipole moment of benzotrifluoride, 2.6 D., is considerably greater than that of fluorobenzene, 1.5 D.,[23] but a larger charge–dipole distance is involved with the CF_3 group. We plan to obtain accurate kinetic acidities for other substituted benzenes and defer more extensive calculations of electrostatic energies until more extensive such data have been obtained.

Finally, we may mention 2,5-difluorotoluene. From the reactivities of *m*- and *o*-fluorotoluene,[6] we can derive an estimated rate ratio, 2,5-difluorotoluene-α-*t*-toluene-α-*t*, of 260 at 50°. This estimate compares well with the value, 350 at 25°, found experimentally in the present work for deuterium exchange.[24]

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[23] Summarized in A. L. McClellan, "Tables of Experimental Dipole Moments," W. H. Freeman and Co., San Francisco, Calif., 1963. The "best value" cited for fluorobenzene is in error and has been revised to 1.463 D. (Dr. McClellan, personal communication).

[24] Note that if the relative rate is due entirely to an energy of activation difference the extrapolated relative rate for tritium exchange at 25° is 414.

Acidity of Hydrocarbons. XXVIII. Hydrogen Isotope Exchange of Polyfluorobenzenes with Sodium Methoxide in Methanol[1,2]

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Abstract: Rates are reported for hydrogen isotope exchange with sodium methoxide–methanol of pentafluorobenzene, the tetrafluorobenzenes, and 1,3-difluorobenzene. The results give the following log partial rate factors at 40°: *ortho*, 5.25; *meta*, 2.07; *para*, 1.13. The exchange reaction of pentafluorobenzene shows no primary isotope effect and appears to be of the internal return type. The relative rates are shown to be approximately equal to the relative equilibrium acidities.

The pronounced effects of fluorine substituents on kinetic acidities of aryl hydrogens[4] prompted a study of the kinetic acidity of pentafluorobenzene which was first synthesized only a decade ago[5] and which has recently become commercially available. Preliminary kinetic studies showed that the compound undergoes hydrogen isotope exchange quite rapidly even with dilute sodium methoxide in methanol; hence, a detailed study was undertaken that included the tetrafluorobenzenes and *m*-difluorobenzene.

Experimental Section

Polyfluorobenzenes. Commercial products (Pierce Chemical Company, Rockford, Ill.) were used without further purification. Purity was monitored by refractive index and by glpc using the following columns: Dow 710, a 1:4 mixture of Bentone 34 and SE 52, and QF 1, each 20% on Chromosorb W. All of the fluorobenzenes showed one peak except for 1,2,4,5-tetrafluorobenzene which contained traces of 1,2,3,5-tetrafluorobenzene and pentafluorobenzene and about 5% 1,2,3,4-tetrafluorobenzene.

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[1] Supported in part by Grant No. GM-12855 of the National Institutes of Health, U. S. Public Health Service, and by a grant from the Petroleum Research Fund of the American Chemical Society.

[2] Paper XXVII: A. Streitwieser, Jr., and F. Mares, *J. Am. Chem. Soc.*, **90**, 644 (1968).

[3] On leave from the Institute of Chemical Process Fundamentals, Prague, Suchdol 2, Czechoslovakia.

[4] *Cf.* summaries by A. I. Shatenshtein, *Advan. Phys. Org. Chem.*, **1**, 187 (1963), and A. Streitwieser, Jr., and J. H. Hammons, *Progr. Phys. Org. Chem.*, **3**, 41 (1965).

[5] R. Stephens and J. C. Tatlow, *Chem. Ind.* (London), 821 (1957).

Table I. Summary of Tritium Exchange Rates for Polyfluorobenzenes with Methanolic Sodium Methoxide

Run	Perfluorobenzene	10^2[CH$_3$ONa]	Temp, °C	$k_{\text{exp}}10^5$,[a] sec^{-1}	$k_2 10^4$,[b] l. mol^{-1} sec^{-1}
1	Pentafluoro	4.51	0.00	4.29 ± 0.27	9.73 ± 0.6
2	Pentafluoro	24.9	0.00	26.8 ± 0.7	10.8 ± 0.8
3	Pentafluoro	4.74	24.95	113.0 ± 8	238.0 ± 17.0
4	Pentafluoro-*d*(*t*)[c]	4.75	25.02	115.0 ± 5[c]	242.0 ± 10.0[c]
5	Pentafluoro	3.71	25.00	96.0 ± 6.5	259.0 ± 17.0
6	Pentafluoro	0.593	40.08	80.5 ± 2.1	1360.0 ± 30
7	Pentafluoro	0.321	59.94	189.0 ± 4.0	5890.0 ± 120
8	1,3-Difluoro	14.5	40.00	0.00840 ± 0.00028	$(5.79 \pm 0.19) \times 10^{-3}$
9	1,3-Difluoro	43.15	39.99	0.0274 ± 0.0008	$(6.35 \pm 0.18) \times 10^{-3}$
10	1,3-Difluoro	13.05	59.80	0.121 ± 0.004	$(92.7 \pm 3.1) \times 10^{-3}$
11	1,2,3,4-Tetrafluoro	65.0	25.00	0.0036 ± 0.0001	$(5.57 \pm 0.15) \times 10^{-4}$
12	1,2,3,4-Tetrafluoro	23.76	40.00	0.0126 ± 0.0001	$(53.0 \pm 0.4) \times 10^{-4}$
13	1,2,4,5-Tetrafluoro	6.67	39.99	39.4 ± 2.3	59.1 ± 3.4
14	1,2,4,5-Tetrafluoro	7.64	39.96	43.7 ± 2.0	57.2 ± 2.6
15	1,2,3,5-Tetrafluoro	21.3	39.99	12.5 ± 0.5	5.88 ± 0.30
16	1,2,3,5-Tetrafluoro	21.17	39.96	11.2 ± 0.5[d]	5.31 ± 0.20[d]
17[e]	Pentafluoro	3.43	25.00	193[e]	562[e]

[a] Pseudo-first-order rate constant. [b] k_{exp}/[NaOCH$_3$]. [c] Rate given is for loss of T; for loss of D, $k_{\text{exptl}} = (115 \pm 5) \times 10^{-5}$ sec^{-1}, $k_2 = (242 \pm 10) \times 10^{-4}$ l. mol^{-1} sec^{-1}. [d] From isolation and counting of substrate. Direct counting of toluene extracts gave $k_{\text{exptl}} = (11.8 \pm 0.7) \times 10^{-5}$ sec^{-1}, $k_2 = (5.57 \pm 0.33) \times 10^{-4}$ l. mol^{-1} sec^{-1}. [e] This rate is for incorporation of T from methanol-*d*(*t*). For incorporation of D, $k_{\text{exptl}} = 182 \times 10^{-5}$ sec^{-1}, $k_2 = 536 \times 10^{-4}$ l. mol^{-1} sec^{-1}.

Kinetics. In most of the runs, exchange was followed by tritium incorporation into the fluorinated benzene from methanol-*t*. Methanol was dried by distillation from magnesium methoxide and was tritiated by the addition of a small amount of highly active tritiated water. Sodium methoxide stock solutions were prepared by dissolving clean sodium in dry methanol. In slow kinetic runs a mixture of methanol-*t*, sodium methoxide stock solution, and the substrate was sealed into tubes and maintained at temperature for appropriate times. For faster kinetics a solution was prepared and brought to temperature in a thermostat and the reaction was initiated by injecting either the substrate or tritiated methanol. Aliquots were then removed at intervals with a syringe, quenched with excess HCl, and extracted with toluene. In many runs the polyfluorobenzene was removed from toluene by glpc and a weighed amount was dissolved in the scintillation solution and counted on a Nuclear Chicago Model 720 or Mark I[6] liquid scintillation counter. This method was found to give the same results as a simpler procedure in which a known amount of kinetic sample was extracted with a known amount of toluene, and the washed and dried toluene extract was counted directly. In a few cases, the substrate content in the toluene extract was determined by adding a known standard followed by a glpc analysis.

For those runs which were so slow that an infinity value could not be obtained directly, part of the tritiated sodium methoxide solution was used with pentafluorobenzene and the equilibrium isotope content of the latter was used to determine the infinity of the kinetic run. In all cases the rate constants were obtained as experimental pseudo-first-order rate constants using DeTar's LSKIN1 program.[7] The results are summarized in Table I.

Results and Discussion

Polyfluorobenzenes are known to undergo nucleophilic substitution with methanolic sodium methoxide but the exchange reaction for pentafluorobenzene is much faster. Titration during the course of a kinetic run showed no change in base concentration even when complete isotopic equilibration had been reached. At long times some base was lost indicative of nucleophilic substitution but at a rate several orders of magnitude slower than exchange. Hollyhead, *et al.*,[8] have recently determined the rates of nucleophilic substitution of polyfluorobenzenes by sodium methoxide; their results are summarized in Table II for comparison with the tritium exchange rates. Only for 1,2,3,4-tetrafluorobenzene are the two rates of comparable magnitude; for this compound, tritium exchange is so slow that only an initial rate was determined. Consequently, even for this case nucleophilic displacement was not a complicating problem.

Table II. Comparison of Tritium Exchange with Nucleophilic Displacement with Methanolic Sodium Methoxide

	$10^4 k_2$, l. mol^{-1} sec^{-1}	
Polyfluorobenzene	Exchange, 40°	Displacement, 50°[a]
Pentafluoro	1360	1.05
1,2,3,4-Tetrafluoro	0.0053	0.018
1,2,4,5-Tetrafluoro	58	$<10^{-4}$
1,2,3,5-Tetrafluoro	5.6	0.049
1,3-Difluoro	0.0061	

[a] Reference 8.

It should also be emphasized that there is no common pattern between the effect of structure on the exchange rate and on the displacement rate. This comparison effectively rules out any mechanism for exchange that starts with nucleophilic addition to the ring.

The simplest alternative mechanism is that of a direct attack of methoxide ion on the phenyl hydrogen to give an intermediate polyfluorophenyl anion. For such a mechanism the primary isotope effect is of especial interest. In run 4, pentafluorobenzene-*d*(*t*), which had been prepared by exchange with methanol-*d*(*t*), was allowed to exchange with methanolic sodium methoxide and the loss of both deuterium and tritium was followed. Note that the rate of loss of tritium is the same as the rate of incorporation of tritium from methanol-*t* as expected in the absence of any significant equilibrium isotope effect. Run 4 gave $k_D/k_T = 1.00 \pm 0.06$. This absence of any significant primary isotope effect is diagnostic of extensive internal return.[9-11] The mechanism is given as

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(6) This instrument was purchased with a grant from E. I. du Pont de Nemours and Co.

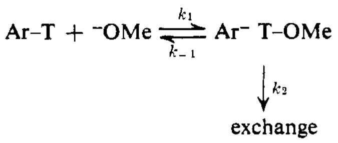
(7) We are indebted to Professor D. DeTar for a copy of his program and to H. A. Hammond for appropriate modifications.

(8) J. Burdon, W. B. Hollyhead, C. R. Patrick, and K. V. Wilson, *J. Chem. Soc.*, 6375 (1965); J. Burdon and W. B. Hollyhead, *ibid.*, 6326 (1965); W. B. Hollyhead, Dissertation, University of Birmingham, 1965.

(9) D. J. Cram, C. A. Kingsbury, and B. Rickborn, *J. Am. Chem. Soc.*, **83**, 3688 (1961).

(10) A. Streitwieser, Jr., and H. F. Koch, *ibid.*, **86**, 404 (1964).

(11) J. E. Hofmann, A. Schriesheim, and R. E. Nichols, *Tetrahedron Letters*, 1745 (1965).



$$\begin{tikzcd}\text{Ar-T} + {}^{-}\text{OMe} \arrow[r, shift left, "k_1"] & \text{Ar}^{-}\ \text{T-OMe} \arrow[l, shift left, "k_{-1}"] \arrow[d, "k_2"] \\ & \text{exchange}\end{tikzcd}$$

for which

$$k_{\text{exptl}} = k_1k_2/(k_{-1} + k_2) \qquad (1)$$

For $k_{-1} \gg k_2$, this expression reduces to

$$k_{\text{exptl}} = k_1k_2/k_{-1} = Kk_2 \qquad (2)$$

which is the product of an equilibrium constant having essentially no isotope effect and a rate constant for diffusion away of a solvent molecule, a step which should also have a negligible isotope effect.

It seems reasonable to expect the return step, k_{-1}, to be especially fast when the conversion of anion to hydrocarbon requires little structural reorganization. The present example is such a case—the structure of the aryl anion is expected to be essentially that of the hydrocarbon with a proton removed. By contrast, k_D/k_T has the value of 2.2 for the exchange of fluorene-9-*d*(*t*) with methanolic sodium methoxide.[12] This value extrapolates to a relatively normal k_H/k_D of 5.8 using the equation for interrelating isotope effects of Swain, *et al.*[13] It should also be noted that the rate of this exchange of fluorene is of the same magnitude as the reactions discussed here: $k_T = 43 \times 10^{-5}$ l. mol^{-1} sec^{-1} at 45°.[14] In the fluorene case internal return is not important;[15] in conversion of the anion to the hydrocarbon, the 9-hydrogen must bend out of the plane of the molecule and the loss of the delocalized charge is probably accompanied by other structural reorganizations—bond lengths and angles, etc.

In the exchange of pentafluorobenzene with methanol-*d*(*t*) (run 17 in Table I), the rate of incorporation of deuterium can be shown to be equal to the rate of loss of hydrogen or k_H *in this solvent* whereas the rate of incorporation of tritium is the same as the rate expected for loss of tritium from pentafluorobenzene-*t* in methanol-*d*, k_T. The resulting $k_H/k_T = 0.94$ is not only small but slightly inverse. Incidentally, comparison of k_T for this case with that in methanol-*t* gives $K(\text{MeOD})/k(\text{MeOH}) = 2.3$, a normal value for such a solvent isotope effect that undoubtedly reflects the effectively greater basicity of methoxide ion in methanol-*d* as compared to methanol.

A primary isotope effect is also available for 1,3-difluorobenzene. Hine and Langford[17] studied the loss of deuterium with methanolic sodium methoxide and found $k_D = 2.9 \times 10^{-6}$ l. mol^{-1} sec^{-1} at 50°. Our interpolated rate for k_T at 50° is 2.5×10^{-6} l. mol^{-1} sec^{-1}. These numbers are probably within the combined experimental errors and for this compound also there is no significant primary isotope effect.

From the data at different temperatures the activation parameters were calculated and are summarized in Table III. Note that the 2×10^5 greater reactivity of pentafluorobenzene over 1,3-difluorobenzene comes primarily from the energy of activation term. The entropies of activation are all not far from zero and not at all the ~−20 eu expected for a normal second-order reaction. For the internal return mechanism the experimental ΔS^* is a composite of the entropies of several steps, an equilibrium, for which ΔS^* might well be small in magnitude, and a diffusion step, for which ΔS^* might also have a small magnitude. Consequently, such activation parameters may also be diagnostic of internal return.

Table III. Activation Parameters for Tritium Exchange with Methanolic Sodium Methoxide

Compound	ΔH^*, kcal/mol	ΔS^*, eu
Pentafluorobenzene	18.0 ± 0.2	−5.6 ± 0.7
1,2,3,4-Tetrafluorobenzene	27.3 ± 1.4	−0.3 ± 4.6
1,3-Difluorobenzene[a]	27.9 ± 0.7	1.9 ± 2.3

[a] Hine and Langford[17] report $\Delta H^* = 27.3 \pm 0.2$ kcal/mol and $\Delta S^* = -4.3 \pm 0.7$ eu for deuterium exchange.

If we assume partial rate factors to apply to these polyfluorobenzenes we may write

$$\log k(\text{F}_5) = 2o_f + 2m_f + p_f + \log k_0$$

$$\log k(1,2,4,5\text{-F}_4) = 2o_f + 2m_f + \log k_0$$

$$\log k(1,2,3,4\text{-F}_4) = o_f + 2m_f + p_f + \log k_0$$

$$\log k(1,2,3,5\text{-F}_4) = 2o_f + m_f + p_f + \log k_0$$

$$\log k(1,3\text{-F}_2) = 2o_f + \log k_0$$

The best least-squares fit to the overdetermined set of parameters is: o_f, 5.25 ± 0.32; m_f, 2.07 ± 0.16; p_f, 1.13 ± 0.26; log k_0, −16.80 ± 0.71, at 40°, with a standard deviation of fit of ±0.25.[18] From eq 2, the relative rates are given by

$$k_{\text{exptl}}/k^0_{\text{exptl}} = (K/K^0)(k_2/k_2^0) \qquad (3)$$

If the diffusion rate constant, k_2, has the same value for each of a series of related compounds the last term in eq 3 reduces to unity and the experimental relative rates are directly the ratio of equilibrium acidities. With this approach the logarithms of the relative exchange rates summarized in Table II become relative p*K* values and the derived logarithms of partial rate factors above become p*K* differences between benzene and fluorobenzene hydrogens in methanol. However, these numbers are remarkably similar to the log relative exchange rate determined for *o*-, *m*-, and *p*-tritiums in fluorobenzene toward lithium cyclohexylamide (LiCHA) in cyclohexylamine; namely, *ortho*, 5.65; *meta*, 1.93; *para*, 0.96.[2] This correspondence implies a Brønsted correlation with a proportionality constant, β, of unity. Such a result is normally expected only when the reverse step of reaction of carbanion with solvent is diffusion controlled.[19] The LiCHA-catalyzed exchange is undoubtedly not of

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(12) A. H. Pudjaatmaka, results to be published.

(13) C. G. Swain, E. C. Stivers, J. F. Reuwer, Jr., and L. J. Schaad, *J. Am. Chem. Soc.*, **80**, 5885 (1958).

(14) A. Streitwieser, Jr., A. P. Marchand, and A. H. Pudjaatmaka, *ibid.*, **89**, 694 (1967).

(15) Ford, Graham, and Cram[16] have shown recently by ingenious kinetic and stereochemical techniques that some internal return does occur even with fluorene derivatives. The amount of such return is comparatively small and does not significantly affect the argument that follows.

(16) W. T. Ford, E. W. Graham, and D. J. Cram, *J. Am. Chem. Soc.*, **89**, 689, 690 (1967).

(17) J. Hine and P. B. Langford, *J. Org. Chem.*, **27**, 4149 (1962).

(18) This calculation was carried out by Dr. E. Dart using a standard computer program.

(19) M. Eigen, *Angew. Chem. Intern. Ed. Engl.*, **3**, 1 (1964).

this type because of observed large primary isotope effects. Diffusion-controlled protonations are expected to have small primary isotope effects; since the equilibrium isotope effect is also normally close to unity, the primary isotope effect for the forward reaction for such cases is likewise small.

The cyclohexylamine reactions are ion-pair reactions. The reverse reaction of protonation of a carbanion is actually a protonation of a lithium aryl with formation of a LiCHA ion pair. The involvement in this reaction of a lithium cation with its solvation demands undoubtedly accounts for the comparative slowness of the reverse protonation. The lithium cation is the moiety undergoing structural reorganization! For such a case the rate of the reverse protonation might well be approximately independent of structure for a related series and would give the effect of a Brønsted correlation for exchange with $\beta = 1$. Note that this type of argument probably does not apply to exchange reactions that involve delocalized carbanions, for here the organic moiety is also undergoing structural reorganization; that is, such different types of C–H bonds cannot be put on the same Brønsted scale.

Finally, we note that the log k_0 derived above, -16.8 at 40°, represents an extrapolated exchange rate of benzene itself. We are currently measuring this quantity directly and defer further comment until the completion of such studies.

Acknowledgment. We are pleased to acknowledge helpful conversations with Professors D. J. Cram and C. D. Ritchie.

Selective Catalytic Effects of Strongly Ionizing Polycations on Ester Solvolysis[1]

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Received September 8, 1967

Abstract: The cationic polyion, poly(1-vinyl-3-methylimidazolium iodide) (PVMI), was found to enhance the solvolytic rates of the negatively charged esters 4-acetoxy-3-nitrobenzoic acid and sodium 4-acetoxy-3-nitrobenzenesulfonate in 28.5% ethanol–water solutions. These rates were unaffected by the monomeric analog of the polyion, 1,3-dimethylimidazolium iodide. At high pH values, the enhanced solvolytic rates of the negatively charged esters may be accounted for by the high local concentration of hydroxyl counterions in the vicinity of the polymer chain. At intermediate pH values, the large catalytic effect is attributed to an enhanced susceptibility of the anionic ester to direct water attack in the vicinity of the polycation. The solvolyses of uncharged esters (*p*-nitrophenyl acetate and *p*-nitrophenyl hexanoate) were not accelerated, indicating that hydrophobic forces are insufficient to concentrate the esters in the neighborhood of the polyion. A copolymer of 1-vinyl-3-methylimidazolium iodide (VMI) containing 86 mol % vinyl alcohol residues catalyzed the hydrolysis of anionic esters with an efficiency similar to that of the PVMI homopolymer, if the two were compared at equal stoichiometric concentrations of imidazolium groups. On the other hand, a copolymer of VMI containing 63 mol % *p*-vinylphenol residues was about twice as effective as PVMI. This is due to the fact that the polycation draws the anionic ester into a region with a high local concentration of the phenoxide nucleophile.

Effects of polyions on the rates of reactions of low molecular weight species may be classified into two categories. (1) The large local fluctuations in the electrostatic potential characterizing polyelectrolyte solutions lead to corresponding fluctuations in the local concentration of charged species. If a reaction involves two ionic species with a charge of the same sign, a polyion of opposite sign will tend to concentrate these reagents in its immediate neighborhood and enhance the rate of their collisions with each other. The polyion will then act as a catalyst. Conversely, if the two reagents carry charges of opposite sign, the polyion will attract one and repel the other, and it will act as an inhibitor. The quantitative analysis of this effect may be used to characterize the distribution of the electrostatic potential in polyelectrolyte solutions.[3] As typical examples of these phenomena we may cite the powerful catalysis of the benzidine rearrangement (involving an attack of a hydrogen ion on the conjugate acid of hydrazobenzene) by a polysulfonic acid[4] and the inhibition of the hydroxide ion catalyzed hydrolysis of a cationic ester in the presence of polymeric acids.[3b] (2) A second effect will be observed if the polyion carries catalytically active substituents. In this case, a reactive counterion may be attracted to the polyion where it finds itself in a region of high local concentration of the catalytic groups. Such a sit-

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(1) This paper comprises a portion of the dissertation submitted by J. C. Salamone in partial fulfillment of the requirements for the degree of Doctor of Philosophy in the Graduate School of the Polytechnic Institute of Brooklyn. A part of this work has been published in a preliminary form: H. Morawetz, C. G. Overberger, J. C. Salamone, and S. Yaroslovsky, *J. Polymer Sci.*, **B4**, 609 (1966).

(2) (a) The University of Michigan, Ann Arbor, Mich. (b) Deceased, July 1967.

(3) (a) H. Morawetz, *J. Polymer Sci.*, **42**, 125 (1960); (b) H. Morawetz and J. A. Shafer, *J. Phys. Chem.*, **67**, 1293 (1963).

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