each center until self-consistency is obtained. The coulomb integrals were adjusted according to $\alpha_r = \alpha_0 + \omega\beta_0(1 - q_r)$, where $\omega = 0.8$ and $q_r$ is the electron density at atom $r$ determined from the previous iteration. The resonance integrals were readjusted according to $\beta_{rs} = \beta_0(S_{rs}/0.275)$, where $S_{rs}$ represents an approximate overlap integral between atoms $r$ and $s$. The overlap integral was assumed to be proportional to the total bond order between atoms $r$ and $s$ and was calculated using $S_{rs} = \mathrm{EL}\ P_{\mathrm{TOTAL}} + \mathrm{CAY}$, where $P_{\mathrm{TOTAL}}$ equals the sum of the $\pi$ and $\sigma$ bond order, and EL = 0.08 and CAY = 0.115 as suggested by Mulliken.[38]

(40) J. Pople and D. Beveredge, "Approximate M.O. Theory," McGraw-Hill, New York, N. Y., 1970, and papers referenced therein. The computer program is the INDO portion of QCPE 141 available from QCPE, Indiana University.

(41) J. W. Crickshank, *Acta Crystallogr.*, **10**, 504 (1957).

(42) M. A. Lasheen, *Acta Crystallogr.*, **5**, 593 (1952).

(43) F. H. Allen and J. Trotter, *J. Chem. Soc. B*, 916 (1970).

(44) E. B. Fleischer, *J. Amer. Chem. Soc.*, **86**, 3889 (1964).

(45) G. L. Hardgrove, L. K. Templeton, and D. H. Templeton, *J. Phys. Chem.*, **72**, 668 (1968).

(46) Optimum values for the parameters $h_1$ and $h_2$ could have been obtained by the perturbation method of Moss, *et al.*[47]

(47) R. E. Moss, N. A. Ashford, R. G. Lawler, and G. K. Fraenkel, *J. Chem. Phys.*, **51**, 1765 (1969).

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# Secondary Valence Force Catalysis. XV. Polysoap Catalysis for the Alkaline Hydrolysis of *p*-Nitrophenyl Hexanoate[1]

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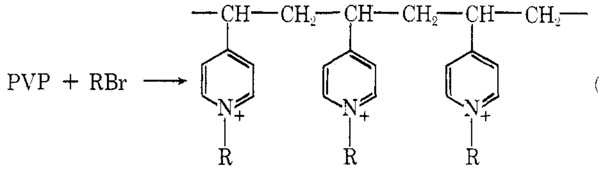
The alkaline hydrolysis of *p*-nitrophenyl hexanoate is subject to catalysis by polysoaps synthesized through alkylation of polyvinylpyridine with ethyl and dodecyl bromides. The polysoaps are increasingly effective as catalysts for this reaction with increasing ratio of dodecyl to ethyl groups, suggesting that the hydrophobic properties of these polymers, as well as their net charge, are important in determining catalytic efficacy. The polysoap-catalyzed hydrolysis of this ester is inhibited by several anions: fluoride is the weakest inhibitor studied and nitrate is the most potent. High concentrations of nitrate convert the polysoap-catalyzed process into a polysoap-inhibited one. In contrast to other anions, azide elicits an increased rate of ester disappearance, strongly suggesting that polysoaps catalyze the attack of this nucleophile on *p*-nitrophenyl hexanoate.

It is well established that both polyelectrolytes[2] and micelles[3] may be effective catalysts for organic reactions, including ester hydrolysis. These reactions are of particular interest since they exhibit many features, including limited substrate specificity, concentration-rate profiles, inhibition by structurally related substances, and inhibition by salts, which are reminiscent of enzymatic reactions.[2,3] Catalysis in such systems generally derives from association of substrate and catalyst through electrostatic and/or hydrophobic interactions with formation of a complex having enhanced reactivity compared to the substrate free in solution. This enhanced reactivity may either reflect stabilization of the transition state relative to the ground state for reaction with some group from the bulk phase or incursion of an entirely novel pathway through reaction with some functionality of the micelle or polyelectrolyte. The *n*-alkyltrimethylammonium bromide catalyzed hydrolysis of *p*-nitrophenyl esters provides one example in the former category[4] and catalysis of carboxylic and sulfate ester hydrolysis by polyethylenimines substituted with imidazole functions provides examples in the latter one.[5]

In general, micelles formed from ionic surfactants are organized and compact structures having the capacity to interact favorably with both organic molecules and ions. In contrast, most polyelectrolytes do not form well-organized structures in aqueous solution and usually rely on electrostatic forces more than hydrophobic ones for complexation with small molecules and promotion of their reactions. A class of compounds has been known for several decades which incorporates both those features of polyelectrolytes and those of micelles into a single structure; they are usually termed polysoaps.[6-8] A typical polysoap can be formed through treatment of polyvinylpyridine with dodecyl bromide or other hydrophobic alkylating agent. The result is a polyelectrolyte having marked hydrophobic properties and such structures are known to form "intramolecular micelles" having a structural organization related to micelles formed from simple surfactants.[6]



(1)

Very few studies of the catalytic capacities of polysoaps have been undertaken. It has been established that polyvinylpyridine partially alkylated with benzyl chloride promotes the hydrolysis of *p*-nitrophenyl esters through a route involving transient acylation of unsubstituted pyridine nitrogens of the polymer.[9] More recently, alkylated polyvinylpyridines have been shown to be catalysts for fading of triphenylmethyl cations such as crystal violet.[10] The work of Klotz and his collaborators has employed polyethylenimines possessing substantial hydrophobic character[5] and Ise, *et al.*, have shown that related polymers catalyze hydrolysis of phenyl phosphate dianions.[11] However, in these latter cases, there is no evidence to suggest that the polyelectrolytes actually form micelle-like structures in solution.

Since the catalytic properties of polysoaps remain largely unexplored, we have initiated a series of investigations in this area. The first of these, dealing with hydrolysis of *p*-nitrophenyl hexanoate, is described in this manuscript.

## Experimental Section

**Materials.** *p*-Nitrophenyl hexanoate was prepared and purified as previously described.[4] Poly-4-vinylpyridine was obtained from Polysciences, Inc.; the supplier provided an approximate molecular weight of 300,000.

Copolymers of 4-vinyl-*N*-ethylpyridinium bromide and 4-vinyl-*N*-dodecylpyridinium bromide were prepared by a minor modifi-

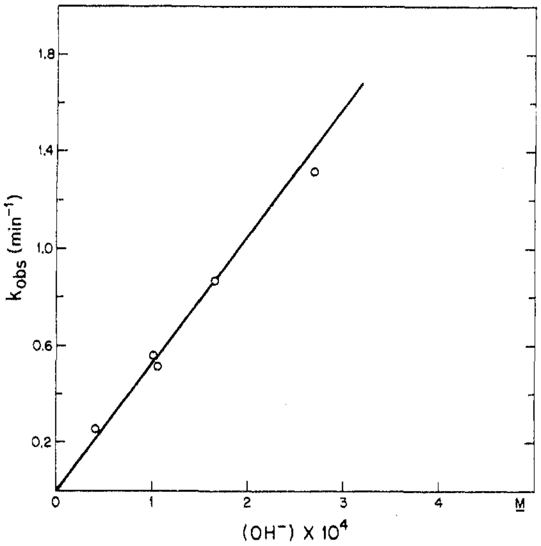


**Figure 1.** First-order rate constants for hydrolysis of *p*-nitrophenyl hexanoate in the presence of 6.5 × $10^{-7}$ *M* polysoap (28% alkylated with dodecyl groups, PS-28) plotted as a function of the concentration of hydroxide ion at 30°. Values of pH were maintained with dilute triethylamine buffers.

cation of the method of Strauss and Gershfeld.[6] One gram of poly-4-vinylpyridine was dissolved in 28 ml of a 1:1 nitromethane–nitroethane mixture and traces of insoluble material were removed by filtration; 50 ml of redistilled dodecyl bromide was then added and the reaction was permitted to proceed for various time periods, depending on the extent of alkylation desired, at a temperature of 50°. The reaction mixture was then poured into three volumes of cold ethyl acetate and the precipitated polymer was collected by filtration and dried *in vacuo*. This preparation was dissolved in hot butanone to which was added sufficient ethanol to effect total solution and the polymer was reprecipitated by diluting this solution into four volumes of ethyl acetate at 0°. This procedure was repeated. This partially alkylated polymer was treated with excess ethyl bromide in nitromethane at 50° for 72 hr. At the completion of the reaction, the mixture was poured into three volumes of ethyl acetate at 0° and the precipitated polysoap was collected by centrifugation and purified through repeated precipitation as described above.

The extent of total alkylation of the polysoaps was estimated through potentiometric titration of bromide ion. Extent of alkylation with dodecyl groups was measured as described by Strauss, *et al.*[6] Three polysoaps were prepared and analyzed as described; these contained 15, 28, and 38% alkylation with dodecyl groups and at least 95% total alkylation. For convenience, these have been termed PS-15, PS-28, and PS-38.

**Kinetic measurements** were made spectrophotometrically employing a Zeiss PMQ II spectrophotometer equipped with a cell compartment through which water from a bath maintained at 30° was continuously circulated.[4] All measurements were made on solutions containing approximately 5 × $10^{-5}$ *M* *p*-nitrophenyl hexanoate at a wavelength of 400 nm. Good first-order rate behavior was observed throughout. First-order rate constants were calculated from plots of the logarithm of the difference between optical density at infinite time and optical density at various times against time, in the usual manner. Values of pH were maintained with dilute triethylamine-ammonium buffers and were measured with the aid of a Radiometer pH meter.

## Results

First-order rate constants for hydrolysis of *p*-nitrophenyl hexanoate at 30° in the presence of 6.5 × $10^{-7}$ *M* PS-28 were measured as a function of the concentration of hydroxide ion; results are presented graphically in Figure 1. A satisfactory straight-line relationship is obtained, indicating that the rate law for ester hydrolysis under these conditions is simply $k_{obsd} = k_2(OH^-)$. That is, the rate of ester hydrolysis is quantitatively accounted for by reaction with hydroxide ion, ruling out significant reaction with the small percentage of unalkylated sites on the polysoap or the incursion of a polysoap-directed alternative reaction pathway. The data in Figure 1 yield a calculated second-order rate constant of 5.3 × $10^3$ $M^{-1}$ min$^{-1}$; this is substantially greater than the corresponding value of 3.4 × $10^2$ $M^{-1}$ min$^{-1}$ for the same reaction in the absence of polysoap,[4] indicating that the polysoap is a catalyst for *p*-nitrophenyl hexanoate hydrolysis.

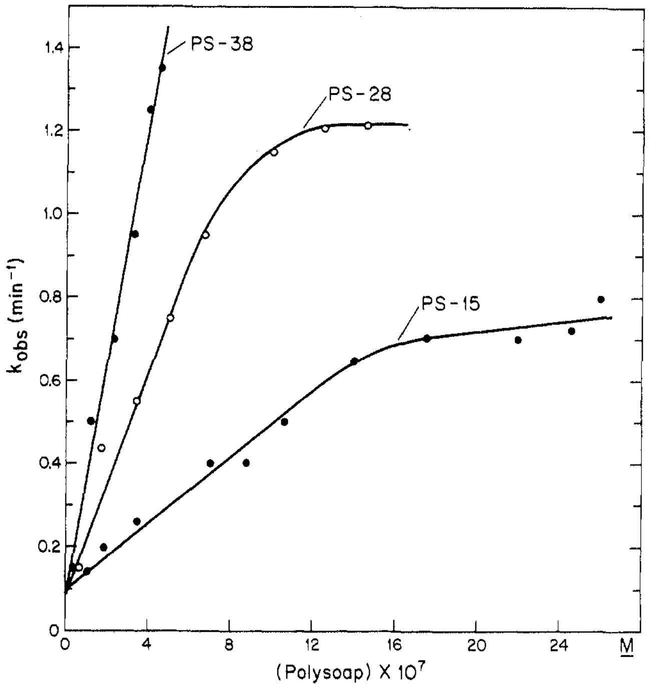


**Figure 2.** First-order rate constants for hydrolysis of *p*-nitrophenyl hexanoate plotted as a function of the concentration of polysoaps alkylated to the extent of 15, 28, and 38% with dodecyl groups (PS-15, PS-28, and PS-38). Values of pH were maintained between 9.95 and 10.1 with dilute triethylamine buffers.

In order to quantitate the apparent catalysis of ester hydrolysis by polysoaps, first-order rate constants were measured at 30° and values of pH near 10 as a function of the concentration of PS-15, PS-28, and PS-38. Results are provided in Figure 2. Note that each of the polysoaps is a catalyst for the hydrolytic reaction. At low polysoap concentrations, rate constants increase linearly with this variable. However, for the cases of PS-15 and PS-28 at least, the rate constants eventually level off and become independent, or nearly so, of polysoap concentration. The relative insolubility of PS-38 prevented examination of this point for that polysoap. Note that the more hydrophobic the polysoap, the more effective the catalysis, whether efficacy is judged in terms of the initial slope of rate constant against polysoap concentration or in terms of the maximal rate constants observed.

In Figure 3, first-order rate constants for *p*-nitrophenyl hexanoate hydrolysis at 30° are plotted as a function of the concentration of PS-28 in the absence and presence of two concentrations of added potassium chloride. Note that the extent of the catalysis, as judged by the slope of rate constant against catalyst concentration, is markedly reduced by the addition of 0.05 *M* KCl and is nearly abolished in the presence of 0.5 *M* KCl. This finding suggests that salts, almost certainly the anionic component, are inhibitors of polysoap-catalyzed ester hydrolysis.

Salt inhibition of the polysoap-catalyzed reaction was further pursued by examination of the effects of several sodium and potassium salts. First-order rate constants for *p*-nitrophenyl hexanoate hydrolysis at 30° in the presence of 6.5 × $10^{-7}$ *M* PS-28 at pH 10.1–10.2 as a function of the concentration of these salts are collected in Figure 4. Fluoride, chloride, sulfate, and nitrate are all observed to

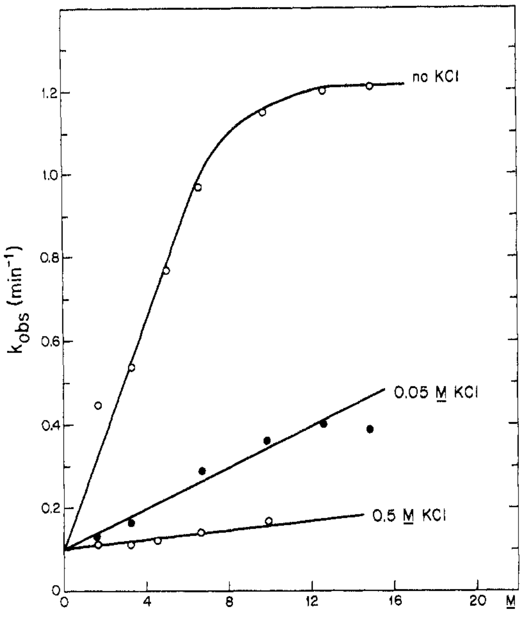


**Figure 3.** First-order rate constants for the hydrolysis of *p*-nitrophenyl hexanoate at 30° plotted as a function of the concentration of polysoap alkylated to the extent of 28% with dodecyl groups (PS-28) in the absence and the presence of two concentrations of potassium chloride. Values of pH were maintained at 10.1-10.2 through use of dilute buffers of triethylamine.

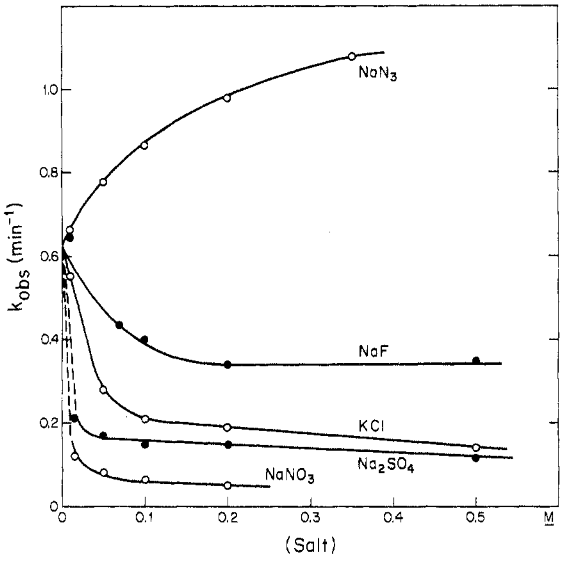


**Figure 4.** First-order rate constants for hydrolysis of *p*-nitrophenyl hexanoate at 30° in the presence of $6.5 \times 10^{-7}$ *M* polysoap alkylated to the extent of 28% with dodecyl groups (PS-28) plotted as a function of the concentration of several salts. Values of pH were maintained between 9.95 and 10.05 through use of dilute triethylamine buffers.

be effective inhibitors, in that order of efficiency. Note that a concentration of nitrate ion of 0.02 *M* is sufficient to reduce the observed rate constant to a value near that which is observed in the absence of polysoap, *ca.* 0.1 min$^{-1}$. At higher concentrations of this anion, rate constants are actually less than this value so that the polysoap-catalyzed reaction is converted into a polysoap-inhibited one.

Among the anions studied, azide has the effect of increasing the rate of ester disappearance in the presence of polysoap (Figure 4). This nucleophile is quite reactive against *p*-nitrophenyl esters[12] so that this observation almost certainly reflects direct nucleophilic attack of azide on *p*-nitrophenyl hexanoate, very probably catalyzed by the polysoap (see below).

## Discussion

Data presented in Figures 1 and 2 establish that copolymers of 4-vinyl-*N*-ethylpyridinium bromide and 4-vinyl-*N*-dodecylpyridinium bromide are effective catalysts for the attack of hydroxide ion on *p*-nitrophenyl hexanoate. In the case of two of the polysoaps studied, saturation-type kinetics were observed (Figure 2), strongly suggesting the formation of a complex between polysoap and ester preceding bond-changing reactions. Similar kinetic behavior has been observed for catalysis of the hydrolysis of the same ester by cationic surfactants.[4] If this interpretation is correct, it follows that the shape of the rate-concentration profiles can be accounted for on the basis of the equilibrium constants for association of ester with polysoap

$$\text{ester} + \text{polysoap} \overset{K_{eq}}{\rightleftharpoons} \text{ester} \cdot \text{polysoap} \tag{2}$$

and rate constants for decomposition of the complexes

$$\text{ester} \cdot \text{polysoap} + \text{OH}^- \xrightarrow{k_2} \text{polysoap} + \text{products} \tag{3}$$

Values for $K_{eq}$ and $k_2$ for reactions involving PS-15 and PS-28 can be crudely estimated from the data in Figure 2. For PS-15, values of $K_{eq} = 1.2 \times 10^6\ M^{-1}$ and $k_2 = 7.5 \times 10^3\ M^{-1}\ \text{min}^{-1}$ are obtained; for PS-28, the corresponding values are $K_{eq} = 2.5 \times 10^6\ M^{-1}$ and $k_2 = 1.2 \times 10^4\ M^{-1}\ \text{min}^{-1}$. Since limits of solubility precluded observing saturation kinetics with PS-38, one can only state qualitatively that values for both constants must be larger than those observed for the less hydrophobic polysoaps.

Polysoap catalysis for alkaline ester hydrolysis must be ascribed in part at least to electrostatic stabilization of the anionic transition state compared to the ground state, uncharged ester associated with the polysoap plus hydroxide ion free in solution, by the cationic polysoap. A similar explanation has been provided for catalysis of the same reaction by cationic micelles.[4] However, hydrophobic interactions appear important to catalytic efficacy as well. The increasing slope of rate-concentration profiles with increasing polysoap hydrophobicity can be accounted for in terms of stronger binding of ester to polysoap, as one might expect. However, the observation that the maximal rate constants achieved are also increased by increasing polysoap hydrophobicity indicates that hydrophobic interactions contribute either directly or indirectly (perhaps through accentuation of the electrostatic interactions) to activation energies.[13,14] Similar observations have been made for surfactant-catalyzed hydrolysis of *p*-nitrophenyl hexanoate: the more hydrophobic the *n*-alkyltrimethylammonium salt, the greater the extent of catalysis observed.[4] Moreover, increasingly hydrophobic polysoaps have been observed to be increasingly effective catalysts for the fading of cationic triphenylmethyl dyes.[10] In related work, polyelectrolyte catalysis for hydrolysis of carboxylic and sulfate esters[5] and phenyl phosphate dianions[11] is accentuated by the hydrophobic character of the catalysts.

From the foregoing account, it is clear that catalysis of the attack of hydroxide ion of *p*-nitrophenyl hexanoate by cationic micelles and by polysoaps have much in common. We may ask which system is the more effective catalyst. The maximum catalysis observed for the polysoaps is 14-fold, achieved in the presence of $4.5 \times 10^{-7}$ *M* PS-38 (Figure 2). Under very similar conditions, dodecyltrimethylammonium bromide, at an optimal concentration of 0.02 *M*, increases the hydrolytic rate by less than threefold.[4] The best surfactant catalyst known, octyltrimethylammonium bromide, increases the rate five- to sixfold at a concentration of $6 \times 10^{-4}$ *M*.[4] By way of comparison, a simple calculation reveals that the concentration of dodecyl groups in a $4.5 \times 10^{-7}$ *M* solution of PS-38 is about $5 \times 10^{-4}$ *M*. By whatever criterion the comparison is made, polysoaps are better catalysts for ester hydrolysis than are simple cationic surfactants. However, it should be noted that catalysis for phenyl ester hydrolysis more marked than that of polysoaps has been observed for both normal and inverted micelles of hexadecyltrimethylammonium bromide in alcohol–water mixtures.[15]

Like micelle-catalyzed phenyl ester hydrolysis,[4] the polysoap-catalyzed reaction is subject to marked inhibition by anions (Figures 3 and 4). The order of effectiveness of the various ions studied increases with increasing anion hydrophobicity and parallels the order previously established for micellar systems.[4] Anion inhibition for polysoap-catalyzed hydrolysis of *p*-nitrophenyl hexanoate hydrolysis may result from (i) neutralization of surface charge on the polysoap micelle, reducing the electrostatic facilitation of the reaction; (ii) changes in shape of the polysoap in solution; and (iii) exclusion of the ester substrate from association with the polysoap; or any combination of these (mutually dependent to some extent) possibilities.

There is evidence to suggest that the bulk of the anion inhibition effect does not derive from disruption of the polysoap–ester complex. The clearest case is provided by nitrate: it is difficult to account for a *net* inhibition of the reaction on the basis of dissolution of the complex. Were this the case, the maximum effect that could be observed would be the return of the rate constant to that value characteristic of a polysoap-free medium. The fact that, in the presence of 0.1 *M* or more of nitrate, the polysoap actually decreases the rate suggests some form of polysoap–ester interaction.

It has been well established that those polysoaps containing 13.6% or more alkylation with dodecyl groups undergo shape changes with increasing electrolyte concentration.[6] Specifically, at modest concentrations of bromide, polysoaps form structures more compact than those which can be achieved by ordinary polymers or polyelectrolytes. It is likely that formation of more compact structures may account in part at least for the anion inhibition of the polysoap-catalyzed reactions. It is, of course, not possible to dissect the effect of the shape change *per se* on reaction rate from the effect of charge neutralization, which must underlie the shape change, on reaction rate. In analogy to results in micellar systems, it seems quite likely that charge neutralization would diminish the catalytic effect through reduction of electrostatic stabilization of the transition state even if no shape change were to occur. However, the change in shape may modify the affinity of polysoap for ester, may modify the access of hydroxide to complexed ester, and may modify the site of localization of ester with respect to charged groups of the polysoap, changes which all should be reflected in alterations in reaction rate.

The increased rate of *p*-nitrophenol release from *p*-nitrophenyl hexanoate in the presence of azide is most easily interpreted as polysoap-catalyzed attack of azide on ester. The alternative involving azide-promoted displacement of ester from the complex with polysoap combined with azide attack on the free ester appears unlikely. First, were the latter alternative to obtain, one would expect the reaction rate to increase at least with the first power of azide concentration. In fact, the rate increases more slowly than that and appears to approach a plateau value (Figure 4). Second, the rate constants observed are somewhat larger than would be predicted on the basis of the reactivity of azide ion toward *p*-nitrophenyl acetate.[12] Finally, since it has been established that polysoaps catalyze the attack of hydroxide ion on the ester, there is no reason to suppose that such catalysis would not be observed for attack of azide ion as well.

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**Registry No.**—*p*-Nitrophenyl hexanoate, 956-75-2.

## References and Notes

(1) Supported by Grant GM-08232 from the National Institutes of Health and by the Facultad de Ciencias, Universidad Central.

(2) For a review of polyelectrolyte catalysis for organic reactions, see H. Morawetz, *Advan. Catal. Relat. Subj.*, **20**, 341 (1969); H. Morawetz, *Accounts Chem. Res.*, **3**, 354 (1970); C. G. Overberger and J. C. Salamone, *ibid.*, **2**, 217 (1969); N. Ise, *Advan. Polym. Sci.*, **7**, 536 (1971).

(3) For a review of micelle catalysis for organic reactions, see E. H. Cordes and R. B. Dunlap, *Accounts Chem. Res.*, **2**, 329 (1969); E. J. Fendler and J. H. Fendler, *Advan. Phys. Org. Chem.*, **8**, 271 (1970); E. H. Cordes and C. Gitler, *Prog. Bioorg. Chem.*, **2**, 1 (1973).

(4) L. R. Romsted and E. H. Cordes, *J. Amer. Chem. Soc.*, **90**, 4404 (1968); M. T. A. Behme, J. G. Fullington, R. Noel, and E. H. Cordes, *ibid.*, **87**, 266 (1965).

(5) I. M. Klotz and V. H. Stryker, *J. Amer. Chem. Soc.*, **90**, 2717 (1968); I. M. Klotz, G. P. Royer, and I. S. Scarpa, *Proc. Nat. Acad. Sci. U. S.*, **68**, 263 (1971); H. C. Kiefer, W. I. Congdon, I. S. Scarpa, and I. M. Klotz, *ibid.*, **69**, 2155 (1972).

(6) U. P. Strauss and L. H. Layton, *J. Phys. Chem.*, **57**, 352 (1953); U. P. Strauss and N. L. Gershfeld, *ibid.*, **58**, 747 (1954); U. P. Strauss, N. L. Gershfeld, and E. H. Crook, *ibid.*, **60**, 577 (1956); U. P. Strauss and S. S. Slowata, *ibid.*, **61**, 411 (1957); E. G. Jackson and U. P. Strauss, *J. Polym. Sci.*, **6**, 649 (1951); U. P. Strauss, S. J. Assony, E. G. Jackson, and L. H. Layton, *ibid.*, **9**, 509 (1952).

(7) S. K. Sinha, A. I. Medalia, and D. R. Harrington, *J. Amer. Chem. Soc.*, **79**, 281 (1957).

(8) H. H. Freedman, J. P. Mason, and A. I. Medalia, *J. Org. Chem.*, **23**, 76 (1958).

(9) Y. E. Kirsh, V. A. Kabanov, and V. A. Kargin, *Proc. Acad. Sci. USSR*, **177**, 976 (1967).

(10) T. Okubo and N. Ise, *J. Amer. Chem. Soc.*, **95**, 2293 (1973).

(11) T. Ueda, S. Harada, and N. Ise, *Polym. J.*, **3**, 476 (1972).

(12) W. P. Jencks and J. Carriuolo, *J. Amer. Chem. Soc.*, **82**, 1778 (1960).

(13) J. Baumrucker, M. Calzadilla, M. Centeno, G. Lehrmann, M. Urdaneta, P. Lindquist, D. Dunham, M. Price, B. Sears, and E. H. Cordes, *J. Amer. Chem. Soc.*, **94**, 8164 (1972).

(14) C. Gitler and A. Ochoa-Solano, *J. Amer. Chem. Soc.*, **90**, 5004 (1968).

(15) S. Friberg and S. I. Ahmad, *J. Phys. Chem.*, **75**, 2001 (1971); S. I. Ahmad and S. Friberg, *J. Amer. Chem. Soc.*, **94**, 5196 (1972); S. Friberg, L. Rydhag, and G. Lindblom, *ibid.*, in press.