even quite efficient scavengers. It is commonly, though erroneously, assumed that even appreciable concentrations of scavenger cannot intervene in the "cage effect." This present work is consistent with our previous observations[9,10] that iodine at moderate concentrations intervenes in step (5) to react with free radicals still capable of "geminate recombination"[13] (step 4)).

Employing methods previously described[9,10] let the probability of non-recombination be identified with one-half the yield of organic iodide (steps (3) plus (5)). Then

$$\log \Phi = 0.31 P_1 P_2^{1/2} \gamma^{-2} X^{1/2} - 0.17 P_1 \gamma^{-1} \rho_0^{-1}$$

where P_1 and P_2 are probabilities of reaction upon encounter for steps (4), (5); γ and ρ_0 are relative mean free path and initial separation of radicals R· in units of collision diameters. From the diffusion equation and Fig. 1 we find $P_1\gamma^{-1}\rho_0^{-1} = 6$ and $P_1 \cdot P_2^{1/2}\gamma^{-2} = 19$. For the ratio $P_2^{1/2}\gamma^{-1}\rho_0 = 3$. No great error can arise from taking $P_2 = 0.25$, for which the initial separation R . . . . R is six mean free paths. If the initial process for photodecomposition is that postulated in step (1), then the mean free path for diffusion is rather small but is consistent with our previous work.[9,10] It is also consistent with very recent considerations of Noyes.[14]

Since it appears from the photochemical measurements that steps (4) and (5) compete effectively it follows that absence of similar diffusion effects for thermal decomposition of AIBN (Table I) indicates different primary products for the two modes of decomposition. Whatever free radicals are produced by thermal decomposition, they appear to be incapable of recombination upon one or two encounters. There appear to be two mechanisms for thermal decomposition, one of which corresponds to the constant ratio RI/N_2 = 2, whose relative contributions do not vary up to appreciable iodine concentrations. The second mechanism, applying to 40% of the reaction under our experimental conditions but somewhat dependent upon the solvent,[8] produces nitrogen but apparently produces no free radicals. It has been proposed that this mechanism consists of rearrangement to tetramethylsuccinodinitrile with elimination of nitrogen.[8]

Other recent work[15] demonstrates that thermal decomposition of AIBN at 80° in toluene produces an unstable intermediate amounting to about one-third of the whole reaction and eventually yielding tetramethylsuccinodinitrile. Under the same conditions but with added iodine we find an efficiency of 60% which indicates that free radicals are not involved in the formation of the intermediate or in its decomposition as postulated.[15]

Acknowledgment.—We are indebted to Professor R. M. Noyes for the opportunity to read an unpublished manuscript.

NOTRE DAME, INDIANA

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[14] R. M. Noyes, forthcoming publication.

[15] M. Talât-Erben and S. Bywater, THIS JOURNAL, **77**, 3710, 3712 (1955).

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY, COLUMBIA UNIVERSITY]

Rate Measurements on Fast Reactions in the Stirred Flow Reactor; the Alkaline Hydrolysis of Methyl and Ethyl Formate[1]

BY HARRIE M. HUMPHREYS AND LOUIS P. HAMMETT

RECEIVED AUGUST 29, 1955

The stirred flow reactor technique for measuring reaction rates in solution gains considerably in convenience and in generality of application without losing in precision when one of the reactants is injected into the reactor as a pure liquid or as a concentrated solution by a hypodermic syringe driven by a screw rotating at constant speed. The technique is found to be applicable to reactions with half times as short as 0.6 sec. Measurements by this method on the alkaline hydrolysis of ethyl formate in 85% ethanol and of methyl formate in water show that compared with acetate the several hundred-fold higher rate of reaction of the formate is essentially the result of a more positive entropy of activation, the enthalpies of activation for acetate and formate differing by unimportant amounts. This result, taken with available data on the acid-catalyzed esterification and on the ionization of acetic and formic acids appears to justify the generalization that the conversion of an electrically neutral acetic acid derivative to an electrically charged entity involves a considerably greater decrease in entropy than does the analogous process in the case of a formic acid derivative.

Thanks to the excellent work of H. A. Smith and his students[2] precise data are available on the temperature coefficients of the alkaline hydrolysis rates of a wide range of ethyl esters in 85% ethanol. But this list does not include the outstandingly important case of the formate ester, presumably because of the inconveniently high rate of reaction. Leimu, *et al.*,[3] have obtained data on several formates in water solution by a highly refined adaptation of the usual batch technique and find that the 200-fold difference in reaction rate between methyl formate and methyl acetate results almost exactly from a difference in entropy of activation, the activation energies being identical within experimental error. If their results are essentially correct, the matter is of very considerable theoretical interest, but the danger of systematic error in measurements of this sort on a reaction with a half-time measured in seconds is so large that a verification by a different technique is clearly desirable. It

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[1] Based on a dissertation submitted by Harrie M. Humphreys in partial fulfillment of the requirements for the degree of Doctor of Philosophy in the Faculty of Pure Science, Columbia University. The work was carried out as project NR 056-062, under contract N6-onr-271 between the Office of Naval Research and Columbia University. Reproduction in whole or in part permitted for any purpose of the United States Government.

[2] (a) H. A. Smith and H. S. Levenson, THIS JOURNAL, **61**, 1172 (1939); (b) H. S. Levenson and H. A. Smith, *ibid.*, **62**, 1556 (1940).

[3] R. Leimu, R. Korte, E. Laaksonen and U. Lehmuskoski, *Suomen Kemistilehti*, **19B**, 93 (1946).

has been shown[4] that the stirred flow reactor technique can give rate measurements of high precision in relatively rapid reactions, and we have therefore employed it in the study of the formate hydrolysis. We could not, however, operate as we have done previously by flowing into the reactor two dilute solutions in the reaction solvent, one of ester, the other of alkali, because the rapid hydrolysis of the formate ester would make the ester solution unstable. We have therefore modified the technique by introducing the undiluted ester directly into the reactor by a screw driven hypodermic syringe. The method turns out not merely to be precise, but also to offer significant advantages in convenience over those we have used previously.

Experimental

Apparatus.—Our reactor was a vertical cylinder of 32.60 ± 0.03-ml. volume. On the top a ground glass stopper carried a Western Electric 14-A thermistor cemented into place with Apiezon W. The electrical circuits used with the thermistor for measuring the reactor temperature were those of Rand and Hammett.[4c] The reactor contents were stirred by a plastic coated magnetized bar rotated by a rotating permanent magnet below the cell. The sodium hydroxide solution entered through a 1-mm. diameter tube sealed through the reactor wall half way up one side; the ester entered through a piece of 21 gage stainless steel tubing (inside diameter 0.02 inches) which was led through a rubber stopper in an opening in the reactor opposite the alkali inlet. The 1.5 mm. diameter outlet was attached to the upper part of the side of the reactor, was 12 cm. long and led to a continuous flow titration chamber which was essentially identical with that used by Saldick and Hammett.[4b] The flow of the sodium hydroxide solution was controlled and measured by Rand's[4c] constant flow pipet and by a capillary of suitable size between the pipet and the reactor. The flow of the hydrochloric acid used in the continuous titration was controlled and measured in the same way.

The flow of the ester was controlled by a hypodermic syringe whose plunger was driven by an accurate screw which was rotated at constant but adjustable speed by a synchronous motor operating through a change speed gear box. The positive control of the ester flow thus obtained eliminated much of the inconvenience and uncertainty associated with systems in which two flows are controlled by constant pressure-choke systems and still furnished a surprisingly constant flow rate. Careful calibrations showed variations of only 0.1% with a 2-ml. syringe at 0.00045 ml./sec. flow rate and 0.7% with a 20-ml. syringe at 0.01 ml./sec. when the plunger was started at different positions in the syringe barrel. With the very volatile formate esters it was necessary to prevent evaporation through the annulus between plunger and barrel of the syringe, and we found that a small amount of "Nonaq" stopcock grease applied to the rear portion of the plunger accomplished this. The standard 25 gage syringe needle (outside diameter 0.02 inches) was soldered with stainless steel solder into a 50 cm. length of 0.02 inch inside diameter stainless steel tubing which led directly to the reactor. This was immersed in a thermostat along with the reactor and the alkali control system.

To convert the volume flow rates of ester into mass flow rates, we used values of density as a function of temperature from the literature[5] which we checked in a few cases with agreement to 0.20%. The temperature used was the average temperature near the syringe. This was found to be constant to ±0.4°, which corresponds to ±0.1% in the rate constant.

With only one flow controlled by a constant pressure-choke system it was practicable to have significant flow resistance in the reactor outlet, which was not the case when two flows are controlled by this method.[4b] Consequently the diameter of the outlet tube could be reduced to 1.5 mm., and the resulting decrease in volume of this tube (its length was 12 cm.) brought the correction for reaction in the outlet[4b] to a satisfactorily low figure. Even so an error of 0.05 ml. in determining the volume of the outlet would lead to an error of 1% in the rate constant for methyl formate at 10°.

Reactor, alkali pipet, tubes connecting pipet to reactor, and the thin metal tube carrying ester to the reactor were immersed in a thermostat. With the rapid and exothermic reaction involved the reactor temperature was from 0.01 to 0.37° higher than the thermostat temperature. The reproducibility of the reactor temperature measurement by the thermistor was ±0.015°. The fundamental temperature standard was a set of National Bureau of Standards certified thermometers graduated to 0.02°.

Materials.—Methyl and ethyl formates were distilled at 35:1 reflux ratio through a glass helix packed column 1 meter long, a 0.3° cut being taken. Indices of refraction agreed closely with literature values. No measurable free acid could be detected. Saponification values were within 1% of theoretical for the methyl ester; for the ethyl ester they ran 2% low, a fact which we attribute to difficulty in getting the weighed sample of the not very soluble but highly volatile ester into solution. Ethanol (95%) was distilled through the same column at a 15:1 reflux ratio and was made up with boiled distilled water and concentrated carbonate-free aqueous sodium hydroxide to the desired composition. Batch 1 of this solution used in the measurements at 20 and 30°, contained 85.10% by weight ethanol and was 0.01203 *N* in sodium hydroxide; batch 2, used in the measurements at 10° contained 84.85% ethanol and was 0.01174 *N* in alkali.

The indicators, thymol blue for the reaction in water solution, brom thymol blue for the alcoholic solutions, were added to the standard hydrochloric acid used in the continuous titration.

Procedural Points.—In most of the experiments the input base concentration was 0.012 *M*, while input ester concentrations varied from 0.013 to 0.017 *M;* for methyl formate at 20° data were also obtained at input concentrations of approximately 0.06 *M*. At each temperature and for each ester the flow rate of alkali was varied in successive experiments over a range of from 20 to 30%, the ester flow remaining constant. This altered the per cent. excess of ester in the input by a factor of from 2 to 5. No consistent change in specific rate with changing flow rate was observed. Actual flow rates were in the range from 0.3 to 0.4 ml./sec. for ethyl formate and from 0.5 to 0.6 for methyl formate. The degree of conversion varied from 35–49% for ethyl formate at the lowest temperature; it was in the neighborhood of 90% for methyl formate both at 30° and 0.01 *M* and at 20° and 0.06 *M*.

Experiments were made in pairs for one of which the end-point color was definitely on the blue side, while the other was definitely on the yellow side. Analysis of the variance[6] for the individual sets of data showed that the error due to the alternating end-points is smaller than that due to flow rates or other experimental errors. The specific rates were calculated according to the method of Saldick,[4b] correcting for reaction in the outlet tube of the reactor. The probable errors were calculated using the average of the yellow and blue end-point values at a single flow rate as a single measurement. If they were treated as independent measurements, the per cent. probable errors would be smaller by from 0.2 to 0.5. At least 6 measurements were made for each ester at each temperature and at each concentration studied.

Error Considerations.—The most serious cause for worry in the use of the stirred flow reaction for the study of a reaction whose half time is as low as 0.6 sec. (in the case of methyl formate at 0.06 *M*) is the possibility that stirring is insufficient to maintain the necessary degree of homogeneity in the reactor. Our evidence for adequate stirring is: (1) freedom of the end-point color from temporal fluctuations; (2) independence of end-point setting from changes in stirring rate over a wide range; (3) agreement to 2.1% between the specific rates obtained at input concentrations of 0.06 and 0.01 *M* (in this connection the fact that alkaline ester hydrolysis is insensitive to changes in concentration of

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(4) (a) H. H. Young, Jr., and L. P. Hammett, THIS JOURNAL, **72**, 280 (1950); (b) J. Saldick and L. P. Hammett, *ibid.*, **72**, 283 (1950); (c) M. J. Rand and L. P. Hammett, *ibid.*, **72**, 287 (1950).

(5) "International Critical Tables," McGraw-Hill Book Co., New York, N. Y., Vol. III, p. 28.

(6) See W. J. Youden, "Statistical Methods for Chemists," John Wiley and Sons, Inc., New York, N. Y., 1951, Ch. 6, for the method used.

reactants or of reaction products is important). In our present techniques the limiting factor can be insufficient solubility rather than too great reactivity. This is shown by our inability to get satisfactory measurements on the hydrolysis of propyl formate in water, this ester being less reactive but also less soluble than methyl formate.

The reaction of the ester with hydroxyl ion is of course accompanied by a reaction with water. From the data of Poethke[7] it may however be estimated that the maximum rate of this reaction at 20° is 2.5×10^{-8} mole $l.^{-1}$ $sec.^{-1}$. This is only 0.1% of the slowest hydrolysis rate we have measured and may therefore be neglected.

Results

The specific rates and the enthalpies and entropies of activation calculated from them are tabulated in Tables I and II. As previously indicated the probable errors in the specific rates measure the scatter of some half dozen or more independent measurements; the probable errors in enthalpy and entropy are those which would result from these uncertainties in the specific rates. The Arrhenius equation applies to the temperature dependence of these specific rates to good precision; the maximum difference between observed specific rate and the value calculated from the enthalpies and entropies of activation reported is 0.52%.

Our specific rates for the hydrolysis of methyl formate in water are slightly higher than those of Leimu, *et al.*,[3] at the higher temperatures (by 4.5% at 25°) and slightly lower at the lower temperatures. As a result our value for the enthalpy of activation is 0.85 kcal. higher than theirs and our entropy of activation is 2.9 cal./deg. more positive. The agreement is in fact remarkably good for so fast a reaction.

Our value for the enthalpy of activation for the hydrolysis of ethyl formate in 85% ethanol differs by only 0.2 kcal. from the 14.1 kcal. found by Smith and Levenson[2] for the hydrolysis of ethyl acetate in the same medium. The 400-fold difference in rate of hydrolysis of the two esters is therefore essentially the result of the 11.3 cal./deg. greater entropy of activation of the formate reaction. We have recalculated the data of Fairclough and Hinshelwood[8] for the hydrolysis of methyl acetate in water and find that under these conditions the enthalpy of activation is 0.19 kcal. greater for formate than for acetate and that the entropy of activation is 11.7 cal./deg. greater for the formate.

TABLE I

ALKALINE HYDROLYSIS OF ETHYL FORMATE IN 85% ETHANOL

Av. temp., °C.	9.85	19.76	29.86
k ($l.$ $mole^{-1}$ $sec.^{-1}$)	0.835	1.98	4.53
% probable error	0.77	1.3	1.2

$\Delta H^{\ddagger} = 13.86 \pm 0.14$ kcal.; $\Delta S^{\ddagger} = -9.83 \pm 0.45$ cal./deg.; $k = 2.82$ $l.mole^{-1}$ $sec.^{-1}$, all at 25°.

TABLE II

ALKALINE HYDROLYSIS OF METHYL FORMATE IN WATER

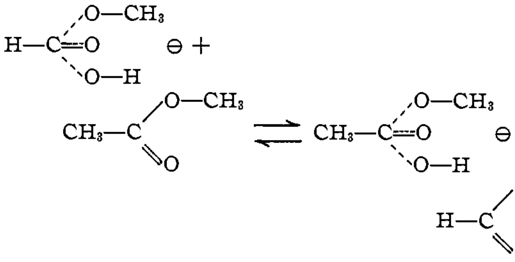
Av. temp., °C.	10.04	20.15	29.97	20.08
Approx. concn. (mole/l.)	0.01	0.01	0.01	0.06
k ($l.$ $mole^{-1}$ $sec.^{-1}$)	15.1	28.6	50.9	27.9
% probable error	0.58	0.24	2.5	1.6

$\Delta H^{\ddagger} = 9.81 \pm 0.22$ kcal.; $\Delta S^{\ddagger} = -18.41 \pm 0.77$ cal./deg.; $k = 38.4$ $l.mole^{-1}$ $sec.^{-1}$, all at 25°.

Discussion

In terms of a rather widely held point of view the difference between the alkaline hydrolysis rates of a formic ester and the corresponding acetic ester is a typical, perhaps the typical example of the electron repelling properties which methyl undoubtedly possesses compared with hydrogen. According to this view the increase in electron density on the COOR group produced by this electron shift repels the negatively charged hydroxyl ion and thus increases the energy of activation. Since the energies of activation of the two esters are in fact nearly identical, the large difference in rate deriving from a difference in entropy of activation, this point of view is at the least an oversimplification.

From our results in water solution we may write in cal. or cal./deg. for the process.

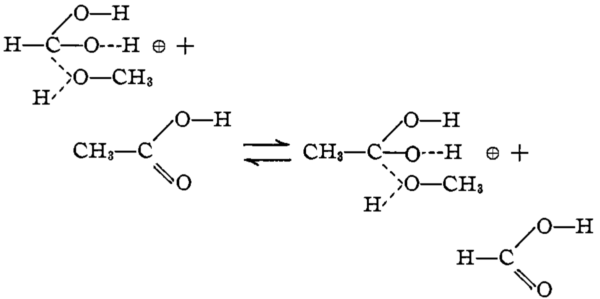


$\Delta F^{\ddagger} = 3440$; $\Delta S^{\ddagger} = -11.7$; $\Delta H^{\ddagger} = -190$

and for the corresponding process involving the ethyl esters in 85% ethanol

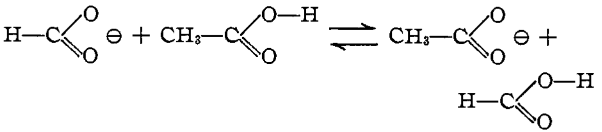
$\Delta F^{\ddagger} = 3560$; $\Delta S^{\ddagger} = -11.3$; $\Delta H^{\ddagger} = 200$

From the data of Smith[9] on the acid-catalyzed esterification of formic and acetic acids in methanol we may write for the process



$\Delta F^{\ddagger} = 1400$; $\Delta S^{\ddagger} = -5$; $\Delta H^{\ddagger} = 0$

Entropy effects of similarly large magnitude appear in both the acid- and the base-catalyzed transesterification of menthyl formate and acetate to the corresponding methyl esters.[10] Finally Harned's[11] precise measurements of the ionization constants of formic and acetic acids show that for the process



$\Delta F = 1370$; $\Delta S = -4.94$; $\Delta H = -100$

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[7] W. Poethke, *Ber.*, **68B**, 1031 (1935).

[8] R. A. Fairclough and C. N. Hinshelwood, *J. Chem. Soc.*, 538 (1937).

[9] H. A. Smith, THIS JOURNAL, **61**, 254 (1939).

[10] R. W. Taft, Jr., private communication.

[11] H. S. Harned and N. D. Embree, THIS JOURNAL, **56**, 1042 (1934); H. S. Harned and R. W. Ehlers, *ibid.*, **55**, 652 (1933).

There is therefore considerable basis for the generalization that the conversion of an electrically neutral acetic acid derivative to an electrically charged entity, no matter whether this be positively charged or negatively charged, involves a considerably greater decrease in entropy than does the analogous process in the case of a formic acid derivative, whereas the enthalpy effects are likely to be nearly the same for the acetic and formic acid derivatives.

In the gaseous state an acetic acid derivative probably possesses 10 or 11 cal./deg. more entropy than does the corresponding formic acid derivative. (The entropy of gaseous ethanol is 10.6 units greater than that of gaseous methanol, the corresponding difference for acetaldehyde and formaldehyde is 11.2.[12]) This difference can reasonably be assigned to the entropy associated with the internal motions of the methyl group, the mass effect being trivial. A considerable portion of the difference probably disappears on solution in water since the entropy of solution of gaseous ethanol is 6 units more negative than that of gaseous methanol.[13] Some of this effect is no doubt associated with a partial freezing out in solution of the internal motions of the methyl group, but there are indications[13,14] that interference with the freedom of motion of the solvent molecules is also involved. It seems likely therefore that the entropy of a dissolved acetic acid derivative is greater than that of the corresponding formic acid derivative by some 4 to 6 cal./deg. On this basis we may expect that the entropy of acetate ion is approximately the same as that of formate ion, but that the transition state for the alkaline hydrolysis of methyl acetate is several cal./deg. more negative than that for methyl formate.

One could account for equality of the entropies of the charged molecular species derived from acetic acid and from formic acid by supposing merely that all of the internal motion of the methyl group is frozen out by the charge through an effect of the sort suggested by Kirkwood[15] but in order to account for a more negative entropy in the acetic acid derivative, one must certainly assume that some hampering of the motions of the solvent molecules must result from the substitution of methyl for hydrogen in the solute ion. This is of course exactly opposite to what would be expected from the picture in which the two ions are thought of as conducting spheres immersed in a dielectric medium with the acetate having the greater volume, a picture which accounts satisfactorily[16] for the fact that the conversion of methylamine to methylammonium ion involves an entropy change 4.2 cal./deg. more *positive* than does the conversion of ammonia to ammonium ion. It is not however inconsistent with a picture in which the charge is confined to a limited portion of the charged molecule, the COOR structure, but is displaceable within that region. The substitution of a methyl group for hydrogen would then represent the attachment to this conducting region of a region of low dielectric constant. The charge would under such conditions tend to move away from the methyl group and toward the region of high dielectric constant represented by the solvent. This would increase the binding to the ion of the solvent molecules and would thereby decrease the entropy associated with the motions of the solvent molecules.[17]

[12] "Selected Values of Chemical Thermodynamic Properties," Circular of the National Bureau of Standards 500.

[13] R. E. Powell and W. M. Latimer, *J. Chem. Phys.*, **19**, 1139 (1951).

[14] H. S. Frank and M. W. Evans, *ibid.*, **13**, 507 (1945).

[15] See F. P. Price and L. P. Hammett, THIS JOURNAL, **63**, 2392 (1941).

[16] A. G. Evans and S. D. Hamann, *Trans. Faraday Soc.*, **47**, 34 (1951).

[17] This point of view developed as a result of discussion of this problem with Dr. L. Onsager.

NEW YORK, N. Y.

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY, COLUMBIA UNIVERSITY]

Structural Studies on Ester Hydrolysis by Strong Base Ion Exchangers[1]

BY HAROLD SAMELSON AND LOUIS P. HAMMETT

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The rates of the alkaline hydrolysis in 85% ethanol of three typical esters, ethyl acetate, ethyl *n*-hexanoate and ethyl phenylacetate by seven ion exchangers of the quaternary ammonium hydroxide type which differed in the alkyl groups attached to nitrogen follow the second-order rate law to considerable precision. The hydrolysis rate of hexanoate is materially lower on all the resins studied than it is in homogeneous strong base solution. This effect is presumably to be attributed to the greater loss of entropy which this entropy-rich molecule undergoes in the formation of the transition state of the reaction. Rates of reaction of all esters with the resin in which the functional group is $\mathrm{RN(CH_3)_3^+}$ are considerably lower than they are with resins whose functional groups contain ethyl, *n*-hexyl and benzyl groups; the $\mathrm{RN(CH_3)_3^+}$ resin is much less swollen by water or aqueous ethanol than are the other resins. In other respects changes in the structure of the functional group have only small effects on the hydrolysis rates.

In an extension of our studies[2] on the catalytic properties of ion exchange resins we have investigated some of the structural factors which affect the rate of hydrolysis of an ester by a resin of the quaternary ammonium hydroxide type. The reaction is

$$\mathrm{-PCH_2NR_1R_2R_3^+ + OH^- + R_4COOC_2H_5 \longrightarrow -PCH_2NR_1R_2R_3^+ + R_4COO^- + C_2H_5OH}$$

where $\mathrm{-PCH_2NR_1R_2R_3^+}$ is a unit of the structure of a bead-form cross-linked polymer obtained by (1) suspension polymerization of styrene together with

[1] The work reported herewith was carried out as project NR 056-062 under contract N6onr-271 between the Office of Naval Research and Columbia University. Reproduction in whole or in part permitted for any purpose of the United States Government.

[2] (a) V. C. Haskell and L. P. Hammett, THIS JOURNAL, **71**, 1284 (1949); (b) S. A. Bernhard and L. P. Hammett, *ibid.*, **75**, 1798 (1953); (c) **75**, 5834 (1953); (d) S. A. Bernhard, E. Garfield and L. P. Hammett, *ibid.*, **76**, 991 (1954); (e) P. Riesz and L. P. Hammett, *ibid.*, **76**, 992 (1954).