966. *Energy–Entropy Relations in Acylations. Part II.* Medium Effects.*

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Arrhenius parameters have been determined at various temperatures for the acylation of substituted anilines with two different acid chlorides in benzene–nitrobenzene and benzene–phenyl cyanide mixtures. In each example the rate constant increases steadily with the proportion of the polar constituent. Up to about 25% by weight, E and log A fall steeply and over this range are approximately linearly related. E is insensitive to further increase of the polar component but log A rises almost linearly.

The correlation between E and log A is discussed in terms of changing solvation of the polar transition complex. A rough model derived on this basis shows qualitative agreement with the observed variations in the Arrhenius parameters.

IN the acylation of substituted anilines the parameters of the Arrhenius equation, $k = Ae^{-E/RT}$, respond in various ways to changes in the reacting molecules: A, the entropy term, may remain constant, while changes in k are determined solely by the activation energy. This is true especially for ring substitution and perhaps for purely polar effects in general. Structural modification in the acylating agent is associated with a rather complex correlation between the variations in the energy and entropy factors. The interplay of these effects is discussed in the preceding paper.

The rate of acylation is markedly dependent on the solvent, which can exert polar effects and can also modify profoundly the structure of the transition complex. The way in which solvents alter the values of A and E is thus closely relevant to the general understanding of energy–entropy correlations. Since the change from one solvent to another is difficult to represent by any quantitative parameter, the present study has been made with binary mixtures of polar and non-polar solvents in which the composition could be continuously varied.

Previous studies have included the benzoylation of *m*-nitroaniline in several solvents,[1] and the benzoylation of substituted anilines in benzene [2,3] and nitrobenzene.[4] The reactions are bimolecular and of the second order in all the solvents studied and go to completion. With monosubstituted anilines the solvent does not alter the transmission of electronic substituent effects to the nitrogen atom. The reactions are generally faster in polar solvents, but no unambiguous correlation has been found between the rate and the dielectric constant. The energy and frequency factors are both lower in the more polar solvents.

In the present work various combinations of four *meta*-substituted anilines with two different acid chlorides were studied in a series of mixed solvents consisting of benzene and either nitrobenzene or phenyl cyanide.

EXPERIMENTAL

Materials.—The substituted anilines and acid chlorides were purified by crystallization or distillation: " AnalaR " benzene and " AnalaR " nitrobenzene were used throughout. Commercial phenyl cyanide was dried for several days and then fractionated (b. p. 188—190°).

Solvent mixtures were made up in bulk from the required volumes of the two components.

Technique.—The reaction between the amines and the acid chlorides used is simple and complete. The method described by Bose and Hinshelwood [4] was adopted for following the

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* Part I, preceding paper.

[1] Pickles and Hinshelwood, *J.*, 1936, 1353.

[2] Williams and Hinshelwood, *J.*, 1934, 1079.

[3] Stubbs and Hinshelwood, *J.*, 1949, 71.

[4] Bose and Hinshelwood, *J.*, 1958, 4085.

reactions in polar solvents and solvent mixtures, the other details being as described in earlier papers.[2,5]

The concentrations of the reagents used in all experiments were: acid chloride, M/200; aniline, M/100. Between 13° and 60° electrically regulated thermostats (±0·05°) were used. Experiments at 0° or 5·1° were carried out in a Dewar vessel containing crushed ice or melting benzene (±0·2°). The rate constants are uncorrected for the thermal expansion of solvents. Satisfactory second-order constants were given in all cases, and are recorded in l. mole^{-1} sec.$^{-1}$ (see preceding paper).

The activation energies were calculated by the method of least squares from the rate constants at 4 or 5 different temperatures, the plots of log k against $1/T$ being linear in all cases.

RESULTS

The results of the kinetic measurements, rate constants at 25° together with E and log A, are summarised in Tables 1—4. The composition of the medium is given as weight per cent of the polar solvent.

TABLE 1. *Isobutyrylation of* m-*nitroaniline in benzene–nitrobenzene.*

Nitrobenzene (%)	10^2k_{25}	E	log A	Nitrobenzene (%)	10^2k_{25}	E	log A
0	0·513	10,400	5·31	82·0	11·68	6600	3·99
13·2	1·514	7700	3·82$_5$	92·5	16·57	6700	4·22
31·2$_5$	2·910	6800	3·45	96·3	21·84	6650	4·26
57·7$_5$	7·325	6700	3·86$_5$	100	27·92	6600	4·28$_5$

TABLE 2. *Benzoylation of* m-*nitroaniline in benzene–phenyl cyanide.*

Phenyl cyanide (%)	10^2k_{25}	E	log A	Phenyl cyanide (%)	10^2k_{25}	E	log A
0	0·0438	10,550	4·38	43·2	1·836	6650	3·14
5·7	0·162	7550	2·76	53·3	3·036	6700	3·39
11·3	0·338	7200	2·81	77·4	5·426	6700	3·65
22·2	0·651	6700	2·72$_5$	91·1	6·346	6750	3·80
32·9	1·140	6650	2·93	100	6·943	6800	3·83

TABLE 3. *Benzoylation of* m-*nitroaniline in benzene–nitrobenzene.*

Nitrobenzene (%)	10^2k_{25}	E	log A	Nitrobenzene (%)	10^2k_{25}	E	log A
0	0·0438	10,550	4·38	82·0	0·924	7250	3·28
10	0·137	8250	3·19	88·5$_5$	1·076	—	—
13·2	0·156	7700	2·84	92·5	1·21	7250	3·40
31·2$_5$	0·342	7450	3·00	96·3	1·418	—	—
57·7$_5$	0·640	7350	3·16	100	1·71	7200	3·51

TABLE 4. *Benzoylation of* m-*chloroaniline in benzene–nitrobenzene.*

Nitrobenzene (%)	10^2k_{25}	E	log A	Nitrobenzene (%)	10^2k_{25}	E	log A
0	0·569	9350	4·61	57·7$_5$	4·813	5850	2·97
2·7	0·676	8500	4·06	82·0	7·741	5900	3·29
6·4	0·845	8250	3·98	92·5	9·093	5950	3·32
13·2	1·392	7000	3·27	96·3	10·29	5900	3·34
19·4	1·732	6250	2·82	100	12·54	5800	3·35
36·9	2·822	5850	2·85				

Comparison of the results for the four sets of reactions in mixed solvents reveals the following general picture irrespective of the acid chloride or the amine used.

(1) The addition of either nitrobenzene or phenyl cyanide to benzene results in a steady increase in rate (Fig. 1).

(2) The curves giving the Arrhenius parameters, E and log A, as a function of the composition of the solvent have the same form for all the 4 reactions, as shown in Figs. 3 and 4.

(3) In all the examples E decreases as an approximately linear function of log A up to about

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[5] Grant and Hinshelwood, *J.*, 1933, 1351.

25% of the polar solvent (Fig. 5). Beyond this composition there is a rapid change, E remaining almost constant while log A increases and the rate rises.

Somewhat similar behaviour has been found for quaternary ammonium salt formation in the same solvent mixtures [6] and for the hydrolysis of esters in different mixed solvents.[7]

DISCUSSION

Correlation of Rate with the Composition of the Solvent.—The variation of rate with composition is generally similar for different systems as is illustrated in Fig. 1 where log k_{25} is plotted against the percentage of the polar component. It is perhaps surprising that the rates are higher in phenyl cyanide mixtures despite the lower dielectric constant. Similar results have been found in solvolysis,[8] hydrolysis,[9] and quaternary ammonium salt formation.[10] They indicate that dielectric effects are not of prime importance [11] and may indeed play a minor rôle.[12] Leffler [13] has pointed out that even the ionizing power of a solvent is not due to its dielectric constant alone but also to a specific chemical effect.

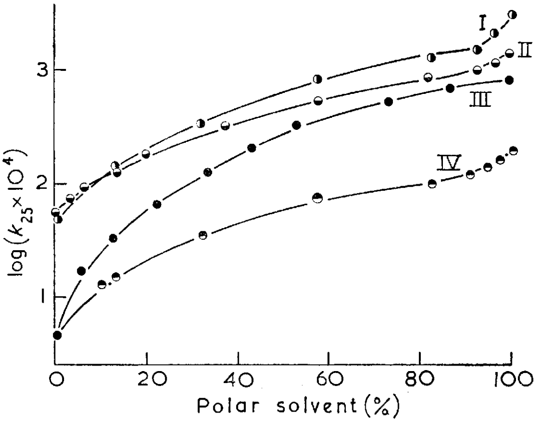

FIG. 1.
I, m-*Nitroaniline, isobutyryl chloride; benzene–nitrobenzene.*
II, m-*Chloroaniline, benzoyl chloride; benzene–nitrobenzene.*
III, m-*Nitroaniline, benzoyl chloride; benzene–phenyl cyanide.*
IV, m-*Nitroaniline, benzoyl chloride; benzene–nitrobenzene.*

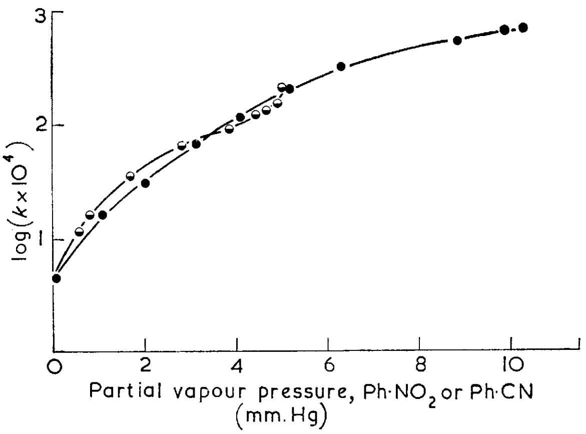

FIG. 2. m-*Nitroaniline, benzoyl chloride:* ● *benzene–phenyl cyanide;* ◒ *benzene–nitrobenzene.*

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[6] Raine and Hinshelwood, *J.*, 1939, 1378.

[7] Tommila, Koivisto, Lyyra, Antell, and Heimo, *Ann. Acad. Sci. Fenn.*, 1952, *A*, II, **47**; Tommila, *Suomen Chem.*, 1952, **25**, B, 37; Tommila and Hietala, *Acta Chem. Scand.*, 1954, **8**, 257.

[8] Hammond, Reeder, Fang, and Kochi, *J. Amer. Chem. Soc.*, 1958, **80**, 569.

[9] Hudson and Brown, *J.*, 1953, 883.

[10] Watanabe and Fuoss, *J. Amer. Chem. Soc.*, 1956, **78**, 527.

[11] Pearson, *J. Chem. Phys.*, 1952, **20**, 1479.

[12] Streitwieser, *Chem. Rev.*, 1956, **56**, 603.

[13] Leffler, *J. Org. Chem.*, 1955, **20**, 1202.

An interesting empirical relation is shown in Fig. 2 where log k_{25} is plotted against the partial vapour pressure [14,15] of the polar component, and the curves for the two systems are seen to be almost coincident. This partial vapour pressure is a measure of the tendency for the polar component to escape from the bulk medium and may be related to the ease with which its molecules can be disengaged to arrange themselves around the activated complex as a solvation sphere, and so, in turn affect the reaction rate. This problem, however, is not a very tractable one.

Activation Energies and Frequency Factors.—The dependence of E and log A on the composition of the solvent is represented in Figs. 3 and 4. As the proportion of the polar component is increased, E decreases rapidly, reaching a minimum value at about 25%, beyond which it is insensitive to further addition of nitrobenzene or phenyl cyanide. Log A also falls rapidly to a minimum at about the same point as E but beyond this it increases more or less linearly. (A similar change is observed in the racemization of a biphenyl derivative in dioxan–water and dimethyl sulphoxide–water mixtures.[16]) The plots of log A against E for all the reactions show an approximately linear relation as far

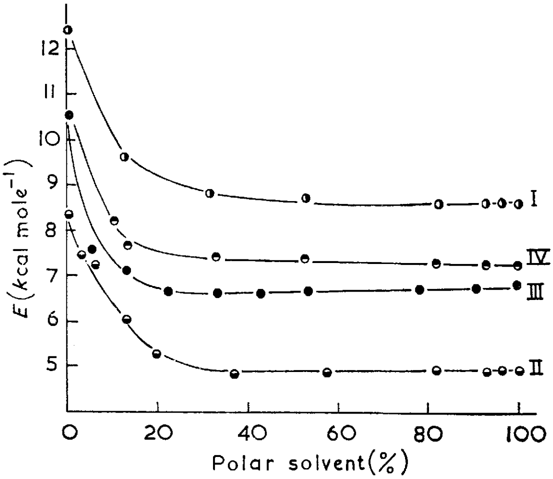

FIG. 3. *The numbering of curves and the marking of points correspond to Fig. 1. (The ordinates of curve I are displaced upwards by 2 kcal., and those of curve II downwards by 1 kcal.)*

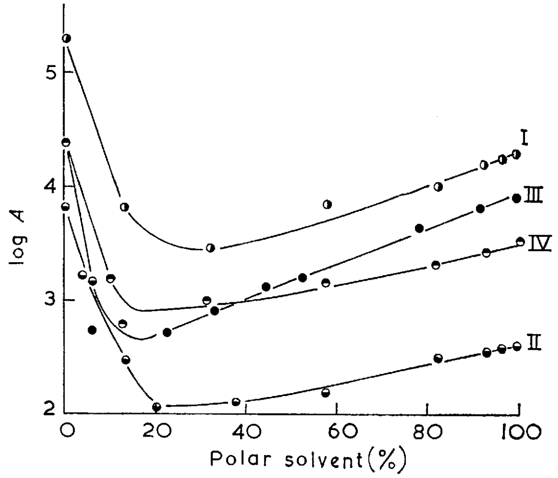

FIG. 4. *Numbering and marking are as in Figs. 1 and 3. Ordinates of curve II are displaced downwards by 0·8 unit.*

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[14] Martin and Collie, *J.*, 1932, 2658.
[15] Martin and George, *J.*, 1933, 1413.
[16] Leffler and Graham, *J. Phys. Chem.*, 1959, **63**, 687.

as the composition at which the constant value of E is attained, the linear portions having a slope of about 0·6.

The reaction between the amine and the acid chloride involves an increase in polarity in the activated complex, and a zwitterion structure has been suggested for the transition state.[17] There will be a considerably greater charge separation than in the reactant molecules separately and the complex will orient more solvent molecules.[18] The influence of the medium on such reactions is well known.[19-21] The rates are higher in nitrobenzene and phenyl cyanide, and as expected [18] E and log A are both low. Up to 25% of the polar solvent, the fall in E may well be due to the gradual increase in the solvation of the transition complex as more molecules of the polar solvent become available. When there are enough to surround this complex, E will not suffer much further change. The low value of the entropy in this region is explicable by that lost when solvent molecules become frozen in the polar transition state.[22] The fact that further addition of them actually increases the entropy term seems at first sight surprising. To explain such effects various factors have been invoked, such as the internal structure of the solvent and the ability of the solvent to prevent relapse of the transition complex into its original state.[22]

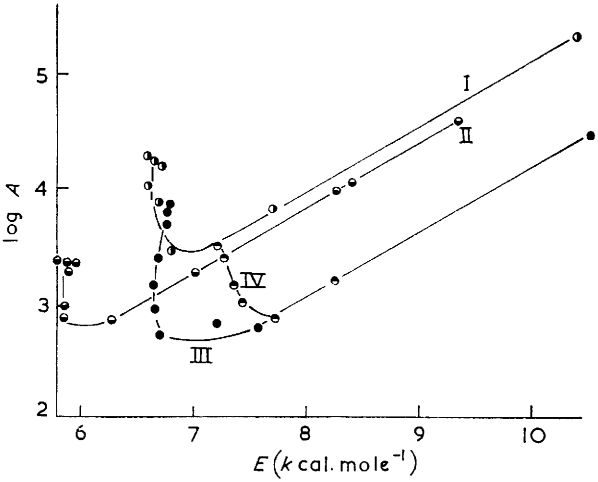

FIG. 5. *Numbering and marking as in Figs. 1, 3, and 4.*

For variations in the rates of reactions under high pressures [23] linear relations have been found between the energy and entropy factors. Laidler [24] has suggested that factors tending to increase the binding of the solvent molecules during the activation process bring about a decrease in both energy and entropy and there appears to be almost exact compensation between these two factors. This is paralleled by the present results in solvent mixtures containing up to 25% of the polar solvent.

Rough Model of E–*log* A *Relations*.—The following model is intended to illustrate qualitatively the variation of log A with composition, and in particular to reproduce the continued rise in the entropy factor after E has fallen to its steady minimum value.

Let the transition complex be capable of accommodating m molecules of the polar solvent around it, m being comparable with a co-ordination number. Let the mole fraction of the polar solvent be α. Then, if pure chance determined whether a given one of the

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[17] Elliot and Mason, *Chem. and Ind.*, 1959, 488.
[18] Bell, *J.*, 1943, 629.
[19] Wynne-Jones and Eyring, *J. Chem. Phys.*, 1935, **3**, 492.
[20] Ogg and Rice, *J. Chem. Phys.*, 1937, **5**, 140.
[21] Laidler and Eyring, *Ann. New York Acad. Sci.*, 1940, **40**, 303.
[22] Tommila and Kauranen, *Acta Chem. Scand.*, 1954, **8**, 1152.
[23] Burris and Laidler, *Trans. Faraday Soc.*, 1955, **51**, 497.
[24] Laidler, *Discuss. Faraday Soc.*, 1956, **22**, 88.

m places were occupied by the polar or the non-polar component, the probability that n out of the m places is occupied by polar molecules would be

$$^{m}C_{n}\,.\,\alpha^{n}(1-\alpha)^{m-n}$$

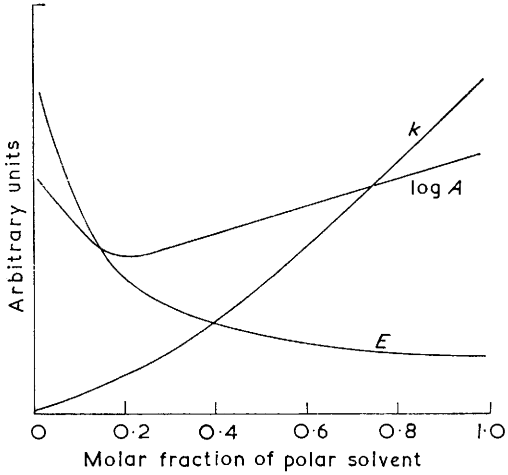
If, as is almost certain, there is an energy difference for solvation by polar and non-polar molecules, the chance for each of the former is increased by a Boltzmann factor which virtually increases α to a higher value which can be designated α'.

Let E_0 be the activation energy in the non-polar solvent. This will be lowered by each polar molecule entering the solvation zone of the complex, being $E_0 - \mathrm{f}(n)$ when the number is n. Each of the solvated complexes will make its own contribution to the total reaction rate, the total being proportional to

$$\sum_{0}^{m} {}^{m}C_{n}\,.\,\alpha'^{n}(1-\alpha')^{m-n}\,.\,\exp\left[-\{E_0-\mathrm{f}(n)\}/\boldsymbol{R}\mathrm{T}\right]$$

Calculations can now be made as follows. A value of 6 may be assumed for m and will not be far from correct. A trial assumption is next made about $\mathrm{f}(n)$, for example, that each polar molecule in the solvation zone reduces E_0 by an equal amount, until the value for $n = 6$ becomes equal to that for the polar medium itself. The relative contribution to the rate from each of the complexes corresponding to values from $n = 0$ to $n = 6$ is then calculated, and the total found. This process is repeated for a series of values of α'. As α' increases from 0 to 1, the value of n from which the maximum contribution comes shifts from 0 to 6. The whole calculation is then made again for a higher temperature, the contribution due to each value of n having of course a different temperature coefficient. From the total calculated rates at the two temperatures the observable value of the activation energy for each value of α' can be found by application of the Arrhenius equation. We now have the total calculated rates, each one relative to that in the non-polar medium, and the observable activation energies for each composition, and from these the corresponding values of log A can be calculated.

FIG. 6.



To illustrate the general form of the results we may take α' to be given nearly enough by α, the actual mole fraction. With the above assumption about $\mathrm{f}(n)$ the activation energy falls too gradually: nevertheless, the most important characteristic of A is predicted, namely, the fall to a minimum and the subsequent rise.

A different assumption can now be made, namely, that successive polar molecules in the complex produce decreasing effects, and it is not difficult more or less to reproduce the actual course of the E–composition curve. When this is done the variation of A with composition is as shown in Fig. 6, which shows the essential qualitative characteristics of the experimental curves. In view, however, of the assumed identity of α and α' the concentration scale must really be distorted.

The model does at least give a possible rational explanation for the minimum in the value of A. It may be useful to re-state this in simple terms. At certain intermediate compositions it is far from probable that the complexes with high values of n, the number of polar solvating molecules, will predominate in the solution. Nevertheless, in virtue of their low activation energies and in spite of their relative scarcity, these particular complexes make the largest contribution to the rate. Because they do this, the observable rate is associated with a measurable activation energy corresponding to the high solvation number. When now the mole fraction of polar solvent is increased, the probability of high polar solvation numbers increases, and the rate rises. But since the major contribution to the rate was already coming from the highly solvated complexes, the activation energy cannot be much more reduced. Hence the tendency for E to reach a limit while A steadily rises.

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