135. *The Kinetics of the Esterification of Normal Fatty Acids.*

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Rates of esterification of series of normal fatty acids have been determined for the following systems :

(*a*) In ethyl alcohol : catalysed by undissociated acid—reaction bimolecular with respect to acid.

(*b*) In methyl alcohol : catalysed by hydrion from hydrogen chloride.

(*c*) In *cyclo*hexanol : catalysed by undissociated acid.

From butyric acid upwards the activation energies show only small random variations. The true values are therefore assumed constant, and k is taken as a measure of PZ in the Arrhenius equation.

It is found to rise with the higher acids appreciably in (*a*), to a smaller extent in (*c*), and little, if at all, in (*b*).

The rise is ascribed to an increase in collision number with increasing incompatibility of packing of solvent and solute molecules.

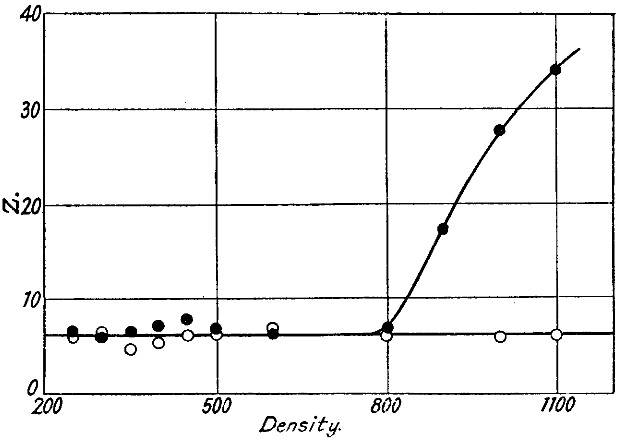
The possibility of this effect had been suggested by experiments with a mechanical model in which steel balls represented molecules.

The velocity constant of a reaction in solution being expressed by the modified Arrhenius equation $k = PZe^{-E/RT}$, the problem arises of factorising the temperature-independent term PZ, in which Z is the actual rate of collision between reactant molecules, and P the probability that collisions between activated molecules lead to reaction. General theoretical arguments, and direct experimental comparisons between reactions in the gas phase and in solution indicate that the order of magnitude of Z in solution is that calculated from the kinetic theory of gases. Although the solvent impedes diffusion and thus cuts down the number of collisions which a given solute molecule makes in unit time with fresh ones, it hems in pairs of solute molecules and causes them to make repeated collisions with each other, the two effects probably nearly balancing. For the purposes of ordinary slow chemical reactions the repeated collisions are as effective as fresh encounters. Thus the large changes in PZ over several powers of ten which are often found must be due to changes in P. But the smaller changes within a given order of magnitude may, as far as theory can decide, equally well be due to changes in Z.

In some recent experiments (J., 1937, 1573; 1938, 236) we have found changes in PZ with variation of solvent, which, with reasonable certainty, can be analysed into two parts, one showing the well-known correlation with E, the other being apparently a function of the polarity of the solvent and reactants. It was suggested that part of this variation might be due to changes in Z. The most probable source of these variations appears to be the " squeezing out " to a greater or smaller extent of solute molecules from the structure of the solvent, whereby the concentration virtually changes, with consequent increase of Z.

The increase in the number of repeated collisions as the number of fresh encounters diminishes with increasing density of solvent has been illustrated on a mechanical model by Rabinovitch and Wood (*Trans. Faraday Soc.*, **1937**, **33**, 1225). In order to investigate further some purely geometrical aspects of the collision problem, we had a somewhat similar model constructed, in which the behaviour of small steel balls could be observed under various conditions. In this model steel balls of given sizes represented solvent and solute molecules. In the first series of experiments solute and solvent balls were of the same size, and the effect of increasing solvent density was investigated. The above-mentioned result was confirmed : as the solvent became more dense the number of sets of repeated collisions grew at the expense of the number of fresh encounters. The total collision rate, however, remained constant. The same result was found when the solute balls were smaller than those representing the solvent. When, however, they were larger, a new phenomenon appeared. At moderate densities of solvent, the total collision rate was constant as before, but at very high densities the total solute collision rate rose sharply. The solvent exhibited a very strong tendency to assume a close packed " structure," into which the larger solute balls did not fit, but tended to become " squeezed out " in groups, with a consequent large increase in the number of repeated collisions (see Fig. 1).

FIG. 1.



Open circles = *solute diameter* 3/16 *inch.*
Black ,, = ,, 5/16 ,,

The question arises as to how far results for the behaviour of steel balls can be applied to real molecular systems. Although the latter would certainly show greater complexity of behaviour, leading perhaps to new sources of variation in Z, an effect depending upon purely geometrical factors such as the packing of the solvent and solute might well appear. To investigate this matter we need a bimolecular reaction involving two like molecules the size of which can be varied by passage up a homologous series, the solvent being kept constant. The analogy would naturally be far from exact, since molecules are not steel balls, nor would the higher members of the series be even approximately spherical. Yet if we reflect that the increased collision rate observed with the larger balls in the model is less due to their size than to the increasing incompatibility of the packing systems preferred by solute and solvent as the solute becomes larger, we see that the analogy in this particular respect would be quite close. Systems composed of higher members of a homologous series dissolved in a solvent of low molecular weight may, from this point of view, prove to be analogous to those containing the large and small steel balls.

One of the few chemical reactions fulfilling the condition required is the esterification of a long-chain fatty acid in one of the lower alcohols. The undissociated molecule of the acid acts as its own catalyst, and the reaction is kinetically bimolecular with respect to the acid, there being excess of alcohol all the time. By using a number of normal fatty acids of different molecular weight, the size of the solute molecules can be varied without appreciable change in their chemical properties.

The kinetics of this reaction were therefore investigated for a series of fatty acids having straight chains of from 2 to 22 carbon atoms, the solvent being ethyl alcohol. The results are more fully discussed in a later section, and it is sufficient to state here that a steady increase in PZ of the kind discussed was observed from butyric to behenic acid.

Certain results of Goldschmidt (*Z. physikal. Chem.*, 1920, **94**, 233) suggested that with one only of the two reactants increasing in size along the series, the incompatibility of packing might not lead to any change in the collision rate, there being no extrusion of the smaller reactant from the solvent structure. The kinetics of esterification of the same series of acids was therefore investigated in methyl alcohol, with hydrion as catalyst. Here no significant increase in PZ with the size of fatty acid was observed.

To complete the series of experiments, the esterification of the same acids should be investigated in a long-chain alcohol as solvent, where no increase in Z due to incompatibility of packing would be expected. Unfortunately, such alcohols do not dissolve water, and therefore as soon as any reaction occurred the system would become heterogeneous, so that accurate measurements could not be made. For this reason *cyclo*hexanol was chosen as the largest solvent molecule which it was practicable to use.

EXPERIMENTAL.

Experiments with the Mechanical Model.—The model consisted of a polished wooden annular tray, the radial cross-section of the annular space being approximately 6″. The tray was mounted eccentrically on a geared axle driven by an electric motor, and could be rotated at constant speeds varying from approximately 50 to 150 r.p.m. A periscopic arrangement enabled a circular area of diameter $2\frac{1}{4}''$ to be observed. Steel balls were used to represent molecules, those representing the solute being painted. The total number of collisions between solute balls in a given time in the area of observation was counted. Each time of counting was long enough to ensure only small fluctuations from the mean. Since the balls received their kinetic energy from collisions with the walls of the tray and with each other, " thermal equilibrium " was only set up inside the system when the number of collisions between balls was greater than that of balls with the walls of the tray. In other words, the experiments were only significant when the mean free path of the balls was considerably less than the dimensions of the tray. The steel balls varied in size from $\frac{1}{8}''$ to $\frac{5}{16}''$.

Fig. 1 shows the results of a set of experiments in which the solvent balls (diameter $\frac{3}{16}''$) were varied in density from the minimum compatible with " thermal equilibrium " almost to the maximum which the tray would accommodate. When the solute and solvent balls were the same size the total collision rate was independent of the solvent density. When the diameter of solute was raised to $\frac{5}{16}''$ the collision rate exhibited a sharp rise when the solvent density was increased beyond a certain point. Z, the collision rate between solute balls, is expressed in min.$^{-1}$ and the density of the solvent is represented by the total number of balls on the tray; in each case there were ten solute balls present. Inspection of the model in operation showed that in the flat part of the curve the ratio of repeated collisions to fresh encounters was increasing with solvent density. In the region of abnormal increase in collision rate the solvent was so tightly packed that it took on a definite structure, from which the solute balls tended to be excluded.

On repetition of the experiments with solute and solvent balls both of the larger size, or with solute smaller than solvent, the rate was again found to be independent of density at all densities, since under these conditions the solute balls were easily included in the packing arrangement.

Purification of Materials.—Ethyl alcohol was refluxed for several hours over freshly burnt quicklime, and distilled. Methyl alcohol was fractionated, before and after treatment, with aluminium amalgam. *cyclo*Hexanol was dried over anhydrous sodium sulphate and fractionated (b. p. 159°; m. p. 23°). Acetic acid was frozen out six times, propionic, butyric, hexoic, and pelargonic acids were fractionated, and myristic and palmitic acids were recrystallised several times from alcohol. The last two were free from unsaturated acids and had setting points of 52° and 62° respectively.

Behenic acid was prepared from " Kahlbaum " erucic acid by hydrogenation of the ester in presence of colloidal palladium, the details of technique being kindly supplied by Professor F. Francis. The product was free from erucic acid, and after two recrystallisations from glacial acetic acid had a setting point of 79°. All fractionations and distillations were carried out in all-glass apparatus. Thermometers were checked against N.P.L. standards. All constants are corrected for solvent expansion and expressed in l. (g.-mol.)$^{-1}$ sec.$^{-1}$.

Catalysis by Undissociated Acid in Ethyl Alcohol.—For the main series of experiments, N-solutions were used, and measurements were made at four temperatures, 75·5°, 100°, 120° and 154°, vapour-bath thermostats being employed.

Solutions were sealed up in 5-c.c. glass bulbs, preheated, and, at suitable intervals, titrated with alkali, cresol-red being the indicator.

Solutions of acids up to pelargonic were buffered with N/100-sodium or ammonium salt. With butyric acid, the effect of varying the concentration of acid and of buffer was investigated at 100°. Change of acid concentration to 0·25N produced a very slight increase in rate, as did absence of buffer, or increase of buffer concentration to 0·025N (see Table I).

TABLE I.

Influence of Concentration and Buffer on Butyric Acid at 100°.

Acid	1·0N	1·0N *	1·0N	0·25N
Buffer	0·0	0·01N	0·025N	0·01N
$k \times 10^{-6}$	336·1	309·8	317·0	336·1

* Main series of experiments.

Myristic acid was investigated in buffered and in unbuffered solution, no difference in rate being observed, indicating that its dissociation in ethyl alcohol is negligible; the same was taken to apply to palmitic and behenic acids. With myristic acid, k was constant between 1·0N and 0·25N, and for the main series of experiments with this acid N/2-solutions were used. Owing to the low solubility of palmitic acid, N/2-solutions were used, and even at this concentration solutions had to be made up at 30°, which was allowed for in correcting for solvent expansion.

The extremely low solubility of behenic acid made experiments with it rather difficult, and probably less accurate. The maximum concentration which could be obtained was 0·0785N, solutions being made up and pipetted out at 50°. In each case bimolecular constants were found, and the Arrhenius equation was obeyed (Tables II and III).

TABLE II.

Acetic acid, 1·0N.
Temp. 151·0°.

Mins.	%.	$k \times 10^6$.
23	6·78	52·9
29	8·76	55·2
53	14·17	52·2
77	20·55	55·8
228	42·72	54·5
254	45·91	55·7
265	47·21	56·0
284	49·00	56·4
349	53·18	54·3

Mean k (uncorr.) = $54·9 \times 10^{-6}$.

TABLE III.

Palmitic acid, 0·5N.
Temp. 153·9°.

Mins.	%.	$k \times 10^6$.
58	8·46	53·5
85	10·94	48·2
108	14·43	52·2
140	17·16	49·3
170	19·65	47·8
199	23·14	50·4
287	29·10	47·6
350	34·82	50·8
408	36·32	46·6

Mean k (uncorr.) = $49·6 \times 10^{-6}$.

In Table IV are given the values of k at 100·0° calculated from those of E and log PZ determined by the method of least squares.

Catalysis by Hydrion in Methyl Alcohol.—Hydrogen chloride was used as catalyst. Dry filtered gas, tested for freedom from sulphuric acid spray, was passed into methyl alcohol, and the concentrated solution was diluted to the required strength. The reaction was studied at 0°, 20°, 30°, and 45°, at least two runs at each temperature being made. Owing to the low

TABLE IV.

Summary of Results for Esterification in Ethyl Alcohol.

Acid.	E.	log PZ.	$k_{100} \times 10^6$.
Acetic	15,200	3·69	5·70
Propionic	15,000	3·47	4·47
Butyric	14,800	3·18	3·00
Hexoic	15,100	3·41	3·39
Pelargonic	14,900	3·34	3·80
Myristic	14,800	3·30	3·99
Palmitic	15,100	3·51	4·27
Behenic	15,100	3·51	4·27

solubility of palmitic acid, it was only possible to work at the three higher temperatures. Behenic acid is too insoluble in methyl alcohol to be used at all. Solutions of the fatty acid and catalyst were preheated before mixing, and 5 c.c. portions were withdrawn from time to time and titrated against N/50-sodium hydroxide. (Methyl alcohol was used instead of ethyl alcohol for these experiments with hydrion, which are much more sensitive to minute traces of impurity, presumably basic, in the alcohol. It was found easier to obtain methyl alcohol of the standard of purity required for the most accurate work.)

In the main series of experiments N/10-fatty acid, and hydrogen chloride from 0·002N to 0·006N were used. The course of the reaction was found to be exactly described by Goldschmidt's equation (*Z. physikal. Chem.*, 1907, **60**, 728), *viz.*,

$$kct = (1 + a/r)\ln a/(a - x) - x/r$$

where c is the concentration of hydrion, a the initial concentration of fatty acid, and x the amount esterified in time t. The empirical constant r has a single value for a given acid at a given temperature, but varies with temperature, and with the nature and concentration of the fatty acid. Its value for each run was determined from the ratio of the values of t corresponding to 25% and 50% change, read from a smooth curve of the experimental points. The values of r and of the activation energy for acetic acid in N/10-solution differ from those found by Williamson and Hinshelwood for N/2-solution, but repetition of the experiments using their concentration confirmed their results for r and gave a value of E fairly close to theirs; E thus varies slightly with concentration. Tables V and VI show typical runs for acetic and palmitic acids at 20·0°.

TABLE. V.

Acetic acid.

$c = 0{\cdot}00349$; $a = 0{\cdot}0997$.

Mins.	x.	k.
18·5	0·01610	0·0467
30·8	0·02400	0·0453
39·0	0·02920	0·0450
55·0	0·03840	0·0466
78·0	0·04793	0·0454
88·0	0·05160	0·0451
113·0	0·05924	0·0459
208·0	0·07920	0·0453
230·0	0·08200	0·0451

($r = 0{\cdot}154$.)

Mean $k = 0{\cdot}0455$.

TABLE VI.

Palmitic acid.

$c = 0{\cdot}00553$; $a = 0{\cdot}100$.

Mins.	x.	k.
33	0·0176	0·0190
39	0·0222	0·0211
53	0·0282	0·0210
67	0·0343	0·0215
90	0·0421	0·0215
105	0·0468	0·0219
123	0·0516	0·0208
213	0·0697	0·0223
250	0·0748	0·0224

($r = 0{\cdot}16$.)

Mean $k = 0{\cdot}0213$.

The values of r which were found to fit the Goldschmidt equation are given in Table VII, and Table VIII shows the values of E and log PZ and of k at 20·0°.

TABLE VII.

Acid.	45°.	30°.	20°.	0°.
Acetic (N/2)	0·345	0·21	0·17	0·125
„ (N/10)	0·11	0·16	0·19	0·27
Propionic	0·13	0·19	0·18	0·27
Butyric	0·13	0·14	0·17	0·16
Hexoic	0·11	0·10	0·14	0·21
Pelargonic	0·25	0·27	0·39	0·32
Palmitic	0·11	0·14	0·13	—

TABLE VIII.

Acid.	E.	log PZ.	k_{20}.
Acetic (N/10)	12,450	7·955	0·0442
Propionic (N/10)	11,600	6·960	0·0384
Butyric (N/10)	11,800	7·183	0·0228
Hexoic (N/10)	11,700	7·073	0·0208
Pelargonic (N/10)	11,300	6·750	0·0198
Palmitic (N/10)	11,800	7·176	0·0221
Acetic (N/2)	10,800	6·710	0·0447

Catalysis by Undissociated Acid: Experiments in cyclo*Hexanol.*—In preliminary experiments it was found that the bimolecular constants for the esterification of acetic acid at 120° fell as the reaction proceeded. This was shown to be due to a retarding influence of the water formed in the reaction, the initial constant for N-acetic acid when the *cyclo*hexanol contained 0·5 mole/l. of added water being the same as the instantaneous value of k at 50% change in pure *cyclo*hexanol. The constant was independent of the initial concentration of the acetic acid and depended only on the water content of the solvent. The retardation was most marked at lower temperatures, and in some cases no drift in the bimolecular constants could be seen at the two highest temperatures. Constants were calculated from initial rates when the retardation was observed. Acetic acid was studied at 90°, 100°, 110°, 120°, 138°, and 154°, and the other acids at the five highest temperatures only. The Arrhenius equation was obeyed. N-Solutions were used, except for the less soluble palmitic and behenic acids, for which 0·5N- and 0·08N-solutions respectively were used.

Since a five-fold decrease in the concentration of palmitic acid itself caused a change in rate equal to the difference between palmitic and behenic acids, the results for the latter are, unfortunately, not directly comparable with the rest of the series. The results for palmitic acid at 0·5N are, however, closely comparable with the main series, since halving the concentration of acetic acid made no difference to the rate. Results are summarised in Table IX.

TABLE IX.

Acid.	E.	log PZ.	$k_{100} \times 10^6$.
Acetic	16,600	4·262	3·18
Propionic	16,500	4·038	2·24
Butyric	15,700	3·431	1·59
Hexoic	15,000	3·095	1·89
Pelargonic	15,400	3·335	1·91
Palmitic (N/2)	15,400	3·375	2·09
Behenic (N/12)	14,300	2·800	2·49

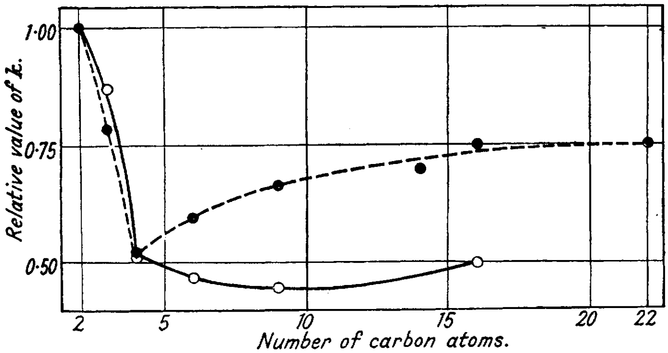
DISCUSSION.

The activation energies are recorded in Tables IV, VIII, and IX. When the undissociated acid is the catalyst and ethyl alcohol the solvent, the activation energy is constant within the limits of experimental error throughout the whole series. With hydrion as the catalyst in methyl alcohol, E falls from acetic to propionic acid, after which only small irregular variations occur. In the *cyclo*hexanol series there are decreases in the earlier members, followed by small irregular variations. We should expect from the electronic theory of organic reactions that the first few members of the fatty acid series would show progressive changes in the energy of activation for esterification. Since, however, a saturated carbon chain is not a good conductor for the inductive effect which determines this change, less and less influence of further substitution should be observed as the chain length is increased. This is in accordance with the experimental results, since no progressive change in E is observed after butyric acid, in any of the three cases. It is difficult to imagine any mechanism other than induction by which the activation energy could be changed. We therefore assume that, as the small observed variations in E, after butyric acid, are quite random, the real activation energy is exactly constant, and the variations are due to experimental error.

If this is admitted in principle, PZ becomes directly proportional to k. Since k can be measured more accurately than PZ itself, we shall take its variation in the series as giving the best indication of the true changes in PZ from butyric acid onwards.

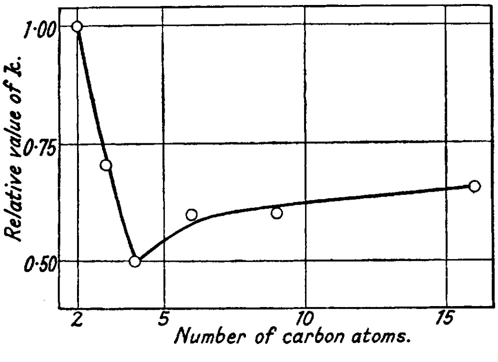
In Fig. 2 the relative rates of esterification in ethyl alcohol at 100° (catalyst undissociated acid) and in methyl alcohol at 20° (catalyst hydrion) are plotted against the number of carbon atoms in the fatty acid. In the former series there is a minimum at butyric acid, after which the rate rises steadily. In the latter series, however, the rate continues to decrease slightly, and never rises again above the value for butyric acid. If we accept the validity of the foregoing argument as to the constancy of *E*, we may regard these changes as proportional to those in *PZ*. The hydrocarbon chains are probably quite inert; the greater apparent collision area of the larger acids is irrelevant, because only collisions with carboxyl groups lead to reaction; in gases, increasing molecular weight would decrease *Z*; we must therefore seek for some explanation of the greater efficiency, or number, of collisions in the reactions which involves two molecules of the larger acids. The efficiency of collisions between the carboxyl groups of these cumbrous molecules would if anything decrease as their size increased. On the other hand, if the incompatibility of packing patterns of solvent and solute were great enough to lead to exclusion of the solute from certain regions of the solvent, we should expect *Z* to show the increase indicated by the experiments with the model. This effect would increase as the incompatibility became greater, *i.e.*, on ascending the homologous series.

FIG. 2.



Open circles = reaction catalysed by hydrion.
Black „ = „ „ undissociated acid.

FIG. 3.



The fact that no comparable increase in *PZ* is found with the hydrion-catalysed reaction (Fig. 2) lends some support to this view of the mechanism. Changes in the chemical properties of the acids might affect the values of *PZ* in both series, but the incompatibility of packing would operate differently with the hydrion-catalysed reactions. Although fatty acid molecules would still be excluded from the pattern of the solvent, hydrions would not; the latter would thus be free to move about the solution, without undergoing abnormally long sets of repeated collisions with the fatty acid molecules.

The relative rates of reaction at 100° for the various acids in *cyclo*hexanol are given in Fig. 3. Here the position is not quite so clear, since the total variations are smaller, and the measurements probably less accurate; but, on the assumption as before that the activation energy is really constant after butyric acid, there is a slight rise in *PZ* in the latter part of the series. This might be taken to indicate that in *cyclo*hexanol solution molecules of the higher acids become less able to fit in with the solvent pattern, although the effect is less marked than with smaller solvent molecules of ethyl alcohol.

With whatever reserve the mechanical model may be regarded, it did in fact suggest the whole of the chemical investigation, and even without it, the interpretation in terms of collision numbers appears reasonable.

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