Anal. Calcd. for $C_5H_7NO_2$: C, 53.1; H, 6.24; N, 12.38. Found: C, 52.7, 53.0; H, 6.22, 6.42; N, 11.76, 11.82.

The polymer was insoluble in, but swollen by, water. It was soluble in formamide, mineral and organic acids, and 37% aqueous formaldehyde. The polymer was insoluble in the common organic solvents such as alcohols, esters, ethers, hydrocarbons.

(*2*) *Polymerization in bulk.* A thick-walled, Pyrex, cappable tube was charged with 2.16 g. of monomer and 0.022 g. of azobisisobutyronitrile. The tube was capped, and polymerization carried out under nitrogen as before. After 48 hr., the contents of the tube had gelled. Isolation afforded 1.04 g. (46%) of polymer, η_{rel} (1% in formamide) 1.44.

(*3*) *Polymerization in an organic solvent.* A solution of 3.0 g. of III in 2.0 ml. of anhydrous benzene was polymerized at 70 over a 36-hr. period, with 0.03 g. of benzoyl peroxide as catalyst. Yield: 1.71 g. (57%) of polymer, η_{rel} (1% in formamide) 1.23.

Attempted polymerization of III *with titanium trichloride-hydroxylamine hydrochloride.* To 2.0 g. of III in 25 ml. of hexane, at 22°, were added, with stirring, 0.02 g. of hydroxylamine hydrochloride and 0.2 ml. of 20% titanium chloride solution (Lamotte Chemical Co.). In a few minutes, the temperature rose to 26°, and an oily product deposited on the sides of the flask. The product hardened in 15 min. to a solid, soluble in water and in methanol. It was purified by solution in methanol, and precipitation in excess ether. The product (0.85 g.), after recrystallization from methanol ether, melted at 146–148°. Elementary analysis (C, 47.95, 47.92; H, 5.86, 6.05; N, 14.09, 14.18) and a mixed melting point with V confirmed the identity of this product as V.

Hydrolysis of poly(N-vinyl-2-oxazolidinone). Poly(*N*-vinyl-2-oxazolidinone) 2.0 g., η_{rel} (1% in formamide) 2.5, was dissolved in 25 ml. of concd. hydrochloric acid. The solution was refluxed for 3 hr., while carbon dioxide evolved. The cooled, viscous solution, added slowly with stirring to 150 ml. of methanol, deposited a white, granular precipitate. After being filtered and washed with methanol, the hydrolyzate was redissolved in 25 ml. of water and reprecipitated in methanol. The specific viscosity of the hydrolyzed polymer in water was 15.7. Its intrinsic viscosity $[\eta]$ at 25° in 0.1*N* sodium chloride was 0.185.

WILMINGTON, DEL.

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY OF THE AMERICAN UNIVERSITY OF BEIRUT]

Kinetics of the Reaction of Phenyl Isocyanate with Aniline[1]

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Received December 5, 1960

The kinetics of the reaction between phenyl isocyanate and aniline at 25° in benzene and chlorobenzene as solvents has been studied as a function of the initial concentrations of the reactants. When the initial concentrations of the two reagents are approximately equal the rate expression is given by Rate = $k_3(C_6H_5NH_2)^2(C_6H_5NCO)$. The rate constant, k_3, depends on the initial concentrations. When aniline is initially in substantial excess, the rate expression is the same and $k_3 = 0.021 \pm 0.001$ l.^{2}/mole^2-sec. When phenyl isocyanate is in substantial excess the rate law becomes Rate = $k_4(C_6H_5NH_2)^2(C_6H_5NCO)^2$, where $k_4 = 1.0 \pm 0.2$ l.^{3}/mole^3-sec. The measurements were made on solutions saturated with the product, diphenylurea. A mechanism is proposed which involves two intermediate complexes of the reagents.

The reaction of isocyanates with amines has been studied previously by several investigators.[3–5] Reviews of the early work on this reaction and also of that on related reactions of isocyanates, especially with alcohols, have been published.[6,7] The results of Naegeli *et al.*[4] showed that an increase in the rate of reaction occurs when electron-attracting substituents are present on the isocyanate molecule or when electron-donating groups are attached to the amine. They further observed that carboxylic acids and tertiary bases, such as pyridine, are catalysts for the reaction.

Studies of the kinetics of the reaction between isocyanates and alcohols have led to the proposal that the first step in the mechanism is the formation of a complex between the isocyanate and the alcohol (or a base catalyst, if present) followed by the reaction of the complex with the alcohol to give the urethane product.[8] The kinetics of this reaction are between second and third order. Arnold *et al.*[7] suggested that the reaction between isocyanates and amines follows an analogous mechanism with the amine taking the place of the alcohol. They report some unpublished experiments of Craven which show that the amine-isocyanate reaction is not simply first-order in each reactant, as had been previously assumed.[3] The order with respect to amine was seen to vary from first to second, depending on the particular amine used.[7]

Baker and Bailey,[5] as a result of their extensive kinetic studies, conclude that the mechanism of the isocyanate-amine reaction is analogous to that for the isocyanate-alcohol reaction. As expected, the

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[1] Abstracted from the thesis of Marie H. Kasparian, submitted in partial fulfillment of the requirements for the Master of Science degree at the American University of Beirut, June 1960.

[2] Chemistry Department, University of California, Berkeley, Calif.

[3] T. L. Davis and F. Ebersole, *J. Am. Chem. Soc.*, **56**, 885 (1934).

[4] C. Naegli, A. Tyabji, and L. Conrad, *Helv. Chim. Acta*, 21, 1127 (1938).

[5] (a) J. W. Baker and D. N. Bailey, *J. Chem. Soc.*, 4649 (1957); (b) 4652 (1957); (c) 4663 (1957).

[6] J. H. Saunders and R. J. Slocombe, *Chem. Revs.*, **43**, 203 (1948).

[7] R. G. Arnold, J. A. Nelson, and J. J. Verbanc, *Chem. Revs.*, **57**, 47 (1957).

[8] (a) J. W. Baker and J. B. Holdsworth, *J. Chem. Soc.*, 713 (1947); (b) J. W. Baker and J. Gaunt, *J. Chem. Soc.*, 9 (1949); (c) 19 (1949); (d) J. W. Baker, M. W. Davies and J. Gaunt, *J. Chem. Soc.*, 24 (1949); (e) J. W. Baker and J. Gaunt, *J. Chem. Soc.*, 27 (1949).

amines and the product ureas are much better catalysts for the reaction than are the alcohols and product urethans, respectively. Their experimental observations were complicated by the fact that the ureas, which are only slightly soluble in the solvent benzene, tended to supersaturate easily. Even reactions run in solutions initially saturated with the product urea did not remove this difficulty. A further complication results from their observation that the solubility of ureas is greatly affected by the concentrations of the reagents, especially the amine, present.[5a] These factors resulted in considerable complexity in their experimental findings and consequent difficulties in the interpretation of the results.

The object of the present study is to investigate the isocyanate-amine reaction under conditions which minimize the difficulties observed by Baker and Bailey. The diphenylurea resulting from the reaction of phenyl isocyanate with aniline is less soluble than are the ureas resulting from the reaction of phenyl isocyanate with benzocaine or of cyclohexyl isocyanate with aniline, the two reactions most extensively reported by Baker and Bailey. Furthermore, the solubility of diphenylurea is not nearly so strongly affected by the concentrations of the reactants as is the case in the former reactions. In most of the runs reported in this study the solutions were initially saturated with product; the slowness with which the reactions proceed and regular agitation serve to keep the urea from seriously supersaturating the solutions. In those runs where initial saturation with product was not carried out, the crystals appeared early in the course of the reaction. The test of the effectiveness of the procedures is the fact that only in a few instances were discontinuous or sigmoid curves of the type reported by Baker and Bailey observed.

EXPERIMENTAL

Materials. All chemicals used were Eastman, White Label, unless specified otherwise. Phenyl isocyanate, *n*-butylamine, di-*n*-butylamine and aniline (Merck, extra pure) were each freshly distilled before using. Stock solutions of each of these reagents in benzene (Spectro-grade) or chlorobenzene were prepared within a few hours of the start of each run. The acid titrant solution was prepared by dissolving sufficient 70% perchloric acid (Merck, extra pure) in glacial acetic acid (Merck, extra pure) to make up approximately 0.05*M* solution. Sufficient acetic anhydride was added to react with 90% of the theoretical amount of water present.

Procedure. In order to prevent the reaction of phenyl isocyanate with moisture, all the glass apparatus was dried at 125° before using, all solutions were kept in glass-stoppered vessels and these were exposed to atmospheric moisture as little as possible.

Separate stock solutions of the reactants were prepared and, in most cases, were saturated with *N*,*N*′-diphenylurea at the temperature of the experiment: 25 ± 0.2°. After equilibrium became established portions of the reactants were mixed at the zero time for the experiment. The change of concentration of the reagents with time was followed in two different manners. When the initial concentrations of the reactants were approximately the same or when isocyanate was in excess (Tables II and III) the aniline concentration was followed by a titration method.[9] Samples of 1–10 ml. withdrawn from the reaction mixture were pipetted into 40–50 ml. of chilled glacial acetic acid. Methyl Violet indicator was added and the sample was titrated with standard 0.05*N* perchloric acid in acetic acid. The initial concentration of the aniline solution was measured in an analogous manner. The color of the indicator at the equivalence point in the mixed solvents was checked potentiometrically. The titrant solution was standardized against potassium acid phthalate. The initial concentration of the isocyanate solution was determined by a procedure similar to that proposed by Stagg.[10] A known volume of the isocyanate solution was allowed to react to completion with an excess of a standard *n*-butylamine or, preferably, di-*n*-butylamine solution in benzene, and the excess amine was then back-titrated with the standard perchloric acid. Checks using this method showed that the solutions of the reagents in benzene were stable for at least 1 week, that the presence of isocyanate does not interfere with the direct titration of the aniline and that the isocyanate does not react with the product, diphenylurea.

For those runs where aniline was initially in excess (Table I), the change in the isocyanate concentration was followed using the strong infrared absorption band at 2260 cm.$^{-1}$ of phenyl isocyanate.[11] Since the Beer-Lambert Law is not followed over the concentrations used (apparent molar absorption coefficients are 930 and 900 l.-mole^{-1} cm.$^{-1}$ at 6.4 × 10^{-3} and 7.7 × 19^{-2} m./l., respectively), a calibration curve was prepared which allowed concentrations to be determined to ±2%. Samples of the reaction mixture were withdrawn at appropriate intervals and placed in a cell with sodium chloride windows. The spectrum was scanned from 2200 to 2500 cm.$^{-1}$ The time was recorded when the instrument went through the peak at 2260 cm.$^{-1}$ The time at which the sample was at room temperature was negligible in comparison with the duration of the reaction intervals. It was demonstrated that neither aniline nor diphenylurea interfered with this analytical procedure.

RESULTS

The experiments reported here fall into three principal categories: those where aniline is in large excess in the reaction mixture, those where isocyanate is in large excess and those where the initial concentrations of the two reactants are nearly equal. These conditions are chosen so as to simplify the mathematics of treating a reaction of uncertain order. For runs which fall in one of the first two categories, the concentration of the reagent in excess can be assumed, to a first approximation, to remain constant throughout the course of the reaction. The data can then be plotted as a pseudo-order reaction in the reagent in lesser concentration. The additional assumption is made that the reaction follows a simple rate law of the form where x is

$$\frac{dx}{dt} = k_n a^{y} i^{z} \qquad (1)$$

the extent of the reaction and a and i are the instantaneous concentrations of aniline and isocyanate, respectively.

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[9] J. S. Fritz, *Acid-Base Titrations in Non-Aqueous Solvents*, G. Frederick Smith Chemical Co., Columbus, Ohio (1952).

[10] H. E. Stagg, *Analyst*, **71**, 557 (1946).

[11] M. E. Bailey, V. Kirss, and R. G. Spaunburgh, *Ind. Eng. Chem.*, **48**, 794 (1956).

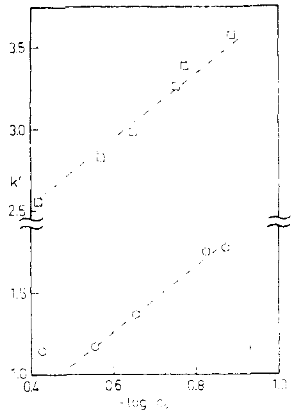

Fig. 1. Pseudo-rate constants for the reaction of phenyl isocyanate with aniline, in benzene at 25°, at various initial concentrations of the reagent in excess. □, $-\log k_1'$ *vs.* $-\log a_0$, where a_0 is the initial concentration of aniline (for runs in Table I): ○, $-\log k_2'$ *vs.* $-\log i_0$, where i_0 is the initial concentration of phenyl isocyanate (for runs in Table II). Dashed lines are drawn with slopes equal to 2.0

Aniline in excess. When aniline is present in large excess, the data on the change of phenyl isocyanate concentration with time are found to follow a first-order dependence; *i.e.* z in Equation 1 is equal to unity. The pseudo-first-order rate constants, k_1', for these runs are then

$$k_1' = k_n a_0^y \tag{2}$$

By writing Equation 2 in the logarithmic form

$$\log k'_1 = \log k_n + y \log a_0 \tag{3}$$

values for k_n and y can be determined by plotting $\log k_1'$ *versus* $\log a_0$ for experiments with different values of the initial concentration of aniline. The resulting plot is shown in Fig. 1 (upper curve), from which it can be readily seen that y, the order in aniline, is very nearly second (dashed line). Thus the overall order of the reaction is third.

With this information the more exact equation

$$\frac{dx}{dt} = k_3(a_0 - x)^2(i_0 - x) \tag{4}$$

can be applied. Upon integration, Equation 4 yields the result

$$k_3t + \text{const.} = \frac{1}{i_0 - a_0}\left(\frac{1}{a} + \frac{1}{a_0 - i_0}\ln\frac{i}{a}\right) \tag{5}$$

When the measured values for i and the corresponding calculated values for a are put into the right-hand term of equation 5 and the results are plotted against time, curves of the type shown in Fig. 2 are obtained. The values of k_3 obtained from the slopes of these plots are tabulated in Table I. For an initial concentration of aniline greater than about 0.15 mole/l. the third-order rate constant is 0.021 ± 0.001 l.^{2}/mole^{2-} sec. For initial concentrations of aniline less than 0.15 mole/l., the rate constant appears to rise with decreasing initial aniline concentrations.

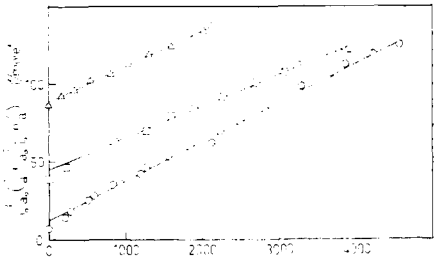

Fig. 2. Typical experimental curves; reaction of phenyl isocyanate with excess aniline, in benzene at 25°. For clarity, the origin of the lowest curve (circles) has been displaced downward 50 units and that of the uppermost curve (triangles) upward by 50 units. The point furthest to the right on each curve corresponds to 72–82% of completion of the reaction

TABLE I

RATE CONSTANTS FOR THE REACTION WITH EXCESS ANILINE IN BENZENE SATURATED WITH DIPHENYLUREA, AT 25°

$(C_6H_5NH_2)_0$ mole/l.	$(C_6H_5NCO)_0$ mole/l.	$k_3 \times 10^2$ l.^{2}/mole^2-sec.
0.382	0.0339	1.97
0.270	0.0339	2.02
0.225	0.0282	2.13
0.177	0.0432	2.20
0.169	0.0441	2.01
0.130	0.0444	2.5
0.1125	0.0282	3.1
0.094	0.0441	5.6[a]

[a] No diphenylurea added initially.

Isocyanate in excess. By applying an analogous procedure to the experiments where isocyanate was the reactant in excess, the data on the change of aniline with respect to time are found to follow a second-order dependence. Utilizing Equation 1, the pseudo-second-order rate constants can be written

$$k_2' = k_n i_0^z \tag{6}$$

The data plotted according to the logarithmic form of this equation are shown in Fig. 1 (lower curve) and the order in isocyanate appears to be second as well. The overall order of the reaction with excess isocyanate present is, therefore, fourth and follows the rate law

$$\frac{dx}{dt} = k_4(a_0 - x)^2(i_0 - x)^2 \tag{7}$$

Integration of this expression yields

$$k_4t + \text{const.} =$$

$$\frac{1}{i_0 - a_0}\left[\frac{2}{i_0 - a_0}\left(\frac{1}{a} + \frac{1}{a_0 - i_0}\ln\frac{i}{a}\right) - \frac{1}{ai}\right] \tag{8}$$

which in turn gives curves of the type shown in Fig. 3 for individual experiments using excess

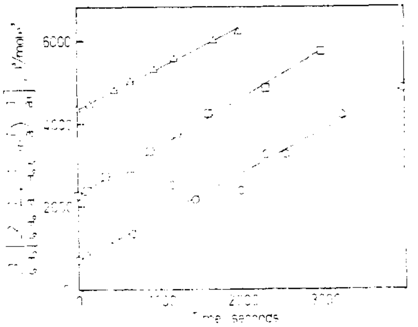

Fig. 3. Typical experimental curves; reaction of aniline with excess phenyl isocyanate, in benzene at 25°. For clarity the origin of each successive curve is displaced upward by increments of 2000 units. Reaction range up to 72–88% of completion

isocyanate. Values of k_4 obtained from plots using Equation 8 are given in Table II. For an initial phenyl isocyanate concentration greater than about 0.15 mole/l. the value of k_4 is 1.0 ± 0.2 l.^{3}/mole^3-sec. At the lower initial concentrations of phenyl isocyanate the measured rate constants are larger.

TABLE II

Rate Constants for the Reaction with Excess Phenyl Isocyanate in Benzene Saturated with Diphenylurea, at 25°

$(C_6H_5NH_2)_0$ mole/l.	$(C_6H_5NCO)_0$ mole/l.	k_4 l.^{3}/mole^3-sec.
0.0362	0.374	0.66
0.0356	0.280	1.16
0.0354	0.222	0.98
0.0364	0.150	1.07
0.0356	0.134	1.47
0.0365	0.068	2.62
0.0470	0.088	5.4[a]

[a] No diphenylurea added initially.

Equal initial concentrations. Since the reaction between isocyanate and aniline is observed to be third-order when aniline is in excess and fourth-order when phenyl isocyanate is in excess, it might be expected that when the initial concentrations are about the same, the overall order might be intermediate between the two. Preliminary plots of the data indicated that third-order is the more correct. The technique used was to average the concentrations of the two reagents, which were never exactly equal to each other but usually differed by 3 to 5%, and plot the appropriate functions *versus* time.

Once the reaction was determined to be approximately third-order, corrections were made for the nonequality of the reactant concentrations by using a mathematical technique originally suggested by Arrhenius.[12] The kinetics were assumed to follow Equation 4. By letting $a_0 = c_0 - s$ and $i_0 = c_0 + s$, where c_0 is the average initial concentration and $2s$ is the excess of i_0 over a_0, Equation 5 can be expanded to a more useful form. Substituting into Equation 5, one obtains

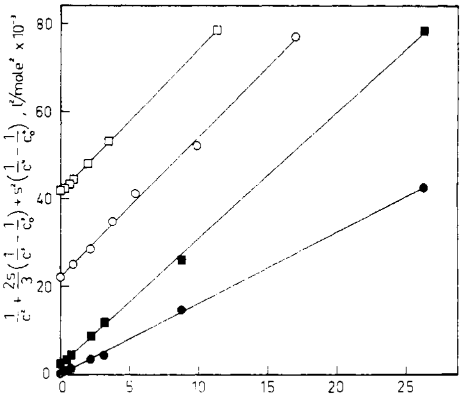

Fig. 4. Typical experimental curves; reaction with approximately equal initial concentrations of phenyl isocyanate and aniline at 25°. Open symbols: benzene as solvent. Solid symbols: chlorobenzene as solvent. Coordinates apply to solid symbols. For open circles: origin is displaced upward by 20 units; vertical scale should be multiplied by 0.1 and time scale multiplied by 0.16. For open squares: origin is displaced upward by 40 units and scale factors of 0.5 and 0.8, respectively, should be applied. (Diphenylurea was not added to these solutions initially.) Data shown represent the reaction from 0 to between 79 and 93% of completion of the reaction

$$k_3t = \frac{1}{2s}\left[\frac{1}{c - s} - \frac{1}{c_0 - s} + \frac{1}{2s}\ln\frac{(c_0 - s)(c - s)}{(c_0 - s)(c - s)}\right] \quad (9)$$

Series expansion of equation 9 followed by collection of terms gives

$$k_3t = \frac{1}{2}\left[\frac{1}{c^2} - \frac{1}{c_0^2} + \frac{2s}{3}\left(\frac{1}{c^3} - \frac{1}{c_0^3}\right) + s^2\left(\frac{1}{c^4} - \frac{1}{c_0^4}\right) + \ldots\right] \quad (10)$$

When s is zero, this equation reduces to the third-order law expected for reactants of equal concentration. The correction terms shown in Equation 10 were applied to the data, and corresponding plots of typical runs in benzene or in chlorobenzene as solvent are shown in Fig. 4. The linearity of the plots testifies as to the correctness of the conclusion that the reaction is close to third-order. Third-order rate constants calculated from experimental curves of the type shown in Fig. 4 are assembled in Table III for the various runs in the two solvents. These experiments were run without any initial addition of diphenylurea. Inspection of the results in Table III shows that the values of k_3 are not constant, but tend to decrease with increasing initial concentration of the reactants. There is no appreciable difference between the rate constants in benzene and

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(12) S. Arrhenius, *Z. physik. Chem.*, 1, 110 (1887); A. A. Frost and R. G. Pearson, *Kinetics and Mechanism*, John Wiley and Sons, Inc., New York, N. Y., 1953, pp. 18–19.

TABLE III

Rate Constants for the Reaction with Nearly Equal Initial Concentrations, at 25°

$(C_6H_5NH_2)_0$ mole/l.	$(C_6H_5NCO)_0$ mole/l.	k_3 l.²/mole²-sec.
Benzene solvent		
0.0706	0.0662	0.101
0.0353	0.0331	0.106
0.0360	0.0328	0.092
0.0235	0.0221	0.210
0.0240	0.0218	0.155
Chlorobenzene solvent		
0.0643	0.0602	0.080
0.0319	0.0352	0.097
0.0318	0.0298	0.128
0.0213	0.0235	0.155
0.0211	0.0197	0.144

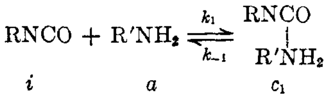
those in chlorobenzene, for the same initial concentrations.

DISCUSSION

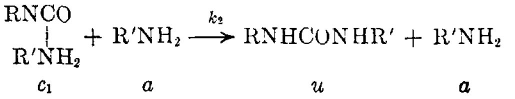
The reactions of amines with isocyanates are known[5] to be base-catalyzed by the amines as well as autocatalyzed by the weakly basic product ureas. Using the value reported by Arnold *et al.*[7] for the ratio of the rate constant for autocatalysis by diphenylurea to the rate constant for catalysis by aniline in dioxane as solvent, we can estimate that the corresponding rate in benzene should not be greater than 0.1 and may well be less than this value. The solubility data at 20° of Baker and Bailey[5a] indicate that the maximum concentration of diphenylurea present in the reacting medium in the present study is between 0.5 and 1.0×10^{-3} mole/l., depending principally on the concentration of aniline present. This assumes that the solutions are not supersaturated. Calculations based on these estimates readily show that catalysis by diphenylurea should not be more than a few tenths of a per cent compared with catalysis by aniline throughout most of each reaction studied, and would rise no higher than 1.5% even toward the end of the reaction where the aniline concentration has been depleted.

Baker and Bailey[5b,c] found on the other hand, that most of their reaction mixtures did seriously supersaturate, giving rise to sigmoid curves when the data were plotted. In the present study supersaturation was not a serious problem, due to the fact that most of the solutions were pre-saturated with diphenylurea and also due to the microcrystalline nature of the product, which formed within a few minutes of the start of any reaction and remained suspended throughout the body of the reaction mixture. The ultimate test of the validity of this conclusion is the fact that breaks and sigmoid curves of the type found by Baker and Bailey were not observed here, as can be seen in the figures. On the basis of these observations the subsequent discussion will be carried out assuming that product-catalysis is not an important factor in these studies.

In the paper of Arnold *et al.*[7] a mechanism is proposed for the reaction of isocyanates with amines which for our purposes may be written



$$\underset{i}{\text{RNCO}} + \underset{a}{\text{R'NH}_2} \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} \underset{c_1}{\begin{matrix}\text{RNCO}\\ | \\ \text{R'NH}_2\end{matrix}} \qquad (11)$$



$$\underset{c_1}{\begin{matrix}\text{RNCO}\\ | \\ \text{R'NH}_2\end{matrix}} + \underset{a}{\text{R'NH}_2} \xrightarrow{k_2} \underset{u}{\text{RNHCONHR'}} + \underset{a}{\text{R'NH}_2} \qquad (12)$$

This scheme is based on the earlier one of Baker and Gaunt,[8] for the reaction of isocyanates with alcohols. Applying steady state kinetics to the intermediate complex, the rate of the reaction is given by

$$\text{Rate} = \frac{k_1k_2a^2i}{k_{-1} + k_2a} \qquad (13)$$

A consequence of Equation 13 is the fact that for $k_{-1} > k_2a$ the reaction should be third-order overall and for $k_{-1} > k_2a$ it should be second-order overall. The fact that the reaction studied here was observed to be third-order, even in excess of aniline, leads to the conclusion that $k_{-1} > k_2a$ for the range studied.

The mechanism of Equations 11 and 12, however, is unable to account for the observation of the present study that, when phenyl isocyanate is present in excess, the kinetics correspond to a fourth-order reaction. For this purpose an extension of the above mechanism is required. The following is proposed as a simple, plausible scheme which is consistent with the observed kinetics.

$$i + a \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} c_1 \qquad (11)$$

$$c_1 + a \underset{k_{-2}}{\overset{k_2}{\rightleftharpoons}} c_2 \qquad (14)$$

$$c_2 \xrightarrow{k_3} u + a \qquad (15)$$

$$c_2 + i \xrightarrow{k_4} 2u \qquad (16)$$

The essence of this scheme is that complex 1 reacts reversibly with a molecule of aniline to produce a second complex, c_2, which has a definite lifetime and, hence, concentration in the reaction mixture. The first two steps can be considered to come quickly to equilibrium, in which case Equation 15 will represent the rate-determining step in excess aniline and Equation 16 the rate-determining step in excess phenyl isocyanate. Under these conditions

$$c_1 = \frac{k_1}{k_{-1}}\,ia = K_1ia \qquad (17)$$

and

$$c_2 = \frac{k_2}{k_{-2}}\,c_1a = \frac{k_1k_2}{k_{-1}k_{-2}}\,ia^2 = K_1K_2ia^2 \qquad (18)$$

Therefore, the reaction rate equations become, in excess aniline

$$\frac{dx}{dt} = K_1K_2k_3ia^2 \tag{19}$$

and in excess phenyl isocyanate

$$\frac{dx}{dt} = K_1K_2k_4i^2a^2 \tag{20}$$

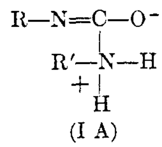
in agreement with the experimental observations. These can be expressed in a single equation

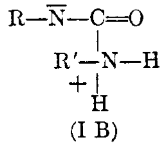
$$\frac{dx}{dt} = K_1K_2ia^2(k_3 + k_4i) = K_1K_2i^2a^2\left(\frac{k_3}{i} + k_4\right) \tag{21}$$

Using the experimentally determined values of the rate constants we get the result that $K_1K_2k_3$ is 0.021 l.^{2}/mole^2-sec. and $K_1K_2k_4$ is 1.0 l.^{3}/mole^3-sec. Accordingly, the transition from third-order to fourth-order kinetics should come at a phenyl isocyanate concentration of about 0.02 mole/l. Table II does in fact show that, as the initial concentration of phenyl isocyanate is reduced in the reactions where it is the reagent in excess, the apparent fourth-order rate constant increases, as would be expected from Equation 21. On the other hand Tables I and III show trends for the apparent third-order rate constants which are in the opposite direction from that predicted by Equation 21. A possible explanation would be that when aniline is in relatively low concentration one of the steps leading up to the formation of Complex 2 becomes at least partially rate-determining. Thus, the mechanism of Equations 11, 14, 15, and 16 is only a suggestive one which may need modification in the light of further studies.

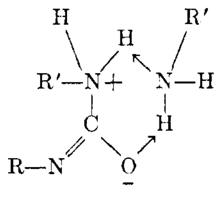
It might be argued that the observed fourth-order kinetics arise instead from the increased polarity of the medium due to the presence of the high concentration of phenyl isocyanate. The data of Table III, however, show clearly that not only the order of the reaction but even the numerical values of the rate constants remain unchanged when the solvent is changed from benzene (dielectric constant, $\epsilon = 2.27$)[13] to chlorobenzene ($\epsilon = 5.61$).[13] In view of this result it is highly unlikely that a benzene solution containing less than 3.5 mole % of phenyl isocyanate ($\epsilon = 8.8$)[13] would be of sufficiently different polarity to appreciably alter the kinetics. Furthermore, the solutions containing excess aniline ($\epsilon = 7.0$)[13] would be expected to show a similar effect of nearly equal magnitude.

The complexes proposed in the mechanism of Equations 11, 14, 15, and 16 are not unreasonable ones to suppose. Complex 1 undoubtedly results from interaction between the lone pair of electrons on the nitrogen of the amine and the electrophilic carbon of the isocyanate molecule.

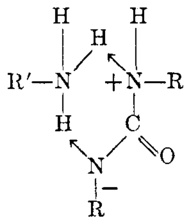

(I A)


(I B)

The studies of the effects of substituents on the relative reaction rates[4] strongly support this conclusion. The two structures above would lead to the formation of different complexes upon addition of a molecule of amine.

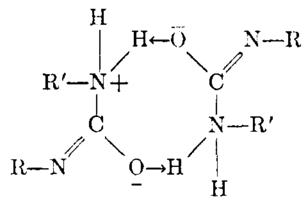

(II A)

or

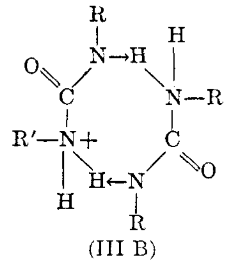

(II B)

Each of these involves the transfer of protons. Complex II A could break up to form the enol form of the product urea and a molecule of amine; Complex II B would give the product urea and a molecule of amine directly.

In the presence of excess isocyanate, however, these six-membered ring complexes may react with another molecule of isocyanate, through intermediates of the type


(III A)

or


(III B)

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(13) *Landolt-Bornstein Tabellen*, 6th ed., Vol. II, part 6, Springer-Verlag, Berlin, 1959, p. 613.

Either of these intermediates can then break up to give two molecules of the product area (or one molecule of the urea and one molecule each of phenyl isocyanate and aniline). Whether series A or series B is the more likely is difficult to predict; there is no direct experimental evidence to support one over the other.

Further studies are indicated in the direction of obtaining such direct evidence of the existence of these complexes. Perhaps spectroscopic studies of the system during the course of the reaction would show their presence. More extensive kinetic studies in the region of low concentrations of both reactants and at even higher excess concentrations of aniline than those studied here are called for.

Acknowledgment. The authors are indebted to Dr. Costas Issidorides for helpful discussions concerning the mechanism of the reaction, and to Messrs. Abdu Bardawil and Ismat Abu-Isa, who worked on some of the preliminary experiments and contributed to the development of techniques. This work was supported by grants from the Petroleum Research Fund (PRF #421-A) and from Sigma Xi-RESA. Funds for the purchase of equipment and supplies were available from the A.U.B.-Rockefeller Foundation Research Fund.

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Notes

A department for short papers of immediate interest.

1,3-Diaxial Methyl Hydroxyl Interaction

ERNEST L. ELIEL AND H. HAUBENSTOCK

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The 1,3-diaxial methyl hydroxyl interaction has been estimated to be 2.15 kcal./mole by the application of Equation 1 to the rate of acetylation of 3,3-dimethylcyclohexanol with acetic anhydride in pyridine.[1,2] In this equation K is the conformational equilibrium constant shown in Fig. 1, k is the specific rate constant for 3,3-dimethylcyclohexanol, k_e is the specific rate constant for *trans*-4-*t*-butylcyclohexanol, assumed to be identical with that of conformation E in Fig. 1, and k_a is taken as zero, the acetylation rate of conformation A in Fig. 1 being assumed negligible.

$$K = (k_a - k)/(k - k_e) \qquad (1)$$

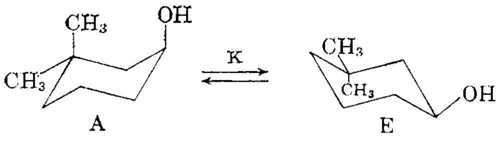
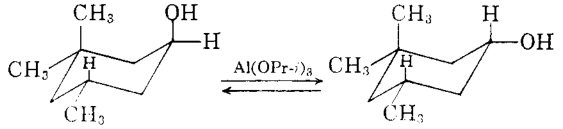


Figure 1



trans (A) *cis* (E)

Figure 2

The present communication reports an independent determination of the 1,3-diaxial methyl hydroxyl interaction by means of the chemical equilibration of *cis*- and *trans*-3,3,5-trimethylcyclohexanol (Fig. 2).

trans-3,3,5-Trimethylcyclohexanol[3] was prepared by the hydrogenation of isophorone or of 3,3,5-trimethylcyclohexanone (dihydroisophorone) in glacial acetic acid at room temperature over a platinum oxide catalyst. The catalytic reduction of dihydroisophorone under these conditions gave a mixture containing approximately 95% of the *trans* isomer. The *cis*-alcohol was prepared from commercial 3,3,5-trimethylcyclohexanol by purification through the acid phthalate.[4]

Equilibration of the 3,3,5-trimethylcyclohexanols was accomplished by boiling either isomer with aluminum isopropoxide and a small amount of acetone in anhydrous isopropyl alcohol. The equilibrium composition at 84.5° was determined by gas chromatography to be 94% *cis* and 6% *trans* starting from either isomer. The equilibrium constant for the equilibrium shown in Fig. 2 is 15.7 at 357.7°K., corresponding to a $\Delta F°$ of -1.96 kcal./mole. In isomer A (Fig. 2) the following 1,3-diaxial interactions are present: CH_3—OH, CH_3—H, and OH—H. The concentration of the conforma-

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(1) E. L. Eliel, *J. Chem. Ed.*, **37**, 126 (1960).

(2) For the derivation of an alternative form of equation (1) see S. Winstein and N. J. Holness, *J. Am. Chem. Soc.*, **77**, 5562 (1955).

(3)(a) E. G. Peppiatt and R. J. Wicker, *J. Chem. Soc.*, 3122 (1955); (b) I. Alkonyi, *Chem. Ber.*, **92**, 1130 (1959).

(4) R. S. Ro, unpublished results.