evidence to substantiate any predictions concerning the details of this breakdown process.

## Experimental

***t*-Amyl Alcohol.**—The *t*-amyl alcohol (Eastman Kodak Co.) was dried over potassium carbonate and then fractionated using a four-foot helix packed column. A cut of b.p. 100.9–101.0° (760 mm.) was retained for use in the kinetic experiments.

**Kinetic Experiments.**—Dilute solutions of bromine (Mallinckrodt Reagent grade) in carbon tetrachloride (Eastman Kodak Co., white label) were prepared at 25° and standardized iodometrically. For the rate studies samples of the bromine solutions were mixed with known amounts of *t*-amyl alcohol, and the resultant solutions were stored in the constant temperature bath at 25°. Samples of these solutions were removed from time to time for analysis by spectrophotometric or volumetric methods.

In the spectrophotometric procedure the samples were transferred to 1-cm. absorption cells for optical density measurements against a carbon tetrachloride blank. In general the wave length at which readings were made in any given run was chosen so that initial optical density readings were less than 1.0 and greater than 0.30. The rate constants based on data obtained in following an individual run by making a series of measurements at two different wave lengths were closely similar. For the fast runs (1–10) there was not sufficient time to take fresh samples for each optical density determination. In these runs the optical measurements were made using a single sample kept in the 1-cm. cell which in turn was stored at 25° within the cell housing of the spectrophotometer.

In the volumetric analytical procedure samples of known volume were removed and added to aqueous potassium iodide. The resultant mixture was then titrated with standard sodium thiosulfate solution.

**Test for Hydrogen Bromide in the Reaction Products.**—To 3 cc. of *t*-amyl alcohol was added 3 cc. of 0.38 *M* bromine in carbon tetrachloride and the resultant solution (*cf.* run 3) was allowed to decolorize completely. A few drops of phenolphthalein were added, and the solution was titrated with 0.1 *N* sodium hydroxide. Addition of several drops of base caused a temporary pink coloration which faded when the mixture was shaken vigorously. Further additions of base caused a similar behavior. A suspension of trimethylethylene dibromide in water responded in a similar fashion to titration with sodium hydroxide. Results of these experiments would indicate that little or no hydrogen bromide was present in the reaction medium after the bromine had been decolorized.

**The Reaction of *t*-Amyl Alcohol with Bromine in Carbon Tetrachloride Solution.**—Ten cc. of *t*-amyl alcohol (0.091 mole) was added to a solution of 14.5 g. of bromine (0.091 mole) in 220 cc. of carbon tetrachloride. The resulting solution was allowed to stand in diffuse light at room temperature. Separation of a water phase became noticeable within one-half hour. Two days later the bromine color had disappeared entirely. The reaction mixture was then washed with aqueous sodium bicarbonate and dried over calcium chloride. The carbon tetrachloride was removed under reduced pressure, and the residue was distilled from a modified Claisen flask. Five and one-half grams of trimethylethylene dibromide ((b.p. 100–103°) (101 mm.) and $n^{27}\text{D}$ 1.5070)[15] was collected. About 3 g. of lachrymatory residue which underwent decomposition on further attempts at distillation remained in the distilling flask. This residue probably contained a substantial quantity of the dibromide.

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[15] *Cf.* C. M. Suter and H. D. Zook, THIS JOURNAL, **66**, 741 (1944).

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[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY, UNIVERSITY OF CALIFORNIA AT DAVIS]

# Equilibrium Constants for the Interaction of Halogens with Alcohols and Ethers

BY R. M. KEEFER AND L. J. ANDREWS

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The visible absorption of iodine in cyclohexane or carbon tetrachloride solutions is markedly altered by the addition to the solutions of small amounts of isopropyl, *t*-butyl or *t*-amyl alcohol or *n*-butyl ether. These changes are explained on the assumption that 1:1 halogen–alcohol or ether complexes are formed in these solutions. Spectrophotometric methods have been used to evaluate the equilibrium constants for the interaction of iodine with certain of these oxygenated substances in cyclohexane or carbon tetrachloride at 25°. Data on the interaction of bromine with *t*-butyl alcohol are also presented. Solutions of widely varying alcohol concentration were used in the studies of the *t*-butyl alcohol–iodine complex in cyclohexane. The results have been examined to determine the range of alcohol concentration over which the equilibrium constants for complex formation remain constant.

The results of a kinetic study of the reaction of *t*-amyl alcohol with bromine to form trimethylethylene dibromide in carbon tetrachloride solution indicate that a 1:1 alcohol–bromine complex is formed as a reaction intermediate.[1] From the kinetic data, obtained for the most part by measurements on solutions in which the *t*-amyl alcohol concentration was 1 *M* or less, it was calculated that the value of the concentration equilibrium constant $K_c$, as defined by equation (1), was 2.2 mole$^{-1}$ l. at 25°.

$$K_c = (\mathrm{ROH{\cdot}X_2})/(\mathrm{ROH})(\mathrm{X_2}) \tag{1}$$

It has seemed desirable to check the magnitude of this constant by direct determination of the concentration of the complex in the alcohol–bromine solutions. Unfortunately the reaction of *t*-amyl alcohol to form trimethylethylene dibromide occurs too rapidly to permit such a determination. However dilute carbon tetrachloride solutions of bromine in *t*-butyl alcohol and of iodine in *t*-amyl alcohol are relatively stable. Equilibrium constants for halogen–alcohol complex formation in these solutions have therefore been determined on the assumption that in order of magnitude they should be comparable with the equilibrium constant for the interaction of *t*-amyl alcohol with bromine.

The constants reported in this communication have been evaluated through studies of the visible absorption of the solutions using procedures similar to those previously employed to study complex formation of iodine with alcohols and ethers[2] and with dioxane[3] in inert solvents. This method is feasible since the visible absorption maxima of the complexes lie further toward the ultraviolet than do those of the free halogens. With bromine

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[1] L. J. Andrews and R. M. Keefer, THIS JOURNAL, **75**, 3557 (1953).

[2] J. S. Ham, *J. Chem. Phys.*, **20**, 1170 (1952).

[3] J. A. A. Ketelaar, C. van de Stolpe and H. R. Gersmann, *Rec. trav. chim.*, **70**, 499 (1951).

solutions the accuracy of the method is limited by the fact that the differences in the absorption curves for free and complexed bromine are not as marked as those noted for iodine solutions. Although it has been shown that measurements of the ultraviolet absorption peak of *t*-butyl alcohol-iodine solutions at 236 mμ can be utilized for evaluation of interaction constants,[2] measurements in this region of the spectrum have been avoided in the present work. In the bromine solutions, at least, the ultraviolet absorption of the complex appears to be masked by the intense absorption of traces of trihalide ions.

A series of carbon tetrachloride solutions in which the alcohol concentrations varied from 1 to 0.1 *M* and the halogen concentrations varied from $10^{-4}$ to $10^{-3}$ *M* were investigated spectrophotometrically at 25° against the halogen-free alcohol solution as a blank over a wave length range of 440–600 mμ. As shown recently[4] in an investigation of dioxane–iodine complex formation the absorption of such solutions at any given wave length can be accounted for by equation (2)

$$\frac{1}{\epsilon_a - \epsilon_{X_2}} = \frac{1}{\epsilon_c - \epsilon_{X_2}} \times \frac{1}{K_c C} + \frac{1}{\epsilon_c - \epsilon_{X_2}} \qquad (2)$$

in which $\epsilon_a = \log_{10}(I_0/I)/l(X_2)_t$, $l$ = the length of the light path in cm., $(X_2)_t$ = total molar iodine concentration, $C$ = the molar concentration of the alcohol, $K_c$ is the equilibrium constant and $\epsilon_{X_2}$ and $\epsilon_c$ are the molar extinction coefficients of free and complexed halogen. Since $\epsilon_{X_2}$ can be evaluated independently from measurements of the halogen spectrum in the inert solvent, it follows from equation (2) that one can evaluate $K_c$ by graphical interpretation of the spectrophotometric data. The values of $K_c$ and of $\epsilon_c$ obtained from the slopes of the straight lines obtained by plotting $1/(\epsilon_a - \epsilon_{X_2})$ *versus* $1/C$ values for a series of solutions are given in Table I.

TABLE I

EQUILIBRIUM CONSTANTS (25°) FOR THE COMPLEXES

| Complex | λ, mμ | $\epsilon_{X_2}$ | $\epsilon_c$ | $K_c$ |
|---|---|---|---|---|
| *t*-Amyl alcohol–$I_2$[a] | 440 | 87 | 817 | 0.91 |
| | 460 | 248 | 865 | 1.17 |
| | 540 | 789 | 179 | 0.93 |
| *t*-Butyl alcohol–$Br_2$[a] | 500 | 96 | 51 | 0.59 |
| | 520 | 65 | 24.3 | 0.50 |
| Isopropyl alcohol–$I_2$[b] | 460 | 188 | 1021 | 1.03 |
| | 510 | 870 | 359 | 1.32 |
| | 520 | 937 | 283 | 0.99 |
| *t*-Butyl alcohol–$I_2$[b] | 440 | 61 | 686 | 1.37 |
| | 460 | 188 | 805 | 1.60 |
| | 520 | 937 | 437 | 1.40 |
| | 540 | 878 | 272 | 1.40 |
| | 560 | 601 | 185 | 1.38 |
| *n*-Butyl ether–$I_2$[b] | 280 | 130 | 2410 | 0.71 |
| | 290 | 73 | 1235 | .74 |
| | 440 | 61 | 622 | .74 |
| | 460 | 188 | 958 | .74 |

[a] In carbon tetrachloride as a solvent. [b] In cyclohexane as a solvent.

Included in this table are the results of similar measurements to determine equilibrium constants for the interaction of iodine with isopropyl and *t*-butyl alcohols and with *n*-butyl ether in cyclohexane solutions. These alcohols were investigated to see whether the pronounced observed difference in tendency for the formation of iodine complexes of isopropyl iodide and *t*-butyl iodide[5] is also apparent for the formation of the corresponding alcohol complexes. The ether was investigated, since its iodine solutions were unquestionably free of triiodide ion, to check the equilibrium constants for 1:1 iodine complex formation obtained from light absorption measurements in the visible region against those obtained by studies in the ultraviolet region.

**The Validity of the Experimental Method.**—In the derivation of equation (2) it was assumed that the activities of the halogen–alcohol complex, the alcohol and the free halogen were proportional to their concentrations. On the basis of vapor pressure measurements it has been concluded that in cyclohexane solution, *t*-butyl alcohol begins to deviate markedly from Henry's law when its concentration increases beyond 0.6 *M*.[6] Such deviations, which should be considered in interpreting the experimental data on the halogen–alcohol complexes, probably result from aggregation of the *t*-butyl alcohol molecules.[7]

To determine that range of alcohol concentration over which $K_c$ values remain constant spectrophotometric studies of the *t*-butyl alcohol–iodine complex were made on cyclohexane solutions the alcohol concentration of which varied from 0.065–10 *M*. The graphical analysis according to equation (2) of data taken on these solutions at several different wave lengths is given in Fig. 1. The data taken at alcohol concentrations up to about 1 *M* conform reasonably well to a linear plot. This fact suggests that Henry's law applies over this concentration range.

This point was further tested by employing the extinction coefficients for the complexes, $\epsilon_c$, determined from the ordinate intercepts of Fig. 1, for the calculation of concentrations of the complex and of the free halogen and of $K_c$ values at each alcohol concentration. Typical $K_c$ values obtained in this fashion are listed in Table II. It is obvious that the $K_c$ values remain constant as the alcohol concentration is increased to 0.9 *M*, and the corresponding $K_c$ values obtained at different wave lengths do not differ significantly. For 1.62 *M* *t*-butyl alcohol solutions $K_c$ values are about 30% higher than those for the more dilute alcohol solutions. For solutions above 1.62 *M* in alcohol concentration, $K_c$ values increase rapidly with increasing alcohol concentration It should be noted that these results are based on data taken both in the wave length region (460 mμ) in which the measured light absorption of the solutions results primarily from the complex and in the region (520–560 mμ) in which the major absorption is produced by free iodine.

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(4) J. A. A. Ketelaar, C. van de Stolpe, A. Goudsmit and W. Dzcubas, *Rec. trav. chim.*, **71**, 1104 (1952).

(5) R. M. Keefer and L. J. Andrews, THIS JOURNAL, **74**, 1891 (1952).

(6) L. Prigogine and A. Desmyter, *Trans. Faraday Soc.*, **47**, 1137 (1951).

(7) E. G. Hoffmann, *Z. physik. Chem.*, **B53**, 179 (1943).

TABLE II

THE EFFECT OF CHANGING WAVE LENGTH AND ALCOHOL CONCENTRATION ON $K_c$ FOR THE *t*-BUTYL ALCOHOL–IODINE COMPLEX IN CYCLOHEXANE

| (ROH) mole/l. |  | $K_c$ |  |
|---|---|---|---|
|  | $\lambda$ = 440 m$\mu$ | $\lambda$ = 520 m$\mu$ | $\lambda$ = 560 m$\mu$ |
| 0.0654 | 1.30 | 1.15 | 0.98 |
| .100 | 1.32 | 1.36 | 1.40 |
| .151 | 1.37 | 1.47 | 1.46 |
| .301 | 1.37 | 1.40 | 1.43 |
| .602 | 1.37 | 1.37 | 1.44 |
| .879 | 1.36 | 1.38 | 1.51 |
| 1.62 | 1.94 | 1.78 | 1.90 |
| 3.16 | 21.4 | 7.2 | 6.6 |

It was also assumed in applying equation (2) to the experimental data that $\epsilon_c$ and $\epsilon_{I_2}$ values did not change with changes in the alcohol concentration of the solvent. It seems likely that this assumption is valid for *t*-butyl alcohol–cyclohexane solutions of alcohol concentration no greater than 1 *M*, since $K_c$ values determined for such solutions were independent of the alcohol concentration of the medium.

Another test of the constancy of $\epsilon_c$ and $\epsilon_{I_2}$ values is based on the consideration of the fact that at some wave length between the absorption maxima of free and complexed iodine the extinction coefficients of the two molecular species should be equal. If, then, $\epsilon_c$ and $\epsilon_{I_2}$ values do not vary with changes in alcohol concentration of the solvent, $\epsilon_a$ values for a series of solutions of varying alcohol concentration should be identical at some fixed wave length at which $\epsilon_c$ and $\epsilon_{I_2}$ are equal. Actually $\epsilon_a$ values for *t*-butyl alcohol–cyclohexane solutions of iodine show a constant value of 668 at 493 m$\mu$ as the alcohol concentrations vary from 0.065–1.62 *M*. In more concentrated alcohol solutions (3–10 *M*) the $\epsilon_a$ values at 493 m$\mu$ decrease with increasing alcohol concentration of the solvent.

It is obvious that equation (2) is applicable in the interpretation of spectrophotometric data on the *t*-butyl alcohol–iodine complex solutions in cyclohexane only when the alcohol concentration of the solution is below 1.62 *M*. In that range of alcohol concentration over which the calculated $K_c$ values remain constant it must also be true that the activity coefficient ratio $\gamma_{ROH \cdot I_2}/\gamma_{ROH}\gamma_{I_2}$ remains constant. The constancy of the activity coefficient ratio may result from a fortunate cancellation of the effects of changes in individual activity coefficients. It seems more likely, however, that Henry's law is applicable for each of the molecular species involved in the equilibrium to form the complex in the more dilute (1 *M* or less) alcohol solutions.

It should be recalled that data for other complexes reported in this investigation were obtained from measurements on solutions in which the alcohol concentration did not exceed 1 *M*. As these data conformed to linear plots when interpreted graphically according to equation (2), it is presumed that the $K_c$ values thus evaluated were not subject to errors resulting from non-ideality of the solutions.

**The Relative Magnitudes of the Equilibrium Constants.**—The equilibrium constants for the formation of *t*-amyl alcohol–iodine and *t*-butyl alcohol–bromine complexes (Table I) are somewhat lower than that calculated from kinetic data for the *t*-amyl alcohol-bromine complex. One would anticipate in view of what is known concerning interactions of halogens with electron donor reagents[5] that the *t*-amyl alcohol–bromine complex would lie in stability between *t*-amyl alcohol–iodine and *t*-butyl alcohol–bromine complexes. However the orders of magnitude of the constants for the three complexes are comparable, a fact which is gratifying in view of the uncertainties inherent in the estimation from kinetic data of the constant for the *t*-amyl alcohol–bromine complex.

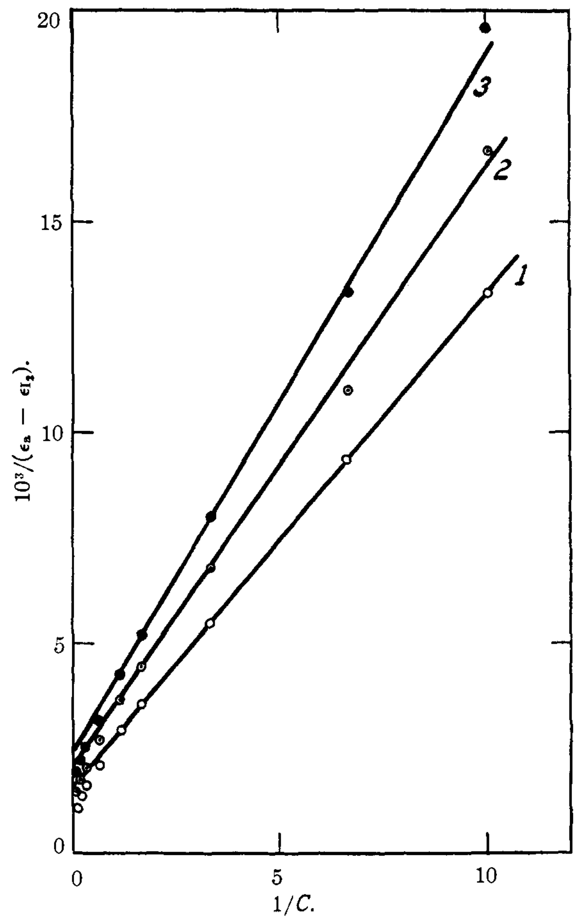

Fig. 1.—The graphical evaluation of $K_c$ and $\epsilon_c$ for the *t*-butyl alcohol–iodine complex in cyclohexane: curves 1, 2, 3 represent data taken at $\lambda$ = 440, 520 and 560 m$\mu$, respectively. For curves 2 and 3 multiply the ordinate values by −1 to obtain the correct readings.

The constant for the *t*-butyl alcohol–iodine complex is approximately the same as that for the interaction of isopropyl alcohol and iodine. The stabilities of these complexes appear not to be as sensitive to structural changes in the organic molecule as are those of the alkyl iodide–iodine complexes. The constant for the *t*-butyl alcohol–iodine complex is higher than that obtained by Ham[2] from measurements at 21.5° in which *n*-heptane was used as a solvent. The extinction coefficients reported by Ham for this complex are somewhat higher than those observed in the present investigation. The experimental details of Ham's investigations are not reported.

The equilibrium constants for the *n*-butyl ether–

iodine complex obtained from measurements in the visible and ultraviolet region correspond closely. This constant based on the expression of the ether in mole fraction rather than molar concentration units is 4.3, a value which is somewhat less than that (4.9) reported[2] for the diethyl ether-iodine complex in *n*-heptane at 21.5°.

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# The Carbon Isotope Effect in the Acid Hydrolysis of Urea

BY JOHN A. SCHMITT AND FARRINGTON DANIELS

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Urea containing approximately 1% each of $C^{13}$ and $C^{14}$ was subjected to acid hydrolysis at 100° in phosphate buffer. The relative rates of decomposition were determined by mass spectrometric analysis of the carbon dioxide evolved at different time intervals during the reaction. The $C^{12}$-urea is hydrolyzed 10.1% faster than $C^{14}$-urea and 5.5% faster than $C^{13}$-urea. The ratio of these values agrees with theory.

The relative rates of enzymatic hydrolysis of urea at 30° and *p*H 5.0 were reported recently.[1] Values of 1.010 ± 0.001 and 1.032 ± 0.002 were obtained for $k_{C^{12}\text{-urea}}/k_{C^{13}\text{-urea}}$ and $k_{C^{12}\text{-urea}}/k_{C^{14}\text{-urea}}$, respectively.

Theoretical considerations[2,3] indicate that in competing isotopic reactions the value of $(k_{C^{12}}/k_{C^{14}} - 1)/(k_{C^{12}}/k_{C^{13}} - 1)$ should be slightly less than 2.0, although the results described for the enzymatic hydrolysis of urea, and those of some other reactions[4,5] have given a value closer to 3.0. Bigeleisen[6] has discussed these discrepancies in a recent review of the isotope effect.

In view of the obvious complication in the reaction caused by the enzyme, it was decided to carry out the hydrolysis in the absence of an enzyme. The reaction is too slow for measurement at 30°, but it is satisfactory at higher temperatures. Urea containing approximately 1% each of $C^{13}$ and $C^{14}$ was hydrolyzed at 100° in the presence of a phosphate buffer which gave a *p*H of 5.0 at 30°. The urea decomposed into ammonia and carbon dioxide and the relative rates of hydrolysis of urea containing $C^{12}$, $C^{13}$ and $C^{14}$ were measured by determining the isotopic ratios with the mass spectrometer, as in the preceding work. The light isotopes reacted faster and the later fractions became enriched in the heavier isotopes.

## Experimental Procedure

The decomposition was carried out in a simple reaction vessel immersed in a bath at 100.0°. $CO_2$-free nitrogen was bubbled through the solution at the rate of 80-100 ml. per minute using a finely fritted glass tube, and the exit gas carrying along the $CO_2$ evolved was passed through a water condenser and then into standardized $NaOH-BaCl_2$ solutions.

Sixty ml. of phosphate buffer, *p*H 5.0 at 30°, was first brought to temperature and the nitrogen bubbled through for 1 hour. Then the urea in buffer was added, using a total of 10 ml. including rinses. In each experiment, 31.0 mg. of $C^{14}$-urea containing 1 mC. activity was diluted with Mallinckrodt A. R. urea to give approximately 1% $C^{14}$ and total weight of urea of 75 mg.

The $C^{14}$-labeled urea was purchased from the Atomic Energy Commission at Oak Ridge. Successive fractions of $CO_2$ produced were collected in the manner described previously.[1] The mass analysis was carried out in the laboratory of Professor R. H. Burris with a mass spectrometer given by the Thomas E. Brittingham Foundation as in the earlier work.

## Results

The carbon isotope ratios of the successive fractions of $CO_2$ collected in the two experiments are given in Table I.

TABLE I

ISOTOPIC RATIOS OF CARBON IN THE CARBON DIOXIDE RELEASED DURING THE COURSE OF THE NON-ENZYMATIC HYDROLYSIS OF UREA AT 100°

| Fraction of reaction | $C^{13}/C^{12}$ | $C^{14}/C^{12}$ |
|---|---|---|
| | Experiment I | |
| 0.000–0.168 | 0.010493 | 0.009809 |
| .168– .335 | .010624 | .011047 |
| .335– .503 | .010766 | .011284 |
| .503– .671 | .010923 | .011551 |
| .671– .837 | .011279 | .012215 |
| | Experiment II | |
| 0.240–0.360 | 0.010667 | 0.011034 |
| .360– .480 | .010762 | .011279 |
| .480– .600 | .010905 | .011515 |
| .600– .720 | .011061 | .011812 |

The simplest way to determine the relative rates of decomposition of the $C^{12}$, $C^{13}$ and $C^{14}$-urea from these data is by the use of an equation derived by Downes[7] which is a modified form of the one used in this earlier work on the enzymatic hydrolysis.[1]

According to this equation

$$W/N_{X^0} = \epsilon(1 - f)^{\epsilon - 1} \quad (1)$$

or

$$\log W = (\epsilon - 1) \log (1 - f) + \epsilon \log N_{X^0} \quad (2)$$

where

$f$ = fraction of total reaction completed
$W$ = $C\ddagger/C^{12}$ of the $CO_2$ being produced at $f$, where $C\ddagger$ is the heavier isotope, $C^{13}$ or $C^{14}$
$N_{X^0}$ = $C\ddagger/C^{12}$ of the total product
$\epsilon$ = $k\ddagger/k$ where $k\ddagger$ and $k$ are the specific reaction rate constants for the hydrolysis of urea for $C^{13}$ or $C^{14}$ and for $C^{12}$, respectively.

Thus, a plot of log $(1 - f)$ *vs.* log $W$ will give a straight line with slope equal to $\epsilon - 1$. As pointed

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(1) J. A. Schmitt, A. L. Myerson and F. Daniels, *J. Phys. Chem.*, **56**, 919 (1952).

(2) J. Bigeleisen, *J. Chem. Phys.*, **17**, 675 (1949).

(3) J. Bigeleisen and T. L. Allen, *ibid.*, **19**, 760 (1951).

(4) W. H. Stevens, J. M. Pepper and M. Lounsburg, *ibid.*, **20**, 192 (1952).

(5) P. E. Yankwich, E. C. Stivers and R. F. Nystrom, *ibid.*, **20**, 344 (1952).

(6) J. Bigeleisen, *J. Phys. Chem.*, **56**, 823 (1952).

(7) A. M. Downes, *Australian J. Scientific Research*, **5**, 521 (1952).