***cis*-Bicyclo[6.1.0]nonan-3-ol.**—By the same procedure used for *cis*-bicyclo[6.1.0]nonane, 2.57 g. of 3-cycloocten-1-ol[46] yielded 1.93 g. of a mixture found by gas chromatography (silicone grease, 145°) to consist of the original 3-cycloocten-1-ol (20%) and *cis*-bicyclo[6.1.0]nonan-3-ols (80%) in addition to a few minor components. Chromatography on alumina (100 g., activity III) removed the minor components (eluted with pentane). The mixture of alcohols (eluted with 4% ether in pentane) was collected in several fractions. The early alcohol fractions were enriched with a mixture of the bicyclic alcohols which were separated from 3-cycloocten-1-ol by gas chromatography (silicone grease, 140°). 3-Cycloocten-1-ol was identified by a comparison of infrared spectra.

The bicyclic alcohols (*endo* and *exo* mixture) could not be separated on silicone grease or TCEP (140°). However, the infrared spectra of the various alcohol fractions were different. A partial separation of the isomeric phenylurethans was accomplished.[47]

Hydrogenation of *cis*-Bicyclo[6.1.0]nonane.—*cis*-Bicyclo[6.1.0]nonane (90 mg.) was hydrogenated at atmospheric pressure in 5 ml. of glacial acetic acid containing 420 mg. of prereduced platinum oxide. No hydrogen uptake was observed after 7 hr. at room temperature. When the temperature was increased to 75°, uptake was 55% of the theoretical amount after 48 hr. The solution was filtered, diluted with 50 ml. of pentane, washed with water and 5% sodium bicarbonate solution, and dried over magnesium sulfate. The product (59 mg.) was shown by gas chromatography (silicone oil, 70°, and TCEP, 70°) to consist of methylcyclooctane (50%) and unchanged *cis*-bicyclo-[6.1.0]nonane (50%), identified by comparison of infrared spectra with authentic samples.

Attempted Hydrogenation of *cis*-Bicyclo[6.1.0]nonan-2-ol.—When *cis*-bicyclo[6.1.0]nonan-2-ol (80 mg.) was hydrogenated as described above, no appreciable uptake was observed at 63° after 2 days. The alcohols were isolated through acetates which were collected by gas chromatography (silicone grease, 140°) and treated with lithium aluminum hydride. The regenerated alcohol (38 mg.) consisted largely of starting material, collected by gas chromatography (TCEP, 140°), and identified by a comparison of retention times and infrared spectra. A peak (less than 7%) corresponding to that of the 2-methylcyclooctanols was also present, but none corresponding to that of 3-methylcyclooctanols or cyclononanol was observed.

---

[46] A. C. Cope and P. E. Peterson, *J. Am. Chem. Soc.* **81**, 1643 (1959).

[47] G. L. Woo, Ph.D. Thesis, M. I. T., 1962.

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY, SAINT LOUIS UNIVERSITY, ST. LOUIS 4, MO.]

Solvents of Low Nucleophilicity. III. The Effect of Remote Substituents in the Addition of Trifluoroacetic Acid to Substituted Alkenes[1,2]

BY PAUL E. PETERSON AND GEORGE ALLEN

RECEIVED MARCH 22, 1963

First-order rates of addition of trifluoroacetic acid to 5-substituted 1-hexenes at 35 and 60° were determined for the substituents chloro, bromo, iodo, methoxy, acetoxy, trifluoroacetoxy, and cyano. Rate decreases by factors of 32.6, 167, and 311 compared to the rate for 1-hexene were observed for the reactions of the acetoxy, trifluoroacetoxy, and cyano substituted compounds, respectively, at 60°. Notable rate decreases were also observed in the reactions of *trans*-4-hexen-2-yl trifluoroacetate, 6-hepten-2-yl trifluoroacetate, and 7-octen-2-yl trifluoroacetate. The possibility that the large rate effects can be attributed at least in part to an enhancement of the inductive effect of substituents which form hydrogen bonds with the trifluoroacetic acid solvent is discussed.

The rates of addition of trifluoroacetic acid to a number of unbranched alkanes and cycloalkenes at 35° have recently been reported.[3a,b] The reactions may be regarded as typical carbonium ion reactions, which, in the case of terminal olefins, involve Markovnikov addition of a proton to the terminal carbon to give a secondary carbonium ion or its equivalent as an intermediate.[4] Accordingly, the addition of trifluoroacetic acid to olefins would be expected to be retarded by electron-attracting substituents near the potential cationic carbon atom, as is the case in other carbonium ion reactions.[5a,b] Since the sensitivity of the rates of carbonium ion reactions to electron-attracting substituents tends to be higher in solvents of low nucleophilicity, a study of such reactions in trifluoroacetic acid with its presumed very low nucleophilicity is of particular interest. However, the work described in this paper was prompted originally not by such considerations but by the chance observation that 5-hexen-2-yl trifluoroacetate (**1g**, Fig. 1) was surprisingly unreactive toward trifluoroacetic acid. As may be seen from inspection of Fig. 1, the trifluoroacetoxy group in the presumed intermediate cation is on the fourth carbon from the positively charged carbon (counting the positive carbon as number one since it could conceivably have an attached substituent other than hydrogen).

Since it appeared that the finding of a large rate effect due to the presence of a functional group at a large distance from the reaction center would be of unusual interest, a study of the first-order rates of addition of trifluoroacetic acid to the trifluoroacetate **1g** and to other 5-substituted 1-hexenes at 35 and 60° was undertaken. Rates of reaction of *trans*-4-hexen-2-yl trifluoroacetate, 6-hepten-2-yl trifluoroacetate, 7-octen-2-yl trifluoroacetate, and of several unsubstituted alkenes were also determined. The results, given in Table I, confirmed the remarkably large rate retardation due to the presence of the trifluoroacetoxy group. The cyano and acetoxy groups were also found to cause large rate decreases.

Product Studies.—The products of the reaction of trifluoroacetic acid with 5-hexen-2-yl trifluoroacetate were shown to be those indicated in Fig. 2. On the basis of this and other studies, some of which are reported in the Experimental section, and also as a consequence of arguments based on the observed reaction rates which rule out various possible rearrangement reactions, it may be concluded that the main reaction observed for the substituted alkenes of Table I is in fact the addition of trifluoroacetic acid to the unrearranged double bond. The role of the rearrange-

---

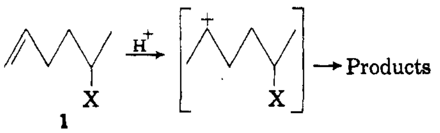
[1] (a) Presented in part at the 141st National Meeting of the American Chemical Society, Washington,D. C., March, 1962; (b) for a preliminary communication see P. E. Peterson and G. Allen, *J. Org. Chem.*, **27**, 2290 (1962).

[2] This paper is based on the Ph.D. research of G. Allen. The planning of the work was aided greatly by the opportunity which was afforded P. E. Peterson to spend the summer, 1961, in Professor R. W. Taft's research group at Pennsylvania State University where studies of inductive effects upon fluorine n.m.r. spectra were in progress which could be correlated with the work reported here. We are indebted to the donors of the Petroleum Research Fund of the American Chemical Society for a grant (PRF-790a) which made this phase of the work possible. G. Allen gratefully acknowledges a research assistantship provided by the Allied Chemical Corporation during the summer, 1962.

[3] (a) P. E. Peterson and G. Allen, *J. Org. Chem.*, **27**, 1505 (1962); (b) P. E. Peterson, *J. Am. Chem. Soc.*, **82**, 5834 (1960).

[4] The postulation of an intermediate as opposed to a positive transition state is primarily a matter of convenience. It may be noted, however, that the differences between addition and solvolysis reactions in trifluoroacetic acid were conveniently rationalized in terms of solvated cationic intermediates (*cf.* ref. 3b).

[5] (a) A. Streitwieser, Jr., *Chem. Rev.*, **56**, 571 (1956); (b) A. Streitwieser, *J. Am. Chem. Soc.*, **78**, 4935 (1956).



1

a, X = H
b, X = Cl
c, X = Br
d, X = I
e, X = OCH_3
f, X = O_2CCH_3
g, X = O_2CCF_3
h, X = CN

Fig. 1.—Formation of the cationic intermediate in the addition of trifluoroacetic acid to 5-substituted 1-hexenes.

ment reaction, which occurs to a smaller extent, is discussed in the next section. The possible occurrence in certain instances of products due to neighboring group participation is under investigation in our laboratory in the case of various terminally substituted alkenes.

TABLE I

RATES OF ADDITION OF TRIFLUOROACETIC ACID TO SUBSTITUTED ALKENES AND DERIVED QUANTITIES

Substituent	$k \times 10^6$, sec.$^{-1}$, 35°	$k \times 10^6$, sec.$^{-1}$, 60°	$\Delta H^{\ddagger}$, kcal./mole	$-\Delta S^{\ddagger}$, e.u.
H[a]	235	1470	14.3	29
Cl[a]	60.3	356	13.9	33
Br[a]	73.3	439	14.0	32
I[a]	72.2	420	13.8	33
OCH_3[a]	63.0	502	16.3	25
O_2CCH_3[a]	5.60	45.1	16.4	29
O_2CCF_3[a]	1.12	8.79	15.5	35
CN[a]	0.39	4.72	19.7	24
H[b]	275	1540	13.4	32
O_2CCF_3[c]	6.42	36.4	13.6	38
H[d]	287	1680	13.8	30
O_2CCF_3[e]	26.3	200	15.9	28
O_2CCF_3[f]	. .	1.50		
$(CH_3)_2$[g]	84.4	. . .		
Cl[h]	. .	217		

[a] Compound = 5-substituted 1-hexene. [b] Compound = 1-heptene. [c] Compound = 6-hepten-2-yl trifluoroacetate. [d] Compound = 1-octene. [e] Compound = 7-octen-2-yl trifluoroacetate. [f] Compound = *trans*-4-hexen-2-yl trifluoroacetate. [g] Compound = 4,4-dimethyl-1-pentene. [h] Compound = 5-chloro-1-pentene.

Evaluation of Rate Constants.—Although plots of the natural logarithm of the total alkene concentration *vs.* time gave lines showing only a slight curvature through approximately 50% reaction, an interpretation of the slope in terms of the values of k_1, k_2, and k_3 (Fig. 2) was desirable. Accordingly the integrated rate equations for reaction according to the scheme of Fig. 2 were obtained. A computer was used to prepare tables of the logarithm of the alkene concentration as a function of time for the range of k_2/k_1 and k_3/k_1 ratios which were possible in our study, as judged from product studies and other qualitative arguments. Inspection of the tables showed that the slopes in a plot of the natural logarithm of the total alkene concentration *vs.* time did in fact have values within 7% or less of k_1 in all instances.[6] The procedure adopted for obtaining the rate constants, illustrated by an example in the Experimental section, depends upon applying a correction to the slopes of the experimentally determined plots in order to obtain k_1 values. For most of the substituted hexenes k_3/k_1 ratios were assumed to be equal to the measured ratio for the trifluoroacetoxy compounds **1g** and *trans*-**3**. The k_2/k_1 ratios could be estimated by comparing the amount of alkene present after approximately seven half-lives with the tabulated values for various k_2/k_1 ratios.

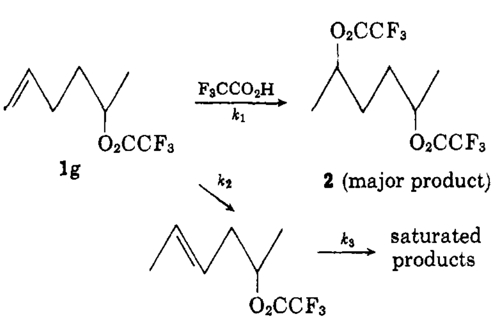
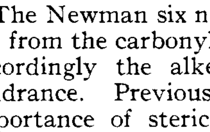


Fig. 2.—Products from reaction of trifluoroacetic acid with 5-hexen-2-yl trifluoroacetate.

For the unsubstituted alkenes, inspection of the tables indicated that a fortunate cancellation of effects occurs which results in plots of the natural logarithm of the alkene concentration *vs.* time which have almost no curvature and whose slopes are almost exactly equal to k_1. Accordingly, the rate constants reported previously for unsubstituted acyclic alkenes[3a] are k_1 values, and may be compared with the values in Table I which are also k_1 values.

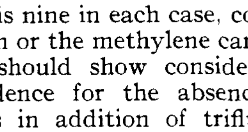
In view of the fact that nearly the same percentage (15%) of rearranged alkene was present after approximately seven half-lives for all substituted hexenes except two (these exceptions may possibly indicate the involvement of anchimeric assistance), it seems reasonably likely that in essentially all cases approximately the same amount (one-third) of the total reaction products formed consists of rearranged olefin. Accordingly, in terms of the rate constants: $k_2/(k_1 + k_2) = {}^1/_3$ or $k_2/k_1 = 0.5$. Although the conclusions regarding these magnitudes of the k_2/k_1 and k_3/k_1 ratios are at best tentative, the k_1 values obtained by our methods are fortunately only slightly dependent on the magnitudes of these ratios within the range of possible values for the ratios. From the standpoint of the following discussion, the main point to be made regarding the accuracy of the k_1 values is that the possible errors in k_1 are of negligible magnitude in comparison with the large variations in the magnitude of k_1.

Influence of Steric Effects upon the Rates.—As a first step in interpreting the rates in Table I, it may be noted that the rate of reaction of 4,4-dimethyl-1-pentene (**4**) was retarded by a factor of only 2.8 at 35° compared to the rate for 1-hexene. This observation is regarded as the best evidence yet obtained that the rate of addition of trifluoroacetic acid to alkenes is relatively insensitive to steric hindrance, since the double bond is located in a position comparable to that of the carbon–oxygen double bond in 3,3-dimethylbutanoic acid (**5**) whose rate of esterification at 40° was found to be retarded by a factor of 42.7 presumably because of steric hindrance.[7]



$$CH_3-\overset{CH_3}{\underset{CH_3}{C}}-CH_2-\overset{CH_2}{\underset{H}{C}}$$

4



$$CH_3-\overset{CH_3}{\underset{CH_3}{C}}-CH_2-\overset{O}{\underset{OH}{C}}$$

5

The Newman six number is nine in each case, counting from the carbonyl oxygen or the methylene carbon; accordingly the alkene **4** should show considerable hindrance. Previously evidence for the absence of importance of steric effects in addition of trifluoroacetic acid to cyclooctene was discussed.[3a] Also it has been concluded that the transition state for hydration of branched cyclic olefins resembles the olefin in geom-

---

(6) The slopes were also tabulated by the computer. The slope was calculated based on the initial total alkene concentration and on the concentration at a time when the alkene content had dropped to about three-quarters of its initial value. At shorter times the slopes approached k_1 more closely.

(7) M. S. Newman in "Steric Effects in Organic Chemistry," M. S. Newman, Ed., John Wiley and Sons, Inc., New York, N. Y., 1956, p. 205.

etry[8]; accordingly, steric interactions in the transition state may be similar to those in the starting material. Finally, examination of models indicated that the alkenes **1b**–**1h** possessed a number of conformations in which the double bond was not hindered, in addition to conformations in which steric hindrance was present. Steric hindrance therefore appears unlikely to be the dominant factor governing the rates reported in Table I, although whether or not it contributes measurably to the over-all rate pattern cannot be judged at present.

Influence of Inductive Effects, Hydrogen Bonding, and Anchimeric Assistance.—The pattern of our rate effects may be seen by examination of Table II where k_H/k_X values at 35 and 60° are given, k_H and k_X being k_1 values for unsubstituted and substituted alkenes, respectively, which have carbon chains of the same length. Values of log k_H/k_X are given for comparison with Taft's σ_I values[9] or with the "σ_I values in trifluoroacetic acid," to be discussed later. Previously reported rate constants for 5-substituted 2-bromopentane formolysis[10] are shown to illustrate the striking contrast between the observed substituent effects and our own results.

Although our results initially may appear puzzling and perhaps surprising, as they were to us, we are now able to present an interpretation of our rates in terms of two hypotheses: (1) The large rate decelerations associated with the presence of acetoxy, trifluoroacetoxy, and cyano substituents are the result of enhanced inductive effects of these substituents which in turn arise as a consequence of strong hydrogen bonding between trifluoroacetic acid molecules and the substituents. (2) The relatively small rate effects exhibited by the chloro, bromo, iodo, and methoxy substituents are the result of anchimeric assistance which partially masks the inductive influence of these substituents.

The hypothesis concerning strong hydrogen bonding is fairly well substantiated. Trifluoroacetic acid has been reported to form remarkably stable complexes or maximum boiling mixtures with ethers, ketones, esters, and nitriles[11] which, although not investigated in detail, are most probably hydrogen-bonded complexes. The inference that strong solvent interactions of this type lead to a very appreciable enhancement of the inductive effect when the hydrogen bonded group serves as a substituent has been made by Taft and coworkers on the basis of n.m.r. studies.[12] In these recent, extensive studies the fluorine resonance frequencies of over 30 *m*-substituted fluorobenzenes dissolved in aqueous methanol compared to the frequency of fluorobenzene itself were found to give a remarkably precise linear correlation with σ_I values. The chemical shifts of over 20 substituted fluorobenzenes were studied in trifluoroacetic acid. The modified shifts observed in trifluoroacetic acid were converted to a new set of σ_I values,[12] some of which are listed in Table II. Oxygen- and nitrogen-containing substituents invariably show considerably increased σ_I values in trifluoroacetic acid as measured by the n.m.r. method, although whether these modified values will prove qualitatively or quantitatively significant for reactivity studies in trifluoroacetic acid can only be answered by experiments. Calculated estimates presented in the next section provide some support for our postulate that enhanced substituent effects are in fact involved in our study, as do other preliminary results obtained in our laboratory. Published estimates of the acidity of trifluoroacetic acid (H_0 = −4.4 and −3.03),[13] and the reported pK_a values of aliphatic ketones (−7.1 to −7.2)[14a] and ethers (−3.83 to −6.2)[14b,c] suggest that protonated substituents are not involved in our study.

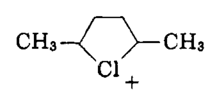
TABLE II

COMPARISON OF RELATIVE RATES WITH VARIOUS QUANTITIES WHICH ARE MEASURES OF INDUCTIVE EFFECTS

Substituent	k_H/k_X (35°)	k_H/k_X (60°)	log k_H/k_X (60°)	$k \times 10^5$, sec.$^{-1}$, for bromopentane formolysis[a]	σ_I[b]	σ_I in F_3CCO_2H from n.m.r. data
H[e]	1.00[j]	1.00	0.0	6.50	0.0	0.0
Cl[e]	3.90	4.13	.62	..	.47	..
Br[e]	3.21	3.35	.53	6.3	.45	0.44[c]
I[e]	3.26	3.50	.54	..	.38	..
OCH_3[e]	3.73	2.93	.47	189	.25	0.51[c]
O_2CCH_3[e]	42.0	32.6	1.51	6.4	.39	.60[d]
O_2CCF_3[e]	192	167	2.22	..	.67[d]	.72[d]
CN[e]	602	311	2.49	1.4	.56	.74[c]
H[f]	1.00[j]	1.00	..	..	..	..
O_2CCF_3[g]	42.8	42.3	1.63	..	..	..
H[h]	1.00[j]	1.00	..	..	..	..
O_2CCF_3[i]	10.9	8.40	0.92	..	..	..

[a] From ref. 10. [b] *Cf.* ref. 9a. [c] From ref. 12. [d] From ref. 1b. [e] Compound = 5-substituted 1-hexene. [f] Compound = 1-heptene. [g] Compound = 6-hepten-2-yl trifluoroacetate. [h] Compound = 1-octene. [i] Compound = 7-octen-2-yl trifluoroacetate. [j] k_H/k_X ratios refer to compounds having the same carbon chain; *e.g.*, k_H for heptene is used in calculating k_H/k_X for 6-hepten-2-yl trifluoroacetate.

Anchimeric assistance, which could lead to cyclic intermediates of the type **6**, was at first thought not to be involved in the case of our halogen substituents because of the essentially identical rate effects observed for chloro, bromo, and iodo substituents. This inference was based on the accepted view that in the acetolysis of *trans*-2-halocyclohexyl brosylates, iodine is more effective than bromine as a participating group by a factor greater than 1000 (the exact magnitude depends on the interpretation adopted) while bromine is more effective than chlorine by a factor somewhat less than 1000.[5a] If in our own study the various halogens participate to an equal extent, as postulated, the contrast with the previous findings can be attributed to any combination of the following circumstances: (a) The relative abilities of *neighboring* halogens to participate with concomitant formation of three-membered rings may not be paralleled in reactions which would involve formation of larger rings, including the five-membered rings postulated to be formed in our study.[15] (b)



6

---

[8] P. Riesz, R. W. Taft, Jr., and R. H. Boyd, *J. Am. Chem. Soc.*, **79**, 3724 (1957).

[9] (a) *Cf.* R. W. Taft, Jr., ref. 7, p. 556; (b) in this paper σ_I values are taken from ref. 12 or calculated from the relationship $\sigma_I = 0.45\sigma^*$. In the present paper the term "inductive effect" is defined as the entire effect arising from the electron attracting or releasing properties of the substituents, although other definitions have been used. Recent references to inductive effects in carbonium ion reactions are: (c) H. Kwart and L. J. Miller, *J. Am. Chem. Soc.*, **83**, 4552 (1961); (d) D. S. Noyce, B. R. Thomas, and B. N. Bastian, *ibid.*, **82**, 885 (1960); (e) D. S. Noyce and R. S. Monson, *Tetrahedron Letters*, **19**, 863 (1962).

[10] S. Oae, *J. Am. Chem. Soc.*, **78**, 4032 (1956).

[11] (a) M. Hauptschein and A. V. Grosse, *ibid.*, **73**, 5131 (1951); (b) J. Lichtinberger, *Bull. soc. chim. France*, 687 (1954); (c) W. D. Emmons, *J. Am. Chem. Soc.*, **76**, 3470 (1954); (d) W. Durrel, J. Young, and R. Dresdner, *J. Org. Chem.*, **28**, 831 (1963).

[12] R. W. Taft, E. Price, I. R. Fox, I. C. Lewis, K. K. Anderson, and G. T. Davis, *J. Am. Chem. Soc.*, **85**, 709 (1963).

[13] (a) F. A. Bovey and G. V. D. Tiers, *ibid.*, **81**, 2870 (1959); (b) H. H. Hyman and R. A. Garber, *ibid.*, **81**, 1847 (1959).

[14] (a) H. J. Campbell and J. T. Edward, *Can. J. Chem.*, **38**, 2109 (1960); (b) E. M. Arnett, C. Y. Wu, J. N. Anderson, and R. D. Bushick, *J. Am. Chem. Soc.*, **84**, 1674 (1962); (c) J. T. Edward, *Chem. Ind.* (London), 489 (1963).

Halogen participation in olefin addition reactions may not parallel participation in solvolyses. (c) The extent of participation by the various halogens may not be the same for reactions in trifluoroacetic acid, as compared to that observed in acetic acid.

Preliminary studies of possible halogen and methoxyl participation in addition of trifluoroacetic acid to terminally substituted alkenes strongly support our position. For example, rate effects due to nonparticipating substituents fall off smoothly with distance (*cf.* ref. 19), but 5-methoxy 1-pentene and 5-bromo-1-pentene react at approximately the same rate as the corresponding 6-substituted 1-hexenes,[16] suggesting that participation partially counteracts inductive rate depressions in the case of the pentenes.

Magnitude of Rate Effects Predicted by the Hammett–Taft Treatment.—The comparison of our rate effects with those found in the bromopentane formolyses[10] suffers from the fact that the formolyses were carried out in a solvent (80% formic acid–20% water) whose nucleophilic character may have influenced the reaction rates, as well as from the fact that bromine is not a very favorable leaving group for the development of carbonium ion character in a reaction. Also, possible limitations involved in comparing our larger rate effects with the effects due to halogen substituents have been indicated.[17] A third approach to the question of whether the rate effects which we observed are in fact larger than might have been expected based on available information involves the comparison of our results with estimates of inductive effects based on the Hammett–Taft relationship, $\log k_X/k_H = \rho^*\sigma^*$. For carbonium ion reactions involving neighboring substituents $-\rho^*$ is usually not appreciably greater than 4,[9a] while the largest common σ^* value is perhaps that of the cyano group, 1.3. The inductive effect is frequently assumed to decrease by a factor of $^1/_2$ to $^1/_{2.8}$ per additional methylene group interposed between the substituent and the reaction center.[12] Using the larger factor (0.5) in conjunction with the other parameters given, the values of k_H/k_X given in Table III were calculated for the reaction of compounds having a cyano substituent (or another substituent having a σ^* value of 1.3) on the fourth, fifth, and sixth carbons, respectively, from the potential carbonium ion center. (These compounds have 2, 3, and 4 interposed methylene groups, respectively.) Since the values of k_H/k_X determined in our study (*cf.* Table II) indicate that a trifluoroacetoxy group and a cyano group exert quantitatively similar effects, the calculated values may be compared with the observed values of k_H/k_X for 5-hexen-2-yl trifluoroacetate, 6-hepten-2-yl trifluoroacetate, and 7-octen-2-yl trifluoroacetate, respectively, which have substituents at the positions assumed in the calculations.

TABLE III

CALCULATED AND OBSERVED SUBSTITUENT EFFECTS

n (number of interposed CH_2 groups)	$-\rho^*\sigma^*$ (calcd.)	k_H/k_X (calcd.)	$-\rho^*\sigma^* =$ $\log k_H/k_X$ (obsd.)	k_H/k_X (obsd.)
2	1.3	19.9	2.22	167
3	0.65	4.47	1.63	42.3
4	.325	2.11	0.92	8.40

The calculated values of $-\rho^*\sigma^*$ are found to be smaller than the values based on the observed rate data by factors of 1.7 to 2.8. Use of the trifluoroacetic acid-enhanced σ_I values estimated from nuclear magnetic resonance data would lead to a calculated $\rho^*\sigma^*$ value not far from the observed value in the case of the 5-substituted 1-hexene, but in the case of the alkenes having more distant substituents the observed rates still lead to $-\rho^*\sigma^*$ values larger than those calculated. Accordingly, the factor by which the inductive effect changes per additional interposed methylene group is larger than 0.5, the largest value commonly used in the literature.[18,19]

In addition to the calculated k_H/k_X values in Table III the value for halogen (10, based on $\sigma^* \cong 1$ and $n = 2$) is of interest. Comparison with the observed values (which vary from 4.13 to 3.35) shows that $-\rho^*\sigma^*$ is overestimated in this calculation in contrast to the underestimation for the nitrile substituent. This result, which is another manifestation of the large differences between the effects of the halogens and the effects of the other substituents, lends further support to our postulate of halogen participation. It may be concluded that although the magnitudes of the rate effects observed in our study appear to be remarkably large from some standpoints, the observed effects are not greatly in excess of the largest effects of inductive origin which are predicted by use of a Hammett–Taft approach provided allowance is made for the sizable increases of σ-values in trifluoroacetic acid and for the surprisingly small decrease of the inductive effect per additional methylene group.

Comparison with Observations in Other Systems.—The demonstration of large rate effects due to remote substituents for carbonium ion reactions in trifluoroacetic acid provides a basis for interpreting some observations in the literature. For example, although trifluoroacetic acid adds to cyclooctene with a half-life of only 2.8 min. at 35°, 3-cycloocten-1-ol and 4-cycloocten-1-ol (or the corresponding trifluoroacetates) were isolated from the reaction of *cis*-cyclooctene oxide with trifluoroacetic acid at 0–5° for 3 hr.[20] Also, the trifluoroacetates of 3-cycloocten-1-ol and *trans*-2-vinylcyclohexanol survived 70 min. of stirring in trifluoroacetic acid at 30–34° as shown by their isolation in 60.5% combined yield (as the corresponding alcohols) from the solvolysis of 3-cycloocten-1-yl *p*-bromobenzenesulfonate.[21] It is now apparent that the substituted alkene products obtained in the studies just mentioned survived the reaction conditions, at least in the second instance, only because the rate of addition of trifluoroacetic acid was decreased greatly by the inductive effect of the substituent. The inductive effect of the brosylate group would similarly protect the double bond in the starting material in the case of a brosylate which does not undergo rapid solvolysis. As an illustration of the magnitude of the rate effects which may be exerted by remote substituents in the cyclooctyl system, we have shown in a preliminary determination that the rate of addition of trifluoroacetic acid to 4-cycloocten-1-yl acetate at 35° was decreased by a factor of 330, compared to the rate for cyclooctene.

---

[15] Unpublished studies by S. Winstein and R. E. Glick, involving solvolysis of halogen substituted primary tosylates indicate that the driving force for formation of 5-membered halonium rings may not be large. Evidence from special salt effects and rates, however, suggests that the familiar order of participating ability, I > Br > Cl, was observed. We are indebted to Prof. Winstein for calling our attention to this work, reported in the Ph.D. thesis of R. E. Glick, University of California at Los Angeles, 1954.

[16] P. E. Peterson, E. Tao, and A. Agtarap, unpublished work.

[17] Such comparisons provided the basis of our earlier interpretation of our rates (*cf.* ref. 1b).

[18] J. C. McGowan, *J. Appl. Chem.*, **10**, 312 (1960).

[19] Other studies in our laboratory indicated that a factor as large as 0.66 is sometimes required; see P. E. Peterson, C. Casey, and G. Allen, Abstracts, 144th National Meeting of the American Chemical Society, Los Angeles, Calif., April, 1963, p. 57M.

[20] A. C. Cope, J. M. Grisar, and P. E. Peterson, *J. Am. Chem. Soc.*, **81**, 1640 (1959).

[21] A. C. Cope, J. M. Grisar, and P. E. Peterson, *ibid.*, **82**, 4299 (1960).

Although our study was concerned with aliphatic compounds with the exception of 4-cycloocten-1-yl acetate, the results may be compared with certain observations in the literature of rate effects due to remote substituents in cyclic and bicyclic systems. Our effects presumably are not related to the long range rate effects in steroidal compounds which have been ascribed to conformational transmission of angle (or other) strain.[22] However, recent communications have reported the importance of long range inductive effects in determining the rates of bromination of unsaturated steroids[23a,b] although the magnitudes of the effects are not necessarily greater than those predicted using the Hammett–Taft approach.[23c]

Conclusion.—The discovery of the slow fall-off of the inductive effect with distance in our system and of the apparent ability of trifluoroacetic acid to increase the inductive effect of various common substituents opens the way for the study of molecules where the effect would until now, have been considered too small to be of importance. The use of trifluoroacetic acid as a solvent for the purpose of influencing the relative rates of competing reactions in the case of compounds having several inductive groups is suggested. Virtually all natural products are compounds which should show interesting reactivity effects, including various steroids, carbohydrates, and alkaloids.[24] From the standpoint of the theoretical chemist the possibility suggested by our study of isolating the inductive effect from steric and other effects by the use of truly remote substituents is particularly attractive.

Experimental

The techniques used in measuring rates have been described previously.[3a] Compounds employed were purchased or prepared by standard methods.[25]

Identification of the Products from the Reaction of Trifluoroacetic Acid with 5-Hexen-2-yl Trifluoroacetate. (a) Addition Product.—5-Hexen-2-yl-trifluoroacetate (4.9 g., 0.025 mole) was dissolved in trifluoroacetic acid, 0.125 *M* in sodium trifluoroacetate, to give a solution 0.25 *M* in ester. This solution was kept at 35° for 49 days. This time constitutes 7.5 half-lives for a 0.1 *M* solution; for the solution used, a slightly greater half-life would be expected.

One hundred ml. of water and 50 ml. of ether were then added. Sodium carbonate was added in portions until the evolution of carbon dioxide ceased. Ether extraction followed by distillation resulted in 1.7 g., b.p. 93–95° (15 mm.), of product.

A sample of the distilled product was gas chromatographed on a silicone oil column (Dow Corning 550) at 141° along with a known sample of 2,5-hexanediol bistrifluoroacetate. Both samples had the same retention time and both split into two peaks which were attributed to the two diastereomers.

One gram of the products was added to 10 ml. of 5 *M* sodium hydroxide, and the mixture was shaken mechanically for 15 min. It was then allowed to stand for 15 min. at room temperature. The water was removed, and 25 ml. of ether was added to the resulting viscous solution. The ether was decanted from the solid (presumably sodium hydroxide and sodium trifluoroacetate). Removal of the ether left a viscous material which was gas chromatographed on a silicone oil column (Dow Corning 550) at 156°. The material had the same retention time as an authentic sample of 2,5-hexanediol.

The hydrolyzed product was oxidized to the diketone by the improved method reported by Brown and Garg.[26] The ether extract contained material having a gas chromatographic retention time (Dow Corning 550 silicone oil) identical with that of acetonylacetone. A bissemicarbazone of the material was too insoluble in a variety of solvents to allow recrystallization, but the melting point (222° dec.) was in agreement with that reported for the bissemicarbazone of acetonylacetone (220°),[27] and a mixture melting point showed no depression.

(b) Isomerization Product.—5-Hexen-2-yl trifluoroacetate (50 g.) was refluxed with trifluoroacetic acid (450 g.) and sodium trifluoroacetate (1 g.) for 20 hr. An isolation procedure similar to that described for the addition product gave an early distillation fraction having a lower boiling point (62.0–62.5°, 59 mm.) than the addition product. Gas chromatography (Dow Corning 550 silicone oil) gave two peaks attributable to *trans*-4-hexen-2-yl trifluoroacetate and to *cis*-4-hexen-2-yl trifluoroacetate and/or 5-hexen-2-yl trifluoroacetate.

Lithium aluminum hydride reduction of the trifluoroacetate mixture gave an alcohol mixture which was resolved by gas chromatography on a 20-ft. 20% silver nitrate–Carbowax 200 column at 110° into three peaks attributable to *trans*-4-hexen-2-ol (58%), *cis*-4-hexen-2-ol (18%), and 5-hexen-2-ol (24%). The alcohol component thought to be *trans*-4-hexen-2-ol was collected by gas chromatography on a neopentyl glycol succinate column and identified as the expected compound by comparison of its infrared spectrum with that of an authentic sample.

The Addition Product from the Reaction of Trifluoroacetic Acid with 5-Chloro-1-pentene.—5-Chloro-1-pentene (1.32 g., 0.013 mole) and some sodium trifluoroacetate were made up to 50 ml. with trifluoroacetic acid and maintained at 60° for 5.2 hr. An isolation procedure similar to that used in the other product studies yielded 0.46 g. of material, b.p. 62° (10 mm.), whose infrared spectrum showed the bands expected for the spectrum of 5-chloro-2-pentyl trifluoroacetate. Lithium aluminum hydride reduction, followed by oxidation[26] gave 5-chloro-2-pentanone, identified as the 2,4-dinitrophenylhydrazone, m.p. 125.5–126.0°, undepressed upon admixture with an authentic sample prepared from 5-chloro-2-pentanone obtained from Aldrich Chemical Co.; reported[28] for 5-chloro-2-pentanone dinitrophenylhydrazone, m.p. 126.5–127.5°.

Rate Determinations.—Rates of addition of trifluoroacetic acid were determined by the iodine monobromide titrimetric method used previously.[3a] Again trifluoroacetic acid solutions 0.125 *M* in sodium trifluoroacetate and 0.1 *M* in alkene were employed for reasons mentioned previously.[3a]

The approach taken in evaluating k_1 values (*cf*. Fig. 2) has been outlined in the Discussion section. Additional details and an example are given here. If the concentration of substituted alkene (*e.g.*, **1g**, Fig. 1) is designated as *A* and the concentration of rearranged alkene (*e.g.*, **3**, Fig. 2) is designated as *B*, the reaction scheme of Fig. 1 leads to the rate expressions

$$\mathrm{d}A/\mathrm{d}t = -(k_1 + k_2)A \tag{1}$$

$$\mathrm{d}B/\mathrm{d}t = k_2A - k_3B \tag{2}$$

Integrating 1 gives 3.

$$A = A_0e^{-(k_1+k_2)t} \tag{3}$$

where A_0 is the value of *A* when the time $t = 0$. Substitution of *A* from 3 into 2 and integrating gives 4.

$$B = k_2A_0/k_3 - k_1 - k_2[e^{-(k_1+k_2)t} - e^{-k_3t}] \tag{4}$$

Adding 3 to 4 gives an expression for $A + B$ as a function of time. Since *A* and *B* are both alkenes, it is $A + B$ which is determined by the iodine monobromide titrimetric method.[29]

As has been mentioned in the Discussion section a plot of the natural logarithm of $A + B$ *vs*. time is a curve.[30] However, a line which passes through the point at $t = 0$ and another point taken at time $= t$ has a slope, *m*, which approaches k_1 as *t* approaches zero. This trend is illustrated in Table IV which gives selected values of *m* as a function of time when $k_1 = 1$ and the other constants k_2 and k_3 have the values shown. For such a case a "half-life" may be defined as $0.693/m$.

The use of the tables may now be illustrated: As a specific example we may assume that a certain substituted alkene shows 7% of the initial double bond content after the addition of tri-

---

[22] D. H. R. Barton, F. McCapra, P. J. May, and F. Thudium, *J. Chem. Soc.*, 1297 (1960), and previous papers.

[23] (a) V. Schwarz, S. Hermanek, and J. Trojanek, *Chem. Ind.* (London), 1212 (1960); (b) V. Schwarz and S. Hermanek, *Tetrahedron Letters*, 809 (1962); (c) P. E. Peterson, *ibid.*, 181 (1963).

[24] It would be interesting to know whether the catalytic action of enzymes is sometimes in part a consequence of important inductive interactions. Our study suggests that such effects could be appreciable even when the inductive groups are at some distance from the site of reaction. In particular, the effect of a charged group might be transmitted for a considerable distance through a protein "cavity" of low dielectric constant.

[25] All new compounds gave satisfactory analyses for carbon and hydrogen. Details of synthesis are available in the Ph.D. thesis of G. Allen, which has been accepted by St. Louis University and will be available from University Microfilms, Ann Arbor, Mich. An experimental section reporting syntheses was made available to the referees of this paper prior to possible publication elsewhere.

[26] H. C. Brown and C. P. Garg, *J. Am. Chem. Soc.*, **83**, 2952 (1961).

[27] R. L. Shriner, R. Fuson, and D. Curtin, "The Systematic Identification of Organic Compounds," John Wiley and Sons, Inc., New York, N. Y., 1959, p. 218.

[28] J. D. Cawley, British Patent 728,446 (1950); *Chem. Abstr.*, **50**, 5722 (1956).

[29] Equations similar to or identical with those above have been used in connection with problems involving radioactive decay. In the present instance we are indebted to Mr. Charles Casey, N. S. F. undergraduate research participant, for independently deriving and integrating the equations.

[30] The curvature was sometimes not noticeable in actual plots of experimental data until after 30 to 50% reaction.

TABLE IV

QUANTITIES COMPUTED FROM INTEGRATED RATE EQUATION USING ASSUMED k-VALUES

Time	$A + B$	m	$A + B$ at 7 half-lives	Time	$A + B$	m	$A + B$ at 7 half-lives
First group: $k_2 = 0.2$, $k_3 = 0.2$				Third group: $k_2 = 0.5$, $k_3 = 0.167$			
0.0	0.1000		0.00781[a]	0.0	0.1000		0.0167[a]
.3	.0746	0.974	.00759	.2	.0826	0.957	.0161
.6	.0567	.946	.00734	.4	.0694	.914	.0155
.9	.0439	.915	.00706	.6	.0593	.870	.0148
Second group: $k_2 = 0.6$, $k_3 = 0.4$							
0.0	0.1000		0.00720[a]				
.3	.0753	0.946	.00644				
.6	.0585	.894	.00571				
.9	.0467	.845	.00504				

[a] Based on $m = 1$

fluoroacetic acid has been allowed to proceed for seven "half-lives" based on an early value of m. The first two groups of values in Table IV approximately correspond to the situation observed experimentally since values of $A + B$ at seven "half-lives" based on m when $t = 0.3$ are 7.59 and 6.44% of the initial alkene content, respectively. If other evidence leads us to use a k_3 value of 0.2 ($k_3/k_1 = 0.2$), the tables show that the k_2 value which fits the experiment is also approximately 0.2. The slope at $t = 0.3$ is 0.974. Multiplying the m-value by 1000/974 gives k_1. Accordingly, the experimentally determined slope is corrected by this amount. If we incorrectly assign $k_3 = 0.4$, the slope is 0.946 and a correction factor of 1000/946 would be applied to the experimentally determined slope. Accordingly, the tables lead us to the satisfying conclusion that the corrected slopes have values which are not far from k_1 even if the estimated k_3/k_1 ratios are grossly inaccurate.

The third group of values with $k_2 = 0.5$ and $k_3 = 0.167$ illustrates the behavior of m when approximately 16% of the initial alkene is present after seven half-lives and $k_3 = 0.167$. It may be seen that multiplying m by 1000/957 gives k_1 at $t = 0.2$ and that multiplying m by 1000/914 gives k_1 at $t = 0.4$. Since 16% was approximately the maximum amount of alkene found at seven half-lives, this portion of Table IV illustrates the maximum corrections used in obtaining k_1 values. All of the substituted hexenes (Fig. 1) except the cyano- and methoxyhexenes showed approximately 16% double bond content after seven "half lives." In practice the tables for $k_1 = 1$ may be used for any rate determination since only the k_3/k_1 and k_2/k_1 ratios affect the rate plots for reactions followed for a specific portion of a half-life, as judged from the amount by which the alkene concentration ($A + B$) decreases during the time the reaction is followed.

Isomerization Products from Reaction of 1-Hexene with Trifluoroacetic Acid.—1-Hexene was allowed to react with a pure sample of trifluoroacetic acid, obtained by hydrolysis of the anhydride, for exactly $^1/_6$ of a half-life at 35°. A sample of the reaction mixture was immediately analyzed by gas chromatography on a 20-ft. tricresyl phosphate column. The amounts of *cis*-2-hexene and *trans*-2-hexene were determined by comparison with the gas chromatogram of a 0.1 M solution in acetic acid of a standard mixture consisting of 97% 1-hexene and 3% of a mixture of *cis*-2-hexene and *trans*-2-hexene. Based on the amounts found and the amount of 1-hexene which can be calculated to have undergone addition after $^1/_6$ of a half-life, values of k_2/k_1 (ratio of rearrangement to addition) could be calculated. Two experiments gave $k_2/k_1 = 0.35$ and 0.47, respectively. The *cis* to *trans* ratios in the rearranged alkene were found to be 1.42 and 1.72, respectively. At the short reaction time employed, reaction of the 2-hexenes with trifluoroacetic acid is negligible within the accuracy of the analyses. A short fore-column of firebrick impregnated with potassium carbonate and then coated with Carbowax 20 M was used in the gas chromatographic analysis.

Isomerization of 2-Hexyl Trifluoroacetate.—A 0.1 M solution of 2-hexyl trifluoroacetate in trifluoroacetic acid 0.125 M in sodium trifluoroacetate was placed in a 60° bath. After 12 days, 21% isomerization to 3-hexyl trifluoroacetate had occurred as shown by analysis of the products (as the corresponding alcohols) on a tris-2-cyanoethoxypropane column. At equilibrium, approximately 37% of 3-hexyl trifluoroacetate is present. Detailed consideration of this experiment indicates that substituent isomerization was not important for the trifluoroacetates used in our study.

Acknowledgment.—The comments of Professor R. W. Taft have been particularly helpful. Alkene samples donated by Dr. H. G. Richey and Dr. Sung Moon are gratefully acknowledged.

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY, TEXAS TECHNOLOGICAL COLLEGE, LUBBOCK, TEXAS]

The Decomposition of Acetyl Peroxide in Solution. III. Kinetics and Use of Radical Traps[1]

BY H. J. SHINE, J. A. WATERS,[2] AND D. M. HOFFMAN

RECEIVED SEPTEMBER 28, 1962

The rates of decomposition of acetyl peroxide in a number of olefins, benzene, carbon tetrachloride, and cumene at several temperatures were measured. The rate constants for the first-order process do not differ markedly from solvent to solvent. Energies of activation of decomposition in the olefins are similar to those in benzene, carbon tetrachloride, and toluene. In particular, the rate constants for decomposition in cyclopentene and cyclohexene are essentially the same as each other. The rate of formation of cyclopentyl acetate was found to be about twice as large as that of cyclohexyl acetate. The effect of iodine, water, and 9,10-dihydroanthracene on the rate of decomposition of acetyl peroxide has been studied. Iodine lowers the rate in carbon tetrachloride, water increases the rate, while the effect of 9,10-dihydroanthracene varies with the solvent. Some similar experiments with benzoyl peroxide were carried out. Acetyl peroxide is hydrolyzed when heated in wet carbon tetrachloride. A product of hydrolysis was titrated iodometrically and is assumed to be peracetic acid. The hydrolysis by water makes ambiguous the "trapping" experiments with iodine–water reported earlier. Hydrolysis of benzoyl peroxide is comparatively minor. The effect of diphenylpicrylhydrazyl and galvinoxyl on benzoyl and acetyl peroxides is to decrease the extent of carbon dioxide formation. At high concentrations (near 10^{-2} M) of peroxide and scavenger the effect of diphenylpicrylhydrazyl is greater than that of galvinoxyl. This may be due to induced decomposition by the hydrazyl. Increasing concentrations of the hydrazyl fail to reduce the yield of carbon dioxide to zero. In the case of acetyl peroxide the yield levels off at about 40%. It is argued that this leveling-off is not explainable unless acetoxy radicals are being trapped by the hydrazyl. At low concentrations of acetyl peroxide and diphenylpicrylhydrazyl (near 10^{-4} M) trapping of the acetoxy radical occurs to a small extent, if at all. It appears that the problem of the decomposition of diacyl peroxides in the presence of diphenylpicrylhydrazyl needs to be resolved further. Experiments with both peroxides in solutions of cyclohexene containing 9,10-dihydroanthracene indicate that the acyloxy radicals prefer to add to the double bond rather than abstract allylic hydrogen. It is our view that the rates of decomposition of acetyl peroxide in cyclopentene and cyclohexene and the rates of formation of the cycloalkyl acetates together show the participation of the acetoxy radical in ester formation. Attempts to substantiate this point of view by scavenging the acetoxy radical have given results which are ambiguous.

The decomposition of acetyl peroxide in cyclohexene was shown to be unique[3] in that a substantial part of the acetoxy groups of the peroxide was recovered as cyclohexyl acetate. The formation of the saturated ester was interpreted[3] as a radical addition reaction and

(1) Presented in part at the 142nd National Meeting of the American Chemical Society, Atlantic City, N. J., September, 1962. Abstracted in part from the M. S. thesis of D. M. Hoffman, Texas Technological College, August, 1962.

(2) Robert A. Welch Foundation post-doctoral fellow.

(3) H. J. Shine and J. R. Slagle, *J. Am. Chem. Soc.*, **81**, 6309 (1959).