tral comparison. Both ir and 100-Mcps nmr spectra were superimposable; for details see the tables.

A control experiment was carried out to see whether heating could or could not have caused N inversion of the methobromide-d_3. No change of the positions of methyl and H-1,5 signals could be observed under these conditions.

N_b-Methoxycarbonylmethyl-d_2-3α,6β-dihydroxy-d_2-tropanium bromide (7a-d_4) was prepared in a completely analogous manner from **7a** as the tropine derivative. Hydrolysis to the betaine **7e**-d_4 by silver oxide in D_2O proceeded normally. However, evaporation of the carboxylic acid bromide **7a**-d_5 had to be carried out with freeze drying in order to avoid spontaneous lactonization. This way the yield was quantitative. Decarboxylation of 0.5 g (0.0015 mol) of **acid 7d**-d_5 was carried out in 35 ml of quinoline, by refluxing under N_2 with 0.4 g of cupric bromide for 2.5 hr. 84% of the calculated amount of CO_2 was trapped as $BaCO_3$ and the mixture worked up as given for the tropine derivative. Precipitation with 0.52 g of $AgSbF_6$ in 10 ml of nitromethane afforded 0.462 g of N_b-methyl-d_3-3α,6β-dihydroxytropanium hexafluoroantimonate (69% over-all yield), mp 263-265°.

The axial N-Me signal is at 3.01 ppm at 60 Mcps (in acetonitrile-d_3) of this authentic specimen; its position is same as that of the major product of deuteriomethylating 3α,6β-tropanediol. For more spectral and analytical data see the tables. Control experiments have shown no change in N-Me chemical shift when **21** (bromide) was heated in quinoline under the same conditions under which decarboxylation of **7d**-d_5 (bromide) had previously been carried out.

Thus it was definitely proven that deuteriomethylation of tropine and of 3α,6β-tropanediol has taken the same steric course as other N alkylations.

Conversion of Tropine Ethobromide 2 and of Pseudotropine Ethobromide 22 into Tropane Ethobromide 23. (a) Following procedures used for conversion of **3a, 7a,** and **12a** into **16d,** 1.45 g (0.0058 mol) of tropine ethobromide in 10 ml of benzene with 5 ml of thionyl chloride and 0.12 ml of DMF gave 0.927 g of 3β-chlorotropane ethobromide, mp 210°. Hydrogenolysis of 0.537 g (0.002 mol) of the same specimen in 15 ml of water over 2 g of Raney-Ni and 0.414 g (0.002 mol + 10% excess) of potassium carbonate led to 0.735 g of tropane ethobromide, mp 350° dec; for analysis and spectral data see the tables.

(b) Pseudotropine ethobromide (**22**), 2.3 g (0.0094 mol), 9 ml of $SOCl_2$, and 0.4 ml of DMF as catalyst gave by refluxing, 5 hr, 0.529 g of 3α-chlorotropane ethobromide, mp 182°, that by hydrogenolysis led to 240 mg of tropane ethobromide (**23**), mp 346° dec; ir and nmr spectra were superimposable with those obtained from 3α-tropanol ethobromide. Tentative configurations at C-3 had been ascribed to the chloro derivatives based upon relative chemical shifts[29,30] of H-3 (3.43 ppm for α and 4.30 for β).

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Vinyl Ether Hydrolysis. III. Some Brønsted Relations and Transition State Structure[1]

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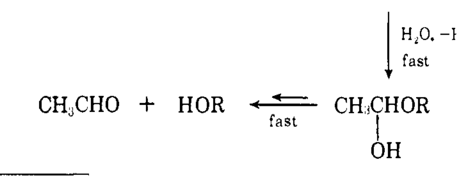
Abstract: Brønsted relations for the acid-catalyzed hydrolysis of seven vinyl ethers are constructed using a homogeneous set of carboxylic acid catalytic coefficients. Small but real systematic deviations from these relationships by electronegatively substituted carboxylic acids are detected, and these are attributed to intermolecular interactions between catalyst and substrate in the transition state. The presence of these interactions implies that the exponents of these Brønsted relations overestimate the extent of proton transfer at the transition state of these reactions by at least 0.1; this conclusion is consistent with the result of isotope effect studies. Additional evidence for intermolecular effects in proton transfer reactions is adduced from other deviations from Brønsted relations, and it is suggested that some of the large negative deviation usually found for the hydronium ion in rate-determining proton transfer to a substrate is the result of such effects.

Proton transfer from catalyst to substrate is an essential component of all acid-catalyzed reactions. In most cases, however, this elementary reaction step takes place in a rapidly established equilibrium preceding the rate-determining step; this makes it inaccessible to direct kinetic investigation by ordinary methods. In the acid-catalyzed hydrolysis of simple vinyl ethers, on the other hand, proton transfer from catalyst to substrate is itself rate-determining (eq 1).[1b,3-8] Since proton transfer is also the first step



$$\begin{tikzcd} CH_2{=}CHOR + HA \arrow[r, "slow"'] & CH_3CHOR^+ + A^+ \arrow[d, "H_2O,\ -H^+", "fast"'] \\ CH_3CHO + HOR & \underset{\displaystyle OH}{CH_3CHOR} \arrow[l, shift left, "fast"'] \arrow[l, shift right, leftarrow] \end{tikzcd}$$

(1)

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in this reaction, and since the reaction is not subject to catalysis by bases, this important elementary reaction step is isolated here in particularly simple form. Thus, vinyl ether hydrolysis presents a good opportunity for studying the proton transfer process.

One aspect of proton transfer which has received increasing attention lately is the structure of its transition state, especially the extent of proton transfer in this state.[9] This is the property of the transition state which, to a large degree, governs the magnitude, and therefore the mechanistic interpretation, of hydrogen isotope effects;[10] it also plays a prominent role in determining the kinetic acidity dependence of proton transfer reactions in concentrated acid solution.[11] Transition states, of course, cannot be observed, and transition state structure cannot be determined directly; it must be deduced indirectly. There is some reason to believe that the Brønsted relation[12] may be capable of doing this, and that the exponent in this relation might even be numerically equal to the degree of proton transfer at the transition state.[13a] As part of our general intention to examine this hypothesis, we have constructed Brønsted relations for the hydrolysis of seven vinyl ethers. We present the relevant data here and show how these bear on the usefulness of Brønsted exponents as indicators of transition state structure.

Results

The Brønsted relation (eq 2) correlates catalytic coefficients, k_{HA}, for a given reaction with acid dissociation constants of the catalysts, K_{HA}. In order to minimize deviations from this relationship known to occur when acids of different structure are used,[14,15] only carboxylic acids were employed as catalysts in the present study.

$$k_{\mathrm{HA}} = G(K_{\mathrm{HA}})^{\alpha} \qquad (2)$$

No attempt was made at first to use the same catalysts for each substrate, but, as the work progressed, consistent deviations for individual acids were noticed and a uniform set of catalysts was therefore adopted. This set is cyanoacetic, chloroacetic, methoxyacetic, formic, glycolic, acetic, and propionic acids; some catalytic coefficients for phenoxyacetic, phenylacetic, and pivalic acids were also obtained. Rate constants for catalysis by the hydronium ion were needed to analyze kinetic data obtained in solutions of the stronger carboxylic acids, and these were therefore measured as well (except in the case of ethyl vinyl ether where the value was available from our earlier work[1b,5]).

Measurements were made on seven substrates; these encompassed six orders of magnitude in reactivity and ranged from ethyl isopropenyl ether (most reactive) to phenyl vinyl ether (least reactive). The other five, in order of decreasing reactivity, were ethyl cyclopentenyl ether, ethyl cyclohexenyl ether, methyl cyclohexenyl ether, phenyl isopropenyl ether, and ethyl vinyl ether.

Rate constants were determined for the most part by monitoring the decrease in the strong end absorption of vinyl ethers at 210–230 nm.[5] When half-lives were longer than *ca.* 20 sec, measurements were made with a conventional recording spectrometer (Cary Model 11). A Durrum–Gibson stopped-flow spectrometer was used for faster reactions, *i.e.*, for all of the hydrolyses of ethyl isopropenyl ether except for some runs catalyzed by formic and acetic acids and for the hydrolyses of ethyl cyclopentenyl ether catalyzed by the hydronium ion and by cyanoacetic and chloroacetic acids.

The hydrolysis of phenyl vinyl ether catalyzed by carboxylic acid was too slow to allow convenient measurement by these methods; for example, the half-life of this reaction in acetic acid buffers is approximately 10^2 days. An initial-rate method was therefore used instead. The technique adopted employs a radioactive tracer to measure small extents of reaction and depends upon the fact that labeled product, phenol-*4-t*, obtained in the hydrolysis of phenyl-*4-t* vinyl ether, can be separated cleanly from unreacted starting material by basic extraction. Control experiments showed that neither phenol nor phenyl vinyl ether undergo ring-hydrogen exchange under hydrolysis and extraction conditions; the radioactivity of the basic extract, therefore, is directly proportional to the extent of the hydrolysis reaction. Since the radioactivity of this extract rises from zero (background) to some level determined only by the molar activity of the substrate (and the extent of the reaction), the method is capable of great sensitivity, limited essentially only by the level of radioactivity which it is practical to place in the substrate. In the present work, good precision was obtained with activities of the order of 0.5 Ci mol^{-1} when measurements were made over the first 2–3% reaction.

Zero-order plots of counts per minute (cpm) *vs.* time were constructed using data obtained in this way, and the slopes of these plots were then converted into first-order rate constants by dividing by infinite-time radioactivities. The latter were gotten through accelerating hydrolysis in a portion of reaction mixture by adding a known amount of strong acid, waiting for 10–12 half-lives to elapse under the new conditions, and then performing the radioassay as before.

This method produced results whose precision is equal to that realized by the spectroscopic method using a conventional spectrometer; rate constants could be reproduced to *ca.* 2–3% (standard deviation). Somewhat inferior results were obtained with the stopped-flow spectrometer, but reproducibility here was still of the order of 10%. Intercomparison of the three methods gave agreement within these precision limits.

Catalysis by the Hydronium Ion. Hydronium ion catalytic coefficients were determined in perchloric acid solution. Observed rate constants were accurately first order, and these first-order constants were directly proportional to acid concentration in all cases; the reaction is therefore second order overall. The relevant data are presented in Table I.

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[10] For a recent review of the subject and references to the previous literature, see R. A. More O'Ferrall and J. Kouba, *J. Chem. Soc. B*, 985 (1967).

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[13] (a) See, for example, J. E. Leffler and E. Grunwald, "Rates and Equilibria of Organic Reactions," Wiley, New York, N. Y., 1963, p 235; (b) p 24.

[14] R. P. Bell, "Acid-base Catalysis," Oxford University Press, New York, N. Y., 1941, Chapter V.

[15] R. P. Bell, "The Proton in Chemistry," Cornell University Press, Ithaca, N. Y., 1959, (a) Chapter X; (b) p 171; (c) p 172.

Table I. Rates of Hydrolysis of Vinyl Ethers in Perchloric Acid Solution at 25° [a]

$10^2[\mathrm{H}^+]$, M	$10^2 k_{\mathrm{obsd}}$, sec^{-1}	k_{H^+}, M^{-1} sec^{-1}
Ethyl isopropenyl ether		
0.097	37.1, 38.4, 47.1, 48.6, 53.3, 64.9	497 ± 43
0.299	160, 165, 167, 175, 182	570 ± 12
0.507	267, 267, 272, 283, 297, 299, 305, 312	566 ± 13
0.551	310, 312, 351, 339, 388	620 ± 25
0.795	448, 474, 499, 515, 590	635 ± 30
0.963	515, 516, 522, 541, 585, 621	571 ± 19
0.984	564, 587, 588, 611, 631, 641	613 ± 12
	Grand mean	579 ± 11
Phenyl vinyl ether		
2.96	0.0096	0.00325
3.94	0.0127, 0.0129	0.00325 ± 0.00003
5.91	0.0196, 0.0197	0.00332 ± 0.00001
7.97	0.0260	0.00327
10.2	0.0341	0.00336
10.4	0.0332, 0.0341	0.00325 ± 0.00005
	Grand mean	0.00328 ± 0.00002
Phenyl isopropenyl ether		
0.130	73.7, 75.0, 75.1, 75.6, 76.1, 76.5, 78.1, 78.2, 78.7, 79.8	5.89 ± 0.05
0.195	112, 116, 118, 119, 121, 121, 122	6.06 ± 0.07
0.260	147, 148, 152, 152, 155, 155	5.80 ± 0.05
0.326	197, 203, 203, 204, 204, 205	6.23 ± 0.04
	Grand mean	5.98 ± 0.04
Ethyl cyclopentenyl ether		
0.60	263, 288, 297	428 ± 16
0.72	339, 345, 345	478 ± 03
1.00	435, 445, 491	457 ± 17
1.50	691, 691	461
2.00	864, 880, 953	450 ± 14
	Grand mean	454 ± 17
Methyl cyclohexenyl ether		
0.049	2.15, 2.17	44.1 ± 0.3
0.082	3.48, 3.48, 3.48	42.5
0.131	5.38, 5.38, 5.62	41.7 ± 0.7
0.205	8.06, 8.54, 8.77	41.2 ± 1.0
0.276	12.9, 12.9	46.7
0.375	14.8, 15.3, 15.4	40.5 ± 0.5
2.25	89.8, 90.7, 92.1	41.7 ± 1.4
3.00	124, 126, 130	42.2 ± 0.6
3.75	155, 155, 164	42.1 ± 0.8
	Grand mean	42.3 ± 0.4
Ethyl cyclohexenyl ether		
0.60	44.4, 48.8, 52.5	80.9 ± 3.9
0.80	55.1, 63.8, 69.1	78.3 ± 5.1
1.00	80.0, 80.4, 85.9	82.1 ± 1.9
1.50	115, 116, 123	79.9 ± 1.7
2.00	157, 160, 162	78.8 ± 0.9
	Grand mean	80.0 ± 1.2

[a] Ionic strength = 0.040 M except for phenyl vinyl and phenyl isopropenyl ethers where it was the same as the perchloric acid concentration. Error estimates are standard deviations of mean values.

Table II. Rates of Hydrolysis of Ethyl Vinyl Ether in Aqueous Buffer Solutions at 25° [a]

$10^2[\mathrm{HA}]$, M[b]	$10^3 k_{\mathrm{obsd}}$, sec^{-1}
HA = $\mathrm{CNCH_2CO_2H}$, [HA]/[NaA] = 0.995	
0.79	5.78, 4.80
1.59	6.27, 6.28
1.99	6.83, 6.85, 6.90
2.39	7.20, 7.25
3.18	8.05, 8.05
3.98	8.60, 8.60, 8.70
$k_1\,(\mathrm{sec}^{-1}) = (7.05 \pm 0.03) \times 10^{-3} + (4.42 \pm 0.13) \times 10^{-2}[\mathrm{HA}]$[c]	
$k_{\mathrm{H}^+} = 1.80\ M^{-1}\,\mathrm{sec}^{-1}$	
HA = $\mathrm{ClCH_2CO_2H}$, [HA]/[NaA] = 2.00	
0.80	3.69, 3.70
1.20	4.41, 4.43
1.60	4.76, 4.76, 4.83, 4.85
2.00	5.21, 5.24, 5.32
2.80	5.85, 5.90
3.20	6.18, 6.20, 6.25
$k_1\,(\mathrm{sec}^{-1}) = (5.25 \pm 0.03) \times 10^{-3} + (3.24 \pm 0.16) \times 10^{-2}[\mathrm{HA}]$[c]	
$k_{\mathrm{H}^+} = 1.79\ M^{-1}\,\mathrm{sec}^{-1}$	
HA = $\mathrm{CH_3OCH_2CO_2H}$, [HA]/[NaA] = 2.50	
0.71	1.30, 1.31, 1.35, 1.38
1.42	1.54, 1.61, 1.63, 1.66
2.13	1.77, 1.77, 1.78, 1.79
2.84	1.76, 1.76, 1.77, 1.77
3.55	1.95, 1.97, 1.99, 2.00
4.26	2.04, 2.07, 2.12, 2.13
$k_1\,(\mathrm{sec}^{-1}) = (1.60 \pm 0.02) \times 10^{-3} + (1.11 \pm 0.09) \times 10^{-2}[\mathrm{HA}]$[c]	
$k_{\mathrm{H}^+} = 1.77\ M^{-1}\,\mathrm{sec}^{-1}$	
HA = $\mathrm{HCO_2H}$, [HA]/[NaA] = 2.40	
1.30	1.06, 1.07
1.95	1.18
3.90	1.32, 1.33
4.55	1.41, 1.41
5.20	1.46
$k_1\,(\mathrm{sec}^{-1}) = (1.10 \pm 0.01) \times 10^{-3} + (6.72 \pm 0.42) \times 10^{-3}[\mathrm{HA}]$[c]	
$k_{\mathrm{H}^+} = 1.71\ M^{-1}\,\mathrm{sec}^{-1}$	
HA = $\mathrm{HOCH_2CO_2H}$, [HA]/[NaH] = 3.05	
1.49	1.10, 1.11, 1.13
2.23	1.16, 1.17
3.72	1.28, 1.29, 1.29, 1.31
4.46	1.34, 1.38, 1.39, 1.39
5.20	1.38, 1.41, 1.41
$k_1\,(\mathrm{sec}^{-1}) = (1.12 + 0.01) \times 10^{-3} + (5.53 + 0.38) \times 10^{-3}\,[\mathrm{HA}]$[c]	
$k_{\mathrm{H}^+} = 1.85\ M^{-1}\,\mathrm{sec}^{-1}$	
HA = $\mathrm{CH_3CO_2H}$, [HA]/[NaA] = 2.95	
3.01	0.174, 0.175
5.98	0.211, 0.212
8.96	0.261, 0.266, 0.267
11.9	0.290, 0.294, 0.298
$k_1\,(\mathrm{sec}^{-1}) = (1.34 \pm 0.05) \times 10^{-4} + (1.38 \pm 0.06) \times 10^{-3}[\mathrm{HA}]$[c]	
$k_{\mathrm{H}^+} = 1.83\ M^{-1}\,\mathrm{sec}^{-1}$	
HA = $\mathrm{CH_3CH_2CO_2H}$, [HA]/[NaA] = 4.60	
1.84	0.150, 0.153
2.76	0.159, 0.164
3.68	0.180
4.60	0.185, 0.188
6.45	0.201, 0.201
$k_1\,(\mathrm{sec}^{-1}) = (1.37 \pm 0.03) \times 10^{-4} + (1.04 \pm 0.08) \times 10^{-3}[\mathrm{HA}]$[c]	
$k_{\mathrm{H}^+} = 1.60\ M^{-1}\,\mathrm{sec}^{-1}$	
$k_{\mathrm{H}^+} = 1.76 \pm 0.03\ M^{-1}\,\mathrm{sec}^{-1}$ (average of all buffers)	

[a] Ionic strength = 0.040 M except for $\mathrm{CNCH_2CO_2H}$ buffer solutions where it was 0.044 M. Error estimates are standard deviations of mean values. [b] Stoichiometric concentrations. [c] Relationship obtained by least-squares fit of observed rate constants, adjusted when necessary to the hydronium ion concentration of the most concentrated buffer solution, to undissociated acid concentrations, also adjusted if necessary.

One of these catalytic coefficients, that for the hydrolysis of methyl cyclohexenyl ether, has been measured before in a different laboratory.[8] The value reported, 46.8 M^{-1} sec^{-1} at $\mu = 0.25$ M, is slightly higher than the result obtained here, 42.3 M^{-1} sec^{-1} at $\mu = 0.040$ M; this suggests that the reaction may be subject to a weak positive primary salt effect.

Catalysis by Carboxylic Acids. Carboxylic acid catalytic coefficients were determined by measuring rates of hydrolysis at constant ionic strength in series of buffer solutions of constant buffer ratio but changing total buffer concentration. Observed rate constants were found to be exactly first order; the relevant data are summarized in Tables II–VIII.

Table III. Rates of Hydrolysis of Ethyl Isopropenyl Ether in Aqueous Buffer Solutions at 25° [a]

10^2[HA], M[b]	10^3k_{obsd}, sec^{-1}
HA = $CNCH_2CO_2H$, [HA]/[NaA] = 1.00	
1.60	22.2, 22.5, 23.9
2.20	25.9, 26.3, 26.5, 26.5
3.20	30.5, 31.1, 32.2, 32.2
4.00	34.5, 36.7, 36.8
k_1 (sec^{-1}) = (2.22 + 0.06) + (3.69 + 0.24) × 10^1[HA][c]	
k_{H^+} = 5.68 × 10^2 M^{-1} sec^{-1}	
HA = $ClCH_2CO_2H$, [HA]/[NaA] = 1.00	
0.63	7.83, 8.41, 8.52, 9.67
0.75	8.87, 9.44, 9.44, 10.1
1.00	10.4, 10.9, 10.9, 11.2
1.25	10.9, 11.9
1.99	13.7, 13.9, 14.4, 14.4
k_1 (sec^{-1}) = (9.58 + 0.27) × 10^{-1} + (2.49 + 0.25) × 10^1[HA][c]	
k_{H^+} = 5.92 × 10^2 M^{-1} sec^{-1}	
HA = $CH_3OCH_2CO_2H$, [HA]/[NaA] = 1.00	
0.49	2.65, 2.70
0.62	3.11, 3.17
1.23	3.71, 3.80, 3.80
1.97	4.84, 5.01, 5.18
2.46	5.41, 5.47, 5.58
k_1 (sec^{-1}) = (2.39 + 0.09) × 10^{-1} + (1.31 + 0.06) × 10^1[HA][c]	
k_{H^+} = 6.53 × 10^2 M^{-1} sec^{-1}	
HA = HCO_2H, [HA]/[NaA] = 1.00	
1.00	1.99, 2.01, 2.15, 2.24
1.60	2.48, 2.55, 2.55, 2.69
2.00	2.67, 2.68, 2.70, 2.71, 2.82
2.50	2.95, 3.00, 3.07, 3.20, 3.21
3.00	3.45, 3.49, 3.51, 3.53
4.00	3.92, 3.96, 3.97, 4.03
k_1 (sec^{-1}) = (1.58 ± 0.05) × 10^{-1} + (6.13 ± 0.20)[HA][c]	
k_{H^+} = 6.44 × 10^2 M^{-1} sec^{-1}	
HA = HCO_2H, [HA]/[NaA] = 0.100	
0.055	0.184, 0.185, 0.190
0.101	0.217, 0.220, 0.221
0.202	0.269, 0.269, 0.276
0.291	0.329, 0.330, 0.331
0.301	0.336, 0.337, 0.338
0.402	0.397, 0.409
k_1 (sec^{-1}) = (1.59 ± 0.03) × 10^{-2} + (5.93 ± 0.11)[HA][c]	
k_{H^+} = 6.24 × 10^{-2} M^{-1} sec^{-1}	
HA = $HOCH_2CO_2H$, [HA]/[NaA] = 1.00	
0.50	1.50, 1.51, 1.60
1.25	2.13, 2.22, 2.30
1.86	2.65, 2.71, 2.76
2.34	2.69, 2.81, 2.88, 3.00
k_1 (sec^{-1}) = (1.34 ± 0.08) × 10^{-1} + (6.83 ± 0.49)[HA][c]	
k_{H^+} = 6.45 × 10^2 M^{-1} sec^{-1}	
HA = CH_3CO_2H, [HA]/[NaA] = 1.00	
0.100	0.0303, 0.0305, 0.0317, 0.0323, 0.0330
0.200	0.419, 0.431, 0.450, 0.453, 0.453
0.200	0.423, 0.426, 0.427, 0.453, 0.453
0.250	0.476, 0.479, 0.493, 0.520
0.300	0.546, 0.567, 0.568, 0.573, 0.574
0.400	0.695, 0.699, 0.713, 0.755, 0.759
0.400	0.689, 0.690, 0.704, 0.732
k_1 (sec^{-1}) = (1.70 ± 0.09) × 10^{-2} + (1.33 ± 0.06)[HA][c]	
k_{H^+} = 6.90 × 10^2 M^{-1} sec^{-1}	
HA = CH_3CO_2H, [HA]/[NaA] = 0.270	
0.220	0.0792, 0.0795
0.364	0.0988, 0.0990
0.609	0.134, 0.134
0.910	0.170, 0.173, 0.173
0.962	0.175, 0.177, 0.179
1.22	0.214, 0.216, 0.218
k_1 (sec^{-1}) = (4.95 ± 0.14) × 10^{-3} + (1.35 ± 0.02)[HA][c]	
k_{H^+} = 6.57 × 10^2 M^{-1} sec^{-1}	
HA = CH_3CO_2H, [HA]/[NaA] = 0.100	
0.100	0.0288, 0.0296
0.200	0.0400, 0.0414
0.300	0.0511, 0.0528
0.400	0.0660, 0.0677
k_1 (sec^{-1}) = (1.62 ± 0.12) × 10^{-3} + (1.24 ± 0.04)[HA][c]	
k_{H^+} = 6.57 × 10^2 M^{-1} sec^{-1}	
HA = $CH_3CH_2CO_2H$, [HA]/[NaA] = 1.00	
0.386	0.164, 0.175, 0.175
0.745	0.198, 0.203
0.900	0.216, 0.226, 0.230
1.08	0.230, 0.256, 0.287
k_1 (sec^{-1}) = (1.21 ± 0.14) × 10^{-2} + (1.19 ± 0.17)[HA][c]	
k_{H^+} = 6.51 × 10^2 M^{-1} sec^{-1}	
HA = $(CH_3)_3CCO_2H$, [HA]/[NaA] = 1.00	
0.142	0.117, 0.120
0.284	0.140, 0.143, 0.145
0.427	0.168, 0.170
0.569	0.198, 0.202
k_1 (sec^{-1}) = (9.05 ± 0.02) × 10^{-3} + (1.80 ± 0.06)[HA][c]	
k_{H^+} = 6.80 × 10^2 M^{-1} sec^{-1}	
k_{H^+} = (6.42 ± 0.11) × 10^2 (average of all buffers)	

[a] Ionic strength = 0.040 M. Error estimates are standard deviations of mean values. [b] Stoichiometric concentrations. [c] Relationship obtained by least-squares fit of observed rate constants, adjusted when necessary to the hydronium ion concentration of most concentrated buffer solution, to undissociated acid concentrations, also adjusted if necessary.

With well-behaved buffers, the hydronium ion concentration of each such series of solutions was constant, and the first term in the rate law governing vinyl ether hydrolysis under these conditions (eq 3)[4a,5,6,8] was

$$-d[S]/[S]dt = k_{\text{obsd}} = k_{H_3O^+}[H_3O^+] + k_{HA}[HA] \quad (3)$$

therefore also constant. Consequently, catalytic coefficients, k_{HA}, could be evaluated as slopes of plots of k_{obsd} *vs.* [HA]. At the rather low buffer concentrations employed, however, some of the stronger carboxylic acids gave series of buffer solutions whose hydronium ion concentrations decreased with decreasing buffer concentration. Such buffer failure was compensated for by adjusting observed rate constants within a given series of buffer solutions to a common hydronium ion concentration. These corrections were based on the hydronium ion catalytic coefficients of Table I and ref 5, and on hydronium ion concentrations calculated from the acidity constants listed in Table IX and activity coefficients determined by the Debye–Hückel equation with an ion-size parameter of 6.00 Å. Best values of carboxylic acid catalytic coefficients were obtained by least-squares fitting of first-order rate constants (k_{obsd}, adjusted if necessary) to eq 3; these are summarized in Table IX. Hydronium ion catalytic coefficients were also evaluated from these buffer data using the terms $k_{H_3O^+}[H_3O^+]$ (eq 3) and calculated hydronium ion concentrations; these are listed in Tables II–VIII. On the whole, hydronium ion catalytic coefficients so obtained were consistent with directly measured values; in only one case (ethyl isopropenyl ether) did the difference between the two values exceed their combined standard deviations. This agreement attests to the

Table IV. Rates of Hydrolysis of Phenyl Vinyl Ether in Aqueous Buffer Solutions at 25° [a]

10^2[HA], M[b]	10^6k_{obsd}, sec^{-1}
HA = $ClCH_2CO_2H$, [HA]/[NaA] = 3.14	
2.51	10.6
3.77	12.0
5.03	13.6
5.66	14.3
6.29	14.9
k_1 (sec^{-1}) = $(1.25 \pm 0.01) \times 10^{-5} + (4.04 \pm 0.57) \times 10^{-5}$[HA][c]	
$k_{H^+} = 2.93 \times 10^{-3}\ M^{-1}$ sec^{-1}	
HA = $C_6H_5OCH_2CO_2H$, [HA]/[NaA] = 2.27	
1.07	4.50
2.14	5.64
3.21	6.43
4.28	6.95
5.35	7.37
k_1 (sec^{-1}) = $(5.75 \pm 0.05) \times 10^{-6} + (3.24 \pm 0.15) \times 10^{-5}$[HA][c]	
$k_{H^+} = 3.11 \times 10^{-3}\ M^{-1}$ sec^{-1}	
HA = $CH_3OCH_2CO_2H$, [HA]/[NaA] = 1.00	
4.00	1.73
5.00	1.81
6.00	1.93
8.00	2.16
9.00	2.24
k_1 (sec^{-1}) = $(1.33 \pm 0.03) \times 10^{-6} + (1.02 \pm 0.03) \times 10^{-5}$[HA][c]	
$k_{H^+} = 3.15 \times 10^{-3}\ M^{-1}$ sec^{-1}	
HA = HCO_2H, [HA]/[NaA] = 1.00	
2.00	1.04, 1.16
4.00	1.19
5.00	1.31
6.00	1.31, 1.34, 1.48, 1.54
8.00	1.57, 1.62
10.0	1.66, 1.70
k_1 (sec^{-1}) = $9.64 \pm 0.55) \times 10^{-7} + (7.44 \pm 0.84) \times 10^{-6}$[HA][c]	
$k_{H^+} = 3.46 \times 10^{-3}\ M^{-1}$ sec^{-1}	
HA = $C_6H_5CH_2CO_2H$, [HA]/[NaA] = 2.05	
1.93	0.492
2.90	0.517
3.86	0.542
4.83	0.568
k_1 (sec^{-1}) = $(4.41 \pm 0.01) \times 10^{-7} + (2.63 \pm 0.01) \times 10^{-6}$[HA][c]	
$k_{H^+} = 3.50 \times 10^{-3}\ M^{-1}$ sec^{-1}	
HA = CH_3CO_2H, [HA]/[NaA] = 1.00	
2.00	0.114
4.00	0.139
5.00	0.147
6.00	0.156
8.00	0.185
10.0	0.212
k_1 (sec^{-1}) = $(8.84 \pm 0.33) \times 10^{-8} + (1.21 \pm 0.05) \times 10^{-6}$[HA][c]	
$k_{H^+} = 3.19 \times 10^{-3}\ M^{-1}$ sec^{-1}	
$k_{H^+} = (3.22 \pm 0.09) \times 10^{-3}\ M^{-1}$ sec^{-1} (average of all buffers)	

[a] Ionic strength = 0.100 M except for $C_6H_5OCH_2CO_2H$ buffer solutions where it was 0.028 M and $C_6H_5CH_2CO_2H$ and $ClCH_2CO_2H$ buffer solutions where it was 0.20 M. Error estimates are standard deviations of mean values. [b] Stoichiometric concentrations. [c] Relationship obtained by least-squares fit of observed rate constants, adjusted when necessary to the hydronium ion concentration of the most concentrated buffer solution, to undissociated acid concentrations, also adjusted if necessary.

Table V. Rates of Hydrolysis of Phenyl Isopropenyl Ether in Aqueous Buffer Solutions at 25° [a]

10^2[HA], M[b]	10^3k_{obsd}, sec^{-1}
HA = $CNCH_2CO_2H$, [HA]/[NaA] = 0.207	
0.79	6.89
1.18	7.89
1.58	8.52
1.97	9.15
k_1 (sec^{-1}) = $(6.40 \pm 0.09) \times 10^{-3} + (1.49 \pm 0.06) \times 10^{-1}$[HA][c]	
$k_{H^+} = 6.21\ M^{-1}$ sec^{-1}	
HA = $CH_3OCH_2CO_2H$, [HA]/NaA] = 0.98	
0.79	2.57
1.18	2.78
1.57	2.99
1.97	3.13
2.95	3.48
3.93	4.00
k_1 (sec^{-1}) = $(2.40 \pm 0.03) \times 10^{-3} + (4.02 \pm 0.10) \times 10^{-2}$[HA][c]	
$k_{H^+} = 6.59\ M^{-1}$ sec^{-1}	
HA = HCO_2H, [HA]/[NaA] = 0.98	
0.78	1.68
1.57	1.96
2.35	2.15
3.14	2.37
3.92	2.55
k_1 (sec^{-1}) = $(1.55 \pm 0.02) \times 10^{-3} + (2.52 \pm 0.07) \times 10^{-2}$[HA][c]	
$k_{H^+} = 6.43\ M^{-1}$ sec^{-1}	
HA = $HOCH_2CO_2H$, [HA]/[NaA] = 1.03	
0.82	1.43
1.64	1.67
2.47	1.86
3.29	2.07
4.11	2.28
k_1 (sec^{-1}) = $(1.29 \pm 0.01) \times 10^{-3} + (2.42 \pm 0.04) \times 10^{-2}$[HA][c]	
$k_{H^+} = 6.14\ M^{-1}$ sec^{-1}	
HA = CH_3CO_2H, [HA]/[NaA] = 2.50	
2.00	0.483
4.00	0.615
6.00	0.719
8.00	0.854
k_1 (sec^{-1}) = $(3.69 \pm 0.04) \times 10^{-4} + (6.04 \pm 0.21) \times 10^{-3}$[HA][c]	
$k_{H^+} = 6.02\ M^{-1}$ sec^{-1}	
HA = $(CH_3)_3CCO_2H$, [HA]/[NaA] = 1.50	
1.20	0.198
2.41	0.280
3.61	0.362
4.81	0.445
6.02	5.39
k_1 (sec^{-1}) = $(1.11 \pm 0.05) \times 10^{-4} + (7.03 \pm 0.11) \times 10^{-3}$[HA][c]	
$k_{H^+} = 5.66\ M^{-1}$ sec^{-1}	
$k_{H^+} = 6.18 \pm 0.13$ (average of all buffers)	

[a] Ionic strength = 0.040 M except for $CNCH_2CO_2H$ buffer solutions where it was 0.095 M. Error estimates are standard deviations of mean values. [b] Stoichiometric concentrations. [c] Relationship obtained by least-squares fit of observed rate constants, adjusted when necessary to the hydronium ion concentration of most concentrated buffer solution, to undissociated acid concentrations, also adjusted if necessary.

general validity of this method of treating the experimental data.

Several of the carboxylic acid catalytic coefficients measured here have been determined before in other laboratories. Those reported[8] for the hydrolysis of methyl cyclohexenyl ether by glycolic and acetic acids at $\mu = 0.25\ M$ are *ca.* 10% higher than our values at $\mu = 0.040\ M$; this is similar to the difference in hydronium ion catalytic coefficients for this reaction noted above and may represent a weak primary salt effect. It is unlikely that the much larger difference, approaching 50%, in chloroacetic acid catalytic coefficients for the same reaction can be reconciled in this way; it has been our experience, however, that the chloroacetate ion is easily hydrolyzed to glycolate ion and hydrochloric acid in aqueous solution, and, unless freshly prepared buffer solutions are used, erroneous kinetic results are likely to be obtained. Rate constants for the hydrolysis of ethyl vinyl ether by chloroacetic and formic acids have also been reported;[6] these are 4% higher and 13% lower, respectively, than our values.

Table VI. Rates of Hydrolysis of Ethyl Cyclopentenyl Ether in Aqueous Buffer Solutions at 25° [a]

10^2[HA], M^b	$10k_{\text{obsd}}$, sec^{-1}
HA = CNCH$_2$CO$_2$H, [HA]/[NaA] = 1.00	
0.40	8.91, 8.93, 9.34, 9.53, 9.65, 9.72
1.01	14.3, 14.4, 14.6, 15.0, 15.3, 15.8
2.01	19.6, 19.7, 20.0, 20.2, 20.3, 20.6
3.02	21.9, 22.0, 22.0, 22.0, 22.4, 22.6, 22.6
4.02	25.3, 25.4, 25.6, 26.0, 26.3, 26.5, 26.6, 27.3
$k_1\ (\text{sec}^{-1}) = (1.86 \pm 0.02) + 2.02 \pm 0.09 \times 10[\text{HA}]^c$	
$k_{\text{H}^+} = 4.76 \times 10^2\ M^{-1}\ \text{sec}^{-1}$	
HA = ClCH$_2$CO$_2$H, [HA]/[NaA] = 1.00	
0.400	5.69, 5.92, 5.93, 6.00, 6.10
1.00	8.38, 8.40, 8.46, 8.50, 8.56, 8.67, 8.71
2.00	10.6, 10.7, 10.7, 10.9, 11.5, 11.6
3.00	12.2, 12.2, 12.3, 12.5, 12.7, 12.7, 12.9
4.00	13.1, 13.6, 13.8, 14.0, 14.4, 14.4, 14.6, 15.1
$k_1\ (\text{sec}^{-1}) = (8.66 \pm 0.14) \times 10^{-1} + (1.45 \pm 0.06) \times 10[\text{HA}]^c$	
$k_{\text{H}^+} = 4.96 \times 10^2\ M^{-1}\ \text{sec}^{-1}$	
HA = CH$_3$OCH$_2$CO$_2$H, [HA]/[NaA] = 0.100	
0.080	0.202, 0.207
0.120	0.232, 0.234, 0.235
0.160	0.252, 0.254, 0.255
0.200	0.278, 0.278, 0.278
$k_1\ (\text{sec}^{-1}) = (1.70 \pm 0.03) \times 10^{-2} + (5.51 \pm 0.17)[\text{HA}]^c$	
$k_{\text{H}^+} = 4.60 \times 10^2\ M^{-1}\ \text{sec}^{-1}$	
HA = HCO$_2$H, [HA]/[NaA] = 0.134	
0.108	0.174, 0.176
0.215	0.202, 0.202, 0.209
0.323	0.244, 0.245, 0.245
0.430	0.259, 0.268, 0.272
0.538	0.317, 0.318, 0.320
$k_1\ (\text{sec}^{-1}) = (1.40 \pm 0.05) \times 10^{-2} + (3.19 \pm 0.13)[\text{HA}]^c$	
$k_{\text{H}^+} = 4.26 \times 10^2\ M^{-1}\ \text{sec}^{-1}$	
HA = HOCH$_2$CO$_2$H, [HA]/[NaA] = 0.109	
0.087	0.136, 0.136, 0.137
0.174	0.156, 0.160, 0.161
0.218	0.171, 0.171, 0.172
0.262	0.187, 0.192
0.349	0.218
0.436	0.242
$k_1\ (\text{sec}^{-1}) = (1.12 \pm 0.03) \times 10^{-2} + (2.94 \pm 0.11)[\text{HA}]^c$	
$k_{\text{H}^+} = 4.96 \times 10^2\ M^{-1}\ \text{sec}^{-1}$	
HA = CH$_3$CO$_2$H, [HA]/[NaA] = 0.109	
0.233	0.0885, 0.0886, 0.0888
0.699	0.127, 0.127
0.932	0.140, 0.140, 0.141
1.17	0.163, 0.164
1.40	0.179, 0.179, 0.179
1.63	0.194, 0.194, 0.194
1.86	0.212, 0.214, 0.215
$k_1\ (\text{sec}^{-1}) = (7.19 \pm 0.11) \times 10^{-3} + (7.58 \pm 0.09) \times 10^{-1}[\text{HA}]^c$	
$k_{\text{H}^+} = 5.50 \times 10^2\ M^{-1}\ \text{sec}^{-1}$	
HA = CH$_3$CH$_2$CO$_2$H, [HA]/[NaA] = 1.00	
0.408	0.128, 0.128, 0.129
1.02	0.152, 0.167, 0.174
2.04	0.230, 0.243, 0.254, 0.256, 0.257
3.06	0.320, 0.332, 0.336
4.08	0.327, 0.359, 0.390, 0.406, 0.434
$k_1\ (\text{sec}^{-1}) = (9.79 \pm 1.04) \times 10^{-3} + (7.10 \pm 0.39) \times 10^{-1}[\text{HA}]^c$	
$k_{\text{H}^+} = 5.36 \times 10^2\ M^{-1}\ \text{sec}^{-1}$	
$k_{\text{H}^+} = (4.85 + 0.13) \times 10^2\ M^{-1}\ \text{sec}^{-1}$ (average of all buffers)	

[a] Ionic strength = 0.040 M. Error estimates are standard deviations of mean values. [b] Stoichiometric concentrations. [c] Relationship obtained by least-squares fit of observed rate constants, adjusted when necessary to the hydronium ion concentration of the most concentrated buffer solution, to undissociated acid concentrations, also adjusted if necessary.

Table VII. Rates of Hydrolysis of Methyl Cyclohexenyl Ether in Aqueous Buffer Solutions at 25° [a]

10^2[HA], M^b	10^3k_{obsd}, sec^{-1}
HA = CNCH$_2$CO$_2$H, [HA]/[NaA] = 0.134	
0.107	15.4, 15.4
0.214	20.6, 20.9, 21.0
0.268	22.1, 22.1, 22.3
0.322	23.5, 23.6, 23.7
0.429	25.8, 26.1
0.536	27.5, 27.6, 27.7
$k_1 = (2.23 \pm 0.01) \times 10^{-2} + (1.13 \pm 0.03)[\text{HA}]^c$	
$k_{\text{H}^+} = 3.97 \times 10^1\ M^{-1}\ \text{sec}^{-1}$	
HA = ClCH$_2$CO$_2$H, [HA]/[NaA] = 0.210	
0.085	11.1, 11.6, 11.6
0.170	14.9, 14.9, 14.9
0.254	15.6, 15.6, 16.3
0.339	17.2, 17.3, 17.7
0.424	18.3, 18.8, 18.9
0.509	20.0, 20.0, 20.1
$k_1\ (\text{sec}^{-1}) = (1.58 \pm 0.01) \times 10^{-2} + (8.72 \pm 0.26) \times 10^{-1}[\text{HA}]^c$	
$k_{\text{H}^+} = 4.29 \times 10^1\ M^{-1}\ \text{sec}^{-1}$	
HA = CH$_3$OCH$_2$CO$_2$H, [HA]/[NaA] = 0.275	
0.220	4.88, 4.88
0.440	5.61, 5.66
0.660	6.38, 6.40
0.880	6.82, 6.85
1.10	7.52, 7.52
$k_1\ (\text{sec}^{-1}) = (4.54 \pm 0.05) \times 10^{-3} + (2.73 \pm 0.07) \times 10^{-1}[\text{HA}]^c$	
$k_{\text{H}^+} = 4.43 \times 10^1\ M^{-1}\ \text{sec}^{-1}$	
HA = HCO$_2$H, [HA]/[NaA] = 0.98	
0.78	11.3, 11.4, 11.7, 12.0, 12.0
1.57	12.6, 12.8, 13.0, 13.1
2.35	13.8, 14.0, 14.2, 14.5, 14.6
3.14	15.6, 15.8, 16.1, 16.3
3.92	17.8, 17.8, 18.0
$k_1\ (\text{sec}^{-1}) = (1.05 \pm 0.02) \times 10^{-2} + (1.77 \pm 0.08) \times 10^{-1}[\text{HA}]^c$	
$k_{\text{H}^+} = 4.37 \times 10^1\ M^{-1}\ \text{sec}^{-1}$	
HA = HOCH$_2$CO$_2$H, [HA]/[NaA] = 0.98	
0.78	9.30, 9.56, 9.65
1.56	11.1, 11.2, 11.4, 11.6
2.34	12.1, 12.1, 12.5, 12.6
3.12	13.4, 13.4, 13.5, 13.7, 13.8
3.90	14.7, 14.9, 15.0, 15.3, 15.3
$k_1\ (\text{sec}^{-1}) = (8.77 \pm 0.14) \times 10^{-3} + (1.59 \pm 0.05) \times 10^{-1}[\text{HA}]^c$	
$k_{\text{H}^+} = 4.38 \times 10^1\ M^{-1}\ \text{sec}^{-1}$	
HA = CH$_3$CO$_2$H, [HA]/[NaA] = 0.98	
1.60	1.76, 1.76, 1.77, 1.81
1.99	1.90, 1.91, 1.92, 1.95, 1.97
2.39	2.10, 2.16
3.19	2.38, 2.42
3.99	2.72, 2.79, 2.80
$k_1\ (\text{sec}^{-1}) = (1.11 \pm 0.03) \times 10^{-3} + (4.14 \pm 0.10) \times 10^{-2}[\text{HA}]^c$	
$k_{\text{H}^+} = 4.53 \times 10^1\ M^{-1}\ \text{sec}^{-1}$	
HA = CH$_3$CH$_2$CO$_2$H, [HA]/[NaA] = 4.95	
1.99	4.60, 4.63, 4.75
3.96	5.36, 5.40, 5.50, 5.54
5.93	6.14, 6.14, 6.17
7.91	6.66, 6.73, 6.76, 6.84
9.89	7.27, 7.39, 7.40, 7.40, 7.57
$k_1\ (\text{sec}^{-1}) = (4.14 \pm 0.05) \times 10^{-3} + (3.31 \pm 0.07) \times 10^{-2}[\text{HA}]^c$	
$k_{\text{H}^+} = 4.49 \times 10^1\ M^{-1}\ \text{sec}^{-1}$	
$k_{\text{H}^+} = (4.35 \pm 0.07) \times 10^1\ M^{-1}\ \text{sec}^{-1}$ (average of all buffers)	

[a] Ionic strength = 0.040 M. Error estimates are standard deviations of mean values. [b] Stoichiometric concentrations. [c] Relationship obtained by least-squares fit of observed rate constants, adjusted when necessary to the hydronium ion concentration of the most concentrated buffer solution, to undissociated acid concentrations, also adjusted if necessary.

Brønsted Relations. The carboxylic acid catalytic coefficients determined here are plotted as Brønsted relations in Figure 1. The correlation lines shown are based on least-squares analysis of all of the data in each set except the three pivalic acid points. This acid often deviates appreciably from Brønsted relations based on other carboxylic acids,[8,16] and, since this was also the

Table VIII. Rates of Hydrolysis of Ethyl Cyclohexenyl Ether in Aqueous Buffer Solutions at 25° [a]

10^2[HA], M[b]	10^3k_{obsd}, sec^{-1}
HA = $CNCH_2CO_2H$, [HA]/[NaA] = 0.134	
0.107	28.6, 28.9
0.214	38.1, 38.5
0.268	40.3, 40.9
0.322	42.7, 42.8, 43.0
0.375	44.8, 44.8
k_1 (sec^{-1}) = $(4.16 \pm 0.03) \times 10^{-2} + (1.65 \pm 0.12)$[HA][c]	
$k_{H^+} = 7.39 \times 10^1\ M^{-1}$ sec^{-1}	
HA = $ClCH_2CO_2H$, [HA]/[NaA] = 0.134	
0.107	16.5, 16.6
0.214	20.1, 20.1, 20.1
0.268	21.4, 21.4
0.322	21.9, 22.1
0.377	23.5, 23.6
0.429	23.8, 24.0
k_1 (sec^{-1}) = $(1.84 \pm 0.02) \times 10^{-2} + (1.40 \pm 0.01)$[HA][c]	
$k_{H^+} = 7.70 \times 10^1\ M^{-1}$ sec^{-1}	
HA = $CH_3OCH_2CO_2H$, [HA]/[NaA] = 0.125	
0.100	4.28, 4.32
0.200	4.91, 4.94
0.300	5.30, 5.37, 5.41
0.400	6.00, 6.02
0.500	6.47, 6.54
k_1 (sec^{-1}) = $(3.80 \pm 0.04) \times 10^{-3} + (5.51 \pm 0.12) \times 10^{-1}$[HA][c]	
$k_{H^+} = 8.13 \times 10^1\ M^{-1}$ sec^{-1}	
HA = HCO_2H, [HA]/[NaA] = 0.216	
0.173	4.58, 4.60
0.346	5.29, 5.30
0.520	5.88, 5.90, 5.96
0.693	6.65, 6.74, 6.81
0.866	7.36, 7.38
k_1 (sec^{-1}) = $(4.02 \pm 0.05) \times 10^{-3} (3.88 \pm 0.09) \times 10^{-1}$[HA][c]	
$k_{H^+} = 5.74 \times 10^1\ M^{-1}$ sec^{-1}	
HA = $HOCH_2CO_2H$, [HA]/[NaA] = 0.125	
0.100	2.34, 2.36
0.200	2.68, 2.70
0.300	3.04, 3.05
0.400	3.31, 3.31
0.500	3.72, 3.72
k_1 (sec^{-1}) = $(2.08 \pm 0.02) \times 10^{-3} + (3.24 \pm 0.07) \times 10^{-1}$[HA][c]	
$k_{H^+} = 8.07 \times 10^1\ M^{-1}$ sec^{-1}	
HA = CH_3CO_2H, [HA]/[NaA] = 1.00	
0.415	2.25, 2.26
0.830	2.67, 2.67
1.25	2.96, 2.97
1.66	3.39, 3.41, 3.42
2.08	3.78, 3.83
k_1 (sec^{-1}) = $(1.89 \pm 0.03) \times 10^{-3} + (9.13 \pm 0.20) \times 10^{-2}$[HA][c]	
$k_{H^+} = 8.33 \times 10^1\ M^{-1}$ sec^{-1}	
HA = $CH_3CH_2CO_2H$, [HA]/[NaA] = 1.00	
0.400	1.63, 1.64
0.800	1.96, 1.96
1.20	2.25, 2.25
1.60	2.58, 2.58
2.00	2.82, 2.83
k_1 (sec^{-1}) = $(1.36 \pm 0.02) \times 10^{-3} + (7.46 \pm 0.12) \times 10^{-2}$[HA][c]	
$k_{H^+} = 7.85 \times 10^1\ M^{-1}$ sec^{-1}	
HA = $(CH_3)_3CCO_2H$, [HA]/[NaA] = 1.00	
0.275	1.13, 1.13
0.550	1.43, 1.43
0.825	1.69, 1.69
1.10	1.92, 1.93
1.38	2.21, 2.22
k_1 (sec^{-1}) = $(8.88 \pm 0.11) \times 10^{-4} + (9.62 \pm 0.12) \times 10^{-2}$[HA][c]	
$k_{H^+} = 7.38 \times 10^1\ M^{-1}$ sec^{-1}	
$k_{H^+} = (7.80 \pm 0.13) \times 10^1\ M^{-1}$ sec^{-1} (average of all buffers)	

[a] Ionic strength = 0.040 M except for CH_3CO_2H, $CH_3CH_2CO_2H$, and $(CH_3)_3CCO_2H$ buffer solutions where it was 0.020 M. Error estimates are standard deviations of mean values. [b] Stoichiometric concentrations. [c] Relationship obtained by least-squares fit of observed rate constants, adjusted when necessary to the hydronium ion concentration of the most concentrated buffer solution, to undissociated acid concentrations, also adjusted if necessary.

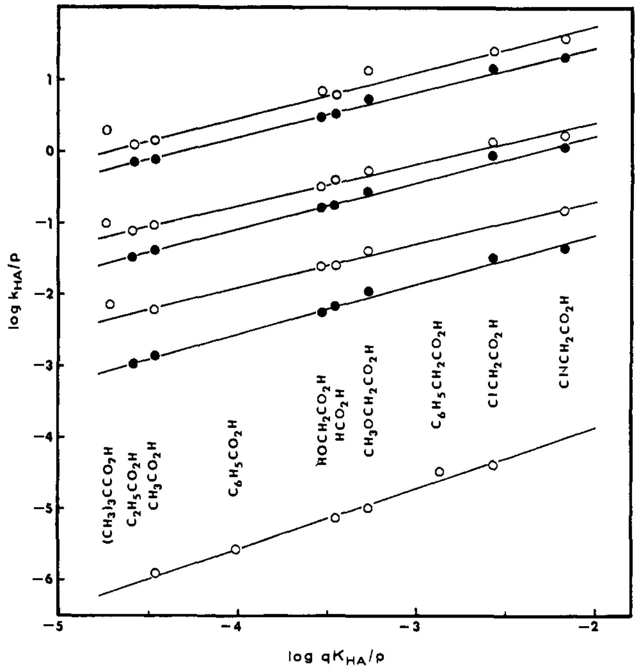

Figure 1. Brønsted relations for the hydrolysis of vinyl ethers; from top to bottom: ethyl isopropenyl ether, ethyl cyclopentenyl ether, ethyl cyclohexenyl ether, methyl cyclohexenyl ether, phenyl isopropenyl ether, ethyl vinyl ether, and phenyl vinyl ether.

case here, these data were not used. Brønsted exponents based on these correlations are listed in Table X.

The reason usually advanced for pivalic acid deviations from Brønsted correlations is hydrophobic bonding between substrate and catalyst.[8,16a] This explanation requires the deviations to increase with increasing bulk of the substrate, and it is curious, therefore, that in the present case the largest deviation occurs with the smallest vinyl ether. The actual differences expressed as per cent deviations from the least-squares lines are: ethyl isopropenyl ether, +90%; phenyl isopropenyl ether, +60%; and methyl cyclohexenyl ether, +45%.

Examination of Figure 1 shows that pivalic acid is not alone in providing systematic departures from the correlation lines. Points for cyanoacetic acid lie below the line in all six reactions for which this catalyst was used, and methoxyacetic acid is a better catalyst than it should be in six out of seven cases; less pronounced systematic deviations also occur with chloracetic acid. These deviations are smaller than the pivalic acid differences noted above; those for cyanoacetic and methoxyacetic acids, for example, average about 20%. This, however, is still four to five times the average experimental standard deviation in catalytic coefficient for each of these acids (as estimated by the goodness of fit in k_1 vs. [HA] correlations), and the high directional consistency of these deviations also argues against their being the results of experimental uncertainty.

Specific effects such as these seem not to have been noticed in Brønsted relations before. Random deviations of the same magnitude, on the other hand, are not uncommon, and this suggests that specific effects

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Table IX. Summary of Catalytic Coefficients

Catalyst	pK_a	Ether, k_{HA}, M^{-1} sec^{-1} Ethyl vinyl	Ethyl isopropenyl	Phenyl vinyl	Phenyl isopropenyl	Ethyl cyclopentenyl	Methyl cyclohexenyl	Ethyl cyclohexenyl
H_3O^+	−1.74	1.75[a]	579	3.28×10^{-3}	5.98	454	42.3	80.0
$CNCH_2CO_2H$	2.47[b]	0.0442	36.9		0.149	20.2	1.13	1.65
$ClCH_2CO_2H$	2.87[c]	0.0324	24.9	4.04×10^{-5}		14.5	0.872	1.40
$C_6H_5OCH_2CO_2H$	3.17[d]			3.24×10^{-5}				
$CH_3OCH_2CO_2H$	3.57[e]	0.0111	13.1	1.02×10^{-5}	0.0402	5.51	0.273	0.551
HCO_2H	3.75[f]	0.00672	5.98	7.44×10^{-6}	0.0252	3.19	0.177	0.388
$HOCH_2CO_2H$	3.83[g]	0.00553	6.83		0.0242	2.94	0.159	0.324
$C_6H_5CH_2CO_2H$	4.31[h]			2.63×10^{-6}				
CH_3CO_2H	4.76[i]	0.00138	1.33	1.21×10^{-6}	0.00604	0.758	0.0414	0.0913
$CH_3CH_2CO_2H$	4.88[j]	0.00104	1.19			0.710	0.0331	0.0746
$(CH_3)_3CO_2H$	5.03[j]		1.89		0.00703			0.0962

[a] A. J. Kresge and Y. Chiang, *J. Chem. Soc. B*, 58 (1967). [b] F. S. Feates and D. J. G. Ives, *ibid.*, 2798 (1956); D. J. G. Ives and P. D. Marsden, *ibid.*, 649 (1965). [c] D. J. G. Ives and J. H. Pryor, *ibid.*, 2104 (1955). [d] N. V. Hayes and G. E. K. Branch, *J. Amer. Chem. Soc.*, **65**, 1555 (1943). [e] E. J. King, *ibid.*, **82**, 3575 (1960). [f] H. S. Harned and N. D. Embree, *ibid.*, **56**, 1042 (1934). [g] L. F. Nims, *ibid.*. **58**, 987 (1936). [h] J. F. Dippy and F. R. Williams, *J. Chem. Soc.*, 161 (1934); G. H. Jeffrey and A. I. Vogel, *ibid.*, 166 (1934); W. L. German and A. I. Vogel, *ibid.*, 912 (1935). [i] H. S. Harned and R. W. Ehlers, *J. Amer. Chem. Soc.*, **55**, 652 (1933). [j] D. H. Everett, D. A. Landsman, and B. R. W. Pinsent, *Proc. Roy. Soc., Ser. A*, **215**, 403 (1952).

Table X. Brønsted Exponents

Ether substrate	α[a]	Log R[b]
Ethyl isopropenyl	0.64 ± 0.04	0.67
Ethyl cyclopentenyl	0.63 ± 0.03	0.39
Ethyl cyclohexenyl	0.58 ± 0.03	−0.58
Methyl cyclohexenyl	0.66 ± 0.03	−0.89
Phenyl isopropenyl	0.61 ± 0.03	−1.73
Ethyl vinyl	0.70 ± 0.03	−2.34
Phenyl vinyl	0.84 ± 0.05	−5.30

[a] Error limits are standard deviations of mean values. [b] R = catalytic coefficient of a hypothetical carboxylic acid with pK_a = 4.00.

may in the past have been masked by experimental uncertainty. This is not unreasonable, for most of the Brønsted relations appearing in the literature refer to reactions which are subject to base catalysis as well as to acid catalysis, and this additional catalytic mode, which of course is absent from vinyl ether hydrolysis, increases the difficulty of extracting accurate individual rate constants from kinetic data.

Systematic deviations of this kind can introduce bias into Brønsted exponents when the number of catalytic coefficients is small and/or the correlation line is short. In order to determine whether the present groups of catalysts were adequate in these respects, correlations were performed for one of the substrates, ethyl vinyl ether, using all possible subsets of the seven available catalytic coefficients. When the points were used in pairs, α ranged from a high of 1.21 (formic and methoxyacetic acids) to a low of 0.34 (chloroacetic and cyanoacetic acids), and its average value was 0.76 ± 0.20 (standard deviation). Using the points in sets of three reduced the average value of α to 0.71 ± 0.09 and provided only one grossly deviant value: $\alpha = 1.17$ for the closely spaced set of catalysts, glycolic, formic, and methoxyacetic acids. When the number of points in a set was greater than three, deviations from mean values were, on the whole, well under 0.10 and standard deviations were also reasonably small: $\alpha = 0.70 \pm 0.04$ (four points per set), $\alpha = 0.70 \pm 0.03$ (five points per set), and $\alpha = 0.70 \pm 0.02$ (six points per set). These calculations suggest that introduction of additional carboxylic acid catalytic coefficients would not alter α appreciably, and that the value based on all of the catalysts, 0.70 ± 0.03, represents the Brønsted exponent for proton transfer from carboxylic acids to ethyl vinyl ether accurately. Since both the identity of the catalysts and the nature of the deviations for ethyl vinyl ether and the other substrates are closely similar, the equivalent is very likely to be true of the other Brønsted exponents listed in column two of Table X.

Discussion

It is convenient, when considering the meaning of Brønsted exponents, to cast the Brønsted relation into a form somewhat different from that shown as eq 2. Consider the general proton transfer process represented by eq 4. The Brønsted relation for this reaction

$$S + HA \rightleftharpoons \left(\overset{\delta+}{S}\text{---}H\text{---}\overset{\delta-}{A}\right)^{\ddagger} \rightleftharpoons SH^+ + A^- \qquad (4)$$

in the forward direction is given by eq 5, but, since the

$$\log k_{HA} = \log G + \alpha \log K_{HA} \qquad (5)$$

equilibrium constant of the overall process, K, is equal to the acidity constant of the catalyzing acid, K_{HA}, divided by the acidity constant of the protonated substrate, K_{SH^+}, eq 5 is equivalent to eq 6. This

$$\log k_{HA} = \log G + \alpha(\log K + \log K_{SH^+}) \qquad (6)$$

expression shows that the Brønsted relation, in addition to correlating rate constants of a proton transfer reaction with acidity constants of the catalysts (eq 5), correlates these rate constants with the equilibrium constants of the proton transfer process itself. It follows from this that the Brønsted exponent α is equal to the effect of a change in catalyst on the free energy of activation of the proton transfer reaction divided by its effect on the free energy of reaction: writing eq 6 for two different catalysts, HA(1) and HA(2), and subtracting one of these expressions from the other leads to eq 7, which is equivalent to eq 8;

$$\log k_{HA(1)} - \log k_{HA(2)} = \alpha(\log K_{(1)} - \log K_{(2)}) \qquad (7)$$

$$\Delta F_{(1)}^{\ddagger} - \Delta F_{(2)}^{\ddagger} = \alpha(\Delta F_{(1)}^{0} - \Delta F_{(2)}^{0}) \qquad (8)$$

$$\alpha = \frac{\Delta F_{(1)}{}^{\ddagger} - \Delta F_{(2)}{}^{\ddagger}}{\Delta F_{(1)}{}^{0} - \Delta F_{(2)}{}^{0}} \qquad (9)$$

$$\alpha = \frac{\delta_R \Delta F^{\ddagger}}{\delta_R \Delta F^{0}} \qquad (10)$$

and the latter, upon rearrangement, gives eq 9. Catalyst changes are usually brought about by changing substituents in the catalyst molecule, and it is convenient, therefore, to reformulate eq 9 in terms of the substituent stabilization operator, δ_R,[18b] as eq 10. A Brønsted exponent, then, is the ratio of substituent effects on the free energy of activation of a proton transfer process to substituent effects on its overall free energy change.

It is readily apparent from this definition of α what conditions must be fulfilled if Brønsted exponents are to be reliable measures of transition state structure. The extent of proton transfer at the transition state of a proton transfer reaction may be conveniently represented by z, the order of the bond being formed between substrate and transferring proton; z is also the fraction of a unit positive charge taken on by the substrate and lost by the catalyst at the transition state. If α is to measure the extent of proton transfer, it must be equal to z; according to eq 10, this requires that $\delta_R\Delta F^{\ddagger} = z\delta_R\Delta F^0$. Since $\Delta F^{\ddagger}$ and ΔF^0 have the same initial limit, and since z must be a fraction, this means that a substituent change in catalyst must alter the free energy of the transition state by an amount intermediate between its effects on the free energies of initial and final states. This would seem to be not unreasonable, for the structure of the transition state will be intermediate between the structures of the initial and final states, and substituent effects on the transition state should therefore also be intermediate between those on initial and final states. Exact correspondence between $\delta_R\Delta F^{\ddagger}$ and $z\delta_R\Delta F^0$, however, requires further that the effect of the substituent on the free energy of the system be linear in z, *i.e.*, that the substituent's perturbation of the free energy of the system change in proportion to the order of the forming bond or the fraction of positive charge transferred. It is likely that this would not be quantitatively true for any real and therefore complicated system; this requirement also leaves no room for effects of the substituent in the transition state which have no counterpart in initial and/or final states, such as the intermolecular interaction between catalyst and substrate recently advanced[17] as an explanation for Brønsted exponents greater than unity and less than zero.[18]

It is likely, in fact, that the systematic deviations from Brønsted relations for vinyl ether hydrolysis visible in Figure 1 have their origin in just such intermolecular effects. At the transition state of the hydrolysis reaction, positive charge is being generated on the oxygen and α-carbon atoms of the vinyl ether, and any polar or dipolar group in the catalyst will interact with this charge. If the interaction is with a group present in every catalyst, such as the partly formed carboxylate anion in a Brønsted correlation based on carboxylic acid catalytic coefficients, it will contribute a uniform amount to each free energy of activation and will therefore have no visible effect upon the correlation. If, on the other hand, interacting groups are present in some catalysts and not in others, this intermolecular effect will show up as systematic deviations from the correlation lines. In the present case, some catalysts contain strongly electronegative groups. Examination of scale models of the hydrolysis transition states shows that conformations are possible in which the electronegative ends of these dipolar groups lie quite close to the substrate atoms bearing positive charge. The fact that the systematic deviations do occur only for acids containing electronegative substituents, and that the deviations are for the most part positive in direction as expected for an attractive and therefore energy-lowering interaction, suggests that such interactions are in fact taking place. There will, of course, be individual differences in the magnitude of these interactions, and the weaker ones will appear as negative deviations from a correlation line dominated by the more strongly interacting systems.

Since electronegative groups also increase the acidity of carboxylic acids, these deviations concentrate near the strong-acid ends of these Brønsted relations. They thus have the effect of raising α, making it greater than z. It is difficult to estimate the magnitude of this increase, for evaluation of the intermolecular interactions is not straightforward. The "nonreacting" ends of these transition states can exist in any of a number of juxtaposed positions with different intercharge distances, and there will be entropy effects in addition to the electrostatic energy differences. It seems possible only to say that the energies involved are likely to be small, in keeping with the difference between α and z inferable from isotope effect studies. The more reactive of the vinyl ethers investigated here show maximum values of the isotope effect $k_{H_3O^+}/k_{D_3O^+}$;[19] presumably, therefore, they have symmetrical transition states with $z = 0.5$. For the same vinyl ethers, α is 0.6–0.65 (Table X). Since the range of catalyst acidities used in these Brønsted relations gives an upper limit of 3.5 kcal/mol to $\delta_R\Delta F^0$, the contribution to $\delta_R\Delta F^{\ddagger}$ which makes α exceed z by 0.1–0.15 cannot be greater than 400–500 cal/mol.

Further evidence for the existence of intermolecular electrostatic interactions such as these may be found in deviations from these and other Brønsted relations produced by catalysts bearing charged groups. Negatively charged catalysts should undergo an additional attractive interaction with substrates to which a proton (and positive charge) is being transferred, and it is significant, therefore, that bisulfate ion and monoanions of carboxylic acids generally provide positive deviations from Brønsted relations based on neutral catalysts (carboxylic acids) for reactions such as these.[9,20] Similarly, positively charged catalysts will experience an energy-raising interaction with substrates receiving positive charge, and the negative deviations often shown by the ammonium ion and especially the hydronium ion may be understood in this way.[6,8,9,16b,20,21]

An alternative explanation[22] is sometimes advanced for the large negative deviations of hydronium ion

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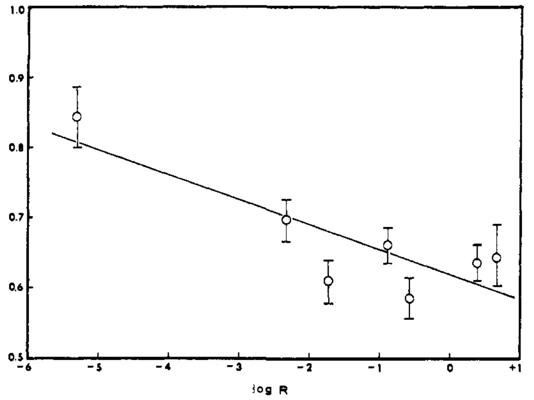

Figure 2. Relationship between Brønsted exponents and relative reactivity in the hydrolysis of vinyl ethers; from left to right: phenyl vinyl ether, ethyl vinyl ether, phenyl isopropenyl ether, methyl cyclohexenyl ether, ethyl cyclohexenyl ether, ethyl cyclopentenyl ether, and ethyl isopropenyl ether.

catalytic coefficients from Brønsted relations based on carboxylic acids. This hypothesis makes use of the fact that the molar concentration of water molecules in pure water, 55 *M*, enters into the evaluation of the conventional acidity constant $pK_a = -1.74$ for the species H_3O^+ in aqueous solution; it suggests that the concentration of free, *i.e.*, non-hydrogen-bonded, water might better be used instead. This would have the effect of lowering the acidity constant of H_3O^+ and moving a negatively deviating catalytic coefficient toward the correlation line. Corollary to this explanation is the suggestion that the true acidity constant of H_3O^+ might be evaluated by fitting hydronium ion catalytic coefficients to Brønsted correlation lines.[20b,21a] The seven reactions investigated in the present study give 0.2 ± 0.1 (standard deviation of mean value) as the pK_a of H_3O^+ estimated in this way. This agrees moderately well with $pK_{H_3O^+} = 0.5$ obtained in the same manner using data for the detritiation of 1,3,5-trimethoxybenzene,[20b] but it is in poor agreement with the value 1.2 deduced from data for the hydrolysis of cyanoketene dimethyl acetal.[16b] The presently derived values, moreover, encompass the considerable range 0.7 (ethyl isopropenyl ether) to −0.2 (ethyl cyclohexenyl ether), which, together with the cyanoketene dimethyl acetal result, show a rather wide variation from reaction to reaction. This spread suggests that other factors, *e.g.*, intermolecular effects, are operative and that the pK_a of H_3O^+ might better be deduced from the fit of hydronium ion catalytic coefficients to Brønsted relations based not on carboxylic acids but rather on positively charged catalysts.

These considerations lead to the conclusion that Brønsted exponents, even if based on a homogeneous set of catalysts such as the carboxylic acids used here, are rather imprecise measures of transition state structure, for systematic intermolecular effects can easily produce a difference of 0.1 or more between α and the degree of proton transfer at the transition state (z). Such intermolecular effects, however, might be expected to be closely similar for a series of similar substrates undergoing the same reaction, and Brønsted exponents along such a series, while not being numerically equal to z, might still provide a good relative measure of transition state structure. It is of interest, therefore, to compare the values of α obtained here with relative reactivity, another property thought to measure transition state structure in a relative way.[23] Such a comparison is shown in Figure 2 as a plot of α *vs.* log R (Table X),[15b] where R is the catalytic coefficient of a hypothetical acid of $pK_a = 4.00$ obtained by interpolating the Brønsted relations of Figure 1. The expected trend of decreasing α with increasing reactivity is clearly visible, but individual variations do stand out and r (correlation coefficient) is only 0.844. This poor fit very probably has its origin in the fact that, whereas intermolecular effects which change from catalyst to catalyst can drop out of such a comparison, new effects which change from substrate to substrate and influence reactivity without affecting z can be introduced (through R). For example, conformational effects make cyclohexenyl ethers less reactive than either cyclopentenyl ethers or analogously substituted open-chain vinyl ethers,[24] and shifting the points for cyclohexenyl methyl and ethyl ethers to the right in Figure 2 to allow for this effect does improve the correlation.

Even though the correlation of Figure 2 is poor, it does serve to illustrate the fact that α is not a very sensitive function of reactivity. The data shown encompass six powers of ten in reaction rate and go beyond the range accessible by conventional kinetic methods; yet this changes α by little more than 0.2. This dependence of Brønsted exponent on reactivity is somewhat weaker than that observed in the base-catalyzed halogenations of ketonic substances,[15c] but the difference is really hardly significant: least-squares analysis of both sets of data gives $d\alpha/d \log R = -0.036 \pm 0.010$ for vinyl ether hydrolysis and $d\beta/d \log R = -0.051 \pm 0.06$ for the halogenation reactions.

It is of interest also that a correlation similar to that of Figure 2 but employing hydronium ion catalytic coefficients predicts a value of 0.86 ± 0.06 for α in the hydration of isobutene. General acid catalysis could not be detected in this reaction,[25] but z has been estimated as 0.84[9] and 0.9[26] on the basis of solvent isotope effects.[26] This good agreement is somewhat surprising, for it was concluded above that the values of α upon which this prediction is based overestimate z by at least 0.1. This suggests that proton transfer to vinyl ethers and ordinary olefins may not conform exactly to a single correlation of α with reactivity.

Experimental Section

Materials. Cyclic vinyl ethers were prepared from the corresponding ketals by *p*-toluenesulfonic acid catalyzed elimination of 1 equiv of alcohol;[27] ethyl cyclopentenyl ether was also obtained by passing the ketal through a heated gas chromatography column.[28] Ethyl and phenyl isopropenyl ethers were synthesized through decarboxylation of β-ethoxy- and β-phenoxycrotonic acids[29] and phenyl vinyl ether was prepared by elimination of hydrogen chloride from β-chloroethyl phenyl ether.[30] Phenyl-*4-t* vinyl ether was

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(30) V. L. Tweedie and B. G. Barrow, *J. Org. Chem.*, **25**, 2023 (1960).

synthesized by treating the Grignard reagent obtained from *p*-bromophenyl vinyl ether,[31] prepared by eliminating hydrogen bromide from β-bromoethyl *p*-bromophenyl ether, with tritiated water. Ethyl vinyl ether was obtained commercially.

The physical properties of these vinyl ethers were in good agreement with literature values; their identity was further confirmed by their nmr spectra. Kinetic samples were purified usually by fractional distillation and occasionally by gas chromatography; purity was monitored by gas chromatography. All other reagents were the best available commercial grades.

Kinetics. Spectroscopic Methods. Reactions with half-lives greater than *ca.* 20 sec were conducted in 1-cm silica cells contained in the sample compartment of a Cary Model 11 spectrometer; thermostating was provided by water from a constant-temperature bath (25.0 ± 0.1°) circulating through a Cary thermostatable cell jacket. Both sample and reference cells were filled with buffer or perchloric acid solution, and sufficient time was allowed for the sample solution to come to temperature; reaction was then initiated by adding vinyl ether, in quantity sufficient to give a final concentration *ca.* 10^{-3} *M*, to the sample cell and shaking vigorously to effect solution. Absorbance at a fixed wavelength in the region 210–230 nm was recorded continuously in the case of faster reactions and at 5- to 10-min intervals in the case of slower runs. Readings were continued for 4–5 half-lives and infinity values were measured after 8–10 half-lives. In some of the slower runs, infinity readings were not made and the data were treated by the method of Guggenheim[32] and/or Swinbourne.[33] Rate constants were evaluated graphically.

For reactions with half-lives shorter than *ca.* 20 sec, a commercial Durrum–Gibson stopped-flow spectrometer was used. This machine has a mixing time of 2 msec; the fastest reactions measured here had half-lives of 70 msec. Thermostating was provided by constant-temperature water (25.0 ± 0.05°) circulating through the reservoir containing the driving syringes and through the mixing and observation block. Substrate solution and buffer or perchloric acid solution were allowed to remain in the driving syringes long enough to equilibrate with the constant temperature fluid; they were then mixed, and changes in transmittance at 220 nm were recorded on a storage oscilloscope. This record was then photographed with a Polaroid camera, and transmittance changes were read off the developed photograph. Transmittance was converted to absorbance, and first-order rate constants were evaluated graphically.

Radiochemical Method. Aqueous solutions of phenyl-*4-t* vinyl ether and buffer, whose combined volume totaled 100 ml, were allowed to equilibrate separately with a constant-temperature bath operating at 25.0 ± 0.02°. These solutions were then mixed, the reaction mixture was replaced in the bath, and 5-ml samples were removed by pipet at appropriate time intervals (all corresponding to less than 3% reaction). These samples were immediately quenched in 2.00-ml portions of 1 *M* aqueous sodium hydroxide, and the quenched aliquots were then extracted with three successive 25-ml portions of toluene; control experiments showed that each extraction removed 99% of the vinyl ether remaining in the aqueous phase. Aliquots (3 ml) of these thrice-washed aqueous solutions were then added to 10.0-ml portions of Bray's counting solution,[34] and radioassay was performed with a Packard Tri-Carb Model 314 EX liquid scintillation spectrometer. Infinite-time radioactivities were determined by adding 10.00 ml of 0.2 *M* perchloric acid to 10.00-ml portions of reaction mixture, allowing the resultant solutions to stand at 25° for at least 6 hr (10 half-lives at 0.1 *M* acid) and then treating 5.00-ml aliquots in the usual way.

Zero-order rate constants were evaluated as slopes of plots of counts per minute *vs.* time; these were converted to first-order rate constants by dividing by the average cpm of the infinite time samples times two (the correction for perchloric acid dilution).

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Hexafluoroacetone Azine. Reaction with Cyclohexane by a Radical Double Chain Mechanism

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Abstract: The thermal reaction of hexafluoroacetone azine with cyclohexane gave three principal products: [2,2,2-trifluoro-1-(trifluoromethyl)ethyl]cyclohexane (**5**), the azo compound **7**, and the hydrazone **8**. Evidence is presented for a unique double radical chain mechanism that involves the intermediate formation of bis(trifluoromethyl)diazomethane (**4**) to account for these products.

We reported earlier that hexafluoroacetone azine (**1**) reacts thermally with cyclohexane to give the carbene insertion product, [2,2,2-trifluoro-1-(trifluoromethyl)ethyl]cyclohexane (**5**).[1] Since both bis(trifluoromethyl)diazomethane and bis(trifluoromethyl)diazirine react with cyclohexane to give primarily nitrogen-containing products, we concluded that the azine may be a better source of bis(trifluoromethyl)carbene than either the diazomethane or the diazirine in reactions with saturated hydrocarbons. Recently, Forshaw and Tipping have shown that the azine **1** also reacts with olefins such as cyclohexene to give products (cyclopropanes) that would be expected from carbene reactions.[2] However, they concluded that initial formation of bis(trifluoromethyl)carbene followed by attack of this carbene on the olefin is unlikely, since the azine is stable to 240°, and, when it is pyrolyzed at 570°, it does not form products expected from the intermediacy of the carbene.

We have reexamined the reaction of the azine **1** with cyclohexane in the temperature range 129–170° and found that other products in addition to **5** are formed, including the azo compound **7** and the hydrazone **8**. We now concur with Forshaw and Tipping that the azine **1** does not react initially to give the carbene. Instead, a

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