fluoride solution,[22] and, although these values cannot be compared directly with the pK_a's evaluated here because the standard states are different, one scale can be translated into the other by virtue of the fact that the pK_a of hexamethylbenzene has also been determined in aqueous sulfuric acid.[26] The pK_B values of benzene and hexamethylbenzene are 9.2 and -1.4, respectively, and the pK_a of hexamethylbenzene is -13.3; adding the pK_B difference between benzene and hexamethylbenzene to -13.3 then gives -23.9 as the pK_a of benzene. This, combined with the previous two values, suggests pK_a = -23 ± 1 as the best estimate of the basic strength of benzene.

The estimates made above of pK_a = -14 and -15 for phenol and anisole protonated on carbon in the para position are of interest in relation to an earlier study of these substances whose results were interpreted in terms of protonation on oxygen.[27] These data indicated that these oxonium ions were formed in 75–80% sulfuric acid, a region somewhat below the 90–95% acid which the presently estimated pK_a's imply to be necessary for carbon protonation. Thus, in a sense, the oxygen atom of phenol and anisole is a more basic site than ring carbon, at least in aqueous solution. As the acidity of the medium is increased, however, these first-formed oxygen-protonated species will give way to carbon-protonated conjugate acids, for protonation on oxygen will have a shallower acidity dependence (*ca.* H_0)[27] than protonation on carbon (*ca.* H_c). In very strongly acidic media, therefore, carbon will be the more basic site.

Experimental Section

Materials. The substrates used were also employed in the study described in the preceding paper;[1b] their synthesis is detailed there. Perchloric acid more concentrated than 70 wt % was prepared by mixing commercial 70% acid with perchloric acid monohydrate (85% acid) which was purchased from the G. Frederick Smith Chemical Co. All other materials were best available commercial grades and were used without further purification.

Indicator Measurements. Sample solutions were prepared by allowing 25-ml aliquots of aqueous acid to equilibrate with a constant temperature bath operating at 25.0 ± 0.1° and then adding 50 μl of *ca.* 10^{-1} *M* aqueous or methanolic stock solution of aromatic base. These 50-μl aliquots were delivered using Lang-Levy micropipets, whose precision, as determined by titration of samples of acid dispensed from them, was better than ±0.2%.

The absorbance of sample solutions was measured against references consisting of acid at the same concentration using a spectrometer (Beckman DK-2) whose cell compartment was thermostated at 25.0 ± 0.1°. (The temperature coefficient of absorbance was determined in a few cases and proved always to be less than 1% per degree.) For solutions whose absorbance changed with time, the time-scan mode of the spectrometer was employed and absorbance was extrapolated linearly back to the time of mixing aromatic base with acid. Replicate measurements on the same solution showed a reproducibility (standard deviation) better than 0.01 absorbance unit in all cases, and better than 0.003 unit for solutions whose absorbance did not change with time.

Acid concentrations were determined as weight per cent by titration with standard base, and values of acidity functions were interpolated using polynomial expressions relating the acidity function to weight per cent acid.

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[26] N. C. Deno, P. T. Groves, and G. Saines, *J. Amer. Chem. Soc.*, **81**, 5790 (1959).

[27] E. M. Arnett and C. Y. Wu, *ibid.*, **82**, 5660 (1960).

Aromatic Protonation. IX. The Kinetic Protonation of Hydroxy- and Alkoxybenzenes[1]

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Abstract: Rates of aromatic hydrogen exchange (protodetritiation) of 1,3,5-trihydroxybenzene and its trimethyl and triethyl ethers, 1,3-dimethoxybenzene, 1,3-dihydroxy-2-methylbenzene and its dimethyl ether, and anisole were measured in concentrated aqueous perchloric acid solutions. The kinetic acidity dependence of these reactions varies markedly from substrate to substrate, but the variations can be understood in terms of two factors: (1) differences in substrate structure, and (2) differences in the extent of proton transfer at transition states. Suitable treatment of these kinetic data, together with the results of equilibrium protonation of the same substrates presented in the previous paper, gives values of α_A, a parameter which measures the extent of transition state proton transfer, which range from 0.5 to 0.9. These same data provide a curved Brønsted relation whose slopes, α_B, are in excellent agreement with the α_A values. This identification of transition state as well as substrate structure as important factors governing kinetic acidity dependence in concentrated acids implies certain limitations on the use of kinetic acidity dependence as a criterion of reaction mechanism.

In the preceding paper of this series,[1b] we described a study of the equilibrium protonation of a group of hydroxy- and alkoxy-substituted benzenes related to phloroglucinol, resorcinol, and phenol. We now add to that an investigation of the kinetics of some of these protonation reactions. The results, just as those of the companion study, in addition to providing specific information about the substances examined, afford insight into the general nature of concentrated acid solu-

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[2] Fellow of the U. S. Steel Foundation, 1962–1964.

tions. In particular, they illustrate how the equilibrium and kinetic acidity dependence of protonation reactions may be combined to provide information on transition-state structure, and, in doing so, show that differences in the extent of proton transfer at the transition state must be taken into account in interpreting kinetic acidity dependence in concentrated acid solutions.

Aromatic hydrogen exchange, in the form of protodetritiation (eq 1), was used to follow the kinetics of

$$\mathrm{TAr} + \mathrm{H}^+ \longrightarrow \mathrm{HAr} + \mathrm{T}^+ \qquad (1)$$

these protonation reactions. The mechanism of this process is now well established as simple protonation–deprotonation of the aromatic substrate, with a benzenonium ion, TArH^+, as the single reaction intermediate (eq 2).[3–5] Since this is an isotope exchange reac-

$$\mathrm{TAr} + \mathrm{H}^+ \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} \mathrm{TArH}^+ \xrightarrow{k_2} \mathrm{T}^+ + \mathrm{ArH} \qquad (2)$$

tion, its two transition states must be of comparable free energy, and neither step can be fully rate determining; observed rate constants therefore contain the isotope effect k_{-1}/k_2 (*i.e.*, $k_\mathrm{H}/k_\mathrm{T}$) in addition to the rate constant for the protonation step, k_1 (eq 3). This iso-

$$k_\mathrm{obsd} = k_1/(1 + k_{-1}/k_2) \qquad (3)$$

tope effect, however, will change only slowly, if at all, with changes in the acidity of the medium,[4h] and the acidity dependence of protonation, *i.e.*, that of k_1, will be essentially that found for k_obsd.

Results

Rates of detritiation were measured by using liquid scintillation counting to monitor the decrease in radioactivity of tritium-labeled substrates. This method, as employed by us[3] and others,[4,5] involves extracting the aromatic substrate from wholly aqueous, quenched aliquots of the reaction mixture. It is convenient to use toluene or xylene as the solvent for this extraction, for these substances are also good media for liquid scintillation counting. Two of the present substrates, 1,3,5-trihydroxybenzene and 1,3-dihydroxy-2-methylbenzene, proved, however, to be too polar to allow efficient removal from water by solvents such as these, and butyl acetate was therefore used instead. Its presence lowered the efficiency of the radioassay for these two substances, but that difficulty was overcome by labeling the substrates to a higher level of radioactivity. In all cases, initial counting rates were at least 10^4 cpm and reactions were routinely followed for 8–10 half-lives; this produced good precision in rate constant, *i.e.*, ±1% or better.

Wholly aqueous perchloric acid was used as the acidic medium in all cases, and measurements were carried out on eight substrates: 1,3,5-trihydroxybenzene and its trimethyl and triethyl ethers, 1,3-dimethoxybenzene labeled in both the 2 and 4 position, 1,3-dihydroxy-2-methylbenzene and its dimethyl ether both labeled in the 4 position, and anisole also labeled in the 4 position. Of these, the first three were too reactive to permit sampling by ordinary pipet in solutions more concentrated than *ca.* 1 *M* acid. This seriously limited the information available, for different acidity dependences are still only weakly differentiated in perchloric acid of this strength. A gas-driven pipet was therefore devised. This reduced sampling times by an order of magnitude and allowed measurements on trimethoxy- and triethoxybenzene to be made up to 3 *M* perchloric acid. Trihydroxybenzene itself is an order of magnitude more reactive than these ethers, and in its case it was also necessary to lower the temperature to get meaningful results. These rate measurements are summarized in Table I.

Table I. Rate Constants for Aromatic Detritiation Measured in Aqueous Perchloric Acid

Wt % acid	*M*	$10^3 k_\mathrm{obsd}$, sec^{-1}	Wt % acid	*M*	$10^3 k_\mathrm{obsd}$, sec^{-1}
1,3,5-Trihydroxybenzene-*2-t*[a]					
9.50	0.993	16.1	23.70	2.665	86.8
9.50	0.993	15.5	23.70	2.665	83.2
13.70	1.483	28.3	26.20	3.020	106.3
13.70	1.483	27.7	26.20	3.020	105.2
18.80	2.079	47.7	28.10	3.250	130.3
20.90	2.324	58.3	28.10	3.250	133.0
1,3,5-Trimethoxybenzene-*2-t*[b]					
7.93	0.824	9.30	17.44	1.918	43.7
10.41	1.097	14.75	19.72	2.200	61.3
12.81	1.369	21.0	21.86	2.473	82.7
15.21	1.650	31.2	23.92	2.742	108.8
1,3,5-Triethoxybenzene-*2-t*[b]					
5.84	0.599	8.32	13.96	1.053	37.7
8.65	0.902	17.0	16.46	1.799	58.8
9.54	1.000	20.2	18.88	2.095	74.2
12.78	1.366	32.2	22.59	2.567	118.7
1,3-Dimethoxybenzene-*2-t*[c]					
18.75	2.079	0.0202	37.28	4.694	0.455
26.91	3.147	0.0680	40.35	5.200	0.864
33.99	4.177	0.236	46.26	6.247	3.70
1,3-Dimethoxybenzene-*4-t*[c]					
9.94	1.045	0.0223	33.95	4.171	1.180
18.82	2.089	0.0970	37.10	4.664	2.26
26.82	3.135	0.335	40.21	5.177	4.44
1,3-Dihydroxy-2-methylbenzene-*4-t*[d]					
9.05	0.946	0.367	32.96	4.021	10.03
17.93	1.978	1.373	39.09	4.989	24.3
24.80	2.860	3.43	45.19	6.052	68.5
1,3-Dimethoxy-2-methylbenzene-*4-t*[b]					
4.70	0.479	0.00737	35.56	4.421	2.08
9.17	0.959	0.0217	37.06	4.659	3.05
17.36	1.908	0.0873	40.14	5.165	4.95
18.40	2.036	0.1058	44.51	5.928	16.83
25.60	2.968	0.337	49.53	6.875	84.0
32.20	4.057	1.240			
Anisole-*4-t*[c]					
43.64	5.77	0.00263	57.46	8.54	0.365
49.78	6.92	0.0160	60.50	9.24	1.46
55.36	8.08	0.133			

[a] 1.40°. [b] 25.10°. [c] 24.69°. [d] 23.90°.

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(3) A. J. Kresge and Y. Chiang, *J. Amer. Chem. Soc.*, **81**, 5509 (1959); (b) *ibid.*, **83**, 2877 (1961); (c) *Proc. Chem. Soc.*, 81 (1961); (d) *J. Amer. Chem. Soc.*, **84**, 3976 (1962); (e) *ibid.*, **89**, 4411 (1967); (f) A. J. Kresge, Y. Chiang, and Y. Sato, *ibid.*, **89**, 4418 (1967); (g) A. J. Kresge, S. Slae, and D. W. Taylor, *ibid.*, **92**, 6309 (1970).

(4) J. Colapietro and F. A. Long, *Chem. Ind.* (*London*), 1056 (1960); (b) F. A. Long and J. Schulze, *J. Amer. Chem. Soc.*, **83**, 3340 (1961); (c) *ibid.*, **86**, 322, 327, 331 (1964); (d) R. J. Thomas and F. A. Long, *ibid.*, **86**, 4770 (1964); (e) *J. Org. Chem.*, **29**, 3411 (1964); (f) B. C. Challis and F. A. Long, *J. Amer. Chem. Soc.*, **85**, 2524 (1963); (g) L. C. Gruen and F. A. Long, *ibid.*, **89**, 1287 (1967); (h) J. L. Longridge and F. A. Long, *ibid.*, **89**, 1292 (1967).

(5) R. D. Batts and V. Gold, *J. Chem. Soc.*, 4284 (1964).

Table II. Slopes of Linear Correlations of Log k with H_0 and H_c

Substrate	$-\mathrm{d}\log k/\mathrm{d}H_0$[a,b]	$-\mathrm{d}\log k/\mathrm{d}H_c$[a,c]
1,3,5-Trihydroxybenzene-*2-t*	0.759 ± 0.009	0.456 ± 0.008
1,3,5-Trimethoxybenzene-*2-t*	0.984 ± 0.012	0.627 ± 0.010
1,3,5-Triethoxybenzene-*2-t*	0.932 ± 0.024	0.619 ± 0.032
1,3-Dimethoxybenzene-*2-t*	1.088 ± 0.013	0.571 ± 0.010
1,3-Dimethoxybenzene-*4-t*	1.117 ± 0.010	0.621 ± 0.010
1,3-Dihydroxy-2-methylbenzene-*4-t*	0.859 ± 0.019	0.472 ± 0.023
1,3-Dimethoxy-2-methylbenzene-*4-t*	1.126 ± 0.009	0.638 ± 0.013
Anisole-*4-t*	1.037 ± 0.010	0.696 ± 0.018

[a] Error limits are standard deviations of mean values. [b] The H_0 values were obtained from A. J. Kresge and H. J. Chen, to be published, and K. Yates and H. Wai, *J. Amer. Chem. Soc.*, **86**, 5408 (1964). Since these differ somewhat from previous values, some of the slopes reported here are slightly different from the values published in a preliminary account (A. J. Kresge, R. A. More O'Ferrall, L. E. Hakka, and V. P. Vitullo, *Chem. Commun.*, 46 (1965)) of this research. [c] The H_c values were obtained from ref 1b and M. T. Reagan, *J. Amer. Chem. Soc.*, **91**, 5506 (1969).

Discussion

Acidity Dependence. Each of the hydrogen exchange reactions examined here shows an acidity dependence considerably steeper than simple proportionality between observed rate constant and hydrogen ion concentration. All eight sets of data, on the other hand, are correlated well by the H_0 and H_c acidity functions in the sense that plots of log k_{obsd} against either of these acidity scales are accurately linear. The correlation coefficients for these linear relationships all lie in the range 0.992–0.999, and it is difficult to say whether H_0 or H_c represents the data better. This is unlike the situation for the equilibrium protonation of some of these same substrates, where H_c is decidedly superior to H_0 in providing linear relationships.[1b] Most of the present measurements, however, refer to acid concentrations outside the region 6–7 *M*, which is the only region of aqueous perchloric acid where H_c is a markedly nonlinear function of H_0,[1b] and where, therefore, the ability of these two scales to provide linear correlations would be different.

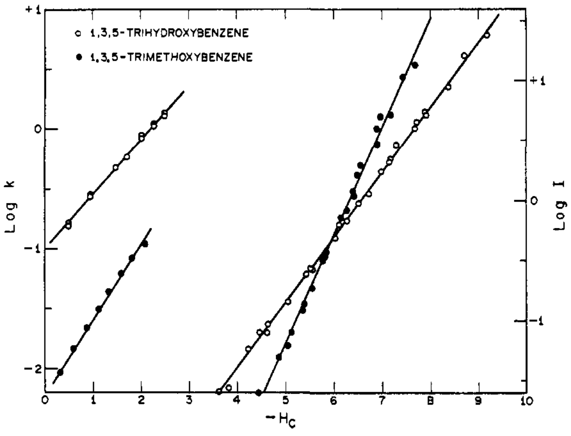

Figure 1. Comparison of the kinetic acidity dependence of the protodetritiation of 1,3,5-trihydroxybenzene-*2-t* and 1,3,5-trimethoxybenzene-*2-t* (left side of figure) with the acidity dependence of the equilibrium protonation of these same substances (right side of figure).

These kinetic protonations, however, do resemble their equilibrium counterparts in that they provide acidity function correlations whose slopes change from substrate to substrate. Table II shows that values of d log $k_{\mathrm{obsd}}/\mathrm{d}H_0$ for the eight sets of data, evaluated by the method of least squares, range from 0.76 to 1.13, and values of d log $k_{\mathrm{obsd}}/\mathrm{d}H_c$, from 0.46 to 0.70. Here, just as in the equilibrium process, moreover, alkoxy-substituted aromatics have steeper acidity dependences than do the corresponding hydroxy compounds. This is illustrated in Figures 1 and 2 where kinetic and equilibrium data are compared for the two pairs of substrates which are most closely related in this way: 1,3,5-trihydroxybenzene and 1,3,5-trimethoxybenzene (Figure 1) and 1,3-dihydroxy-2-methylbenzene and 1,3-dimethoxy-2-methylbenzene (Figure 2). In each case, the methoxy compound has the steeper kinetic acidity dependence by a considerable margin and the steeper equilibrium acidity dependence as well. This suggests that the factor which was identified as the principal influence on the acidity dependence of equilibrium protonation in these systems, *i.e.*, hydrogen bonding between the solvent and positively charged hydroxyl groups in benzenonium ion conjugate acids,[1b] must also be operative in the kinetic process.

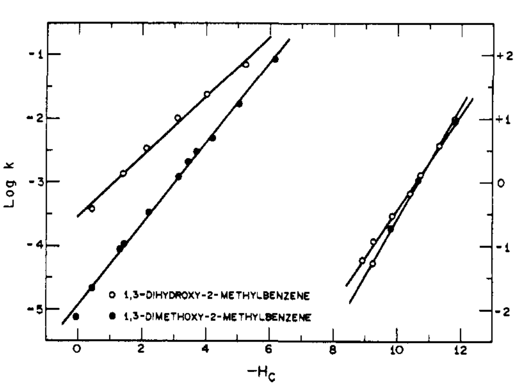

Figure 2. Comparison of the kinetic acidity dependence of the protodetritiation of 1,3-dihydroxy-2-methylbenzene-*4-t* and 1,3-dimethoxy-2-methylbenzene-*4-t* (left side of figure) with the acidity dependence of the equilibrium protonation of these same substances (right side of figure).

It does not necessarily follow from this that the kinetic and equilibrium acidity dependences for protonation of a given substrate will be identical, for the transition state in this process may lie only part way along the reaction coordinate between reactants and products; the difference in solute–solvent interactions between reactants and products, which governs the equilibrium acidity dependence, will then be only partly realized in the corresponding difference involving reactants and

Table III. Correlation of Kinetic and Equilibrium Aromatic Protonation

Substrate	α_A[a]	pK_a[b]	log k_{H^+}[c]	α_B[d]
1,3,5-Trimethoxybenzene	0.554 ± 0.011	−5.7	−2.21[e]	0.56
1,3,5-Triethoxybenzene	0.539 ± 0.044	−4.8	−1.90	0.55
1,3-Dimethoxybenzene (2 position)	0.604 ± 0.007		−5.77	
1,3-Dimethoxybenzene (4 position)	0.639 ± 0.004	−9.0	−5.10[f]	0.62
1,3-Dihydroxy-2-methylbenzene	0.539 ± 0.020	−7.5	−3.45	0.59
1,3-Dimethoxy-2-methylbenzene	0.652 ± 0.009	−9.3	−5.07	0.62
Anisole	0.706 ± 0.024	−15.3	−8.8	0.72
Benzene	0.930 ± 0.041[g]	−23.0	−15.2[g]	0.87
Azulene	0.545 ± 0.014[h]	−1.7[i]	−0.74[j]	0.47
Guaiazulene 2-sulfonate		−0.6[k]	−0.07[k]	0.47
4,6,8-Trimethulazulene		+0.5[k]	+0.60[k]	0.45
Guiazulene		+1.5[k]	+0.79[k]	0.43

[a] (d log k_{obsd}/C_{H^+})/(d log I/C_{H^+}). [b] From A. J. Kresge, H. J. Chen, L. E. Hakka, and J. E. Kouba, *J. Amer. Chem. Soc.*, **93**, 6174 (1971) unless indicated otherwise. [c] Specific rates of protodetritiation except for benzene where k_{H^+} refers to protodedeuteration. [d] d log k_{H^+}/dpK_a. [e] A. J. Kresge and Y. Chiang, *J. Amer. Chem. Soc.*, **89**, 4418 (1967). [f] A. J. Kresge, Y. Chiang, and Y. Sato, *ibid.*, **89**, 4418 (1967). [g] Based on kinetic data provided by D. P. N. Satchell, *J. Chem. Soc.*, 3911 (1956). [h] Based on kinetic (at 7.3°) and equilibrium (at 5.7°) data provided by B. C. Challis and F. A. Long, *J. Amer. Chem. Soc.*, **87**, 1196 (1965). [i] M. T. Reagan, *ibid.*, **91**, 5506 (1969). [j] R. J. Thomas and F. A. Long, *ibid.*, **86**, 4770 (1964). [k] J. L. Longridge and F. A. Long, *ibid.*, **89**, 1292 (1967).

transition state. Comparison of kinetic and equilibrium acidity dependences shows that this is in fact the case. For example, the two substrates shown in Figure 1 have kinetic and equilibrium slopes, as measured against H_c, of 0.46 and 0.55 (1,3,5-trihydroxybenzene) and 0.63 and 0.92 (1,3,5-trimethoxybenzene), and the corresponding numbers for the substrates shown in Figure 2 are 0.47 and 0.73 (1,3-dihydroxy-2-methylbenzene) and 0.64 and 0.88 (1,3-dimethoxy-2-methylbenzene). Similar relationships occur for all of the other substrates examined.

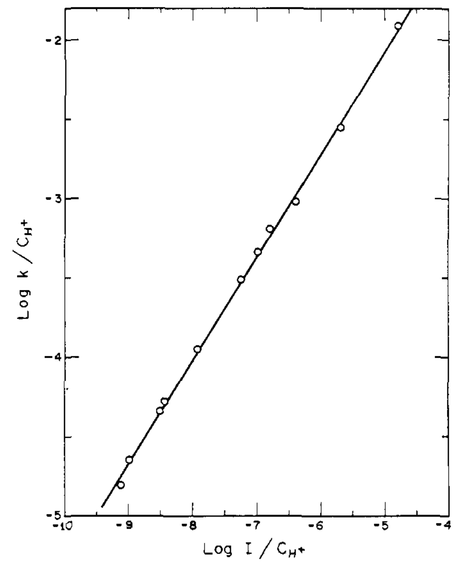

Figure 3. Data for the kinetic and equilibrium protonation of 1,3-dimethoxy-2-methylbenzene plotted in the form of eq 4.

It is of interest to inquire into the quantitative nature of these relationships between kinetic and equilibrium acidity dependence, in particular to determine whether one is a constant fraction of the other or whether the proportion changes from substrate to substrate. This may be done by using a formalism which we developed in a preliminary account of this research.[6] Our treatment expressed the solute–solvent interactions which govern the kinetic process as some fraction, α_A,[7] of the corresponding interactions governing the equilibrium process; stated in terms of activity coefficients, this is $f^{\ddagger}/f_Sf_{H^+} = (f_{SH^+}/f_Sf_{H^+})^{\alpha_A}$, where S = substrate and the other symbols have their usual meanings. This leads to the relationship of eq 4, from which it may be seen that

$$\log k_{obsd}/C_{H^+} = \log k^{\circ}(K_{SH^+})^{\alpha_A} + \alpha_A \log C_{SH^+}/C_SC_{H^+} \quad (4)$$

a plot of log k_{obsd}/C_{H^+} against log I/C_{H^+} (I = indicator ratio = C_{SH^+}/C_S) will be linear as long as α_A is constant, and that α_A may be evaluated as the slope of this plot; its intercept is a function of the medium-independent rate constant k° (rate/$C_{H^+}C_S$ in dilute acid solution) and the acidity constant of the substrate conjugate acid, K_{SH^+}. This method, in essence, places the kinetic acidity dependence on a scale which extends from simple proportionality to C_{H^+} as its lower limit to an acidity dependence identical with that of the equilibrium protonation as its upper limit.

Figure 3 shows the data for 1,3-dimethoxy-2-methylbenzene plotted in this way. The points represent experimental values of k_{obsd} and C_{H^+} together with log I values extrapolated using the linear relationship between log I and H_c reported for this substrate in the preceding paper;[1b] least-squares analysis gives α_A = 0.652 ± 0.009. Similar plots were constructed for all of the other substrates examined here except 1,3,5-trihydroxybenzene, for which the kinetics could not be measured at the same temperature as the equilibrium. Equilibrium data were also not available for anisole, which is too weakly basic to undergo complete protonation in aqueous perchloric acid, but methyl substituents are known not to alter the acidity dependence of this equilibrium reaction,[1b] and the log I *vs.* H_c relationship for 3,5-dimethylanisole was therefore used instead. All of the plots proved to be quite linear with correlation coefficients lying in the range 0.981–0.999, which shows that α_A is in fact constant over the intervals examined. The values of α_A obtained are listed in the second column of Table III.

These results can be supplemented by two additional values of α_A for aromatic protonation which are based

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[6] See Kresge, *et al.*, Table II, footnote *b*.

[7] This parameter was originally[6] called α, but we have now given it the subscript A, for acidity, to distinguish it from the exponent in the Brønsted relation, α_B, with which it is sometimes compared.

on data appearing in the literature. The formation of protonated azulene under conditions where appreciable amounts of this species are present is sufficiently slow to be measured by stopped-flow methods, and rates of approach to equilibrium as well as equilibrium concentrations have been determined in perchloric acid solution.[8] These data give a good linear relationship when plotted according to eq 4, *r* (correlation coefficient) = 0.996 and $\alpha_A = 0.545 \pm 0.014$. Rates of dedeuteration of benzene have also been measured in perchloric acid solution;[9] the combination of these with the expected direct proportionality between log *I* and H_c for this substance gives $\alpha_A = 0.930 \pm 0.041$. These two values of α_A have been placed in Table III following those based on the experiments reported here.[10]

These results show that kinetic acidity dependence in aromatic protonation is not a constant fraction of equilibrium acidity dependence; α_A for the present examples varies from a low of 0.54 to a high of 0.93. It is significant, moreover, that this parameter increases in magnitude with decreasing substrate basicity; the relationship is shown in Figure 4 as a plot of α_A against pK_a of the substrate conjugate acid corrected for differences in the number of basic sites, *q*. A relationship of this kind might follow from the operational definition of α_A, for, as the fractional extent to which changes in solute–solvent interaction attending the equilibrium protonation are accomplished at its transition state, α_A could provide a measure of the position of the transition state along the reaction coordinate. This transition state should, of course, shift from reactant-like to product-like as the substrate becomes less basic and the reaction more endothermic.[13] The fact that α_A does increase in the same sense indicates that this parameter is in fact responding to changes in transition-state structure and is providing a measure of the order of the forming bond, *z*. It is not certain, of course, whether α_A can be directly equated with *z*; this would follow if the changes in solvation energy accompanying protonation were linearly proportional to the extent of positive charge transfer from hydronium ion to substrate, which should itself be equal to *z*.

Brønsted Relations. Another quantity which is thought[14] to measure transition-state structure in the same sense as α_A is the exponent in the Brønsted relation, α_B.[15] Brønsted relations are usually generated by varying the catalyst in an acid- or base-catalyzed reaction while keeping the substrate constant, but there is no fundamental difference between substrate and catalyst in a proton-transfer process, and equally meaningful Brønsted relations can be obtained by varying the substrate and keeping the catalyst constant.[16] The present data lend themselves to the latter kind of treatment.

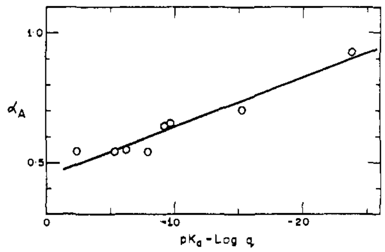

Figure 4. The relationship between α_A and substrate basicity.

Dilute solution specific rates of protonation (protodetritiation), k_{H^+}, were obtained for the purpose of constructing such a Brønsted relation by extrapolating the log k_{obsd} *vs.* *H* (acidity function) relationships described above down to $H = 1.00$; this is the point where H_0 first becomes equal to $-\log C_{H^+}$,[17] and where differences between *H* and $-\log C_{H^+}$ are likely to vanish as well for other acidity functions based on neutral indicators. The H_c function would seem to be the most appropriate scale with which to make these extrapolations, for it is based principally on aromatic hydrocarbon bases very similar to the substrates being examined here. It was found, however, that use of this scale gave extrapolated values two to three times greater than k_{H^+} measured directly in dilute acid solution for the two cases (1,3,5-trimethoxybenzene-*2-t* and 1,3-dimethoxybenzene-*4-t*)[3f] where such comparison is possible, whereas extrapolation using H_0 gave results in agreement with directly measured values to ±10%. This implies that the published[18] H_c scale may be somewhat improperly joined to $-\log C_{H^+}$. This scale, however, is an approximately linear function of H_0 up to *ca.* 6 *M* perchloric acid,[1b] and this difficulty may therefore be circumvented by using H_0 to correlate data for H_c-type bases up to 6 *M* acid. The present extrapolations were consequently performed using H_0 for all substrates except anisole and benzene. These two aromatics give conveniently measurable rates of exchange only in perchloric acid more concentrated than 6 *M* where the linear relationship between H_c and H_0 breaks down,[1b] and these data were therefore extrapolated using H_c. Since these two extrapolations were quite long, the values of k_{H^+} for anisole and benzene thus obtained can be no more than rough estimates, to which it hardly seems worthwhile to apply corrections for a possibly faulty union of H_c with $-\log C_{H^+}$. (The value of k_{H^+} for benzene is also based on rates of protodedeuteration rather than protodetritiation, but, for a substrate as weakly basic as this, isotope effects are likely to be small[4h] and detritiation rates therefore essentially equal to dedeuteration rates.)

The bimolecular rate constants obtained in this way are listed in Table III, together with values available from the literature[4d,h] for the detritiation of azulene and several of its derivatives. These data, together with the pK_a values also listed in Table III, were used to construct the Brønsted plot shown in Figure 5 (open circles). Statistical corrections of $p = 3$ for H_3O^+ and

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(9) D. P. N. Satchell, *J. Chem. Soc.*, 3911 (1956).

(10) Still other aromatic protonation reactions have been examined in concentrated sulfuric acid solution,[11] but only data obtained in perchloric acid were used here because of the uncertainty of C_{H^+} in sulfuric acid as well as the presence of proton transfer from other acidic species.[12]

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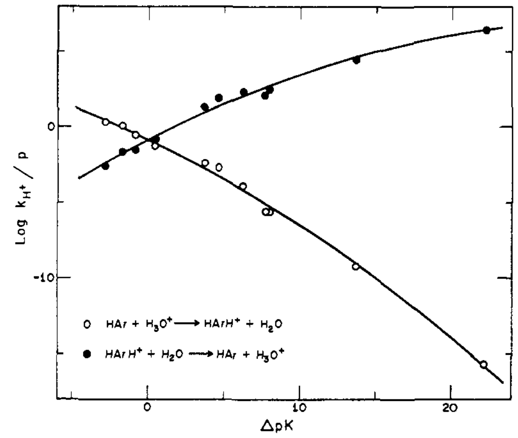

Figure 5. Brønsted relation for the protonation of aromatic substrates by H_3O^+ and for the deprotonation of benzenonium ion conjugate acids by H_2O.

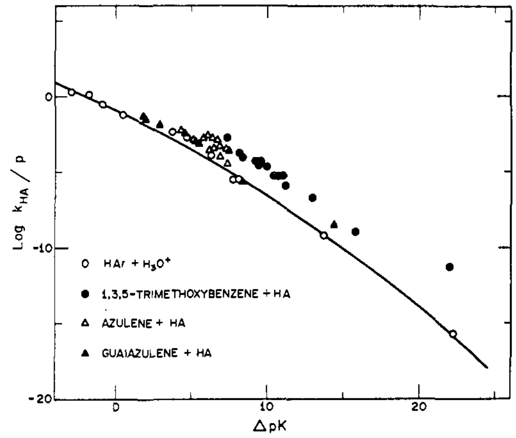

Figure 6. Comparison of specific rates of protonation of 1,3,5-trimethoxybenzene, azulene, and guaiazulene by general acids with the Brønsted relation defined by protonation of these and other aromatic substrates by H_3O^+.

$p = 2$ for $HArH^+$, and $q = 1$ for H_3O^+ and q = (the number of equivalent basic sites in HAr) were used. The abscissa in this plot, instead of being pK_{HArH^+}, is ΔpK, the difference in pK_a (statistically corrected) between proton donor (H_3O^+) and proton acceptor ($HArH^+$).[19] This convention conveniently gives the horizontal coordinate a value of zero for proton transfer between two bases of equal strength, and it puts "downhill" (exothermic) reactions to the left and "uphill" (endothermic) reactions to the right.

This Brønsted plot is quite decidedly nonlinear. The best (least-squares) straight line through the data runs below the points near the center and above those at either end, and gives an error (standard deviation) in $\log k_{H^+}/p$ of 0.65. This error is reduced to 0.35 when the data are fitted to a quadratic expression, which is the line shown in Figure 5 and given as eq 5. Use of

$$\log k_{H^+}/p = -0.846 + 0.481\Delta pK - 0.00866(\Delta pK)^2 \quad (5)$$

higher order expressions produces no further improvement in fit.

Other Brønsted plots are available for aromatic hydrogen exchange,[3g,4d] but none of these is detectably curved. These other plots, however, either cover only a few pK_a units when constructed from a homogeneous set of acids, or, if more extensive, use acids of several different types which cannot be expected to conform to a single Brønsted relation. The curvature in the present case is certainly not very sharp, and its detection is a direct consequence of the fact that the data cover 25 pK_a units.

Brønsted exponents, $\alpha_B = d(\log k_{H^+}/p)/d\Delta pK$, were evaluated from this Brønsted relation by taking the derivative of eq 5 and inserting into the result values of ΔpK for each substrate. The results, listed as the last column of Table III, can be seen to be in remarkably good agreement with the corresponding values of α_A. The average difference between the two parameters is only 0.03, with the largest discrepancies occurring in the case of azulene, for which α_A refers to 6–7° and α_B to 25°, and in the case of benzene, for which the extrapolations required to evaluate k_{H^+} and pK_a were the most uncertain. This good consistency adds support to the interpretation of α_A as a measure of transition-state structure.

It is of interest to compare Brønsted exponents determined in this way with those based on the Brønsted relations for aromatic detritiation reported in the literature, which were constructed by keeping the substrate constant and varying the catalyst acid. Such correlations exist for 1,3,5-trimethoxybenzene,[3g] azulene,[4d] and guaiazulene[4d] and values of α_B based on homogeneous sets of carboxylic acid catalysts are 0.59 ± 0.05, 0.61 ± 0.07, and 0.52 ± 0.10, respectively. These are in reasonably good agreement with the values 0.56 (trimethoxybenzene), 0.47 (azulene), and 0.43 (guaiazulene) listed in Table III. Exact correspondence, however, should not be expected, for α_B depends on ΔpK, and the values of α_B listed in Table III are for an acid, H_3O^+, which is some six pK units stronger than the carboxylic acids used in the literature correlations. Values of α_B calculated for a carboxylic acid of $pK_a = 4.00$ are 0.65 for trimethoxybenzene, 0.57 for azulene, and 0.52 for guaiazulene, each of which is within experimental uncertainty of the measured quantity.

The actual catalytic coefficients upon which these other Brønsted relations are based agree much less well with the present relationship. As Figure 6 shows, these other data, with but one exception (guaiazulene and the ammonium ion), all lie above the correlation line obtained here using hydronium ion catalytic coefficients; the difference often amounts to one or two orders of magnitude. This discrepancy is consistent with the fact that most of these other catalysts are either neutral or negatively charged, and they will therefore undergo an attractive or rate-facilitating interaction with partly formed, positively charged benzenonium ions in the transition states of these reactions. The hydronium ion, on the other hand, is positively charged, and it will experience a repulsive or rate-lowering transition-state interaction. Similar deviations from Brønsted relations have recently been detected in the hydrolysis of vinyl ethers.[20] Another factor which may contribute to the dispersion seen in Figure 6 is the fact that the conventional pK_a of H_3O^+, −1.75, was used in evaluating

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ΔpK; it is likely that this number overestimates the acid strength of this species.[20,21]

The specific rates and acidity constants upon which the present Brønsted relation is based may also be used to calculate rate constants, k_{rev}, for the reverse reactions, deprotonation of the substrate conjugate acids ($k_{rev} = k_{H^+}/p - \Delta pK$ = the value of $[k_{-1}/(1 + k_{-1}/k_2)]$ of eq 2 extrapolated to dilute solution and divided by C_{H^+}). The results give a Brønsted relation themselves, eq 6 and Figure 5 (closed circles), which is constrained

$$\log k_{rev} = -0.846 - 0.519\Delta pK - 0.00866(\Delta pK)^2 \quad (6)$$

by the mathematics of the situation to differ from that for the forward reaction only in the coefficient of the linear term; the linear term coefficient for protonation minus that for deprotonation must in fact be unity. It is hardly significant, therefore, that the Brønsted relation for the reverse reaction is also curved, but it is of some consequence that the slopes of these relationships at $\Delta pK = 0$, α_B (forward) = 0.481 ± 0.036 and α_B (reverse) = 0.519 ± 0.036, are each within experimental uncertainty of being one-half, as demanded by theory.[19]

It is also significant that this Brønsted relation for aromatic deprotonation predicts a limiting rate constant for this reaction, 7×10^5 sec^{-1}, which is far short of the value, 10^{10}–10^{11} sec^{-1},[19] for diffusion apart of two molecules in aqueous solution. It has been suggested[22] that proton transfer to and from carbon in aqueous solution might in general have limiting rate constants several orders of magnitude below diffusion-controlled limits because of the work which must be done to bring the substrate and the acid (or base) together and then break hydrogen bonds between the acid (or base) and the solvent. This work term, W^r, was in fact estimated at 8 ± 2 kcal mol^{-1} for the reaction of diazo acetate ion with acids by analyzing curvature in the Brønsted relation for this reaction through use of a recently proposed formalism describing proton transfer reactions.[14b] A similar treatment of the present data gives $W^r = 12$ kcal mol^{-1} for protonation and $W^r = 8$ kcal mol^{-1} for deprotonation. It is difficult to estimate the uncertainty in these quantities, but it is likely to be at least 2 kcal mol^{-1}, which puts the present results in good agreement with the diazo acetate data. The fact that W^r for protonation in the present case is greater than that for deprotonation may also be significant, for, since one of the reaction partners in this deprotonation is the solvent, the work of bringing the reactants together need not be done.

Conclusions

The results presented here show that kinetic acidity dependence in concentrated acids may be used to furnish reliable estimates of transition-state structure as measured in terms of the extent of proton transfer. In the present examples, the extent of proton transfer at the transition state (α_A or α_B) covered the considerable range 0.5–0.9, and, as a consequence, the kinetic acidity dependence of these reactions showed appreciable variation. This, of course, is in addition to the variation in kinetic acidity dependence produced by differences in substrate structure, which affect acidity dependence through changes in specific substrate–solvent interactions, principally hydrogen bonding.

Identification of these two factors, transition-state structure and substrate structure, as major contributors to differences in kinetic acidity dependence would seem to imply rather severe limitations upon the use of kinetic acidity dependence as a criterion of reaction mechanism. Attempts to use kinetic acidity dependence as a mechanistic tool have traditionally presumed[23] that different reaction mechanisms will have characteristic, and hopefully different, kinetic acidity dependences. The present results show, however, that differences in substrate and transition-state structure *within a single reaction mechanism* can, in themselves, produce enough variation in kinetic acidity dependence to cover a good portion of the available range. Such behavior, if general, could of course erase enough of the distinction resulting from differences in reaction mechanism to make mechanistic assignment on the basis of kinetic acidity dependence impossible.

This potential loss of kinetic acidity dependence as a mechanistic tool is nevertheless compensated for by the rather detailed picture of bonding in transition states which it can supply, for the structure of a transition state can very well be of greater fundamental significance than the route by which it was reached.

Experimental Section

Materials. Anisole-*4-t* and 1,3-dimethoxybenzene-*4-t* were prepared by decomposing the corresponding arylmagnesium bromides with tritiated water. 1,3-Dimethoxybenzene-*2-t* was obtained from a similar reaction using 2,6-dimethoxyphenyllithium, which was synthesized by treating 1,3-dimethoxybenzene with butyllithium;[24] the fact that acid-catalyzed detritiation of 1,3-dimethoxybenzene-*2-t* obtained in this way was accurately first order over ten half-lives shows that lithiation under these conditions occurs exclusively (>99.9%) at the position between the two methoxy groups. All other tritiated aromatics were prepared by acid-catalyzed exchange in tritiated water or tritiated water–dioxane solution; we have already shown[3] that labeled aromatics prepared in this way give exchange rates identical with those prepared *via* organometallic intermediates.

Labeled substrates were purified to constant specific activity and accepted literature values of physical properties by using standard methods. Deionized water was distilled from alkaline permanganate in glass apparatus. All other materials were best available commercial grades and were used without further purification.

Kinetic Procedure. Rates of loss of tritium from aromatic substrates were determined in wholly aqueous solution at substrate concentrations well below saturation levels. The kinetic technique was essentially that already described in some detail,[3] *i.e.*, aliquots of reaction mixture were quenched and extracted, and portions of the dried extracts were assayed for radioactivity by liquid scintillation counting using toluene-based counting solution. Since the detritiation of phenols is catalyzed by bases as well as by acids, sodium bicarbonate was substituted for sodium hydroxide as the quenching agent for substrates containing free hydroxyl groups. Toluene proved to be a poor extraction solvent for phenols, and the more polar butyl acetate was therefore used instead; these butyl acetate extracts were also washed once with saturated salt (sodium sulfate) solution to remove hydroxyl group bound tritium before being dried over magnesium sulfate.

Sampling was performed for the most part by ordinary pipet, which was adequate for reactions with half-lives down to *ca.* 1 min. In the case of trihydroxybenzene and its ethers, however, this barely permitted entry into the concentrated acid region, and

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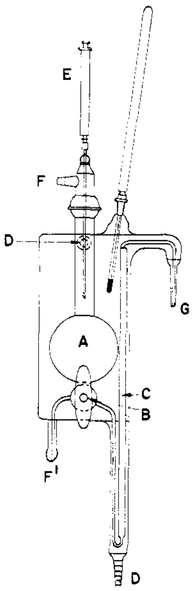

Figure 7. Rapid sampling apparatus.

the gas-driven pipet shown as Figure 7 was used to extend the range of available half-times. This apparatus consists of a reaction chamber, A, connected by way of a 120° stopcock, B, to a pipet of about 5-ml capacity, C; the whole is enclosed in a glass jacket through which constant temperature fluid may be circulated *via* openings D. At the beginning of an experiment, the stopcock is placed so that none of its arms are connected. Acid solution is then introduced into A, and a much smaller amount of substrate solution is placed in the hypodermic syringe, E, which extends into the apparatus through a serum-cap stopper. A supply of nitrogen or some other inert gas at a few pounds per square inch above atmospheric pressure is then connected to openings F and F'. When the acid solution has had sufficient time to come to thermal equilibrium with the water jacket, the reaction is begun by depressing the syringe plunger sharply and shaking the whole apparatus vigorously. Samples are removed by turning the stopcock to connect the reaction chamber to the pipet, allowing the reaction mixture to be driven in to fill the pipet, closing the stopcock, removing excess solution from the pipet tip G, and then expelling the sample into quenching solution through use of gas pressure at F'. Sampling times were noted by a second person, but this could also have been done by the operator speaking into a tape recorder and then timing the record by playing it back. Control experiments showed that the volumes delivered were reproducible to ±0.2% (standard deviation). With some experience using this apparatus, it is possible to take the first sample 15–20 sec after depressing the syringe plunger, and then to remove additional aliquots at 8–10-sec intervals. When used on substrates whose specific activity is sufficient to give an initial sample with counting rate 10^3 times background (so that ten half-lives of reaction may be observed), this method makes accessible specific rates of exchange up to 10^{-1} sec^{-1}.

A Potential Surface for the Insertion of Singlet Methylene into a Carbon–Hydrogen Bond

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Abstract: The potential energy surface for the concerted insertion of singlet methylene into a C–H bond of methane to form ethane is studied by the extended Hückel method. The reaction path is shown in Figures 5, 6, and 7. The reaction begins by an abstraction-like electrophilic attack of a methylene with its empty p orbital impinging in a slightly nonlinear geometry on the H atom. In the intermediate stage of the reaction the H atom is transferred to the methylene while the C–C distance changes little. The reaction terminates in a collapse to the equilibrium ethane structure. The reaction path resembles qualitatively that suggested previously by Benson. There is no indication on our surface of a competing pathway involving true abstraction.

A ubiquitous reaction of singlet methylenes is their rapid and generally indiscriminate insertion into C–H bonds.[1] From the time of the discovery of this reaction by Doering and coworkers[2] there has existed the question whether the mechanism of the reaction is a direct concerted one (eq 1) or whether a two-step abstraction–recombination path (eq 2) is followed. It

$$:CH_2 + R\text{-}H \longrightarrow R\text{-}CH_2\text{-}H \qquad (1)$$

$$\left.\begin{aligned} CH_2 + RH &\longrightarrow CH_3\cdot + R\cdot \\ CH_3\cdot + R\cdot &\longrightarrow RCH_3 \end{aligned}\right\} \qquad (2)$$

appears that both mechanisms are operative but that direct insertion predominates. Thus Doering and Prinzbach[3] found that photolysis of diazomethane in 2-methylpropene-*1*-^{14}C leads mainly to ^{14}C retention at the 1 position in the 2-methylbutene product. Retention of configuration results when the insertion occurs at an asymmetric carbon atom.[4,5] Predominant one-step

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