step.[20] The present findings strengthen the possibility of a direct role for sulfur in catalysis by adenosine deaminase.

Deficiencies of the sulfur nucleophiles, which could be satisfied by the intervention of other catalytic groups, are their relatively slow rates of reaction compared with enzymes and the direction of cleavage of the glutathione derivative in alkali. This latter reaction results in a net transfer of sulfur from glutathione to 6-mercaptopurine ribonucleoside, a reaction which may provide a mechanistic analogy for the enzymatic transfer of sulfur from cysteine to the sulfur-containing nucleotides of soluble ribonucleic acid.[21]

Acknowledgment. We thank Mr. J. B. Macon for his analysis of the rates and products of alkaline hydrolysis of 6-chloropurine ribonucleoside.

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The Mechanism of the Acid-Catalyzed Hydration of Phenylpropiolic Acid[1,2]

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Abstract: In 62–70% sulfuric acid at 25° phenylpropiolic acid is hydrated at a convenient rate to benzoylacetic acid, which decarboxylates more slowly to acetophenone. Rate constants, obtained spectrophotometrically and corrected for decarboxylation, for *p*-methoxy-, *p*-methyl-, unsubstituted, *p*-chloro-, *m*-chloro-, and 3,4-dichlorophenylpropiolic acids show strict proportionality to h_0, plots of log k *vs.* H_0 having essentially unit slope. Hydration rates are quite sensitive to the presence of electron-donating substituents in the phenyl ring; rates for the first five compounds extrapolated to 50% H_2SO_4 at 25° correlate with σ^+ with a ρ of -4.77 ± 0.07. The first four acids hydrate some four times faster in H_2O–H_2SO_4 than in D_2O–D_2SO_4 of the same mole fraction sulfate at 25°. Rates determined at 25 and 44° for *p*-chlorophenylpropiolic acid correspond to an activation energy of 23.7 ± 1.2 kcal/mole and an entropy of activation of -23 ± 4 eu. These results are most consistent with a mechanism involving rate-determining protonation of the α-carbon of phenylpropiolic acid to give a benzyl vinyl cation. A mechanism involving the formation of phenyltrihydroxyallene by 1,4 addition of water, followed by rate-limiting carbon protonation of the allene, is not in agreement with these experimental observations, but is not rigorously excluded by them.

In 1882, Baeyer reported that ethyl phenylpropiolate was converted by cold, concentrated sulfuric acid to another ester, which could be hydrolyzed to an acid.[4] He suggested that the new acid was benzoylacetic acid, formed by hydration of the triple bond. This was confirmed by Perkin, who also prepared benzoylacetic acid directly from phenylpropiolic acid.[5]

Phenylacetylene had been reported to undergo hydration under similar conditions,[6] and Baeyer suggested that this was a general reaction of the triple bond.[4]

Jacobs and Searles[7] studied the rate of hydrolysis of alkyl ethynyl ethers, which hydrolyze more rapidly than alkyl vinyl ethers. Jacobs and Searles suggested that the hydrolysis is initiated by the rate-determining addition of a proton to the triple bond. This conclusion was completely substantiated by the very careful work of Drenth[8–11] and co-workers on alkyl ethynyl ethers and thioethers.

Very recently, Bott, Eaborn, and Walton concluded that the hydration of substituted phenylacetylenes in acetic acid–aqueous sulfuric acid proceeded by the same mechanism.[12]

Two general mechanisms for the acid-catalyzed hydration of α,β-unsaturated carbonyl compounds can be considered: 1,4 addition of water (eq 1–5) or 3,4 addition of water (eq 6–8).

Steps 1, 3, and 5 involve proton transfers to oxygen and are fast equilibrium steps.[13] Steps 2 and 4 may both be rate limiting. In the second general mechanism, (6) is rate determining in aqueous acid, although (7) can conceivably become rate limiting if the activity of water in the medium is low.

Extension of these mechanisms to α,β-acetylenic

(1) Supported in part by a grant from the Petroleum Research Fund administered by the American Chemical Society. Grateful acknowledgment is made to the donors of these funds.

(2) A portion of this material has been published in preliminary form: D. S. Noyce, M. A. Matesich, M. D. Schiavelli, and P. E. Peterson, *J. Am. Chem. Soc.*, **87**, 2295 (1965).

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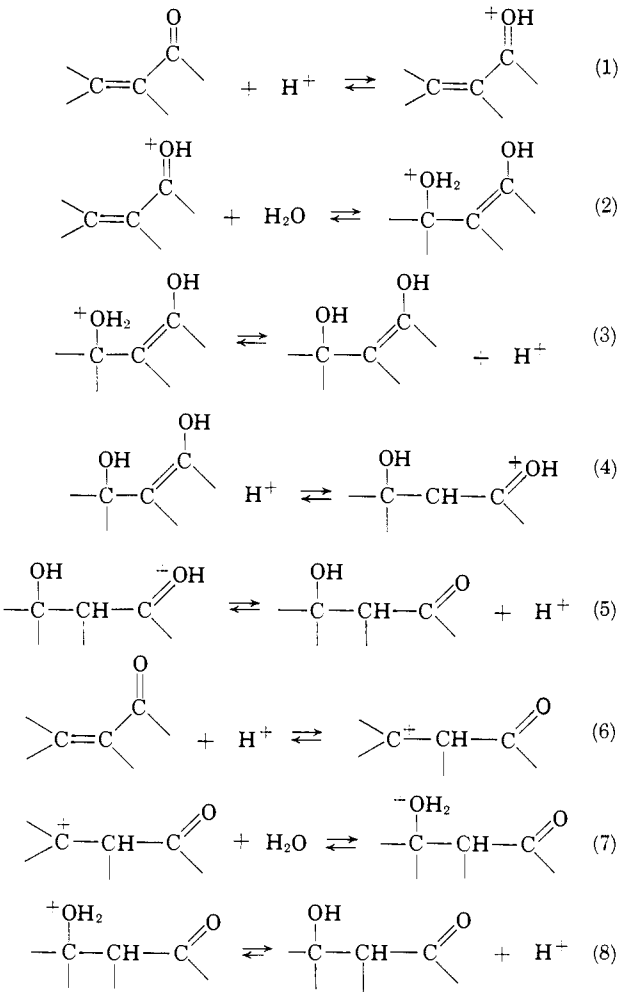

systems provides two reasonable pathways to be considered for the acid-catalyzed hydration of arylpropiolic acids.

Experimental Section[14]

Preparation of Materials. Commercial phenylpropiolic acid was recrystallized from benzene to constant melting point, 137.6–138.5° (lit.[15] 136–137°). *p*-Tolylpropiolic acid was prepared according to the procedure of Newman and Merrill[16] from ethyl *p*-methylcinnamate dibromide. The yield of acid was 50%, mp 148.5–150° (lit.[17] 148° dec). A small portion was recrystallized from chloroform–petroleum ether (bp 30–60°) to give a sample for the kinetic studies, mp 150.9–151.2° dec.

***p*-Methoxyphenylpropiolic Acid.** *p*-Anisoylcarbomethoxymethylenetriphenylphosphorane was prepared by the procedure of Markl[18] from 11.2 g of *p*-anisoyl chloride and 44.1 g of carbomethoxymethylenetriphenylphosphorane, and the product was recrystallized from methanol–water to give 29.7 g (93%) of shining white crystals, mp 184.2–185.3°. Following Markl's procedure, 4.6 g of this material was pyrolyzed at 1.5 mm with a small flame. The yellow material which was produced in the pyrolysis solidified after distilling over to the receiver. It was chromatographed on silica gel with methylene chloride. Slightly yellow methyl *p*-methoxyphenylpropiolate was eluted readily, mp 42.6–43.4° (lit.[19] 45–47°), still contaminated with triphenylphosphine oxide. A solution of 1.96 g of this ester in 30 ml of methanol was stirred at room temperature for 50 min with 15 ml of 1 *N* NaOH. The methanol was removed under vacuum and 50 ml of water added. The cloudy solution was filtered, then acidified with 10% H_2SO_4. The crude acid was filtered and recrystallized from benzene to give 0.970 g (56%) of *p*-methoxyphenylpropiolic acid, mp 139.5–140.0° dec. Two more recrystallizations afforded shining photochromic crystals, mp 138.8–139.3° dec.

Anal. Calcd for $C_{10}H_8O_3$: C, 68.18; H, 4.58; neut equiv, 176.2. Found: C, 67.91; H, 4.34; neut equiv, 174.5.

The crystals slowly turn yellow when exposed to light for several hours, but lose their color if kept several hours in the dark. This characteristic of *p*-methoxyphenylpropiolic acid does not seem to have been reported previously.

***p*-Chlorophenylpropiolic Acid.** From 1.7 g of *p*-chlorobenzoyl chloride and 6.7 g of carbomethoxymethylenetriphenylphosphorane there was obtained 4.2 g (91%) of *p*-chlorobenzoylcarbomethoxymethylenetriphenylphosphorane after recrystallization from methanol–water, mp 149.5–150.9° (lit.[18] 147–149°). Pyrolysis, followed by chromatography, gave 1.8 g of colorless methyl *p*-chlorophenylpropiolate, mp 92.7–94.5° (lit.[19] 92–94°). Saponification, followed by the usual work-up and recrystallization, gave 1.17 g (67%, based on acid chloride) of *p*-chlorophenylpropiolic acid. This acid darkened at 185–186°, then melted sharply 191.8–192.7° dec, in good agreement with that described by Newman and Merrill.[16] A second sample, prepared by the phosphorus oxychloride procedure of Markl,[20] melted at 185–186°, as reported by Markl, but was heavily contaminated with *p*-chlorobenzoic acid.

***m*-Chlorophenylpropiolic Acid.** From 2.0 g of *m*-chlorobenzoyl chloride and 6.8 g of carbomethoxymethylenetriphenylphosphorane a yellow oil was obtained which was pyrolyzed and the product chromatographed as usual to give 1.4 g of a light brown oil whose infrared spectrum corresponded to that expected for methyl *m*-chlorophenylpropiolate. Saponification and recrystallization gave a total yield of 0.80 g (43.6%) of *m*-chlorophenylpropiolic acid, mp 143–146.5° (lit. 144.3–145.1°,[16] 140–141°[21]). A small portion was recrystallized from benzene and from chloroform–petroleum ether to give a sample suitable for kinetic studies, mp 145.7–147.0°.

Anal. Calcd for $C_9H_5O_2Cl$: C, 59.85; H, 2.79; Cl, 19.70; neut equiv, 180.6. Found: C, 60.02; H, 3.02; Cl, 19.88; neut equiv, 179.

3,4-Dichlorophenylpropiolic Acid. In the same fashion, from 3,4-dichlorobenzoyl chloride there was obtained 3,4-dichlorophenylpropiolic acid in 61% yield as fine white needles, mp 182.4–183.5° dec (lit.[22] 182.3–183.4° dec).

Kinetic Measurements and Results. Three absorbing species may be present in any kinetic run: the acetylenic acid, the β-keto acid, and its decarboxylation product, the ketone. The spectra of the three species change considerably in media of differing acidity, due to medium shifts,[23,24] protonation of the acetophenones,[25] and protonation of the benzoylacetic acids. The latter effect is noted in the marked change in the spectra of the hydration products of *m*-chloro- and 3,4-dichlorophenylpropiolic acids in 80–85% H_2SO_4. No striking changes in the spectra of arylpropiolic acids attributable to protonation were noted in this work. It was always possible to find a wavelength at which the molar absorptivities of acetylenic acid and β-keto acid were sufficiently different that the progress of the reaction could be followed spectrophotometrically.

Two methods of making up kinetic solutions were used.

I. A solution (1 ml) of substrate in water or deuterium oxide was diluted to 25 ml with aqueous sulfuric acid or deuteriosulfuric acid which had been cooled slightly so that the heat of mixing would result in a final temperature near 25°. The final concentration of substrate was between 5 and 15 × 10^{-5} *M*. After mixing, the solution was placed in a 1-cm cell in the thermostated cell block of a Beckman DU spectrophotometer and readings were initiated.

II. An aliquot of substrate dissolved in dioxane or 95% ethanol was pipetted into a 50-ml volumetric flask and the solvent evaporated by a dry stream of nitrogen. The flask was then filled with the desired aqueous sulfuric acid solution and mixed. The final con-

(14) Melting points were determined in a liquid bath. Melting points and boiling points are uncorrected. Infrared spectra were obtained using a Perkin-Elmer Infracord or PE-237; ultraviolet spectra were recorded on a Perkin-Elmer 202 spectrophotometer; nmr spectra were obtained using a Varian A-60 instrument. Analyses were performed by the Microanalytical Laboratory, University of California, Berkeley, Calif.

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centration of substrate was about 8×10^{-6} *M*. With fast runs mixing was not always continued until all the solid adhering to the sides of the flask was in solution. The kinetic solution was placed in a 10-cm silica cell in a thermostated cell block in the Beckman DU and readings were initiated.

Method I was used for most phenylpropiolic acid hydrations. Method II was used for *m*-chlorophenylpropiolic acid and 3,4-dichlorophenylpropiolic acid in order to avoid the heat of mixing of the very concentrated sulfuric acid with water.

Constant temperature baths were calibrated with NBS-calibrated thermometers and temperature variations read with Beckmann differential thermometers. The temperature was 24.90 ± 0.05° for all hydrations at 25°, the maximum deviation in any given run being no greater than ±0.02°. At 44° the temperature in the cell block was 43.69 ± 0.04°.

Weighed portions of kinetic solutions were titrated to the phenolphthalein end point with standardized 1 *N* sodium hydroxide, in triplicate. H_0 values below 60% H_2SO_4 were taken from Paul and Long;[26] those above 60% H_2SO_4 from Jorgenson and Hartter.[27] Temperature corrections interpolated from the data of Gelbshtein, Shcheglova, and Temkin[28,29] were applied to determine H_0 at 44°. For solutions in D_2SO_4, H_0 values used were those corresponding to H_2SO_4 solutions of the same mole fraction sulfate. The advantages of this method for comparing acidity in H_2SO_4 and D_2SO_4 have been discussed by Noyce, Avarbock, and Reed.[30]

In most cases absorbance data were collected beyond ten half-lives for hydration in order to assess the extent of which decarboxylation[31] was affecting the infinity absorbance. Occasionally, a portion of the kinetic solution was heated to complete decarboxylation so that final infinity readings could be taken. The ultimate ultraviolet spectrum was shown to be identical with that of the corresponding acetophenone in the kinetic medium for each compound studied. Kinetic data were treated according to the appropriate one of the four following methods.

A. The subsequent decarboxylation was followed for a sufficiently long time after hydration that a rate constant could be determined, using a calculated or experimental stable infinity absorbance, from the final linear portion of a plot of log $(A_t - A_\infty)$ *vs.* time. The appropriate corrections for the subsequent decarboxylation were then made.

B. The wavelength chosen to follow the disappearance of the arylpropiolic acid was that at which the absorptivities of the β-keto acid and the ketone were the same. In this fashion correction of the observed absorbance during the course of the kinetic run for subsequent decarboxylation was minimized.

C. A calculated infinity absorbance gave a good straight line for at least 95% consumption of the acetylenic acid.

D. The subsequent decarboxylation reaction was followed long enough to obtain the rate of change of absorbance per unit time, but was not followed to completion.

Method B is particularly useful when the rate of hydration of the arylpropiolic acid is low, and hence the rate of the subsequent decarboxylation is similar in magnitude. In the more concentrated sulfuric acid media, the correction for the subsequent decarboxylation presents much less of a problem.

Product Isolation Studies. Isolation of Benzoylacetic Acid. A solution of 149 mg of phenylpropiolic acid in 10 ml of ethanol was diluted to 1000 ml with 72.5% sulfuric acid and kept at 25° for 75 min (about 15 half-lives). The solution was poured onto ice and water and extracted four times with ether. The ether extract was concentrated to 100 ml (no heat), washed with acidified water, swirled a few minutes with anhydrous magnesium sulfate, filtered, and evaporated cold to dryness. The slightly yellow material recovered amounted to 126 mg (75% of the theoretical yield of benzoylacetic acid), mp 96–97.5° dec. A mixture melting point with an authentic sample[31] was not depressed, and the infrared spectrum of the product compared well with that of an authentic sample.

Isolation of *p*-Anisoylacetic Acid. A solution of 202.4 mg of *p*-methoxyphenylpropiolic acid in 10 ml of ethanol was diluted to 1000 ml with 40% sulfuric acid. After 68 min (14 half-lives) the solution was poured onto ice and water and extracted with four portions of ether. After the ether was concentrated cold to 75 ml, it was washed with acidified water, swirled with anhydrous magnesium sulfate, filtered, and evaporated cold to dryness. The yield of crude product was 216.9 mg, 97% of the total theoretical yield of *p*-anisoylacetic acid, mp 61.5–66° dec; neut equiv calcd: 194; found: 220; corresponding to 88% β-keto acid and 12% ketone. Attempts at further purification were attended by very severe losses. The nmr spectrum was consistent with the structure assigned.

When a similar solution of *p*-methoxyphenylpropiolic acid was held at 25° for 3 days an excellent yield of *p*-methoxyacetophenone was isolated.

Isolation of 3,4-Dichlorobenzoylacetic Acid. Similarly, 3,4-dichlorophenylpropiolic acid in 85% sulfuric acid for 47 min afforded 3,4-dichlorobenzoylacetic acid, mp 99.2–99.7°, in 53% yield.

A small portion of 3,4-dichlorobenzoylacetic acid was dissolved in acetone and kept at 45° overnight. After the acetone was removed, the solid residue was crystallized from methylene chloride, to give 3,4-dichloroacetophenone, mp 72.3–73.7°. The infrared spectrum was identical with that of an authentic sample, and a mixture melting point showed no depression. Another sample of 3,4-dichlorobenzoylacetic acid in acetone was held at 45° and swept with purified nitrogen. The nitrogen sweep was passed through barium hydroxide solution, and the precipitated barium carbonate was collected and dried. An 85% yield of barium carbonate (37 mg) was obtained.

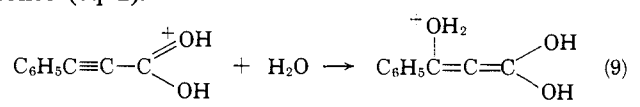
The preparation of 3,4-dichlorobenzoylacetic acid has not been previously reported. Due to the instability of the compound, purification for a satisfactory analysis was not achieved. However, its facile quantitative decarboxylation to 3,4-dichloroacetophenone and its method of formation are considered adequate proof of structure.

Results and Discussion

Kinetic Behavior of Phenylpropiolic Acid Hydration. For the six arylpropiolic acids studied, the hydration rates (corrected for subsequent decarboxylation) generally gave excellent pseudo-first-order kinetics even beyond 95% reaction. No induction periods or other consistent deviations from first-order behavior were observed. Rate constants (collected in Table I) were independent of the wavelength at which they were determined. These observations place certain limitations upon possible reaction mechanisms.

The reaction rate for all the compounds studied correlates very well with h_0, a plot of log *k vs.* H_0 having essentially unit slope in each case. These observations suggest several possible rate-limiting steps which should be carefully considered. Eliminating as unlikely any step involving proton transfer between oxygen atoms as being extremely rapid,[13] the following possibilities are consistent with the observed first-order kinetics and acid catalysis:

A. Rate-limiting attack of water at the β-carbon of the protonated carboxylic acid (eq 9), analogous to the 1,4 addition mechanism for α,β-unsaturated ketones (eq 2).[32]



(9)

B. Rate-limiting protonation of the central carbon of an hypothetical trihydroxyallene present in steady-state concentration (analogous to eq 4).

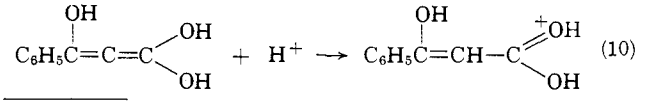


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Table I. Phenylpropiolic Acid Hydration in H_2SO_4 at 24.9°

$$XC_6H_4C{\equiv}CCO_2H \xrightarrow{k_{H_2O}} XC_6H_4\overset{O}{\overset{\|}{C}}CH_2CO_2H \xrightarrow{k_{-CO_2}} XC_6H_4\overset{O}{\overset{\|}{C}}CH_3 + CO_2$$

X	Wt % H_2SO_4	$-CO_2$,[a] cor	$10^5k_{H_2O}$, sec^{-1}	$-\log k$, sec^{-1}	$-H_0$[b]	k_{H_2O}/k_{-CO_2}[c]
p-OCH_3	23.62	A	12.15	3.915	1.27	4.16
	23.63	A	12.15	3.915	1.27	3.92
	27.03	A	21.5	3.667	1.51	7.47
	27.12	A	20.1	3.696	1.52	7.21
	33.05	A	62.6	3.204	1.93	24.6
	33.33	A	65.4	3.184	1.95	26.8
	33.36	A	67.7	3.169	1.95	27.0
	36.05	A	111	2.954	2.13	46.8
	37.92	D	166	2.780	2.27	...
	39.42	D	211.5	2.675	2.37	...
	40.13	B	241	2.618	2.42	...
	40.38	D	266	2.575	2.45	...
	43.81	B	492	2.308	2.74	...
p-CH_3	48.44	A	9.27	4.033	3.21	5.98
	52.72	A	25.7	3.590	3.67	18.2
	57.72	B	90.2	3.045	4.21	...
	61.31	C	224	2.650	4.62	...
H	62.33	A	9.19	4.037	4.74	21.5
	65.56	..	21.8	3.662	5.15	...
	67.05	..	37.9	3.422	5.36	...
	67.40	A	40.0	3.398	5.41	98.4
	68.92	..	61.8	3.209	5.63	...
	69.95	..	96.3	3.017	5.79	...
	70.36	A	107.5	2.969	5.85	310
p-Cl	64.85	C	6.99	4.156	5.05	...
	67.80	B	18.5	3.737	5.47	...
	70.04	B	36.4	3.438	5.80	...
	72.85	B	99.2	3.003	6.23	...
	73.36	B	112	2.951	6.31	....
	75.55	B	255	2.593	6.64	...
m-Cl	74.38	A	8.69	4.061	6.47	37.7
	76.72	A	20.04	3.698	6.83	105
	78.11	A	33.9	3.470	7.05	199
	80.74	D	84.9	3.071	7.46	...
	82.98	B	200	2.699	7.81	...
	83.84	B	265	2.577	7.94	...
	84.74[d]	D	400	2.398	8.10	...
	84.75	B	338	2.472	8.10	...
	84.81[d]	D	438	2.359	8.11	...
3,4-Cl_2	75.60	D	8.79	4.056	6.65	...
	76.22	D	11.3	3.947	6.75	...
	78.84	B	33.5	3.475	7.16	...
	80.30	C	57.7	3.239	7.39	...
	82.19	B	116	2.936	7.69	...
	84.16[d]	B	253	2.596	8.00	...

[a] Method used for correction for subsequent decarboxylation; see Experimental Section. [b] H_0 values below 60% H_2SO_4 were taken from Paul and Long;[26] those above 60% H_2SO_4 from Jorgenson and Hartter.[27] [c] For rate of decarboxylation see ref 31. [d] These fast runs showed some deviation from first-order kinetics after 95% consumption of acetylenic acid in addition to the usual decarboxylation reaction.

C. Rate-limiting protonation of the α-carbon of the acetylenic acid to give a vinyl carbonium ion analogous to the mechanism for the acid-catalyzed isomerization of *cis*-cinnamic acid.[30]

$$C_6H_5C{\equiv}CCO_2H + H^+ \longrightarrow C_6H_5\overset{+}{C}{=}CHCO_2H \quad (11)$$

D. Rate-limiting attack of water on the vinyl carbonium ion present in steady-state concentration.

$$C_6H_5\overset{+}{C}{=}CHCO_2H + H_2O \longrightarrow C_6H_5\overset{{}^+OH_2}{C}{=}CHCO_2H \quad (12)$$

E. Rate-limiting ketonization of enol formed by any sequence of steps in steady-state concentration.

$$C_6H_5\overset{OH}{C}{=}CHCO_2H + H^+ \longrightarrow C_6H_5\overset{{}^+OH}{C}{-}CH_2CO_2H \quad (13)$$

The last possibility can be largely discounted by the following considerations. The enol content of benzoylacetic acid is appreciable as shown by examination of the nmr spectrum. Rates of keto–enol interconversion have been measured for a number of systems and are much too rapid to account for the observed kinetics in the present situation. The most directly related study is the hydrolysis of 2-ethoxy-1-cyclopentenecarboxylic acid studied by Fife.[33]

In order to distinguish between possibilities A and D or B and C, the results of the hydration of phenylpropiolic acids in deuteriosulfuric acid are helpful. Rate-limiting addition of water to a cationic species present in low concentration would be expected to show a modest inverse solvent isotope effect, k_{H_2O}/k_{D_2O} being about 0.5–0.7. Results of rate measurements in deuterated media are given in Table II.

Solvent Isotope Effects. Solvent deuterium isotope effects were determined for four phenylpropiolic acids and are recorded in Table III. It has been shown that the indicator acidities of H_2O–H_2SO_4 and D_2O–D_2SO_4 are the same above 0.1 *M* at a given molarity.[34] In the present work it was more convenient to determine acidity as weight per cent sulfuric acid. Since the molar volumes of H_2O and D_2O as well as those of anhydrous H_2SO_4 and D_2SO_4 are nearly the same (V_{H_2O}/V_{D_2O} = 0.997 at 30°,[35] $V_{H_2SO_4}/V_{D_2SO_4}$ = 0.996 at 25°[36]), comparisons of the two media on a mole fraction basis are almost entirely equivalent to comparisons at the same molarity and same value of the acidity function. Accordingly, we have chosen to report isotope effects at the same mole fraction sulfate for the two media.

Isotope effects for *p*-methoxyphenylpropiolic acid and *p*-chlorophenylpropiolic acid increase with increasing acidity, whereas those for *p*-tolylpropiolic acid and phenylpropiolic acid itself are independent of acidity. Variations in solvent isotope effect with acidity have frequently been noted in moderately concentrated acid.[30,37–39] Two reasons may be suggested. First, increasing acidity is accompanied by decreasing activity of water and a concomitant change in secondary solvent effects. Second, isotope effects on the activity coefficients of substrates and transition states may change with acidity and with substituent.

Isotope effects have been extrapolated to a common acidity in the last column of Table III. Though there is an apparent trend of the magnitude of the solvent isotope effect with substituent, this apparent trend is effectively an artifact of the extrapolation. No true correlation is observed between isotope effects and substituents or acidity.

The solvent isotope effects observed are not consistent with a rate-limiting reaction of water with a proton-

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Table II. Phenylpropiolic Acid Hydration in D_2SO_4 at 24.9°

$$XC_6H_4C{\equiv}CCO_2D \xrightarrow{k_{D_2O}} XC_6H_4C(=O)CD_2CO_2D \xrightarrow{k_{-CO_2}} XC_6H_4C(=O)CD_3 + CO_2$$

X	Wt % D_2SO_4	N_{SO_4}[a]	$-CO_2$,[b] cor	$10^5 k_{D_2O}$, sec^{-1}	$-\log k$, sec^{-1}	$-H_0$[c]	k_{D_2O}/k_{-CO_2}
p-OCH_3	30.29	0.08697	A	14.55	3.837	1.87	6.63
	34.83	0.09662	A	31.4	3.502	2.19	15.7
	40.10	0.1181	A	88.7	3.052	2.60	54.0
	41.44	0.1240	A	114	2.944	2.72	73.7
	42.65	0.1295	A	145	2.838	2.82	111
	44.34	0.1375	A	209	2.682	3.01	154
p-CH_3	51.18	0.1734	B	7.58	4.121	3.73	...
	54.40	0.1927	D	17.73	3.751	4.07	...
	56.37	0.2054	D	28.9	3.540	4.28	...
	60.70	0.2361	A	94.1	3.026	4.79	129
	60.90	0.2376	D	95.3	3.021	4.81	...
	63.65	0.2593	B	185.5	2.732	5.15	...
H	65.85	0.2784	A	11.4	3.944	5.46	34.5
	66.68	0.2860	A	15.0	3.823	5.58	28.6
	68.64	0.3046	A	25.5	3.593	5.87	65
	70.37	0.3222	D	45.5	3.342	6.12	...
	71.71	0.3365	A	74.7	3.126	6.32	267
	74.21	0.3654	D	173	2.763	6.68	...
p-Cl	66.86	0.2876	D	7.23	4.141	5.61	...
	68.04	0.2988	B	10.00	4.000	5.78	...
	70.08	0.3191	C	18.6	3.730	6.07	...
	73.31	0.3547	B	51.6	3.287	6.55	...
	75.36	0.3797	D	93.0	3.031	6.86	...
	77.59	0.4093	B	179	2.747	7.19	...
	78.74	0.4257	B	298	2.526	7.37	...

[a] Mole fraction of sulfate. [b] Method used to correct for subsequent decarboxylation; see Experimental Section. [c] H_0 value given for that wt % H_2SO_4 of the same mole fraction of sulfate. H_0 values below 60% H_2SO_4 were taken from Paul and Long;[26] those above 60% H_2SO_4 from Jorgenson and Hartter.[27]

Table III. Solvent Deuterium Isotope Effects for Hydration of Phenylpropiolic Acids $XC_6H_4C{\equiv}CCO_2H$, 24.9°

X	Acidity range,[a] wt % H_2SO_4	$-H_0$ range[a]	$(k_{H_2SO_4}/k_{D_2SO_4})_{obsd}$[b]	$(k_{H_2SO_4}/k_{D_2SO_4})_{50\%}$[c]
p-OCH_3	32.2–43.8	1.87–2.74	3.73–4.57	5.30
p-CH_3	53.3–61.3	3.73–4.62	3.73–3.73	3.73
H	67.7–70.4	5.46–5.85	4.07–4.07	4.06
p-Cl	68.7–75.6	5.61–6.64	3.43–4.04	2.40
p-Cl[d]	67.9[d]	5.40[d]	3.15[d]	...

[a] Acidity range over which experimental data are available in both media. [b] Isotope effect compared at same mole fraction SO_4, computed from least-squares correlation lines over acitity range indicated. [c] Isotope effect extrapolated to 50% H_2SO_4, H_0 = −3.38, N_{SO_4} = 0.1552. [d] At 43.7°, rate in H_2SO_4 interpolated using least-squares correlation line; see Table IV.

ated substrate present in small equilibrium concentration. Accordingly, mechanisms A and D (eq 9 and 12) are eliminated as possible rate-determining steps in phenylpropiolic acid hydration.

On the other hand, the solvent isotope effects observed are completely consistent with a rate-determining proton transfer to carbon. Thus, they are consistent with either of the rate-limiting steps B or C (eq 10 and 11).

Temperature of Dependence of *p*-Chlorophenylpropiolic Acid Hydration. For most of the phenylpropiolic acids, it was not possible to obtain hydration rate data of high precision above 25° because decarboxylation of the product benzoylacetic acids was comparable in rate to the hydration. *p*-Chlorophenylpropiolic acid is more satisfactory for a study of temperature coefficients. The reaction may be studied at very high acidities, which retard the subsequent decarboxylation reaction. Moreover, *p*-chlorobenzoylacetic acid and *p*-chloroacetophenone have two isosbestic points near 230 mμ in 60% sulfuric acid. The absorbance of the acetylenic acid is sufficiently different so that its rate of hydration can be followed at these wavelengths; thus hydration rates can be measured without interference from subsequent decarboxylation.

Rate constants for *p*-chlorophenylpropiolic acid hydration in H_2SO_4 at 43.69° are given in Table IV. H_0 values were corrected for temperature.[28,29] Activation parameters were calculated from interpolated and extrapolated data in Table V. Since the H_0 slopes are the same within experimental uncertainty, the activation energy was calculated at H_0 = −5.30 and is independent of acidity. Data at 25° were extrapolated to the standard state H_0 = 0 (a_{H^+} = 1) for calculation of the entropy of activation.

Table IV. Hydration of *p*-Chlorophenylpropiolic Acid at 43.69°

Medium	Wt % H_2SO_4-(D_2SO_4)	$-CO_2$[a] cor	$10^5 k$, sec^{-1}	$-\log k$, sec^{-1}	$-H_0$[b]
H_2SO_4	57.54	B	8.75	4.058	4.15
	59.54	B	16.76	3.776	4.36
	62.06	B	29.8	3.525	4.65
	67.45	B	145.2	2.838	5.33
	68.82	B	214	2.670	5.53
D_2SO_4	66.06	B	52.5	3.280	5.40[c]

[a] See Experimental Section. [b] H_0 values below 60% H_2SO_4 were taken from Paul and Long;[26] those above 60% H_2SO_4 from Jorgenson and Hartter;[27] temperature corrections were taken from Gelbshtein, Shcheglova, and Temkin.[28,29] [c] H_0 value for H_2SO_4 of N_{SO_4} = 0.2803.

Table V. Temperature Dependence of *p*-Chlorophenylpropiolic Acid Hydration in H_2SO_4

Temp, °C	H_0 slope[a]	log $k_{-5.30}$[b]	log k_0[c]
24.88 ± 0.02	0.974 ± 0.012	−3.914 ± 0.015	−9.075 ± 0.069
43.69 ± 0.04	0.993 ± 0.027	−2.881 ± 0.032	
	$E^{\ddagger}$ = 23.73 ± 1.17 kcal/mole,[d] calcd at H_0 = −5.30		
	$\Delta H^{\ddagger}$ = 23.1 ± 1.2 kcal/mole		
	$\Delta S^{\ddagger}$ = −22.5 ± 4.3 eu, [d] calcd from E^* and $k_{25}°$ at H_0 = 0		

[a] Uncertainty is standard deviation of slope. [b] Uncertainty is standard deviation of log k; H_0 = −5.30. [c] Uncertainty calculated from standard deviation in slope; H_0 = 0. [d] Maximum uncertainty calculated using uncertainties listed for all experimental quantities.

These activation parameters are consistent with a rate-determining proton transfer to carbon.[40]

Substituent Effects. An examination of the substituent effects is instructive in attempting to decide between mechanistic possibilities B and C (eq 10 and 11). Rate data have been extrapolated to 50% sulfuric acid solution (Table VI) for comparison. A minimum of extrapolation is thus involved. It is to be noted that the relative rates are acidity dependent, since the slopes of the log k *vs.* H_0 plots are not identical (Table VII).

Table VI. Hydration Rates of Phenylpropiolic Acids $XC_6H_4C{\equiv}CCO_2H$, 50% H_4SO_4, 24.9°

X	k, sec^{-1} [a]	k_{rel}	σ^+ [b]
p-OCH_3	2.84×10^{-2}	6400	−0.778
p-CH_3	1.35×10^{-4}	30.4	−0.311
H	4.44×10^{-6}	1.00	0.000
p-Cl	1.64×10^{-6}	0.369	+0.114
m-Cl	6.51×10^{-8}	0.0147	+0.399
3,4-Cl_2	2.66×10^{-8}	0.0060	(+0.513)[c]

[a] Extrapolated to 50% H_2SO_4 (H_0 = −3.38), using parameters in Table VII. [b] Taken from H. C. Brown and Y. Okamoto, *J. Am. Chem. Soc.*, **80**, 4979 (1958). [c] Not included in least-squares calculation of ρ.

Table VII. Acidity Dependence of Hydration Rates of Arylpropiolic Acids $XC_6H_4C{\equiv}CCO_2H$ at 24.9°

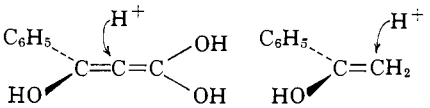
X	Medium	−(d log k/dH_0)[a]
p-OCH_3	H_2SO_4	1.131 ± 0.015
	D_2SO_4	1.030 ± 0.021
p-CH_3	H_2SO_4	0.984 ± 0.008
	D_2SO_4	0.985 ± 0.023
H	H_2SO_4	0.964 ± 0.019
	D_2SO_4	0.964 ± 0.023
p-Cl	H_2SO_4	0.974 ± 0.012
	D_2SO_4	0.904 ± 0.012
	H_2SO_4[b]	0.993 ± 0.027[b]
m-Cl	H_2SO_4	1.011 ± 0.016
3,4-Cl_2	H_2SO_4	1.079 ± 0.013

[a] Least-squares slope; uncertainty is standard deviation. [b] At 43.7°.

A least-squares treatment of the data for the five monosubstituted compounds gives a ρ of −4.77 ± 0.07 when plotted against σ^+. The standard deviation in log k is ±0.06; correlation coefficient r = 0.9997.

The value of ρ is similar to but somewhat more negative than that observed for the isomerization of the *cis*-cinnamic acids;[41] it is more negative than the ρ observed for the hydration of styrenes[42,43] or of phenylacetylenes.[12] This accords very well with mechanism C (eq 11). Taking into consideration the fact that the carbon–carbon bond between the aromatic ring and the side chain is somewhat shorter (sp^2–sp bond) than in the case of *cis*-cinnamic acid, it is perhaps to be expected that there would be increased interaction between the electron-deficient center and the aromatic substituents. Comparison of styrene and phenylacetylene reveals a comparable similarity.

However, the magnitude of ρ is very difficult to rationalize as being consistent with eq 10 as the rate-limiting step. The effective ρ for a reaction sequence with eq 10 as the rate-limiting step can be estimated if one assumes that the steady-state concentration of the hypothetical allenic intermediate is not strongly dependent on the nature of the substituent. A maximum substituent effect for eq 10 itself would be expected if the central carbon of the aryltrihydroxyallene protonated perpendicular to the aryl ring, permitting maximum stabilization of the charge by the ring. Since a subsequent 90° rotation about the central carbon would be required for through-conjugation, the ketene acetal part of the structure would not increase stability in this transition state. Thus, a model for the protonation of allene (I) is the acid-catalyzed ketonization of acetophenone enol (II). ρ for the



I II

latter can be estimated from rate data for the acid-catalyzed enolization of acetophenone[44] (ρ = −1.6) combined with equilibrium data for enol/ketone for 2-aroylcyclopentanones in methanol[45] (ρ = +0.9) to be about −2.5.

However, it is probable that the central allene carbon would protonate preferentially from the other direction, since benzoylketene dimethyl acetal is very rapidly hydrolyzed by dilute acid.[46] In this case, ρ would be near zero since at the transition state the ring is not conjugated with the positive center.

Thus, a ρ of −4.77 is very difficult to rationalize as being consistent with eq 10 as a rate-limiting step. The vinyl cation mechanism is the most satisfactory way to interpret the reaction.

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