

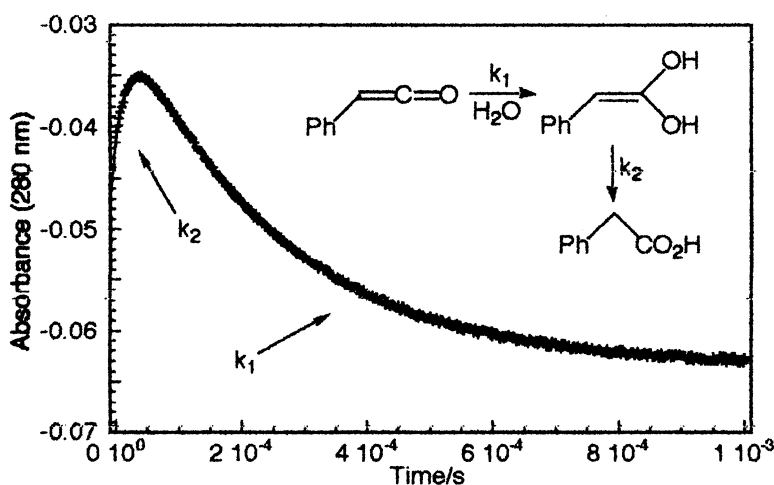
Article

Hydration of Phenylketene Revisited: A Counter-Intuitive Result

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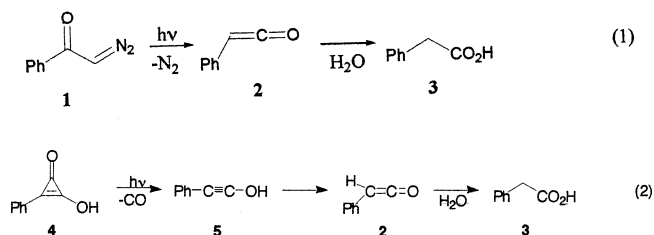
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Abstract: In previous work (*Can. J. Chem.* **1987**, *65*, 1719–1723 and *J. Am. Chem. Soc.* **1995**, *117*, 9165–9171), flash photolysis of diazoacetophenone or phenylhydroxycyclopropanone in aqueous solution was found to produce phenylketene as a short-lived transient species with absorbance at $\lambda \cong 260$ nm, which decayed with single-exponential kinetics. It has now been discovered that, in the acidity region $[H^+] = 0.000\ 01$ to 0.06 M, this decay is preceded by a faster absorbance rise, and that the overall change conforms well to a double exponential rate law. Analysis of the new data produces rate profiles whose general shapes, as well as the numerical values of their constituent rate constants, plus the form of buffer catalysis, indicate that this newly discovered absorbance rise represents ketonization of phenylacetic acid enol, and that the subsequent absorbance decay represents addition of water to phenylketene. The chemistry of the system, however, requires ketene hydration to precede enol ketonization in a time sequence opposite from that of the absorbance changes. This seemingly counter-intuitive result is nevertheless consistent with the rate law that governs the time evolution of the central species in a two-step rise and decay, such as that observed here.

In a previous, somewhat limited study,¹ we found that flash photolysis of diazoacetophenone, **1**, in aqueous solution produced a short-lived transient species with absorbance in the region about $\lambda = 260$ nm, which decayed with single-exponential kinetics in a process attributed to the hydration of phenylketene, **2**, formed by a photo-Wolff reaction, eq 1. This decay was catalyzed by hydroxide ion but not by hydronium ion, which is typical of ketene hydrations,² and gave a rate constant for uncatalyzed hydration that agreed well with one obtained before³ in an investigation using changes in conductivity, produced by ionization of the carboxylic acid hydration product, **3**, to monitor the transient decay. Somewhat later,⁴ we also generated phenylketene by flash photolytic decarbonylation of phenylhydroxycyclopropanone, **4**, followed by tautomeric rearrangement of the phenyllyol, **5**, so formed, eq 2. The kinetics of hydration of phenylketene formed in this way agreed well with those of phenylketene formed from diazoacetophenone.



We have now examined the reaction of phenylketene in somewhat greater detail and have found that, at certain acidities,

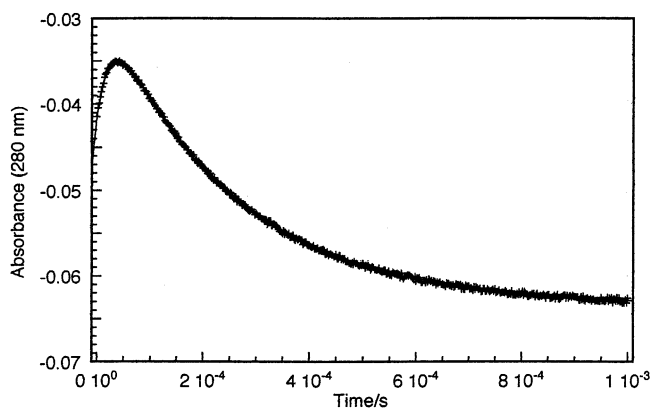


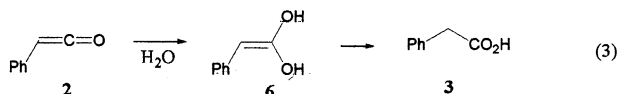
Figure 1. Hydration of phenylketene in an aqueous acetic acid buffer solution at 25 °C; $[\text{HOAc}] = [\text{OAc}^-] = 2.5 \times 10^{-3}$ M; least-squares fitting of a double exponential expression gave the rate constants $k_1 = (4.20 \pm 0.01) \times 10^4 \text{ s}^{-1}$ and $k_2 = (4.11 \pm 0.02) \times 10^3 \text{ s}^{-1}$.

the absorbance decay we saw before is preceded by an absorbance rise, with the overall change conforming well to a double exponential rate law. This is illustrated in Figure 1.

In the present paper, we present evidence showing that this initial absorbance rise is produced by ketonization of the

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- (2) Tidwell, T. T. *Ketenes*; Wiley-Interscience: New York, 1995; pp 576–587.
- (3) Bothe, E.; Meier, H.; Schulte-Frohlinde, D.; von Sonntag, C. *Angew. Chem., Int. Ed. Engl.* **1976**, *15*, 380–381.
- (4) Chiang, Y.; Kresge, A. J.; Popik, V. V. *J. Am. Chem. Soc.* **1995**, *117*, 9165–9171.

carboxylic acid enol intermediate, **6**, formed by the addition of water to the ketene in the hydration reaction, eq 3. This conclusion is, in a sense, counter-intuitive because it attributes the first part of the double exponential absorbance change to the second of the two reactions in eq 3, and the second part of the absorbance change to the first of the two reactions.



Experimental Section

Materials. Diazoacetophenone was prepared by treating benzoyl chloride with diazomethane,⁵ and phenylhydroxycyclopropenone was a sample prepared in a previous study.⁴ All other materials were best available commercial grades.

Kinetics. Flash photolytic rate measurements were made using both a conventional flash lamp (microsecond) system⁶ and an excimer laser (nanosecond) system operating at $\lambda = 248$ nm,⁷ which have already been described. Reaction mixtures were thermostated at 25.0 ± 0.05 °C, and rate constants were calculated by least-squares fitting of exponential functions.

Results

Single-Exponential Rate Measurements. Rates of reaction of phenylketene under single-exponential conditions were determined using substrate generated from diazoacetophenone (eq 1) and also from phenylhydroxycyclopropenone (eq 2); concordant results were obtained with the two substrate sources. Measurements were made in dilute aqueous perchloric acid and sodium hydroxide solutions and in acetic acid, biphosphate ion, tris-(hydroxymethyl)methylammonium ion, and ammonium ion buffers. The ionic strength of these solutions was maintained at 0.10 M by adding sodium perchlorate as required. The data so obtained obeyed the first-order rate law well, and observed first-order rate constants were determined by least-squares fitting of a single-exponential function. These results are summarized in Tables S1–S3.⁸

The rate measurements in buffers were made in series of solutions of constant buffer ratio and constant ionic strength (0.10 M), and therefore constant hydronium ion concentration. As Figure 2 illustrates, observed first-order rate constants determined for a given buffer solution series proved to be linear functions of buffer concentration, and the data were therefore analyzed by least-squares fitting of the linear buffer dilution expression shown as eq 4. The zero buffer-concentration intercepts, k_{int} , obtained in this

$$k_{\text{obs}} = k_{\text{int}} + k_{\text{buff}}[\text{buffer}] \quad (4)$$

way, together with the rate constants determined in perchloric acid and sodium hydroxide solution, are displayed as circles (○) in the lower profile of Figure 3. Hydronium ion concentrations needed to construct this rate profile were determined by calculation, using literature values of thermodynamic acidity

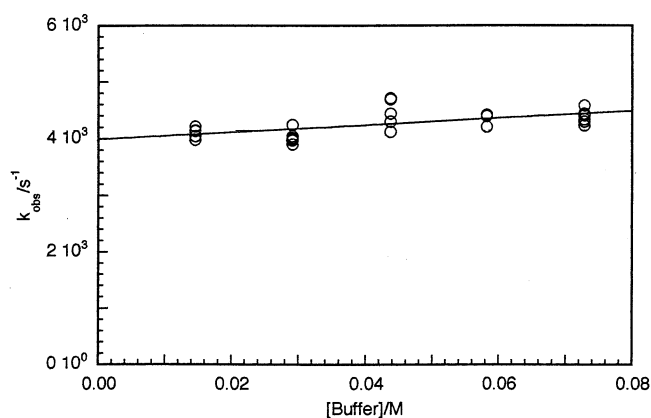


Figure 2. Buffer dilution plot for the reaction of phenylketene in aqueous acetic acid buffer solutions, $[\text{HOAc}]/[\text{OAc}^-] = 0.46$, at 25 °C, based on rate constants determined under single-exponential conditions.

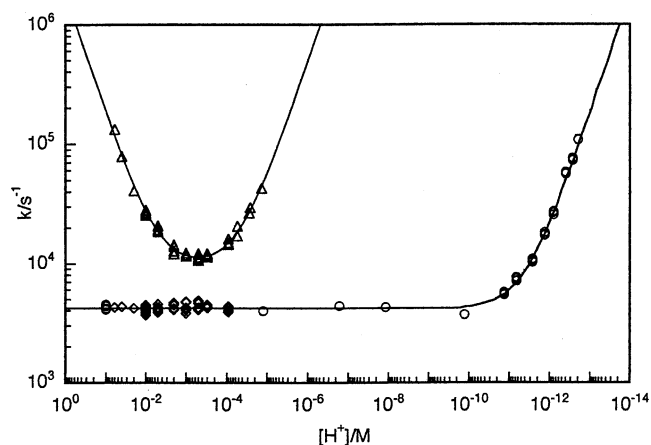


Figure 3. Rate profiles for the addition of water to phenylketene based on rate constants obtained under single exponential (○) and double exponential (◇) conditions, and for the ketonization of phenylacetic acid enol based on rate constants obtained under double exponential conditions (△).

constants of the buffer acids and activity coefficients recommended by Bates.⁹

Double Exponential Rate Measurements. Because the transient signal obtained by flashing phenylhydroxycyclopropenone was stronger than that obtained by flashing diazoacetophenone, only the cyclopropenone precursor was used for the somewhat more difficult double exponential rate determinations. Measurements were made in aqueous perchloric acid solutions over the concentration range 0.06 to 0.000 09 M and in acetic acid buffers that extended the hydronium ion concentration interval provided by these perchloric acid solutions down to 0.000 01 M. Absorbance changes recorded near the middle of this acidity range conformed to a double exponential rate expression fairly well, and reliable values of both a fast and a slower rate constant could be obtained by least-squares fitting of such a function. Near the ends of this range, however, as the faster rate constant became increasingly greater while the slower one remained constant, the initial absorbance rise grew progressively smaller and the double exponential absorbance change morphed into single exponential form; values of the faster rate constant consequently became increasingly less reliable. This difficulty was relieved somewhat for perchloric acid solutions in the range 0.06 to 0.01 M by adding together traces from five

(5) Zeller, K.-P. *Chem. Ber.* **1979**, *112*, 678–688.

(6) Chiang, Y.; Hojatti, M.; Keeffe, J. R.; Kresge, A. J.; Schepp, N. P.; Wirz, J. *J. Am. Chem. Soc.* **1987**, *109*, 4000–4009.

(7) Andraos, J.; Chiang, Y.; Huang, G. C.; Kresge, A. J.; Scaiano, J. *J. Am. Chem. Soc.* **1993**, *115*, 10 605–10 610.

(8) Supporting Information; see paragraph at the end of this paper regarding availability.

(9) Bates, R. G. *Determination of pH Theory and Practice*; Wiley: New York, 1973; p 49.

to six shots and doing a least-squares analysis on the sum. Measurements in the buffer solutions were especially problematic because the faster reaction was strongly catalyzed by the buffer but the slower reaction was not. The absorbance traces obtained there were therefore recorded on a shorter time scale, which expanded the rising portion of the trace but cut off most of the subsequent decay. Under these conditions, the decay could be approximated quite well by a straight line rather than an exponential drop, and the data were consequently analyzed by fitting an exponential plus linear function. In this way, better values of the faster rate constant could be obtained, but the slower rate constant could of course not be determined at all. The results obtained are summarized in Tables S4 and S5.⁸

The rate measurements in buffers were again made in series of solutions of constant ionic strength (0.10 M) and constant buffer ratio. Observed first-order rate constants were once more linear functions of buffer concentration and the data were therefore again analyzed by least-squares fitting of eq 4. The zero concentration intercepts so obtained, together with the rate constants measured in perchloric acid solutions, are displayed in Figure 3, with triangles (Δ) representing the faster rate constants and diamonds (\diamond) representing the slower ones.

Discussion

The double exponential rate behavior observed here for the hydration of phenylketene in certain regions of acidity suggests that this reaction involves the formation and subsequent decay of an intermediate species. It is well-known that the hydration of ketenes, in the absence of acid catalysis, occurs through an enol intermediate, as shown in eq 3. It is tempting, therefore, to assign the initial rising portions of the double exponential kinetic traces to the first step in the sequence of eq 3: addition of water to the ketene forming an enol intermediate, and to assign the subsequent drop to the second step of this sequence: ketonization of the enol into the final carboxylic acid product. These assignments, however, are not at all consistent with the chemistry that may be derived from the individual parts of the double exponential traces.

Figure 3 shows that the slower rate constants given by the slower, second parts of the double exponential traces, represented by diamonds in this figure, fit in very well with rate constants obtained under single exponential conditions, represented by circles. The rate profile provided by the combined data, moreover, is typical of ketene hydrations, which characteristically show large uncatalyzed regions followed by weak hydroxide ion catalysis,^{2,4} as observed here. This rate profile, on the other hand, is quite different from those usually found for enol ketonizations. Ketonization reactions generally show strong acid and strong base catalysis, which produces V- or U-shaped rate profiles,¹⁰ like that generated by the initial parts of the double exponential traces obtained here, represented by triangles in Figure 3. The shapes of the two rate profiles therefore suggest that the slower, lower profile should be assigned to the addition of water to phenylketene to produce the enol of phenylacetic acid and the faster, upper profile should be assigned to the ketonization of this enol to the carboxylic acid product.

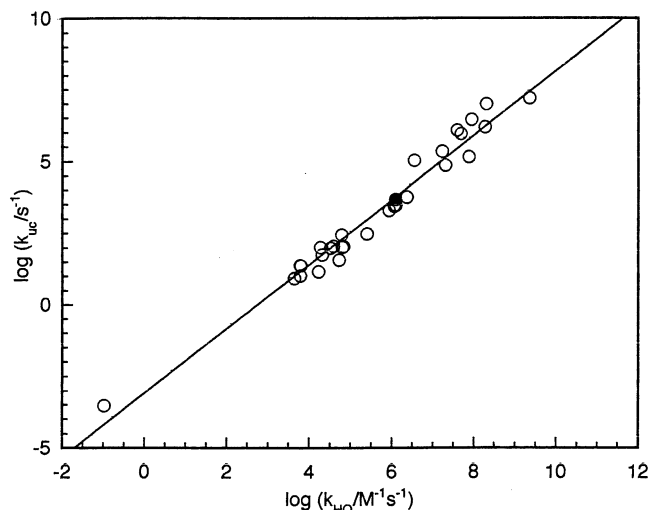


Figure 4. Correlation of rate constants for the addition of water, k_{uc} , and hydroxide ion, k_{HO} , to ketenes in aqueous solution at 25 °C. The filled circle presents presently determined data for phenylketene.

This assignment is supported by the numerical values of the rate constants that make up the rate profiles. The rate law that applies to the lower profile is shown in eq 5

$$k_{obs} = k_{uc} + k_{HO} [HO^-] \quad (5)$$

where k_{uc} refers to the long uncatalyzed region and k_{HO} refers to the hydroxide ion catalyzed reaction that becomes significant only in moderately strongly basic solutions. Least-squares fitting of this expression produced $k_{uc} = (4.20 \pm 0.03) \times 10^3 \text{ s}^{-1}$ and $k_{HO} = (1.19 \pm 0.02) \times 10^6 \text{ M}^{-1} \text{ s}^{-1}$. Water and hydroxide ion operate by similar mechanisms in these reactions: they both serve as nucleophiles that add to the carbonyl carbon atom of the ketene. This similarity leads to a good linear correlation between $\log k_{uc}$ and $\log k_{HO}$, which is illustrated for a group of some 30 ketenes in Figure 4.¹¹ This Figure also shows that the present rate constants, represented by the filled circle in this Figure, conform to this correlation very well. Such good agreement would, of course, be produced only if the bottom rate profile of Figure 3 referred to the addition of water and hydroxide ion to phenylketene and not if the reaction involved were the ketonization of phenylacetic acid enol.

The ketonization of enols occurs by rate-determining proton transfer from a catalyzing acid to the β -carbon atom of the enol or its enolate ion.¹⁰ This produces an acid-catalyzed reaction, such as that represented by the high acidity limb of the upper rate profile shown in Figure 3, which refers to rate-determining carbon protonation of un-ionized enol by hydronium ion with rate constant k_H . There is also an apparent hydroxide ion catalyzed process, such as that represented by the lower acidity limb of the upper rate profile of Figure 3, which occurs through rapid pre-equilibrium acid ionization of the enol to enolate ion, with acidity constant Q_a^E , followed by rate-determining carbon protonation of the enolate ion by water, with rate constant k'_c . This process produces a hydronium ion in the pre-equilibrium step and therefore gives a reaction whose overall rate is inversely proportional to hydronium ion concentration, which appears as hydroxide ion catalysis. These two rate profile limbs are joined by a short uncatalyzed plateau with rate constant k_o , whose

(10) Keeffe, J. R.; Kresge, A. J. In *The Chemistry of Enols*; Rappoport, Z., Ed.; Wiley: New York, 1990; Chapter 7.

(11) Andraos, J.; Kresge, A. J. *Can. J. Chem.* **2000**, *78*, 508–515.

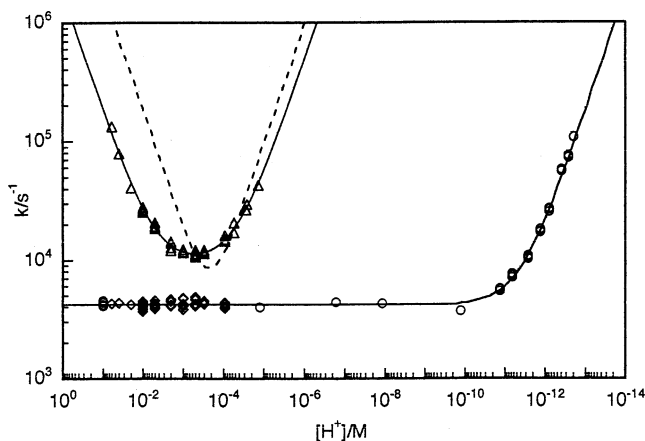


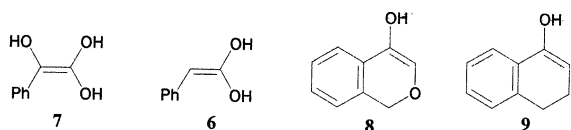
Figure 5. Comparison of rate profile for the ketonization of phenylacetic acid enol estimated as described in the text (dashed line) with rate profiles based on experimentally determined rate constants for the ketonization of phenylacetic acid enol (Δ and upper solid line) and the hydration of phenylketene (\circ , \diamond , and lower solid line).

reaction mechanism may be either rate-determining carbon protonation of un-ionized enol by a solvent water molecule, or pre-equilibrium acid ionization of the enol to enolate plus hydronium ion followed by rate-determining reprotonation of the enolate ion on carbon by hydronium ion. The rate law that applies to the sum of all of these reactions is shown as eq 6, and least-squares fitting

$$k_{\text{obs}} = k_{\text{H}}[\text{H}^+] + k_{\text{o}} + Q_{\text{a}}^{\text{E}}k'_{\text{o}}/[\text{H}^+] \quad (6)$$

of this expression produced the results $k_{\text{H}} = (1.77 \pm 0.05) \times 10^6 \text{ M}^{-1} \text{ s}^{-1}$, $Q_{\text{a}}^{\text{E}}k'_{\text{o}} = (4.74 \pm 0.19) \times 10^{-1} \text{ M s}^{-1}$, and $k_{\text{o}} = (9.48 \pm 0.16) \times 10^3 \text{ s}^{-1}$.

Independent estimates of two of these rate constants may be made from rate constants determined for the ketonization of the enol of mandelic acid,¹² **7**, which differs



from the present enol, **6**, only by the presence of a β -oxygen atom. The effect of this additional oxygen atom may be approximated by comparing rate constants for ketonization of the enols of isochroman-4-one, **8**, and tetralone, **9**,¹³ which also differ only by the presence of a β -oxygen atom. The rate constant k_{H} for the ketonization of tetralone enol is 2000 times greater than that for the ketonization of isochroman-4-one enol, and applying that factor to $k_{\text{H}} = 9.0 \times 10^3 \text{ M}^{-1} \text{ s}^{-1}$ for mandelic acid enol leads to the estimate $k_{\text{H}} = 1.8 \times 10^7 \text{ M}^{-1} \text{ s}^{-1}$ for the enol of phenylacetic acid. Similarly, the product $Q_{\text{a}}^{\text{E}}k'_{\text{o}}$ for tetralone enol is 14 times greater than that for isochroman-4-one enol, and applying that to $Q_{\text{a}}^{\text{E}}k'_{\text{o}} = 7.0 \times 10^{-2} \text{ M s}^{-1}$ for mandelic acid enol leads to $Q_{\text{a}}^{\text{E}}k'_{\text{o}} = 9.9 \times 10^{-1} \text{ M s}^{-1}$ for the enol of phenylacetic acid. These estimates were used to construct the rate profile shown as a dashed line in Figure 5. It may be seen that this rate profile approximates that based on the

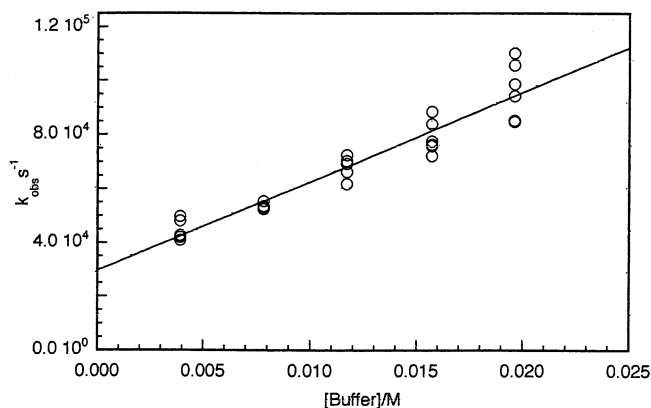


Figure 6. Buffer dilution plot for the reaction of phenylketene in aqueous acetic acid buffer solutions, $[\text{HOAc}]/[\text{OAc}^-] = 0.96$, at 25°C based on faster rate constants determined under double exponential conditions.

measured rate constants assigned to the ketonization of phenylacetic acid enol quite well, especially considering the fact that the analogy between mandelic acid enol and the enol of isochroman-4-one is not exact: in one case the β -oxygen atom is part of a hydroxyl group and in the other case it is an ether oxygen atom. The estimated rate profile, on the other hand, is quite different from the measured rate constants assigned to the addition of water to phenylketene, thus reinforcing the assignments made.

Further support for the assignments made comes from the form of catalysis in acetic acid buffers. The buffer dilution plot shown in Figure 2 was constructed using rate constants determined under single-exponential conditions and whose zero buffer concentration intercept is one of the points making up the lower rate profile of Figure 3. The very weak, almost nonexistent buffer catalysis apparent in Figure 2¹⁴ is characteristic of the behavior of ketenes in acetic acid buffers: in these solutions, the buffer base serves as a nucleophilic adding to the carbonyl carbon atom of the ketene,¹⁵ and acetate ion is a poor nucleophile.¹⁶ The buffer dilution plot shown in Figure 6, on the other hand, was constructed using the faster rate constants determined under double exponential conditions and whose zero buffer concentration intercept is one of the points making up the upper rate profile of Figure 3. The strong buffer catalysis apparent in Figure 6 stands in marked contrast to the weak catalysis demonstrated by Figure 2, and is consistent with the strong catalysis characteristic of enol ketonization, where the buffer species serve as proton-transfer agents.¹⁰

Buffer catalysis may be separated into its general acid, k_{HA} , and general base, k_{B} , components through the use of eq 7, in which $\Delta k_{\text{obs}}/\Delta[\text{buffer}]$ represents the buffer

$$\Delta k_{\text{obs}}/\Delta[\text{buffer}] = k_{\text{B}} + (k_{\text{HA}} - k_{\text{B}})f_{\text{A}} \quad (7)$$

catalytic coefficient, i.e., the slope of a buffer dilution plot, and f_{A} is the fraction of buffer present in the acidic form. Use of this expression to analyze acetic acid buffer catalytic coefficients based upon the faster rate constants obtained under double exponential conditions for the five solution series examined here

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(14) The weakness of this buffer catalysis is reflected in the fairly large uncertainty in its buffer catalytic coefficient, i.e., in the slope of the line shown in Figure 2: $(6.3 \pm 1.7) \times 10^{-3} \text{ M}^{-1} \text{ s}^{-1}$.

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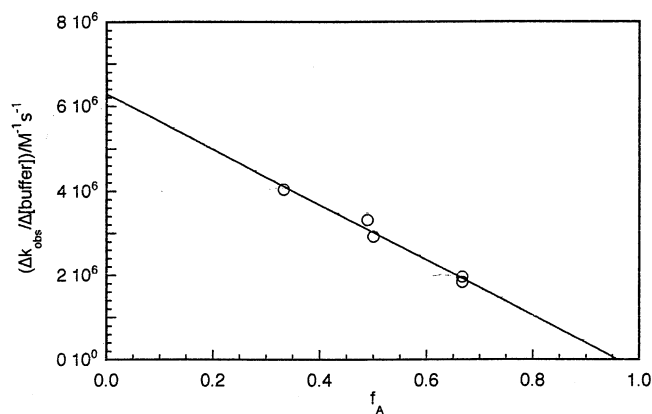


Figure 7. Separation of buffer catalysis into general acid and general base components for the faster reaction observed during the hydration of phenylketene under double exponential conditions.

is illustrated in Figure 7. Least-squares analysis gave $k_{\text{HA}} = -(0.26 \pm 0.27) \times 10^6 \text{ M}^{-1} \text{ s}^{-1}$ and $k_{\text{B}} = (6.29 \pm 0.30) \times 10^6 \text{ M}^{-1} \text{ s}^{-1}$, showing that buffer catalysis is wholly of the general base kind. This is consistent with the fact that all of the buffer series examined had hydronium ion concentrations that lie on the right side of the upper rate profile of Figure 3 where base catalysis is apparent. Ketonization in this region occurs through rapid general base assisted, pre-equilibrium ionization of enol to enolate ion plus general acid, followed by rate-determining carbon protonation of enolate by the general acid so produced; this appears as general base catalysis, as observed.

The arguments made above identify the upper rate profile of Figure 3 as representing the ketonization of phenylacetic acid enol, and the lower rate profile as representing the addition of water to phenylketene. They also assign the initial absorbance rise of the double exponential kinetic traces to enol ketonization and the subsequent absorbance decay to ketene water addition, despite the fact that the chemistry of the system requires ketene water addition to precede enol ketonization. This result may seem to go against common intuition, in the sense that the first part of the absorbance change might be expected to represent the first reaction step and the second part of the absorbance

change to represent the second step, and not the other way around.

This apparent paradox, however, is consistent with the rate law that governs the time evolution of the central species in a two step rise and decay, eq 8, such as that observed here. This rate law, shown as eq 9,¹⁷ contains a positive exponential term



$$[B] = \frac{[A]_0 k_1}{k_2 - k_1} \{ \exp(-k_1 t) - \exp(-k_2 t) \} \quad (9)$$

$\exp(-k_1 t)$, representing the concentration decay, and a negative exponential term, $-\exp(-k_2 t)$, representing the concentration rise. Because of the mathematical form of this expression, the negative exponential term always contains the faster rate constant, and the concentration rise is consequently always faster than the concentration decay. This is a general property of the system and is independent of whether the first or the second reaction in the two-step sequence is the faster one. Since the rate profiles of Figure 3 show the ketonization of phenylacetic acid enol to be faster than the hydration of phenylketene, application of this reasoning to the present system leads to the conclusion that the initial rise in the double exponential kinetic traces observed here represents phenylacetic acid enol ketonization and the subsequent decay represents phenylketene hydration. This seemingly counter-intuitive conclusion consequently follows directly from the mathematics of the situation.

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Note Added after ASAP Publication: The name of one of the authors was misspelled in the version published on the Web July 10, 2004. The spelling of Andrei V. Fedorov is correct in the final Web version published July 14, 2004 and in the print version.

Supporting Information Available: Tables S1–S5 of rate data. This material is available free of charge via the Internet at <http://pubs.acs.org>.

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