Kinetics of Proton Transfer and Tautomerism in Aqueous Diacetylacetone. A Temperature-Jump Investigation

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Abstract: Diacetylacetone (2,4,6-heptanetrione) is a pseudo-acid with a pK of 7.43. Kinetic measurements with the temperature-jump relaxation technique have been carried out in the acidic and basic pH regions at 25° and ionic strength 0.1. Two concentration- and pH-dependent relaxation times were observed in the millisecond time region: one in acidic and one in basic solutions. The nmr spectra of the solutions established that the predominant anionic species had the structure $CH_3COCH_2COCH=CO^-CH_3$ (KHE$^-$). The structures of the main protonated forms could not be deduced from the nmr data. On the basis of the nmr results and the concentration dependence of the acidic relaxation time, the latter was attributed to the process $H^+ + KHE^- \rightleftharpoons KHKH$, where KHKH is the trike-tone form of the acid. The forward rate constant was found to be $2.6 \times 10^6 M^{-1} sec^{-1}$. The mechanism for hydrolysis was found to be considerably more complicated and involved the hydrolysis of two or more anionic species by means of a cyclic mechanism. Only an effective (over-all) rate constant could be measured for the reaction of the molecules with OH$^-$: $k_{eff} = 2.5 \times 10^5 M^{-1} sec^{-1}$. On a Brønsted plot of log k vs. pK_a, both rate constants correlated well with data for other ketones.
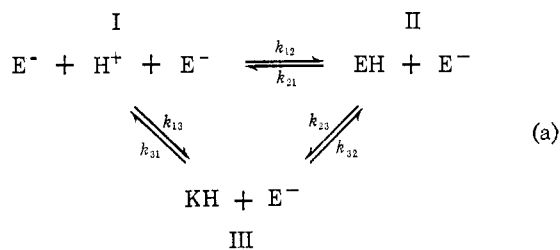
E igen and his co-workers have amply demonstrated that the rate of "normal" proton transfer in simple acid–base systems is diffusion controlled; that is, the rate-determining step is the diffusion of the partners to a critical distance.[2,3] For such reactions, the rate constant for the reaction of H$^+$ with an anion is on the order of 10^{10}–10^{11} M^{-1} sec^{-1}; that for the reaction of OH$^-$ with a protonated species is on the order of 10^{10} M^{-1} sec^{-1}. Exceptions to these generalizations are well understood[3] and occur (a) if steric factors reduce the angle acceptable for successful encounter, (b) in systems which are strongly internally hydrogen bonded (*e.g.*, the reaction of OH$^-$ with the enol of acetylacetone), (c) within the class of acids known as "pseudo-acids," in which a proton is removed directly or indirectly from a carbon atom. It is with an example of this class that this paper is concerned.

Proton transfers in pseudo-acids are characterized by reaction rates[2,3] which range from "very slow" (*e.g.*, the neutralization of $CH_2NO_2^-$ with H$^+$, which can be observed with conventional techniques[4]) to virtually diffusion controlled. Within a series of closely related compounds, the rates of protonation and hydrolysis are a direct function of the pK of the acid.[3,5] An important class of pseudo-acids are ketones which can exist in two or more tautomeric forms. The present paper reports the results of a kinetic study, using the temperature-jump relaxation technique, on the triketone diacetylacetone (2,4,6-heptanetrione). This compound can exist in several tautomeric forms both as the molecule and anion. The possibilities[6] are shown in Table I. The symbols KH and EH refer to keto and enol portions of the molecule, respectively.[7] To date there

has been no detailed investigation of the relative amounts of the various species present nor a study of the mechanism interrelating them.

Reaction Mechanisms

Acidic Region. Eigen, *et al.*,[8] have shown that protonation and deprotonation equilibria in aqueous solutions of tautomeric substances involve a reaction scheme (a) in which the keto and enol forms of the molecule as well as the anion participate. The conversion I–II

$$
\begin{array}{ccc}
\text{I} & & \text{II} \\
E^- + H^+ + E^- & \underset{k_{21}}{\overset{k_{12}}{\rightleftharpoons}} & EH + E^- \\
k_{13}\Big\updownarrow k_{31} & & k_{23}\Big\updownarrow k_{32} \quad\quad (a)\\
& KH + E^- & \\
& \text{III} &
\end{array}
$$

EH = enolic form, KH = keto form, and E$^-$ = enolate anion

involves the recombination of a proton with the enolate to yield the enolic form of the substance. The conversion I–III involves the same partners, but now yielding the ketone. Reaction II–III describes the anion catalyzed keto–enol conversion in which the proton is directly removed from one tautomer by the enolate ion with the generation of the other tautomeric form. This simple mechanism involving only one anion and one each of the keto and enol forms will be shown to be inadequate to explain fully the results obtained in the present work. Nevertheless, the general features of this mechanism and the concentration dependence of the relaxation time apply.

Mechanism a as a whole is characterized by two relaxation times, which may be separated by several orders of magnitude. The shorter of these (τ_1) is associated exclusively with the fast (diffusion-controlled) conver-

(1) Address to which correspondence should be sent.

(2) M. Eigen, *Angew. Chem.*, **75**, 489 (1963); *Angew. Chem. Intern. Ed. Engl.*, **3**, 1 (1964).

(3) M. Eigen, W. Kruse, G. Maass, and L. DeMaeyer, *Progr. Reaction Kinetics*, **2**, 287 (1964).

(4) D. Turnbull and S. Maron, *J. Am. Chem. Soc.*, **65**, 212 (1943).

(5) R. P. Bell, "The Proton in Chemistry," Cornell University Press, Ithaca, N. Y., 1959, pp 160–163.

(6) Some especially unlikely structures (*e.g.*, the trienol) have not been considered.

(7) Note that KHEH₁ signifies the form with a ketone at one end and an enol at the *first* C–O at the other end, *etc.*

(8) M. Eigen, G. Ilgenfritz, and W. Kruse, *Chem. Ber.*, **98**, 1623 (1965); see also G. Ilgenfritz, Diplomarbeit, University of Göttingen, 1963.

Table I. Possible Tautomeric and Resonance Structures for Diacetylacetone and Its Anions

1. Protonated forms		2. Enolate anions[a]	
KHKH	$$\begin{matrix} O & O & O \\ \parallel & \parallel & \parallel \\ CH_3\overset{}{C}CH_2\overset{}{C}CH_2\overset{}{C}CH_3 \end{matrix}$$	KHE⁻	$$\left\{ \begin{matrix} O & O & O^- \\ \parallel & \parallel & \parallel \\ CH_3CCH_2CCH\!=\!\!CCH_3 \\[1em] O & O^- & O \\ \parallel & \parallel & \parallel \\ CH_3CCH_2C\!=\!\!CHCCH_3 \end{matrix} \right.$$
KHEH₁	$$\begin{matrix} O & O\cdots HO \\ \parallel & \parallel \quad \mid \\ CH_3CCH_2CCH\!=\!\!CCH_3 \end{matrix}$$		
KHEH₂	$$\begin{matrix} O\cdots HO & O \\ \parallel \quad \mid & \parallel \\ CH_3CCH\!=\!\!CCH_2CCH_3 \end{matrix}$$	EHE₁⁻	$$\left\{ \begin{matrix} OH\cdots O^- & O \\ \mid \quad \mid & \parallel \\ CH_3C\!=\!\!CHC\!=\!\!CHCCH_3 \\[1em] OH\cdots O & O^- \\ \mid \quad \mid & \mid \\ CH_3C\!=\!\!CHCHC\!=\!\!CCH_3 \end{matrix} \right.$$
EHEH₁	$$\begin{matrix} OH & O\cdots HO \\ \mid & \parallel \quad \mid \\ CH_3C\!=\!\!CHCCH\!=\!\!CCH_3 \end{matrix}$$		
EHEH₂	$$\begin{matrix} OH & OH\cdots O \\ \mid & \mid \quad \parallel \\ CH_3C\!=\!\!CHC\!=\!\!CHCCH_3 \end{matrix}$$	EHE₂⁻	$$\left\{ \begin{matrix} O^-\cdots HO & O \\ \mid \quad \mid & \parallel \\ CH_3C\!=\!\!CHC\!=\!\!CHCCH_3 \\[1em] O & OH\cdots O^- \\ \parallel & \mid \quad \mid \\ CH_3CCH\!=\!\!CCH\!=\!\!CCH_3 \end{matrix} \right.$$

[a] Resonance structures for the anion with the charge localized on carbon not shown.

sion I–II during the restoration of which the slower processes remain virtually unchanged. The longer relaxation time (τ_2) is a function of the kinetic parameters for the two slow steps (I–III, II–III) as well as of the equilibrium parameters of the step I–II. When the two steps are widely separated in time, the two relaxation times are[8]

$$\tau_1^{-1} = k_{12}([H] + [E]) + k_{21} \qquad (1)$$

and

$$\tau_2^{-1} = \frac{k_{13}K_{EH}\left([H] + \dfrac{[E]}{1 + \alpha}\right)}{K_{EH} + [H] + \dfrac{[E]}{1 + \alpha}} + k_{31} +$$

$$\frac{k_{23}[E]}{K_{EH} + [H] + \dfrac{[E]}{1 + \alpha}}\left([H] + \frac{[E]}{1 + \alpha}\right) + k_{32}[E] \quad (2)$$

where the brackets refer to the equilibrium values of the concentrations[9] and the equilibrium constant K_{EH} is defined by
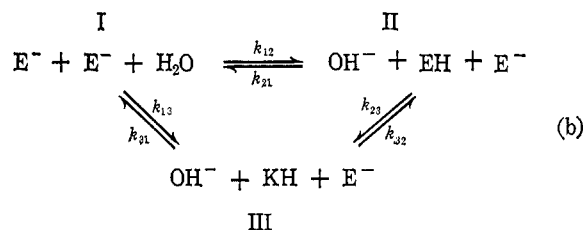
$$K_{EH} = \frac{[H][E]}{[EH]} \qquad (3)$$

The quantity α is a correction term resulting from coupling the slow relaxation time to a rapid indicator equilibrium (to permit the relaxation to be followed by optical means) and is given by

$$\alpha = \frac{[HIn]}{K_{In} + [H]} \qquad (4)$$

where K_{In} is the acid dissociation const   for the indicator.

(9) Charges are omitted for simplicity.

Basic Region. The simplest reaction scheme (*i.e.*, hydrolysis of an anion to yield enol and keto forms) may be represented by

$$\begin{array}{ccc} I & & II \\ E^- + E^- + H_2O & \underset{k_{21}}{\overset{k_{12}}{\rightleftharpoons}} & OH^- + EH + E^- \end{array}$$

$$\begin{array}{c} \searrow k_{13} \qquad \qquad \nearrow k_{23} \\ \nwarrow k_{31} \qquad \qquad \swarrow k_{32} \end{array} \qquad (b)$$

$$OH^- + KH + E^-$$

$$III$$

where the conversion I–II represents the (rapid) protonation of the enolate anion to yield the enol. The two (slower) conversions I–III and II–III involve protonation yielding the keto form and the enolate catalyzed interconversion, respectively. The two relaxation times are

$$\tau_1^{-1} = k_{21}([EH] + [OH]) + k_{12} \qquad (5)$$

and

$$\tau_2^{-1} = k_{31}\left(\frac{[KH]}{1 + \beta}\frac{[OH]}{K'_{EH} + [OH] + \dfrac{[EH]}{1 + \beta}} + [OH]\right) +$$

$$k_{13}\frac{[OH]}{K'_{EH} + [OH] + \dfrac{[EH]}{1 + \beta}} +$$

$$k_{23}[E]\frac{K'_{EH} + [EH]/(1 + \beta)}{K'_{EH} + [OH] + \dfrac{[EH]}{1 + \beta}} + k_{32}[E] \qquad (6)$$

where K'_{EH} is the hydrolysis equilibrium constant[10] of

(10) Throughout this paper, primed symbols will be used to indicate that the equilibrium constant so designated is a hydrolysis equilibrium constant.

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the enol form and β [$\equiv$[HIn]/(K'_{In} + [OH])] is again an indicator correction term.

Experimental Section

Materials. The disodium salt of diacetylacetone was prepared[11] by the hydrolysis of 2,6-dimethyl-γ-pyrone. A solution of 19 g of 2,6-dimethyl-γ-pyrone[12] in 100 ml of absolute ethanol was treated with 19 ml of 16 N NaOH and slowly brought to boiling. The resultant yellow precipitate was collected, repeatedly washed in cold alcohol, and dried. The residue was recrystallized twice from 95% ethanol, vacuum dried, and stored at $-10°$. The spectrum of the fresh salt was found to compare well with that given by Schwarzenbach, et al.[13]

It was found necessary to store the compound in a frozen condition prior to use. A pale yellow aqueous solution at pH 13 turned deep red-brown after 4 days, apparently because of polymerization. Acidic solutions showed no visible evidence of instability over short periods of time. During the course of the experimental work, the compound was prepared freshly for each run.

The solutions were made using sufficient 1 M KNO$_3$ to maintain a constant ionic strength of 0.1. The pK of diacetylacetone at 25° and zero ionic strength[11] (7.43) was converted to the concentration equilibrium constant at 0.1 ionic strength by using the relation $K = K^0/\gamma_\pm^2$ where K^0 is the zero ionic strength value and $\gamma_\pm$ is the mean ionic activity coefficient. The value[14] $\gamma_\pm = 0.79$ was used, yielding $K = 6.1 \times 10^{-8}$. The indicators used in this investigation and their approximate pK values[15] at ionic strength 0.1 were methyl red (4.96), chlorophenol red (6.00), phenol red (8.0), cresol red (8.3), and phenolphthalein (9.7).

Kinetic Measurements. All kinetic runs were carried out on a temperature-jump relaxation spectrometer similar to that described by Czerlinski and Eigen.[16] The solution temperature was adjusted such that the final temperature (following the jump) was 25°. In each (acid and base) pH region only a single relaxation time was observed in the time range accessible (1 sec–5 μsec). The resultant relaxation curves were photographed with a 35-mm camera system. Their relaxation times were evaluated from enlargements made from the negatives. Blank experiments involving only the indicator system, or diacetylacetone solution without indicator, were also carried out.

Results.

Kinetic Measurements. Acidic Region. The results of experiments between pH 5.0 and 6.5 at 25° are shown in Table II. The over-all anion concentration [E]$_\Sigma$ (calculated from the pH and the equilibrium constant) is given in column 4. This number could consist of several different anionic species (cf. Table I).

It is clear that the relaxation time is a function of [H] + {[E]/(1 + α)} and not of [E] or [E]^2/(1 + α). Compare, for example, entries 1 and 5 in Table II, which have different enolate concentrations by a factor of 3 but whose relaxation times are within 10% of each other. The indicator correction term brings the two concentration functions (column 7) to about the same value. The differences in these correction terms lie in the pH differences and the change of the indicator (i.e., pK_{In} differences; see eq 4). A graph of τ^{-1} vs. [H] + {[E]/(1 + α)} yields a straight line, the slope of which is k_{13} = 2.6 $\times$ 10^6 M^{-1} sec^{-1}. An estimate may be made of K_{EH}, the equilibrium constant defined by eq 3, by calcu-

(11) G. Schwarzenbach and K. Lutz, *Helv. Chim. Acta*, **23**, 1162 (1940).

(12) Obtained from EGA-Chemie KG., Keppler and Reif, Steinheim/Albach, West Germany.

(13) G. Schwarzenbach, K. Lutz, and E. Felder, *Helv. Chim. Acta*, **27**, 576 (1944).

(14) See, e.g., C. W. Davies, "Ion Association," Butterworth Inc., Washington, D. C., 1962.

(15) In general, the pK values of the indicators were not available for the exact conditions of this investigation. Since, however, the literature values may vary by 0.2 or 0.3 pK unit, depending on the source, no attempt was made to correct for a slight difference in temperature or ionic strength.

(16) G. Czerlinski and M. Eigen, *Z. Elektrochem.*, **63**, 652 (1959).

Table II. Acid pH Region. Tabulation of Results

No.	$C°$, M	pH[a]	Indi-cator[b]	[E]$_\Sigma$ $\times$ 10^5, M	α	τ, msec	[H] + {[E]$_\Sigma$/(1 + α)} $\times$ 10^5, M
1	1.67 $\times$ 10^{-2}	5.00	MR	8.03	0.40	6.0	7.00
2	8.35 $\times$ 10^{-3}	5.00	MR	4.02	0.40	10.0	4.14
3	8.35 $\times$ 10^{-3}	5.08	MR	4.85	0.48	9.7	4.32
4	8.35 $\times$ 10^{-3}	5.60	CPR	15.7	1.60	5.5	6.4
5	8.35 $\times$ 10^{-3}	5.75	CPR	22.0	2.10	5.5	7.3
6	8.35 $\times$ 10^{-3}	5.85	CPR	27.6	2.80	4.7	7.45
7	8.35 $\times$ 10^{-3}	6.00	CPR	38.3	3.84	4.45	8.0
8	8.35 $\times$ 10^{-3}	6.03	CPR	41.2	4.10	4.4	8.2
9	8.35 $\times$ 10^{-3}	6.25	CPR	66.2	5.76	3.8	9.9
10	8.35 $\times$ 10^{-3}	6.52	CPR	115	7.70	3.1	13.2
11	4.17 $\times$ 10^{-3}	5.00	MR	2.01	0.40	14.0	2.70
12	4.17 $\times$ 10^{-3}	5.05	MR	2.25	0.44	9.0	2.69
13	2.08 $\times$ 10^{-3}	5.10	MR	1.25	0.50	16.9	1.84
14	1.67 $\times$ 10^{-3}	5.08	MR	0.97	0.48	17.0	1.71
15	8.35 $\times$ 10^{-4}	5.00	MR	0.40	0.40	31.5	1.56
16	8.35 $\times$ 10^{-4}	5.00	MR	0.40	0.40	16.4	1.56
17	4.17 $\times$ 10^{-4}	5.00	MR	0.20	0.40	18.0	1.41

[a] [H] was calculated from the measured pH by dividing the hydrogen ion activity by $\gamma_\pm$ ($\cong$0.79). [b] MR = methyl red, CPR = chlorophenol red; indicator concentrations = 2 $\times$ 10^{-5} M.

lating the value which would cause a just detectable deviation of the curve from a straight line. The result is $K_{EH} \geq 1.5 \times 10^{-3}$. The third term in eq 2 (and hence the conversion II–III) apparently does not make a measurable contribution to the relaxation time at the concentrations and pH's used.

Basic Region. The processes occurring in this region were considerably slower than in the acidic region; the measured relaxation times varied between approximately 1 sec and 70 msec (see Table III). The fourth column in Table III tabulates the total concentration of protonated species ([HX]$_\Sigma$) calculated from the over-all concentration and the measured pH. A close examination of eq 6 shows that only the first term fits the results in Table III. The fourth term varies directly as [E], the enolate concentration. Since above pH 8.5 virtually all the compound is present in this form, a pH-independent relaxation time would result. The third term can assume several limiting forms depending on which terms are large in the denominator, but in no case can the experimentally observed form be obtained. On the other hand, the first term in eq 6 (corresponding to the conversion I–III) fits the observed concentration and pH dependence exactly. A graph of τ^{-1} vs. $C°$ (over-all concentration) should yield a straight line at constant pH, the vertical intercept of which should be k_{31}[OH]. Furthermore, if $K'_{EH} \gg$ [EH] + [OH], a series of straight lines with slopes virtually independent of pH (since 1 + $\beta \cong$ 1) will be observed. Figure 1 shows that the linear relationship anticipated clearly holds for the results at pH 8.87. The ratio of the vertical intercept to the hydroxyl ion concentration is 2.1 $\times$ 10^5 M^{-1} sec^{-1}. If straight lines are drawn through the points at pH 9.2 and 9.4, the division of the observed intercepts by the hydroxyl ion concentration yields the values 2.8 and 2.7 $\times$ 10^5 M^{-1} sec^{-1}, respectively. The difference between these numbers is not considered significant; as a result it may be concluded that the of value k_{31} is (2.5 $\pm$ 0.4) $\times$ 10^5 M^{-1} sec^{-1}.

Table III. Basic pH Region. Tabulation of Results

No.	$C°$, M	pH[a]	Indicator[b]	$[HX]_\Sigma$[c] $\times 10^4, M$	β	τ, msec
1	3.35×10^{-2}	8.95	P	8.05	0.22	113,108
2	2.16×10^{-2}	8.12	CR (1×10^{-5})	31.0	1.49	217
3	2.16×10^{-2}	8.42	CR (1×10^{-5})	16.6	0.70	157
4	2.16×10^{-2}	8.72	CR (1×10^{-5})	8.75	0.27	160
5	2.16×10^{-2}	8.87	P	6.21	0.28	112
6	2.16×10^{-2}	9.00	P	4.65	0.26	100
7	2.16×10^{-2}	9.20	P	2.24	0.20	80
8	2.16×10^{-2}	9.37	P	2.02	0.16	65,90
9	1.67×10^{-2}	8.87	P	4.80	0.28	165
10	1.67×10^{-2}	8.93	P	4.22	0.27	125
11	1.67×10^{-2}	9.20	P	1.74	0.20	94
12	1.67×10^{-2}	9.37	P (1.5×10^{-5})	1.56	0.12	80
13	1.67×10^{-2}	9.40	P (2×10^{-5})	1.46	0.15	70
14	1.67×10^{-2}	9.38	P (2.5×10^{-5})	1.52	0.19	72, 74, 85
15	1.67×10^{-2}	9.37	P (4×10^{-5})	1.56	0.32	94, 96, 112
16	1.67×10^{-2}	9.38	P (1×10^{-4})	1.52	0.76	78, 88
17	1.67×10^{-2}	9.39	P (1×10^{-4})	1.49	0.76	74, 90
18	1.67×10^{-2}	9.40	P (2.5×10^{-4})	1.46	1.89	108
19	1.07×10^{-2}	8.30	PR (1×10^{-5})	10.7	0.80	290, 330
20	1.07×10^{-2}	8.28	PR (2×10^{-5})	11.1	1.72	305, 318
21	1.07×10^{-2}	8.35	PR (5×10^{-5})	9.65	3.42	435
22	8.35×10^{-3}	8.00	CR	15.1	3.74	1000
23	8.35×10^{-3}	8.34	CR	7.62	1.75	540
24	8.35×10^{-3}	8.72	CR	3.42	0.53	320
25	8.35×10^{-3}	8.80	P	2.93	0.30	160
26	8.35×10^{-3}	8.87	P	2.40	0.28	190
27	8.35×10^{-3}	8.90	P	2.25	0.28	170
28	8.35×10^{-3}	8.93	P	2.10	0.27	144,174
29	8.35×10^{-3}	9.20	P	0.87	0.20	125
30	8.35×10^{-3}	9.35	P	0.81	0.16	118
31	8.35×10^{-3}	9.40	P	0.73	0.14	83
32	8.35×10^{-3}	9.46	P	0.64	0.13	72
33	4.17×10^{-3}	8.88	P	1.18	0.28	296
34	2.08×10^{-3}	8.87	P	0.60	0.28	405, 350

[a] [OH] was calculated from the measured pH by dividing the hydroxyl ion activity by 0.79. [b] CR = cresol red, PR = propyl red, P = phenolphthalein; indicator concentrations, where not indicated in parentheses, were 2×10^{-5} M. [c] The symbol $[HX]_\Sigma$ is used to designate the total concentration of protonated forms; see Table I.

It can be easily shown that a *single* set of reaction partners (*i.e.*, a single enolate anion and one ketone and enol, respectively) as represented in mechanisms a and b is not sufficient to explain the kinetics in both the acidic and the basic regions. For example, if we assume the reaction in basic solutions is the hydrolysis given by mechanism b (where KH and EH now refer to a pair of related protonated forms of the molecule; see Table I), the slow relaxation time is given essentially by the first term in eq 6 with $K'_{EH} \gg [OH] + \{[EH]/(1 + \beta)\}$. We conclude therefore that $K'_{EH} = 1.4 \times 10^{-3}$ or larger.[17] This gives K_{EH} ($\equiv[H][E]/[EH]$) $\leq 7.1 \times 10^{-12}$.

Now if the same partners are to account for the relaxation process observed in the acidic pH range, we have simply mechanism a with τ^{-1} given by eq 2, and the terms involving k_{23} and k_{32} are negligible. The equilibrium constant K_{EH} for the rapid step I–II is fixed by the basic solution results, and so τ^{-1} reduces to

$$\tau^{-1} = k_{13}K_{12} + k_{31} = k_{31}(1 + K_{12}K_{13}) \cong k_{31} \quad (7)$$

Since $k_{31} = K_{13}k_{13} = k_{13}(6.8 \times 10^{-8})$, $\tau \cong 2 \times 10^7/k_{13}$. If k_{13} is between the (reasonable) values 10^4–10^7 M^{-1} sec^{-1} (see Figure 3), then τ would lie between 2 and 2000 sec and be independent of concentration and pH. Thus, the acidic analog of the basic relaxation time would not yield an observable relaxation in the acidic pH range.

(17) If only part of the protonated form exists as the ketone, [KH] = f[HX]$_\Sigma$ where f is the fraction of HX in the form KH.

Nmr Measurements. The proton resonance spectra of diacetylacetone (determined at room temperature on an A-60 nmr spectrometer) in the basic and acidic
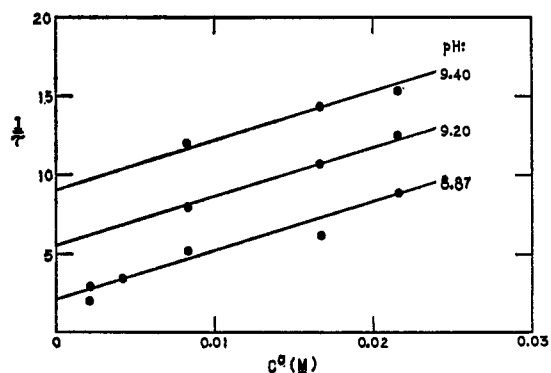
Figure 1. Variation of the relaxation time in basic solutions with the over-all concentration at several pH values.

pH regions are shown in Figure 2. The spectrum of a 1 M solution of acetylacetone at pH 4 was also run for comparison. The latter showed four peaks in the proper ratios to account for 85% ketone and 15% enol in aqueous solution. The resonances for the groups $CH_3C=$, CH_3C-, $-CH_2-$, and $-CH=$ occur at 1.95, 2.14, 3.72, and 5.60 ppm, respectively, relative to TMS. The spectrum for the anions of diacetylacetone

Figure 2. Nmr spectra of 0.5 M diacetylacetone: upper spectrum, pH 4.0; lower spectrum, pH 10.0 (water resonances are not shown).

indicates also only four peaks, associated with the above-mentioned groups in the ratio 3:3:2:1. The only structure consistent with the latter is the anionic form KHE⁻ (see Table I); *i.e.*, of the three anions possible, KHE⁻ is the only one detectable[18] by nmr. The more complicated acidic spectrum shows once more two different methyl peaks, but in addition there are several other peaks, the integrals of which are in no simple proportions to each other. The only conclusion that can be reached is that the protonated species exists as a mixture of various forms.

Discussion

While it is clear that one of the two relaxation times observed is associated with the direct protonation of the enolate anion and the other with the hydrolysis of one or more protonated forms, it is desirable to relate the relaxation phenomena to specific species.

The mechanism interrelating all species in Table I is given for the basic pH region by (c) where the slow

steps are represented by arrows and the rapid equi-

(18) Presumably the species EHE₁⁻ and EHE₂⁻ are also present, but at much smaller concentrations.

libria by = signs. The analogous reaction scheme for the acidic region is (d). Although there are eight equilibria shown in these two mechanisms, there can be no more than seven relaxation times for each, since the circular interconnection reduces the number of independent concentration variables by one. There are only two long relaxation times: one is associated with the isolated slow step I–III (and the equilibria preceding it); the other is associated with the "parallel" pathways[19] II–VII and IV–V. The task remains of making an assignment of the observed relaxation times to specific steps. The following observations are relevant.

1. Only two relaxation times were observed, one each in the basic and acidic pH regions. If, for example, the effect in basic solutions corresponds to the processes II–I–III, it is clear by the analysis above (eq 7 and text) why the exact acidic counterpart ($KHEH_2$ = KHE^- + H^+ $\rightleftharpoons$ $KHKH$) would not be observed in the acidic solutions. Thus, of the four slow steps in mechanisms c and d, all are needed to account for the two times.

2. The magnitude of the relaxation effect is *ca.* tenfold larger in the acidic region than in the basic region. In both pH regions, the size of the effects decreases as the quantities $[HX]_\Sigma$ and $[E]_\Sigma$ decrease relative to [OH] and [H], respectively.[20] In the acidic region, the relaxation effect disappears just as $[E]_\Sigma$ decreases to become virtually identical with [H]. The effect in the basic region, however, is much smaller to begin with and disappears well before the sum of the protonated species becomes small compared to [OH].

Both observations are consistent with the assignment of the relaxation effect in acidic solutions to equilibria involving major constituents; in basic solutions, to equilibria involving minor constituents. The best assignment of relaxation times is that shown by the position of τ_a (acidic) and τ_b (basic) in mechanisms c and d. The final mechanism in the acidic region then becomes

the slow relaxation time (τ_a) of which is given by eq 5, where now

$$\frac{1}{K_{EH}} = \frac{1}{K_{12}} + \frac{1}{K_{14}} = \frac{[KHEH_1] + [KHEH_2]}{[H][KHE]}$$

Under the condition $K_{EH} \gg [H] + [KHE]/(1 + \alpha)$, this reduces to the experimentally observed form. It is also apparent that since KHE⁻ is the only major anion present (see nmr results), its concentration is virtually equal to the total anion concentration. We conclude therefore that the rate constant k_{13} has the value 2.6 × 10^6 M^{-1} sec⁻¹ and $K_{EH} \geq 1.5 \times 10^3$ (see acid results).

The interpretation of the relaxation observed in the basic region is somewhat more involved. If one

(19) They are effectively parallel because of the rapid interconversion between states VII and V on the one hand and between II and IV on the other.

(20) The optimum conditions for observing a reaction coupled to a pH indicator are [OH] << [HX] in the basic region and [H] << [E] in the acidic region, where the symbols HX and E⁻ refer to the reactants which yield the relaxation effect.

$$\tau_{b}^{-1} = \frac{\bar{k}[OH]\left[K'_{14}\frac{[KHEH_1]}{1+\beta}(1+K_{42})(1+K_{57}) + K'_{14}K'_{56} + (1+K_{57})K'_{14}\frac{[EHEH_2]}{1+\beta} + [OH]\right]}{[OH]\left\{(1+K_{57})\frac{[KHEH_1](1+K_{42}) + [EHEH_2]}{1+\beta} + [OH] + K'_{14}(1+K_{42}) + K'_{56}(1+K_{42})\right\} + R} +$$

$$\frac{\bar{k}[OH]\left[\frac{[KHEH_1](1+K_{42}) + [EHEH_2]}{1+\beta} + [OH] + K'_{12}(1+K_{42})\right]}{[OH]\left\{(1+K_{57})\frac{[KHEH_1](1+K_{42}) + [EHEH_2]}{1+\beta} + [OH] + K_{14}(1+K_{42}) + K'_{56}(1+K_{42})\right\} + R} \tag{8}$$

ignores the equilibrium V–Va in mechanism d, the relaxation time τ_b becomes[21] as shown in eq 8 where the equilibrium constants are defined as

$$K'_{14} = \frac{[KHEH_1][OH]}{[KHE]} \qquad K_{57} = \frac{[EH_2E]}{[EH_1E]}$$

$$K'_{56} = \frac{[EHEH_2][OH]}{[EH_1E]} \qquad K_{42} = \frac{[KHEH_2]}{[KHEH_1]}$$

and the effective rate constants (characterizing the parallel pathways IV–V and II–VII) are given by $\vec{k} = k_{45} + K_{42}k_{27}$ and $\bar{k} = k_{54} + K_{57}k_{72}$. The term R in the denominator contains all terms not multiplied by [OH]

$$R = (1 + K_{42}) \times$$

$$\left[\frac{K'_{14}(1 + K_{57})[EHEH_2] + K'_{56}[KHEH]}{1+\beta} + K'_{14}K'_{56}\right]$$

Equation 8 may reduce to several forms depending on the assumptions made concerning the relative values of the terms in the denominator. It is clear that the R term must be large compared to those multiplied by [OH$^-$] in order that the experimentally observed form (eq 6) be obtained. In addition, if $K'_{14}K'_{56}$ is the largest term[22] in the brackets in the expression for R, there results finally

$$\tau_b^{-1} = \frac{\vec{k}[OH]}{1 + K_{42}} \times$$

$$\left[\frac{(1 + K_{57})[EHEH_2] + [KHEH_1] + [KHEH_2]}{K'_{56}(1+\beta)} + 1\right] +$$

$$\frac{k[OH]}{K'_{14}K'_{56}}\left[\frac{[KHEH_1]}{1+\beta} + \frac{[EHEH_2]}{(1+K_{42})(1+\beta)} + K'_{14}\right] \tag{9}$$

This equation now has the correct concentration dependence. From Figure 1 the over-all rate constant $k_{ov} = \bar{k}/(1 + K_{42}) = 2.5 \times 10^5\ M^{-1}\ \text{sec}^{-1}$ and $1 + K_{57}/K'_{56} = 7.5 \times 10^3$. It is not possible to decide unambiguously which of the terms ($k_{45} + K_{42}k_{27}$) in $\vec{k}$ makes the larger contribution.

It might be asked if the assignment of relaxation times in the acidic and basic regions could just as well be reversed; i.e., is there any ambiguity in the assignment of τ_a and τ_b? If one were to interchange the two times, the following is immediately apparent. The time τ_b could indeed be accounted for by step I–III coupled to I–IV and I–II. The acidic τ, however, can easily be shown to be completely inconsistent with the
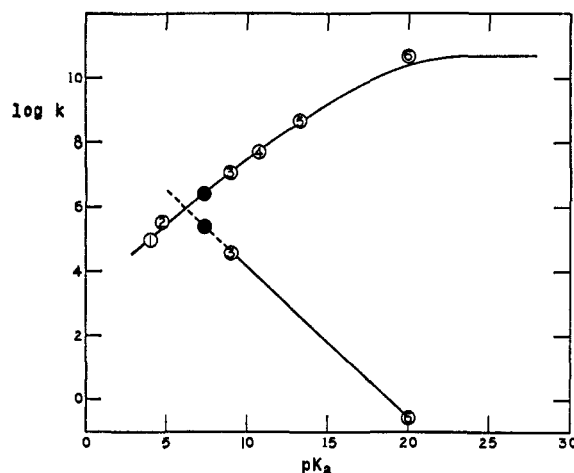

Figure 3. Brønsted plot for acid and base rate constants in some ketones: solid circles, diacetylacetone; others, (1) barbituric acid, (2) Meldrum's acid, (3) acetylacetone, (4) ethylacetylacetone, (5) diethylmalonic acid, (6) acetone. Upper curve is k_a, lower is k_b.

parallel paths V–IV and VII–II. If one again neglects the equilibrium V–Va, an expression for τ_a analogous to eq 8 is obtained where now $\vec{k} = k_{57} + K_{42}k_{27}$. In order to obtain an equation of the proper form, it is necessary to assume that an "R" term[23] in the denominator is negligible. This requires, among others, that $K_{56} \ll [H]$. On the other hand, it is also required that $K_{56} \geq 1.5 \times 10^{-3}$ (see acid results), a result which is irreconcilable with the lower limit of K_{56} necessary in the denominator. Thus we conclude that the assignment of the relaxation times in the acidic and basic regions can be unambiguously made.

It is now well established that a correlation exists between the rate constants for proton transfer in pseudo-acids and the pK_a of the system (Brønsted relation). Figure 3 shows a Brønsted plot for a series of closely chemically related pseudo-acids.[24] Most of the rate constants have been directly measured;[3,5] i.e., they are not obtained from the ionization rate and the pK_a. The acid recombination rate constants all lie on a smooth curve, with the result for diacetylacetone falling precisely where expected. Data are more scanty for the basic reaction (HB + OH$^-$); it is clear, however, that the basic rate constant involving diacetylacetone correlates well with a linear extrapolation of the other points. This correlation is remarkable inasmuch as the value measured was an effective or over-all

(21) See Appendix.
(22) This requires that $K'_{56} \gg (1 + K_{57})[EHEH_2]$ and $K'_{14} \gg [KHEH_1]$.

(23) Containing all terms not multiplied by [H].
(24) If all rate data for acids of varying chemical composition (e.g., trifluoroacetylacetone, nitromethane) were examined, the qualitative correlation would still exist, but several curves would be necessary to adequately represent all the data. For example, the reaction of OH$^-$ with HCN ($pK = 9.1$) is just 10-fold less than diffusion-controlled: J. Stuehr, et al., J. Chem. Phys., 38, 587 (1963).

rate constant (see eq 8). There are at least three explanations: (1) the correlation is purely fortuitous; (2) the pK of the molecules involved is also about 7.4; or (3) the basic reaction involves the same molecule or anion[25] as the acidic. The latter is eliminated, however, by the fact that an examination of mechanism c shows that there is no anion or molecule common to two slow processes. For the second possibility, we must have a single step predominating, a condition which is fulfilled either if $K_{42} \gg 1$ (in which case $k_{eff} \cong k_{27}$) or $K_{42} \ll 1$ ($k_{eff} \cong k_{45}$).

In such a relatively complicated system, it was clearly necessary to have carried out measurements in both pH regions. If measurements had been restricted to acidic solutions, one would have concluded that virtually any two-step mechanism could account for the results.[26] The kinetic parameters for acidic and basic diacetylacetone are summarized in Table IV. Thus, by carrying out and interpreting measurements in both pH regions, the general features of the protonation-deprotonation mechanism could be established with a high degree of certainty.

Table IV. Equilibrium and Rate Constants for Aqueous Diacetylacetone at 25° and 0.1 Ionic Strength

Acidic	Basic
$K_\Sigma = 6.8 \times 10^{-8}$	
$\left.\begin{array}{l} k_{14} \\ k_{12} \\ k_{56} \\ k_{76} \end{array}\right\}$ all $\sim 3 \times 10^{10}\,M^{-1}\,sec^{-1a}$	$\left.\begin{array}{l} k_{65} \\ k_{67} \\ k_{41} \\ k_{21} \end{array}\right\}$ all $\leq 3 \times 10^{10}\,M^{-1}\,sec^{-1b}$
$k_{13} = 2.6 \times 10^8\,M^{-1}\,sec^{-1}$	$\dfrac{k}{1 + K_{42}} = 2.5 \times 10^5\,M^{-1}\,sec^{-1c}$
$K_{14} \geq 1.5 \times 10^{-3}$	
$K_{12} \geq 1.5 \times 10^{-3}$	

[a] Estimated on the basis of diffusion-controlled reactions. [b] Internally hydrogen-bonded molecules hydrolyze with rates ranging from the diffusion-controlled limit to considerably more slowly. [c] Over-all rate constant defined by eq 8.

(25) But not both; see eq 7 and related text.
(26) For example, a mechanism consisting of *one* protonated form and *two* anions ($E_1^- + H^+ = KH \rightleftharpoons E_2^- + H^+$) is easily shown to be compatible with the acid results. The relaxation time for an analogous hydrolysis, however, cannot be fitted to the observed concentration dependence.

Appendix

Derivation of Eq 8. For a small perturbation from equilibrium ($\delta C_i \ll C_i$), the rate of restoration of equilibrium for the parallel paths IV–V and II–VII is given by

$$\frac{d(\delta C_5 + \delta C_6 + \delta C_7)}{dt} = -\frac{d(\delta C_1 + \delta C_4 + \delta C_6)}{dt} =$$
$$k_{45}(C_4 \delta C_{OH} + C_{OH}\delta C_4) + k_{27}(C_2 \delta C_{OH} +$$
$$C_{OH}\delta C_2) - k_{54}\delta C_5 - k_{72}\delta C_7 \quad (A-1)$$

where the symbols C_i refer to the concentrations of the chemical species indicated in mechanism c. This equation may be put in the form

$$\frac{d\delta C_i}{dt} = -\frac{1}{\tau}\delta C_i \quad (A-2)$$

by using the conservation relationships

$$\sum_{i=1}^{i=6} \delta C_i = 0 \,(i \neq 3) \quad \text{(mass conservation)} \quad (A-3)$$

$$\delta C_2 + \delta C_4 + \delta C_6 + \delta C_{H_2O} = 0 \quad \text{(proton conservation)} \quad (A-4)$$

$$\delta C_{OH} + \delta C_{H_2O} = 0 \quad \text{(hydroxyl conservation)} \quad (A-5)$$

and the preequilibrium relationships

$$K_{56}\delta C_5 = C_6 \delta C_{OH} + C_{OH}\delta C_6 \quad (A-6)$$
$$K_{14}\delta C_1 = C_4 \delta C_{OH} + C_{OH}\delta C_4 \quad (A-7)$$
$$\delta C_2 = K_{42}\delta C_4 \quad (A-8)$$
$$\delta C_7 = K_{57}\delta C_5 \quad (A-9)$$

The elimination of all δC_i variables except one from eq A-3–A-9 and substitution into A-1 yields an equation of the form of eq A-2, from which τ^{-1} (eq 8) follows directly.