Metal Ion and Metal Chelate Catalyzed Oxidation of Ascorbic Acid by Molecular Oxygen. II. Cupric and Ferric Chelate Catalyzed Oxidation[1,2]

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Abstract: The kinetics of the oxidation of ascorbic acid, in the presence of cupric and ferric chelate compounds as catalysts, are determined at 25 and 0.4°. In the pH range 2.25–3.45, the catalytic contributions of the hydrolyzed and polymeric forms of the metal chelate compound were found to be negligible. The catalytic activities of the cupric chelates decrease in the order, iminodiacetatocopper(II) > 2-hydroxyethyliminodiacetatocopper(II) > ethylenediaminetetracetatocuprate(II). The catalytic activity of ferric chelates decrease in the order 2-hydroxyethyliminodiacetatoiron(III) ion > nitrilotriacetatoiron(III) > 2-hydroxyethylethylenediaminetriacetatoiron(III) > ethylenediaminetetracetatoferrate(III) > cyclohexanediaminetetracetatoferrate(III) > diethylenetriaminepentacetatoferrate(III). The catalytic activities of the cupric and ferric chelates were found to be independent of oxygen concentration. A linear correlation was found between the stability of the metal chelate compound and its catalytic activity. The differences in the reactivities of the various metal chelate compounds studied are interpreted on the basis of the energetics of the reaction.

In part I of this series[4] cupric and ferric ion catalyzed oxidation of ascorbic acid was reported. The present paper describes similar studies with the chelate compounds of these metal ions. Though the study of metal ion catalyzed oxidation of ascorbic acid has been the subject matter of several investigations, metal chelate catalyzed oxidation has received relatively little attention. Recently, Grinstead[5] described the catalytic effect of the Fe(III)–EDTA chelate on the oxidation of ascorbic acid by H_2O_2. This work was extended[6] to the study of the catalytic effect of a model peroxidase system, consisting of the Fe(III)–EDTA chelate and ascorbic acid, on the oxidation of salicylate by H_2O_2. The purpose of the present work was to examine differences in the reactivities of cupric and ferric chelate compounds as catalysts in ascorbic acid oxidation, to determine the energetics of the reaction, and to provide new information on the reaction mechanism of this important reaction.

Experimental Section

Reagents. The *l*-ascorbic acid used in this investigation was Kodak White Label grade and was used without further purification. The commercial disodium salt of EDTA was recrystallized twice from water, and a measured amount of the resulting dihydrate was dissolved in distilled water to make up a stock solution. The calculated molarity of this solution was checked by titration with standard base. Commercial samples of 2-hydroxyethylethylenediaminetriacetic acid (HEDTA), cyclohexanediaminetetraacetic acid (CDTA), diethylenetriaminepentaacetic acid (DTPA), 2-hydroxyethyliminodiacetic acid (2-HIMDA), iminodiacetic acid (IMDA), and nitrilotriacetic acid (NTA) were recrystallized from water and dried under vacuum. The equivalent weights of the pure acids were determined by potentiometric titration. Solutions of Cu(II) and Fe(III) nitrates were prepared from Fisher analytical grade materials. The Cu(II) solution was standardized both by an iodometric procedure with standard sodium thiosulfate and by titration in ammoniacal solution with standard disodium salt of EDTA, according to the procedure described by Schwarzenbach.[7] The Fe(III) solution was standardized by an oxidation reduction titration with potassium permanganate and also by an EDTA titration[7] with Tiron as indicator. The results of the two methods agree within the experimental error.

Potentiometric Measurements. The dissociation constants of *l*-ascorbic acid at 25 and 0.4° and IMDA, 2-HIMDA, and NTA at 0.4° were determined by potentiometric titration in a medium of 0.10 *M* ionic strength containing potassium nitrate. The formation constants at 0.4° for Cu(II)–NTA, Cu(II)–IMDA, and Cu(II)–HIMDA were determined by titrating solutions containing 1:1 molar ratios of ligand to metal ion, the ionic strength being kept constant at 0.10 *M* by the addition of an appropriate amount of 1.00 *M* potassium nitrate. A Beckman Model G pH meter, fitted with extension glass and calomel electrodes, was used. The pH meter was calibrated in terms of hydrogen ion concentration with acetic acid buffer as well as with standard HCl and sodium hydroxide. The data given by Harned and Owen[8] was used in the acetic acid buffer range to determine the hydrogen ion concentration. The solution of ascorbic acid was prepared in air-free distilled water, and an atmosphere of purified nitrogen was maintained in the titration cell to avoid any disturbing effects resulting from oxidation.

Kinetic Measurements. The pH value of the experimental solution was maintained constant during each run by a Beckman Model K automatic titrator fitted with extension glass and calomel electrodes. It was calibrated with acetic acid buffer and by titration of standard HCl and NaOH solutions as in the case of the equilibrium studies. The ionic strength of the experimental solution was maintained at approximately 0.10 *M* with KNO_3. After adjusting the pH to the desired value, a stream of oxygen was passed through the cell through a fritted glass tube so as to ensure very intimate contact between gas phase and solution. The oxygen was 99% pure and was freed from CO_2 by being passed through an ascarite tube. It was presaturated with water vapor by streaming through a wash bottle maintained at the same temperature and ionic strength as the reacting solution. As the rate of reaction is slow compared

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(4) M. M. T. Khan and A. E. Martell, *J. Am. Chem. Soc.*, 89, 4176 (1967).

(5) R. Grinstead, *ibid.*, 82, 3463 (1960).

(6) R. Grinstead, *ibid.*, 82, 3472 (1960).

(7) G. Schwarzenbach, "Complexometric Titrations," Interscience Publishers, Inc., New York, N. Y., 1958, pp 77–82.

(8) H. S. Harned and B. B. Owen, "Physical Chemistry of Electrolytic Solutions," 3rd ed, Reinhold Publishing Corp., New York, N. Y., 1958, pp 638–752.

Table I. Dissociation Constants of IMDA, 2-HIMDA, and NTA at 0.4° [a]

Ligand	pK_1	pK_2	pK_3	Log K_{CuL} [b]
IMDA	2.84	10.70	...	11.70
2-HIMDA	2.44	9.92	...	12.00
NTA	2.30	3.00	10.76	13.11

[a] μ = 0.10 M (KNO_3). [b] K_{CuL} = formation constant of 1:1 Cu-ligand chelate.

centration of Cu(II) chelate at 25 and 0.4° indicates true catalytic behavior for the Cu(II) chelates. Similar results were obtained at other pH values studied, and at 0.4°. The rate constants given in Table II were calculated from the slopes of the straight lines shown in Figure 1 and from similar curves obtained at other pH values. The data for a particular straight line were obtained by measuring the rate over a wide range of catalyst concentration.

Table II. Rate Constants (M^{-1} sec^{-1}) for Cu(II) Chelate Catalyzed Oxidation of Ascorbic Acid[a]

Temp, °C	−Log [H+]	Cu-EDTA	Cu-HEDTA	Cu-NTA	Cu-HIMDA	Cu-IMDA
25	2.00	0.09×10^1	0.40×10^1	0.10×10^2	0.15×10^2	0.30×10^2
	2.25	0.60×10^1	0.14×10^2	0.18×10^2	0.27×10^2	0.53×10^2
	2.85	2.2×10^1	0.56×10^2	0.68×10^2	1.0×10^2	2.5×10^2
	3.45	0.80×10^2	1.8×10^2	2.3×10^2	3.4×10^2	6.8×10^2
0.4	2.85	1.10×10^1	0.30×10^1	0.60×10^1	1.4×10^1	2.3×10^1
	3.25	0.20×10^1	0.80×10^1	0.16×10^2	2.5×10^1	5.7×10^1
	3.45	0.30×10^1	1.2×10^1	0.20×10^2	3.0×10^1	7.0×10^1
	3.85	1.2×10^1	2.6×10^1	0.55×10^2	8.8×10^1	1.8×10^2

[a] μ = 0.10 M (KNO_3); O_2 pressure = 1.0 atm.

Table III. Rate Constants (M^{-1} sec^{-1}) for the Cu(II) Chelate Catalyzed Oxidation of the Ionic Species of Ascorbic Acid[a]

Metal chelate	25° k_2	25° k_1	0.4° k_2	0.4° k_1
Cu(II)-EDTA (CuL^{2-})	1.0×10^2	0	0.44×10^2	0
Cu(II)-HEDTA (CuL^-)	4.1×10^2	0	1.4×10^2	0
Cu(II)-NTA (CuL^-)	11×10^2	0	2.7×10^2	0
Cu(II)-HIMDA (CuL)	17×10^2	0	4.8×10^2	0
Cu(II)-IMDA (CuL)	33×10^2	0	0.9×10^3	0

[a] μ = 0.10 M (KNO_3); O_2 pressure = 1 atm.

to the rate of dissolution of oxygen, the reacting solution was considered to be saturated with oxygen at all times. The rate of oxidation was measured by the amount of dehydroascorbic acid produced during the course of oxidation.

Procedure. The analytical procedure employed for the estimation of dehydroascorbic acid was that established by Roe and coworkers[9] and is described in detail in a previous paper.[4]

Results

Equilibrium Studies. Dissociation of Ascorbic Acid. The pK values of ascorbic acid were calculated from the titration curves at 25 and 0.4°. The dissociation constants obtained are 9.16×10^{-5} and 3.24×10^{-5} for 25 and 0.4°, respectively, for K_1, and 4.57×10^{-12} and 1.91×10^{-13} at 25 and 0.4°, respectively, for K_2.

Chelate Stability Constants. The dissociation constants of IMDA, HIMDA, and NTA at 0.4° and the formation constants of 1:1 Cu(II) chelates are listed in Table I. These data have not been reported previously. Acid dissociation constants and chelate stability constants at 25° were taken from the tables of stability constants.[10]

Kinetic Studies. Cu(II) Chelate Catalyzed Oxidation. The rate data obtained for the Cu(II) chelate catalyzed oxidation reactions are given in Tables II and III. The variation of rate at 25° with the concentration of Cu(II) chelate at pH 3.45 is indicated in Figure 1. The linear variation of rate with the con-

(9) J. H. Roe, "Methods of Biochemical Analysis," Vol. I, Interscience Publishers, Inc., New York, N. Y., 1954, pp 115–139.
(10) L. G. Sillen and A. E. Martell, "Stability Constants," The Chemical Society, London, 1964.

The rate at a particular pH was corrected for the concentration of free metal ion by the help of the fol-

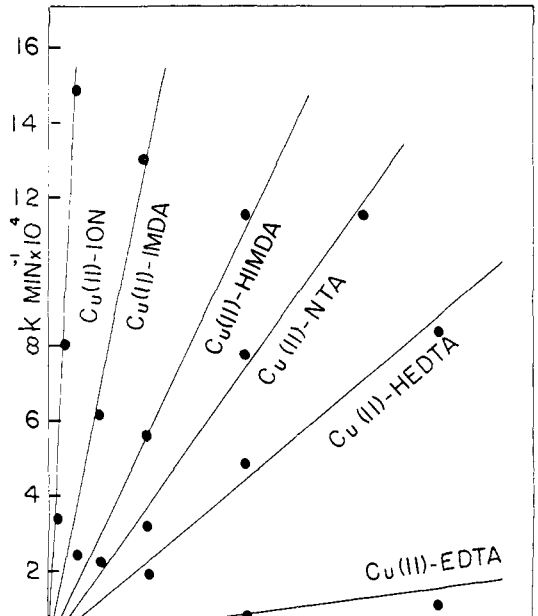


Figure 1. Catalytic effect for the oxidation of ascorbic acid in the presence of Cu(II) chelates at 25°, at a −log [H+] value of 3.45; k = difference between the first-order rate constants in the presence and in the absence of the metal chelate compound; μ = 0.10 M (KNO_3).

lowing equations

$$[M] = \alpha[L] \quad (1)$$

$$[L] = \frac{-\alpha \pm \sqrt{\alpha^2 - 4K_{ML}{}^{M}T_{ML}}}{2K_{ML}{}^{M}\alpha} \quad (2)$$

Table IV. Rate Constants (M^{-1} sec^{-1}) for Fe(III) Chelate Catalyzed Oxidation of Ascorbic Acid[a]

−Log [H+]	Fe-DTPA	Fe-CDTA	Fe-EDTA	Fe-HEDTA	Fe-NTA	Fe-HIMDA
2.00	0.8×10^1	1.2×10^1	1.7×10^1	5.4×10^1	0.7×10^2	0.9×10^2
2.25	1.4×10^1	2.1×10^1	2.8×10^1	1.0×10^2	1.3×10^2	1.8×10^2
2.45	2.3×10^1	3.3×10^1	4.4×10^1	1.6×10^2	2.1×10^2	2.8×10^2
3.00	2.7×10^1	4.1×10^1	5.4×10^1	1.9×10^2	2.9×10^2	3.5×10^2
2.25	0.3×10^1	0.5×10^1	0.7×10^1	2.2×10^1	2.8×10^1	3.4×10^1
2.45	0.9×10^1	1.3×10^1	1.7×10^1	6.0×10^1	0.8×10^2	1.1×10^2
2.85	1.4×10^1	2.0×10^1	2.6×10^1	0.9×10^2	1.1×10^2	1.4×10^2
3.45	5.1×10^1	0.7×10^1	1.0×10^2	3.2×10^2	4.2×10^2	5.1×10^2

[a] μ = 0.10 M (KNO_3); O_2 pressure = 1.0 atm; temperature 25°.

where

$$\alpha = \frac{[H^+]^n}{\prod_{n=1}^{n} K_n} + \frac{[H^+]^{n-1}}{\prod_{n=2}^{n} K_n} + \frac{[H^+]^{n-2}}{\prod_{n=3}^{n} K_n} + \ldots + 1 \quad (3)$$

$K_{ML}{}^M$ = formation constant of the chelate; K_n = acid dissociation constants of ligand; T_L = total concentration of 1:1 chelate species; [L] = concentration of completely ionized ligand species; H_nL = neutral ligand, and [M] = concentration of free metal ion.

The concentration of free metal ion at 25 and 0.4° was calculated from the equilibrium data in the stability constant tables[10] and Table I, respectively. The Cu(II)–EDTA and Cu(II)–HEDTA chelates were completely formed in the pH range studied, the concentration of the free metal ion M being negligible in both cases.

The assignments of the rate constants for the Cu(II) chelate catalyzed oxidation were made to the neutral and the monoionic species of ascorbic acid from the experimental data in the pH range studied (2.25–3.85). This was based on the assumption that each species of ascorbic acid reacts independently with the metal chelate compound. The rate constant k at a particular pH was found to vary linearly with the concentration of the monoionic species of ascorbic acid, the contribution of the un-ionized form being zero. This can be readily verified by a plot of k against the reciprocal of the hydrogen ion concentration (Figure 4). The intercept k_1 was found to be zero. From the slope k_2 was obtained. The constants k_2 and k_1 obtained at 25 and 0.4° are listed in Table III.

In this investigation the catalytic activities of Cu–dipyridyl (1:1), Cu–*o*-phenanthroline (1:1), Cu–EDTA (1:1), Cu–HEDTA (1:1), Cu–NTA (1:1), Cu–HIMDA (1:1), and Cu–IMDA (1:1) were measured and compared. The 1:1 Cu(II)–α,α-dipyridyl and 1:1 Cu(II)–*o*-phenanthroline chelates were found to be completely inactive. This may be due to the stabilization of the lower oxidation state of copper by α,α-dipyridyl and *o*-phenanthroline. In the case of the other chelates studied, the catalytic activity decreases in the order Cu–IMDA > Cu–HIMDA > Cu–NTA > Cu–HEDTA > Cu–EDTA. The rate seems to decrease with increasing stability and increasing negative charge on the metal chelate compounds. In the pH range studied, many of the copper chelates of polyamines and amino acids have a very low degree of formation and their catalytic activities could not be measured.

Ferric Chelate Catalyzed Oxidation. The rate data for the Fe(III) chelate catalyzed oxidation of ascorbic acid are given in Table IV. As in the case of Cu(II) chelate catalyzed oxidation, a linear variation of rate with the concentration of chelate species was observed. The variation of rate with the concentration of Fe(III) chelate at 25° and at −log [H+] = 2.45 is indicated in Figure 2. Similar results were obtained at other pH values studied, and at 0.4°. The rate constants given in Table IV were obtained from the slopes of the straight lines in Figure 2, and similar curves were obtained at other pH values. The results at 0.4° were obtained by a similar procedure.

In all the systems studied, the concentration of free metal ion was found to be negligible compared to the concentration of the chelate species. The concentration of the chelate species was, however, corrected for hydrolysis at a particular pH. For the calculation of the unhydrolyzed chelate species, the hydrolysis constants reported by Vandegaer, *et al.*,[11] Gustafson and Martell,[12] and Sillen and Martell[10] were employed.

The assignments of the specific rate constants for the Fe(III) chelate catalyzed oxidation reactions were made to the un-ionized (k_1) and the monoionic species (k_2) of ascorbic acid from the experimental data in the pH range 2.25–3.45. At a particular pH, the rate constant k was found to vary linearly with the concentration of the monoionic species of ascorbic acid, the contribution of the un-ionized form being zero. This may be readily verified by a plot of k against the reciprocal of the hydrogen ion concentration. The intercept k_1 was found to be zero. From the slope of the straight line k_2 was obtained. The constants k_2 and k_1 obtained at 25 and 0.4° are listed in Table V. At a particular pH the average deviation of the specific rate k from the calculated rates was found to be 4–7%. This may be easily verified from Figure 5. The slightly larger deviation in the specific rates compared to those (2–5%)

Table V. Rate Constants (M^{-1} sec^{-1}) for the Fe(III) Chelate Catalyzed Oxidation of the Ionic Species of Ascorbic Acid at 25 and 0.4° [a]

	25°		0.4°	
Metal chelate	k_2	k_1	k_2	k_1
Fe(III)–DTPA	9×10^2	~0	6×10^2	~0
Fe(III)–CDTA	13×10^2	~0	9×10^2	~0
Fe(III)–EDTA	18×10^2	~0	12×10^2	~0
Fe(III)–HEDTA	60×10^2	~0	39×10^2	~0
Fe(III)–NTA	8×10^3	~0	51×10^2	~0
Fe(III)–HIMDA	11×10^3	~0	61×10^2	~0

[a] μ = 0.10 M (KNO_3); O_2 pressure = 1 atm.

(11) J. Vandegaer, S. Chaberek, and A. E. Frost, *J. Inorg. Nucl. Chem.*, 11, 210 (1959).

(12) R. L. Gustafson and A. E. Martell, *J. Phys. Chem.*, 67, 576 (1963).

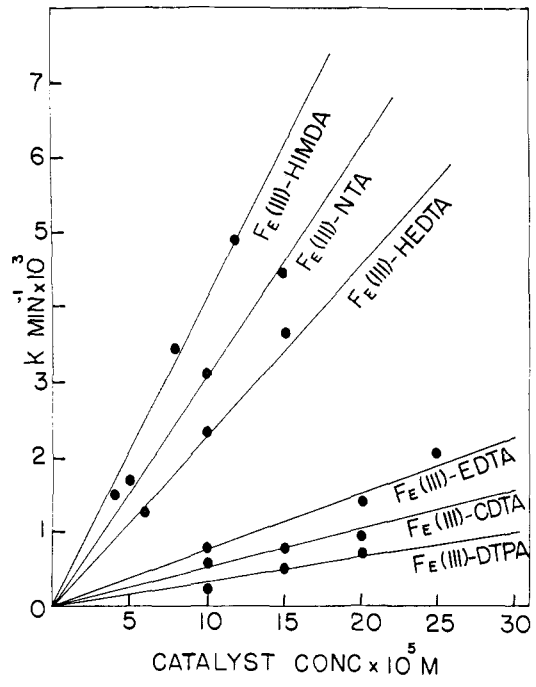


Figure 2. Catalytic effect for the oxidation of ascorbic acid in the presence of Fe(III) chelates at 25°, at a $-\log [H^+]$ value of 2.45; k = difference between the first-order rate constants in the presence and in the absence of the metal chelate compounds; μ = 0.10 M (KNO_3).

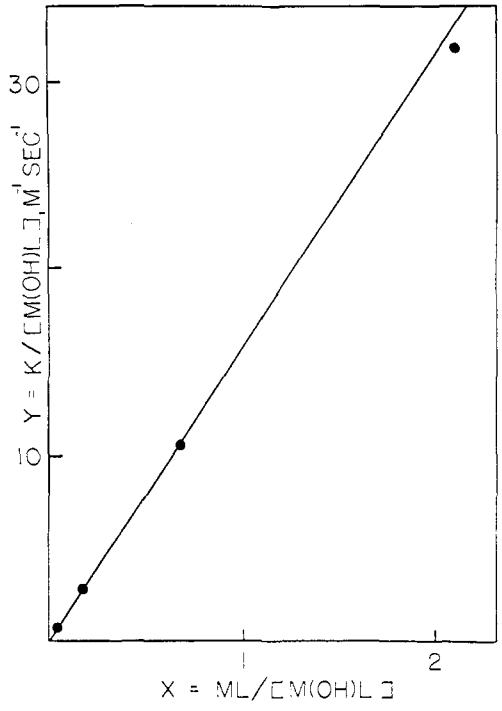


Figure 3. Plot demonstrating the inactivity of the hydroxo species of Fe(III)-HIMDA as a catalyst in the oxidation of ascorbic acid at 25°; μ = 0.10 M (KNO_3).

of Cu(II) chelates (Figure 4) may be due to a slight contribution of the hydrolyzed species in the catalysis of the oxidation reaction. The largest deviations (7%) were found in Fe(III)-HIMDA and Fe(III)-NTA systems for which no precise hydrolysis data are known.[12]

The catalytic activities indicated in Table V decrease in the order Fe(III)-HIMDA > Fe(III)-NTA > Fe(III)-HEDTA > Fe(III)-EDTA > Fe(III)-CDTA > Fe(III)-DTPA. There seems to be a regular decrease in the catalytic activity with the increasing stability and increasing negative charge on the metal chelate compounds.

Activities of Hydrolyzed Chelate Species. In the calculation of the rate constants, the concentration of the hydrolyzed chelate species was subtracted from the total metal chelate concentration. It was found that the hydrolyzed species are inactive as catalysts for the oxidation of ascorbate anion, HA^-, by the following calculation. If k_{ML} and $k_{M(OH)L}$ are the rate constants corresponding to the unhydrolyzed and hydrolyzed species ML and M(OH)L, respectively, then the observed first-order rate constant k_{obsd} may be expressed as

$$k_{obsd} = k_{ML}[ML] + k_{M(OH)L}[M(OH)L] \quad (4)$$

$$\frac{k_{obsd}}{[M(OH)L]} = \frac{[ML]}{[M(OH)L]}k_{ML} + k_{M(OH)L} \quad (5)$$

A plot of $k_{obsd}/[M(OH)L]$ should give a straight line with slope = k_{ML} and intercept $k_{M(OH)L}$. Such a plot is shown in Figure 3 for the 1:1 Fe(III)-HIMDA system. It may be seen from Figure 3 that the intercept is zero, which indicates that the hydroxo species are relatively inactive as catalysts in the oxidation of ascorbate anion. The concentration of the unhydrolyzed species ML of the Fe(III)-HIMDA complexes was calculated by means of the equation

$$[ML] = [ML_T]\left(1 + \frac{K_1}{[H^+]} + \frac{K_1K_2}{[H^+]^2}\right)^{-1} \quad (6)$$

where $[ML_T]$ = total concentration of the metal chelate species. K_1 and K_2 are the first and second acid dissociation constants of the metal chelate species.

The plot in Figure 3 was chosen as an example for checking the inactivity of the hydroxo species of the Fe(III)-HIMDA chelate since it hydrolyzes to a greater extent (pK_1 = 2.46, pK_2 = 5.70) than the other metal chelate compounds studied in this investigation. By the use of this procedure hydroxo species of other metal compounds were also found to be inactive. It should be pointed out, however, that catalysis of the reaction of substrate HA^- by the unhydrolyzed metal chelate ML^{-n} would be kinetically equivalent to catalysis of the undissociated substrate, H_2A, by the hydrolyzed metal chelate $M(OH)L^{-n-1}$.

Oxygen Dependence of the Rate of Metal Chelate Catalyzed Oxidation. The dependence of the rates of the reaction on the partial pressure of oxygen was determined both for ferric and cupric chelate catalyzed oxidations at pH values of 3.00, 3.45, and 3.85. For these studies mixtures of oxygen and nitrogen varying in composition from 80% oxygen to 5% oxygen were employed. The gas mixture was rapidly passed through the experimental solution at a very rapid rate (1 l. of the gas/80 sec). Since the rate of dissolution of oxygen is much faster than any reaction, the amount of oxygen present at a given time is constant and only depends on its partial pressure. The rate of oxidation was followed by measuring the amount of dehydroascorbic acid formed as a function of time. For all the metal chelate compounds studied, the rate of oxidation was found to be independent of the partial pressure of

oxygen. The results for Fe–EDTA and Cu–EDTA at pH 3.00 are presented in Table VI. Similar results were obtained for other metal chelates of Cu^{2+} and Fe^{3+} reported in this paper.

Table VI. Rate Constants (M^{-1} sec^{-1}) for Cu(II)–EDTA and Fe(III)–EDTA Catalyzed Oxidation of Ascorbic Acid as a Function of Oxygen Concentration[a]

Partial pressure of O_2, atm	$k_{Cu-EDTA}$	$k_{Fe-EDTA}$
1.00	2.9×10^1	5.4×10^1
0.81	2.9×10^1	5.3×10^1
0.62	2.8×10^1	5.4×10^1
0.40	2.8×10^1	5.3×10^1
0.19	2.7×10^1	5.2×10^1

[a] Temperature 25°; μ = 0.10 M (KNO_3).

Activation Parameters. The activation parameters of the cupric and ferric chelate catalyzed oxidation of ascorbic acid, calculated from the temperature dependence of rate constants in Tables III and V, are given in Tables VII and VIII, respectively.

Table VII. Activation Parameters of the Cu(II) Chelate Catalyzed Oxidation of Ascorbic Acid[a]

Chelate (1:1)	$\Delta H^{\ddagger}$, kcal/mole	$\Delta S^{\ddagger}$, cal deg^{-1} mole^{-1}	$\Delta F^{\ddagger}$, kcal/mole
Cu(II)–EDTA	+2.0	−42	+14.6
Cu(II)–HEDTA	+2.8	−37	+14.0
Cu(II)–NTA	+3.0	−35	+13.5
Cu(II)–HIMDA	+3.2	−33	+13.0
Cu(II)–IMDA	+3.2	−31	+12.5
None[b]	+10.7	−37	+22.0

[a] μ = 0.10 M (KNO_3). [b] Spontaneous oxidation.

Table VIII. Activation Parameters of Fe(III) Chelate Catalyzed Oxidation of Ascorbic Acid[a]

Chelate (1:1)	$\Delta H^{\ddagger}$, kcal/mole	$\Delta S^{\ddagger}$, cal deg^{-1} mole^{-1}	$\Delta F^{\ddagger}$, kcal/mole
Fe(III)–DTPA	+2.0	−42	+14.6
Fe(III)–CDTA	+2.0	−37	+13.1
Fe(III)–EDTA	+2.2	−36	+13.0
Fe(III)–HEDTA	+2.4	−33	+12.3
Fe(III)–NTA	+2.6	−32	+12.2
Fe(III)–HIMDA	+3.3	−29	+12.0
None[b]	+10.7	−37	+22.0

[a] μ = 0.10 M (KNO_3). [b] Spontaneous oxidation.

Discussion

In the cupric and ferric chelate catalyzed oxidation of ascorbic acid an inverse dependence of specific rate on hydrogen ion concentration was observed. The rate law may be expressed in the form

$$-(dT_A/dt) = kT_A[ML] \quad (7)$$

where T_A = total concentration of unreacted ascorbic acid and [ML] = total concentration of metal chelate catalyst. Equation 7 may be expressed in the form

$$-(dT_A/dt) = k_1[H_2A][ML] + k_2[HA^-][ML] \quad (8)$$

where k_1 and k_1 are constants for the catalytic effect of the metal chelate compound on the neutral and monoionic forms of ascorbic acid, respectively, on the assumption that the two species of ascorbic acid react independently with the metal chelate catalyst. Equation 8 may be rewritten

$$-(dT_A/dt) = k_1[H_2A][ML] + k_2[H_2A]K_1[ML]/[H^+] \quad (9)$$

where

$$[HA^-] = [H_2A]K_1/[H^+] \quad (10)$$

(K_1 is the first dissociation constant of ascorbic acid).

$$-(dT_A/dt) = [H_2A][ML](k_1 + k_2K_1/[H^+]) \quad (11)$$

With the mass balance expression between T_A and $[H_2A]$, eq 7 and 11 give

$$k = \left(k_1 + k_2\frac{K_1}{[H^+]}\right)\left(\frac{[H^+]}{[H^+] + K_1}\right) \quad (12)$$

In acid solution where H_2A is not appreciably ionized, H_2A may be taken as approximately equal to T_A.

$$k = k_1 + k_2K_1/[H^+] \quad (13)$$

A plot of the specific rate constant k against the reciprocal of hydrogen ion concentration gives a straight line with slope equal to k_2K_1 and intercept k_1. In the metal chelate catalyzed oxidation of ascorbic acid k_1 was found to be zero. The values of k_2 listed in Tables III and V were obtained from the slopes of the straight lines in Figures 4 and 5, respectively, and from similar plots at 0.4°. The results indicate that the metal chelate catalyst is totally inactive for the oxidation of the un-ionized species. The observed specific rates apply entirely to the oxidation of the ascorbate monoanion.

Two mechanisms may be considered for the metal chelate catalyzed oxidation of ascorbic acid: (1) a dissociation mechanism, in which the catalytic effect may be considered to be due to the free metal ion in equilibrium with the metal chelate compound; (2) direct participation of the metal chelate species in an electron-transfer process.

As a test of the first mechanism, the concentration of the dissociated metal ion was calculated for the ferric chelates at $-\log [H^+] = 3.00$. From the concentration of the dissociated metal ion and the observed rate for the metal chelate species, a specific rate constant was calculated for each of the ferric chelates studied in this investigation. It may be seen from the results listed in Table IX that the specific rates vary widely in the series of ferric chelates. If a dissociation mechanism was operative, the specific rates calculated in Table IX should be equal to, or a simple function of, the specific rates obtained for the ferric ion catalyzed oxidation of ascorbic acid at $-\log [H^+] = 3.00$. The wide variation of the specific rates in Table IX thus seems to be incompatible with the dissociation mechanism.

The second alternative is the direct participation of the metal chelate species in the oxidation of ascorbic acid. On the basis of the zero dependence of the rate of oxidation on the oxygen concentration, the following mechanism may be considered for this reaction.

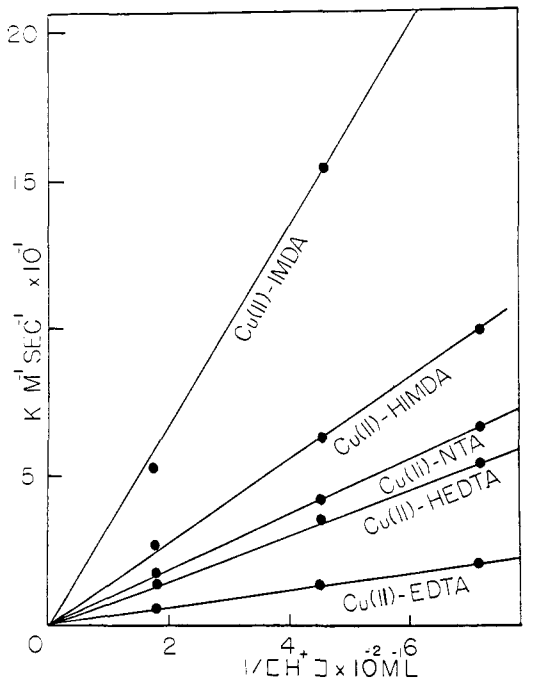


Figure 4. Dependence of the specific rate constants k on the hydrogen ion concentration for Cu(II) chelate catalyzed oxidation of ascorbic acid at 25°; μ = 0.10 M (KNO_3).

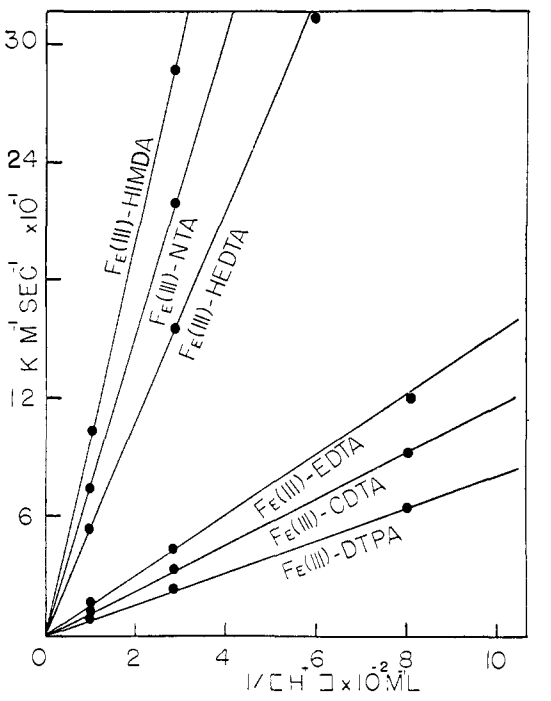


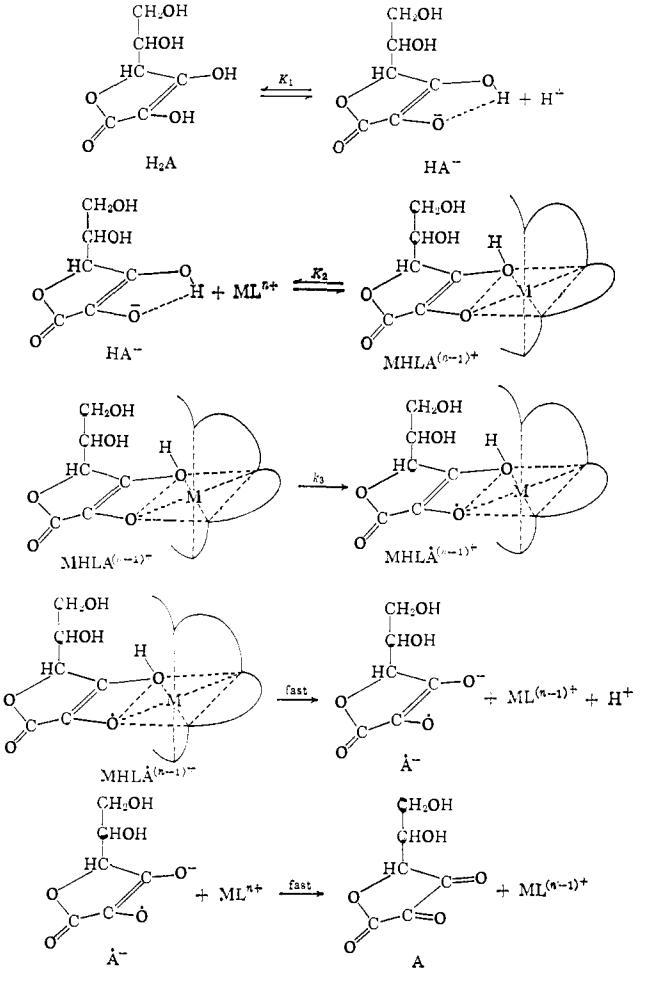
Figure 5. Dependence of the specific rate constants k on the hydrogen ion concentration for Fe(III) chelate catalyzed oxidation of ascorbic acid at 25°; μ = 0.10 M (KNO_3).

In the proposed mechanism, a mixed ligand chelate of ascorbate anion, HA^-, and the metal chelate, ML^{n+}, is formed in a preequilibrium step. This is followed by a rate-determining electron transfer within the mixed ligand chelate, $HAML^{(n-1)+}$, from ascorbate to metal ion. The reduced compound, $HAML^{(n-1)+}$, then dissociates in a fast step to the lower valence metal chelate, $ML^{(n-1)+}$, and the semiquinone. Oxidation of the lower valence metal chelate and semiquinone by molecular oxygen takes place in subsequent fast steps. The fact that the rate of oxidation of ascorbic acid in the presence of Cu(II) and Fe(III) chelates as catalysts is independent of the concentration of oxygen is in agreement with this mechanism. Since the rate-deter-

Table IX. Rate Constants for the Oxidation of Ascorbic Acid as Catalyzed by Free Metal Ion[a]

Chelate (1:1)	T_{MD}, $M \times 10^4$	$[M^{n+}]$, $M \times 10^9$	k_{obsd}, $sec^{-1} \times 10^5$	k_M, M^{-1} sec^{-1}
Fe(III)–DTPA	1.00	1.20	0.53	4.5×10^3
	1.50	1.50	1.0	7.0×10^3
	2.50	1.90	1.8	1.0×10^4
Fe(III)–CDTA	1.00	0.60	0.90	1.6×10^4
	1.50	0.70	1.4	2.2×10^4
	2.50	0.90	2.7	3.0×10^4
Fe(III)–EDTA	2.00	0.50	1.3	2.7×10^4
	2.50	0.70	2.5	3.6×10^4
	3.00	0.80	3.6	4.5×10^4
Fe(III)–HEDTA	0.60	0.50	2.5	5.0×10^4
	1.00	0.60	4.7	7.2×10^4
	1.50	0.80	7.3	1.0×10^5
Fe(III)–NTA	0.50	19.6	3.2	1.6×10^3
	1.00	27.6	6.9	2.5×10^3
	1.20	30.2	8.8	2.9×10^3

[a] μ = 0.10 M (KNO_3); temperature 25°; $-\log [H^+]$ = 3.00. T_{ML} = total concentration of the metal chelate species; $[M^{n+}]$ = concentration of dissociated metal ion; k_{obsd} = rate constant observed for the concentration; T_{ML} of the chelate species; k_M = rate constant attributed to free metal ion.



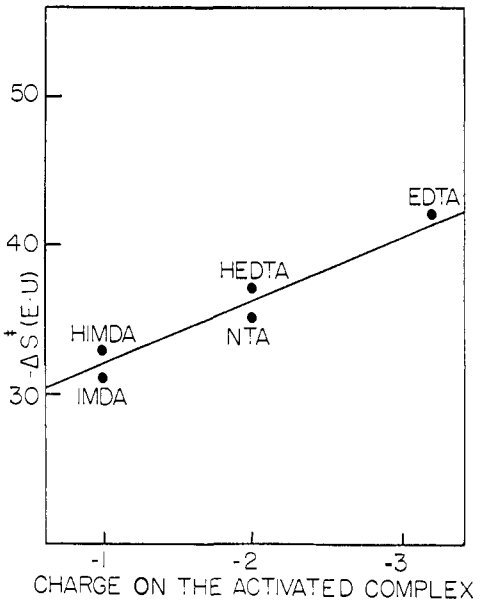


Figure 6. Variation of entropy of activation with charge on the activated complex for Cu(II) chelate catalyzed oxidation of ascorbic acid at 25°; μ = 0.10 M (KNO_3).

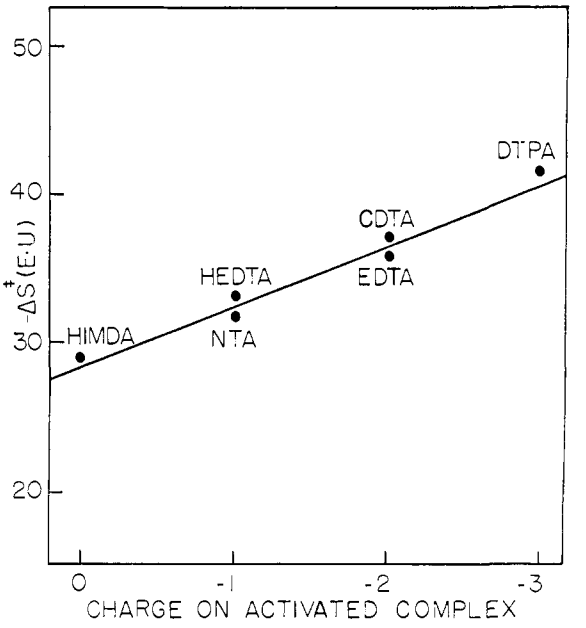


Figure 7. Variation of entropy of activation with charge on the activated complex for Fe(III) chelate catalyzed oxidation of ascorbic acid at 25°; μ = 0.10 M (KNO_3).

mining step in the proposed mechanism involves direct electron transfer to the metal chelate in an activated mixed ligand chelate complex, it should be essentially the same as the oxidation of ascorbic acid by metal chelates in the absence of oxygen. This statement implies that, at least in the initial stages of the reaction, the rates in the presence and in the absence of oxygen should be comparable. This may be verified by comparing the free energies of activation of the oxidation of ascorbic acid by ferric chelates as oxidants and as catalysts (Table X). Further, no hydrogen peroxide

Table X. Comparison of the Free Energies of Activation of the Oxidation of Ascorbic Acid by Fe(III) Chelates as Catalysts and as Oxidants[a]

Chelate (1:1)	$\Delta F^{\ddagger}$, cal[b]	$\Delta F^{\ddagger}$, cal[c]
Fe(III)–DTPA	14.6	13.7
Fe(III)–CDTA	13.1	13.3
Fe(III)–EDTA	13.0	12.6
Fe(III)–HEDTA	12.3	11.3

[a] μ = 0.1 M (KNO_3). [b] Catalyzed oxidation by molecular O_2. [c] Ferric chelates as oxidants, no O_2.

could be detected in both the cupric and ferric chelate catalyzed oxidations in contrast to the metal ion catalyzed reaction. This implies that H_2O_2 produced as an intermediate of the reduction of molecular oxygen reacts further with the lower valence metal chelate, and is reduced to H_2O in subsequent fast steps as suggested by Grinstead.[6]

The feasibility of a mixed ligand chelate HAML in the proposed mechanism is supported by the fact that even in highly stable complexes such as those of EDTA, all the coordinated positions on the metal ion are probably not bound. This is supported by the X-ray work of Hoard, *et al.*,[13] in which the authors have shown coordination number seven for Fe(III) in the crystalline EDTA chelate. In ionic complexes, such as those of Cu(II) and Fe(III), the bond angles are not rigidly fixed. For these labile complexes, the more ionic the complex, the greater is the deviation of the bond angles from the normal octahedral value. Staveley and Randell[14] have discussed the fact that in metal–EDTA complexes, the field acting on the metal ion can only be a crude approach to an octahedral field. Even if four coordinating atoms (two nitrogen and two oxygen atoms) lie roughly in a plane, the other two cannot fall on the fourfold axis through this plane. Thus there may always be one or two positions available on the metal ion to form a mixed ligand chelate of the type HAML with ascorbate anion.

The activation parameters of cupric and ferric chelate catalyzed oxidation are given in Tables VII and VIII, respectively. In a particular series of metal chelate compounds there is not much difference in the values of $\Delta H^{\ddagger}$, but there is a very interesting trend in the values of $\Delta S^{\ddagger}$. The values of $\Delta S^{\ddagger}$ become more negative as the negative charge on the metal chelate catalyst increases. The variation of $\Delta S^{\ddagger}$ with charge on the activated complex for cupric and ferric chelates are shown in Figures 6 and 7, respectively.

The variation of entropies of activation with charge on the activated complex is in accord with the expected variation in the probabilities of forming the activated complex. As the negative charge on the metal chelate ML increases, its affinity for ascorbate anion (to form a mixed ligand chelate, HAML) decreases. This effect would contribute to a more negative entropy of activation.

The enthalpies of activation for the metal chelate catalyzed oxidation are lower than those of the corresponding metal ion catalyzed oxidations.[4] In a particular series of metal chelates the enthalpies of activation are gradually lowered in passing from a weakly

(13) J. L. Hoard, M. Lind, and J. V. Silverton, *J. Am. Chem. Soc.*, 83, 2770 (1961).

(14) L. A. K. Staveley and T. Randell, *Discussions Faraday Soc.*, 26, 157 (1958).

binding to a strongly binding ligand. The difference $\Delta H^{\pm}$ for the metal ion and metal chelate catalyzed oxidation and the variation of $\Delta H^{\pm}$ in a particular series of metal chelate catalysts is in accord with the Franck–Condon principle.[15] As the binding between the metal ion and the chelating agent becomes stronger, the, tendency to form a mixed ligand chelate with ascorbate anion decreases. This results in less rigid bonding of the substrate in the mixed ligand chelate, which is easily deformed on electron transfer. The ease of deformation leaves the mixed ligand chelate species more or less geometrically similar before and after the electron transfer. This geometrical similarity of the oxidized and reduced ions requires very little change in hydration energy during the electron transfer, which is reflected in a lowering of $\Delta H^{\pm}$. A more positive value of $\Delta F^{\pm}$ in these cases is, however, due to a more negative entropy term. In the case of the mixed ligand chelates of weak ligands (more rigid bonding of the substrate) the unfavorable enthalpies are more than compensated for by more positive entropies, which make $\Delta F^{\pm}$ more negative.

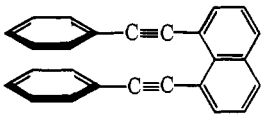
(15) W. F. Libby, *J. Chem. Phys.*, **38**, 420 (1963).

Communications to the Editor

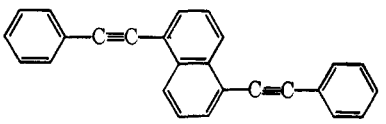
Interactions and Reactions of 1,8-Bis(phenylethynyl)naphthalene

Sir:

We wish to describe 1,8-bis(phenylethynyl)naphthalene (I), a *peri* derivative in which the acetylene groups are essentially parallel.[1]



I

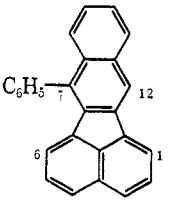


II

Compound I (mp 106°, light yellow) and 1,5-bis(phenylethynyl)naphthalene (II, mp 199°, white) were prepared[2] by treating (phenylethynyl)copper with 1,8-diiodonaphthalene and 1,5-diiodonaphthalene. I is slightly more colored than II. I exhibits acetylenic absorption at 4.48 μ and ultraviolet maxima (ethanol) at 209 (ϵ 47,800), 243 (66,300), 265 (25,800), 344 (29,300), and 365 mμ (26,800); 1-(phenylethynyl)naphthalene[3] has maxima at 256, 268, 317, and 338 mμ. The ultraviolet spectra of I and II are essentially superimposable except that the maxima of I occur at slightly longer wavelengths (1–2 mμ) and the intensity of its phenylethynyl band (265 mμ (ϵ 26,800)) is considerably less than that of II (263 mμ (ϵ 44,000)).

Compound I exhibits nuclear magnetic resonance at τ 2.52–3.08 (multiplet, 12 H) and 2.11–2.40 (multiplet, 4 H); II shows proton resonance at τ 3.15–2.83 (multiplet, 14 H) and 1.39–1.64 (doublet, 2 H). Phenyl proton signals in I (τ 2.87) are centered upfield from those in II (2.57), phenylacetylene (2.68), diphenylbutadiyne (2.67), 1,2-bis(phenylethynyl)benzene[4a] (2.71), 1-phenylnaphthalene[4b] (2.62), and 1,7-diphenylnaphthalene[4b] (2.51); shielding is apparently less than in 1,8-diphenylnaphthalene[4b] (τ 3.15) and [2.2]paracyclophane[4c] (3.63). From the moderate shielding and the absorption, the phenyl groups in I appear to be moved apart, and there may be some interaction between the acetylene groups. What is clear is that there is no marked transannular conjugation in electronically excited I.[4b] The results are also consistent with predictions that cyclobutadienes are not highly delocalized.[5]

Compound I yields 7-phenylbenzo[*k*]fluoranthene (III, mp 167°, pale yellow) when warmed to 100°, photolyzed in pentane, or exposed to aluminum chloride. Compound III does not exhibit infrared absorption for



III

an acetylene group, resists catalytic hydrogenation, and is not oxidized by potassium permanganate. Its ultraviolet spectrum ($\lambda_{max}^{ethanol}$ 216 (ϵ 42,300), 246 (42,900), 270 (17,700), 287 (18,000), 298 (31,200), 309 (38,400), 364 (5250), and 384 mμ (8200)) is similar throughout to that of benzo[*k*]fluoranthene[6a] and 7,12-diphenylbenzo[*k*]fluoranthene.[6b] Its nmr spectrum shows aromatic proton signals at τ 3.47 (doublet, 1 H), 2.49 (multiplet, 14 H), and 1.83 (singlet, 1 H). The naphthyl proton at C-6 is shielded by the twisted phenyl group, giving the doublet at τ 3.47; the proton singlet is that at C-12.

(1) The distance between 1,8-hydrogens in naphthalene is ~2.45 Å; the transannular distances between similar carbon atoms in [2.2]paracyclophane range from 2.83 to 3.09 Å: C. J. Brown, *J. Chem. Soc.*, 3265 (1953).

(2) All new compounds have satisfactory analyses and molecular weights.

(3) S. A. Kandil and R. E. Dessy, *J. Am. Chem. Soc.*, 88, 3027 (1966).

(4) (a) Private communication, H. W. Whitlock, Jr.; (b) H. O. House, R. Magin, and H. W. Thompson, *J. Org. Chem.*, 28, 2403 (1963); (c) D. C. Cram, C. Dalton, and G. R. Knox, *J. Am. Chem. Soc.*, 85, 1088 (1963);[4d] (d) *p*-xylene has aryl proton resonance at τ 2.95.

(5) M. J. S. Dewar and G. J. Gleicher, *J. Am. Chem. Soc.*, 87, 3255 (1965).

(6) (a) M. Orchin and L. Reggel, *ibid.*, 73, 436 (1951); (b) E. Bergmann, *ibid.*, 74, 1075 (1952); (c) structures which are excluded will be discussed in a final manuscript.