Kinetics and Mechanism of the Metal Chelate Catalyzed Oxidation of Pyrocatechols[1]

Charles A. Tyson and Arthur E. Martell*

Contribution from the Department of Chemistry, Texas A & M University, College Station, Texas 77843. Received March 4, 1971

Abstract: Studies of the metal chelate catalyzed oxidations of pyrocatechol (PC) to *o*-benzoquinone and 3,5-di-*tert*-butylpyrocatechol (3,5-DTBP) to the corresponding *o*-quinone (3,5-DTBQ) with manganese(II) 4-nitrocatechol (4-NC), manganese(II) tetrabromocatechol (TBC) and cobalt(II)-4-NC as catalysts are reported and are compared with the free manganese(II)- and base-catalyzed autoxidations. On the basis of equilibrium studies and of dependence of reaction rates on $[H^+]$, the catalytic species are identified as the free metal ion or monohydroxy complex and the 1:1 metal chelate. Rate enhancements of 10^2 and 10^3 over the base-catalyzed autoxidations were observed up to pH 8.0 in unbuffered solutions under an air atmosphere. The manganese(II)-4-NC and manganese-TBC catalyzed reactions conformed to the equations: 3,5-DTBP + $O_2 \rightarrow$ 3,5-DTBQ + H_2O_2; with cobalt(II)-4-NC, 3,5-DTBP + $0.5O_2 \rightarrow$ 3,5-DTBQ + H_2O. The variation of reaction rate with oxygen pressure and $[H^+]$ permits elucidation of the initial steps in the manganese(II) chelate catalyzed mechanism and a probable transition state is suggested.

Few metal-catalyzed oxidations have been studied in detail because of their complexity.[2-4] For this reason the mechanism of the metal-catalyzed oxidation of pyrocatechol (PC) to form *o*-benzoquinone is not known. Gillette, *et al.*,[5] reported an eightfold increase in the rate of oxygen uptake with 2:1 molar ratios of PC to Mn(II) or Co(II), in Veronal buffer at pH 7.9; however, 2 moles of oxygen per mole of PC were consumed. Isaka and Akino[6] found that $Cu(CN)_2$ gives a two- to threefold increase in pH 7.8 phosphate buffer. Jameson and Martell[7] found that the 3,4-dihydroxyphenylalanine (DOPA) autoxidation rate was proportional to [DOPA], [Fe(III)], $[O_2]$, and $[H^+]^{2-}$. Grinstead[8] studied the Mn(II)-catalyzed oxidation of pyrocatechols and suggested a mechanism for the reaction.

By choosing a suitable model pyrocatechol and appropriate experimental conditions it should be possible to elucidate the oxidation mechanism from kinetic studies. With free Mn(II) and Co(II) ions as catalysts 90% or higher yields of 3,5-di-*tert*-butyl-*o*-benzoquinone (3,5-DTBQ) were obtained from 3,5-di-*tert*-butylpyrocatechol (3,5-DTBP) in $KHCO_3$ buffer.[8] The metal-free autoxidation of 3,5-DTBP has been studied[9] and has been found to be a suitable model for pyrocatechol itself. In such studies buffer effects on catalysis can be eliminated by incorporating an automatic pH control unit with a bubble-type manometric apparatus.[9] Chelating agents can be used as auxiliary ligands to maintain metal ion solubility during catalysis. 4-Nitrocatechol (4-NC) and tetrabromocatechol (TBC) are bidentate, leaving a maximum number of coordination sites open, form moderately stable complexes with Mn(II) and Co(II), and are relatively resistant toward autoxidation in mildly alkaline solutions.[10-12] A kinetic investigation of the Mn(II)- and Co(II)-catalyzed oxidation of 3,5-DTBP with 4-NC and TBC as auxiliary ligands was therefore undertaken. It was thought that such a study could be significant in providing new information on the activation of oxygen by metal ions, a subject of interest to coordination chemists,[3,13] synthetic organic chemists,[14] enzymologists,[2,4] and water pollution scientists.[15,16]

Experimental Section

Materials. 3,5-Di-*tert*-butylpyrocatechol and 4-nitrocatechol were obtained from the Aldrich Chemical Co. and used after recrystallization from heptane and water. High purity samples of tetrabromocatechol from Aldrich Chemical Co. and of pyrocatechol from Eastman Organic Chemicals were used without recrystallization. Metal salts and KNO_3 were Fisher Certified Reagent grades, the latter recrystallized once from water. Analyzed oxygen–nitrogen gas mixtures were purchased from the Matheson Chemical Co. Water was deionized and redistilled from glass apparatus. Spectral grade methanol from Fisher Scientific Co. was employed.

Methods. The bubble-type manometric apparatus previously described was used in kinetic experiments.[12] Catalytic quantities of metal ion and auxiliary ligand were added to the reaction solution before equilibration with the atmosphere. Base was added manually through a needle and syringe attachment to the cell cover to initiate the reaction. Thereafter the alkali requirement dropped off rapidly to negligible amounts during the first 15–30 seconds. pH control was held to within ±0.1 unit.

Precipitates appeared within seconds in kinetic runs with 3,5-DTBP in 50% methanol when either Co(II) was the catalyst or when TBC was used in place of 4-NC as the auxiliary ligand, so that in these systems higher methanol percentages were required to maintain solubility. Precipitation of Co(II)-3,5-DTBP complexes had earlier been observed to occur in 50% methanol.

Equilibria were measured in a standard 100-ml potentiometric cell with a Beckman Research pH meter using techniques prescribed

---

[1] This work was supported by a research grant (No. 12050 DIT) from the Federal Water Pollution Control Agency, Department of the Interior.

[2] B. R. Brown, in "Oxidative Coupling of Phenols," W. I. Taylor and A. R. Battersby, Ed., Marcel Dekker, Inc., New York, N. Y., 1967, p 197.

[3] A. E. Martell, *Pure Appl. Chem.*, **17**, 129 (1968).

[4] G. A. Hamilton, *Advan. Enzymol.*, **59**, 55 (1969).

[5] J. R. Gillette, D. Watland, and G. Kalnitsky, *Biochim. Biophys. Acta*, **15**, 526 (1954).

[6] S. Isaka and M. Kaino, *Nature (London)*, **177**, 184 (1956).

[7] See comment: J. E. Gorton and R. F. Jameson, *J. Chem. Soc. A*, 2615 (1968).

[8] R. R. Grinstead, *Biochemistry*, **3**, 1308 (1964).

[9] C. A. Tyson and A. E. Martell, *J. Phys. Chem.*, **74**, 2601 (1970).

[10] C. A. Tyson, *Diss. Abstr. B*, **29**, 3237 (1969).

[11] Y. Murakami and M. Tokunaga, *Bull. Chem. Soc. Jap.*, **37**, 1562 (1964).

[12] C. A. Tyson and A. E. Martell, *J. Amer. Chem. Soc.*, **90**, 3379 (1968).

[13] J. Halpern, *Discuss. Faraday Soc.*, 7 (1968); *Pure Appl. Chem.*, **20**, 59 (1969).

[14] J. Kochi, *Science*, **155**, 415 (1967).

[15] G. M. Fair, J. C. Geyer, and D. A. Okun, "Waste and Wastewater Engineering," Vol. II, Wiley, New York, N. Y., 1968, pp 34-2, 31-28.

[16] R. W. Holcomb, *Science*, **169**, 457 (1970).

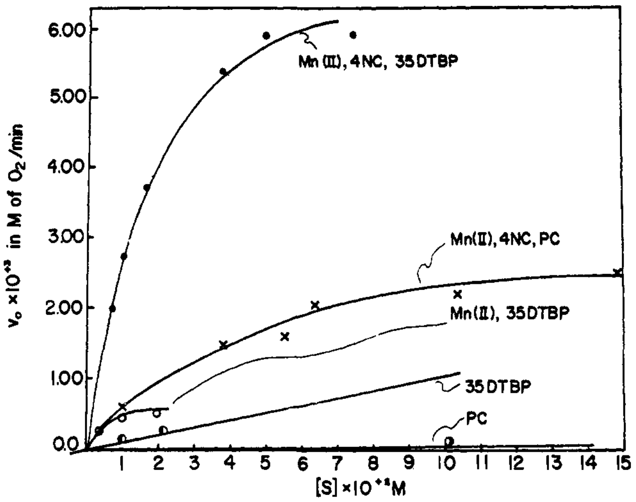

Figure 1. Initial oxygen uptake rate, v_0, in moles/min, *vs.* substrate concentration [S]. Conditions: [Mn(II)], [4NC] = 1.15×10^{-3} *M*, 50% CH_3OH–0.10 *M* KNO_3, $-\log [H^+] = 7.28$, 25°. System: under air, (—·—·—) Mn(II)-4NC-3,5-DTBP; (—○—○—) Mn(II)-3,5-DTBP; (—×—×—) Mn(II)-4NC-PC; under 100% O_2, (—◐—◐—) 3,5-DTBP; (—◑—◑—) PC. v_0 corrected to 760 mm, 25°.

for mixed solvents.[12] The cell was water jacketed for temperature control at 25.0 ± 0.01°.

Rapid flow experiments were carried out with the help of Dr. Helen Brooks on a Varian 4502-12 spectrometer system equipped with a Varian 4548 quartz aqueous flow cell and Varian 4549 liquid flow mixing chamber.[17] Solutions were stored in two 10-l. capacity glass cylinders and forced continuously through the mixing chamber under 20 lb/in.^{2} of N_2 pressure. Fastest resolution time was 7–8 msec after mixing.

Calculations. First-order rate constants for pyrocatechol consumption were calculated by dividing the change in buret reading/min by the total change in fluid level at the end of the reaction. The change in buret reading/min was the initial change, obtained by extrapolation of the experimental points after the 15- or 30-sec mark back to time zero. Since PC solutions consume more than 1 mole of O_2 per mole of substrate, the fluid level displacement corresponding to uptake of 1.0×10^{-2} mole of O_2 was used in the calculation of PC oxidation rate constants.[3,18]

The first-order rate constants are a function of the dissolved oxygen concentration [O_2] since its concentration can always be made sufficiently low for it to influence the rate-determining step. [O_2] is assumed to be directly proportional to the partial pressure of oxygen (P_{O_2}) above the solution (Henry's Law). In order to make rate constants involving O_2 concentration comparable they were corrected for standard conditions by multiplying by the factor $P_{760}/(P_{atm} - P_{solv})$ where P_{atm} is the barometric reading and p_{solv} is derived from $0.485 \times p_{H_2O} + 0.60 \times p_{me}$ for 50% methanol and $0.38 \times p_{H_2O} + 0.72 \times p_{me}$ for 68% methanol at 20°.[9,19] These corrections are less than 5% in aqueous solvent but amount to 10 and 16% on the average in 50 and 68% methanol, respectively. Since solubility data are not available for the solvent media, rate constants were not adjusted to equal [O_2], but it is noted that [O_2] does not vary significantly in mixed solvents up to 50% alcohol.[20]

Equilibrium constants for 1:1 and 1:2 complexes of Mn(II) with 4-NC and with TBC were determined[12] from potentiometric data for the following reactions

$$M^2 + H_2L \rightleftharpoons ML + 2H^+ \qquad {}^*K_1 = [ML][H^+]^2/[M^{2+}][H_2L] \quad (1)$$

$$ML + H_2L \rightleftharpoons ML_2 + 2H^+$$
$$ {}^*K_2 = [ML_2{}^{2-}][H^+]^2/[ML][H_2L] \quad (2)$$

and

$$ {}^*K_n = \frac{(T_L - [H_2L]x - (n-1)T_M)[H^+]^2}{(nT_M - T_L + [H_2L]x)[H_2L]} \quad (3)$$

where $[H_2L] = (2T_L - T_{OH} - [H^+] + [OH^-])\{2 + (K_2{}^H[H^+])^{-1}\}$, T_L and T_M are total analytical concentrations of metal ion and ligand, respectively, T_{OH} = molar concentration of added NaOH and $x = 1 + 1/K_2{}^H[H^+]$. Only the protonation constant $K_2{}^H$ is required, $K_1{}^H$ being outside the region of constant ionic strength and thus subject to large error.

The value of K_w of 6.05×10^{-15} M^2 was taken for 68% methanol from the data of Kaskikallio.[21] Activity coefficient corrections for ionic strength and electrolyte were made from the data of Harned and Owen[22] under the assumption that the relationship for determining activity coefficients for NaCl in mixed electrolytic systems in alcohol-water solvents also holds for KNO_3.[23] Ion pair formation is not detectable in the media used.[24]

Results

Evidence of Catalysis. A 1:1 molar ratio of either Mn(II)-4-NC, Co(II)-4-NC, or Mn(II)-TBC, designated as ML, when added to a solution of 3,5-DTBP or PC, led to an increase in the oxygen uptake rate relative to the substrate alone or to substrate and free metal ion at the same analytical concentration. Plots of initial oxygen uptake rates v_0 *vs.* [S] ([S] is total analytical concentration of 3,5-DTBP or PC) at a given ML concentration were hyperbolic, typical of a catalysis in which an MLS complex is an intermediate. Figure 1 gives an example of this for Mn(II)-4-NC-3,5-DTBP and Mn(II)-4-NC-PC in 50% methanol under an air atmosphere. The uptake rate with Mn(II) alone peaks out at low [3,5-DTBP]/[Mn] and obviously makes only a small contribution to the overall rate with auxiliary ligand present. Under air the turnover of moles of substrate/mole of Mn(II) and 4-NC per min was 400 and 600 times greater, respectively, than with 3,5-DTBP and PC alone. (The uptake rates for substrate alone in the figure are divided by 5 to obtain the rates under air, since autoxidation of each substrate exhibits first-order dependence on [O_2].)[8,25] 4-NC alone exerted no accelerating effect on the initial 3,5-DTBP or PC autoxidation rate.[26,27] Consequently one or more complexes of Mn(II) and the auxiliary ligand must be responsible for catalysis.

It was necessary to establish whether or not the auxiliary ligand is degraded during the catalytic cycle. In the absence of substrate, O_2 uptake with either Mn(II)-4-NC or Mn(II)-TBC was negligible over a 14-min period. No significant change was observed in the 4-NC visible peak after oxidation runs with either Co(II) or Mn(II) and 3,5-DTBP or PC as substrates. Rapid-flow esr showed no alteration of the Mn(II) derivative peaks during 3,5-DTBP oxidation ([3,5-DTBP]/[Mn(II)-4-NC] = 5/1 in 10 m*M* Tris-HCl solution,

---

(17) E. L. Lewis and F. Sicilio, *J. Phys. Chem.*, **73**, 2590 (1969).

(18) Oxygen consumption also depends on the absolute pyrocatechol concentration [molar ratio = 2.5 for PC less than 1×10^{-3} *M*, decreasing to 2.0 for PC = 5×10^{-3} *M*; see W. G. C. Forsyth and V. C. Quesnel, *Biochem. Biophys. Acta*, **25**, 155 (1957)] and on the medium [molar ratio = 0.5 in water-pyridine-acetic acid (27:70:3) with Cu(II) and Fe(III)-1,2-naphthoquinone chelate as catalyst; K. I. Matveev, W. Langenbeck, A. M. Osipov, G. V. Krause, and H. J. Kreuzfeld, *Kinet. Katal.*, **6**, 651 (1965)].

(19) "International Critical Tables," Vol. III, McGraw-Hill Book Co., Inc., New York, N. Y., 1928, p 290.

(20) P. Schlapfer, T. Audykowski, and A. Bukowiecki, *Schweiz. Arch. Angew. Wiss. Tech.*, **15**, 299 (1949); H. Nogami and Y. Kato, *Yakugaku Zasshi*, **79**, 958 (1959).

(21) J. Koskikallio, *Suom. Kemistilehti B*, **30**, 111 (1957).

(22) H. S. Harned and B. B. Owen, "The Physical Chemistry of Electrolytic Solutions," 2nd ed, Reinhold, New York, N. Y., 1958, pp 732, 735.

(23) H. S. Harned and R. A. Robinson, "Multicomponent Electrolyte Solutions," Pergamon Press, London, 1968, p 97.

(24) M. Paabo, R. G. Bates, and R. A. Robinson, *Anal. Chem.*, **37**, 462 (1965); R. G. Bates, M. Paabo, and R. A. Robinson, *J. Phys. Chem.*, **67**, 1833 (1963).

(25) M. A. Joslyn and G. E. K. Branch, *J. Amer. Chem. Soc.*, **57**, 1779 (1935).

(26) G. E. K. Branch and M. A. Joslyn, *ibid.*, **57**, 2388 (1935).

(27) G. Cilento and K. Zinner, *Biochim. Biophys. Acta*, **120**, 84 (1966); **143**, 88 (1967).

Table I. Stoichiometry of Mn(II)-Chelate Catalyzed Oxidation of 3,5-DTBP^a

Catalytic system	Medium	% [O_2] consumed	% [3,5-DTBQ]^b produced	% [H_2O_2]^c produced
Mn(II)-4-NC^d	50% CH_3OH–0.10 *M* KNO_3	100	95–100	85–90
Mn(II)-4-NC^d	68% CH_3OH–0.050 *M* KNO_3	100	90–95	80
Mn(II)-TBC^d	68% CH_3OH–0.050 *M* KNO_3	100	90–95	80
Co(II)-4-NC^e	68% CH_3OH–0.050 *M* KNO_3	50	80	0

^{a} [3,5-DTBP] = 1.0×10^{-2} *M*, pH reading = 7.75, 25°. ^{b} [3,5-DTBQ] determined spectrophotometrically at 404 mμ.^{9} ^{c} [H_2O_2] determined titrimetrically with Ce(IV).^{9} ^{d} Conforms to equation: 3,5-DTBP + $O_2 \rightarrow$ 3,5-DTBQ + H_2O_2 over the pH range 7.5–8.5. ^{e} Conforms to equation: 3,5-DTBP + $0.5O_2 \rightarrow$ 3,5-DTBQ + H_2O.

pH 8.2). At this pH, in the absence of 4-NC, insoluble hydroxo and perhydroxo complexes form and the Mn(II) signal disappears within minutes.[10,28] These experiments demonstrated that the metal–ligand complex was virtually unaltered after catalysis.

Stoichiometry. The products and yields from the Mn(II)-catalyzed oxidation of 3,5-DTBP are presented in Table I for the Mn(II)-4-NC, Mn(II)-TBC, and Co(II)-4-NC catalyzed systems at the same pH reading in each solvent system. With Mn(II) present the oxidation conformed to the equation

$$\text{3,5-DTBP} + O_2 \longrightarrow \text{3,5-DTBQ} + H_2O_2 \tag{4}$$

With Co(II)-4-NC the overall stoichiometry was

$$\text{3,5-DTBP} + 0.5O_2 \longrightarrow \text{3,5-DTBQ} + H_2O \tag{5}$$

The difference in the yield of H_2O_2 was not unexpected, since Co(II) complexes are known to catalyze its decomposition[29] whereas Mn(II) does so only with precipitates present.[30] The immediate product of PC oxidation was assumed to be *o*-quinone, and H_2O_2 is known to be generated but the stoichiometry was not investigated here.[5,25,31]

Equilibrium Constants. Equilibrium constants for 1:1 and 1:2 complexes of Mn(II) with 4-NC and with TBC are necessary to determine the concentration of all complexes in solution, before undertaking a systematic kinetic investigation. Only constants for ML and ML_2 chelate forms could be derived from the potentiometric data. The values for protonated constants K_2^H of the ligands in the different solvent systems and equilibrium constants $^*K_1$ and $^*K_2$ for chelate formation are reported in Tables II and III.

Table II. Log K_2^H Values for Auxiliary Ligands^a

Ligand	Medium	Log $K_2^{H\,b}$
4-NC	Aq 1.0 *M* KNO_3	6.59^c
4-NC	50% (w/w) CH_3OH–0.10 *M* KNO_3	7.27
4-NC	68% (w/w) CH_3OH–0.050 *M* KNO_3	7.77
TBC	68% (w/w) CH_3OH–0.050 *M* KNO_3	6.30

^{a} Determined potentiometrically with Beckman Research pH meter, 25° under N_2. ^{b} K_2^H = [H_2L]/[H^+][HL^-]. ^{c} Reference 11, corrected for *t*, electrolyte, and ionic strength.[22]

Table III. Equilibrium Constants for Metal-Chelate Complexes^a

Reaction	Medium	$-$Log $^*K_1{}^b$	$-$Log $^*K_2{}^c$
Mn(II) + 4-NC	Aq 1.0 *M* KNO_3	10.83^d	12.6^d
Mn(II) + 4-NC	50% (w/w) CH_3OH–0.01 *M* KNO_3	11.65	13.6
Mn(II) + 4-NC	68% (w/w) CH_3OH–0.050 *M* KNO_3	10.4	12.9
Mn(II) + TBC	68% (w/w) CH_3OH–0.050 *M* KNO_3	9.88	13.1

^{a} Determined potentiometrically with Beckman Research pH meter, 25° under N_2. ^{b} $^*K_1$ = [Mn(II)-4-NC][H^+]^2/[Mn(II)]·[4-NC]. ^{c} $^*K_2$ = [Mn(II)-(4-NC)$_2$][H^+]^2/[Mn(II)-4-NC][4-NC]. ^{d} Reference 11, corrected for *t* and ionic strength with data from Harned and Owen.[22]

Dependence on [H^+]. In order to identify the catalytic species a study of the variation of initial oxidation rate with pH was made for the following: with PC, for Mn(II) and Mn(II)-4-NC catalysts in aqueous solution and for Mn(II)-4-NC in 50% methanol; with 3,5-DTBP, for Mn(II) and Mn(II)-4-NC in 50% methanol and for Mn(II)-4-NC, Mn(II)-TBC, and Co(II)-TBC in 69% methanol. First-order rate constants were determined as outlined in the Experimental Section. These were plotted as log k_{obsd} *vs.* log [H^+] and the results are shown in Figure 2 for 3,5-DTBP in 50% methanol, in Figure 3 for 3,5-DTBP in 68% methanol, and in Figure 4 for PC in 50% methanol and aqueous solu-

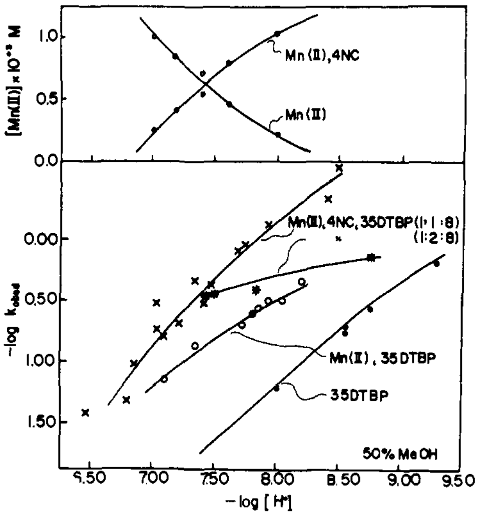

Figure 2. Dependence of metal chelate catalyzed oxidation of 3,5-DTBP in 50% CH_3OH on [H^+]. Lower half: log k_{obsd} *vs.* log [H^+], where k_{obsd} = v_0/[S] at *t* = 0. Conditions: [Mn(II)] = 1.25×10^{-3} *M*; [3,5-DTBP] = 1.0×10^{-2} *M*; 0.10 *M* KNO_3, under 100% O_2, 25°. Systems: (—×—×—) Mn(II)-4-NC-3,5-DTBP (1:1:8); (—*—*—) (1:2:8); (—○—○—) Mn(II)-3,5-DTBP; (—·—·—) 3,5-DTBP. k_{obsd} corrected to 760 mm, 25°. Upper half: concentration of principal Mn(II) species *vs.* $-$log [H^+]. Conditions: same as above; (—○—○—) Mn(II); (—●—●—) Mn(II)-4-NC (1:1) chelate.

---

(28) J. J. Morgan and W. Stumm, *Amer. Chem. Soc., Div. Water Waste Chem., Prepr.*, 3, 13 (1963).

(29) G. Schwarzenbach, *Helv. Chim. Acta*, 32, 839 (1949); R. Kucharkowski and H. G. Doege, *Fresenius Z. Anal. Chem.*, 238, 241 (1968).

(30) D. B. Broughton and R. L. Wentworth, *J. Amer. Chem. Soc.*, 69, 741 (1947).

(31) R. Flitman and E. Frieden, *ibid.*, 79, 5198 (1957).

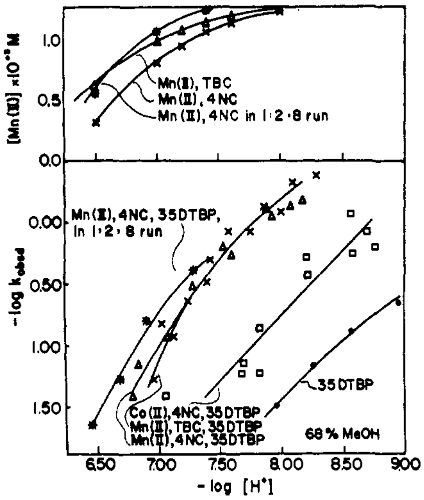

Figure 3. Dependence of metal chelate catalyzed oxidation of 3,5-DTBP in 68% CH_3OH on $[H^+]$. Lower half: log k_{obsd} *vs.* log $[H^+]$ where $k_{obsd} = v_0/[S]$ at $t = 0$. Conditions: [Mn(II)] = 1.25×10^{-3} *M*; [3,5-DTBP] = 1.0×10^{-2} *M*; 0.050 *M* KNO_3, under 100% O_2, 25°. Systems: (—×—×—) Mn(II)-4-NC-3,5-DTBP (1:1:8); (—*—*—) (1:2:8); (—△—△—) Mn(II)-TBC-3,5-DTBP (1:1:8); (—□—□—) Co(II)-4-NC-3,5-DTBP (1:1:8); (—●—●—) 3,5-DTBP. k_{obsd} corrected to 760 mm, 25°. Upper half: concentration of principal Mn(II) species *vs.* pH. Conditions: same as above; (—×—×—) Mn(II)-4-NC (1:1) chelate in 1:1:8 runs; (—*—*—) Mn(II)-4-NC (1:1) chelate in 1:2:8 runs (1:2 chelate is 10^2 lower in concentration); (—△—△—) Mn(II)-TBC (1:1) chelate.

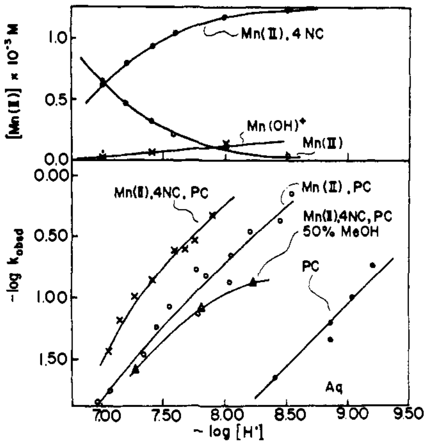

Figure 4. Dependence of metal chelate catalyzed oxidation of PC on $[H^+]$. Lower half: log k_{obsd} *vs.* log $[H^+]$ where $k_{obsd} = v_0/[S]$ at $t = 0$. Conditions: [Mn(II)] = 1.25×10^{-3} *M*; [PC] = 1.0×10^{-2} *M*; under 100% O_2, 25°. Systems: aq 1.0 *M* KNO_3, (—×—×—) Mn(II)-4-NC-PC (1:1:8); (—○—○—) Mn(II)-PC; (—●—●—) PC; 50% CH_3OH-0.10 *M* KNO_3, (—△—△—) Mn(II)-4-NC-PC (1:1:8). k_{obsd} corrected to 760 mm, 25°. Upper half: concentration of principal Mn(II) species *vs.* pH in aq solution; (—○—○—) Mn(II); (—×—×—) $Mn(OH)^+$; (—●—●—) Mn(II)-4-NC (1:1) chelate.

tion, all under 100% O_2 atmosphere. All metal-catalyzed and base-catalyzed systems show increasing k_{obsd} with increasing alkalinity. The falloff in k_{obsd} with Mn(II)-4-NC-3,5-DTBP (1:2:8) and Mn(II)-4-NC-PC (1:1:8) in 50% methanol at higher alkalinity (Figures 2 and 4) is presumably due to formation of inactive, polynuclear metal complexes, as indicated by the formation of precipitates at high pH in the potentiometric titration of Mn(II)-PC systems.[12]

Rate Law. The predominant species under the experimental conditions in Figures 2–4 (upper halves) are found from the equilibrium data in Table III to be the free metal ion and a 1:1 chelate between metal and auxiliary ligand. Both species are 50–100 times greater in concentration than other metal ion containing species up to pH 8.0. Because of this fact, a preliminary fit of the kinetic data (Table IV) to the expression

$$-\frac{d[S]}{dt} = \{k_1[M^{2+}][OR^-] + k_2[ML]\}[S^-] = k_{obsd}[S] \tag{6}$$

where $OR^- = OCH_3^-$ or OH^-, was attempted by calculating $k_1[M^{2+}]$ from the free metal-catalyzed reaction. In 68% methanol the contribution from free metal to the overall reaction is small and may be disregarded. Division of the difference ($k_{obsd} - k_1[M^{2+}][OR^-]$) by [ML] still gave rate constants which increased with increasing alkalinity. A second division by $[OR^-]$ (*i.e.*, $k_2/[OR^-]$) did give fairly constant values in almost every case. Mn(II)-4-NC-3,5-DTBP (1:1:8) in 50% methanol was an exception for reasons which again appear to be associated with polynuclear formation as $[OR^-]$ is increased. Based on the foregoing, it seems appropriate to rewrite eq 6 as

$$-\frac{d[S]}{dt} = \{k_3[OR^-][M^{2+}] + k_4[OR^-][ML]\}[S] \tag{7}$$

The product of $[OR^-]$ and [S] on the right hand side is monoanion, or SH^- (SH_2 and S^{2-} being designated as fully protonated and deprotonated substrate forms, respectively). This implies that the interaction between SH^- and the metal catalysts makes the major contribution to the overall rate. However, the first term on the right-hand side in eq 7 might be equally well satisfied by the expression $k_5 \cdot [(MnOH)^+]$.[32] This exchange would require that k_5 be 50 times greater than k_4 since the $[MnOH^+]/[M^{2+}]$ ratio in the free metal-catalyzed reaction from which k_5 is determined is less than 1:50 over the pH range studied. If $MnOH^+$ is the principal catalytic species in the free metal-catalyzed system, the rate, expressed as turnover of substrate molecules/molecule of $MnOH^+$, min^-, is 2×10^4 greater than the rate in the absence of metal under 100% O_2 atmosphere.

A further test that ML is the main catalytic form of the metal in Mn(II)-4-NC and Mn(II)-TBC systems was made by varying [ML] at constant [S]. The results are shown in Table V. The first-order dependence of k_{obsd} on [ML] is verified in both sets of experiments.

Time Course Kinetics. The metal-catalyzed oxidations almost invariably followed simple first-order kinetics for oxygen uptake over one or more half-lives (Figure 5). The second-order inhibitory term in O_2 uptake with 3,5-DTBP reported for the metal-free systems[9] was missing when metals were present. The runs at high [S] in Figure 1 exhibited zero-order behavior over the first few minutes. These results are ex-

---

(32) The equilibrium constant for an $[MnOR]^+$ complex in 50% methanol has been estimated to be 4×10^4 M^{-1} from potentiometric measurements.[10] Mn^{2+} binds preferentially with H_2O in low methanol content solvents (G. Atkinson and S. K. Kor, *J. Phys. Chem.*, **69**, 128 (1965)) but the solvation sphere is unknown in high methanol content solvents (H. Tsubota and G. Atkinson, *J. Phys. Chem.*, 71, 1131 (1967)). The complex detected[10] is assumed to be $[MnOH^+]$.

Table IV. Rate Constants for Mn(II)–Chelate Catalyzed Oxidation[a]

$-\text{Log}\,[\text{H}^+]$	$10^2 \times k_{\text{obsd}}$,[b] min^{-1}	$10^4\,M \times [\text{Mn(II)}]$[c]	$10^2 \times v_M/[\text{S}]$,[d] min^{-1}	$10^2 \times k_2[\text{Mn(II)-4-NC}]$,[e] min^{-1}	$10^{-2} \times k_2$,[f] $M^{-1}\,\text{min}^{-1}$	$10^{-9} \times k_2/[\text{OR}^-]$,[g] $M^{-2}\,\text{min}^{-1}$
Mn(II)-4-NC-3,5-DTBP in 50% CH_3OH–0.10 M KNO_3						
7.0	13.40	10.1	5.75	7.65	3.20	2.5
7.2	23.50	8.48	6.75	16.75	4.16	1.94
7.4	38.20	7.06	7.10	31.10	5.72	1.72
7.6	60.50	4.68	6.85	53.65	6.88	1.33
8.0	139.10	2.10	5.60	133.50	12.85	0.96
Mn(II)-4-NC-PC in aq 1.0 M KNO_3						
7.0	2.98	6.40	0.64	2.34	0.384	0.23
7.2	7.00	4.60	0.92	6.08	0.770	0.29
7.4	13.40	3.10	1.08	12.32	1.310	0.31
7.6	23.20	2.00	1.15	22.05	2.10	0.31
8.0	58.50	0.90	1.34	57.2	4.93	0.29
Mn(II)-4-NC-3,5-DTBP in 68% CH_3OH–0.050 M KNO_3						
7.0	7.70	4.23		7.70	9.63	8.0
7.2	19.10	3.07		19.10	20.30	10.6
7.4	40.00	2.00		40.0	38.0	12.6
7.6	67.7	1.30		67.7	60.5	12.6
8.0	159.0	0.40		159.0	131.5	11.0
Mn(II)-TBC-3,5-DTBP in 68% CH_3OH–0.5050 M KNO_3[g]						
7.0	10.00	2.72		10.00	10.25	8.6
7.2	20.70	1.80		20.70	19.35	10.1
7.4	40.20	1.20		40.2	35.6	12.1
7.6	67.0	0.70		67.0	56.8	11.8
8.0	140.0	0.10		140.0	113	9.4
Mn(II)-4-NC-3,5-DTBP (1:2:8) in 68% CH_3OH–0.050 M KNO_3						
6.50	2.50	7.00		2.50	4.60	12.1
7.0	17.8	1.90		17.8	17.0	14.0
7.4	50.2	0.30		50.2	40.2	13.6

[a] Under 100% O_2, 25°, corrected to 760 mm and 25°. [b] $k_{\text{obsd}} = v_0$ (observed initial rate of oxygen uptake in M/min) divided by [S]. [c] [Mn(II)] determined from equilibrium data, Table IV, and initial analytical concentration of Mn(II) = 1.25×10^{-3} M. [d] v_M = oxygen uptake rate *via* Mn(II)-catalyzed oxidation under identical experimental conditions; $[\text{Mn(II)}]/1.25 \times 10^{-3}$ M. [e] $k_2[\text{Mn(II)-4-NC}] = (v_0 - v_M)/[\text{S}]$. [f] $k_2[\text{Mn(II)-4-NC}]/[\text{Mn(II)-4-NC}]$ where $[\text{Mn(II)-4-NC}] = 1.25 \times 10^{-3}\ M - [\text{Mn(II)}]$. [g] k_4 in eq 7; $[\text{OR}^-]$ determined from estimated K_w = 1.7, 1.3, and 1.2×10^{-14} M^2 in aq 1.0 M KNO_3–50% CH_3OH–0.10 M KNO_3 and 68% CH_3OH–0.05 M KNO_3, respectively.

Table V. Dependence of Rate on Catalyst Concentration

Catalytic system	Molar ratio[a] components	Medium	$[\text{Mn(II)-4-NC}] \times 10^3$	$10 \times k_{\text{obsd}}$,[b] min^{-1}	$k_{\text{obsd}}/[\text{Mn(II)-4-NC}]$
Mn(II)-4-NC-PC	1:1:8	Aq 1.0 M KNO_3	1.05	2.3	230
			0.50	1.2	240
Mn(II)-4-NC-3,5-DTBP	1:1:8	50% CH_3OH–0.10 M KNO_3	0.78	6.05	775
			0.35	2.8	800

[a] [3,5-DTBP] = 1.00×10^{-2} M; [Mn(II)] = [4-NC] within ± 1%. [b] Determined manometrically from initial rate at pH 7.75, under 100% O_2 atmosphere, corrected to 760 mm, 25°.

pected if the catalytic system conforms to an initial rate law in which the phenolic substrate forms an intermediate complex with the catalytic species and if v_0 (initial velocity of oxygen uptake) > $V_{\max}$ (initial velocity with [3,5-DTBP] saturating)/2.

$1/v$ *vs.* $1/[\text{S}]$. On the basis of the foregoing, and assuming that steady-state conditions hold, a general rate equation for the change in initial velocity as a function of the reaction component concentrations was developed. The resulting expression is analogous to that derived for two substrate enzyme kinetics and is written as

$$v = \frac{k_6[\text{ML}][\text{S}][\text{O}_2]}{K_B\bar{K}_A + K_B[\text{S}] + K_A[\text{O}_2] + [\text{S}][\text{O}_2]} \tag{8}$$

where $\bar{K}_A$ is initially taken as the formation constant for the mixed complex intermediate, $MLSH^-$ or MLS^{2-}, and K_B and K_A are Michaelis constants for O_2 and S, respectively. The term $K_4[\text{OR}^-]$ in eq 7 corresponds to $K_6[\text{O}_2]/\{K_B\bar{K}_A + K_2[\text{O}_2]\}$ when $[\text{S}] \lesssim V_{\max/2}$. Reciprocal plots of v *vs.* [S] at saturating $[O_2]$ were linear with intercepts on the x axis. At the intercept, $1/\bar{K}_A = 1/[\text{S}]$. These intercepts were pH dependent (*e.g.*, Figure 6), which indicated $\bar{K}_A$ was also. A comparison of kinetically determined $\bar{K}_A$ and known MS complex formation constants then was made as in Table VI. Closer agreement exists between $\bar{K}_A$ and K_{MSH^+}' for Mn(II)-4-NC-PC and Mn(II)-PC than between $\bar{K}_A$ and K_{MS}'. Formation constants for chelates are at least an order of magnitude too small compared with $\bar{K}_A$ in all cases. Also $\bar{K}_A$ in Table VI for the Mn(II)-4-NC-3,5-DTBP systems increases almost linearly with $[OH^-]$ and certainly not with $[OH^-]^2$, indicating that the pH dependence of $\bar{K}_A$ is consistent with $MLSH^-$ formation and not MLS^{2-} as the kinetic intermediate.

$1/v$ *vs.* $1/[O_2]$. The good agreement between $\bar{K}_A$ and K_{MSH^+} suggests that either the pyrocatechol is the substrate that first complexes with ML, leading to a sequential mechanism, or that both substrates complex

Table VI. K_M for Substrates Compared with Metal–Substrate Equilibria

Catalytic systems	Conditions[a]	Apparent[b] K_M	K_{MS}'[c]	K_{MHS^+}'[d]
Mn(II)-4-NC-PC	Air, 50% CH_3OH, pH 7.30	22	0.8	
Mn(II)-4-NC-3,5-DTBP	Air, 50% CH_3OH, pH 7.30	27	0.8	
Mn(II)-4-NC-3,5-DTBP	100% O_2, 50% CH_3OH, pH 7.75	42	7.0	
Mn(II)-4-NC-PC	100% O_2, aq, pH 7.60	20	1.0	20
Mn(II)-PC	100% O_2, aq, pH 7.60	22	1.0	20

[a] Aqueous 1.0 *M* KNO_3–50% CH_3OH–0.10 *M* KNO_3, 25°. [b] K_M = [S] from *x* intercept of $1/v_0$ *vs.* 1/[S] plot. [c] $K_{MS}' = K_{MS}/\{1 + K_1^H[H^+] + K_1^HK_2^H[H^+]^2\}$, where $K_{MS} = [MS]/[M^{2+}][S^{2-}]$. Log K_1^H, K_2^H, $^*K_{MS}$(Mn-3,5-DTBP) = 10.35, 14.7, −14.7; (Mn-PC, 50% CH_3OH) = 10.0, 14.1, −14.7; (Mn-PC, aq) = 9.23, 13.1, −14.8, where $^*K_{MS} = [MnS][H^+]^2/[SH_2][Mn^{2+}]$. [d] $K_{MHS^+}' = K_{MHS^+}/\{1 + K_2^H[H^+]\}$, where $K_{MS^+} = [MHS^+]/[M^{2+}][HS^-]$. Log $^*K_{MHS^+} = -6.4$.[11]

independently (randomly).[33–36] If O_2 is the lead substrate, $1/v$ *vs.* $1/[O_2]$ plots with different pyrocatechols under the same experimental conditions should intersect at the same point on the negative *x* axis.[33] This is clearly not the case with the two pyrocatechols under study in either 50 or 68% methanol (Figure 7).

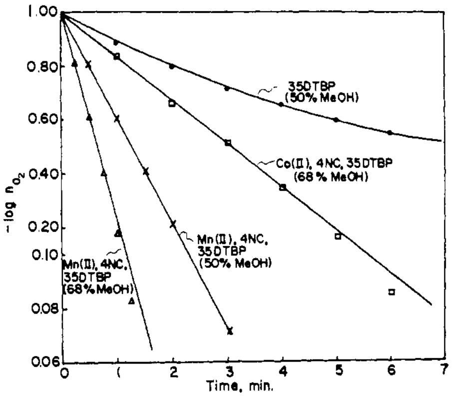

Figure 5. Time course of metal chelate catalyzed oxidation of 3,5-DTBP. First-order plots of oxygen uptake $-\log [n_{O_2}]$ *vs.* time, in min, where n_{O_2} is unreacted moles of O_2. Conditions: [Mn(II)] = 1.25×10^{-3} *M*; [3,5-DTBP] = 1.0×10^{-2} *M*, under 100% O_2, 25°. (—△—△—) Mn(II)-4-NC-3,5-DTBP (1:1:8), 68% MeOH, pH 7.90, 0.050 *M* KNO_3; (—×—×—) same system in 50% CH_3OH, pH 7.72, 0.10 *M* KNO_3; (—□—□—) Co(II)-4-NC-3,5-DTBP (1:1:8), 68% CH_3OH, pH 8.17, 0.050 *M* KNO_3; (—●—●—) 3,5-DTBP, 50% CH_3OH, pH 8.87, 0.10 *M* KNO_3.

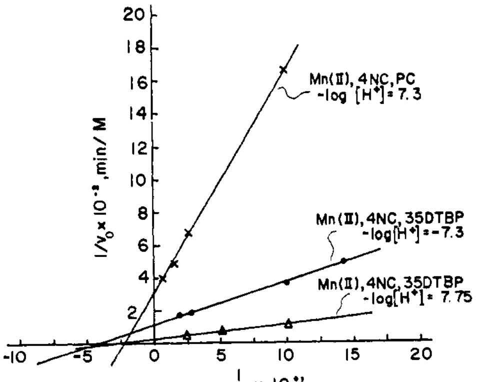

Figure 6. Reciprocal plots of initial velocity, v_0, against substrate concentration, [S]. Conditions: [Mn(II)] = 1.25×10^{-3} *M*; 50% CH_3OH–0.10 *M* KNO_3. (—△—△—) Mn(II)-4-NC-3,5-DTBP (1:1:8), pH 7.75 under 100% O_2; (—·—·—) same system at pH 7.30, under air; (—×—×—) Mn(II)-4-NC-PC (1:1:8), pH 7.30, under air, 25°.

It may be observed that the pyrocatechol concentrations used were not saturating. A further check, however, by observing that $K_A = V_1/k_1E_t$ for the hypothetical case where oxygen is the lead substrate was not constant, was sufficient to rule out this possibility.[37]

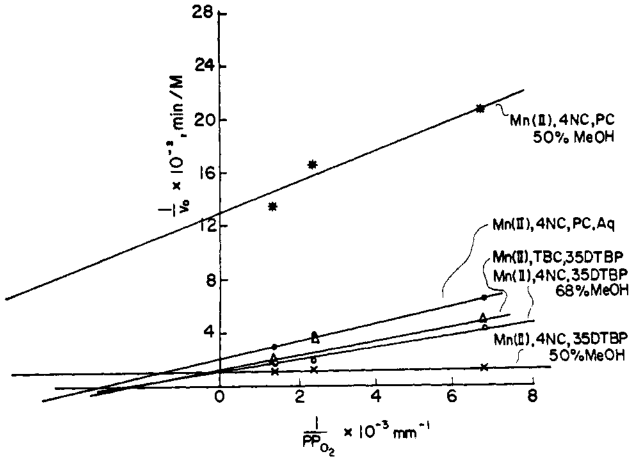

Figure 7. Reciprocal plots of initial velocity, v_0, against 1/partial pressure of O_2. Conditions: [Mn(II)] = 1.25×10^{-3} *M*; [S] = 1.0×10^{-2} *M*, 50% CH_3OH–0.10 *M* KNO_3. (—×—×—) Mn(II)-4NC-3,5-DTBP (1:1:8); (—*—*—) (Mn(II)-4-NC-PC (1:1:8). 68% CH_3OH–0.050 *M* KNO_3: (—○—○—) Mn(II)-4NC-3,5 DTBP (1:1:8); (—△—△—) Mn(II)-TBC-3,5-DTBP (1:1:8). Aqueous 1.0 *M* KNO_3: (—●—●—) Mn(II)-4-NC-PC (1:1:8); pH 7.75, each solvent system, 25°.

Discussion

It appears from the variation in autoxidation rate with alkalinity (Table V) and with metal–auxiliary ligand concentration (Table VI) that the 1:1 chelate is the dominant catalytic entity in mildly alkaline solution. A measure of its catalytic effectiveness in autoxidation reactions may be had by expressing rate turnover numbers derived earlier from Figure 1 in terms of moles of S/mole of MnL·min instead of Mn(II). Rate enchancements on the order of 10^3 molecules of 3,5-DTBP or PC per MnL molecule take place under an atmosphere of air. These turnover numbers are about 10 times faster than with metal ions alone. They are comparable to polymeric Cu(II) catalysts with ascorbic acid and hydroquinone[38,39] and demonstrate that Mn(II)

---

(33) L. L. Ingraham, *J. Amer. Chem. Soc.*, 78, 666 (1956).
(34) H. R. Mahler and J. Douglas, *ibid.*, **79**, 1159 (1957).
(35) C. C. Wratten and W. W. Cleland, *Biochemistry*, 4, 2442 (1965).
(36) C. H. Blomquist, *Acta Chem. Scand.*, **20**, 1747 (1966).
(37) W. W. Cleland, *Biochim. Biophys. Acta*, **67**, 104 (1963).
(38) I. Pecht, A. Levitzki, and M. Anbar, *J. Amer. Chem. Soc.*, **89**, 1957 (1967).
(39) N. A. Vengerova, Y. E. Kirsh, V. A. Kabanov, and V. A. Kargin, *Dokl. Akad. Nauk SSSR*, **190**, 131 (1970).

chelates are very efficient oxidizing catalysts when appropriate ligands are chosen. Substitution of TBC for the 4-NC molecule does not have a marked effect on the autoxidation rate constants (Table IV and Figure 7), which suggests that we may be near the limit of rate enhancement for this class of chelates with Mn(II).

From the pH-dependent studies (Figures 2–4 and Table IV) and variation in x axis intercepts of reciprocal plots with $[OH^-]$ rather than $[OH^-]^2$ (Figure 5) it is concluded that the pyrocatechol species, interacting mainly with the catalyst over the pH range 6–9, is the monoanion (eq 7). Support for this conclusion also comes from better agreement of the kinetically determined formation constant with the potentiometrically determined value for the Mn(II) complex with PC monoanion than with the dianion (Table VI). This finding differs from that predicted by Grinstead[8] in HCO_3^- solution in that the rate-determining step does not involve chelation of the PC dianion either by the manganese(II) chelate, Mn(II)-L, or by the manganese(II) ion. A kinetic investigation had not been carried out in that work.

While rate equations such as 8 are not expressions of a unique reaction mechanism, systematic analysis of x axis intercept variations with different reactant analogs can lead to reasonable conclusions about the order of binding to the catalyst, if an order sequence exists.[33–36,40] Thus, the variation in the x axis intercept of a $1/v$ *vs.* $1/[O_2]$ plot with different analogs at saturating levels indicates that the intercept values do not correspond to the formation constant for an $ML{-}O_2$ complex (is not $1/\bar{K}_A$ in eq 8) and, from this fact, that O_2 is probably neither the lead reactant in an order sequence nor randomly bound in a kinetic sense. It may by inferred that, when S is either PC or 3,5-DTBP, the following sequence of steps occurs

$$SH_2 \overset{K_2^H}{\rightleftharpoons} H^+ + SH^- \qquad (9)$$

$$MnL + SH^- \underset{\bar{K}_A}{\rightleftharpoons} MnLSH^- \qquad (10)$$

$$MnLSH^- + O_2 \rightleftharpoons MnL(SH)O_2^- \qquad (11)$$

$$MnL(SH)O_2^- \longrightarrow MnL + Q + O_2 \qquad (12)$$

The sequence envisions complexation of catalyst and reactants in a ternary complex prior to electron transfer, in line with the observed kinetics (Figures 1, 6, and 7). A probable arrangement of ligands bound to the metal ion in this complex is suggested by Figure 8.

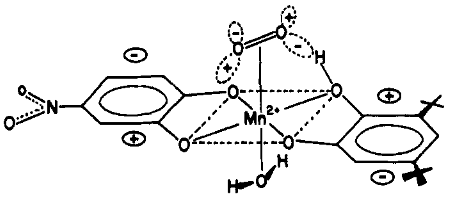

Figure 8. Schematic representation of transition state for Mn(II)-L catalyzed oxidation of pyrocatechols. Circled charges represent wave function symmetry for maximum π electron overlap between substrates.

The kinetics, however, give no information about the number of electrons transferred in the rate-determining step. If one electron, the free radicals created dismute to form quinone and O_2H^-; if two, an ionic mechanism results. This question was insoluble in the case of ascorbic acid but can likely be resolved by fast reaction techniques using more readily autoxidizable pyrocatechols in mildly alkaline solution. In advance of such studies it should not be assumed that the hydrogen ion for charge neutralization on the oxygen comes from the medium.[4] Rabani and Nielson[41] have shown that the dismutation $\cdot O_2H + \cdot O_2^- \rightarrow {}^-O_2H + O_2$ is a much faster route to H_2O_2 in mildly alkaline solution than dismutation of either $\cdot O_2$ or $\cdot O_2H$. Alternatively, intracomplex transfer of a hydrogen atom to bound O_2 (or if kinetically indistinguishable, H^- to O_2) is quite possible with *cis*-diol type compounds, considering the proximity of phenolic or alcoholic hydrogen to oxygen likely to exist in the activated complex. There is no experimental evidence here or in the literature to favor one or the other mechanism for hydrogen atom transfer, for the systems under investigation.

The Co(II)-4-NC catalyzed oxidation was not studied in detail beyond confirming that log K_{obsd} *vs.* log $[H^+]$ was linear (Figure 3) over the pH range 7.2–8.5 by regression analysis. The Co(II)-4-NC catalyzed decomposition of H_2O_2 in 68% CH_3OH occurs in the absence of substrate. The Co(II)-4-NC-3,5-DTBP system is worthy of further study, since it is one of the few nonenzymic reactions so far reported to give *o*-quinone without H_2O_2 accumulation, the same products as those obtained in 2-electron transferase reactions with pyrocatechols and O_2.[4]

(40) J. T-F. Wong and C. S. Hanes, *Can. J. Biochem. Physiol.*, **40**, 763 (1962).

(41) J. Rabani and S. O. Nielson, *J. Phys. Chem.*, 73, 3736 (1969).