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Activation of Hydrogen Peroxide for Oxidation by Copper(II) Complexes

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Copper(II) complexes of the general formula $[LCu(H_2O)_4]^{n+}$ (where L is a bidentate ligand and $n = 1$ or 2) activate hydrogen peroxide for the oxidation of quinaldine blue, an oxidation indicator. The copper(II) complexes of tri- and tetradentate ligands are shown to be inactive, as are the *bis*-complexes of bidentate ligands. The proposed mechanism for peroxide activation involves the formation of a copper(II)–hydroperoxide complex, which then rapidly oxidized the substrate. Comparison of reaction rates with different ligand systems, and different ligand-to-metal ratios, lead to the conclusion that the most active complexes are those in which two equatorial coordination positions are occupied by easily displaced water to form the active catalyst. Rate studies are performed which give an experimental rate law which is first order in copper(II) complex, zero order in substrate, and variable order in peroxide. These kinetics are predicated by the rate law derived from our proposed mechanism. The variable order in peroxide can be explained in terms of Michaelis–Menten-type kinetics, as linear Lineweaver–Burk plots of (rate^{-1}) vs $([O_2H^-]^{-1})$ are obtained from our experimental data. This is consistent with our proposed mechanism, as the derived rate law can be rearranged into the Michaelis–Menton equation. © 1997 Academic Press

INTRODUCTION

Environmental concerns have motivated replacement of aqueous hypochlorite as an oxidant with peroxide and use of water as a reaction medium. Activation of hydrogen peroxide by transition metal complexes for the oxidation of organic substrates has been studied extensively (1–3). Few transition metal complexes catalyze oxidations by hydrogen peroxide in water, and fewer are active above pH 8. Precipitation of the catalytically active metal complex as the oxide or hydroxide at higher pH is a major limitation. Chelating ligands can stabilize metal ions in basic solution and prevent precipitation (4). Thus, a manganese complex of 1,4,7-triazocyclononane (TACN) (5) is an effective peroxide activator in basic aqueous solution. This complex catalyzes the epoxidation of styrene and 4-vinylbenzoic acid with hydrogen peroxide and was commercialized as a bleaching agent for household applications. This discovery has heightened interest in aqueous transition metal-catalyzed peroxide activation with a primary goal of finding complexes that react with hydrogen peroxide to produce active oxidants in basic aqueous solution.

Activation of hydrogen peroxide by metal redox mechanisms (6) almost invariably leads to catalysis of peroxide decomposition. This can be avoided by employing catalysts that function by a Class IVb mechanism (6). In this mechanism peroxide coordinates to the metal but does not change its oxidation state. Substrate attacks the coordinated, activated hydrogen peroxide which transfers an oxygen atom. Coordination can be viewed as replacing the proton of H_2O_2 with a Lewis acid (the metal complex) and, in the process, making MO_2H more electrophilic than HO_2^-. Accordingly, a series of experiments were carried out with complexes that might function as peroxide activators through hydroperoxide coordination.

Bis(ethylenediamine) (7) and bis(bipyridine) (8) copper complexes are reported to be active for peroxide activation, though not in aqueous base. Many copper complexes containing a bound peroxo species are known, although most are dimeric species (9–12). Complexes with cumylhydroperoxide and *tert*-butylhydroperoxide as ligands to a copper(II) monomer have been isolated (13).

This paper examines the activation of hydrogen peroxide by a series of copper(II) complexes. The initial experiments were performed with picolinic acid ligand, which is effective in promoting peroxide activation with iron in a pyridine–acetic acid solvent mixture (2, 3). Finding activity in aqueous base with the copper(II)–picolinic acid complex, the study was expanded to other ligands.

Applications of catalysts for oxidation in basic aqueous solutions require broad base oxidants. This requirement is in contrast to selective oxidations whose objective is to produce a single desired product. Accordingly, a redox indicator, quinaldine blue, was chosen as the substrate for oxidation. Though the oxidation products are unknown, this substrate has several advantageous properties for a study whose primary focus is to ascertain desirable metal complex properties for activation. Initial rates of oxidation of quinaldine blue (I) can be measured conveniently as a function of pH, peroxide concentration, ligand-to-metal ratio,

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and complex concentration. A mechanism for peroxide activation with complexes of this type is proposed.

EXPERIMENTAL

Materials

Copper(II) chloride and sodium tetraborate were used as received from Fisher. Hydrogen peroxide (30%) was used after iodometric titration to verify its concentration. Quinaldine blue, 2-pyridinecarboxylic acid (picolinic acid, PA), 2,4-pentanedione (acetylacetone, AcAc), 1-aminoacetic acid (glycine, Gly), 2-aminopropanoic acid (alanine, Ala), 4-amino-2-methylbutonic acid (leucine, Leu), 2,6-pyridinedicarboxylic acid (dipicolinic acid, DPA), (iminodiacetic acid, IDA), 1,2-diaminoethane (ethylenediamine, EN), *N*,*N*,*N*′,*N*′-tetramethylethylenediamine (TMED), 2,9-dimethyl-1,10-phenanthroline (DMP), 2,2′-dipyridylamine (bipyridylamine, BPA), tris(2-aminoethyl)amine (TREN), bis(1-aminopropyl)amine (IBPA), and 2,2′ : 6′,2″-dipyrdylpyridine (terpyridine, TERPY) were all used as received from Aldrich. Water used for the peroxide solutions was purified by a Barnstead NANOpure unit to eliminate metal–ion contaminants that might cause peroxide decomposition. Its final resistance was 18 megohm $\cdot$ cm^{-1}.

Measurements

UV–Vis measurements were performed on a Perkin–Elmer Lambda-6 spectrophotometer at room temperature. Rate data were obtained by monitoring the decrease in quinaldine blue concentration at 600 nm over time. pH measurements were made with a Fisher Accumet Model 630 pH meter.

Preparation of Reaction Mixtures for Kinetic Measurements

(1) Preparation of the substrate. Quinaldine blue, I, 0.0060 g, was dissolved in a 0.01 *M* aqueous sodium tetraborate buffer solution. This gives a deep purple solution with an absorbance between 1.0 and 1.1 at 600 nm. The initial absorbance was checked before the solution was used to ensure that it was in the above range. The substrate oxidation was performed at pH 9.1 in a 0.1 *M* borax buffer unless otherwise indicated.

(2) Preparation of the catalysts. The catalyst solution used was prepared from 2.5×10^{-3} *M* aqueous copper(II) chloride and 2.5×10^{-3} *M* aqueous solutions of the appropriate ligands, as previously described (2, 8) for *in situ* generation of active catalysts. Use of copper(II) nitrate, rather than the copper(II) chloride, as the copper(II) source did not significantly affect the rates, nor did addition of excess chloride to the system. At these concentrations, precipitation was not observed, and the catalyst solutions did not loose activity upon standing.

(3) Preparation of the peroxide solution. A 0.200 *M* stock solution of hydrogen peroxide was made by adding 10 mL of 30% hydrogen peroxide (whose concentration was checked by iodometric titration) to a 500-ml volumetric flask and diluting with purified water to 500 ml. Subsequent dilutions were made to give the peroxide solutions used. All the rate data were obtained with peroxide solutions that were no more than 48 h old.

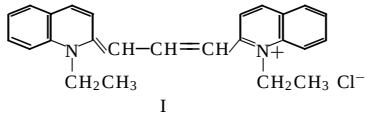
(4) Procedure. In each experiment, 2 mL of the buffered substrate (I) solution, 1 mL of catalyst solution (in blank runs, water), and 0.5 mL of peroxide solution were added, in that order, to a UV–Vis cell. This combination of buffer/substrate solution, copper complex, and hydrogen peroxide has a pH of 9.1. The cell was capped and shaken for 5 s and then placed in the cell holder in the UV–Vis, at which point data collection was initiated. Decreasing absorbance at 600 nm was monitored as a function of time. The initial rates of reaction were determined from data collected during the first 10 to 20 s of reaction. These initial rates are given under Results and Discussion.

RESULTS AND DISCUSSION

Substrate Selection

Redox indicators (15–19) have been used to measure peroxide activation because the reaction can be monitored continuously and conveniently. This permits screening of a large number of systems and simpler evaluation of the inorganic chemistry of peroxide activation. For this study quinaldine blue (pinacyanol chloride), I, is used.

Quinaldine blue is a redox indicator that has convenient spectral properties to monitor peroxide activation.



I

Its absorbance maximum at 600 nm is insensitive to pH's above pH 4. (I) is not oxidized by H_2O_2 at the concentrations used in this work. Stronger oxidants such as oxone (monosodium peroxysulfate) or hypochlorite produce a colorless product with negligible absorbance at 600 nm. Reactions of (I) are readily monitored with UV/Vis spectroscopy.

Copper(II) complexes have been shown to oxidize amine containing buffers in aqueous solution without peroxide (19, 20). To rule out the possibility of this occurring here, the absorbance of a solution containing equimolar concentrations of copper(II) and copper(II) with picolinic acid and (I) showed no change in absorbance over a 2-h period. We therefore conclude that copper(II) does not oxidize I on the timescale of our experiments.

Several products can result from the oxidation of I (14). For elucidating the essential properties of the complex to activate H_2O_2 in basic solution this information is not essential. Catalyst specificity is expected and these conclusions are expected to be substrate dependent (6).

Activation of Peroxide by Copper(II) Chloride and Picolinic Acid

The first set of copper(II) complexes examined as catalysts uses picolinic acid as the ligand. A solution containing 1.0×10^{-4} *M* picolinic acid and copper(II) chloride (a 1 : 1 mole ratio) was found to effectively activate peroxide for the oxidation of I. The absorbance at 600 nm versus time for this reaction is shown in Fig. 1. Picolinic acid itself does not catalyze the oxidation of I by hydrogen peroxide. In the absence of a strong-binding ligand, copper(II) forms the insoluble hydroxide which decomposes H_2O_2 on the surface of the solid particles. Only a small amount of quinaldine blue oxidation occurs in the absence of ligand.

The activity of the copper(II)–picolinic acid system as a function of the ligand : copper(II) mole ratio at constant [copper(II)] is shown in Fig. 2. The decrease noted at ratio of 1.5 to 1 is larger than expected for 1 : 1 mixture of 1 : 1 and 2 : 1 complexes and could be due to a 3 : 2 aggregate. The most active species is formed at a 1 : 1 ratio of copper to ligand. Stability constants predict that $Cu(PA)^+$ is the dominant species under these conditions (21).

Oxidations were studied with different concentrations of $Cu(PA)^+$, peroxide, and I (Table 1). Runs 1–8 in Table 1 give the initial rates of quinaldine blue oxidation with a [Cu(II)] of 2.9×10^{-5} *M* and $[H_2O_2]$ between 5.7×10^{-4} and 2.9×10^{-2} *M*.

The reaction is zero order in $[H_2O_2]$. Repeated rate measurements were reproducible to 5%. Next, the peroxide concentration was held constant (0.029 *M*), while the $Cu(PA)^+$ concentration was varied. These results are shown in Table 1 as runs 9–14. Consistent with our discussion of Fig. 2 the ln(initial concentration) versus ln(rate) plot is nonlinear. This is in contrast to the systems in Table 3 where the reaction is first order in [Cu(II)]. The order with respect to substrate is zero, as determined from runs 13, 15, and 16. A mechanism consistent with these observations will be presented below.

All rate calculations involved initial rates, at low conversions. However, all reactions in Table 1 proceed to complete oxidation of the substrate as shown in Fig. 1. Thus, runs 9–12 in Table 1 are catalytic with the complex undergoing 10 or more turnovers.

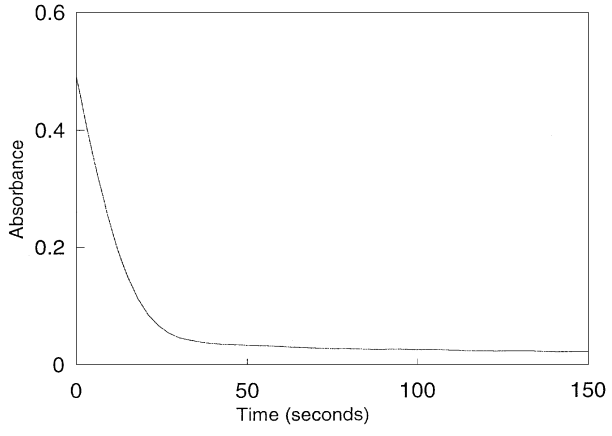


FIG. 1. Oxidation of quinaldine blue with hydrogen peroxide catalyzed by copper(II) chloride + picolinic acid. $[H_2O_2] = 2.9 \times 10^{-3}$ *M*, $[Cu(II)] = 2.9 \times 10^{-5}$ *M*.

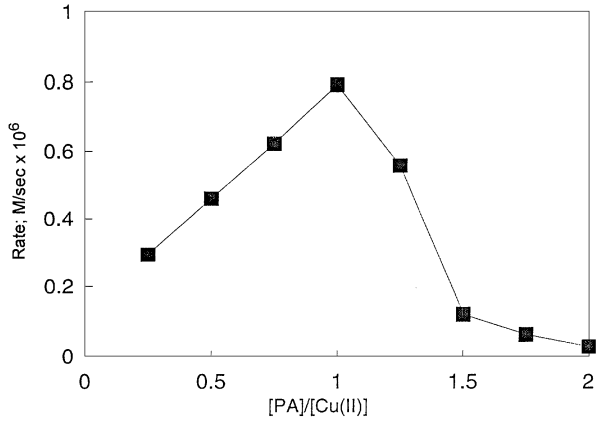


FIG. 2. Effect of ligand : metal ratio on the rate of quinaldine blue oxidation. $[Cu(II)] = 2.9 \times 10^{-5}$ *M*, $[H_2O_2] = 5.7 \times 10^{-3}$ *M*. All rates are expressed as molars per second.

TABLE 1

Dependence of Quinaldine Blue Oxidation Rate on Peroxide, $Cu(PA)^+$, and Substrate Concentrations[a]

Run	$[H_2O_2]$[a]	[Quinaldine blue]	$[Cu(PA)^+]$	Rate[b]
1	5.70×10^{-4}	3.0×10^{-5}	2.9×10^{-5}	8.4×10^{-7}
2	5.70×10^{-4}	3.0×10^{-5}	2.9×10^{-5}	8.3×10^{-7}
3	2.86×10^{-3}	3.0×10^{-5}	2.9×10^{-5}	7.9×10^{-7}
4	2.86×10^{-3}	3.0×10^{-5}	2.9×10^{-5}	7.8×10^{-7}
5	8.59×10^{-3}	3.0×10^{-5}	2.9×10^{-5}	7.9×10^{-7}
6	8.59×10^{-3}	3.0×10^{-5}	2.9×10^{-5}	7.6×10^{-7}
7	1.72×10^{-2}	3.0×10^{-5}	2.9×10^{-5}	7.1×10^{-7}
8	2.86×10^{-2}	3.0×10^{-5}	2.9×10^{-5}	7.3×10^{-7}
9	2.86×10^{-2}	3.0×10^{-5}	1.4×10^{-6}	9.2×10^{-9}
10	2.86×10^{-2}	3.0×10^{-5}	1.8×10^{-6}	2.0×10^{-8}
11	2.86×10^{-2}	3.0×10^{-5}	2.3×10^{-6}	5.4×10^{-8}
12	2.86×10^{-2}	3.0×10^{-5}	1.1×10^{-5}	3.2×10^{-7}
13	2.86×10^{-2}	3.0×10^{-5}	2.9×10^{-5}	7.3×10^{-7}
14	2.86×10^{-2}	3.0×10^{-5}	8.6×10^{-5}	9.3×10^{-7}
15	2.86×10^{-2}	2.0×10^{-5}	2.9×10^{-5}	7.6×10^{-7}
16	2.86×10^{-2}	1.0×10^{-5}	2.9×10^{-5}	7.5×10^{-7}

[a] All concentrations are in molars.
[b] Rates are expressed as molars per second.

Effect of pH on Quinaldine Blue Oxidation Rate

The rate of quinaldine blue oxidation catalyzed by the copper (II)–picolinic acid system was measured at different pH's. The copper complex concentration was held constant at 2.9×10^{-5} *M*, and the peroxide concentration was 2.9×10^{-3} *M*. The pH was adjusted by addition of small amounts of concentrated KOH solution to a solution of the catalyst and sodium tetraborate. Figure 3 shows the rate dependence on pH. The rate of quinaldine blue oxidation is negligible below pH 8.0. The rate of reaction increases up to pH 10.0, where it becomes independent of pH. The pH did not change during the reaction.

Peroxide Activation by Copper(II) Complexes of Anionic Bidentate Ligands

Experiments were performed with AcAc, Gly, Ala, and Leu as ligands, with $[Cu(II)] = 2.8 \times 10^{-5}$ *M* and $[H_2O_2] = 2.8 \times 10^{-2}$ *M*. The oxidation rate was studied with 1 : 1 and 2 : 1 ratios of ligand to copper(II). Initial rates of oxidation are shown as runs 1–10 in Table 2.

All these amino acid complexes catalyze the oxidation of quinaldine blue by hydrogen peroxide, with glycine being the least active and leucine being the most active at a 1 : 1 ratio of ligand to metal. The copper(II) complex of acetylacetone showed the highest activity, with an oxidation rate 15–20% greater than for the copper(II) leucine complex. As with picolinic acid, all of the complexes exhibited a sharp decrease in activity when the ligand to metal ratio was greater than 1.

Peroxide Activation by Copper(II) Complexes of Tridentate Anionic Ligands

Two tridentate anionic ligands, DPA and IDA were examined. The results are shown as runs 11–14 in Table 2. At a 1 : 1 mole ratio, tridentate anionic ligands lead to much poorer catalysts for peroxide activation than bidentate anionic ligands. Complexes with tridentate anionic ligands show little activity for peroxide activation at a 2 : 1 ligand to copper ratios. Lower activity for complexes of anionic tridentate ligands shows that a positively charged complex is needed or that two adjacent equatorial coordination sites must be occupied by water to generate an effective copper(II) catalyst for peroxide activation.

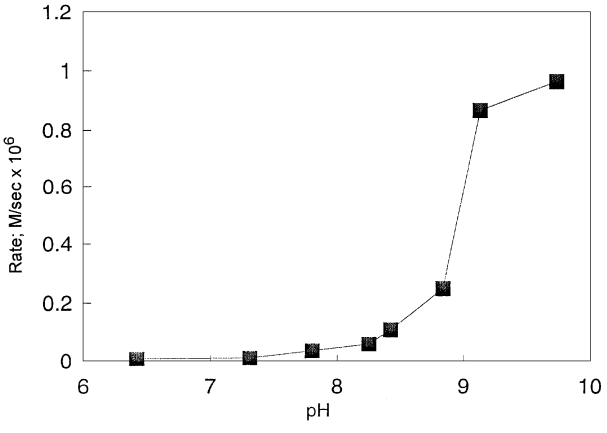


FIG. 3. Effect of pH on the rate of quinaldine blue oxidation. $[Cu(II)] = 5.7 \times 10^{-3}$ *M*]. Rates are expressed in molars per second.

TABLE 2

Rates of Quinaldine Blue Oxidation with Copper(II) Complexes of Anionic Bidentate (Runs 1–10), Anionic Tridentate (Runs 11–14), Neutral Bidentate (Runs 15–24), and Neutral Multidentate (Runs 25–30) Ligands

Run	Catalyst[a]	Rate $\times 10^{7}$ [b]
1	Cu^{II}/1 PA	7.9
2	Cu^{II}/2 PA	0.21
3	Cu^{II}/1 AcAc	11.1
4	Cu^{II}/2 AcAc	1.8
5	Cu^{II}/1 Gly	4.7
6	Cu^{II}/2 Gly	1.9
7	Cu^{II}/1 Ala	5.9
8	Cu^{II}/2 Ala	3.2
9	Cu^{II}/1 Leu	8.8
10	Cu^{II}/2 Leu	2.6
11	Cu^{II}/1 DPA	1.5
12	Cu^{II}/2 DPA	0.014
13	Cu^{II}/1 IDA	2.9
14	Cu^{II}/2 IDA	0.17
15	Cu^{II}/1 EN	5.2
16	Cu^{II}/2 EN	0.10
17	Cu^{II}/1 TMED	6.4
18	Cu^{II}/2 TMED	5.9
19	Cu^{II}/4 TMED	3.9
20	Cu^{II}/1 DMP	10.2
21	Cu^{II}/2 DMP	0.0
22	Cu^{II}/1 BPA	35.1
23	Cu^{II}/2 BPA	23.4
24	Cu^{II}/5 BPA	12.6
25	Cu^{II}/1 TREN	0.0
26	Cu^{II}/1 TREN	0.0
27	Cu^{II}/1 IBPA	0.0
28	Cu^{II}/2 IBPA	0.0
29	Cu^{II}/1 TERPY	0.06
30	Cu^{II}/2 TERPY	0.05

[a] $[Copper(II)] = 2.8 \times 10^{-5}$ *M*; [hydrogen peroxide] $= 2.8 \times 10^{-3}$ *M*. All reaction solutions were buffered at pH 9.1.

[b] Rates are in molars per second.

Peroxide Activation by Copper(II) Complexes of Neutral Bidentate Ligands

Copper(II) complexes of the bidentate neutral ligands EN, TMED, DMP, and BPA were evaluated as possible catalysts. Table 2, runs 15–24, demonstrates that the copper(II) complexes of bidentate neutral ligands are very effective as catalysts for hydrogen peroxide activation. The copper(II) complex of 2,2′ -bipyridylamine has exceptional activity.

Rate variations with the copper(II) complexes of EN and TMED in the presence of excess ligand are interesting. The activity of the EN complex is reduced by a factor of 50 at a 2:1 ratio of ligand to metal. However, the activity of the TMED complex decreased only slightly in the presence of a fourfold excess of ligand. Sterically hindering ligands that decrease the equilibrium constant for binding a second bidentate ligand give copper catalysts that function in the presence of excess ligand. Formation constants for the 1:1 and 2:1 complexes of both EN and TMED show that the formation of the 2:1 complex is much less favorable with TMED (21).

Peroxide Activation with Copper(II) Complexes of Neutral Multidentate Ligands

To complete this study, several copper(II) complexes with neutral tri- and tetradentate ligands were evaluated as possible catalysts. Table 2, runs 25–30, shows the results with copper(II) complexes of TREN, IBPA, and TERPY. The conditions and peroxide concentrations were the same as in previous runs. Since the copper(II) complexes of tri- and tetradentate neutral ligands result in poor or inactive catalysts, we can conclude that having two adjacent coordination positions occupied with easily displaced solvent, and not the complex charge, is the necessary condition for active copper (II) catalysts.

Rate Studies

The following four complexes were selected for further kinetic examination. The copper(II)/leucine catalyst was selected because it gave the best performance for peroxide activation of the amino acid ligands. The copper(II)/acetylacetone complex gave the best performance with anionic ligands. The bipyridylamine copper complex gave the best overall peroxide activation. The copper(II)/*N*,*N*,*N*′,*N*′-tetramethylethylenediamine complex was selected to determine the influence of steric effects on the reaction rates.

Rates with all these ligands are first order in [Cu(II)] (Table 3). The copper(II)–leucine complex obeys Eq. [1]. However, rates with the copper(II) complexes of the other three ligands depended on [peroxide] (Table 4). The catalyst concentration is kept constant at 2.8×10^{-5} *M* for the copper(II)–AcAc and copper(II)–TMED complexes and, because of its higher activity, at 2.8×10^{-6} *M* for the copper(II)–BPA complex.

The Rate Law and Proposed Mechanism

A wide variety of copper(II) complexes activate hydrogen peroxide. Activities decrease dramatically when the ligand to metal ratio is greater than 1:1. Tri- and tetradentate ligand complexes are much less active than those with bidentate ligands. These results suggest that a significant concentration of copper(II) complexes with two adjacent coordinated aquo ligands is required for efficient hydrogen peroxide activation.

The mechanism of hydrogen peroxide activation by copper(II) must account for the following observations:

1. The order with respect to [quinaldine blue] is zero.
2. All the reactions are first order in [Cu(II)].

TABLE 3

Rates of Quinaldine Blue Oxidation with Various Concentrations of Copper(II) Complexes of Leu and AcAc

Run	Ligand	Complex concentration[a]	Rate $\times 10^{7}$ [b]
1	Leu	2.9×10^{-5}	8.8
2	Leu	2.3×10^{-5}	6.8
4	Leu	1.1×10^{-5}	4.4
5	Leu	5.7×10^{-6}	1.7
6	Leu	2.3×10^{-6}	0.67
7	AcAc	2.9×10^{-5}	11.1
8	AcAc	1.7×10^{-5}	8.5
9	AcAc	1.1×10^{-5}	5.9
10	AcAc	5.7×10^{-6}	2.2
11	AcAc	2.3×10^{-6}	0.65

Note. [Peroxide] = 2.8×10^{-2} *M*.
[a] Concentrations are in moles per liter.
[b] Rates are in units of molars per second.

TABLE 4

Rates of Quinaldine Blue Oxidation at Various Peroxide Concentrations

Run	Ligand	[Peroxide][a]	Rate $\times 10^{7}$ [b]
1	BPA	2.86×10^{-2}	3.66
2	BPA	9.16×10^{-3}	2.29
3	BPA	4.66×10^{-3}	1.55
4	BPA	5.73×10^{-4}	0.226
5	AcAc	2.86×10^{-2}	11.2
6	AcAc	9.16×10^{-3}	11.1
7	AcAc	4.66×10^{-3}	9.8
8	AcAc	1.15×10^{-3}	8.9
9	AcAc	5.73×10^{-4}	5.7
10	AcAc	2.86×10^{-4}	4.6
11	TMED	2.86×10^{-2}	5.8
12	TMED	2.86×10^{-2}	5.0
13	TMED	9.16×10^{-3}	4.5
14	TMED	9.16×10^{-3}	4.4
15	TMED	4.66×10^{-3}	3.0
16	TMED	4.66×10^{-3}	3.2
17	TMED	1.15×10^{-3}	1.6
18	TMED	5.73×10^{-4}	1.0
19	TMED	5.73×10^{-4}	0.97
20	TMED	2.86×10^{-4}	0.40
21	TMED	2.86×10^{-4}	0.383

Note. [Cu(II)] =
[a] Concentrations are in molars.
[b] Rate units are in molars per second.

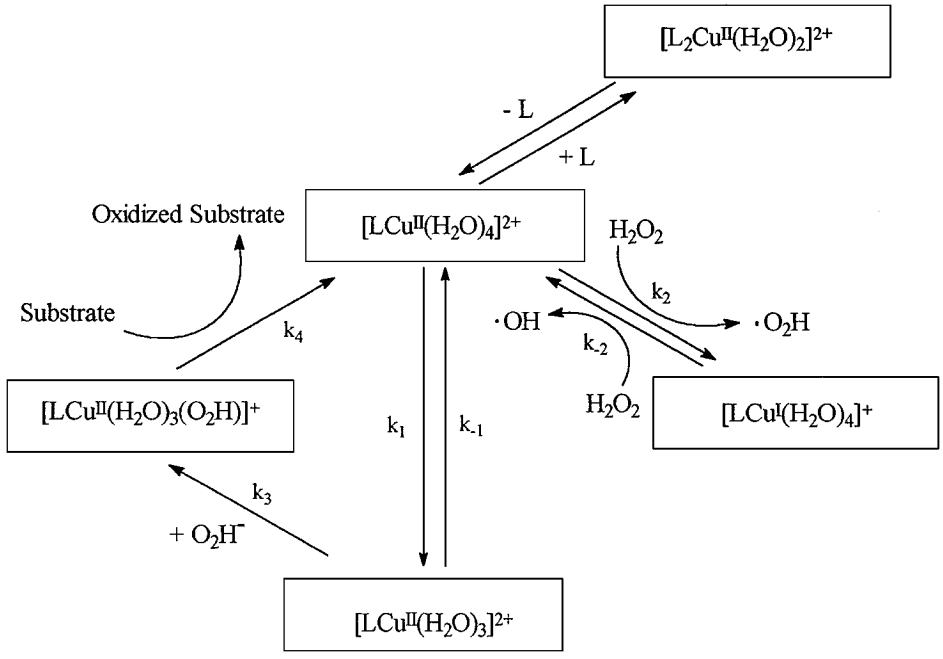


FIG. 4. The proposed mechanism.

3. Addition of ligand above a 1:1 mole ratio decreases activity. Tri- or tetradentate ligands tend to give inactive complexes.

4. The reaction rate is pH dependent.

5. The reaction rate is independent of $[H_2O_2]$ with PA and Leu ligands, but increases with $[H_2O_2]$ in a nonlinear manner with AcAc, TMED, and BPA ligands.

A mechanism consistent with these observations is shown in Fig. 4.

The first step is the reversible dissociation of a water molecule from a distorted octahedral, copper(II) complex. This process is reported to be reasonably easy (22). The next step is coordination of the hydroperoxide anion (O_2H^-) to produce the peroxo or hydroperoxo complex. Once formed, this copper–hydroperoxide species reacts extremely rapidly with the substrate. As a result, the rate of oxidation is determined by the rate of formation of the copper–hydroperoxide complex, giving the zero-order substrate dependence. While dimeric copper complexes cannot be ruled out, their presence is unlikely at the low copper concentrations employed.

Assuming that the rate of formation of the copper–hydroperoxide complex is rate limiting and is followed by rapid oxidation of quinaldine blue, the rate of oxidation catalyzed by the copper(II) complex is given by

$$\begin{aligned} \nu_{ox} &= d[LCu^{II}(H_2O)_3(O_2H)]/dt \\ &= k_3[LCu^{II}(H_2O)_3][O_2H^-]. \end{aligned} \quad [1]$$

A steady-state approximation for the $[LCu^{II}(H_2O)_3]$ concentration leads to the expression

$$[LCu^{II}(H_2O)_3] = \frac{k_1[LCu^{II}(H_2O)_4]}{k_{-1} + k_3[O_2H^-]}. \quad [2]$$

The concentration of the hydroperoxide anion is directly proportional to peroxide concentration and pH; $[O_2H^-] = K_a[H^+][H_2O_2]$. Substituting Eq. [2] into Eq. [1] gives the rate of oxidation as

$$\text{Reaction rate} = \frac{k_1k_3[LCu^{II}(H_2O)_4][O_2H^-]}{k_{-1} + k_3[O_2H^-]}. \quad [3]$$

Equation (3) can be rearranged to

$$\text{Reaction rate} = \frac{k_1[LCu^{II}(H_2O)_4][O_2H^-]}{(k_{-1}/k_3) + [O_2H^-]}. \quad [4]$$

In addition, if we derived the rate law by using two consecutive steady-state approximations for both $[LCu^{II}(H_2O)_3]^{2+}$ and $[LCu^{II}(H_2O)_3(O_2H)]^+$, we obtain the same rate law as in Eq. [4]. The only requirement for this to be valid is the rate of substrate oxidation needs to be greater than the rate of copper(II) hydroperoxide complex formation.

Experimental observations support the above mechanism and suggest that other possibilities are less reasonable. The active species is not a dissociated hydroxyl radical produced by peroxide decomposition. While this pathway exists, as shown in the mechanism (the k_2 pathway), the reaction order in peroxide would not be zero if it were the only reaction. Furthermore, this reaction does not account

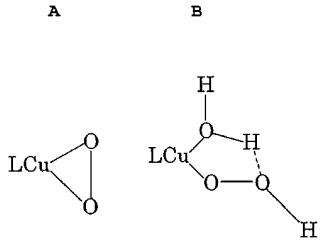


FIG. 5. Proposed structures for: (A) a μ^2 bound peroxide and (B) a hydrogen-bond-stabilized copper–hydroperoxide complex.

for the dramatic decrease in catalytic reactivity for complexes of tri- and tetradentate ligands.

Another possible mechanism involves a high oxidation state of copper as the active oxidant. While copper(II) complexes have been shown to oxidize amine-containing buffers (19, 20), no oxidation of quinaldine blue by Cu(II) occurs in the absence of hydrogen peroxide. Copper(III) species are known in aqueous solutions (23–26). However, a Class V reaction mechanism (6) (which involves a metal center changing its oxidation state in substrate oxidant reactions) involving copper(III) redox cycling cannot explain the need for two available equatorial binding sites.

Dissociation of water occurs from the axial positions. Once the five-coordinate copper complex is formed, the attack by O_2H^- occurs to coordinate the O_2H^- in the strong binding equatorial position (Fig. 5). The strongest binding of hydroperoxide anion occurs when it is bound to the same $d_{x2\text{-}y2}$ orbital that is involved in ligand coordination. Coordination of O_2H^- to the equatorial position is important because the use of this orbital leads to a stronger acceptor than the axial position. Strong Lewis acidity activates hydrogen peroxide in Class IVb mechanisms. When the 2 : 1 ligand to copper(II) complex is formed, the reactivity decreases because there are no strong-binding equatorial positions available to bind the hydroperoxide anion.

Our experiments show that it is necessary to have two equatorial coordination sites occupied by easily displaced solvent. This is demonstrated by the large decrease in reactivity seen with complexes of tri- and tetradentate ligands, compared to the bidentate ligands. The most straightforward explanation requires two coordination sites to bind the peroxide (Fig. 5A). Complexes of copper(II) dimers have been shown to bind O_2^{2-} in a similar manner (9c, 11b–11e, 12b, 13a). Peroxo complexes of other metals are known (1b) and several are useful oxidants and oxidation catalysts (1c).

Another possibility is that the metal–peroxo complex is stabilized by a hydrogen bonding interaction between bound hydroperoxide and an equatorial water. A similar interaction has been used to explain the stability of a vanadium–peroxo complex in aqueous solution (27). This interaction can only occur when one bidentate ligand is coordinated to the copper center, as shown in Fig. 5B. When a tri- or tetradentate ligand is coordinated to the copper(II) complex, or a 2 : 1 complex of the bidentate ligand is formed, stabilization of the hydroperoxide anion through hydrogen bonding does not occur.

Our proposed mechanism leads to a rate law, Eq. [5], that explains our kinetic data. The requirement that our mechanism gives zero-order kinetics for quinaldine blue oxidation is satisfied with the assumption that the copper(II) hydroperoxide species will react very rapidly with quinaldine blue. This is reasonable, as Oxone and sodium hypochlorite oxidize quinaldine blue too rapidly to study with the procedures used here.

The rate of oxidation is predicted, by the derived rate law, to be first order in copper(II) complex. This was experimentally observed. A zero-order dependence of the rate on peroxide concentration was observed for copper(II) complexes of PA and Leu. However, the copper(II) complexes of AcAc, TMED, and BPA all showed significant deviations from zero-order kinetics.

While these results differ from the zero-order dependence seen for copper(II) complexes of picolinic acid or leucine, the rate law derived from the proposed mechanism (Eq. [4]) can be used to explain these differing orders. If we define V_{max} as $k_1[LCu^{II}(H_2O)_2]$ and k_m as k_{-1}/k_3 we obtain the expression

$$\text{Reaction rate} = \frac{V_{max}[O_2H^-]}{K_m + [O_2H^-]}. \qquad [5]$$

Equation [5] is now in the same form as the Michaelis–Menten equation (28). Systems having kinetic behavior described by Eq. [5] give linear plots of $(\text{rate})^{-1}$ versus $[HO_2^-]^{-1}$. Figure 6 shows this plot for the Cu–AcAc system. The plots for Cu–TMED and Cu–BPA are also linear. Thus, these systems follow the kinetics predicted by the derived rate law (Eq. [5]). From these plots we can determine K_m/V_{max} and $(V_{max})^{-1}$ whose values are given in

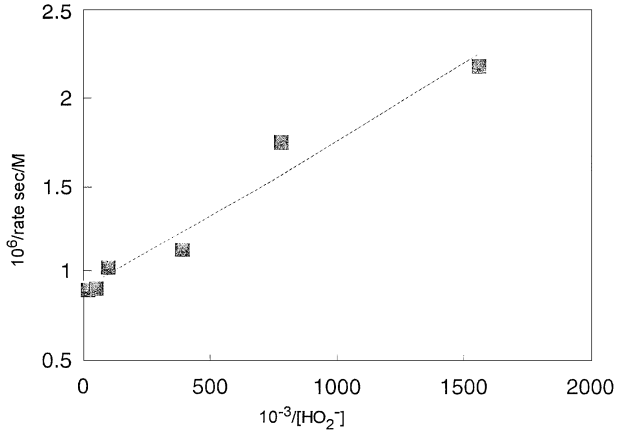


FIG. 6. Rate^{-1} vs $[HO_2^-]^{-1}$ for the Cu–AcAc system.

TABLE 5

Values for V_{max} and K_m as Determined for the Copper(II) Complexes of Various Ligands

Ligand	V_{max}[a]	K_m
PA	7.9×10^{-7}	$<8 \times 10^{-7}$
Leu	8.8×10^{-7}	$<8 \times 10^{-7}$
AcAc	1.1×10^{-6}	9.4×10^{-7}
TMED	9.1×10^{-7}	1.4×10^{-5}
BPA	8.1×10^{-6}	4.8×10^{-5}

[a] V_{max} is expressed for $[Cu(II)] = 2.8 \times 10^{-5}$ *M*.

Table 5. For the copper(II) complexes of picolinic acid and leucine, V_{max} is the observed rate of reaction, because peroxide concentrations low enough to cause deviation from zero order kinetics were not examined. K_m must clearly be lower than $[HO_2^-]$ to give zero-order kinetics in peroxide (see Eq. [5]. Thus K_m values for the picolinic acid and leucine complexes are estimated based on the lowest hydroperoxide anion concentration studied.

In Table 5, the BPA and TMED complexes are positively charged while the AcAc, Leu, and PA complexes have unit positive charge. Interestingly, the di-positively charged copper(II) complexes (neutral ligands) have significantly higher values of K_m for peroxide activation. K_m is equal to k_{-1}/k_3 in the derived rate law, which implies that coordination of water to the dissociated copper(II) complex ($[LCu(H_2O)]$), rather than coordination of the hydroperoxide anion, is more favorable for the more positively charged copper complexes. However, in view of the complex enthalpic and entropic contributions to the coordination of these species, we cannot provide a more detailed explanation of the trends observed in K_m or V_{max}.

Since the rates of oxidation are dependent upon the concentration of the hydroperoxide anion, rather than hydrogen peroxide, changes in pH affect reaction rates. For the copper(II)–picolic acid complex at high pH, the concentration of hydroperoxide anion will be high enough to give zero-order kinetics, where the observed rate is equal to V_{max}. As the pH drops below 9, the concentration of hydroperoxide anion decreases enough to see deviation from zero-order kinetics similar to that seen for the copper(II) complexes of AcAc, TMED, and BPA. This gives rise to a dependence of rate on hydroperoxide anion concentration. If the concentration of $[HO_2^-]$ is calculated for the data in Fig. 3 (in the range from pH 9.7 to 7.3) and a $(\text{rate})^{-1}$ versus $[HO_2^-]$ plot constructed, linear behavior is observed supporting the above conclusion.

CONCLUSIONS

We have found a series of copper(II) complexes that activate hydrogen peroxide for oxidation in basic aqueous solution. In all cases, increasing the amount of ligand above a 1:1 ratio to copper(II) decreases the rate of peroxide activation, and in some cases this occurred dramatically. Complexes of tri- and tetradentate ligands are generally inactive. A mechanism (Fig. 4) is proposed which involves the dissociation of coordinated water from an axial position followed by nucleophilic attack by a hydroperoxide anion, to form an equatorially bound copper(II)–hydroperoxide complex. The more strongly coordinating equatorial position provides the needed Lewis acidity to activate bound hydroperoxide for oxidations. Several theories are examined on why two equatorial coordination sites containing easily displaced solvent are necessary for peroxide activation with these complexes. One possibility is that the peroxide occupies two coordination sites, as shown in Fig. 5A. Another possibility is that the coordinated hydroperoxide is stabilized by hydrogen-bonding interaction with a coordinated water, also shown in Fig. 5A. A first-order dependence of rate on catalyst concentration is observed, which is shown to agree with the rate law derived from our mechanism (Eq. [4]). The observed zero-order dependence of oxidation rate on substrate concentration is also explained within the context of the proposed mechanism.

The observation of zero-order dependence of rate on hydrogen peroxide concentration for PA and Leu and the nonzero-order rate dependence for AcAc, TMED, and BPA are also explained. This is done by rewriting Eq. [4], the derived rate law, in the form of the Menten–Michaelis expression. This expression predicts under which conditions the rate of reaction is equal to V_{max}, where the zero-order rate dependence on peroxide concentration occurs. This also explains the observed dependence of reaction rate upon pH for oxidations catalyzed by the copper(II)–picolinic acid complex. As the pH decreases the concentration of hydroperoxide anion also is decreased. When the pH drops below that needed for sufficient concentration of hydroperoxide anion to give the observed rate as V_{max}, a decreased rate is observed.

The mechanism offered for these systems can be generalized. In order to activate hydrogen peroxide by coordination of the peroxide to a metal center, the complex used must be capable of binding O_2H^{-1} to a strongly acidic coordination position. The stability constants for successive ligand binding can be employed to select metal complexes that have the potential to bind and activate O_2H^{-1}. The sequence of formation constants must allow the existence of aquo complexes with which O_2H^{-1} can react by displacing water from strong binding metal orbitals. For copper(II), the 2:1 ($Cu(L)_2$) formation constant (when L is bidentate) must be low enough relative to 1:1 (Cu(L)) to permit substantial quantities of the 1:1 complex to exist in solution. For metal ions with fewer *d*-electrons, strong axial coordination positions exist. Bidentate ligand complexes, ML, strongly bind O_2H^{-1} and activate peroxide.

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