The Copper-catalysed Oxidation of Ascorbic Acid by Dioxygen. Part 4.[1] The Effect of Chloride Ions on the Kinetics and Mechanism

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The effect of chloride ions on the kinetics and mechanism of the copper-catalysed oxidation of ascorbic acid by dissolved oxygen has been examined over the pH range 2.0–3.5 and chloride concentrations up to 0.2 mol dm^{-3}. The empirical rate law has been established as $-d[O_2]/dt = k_{obs.}[O_2]^{\frac{1}{2}}$ over the whole range. At constant total ascorbate, total copper, and hydrogen ion concentrations, the general relationship (1) was obtained. The constants

$$k_{obs.} = \frac{A + B[Cl^-]}{C + D[Cl^-] + E[Cl^-]^2} \qquad (1)$$

A to *E* were evaluated using data obtained in pure nitrate media to fix *A* and *C*, and the dependence of all the constants on pH, total copper concentration, and total ascorbate concentrations established experimentally. It is also pointed out that a mechanism based on the involvement of copper(I) rather than the copper(III) invoked by us cannot be entirely ruled out.

THE kinetics and mechanism of the copper(II)-catalysed oxidation of ascorbic acid H_2Asc by dissolved dioxygen has been investigated in the presence of chloride ions. The chloride concentration was varied from zero to 0.2 mol dm^{-3} and the ionic strength maintained constant up to 0.100 mol dm^{-3} by addition of the required amount of $K[NO_3]$, but then allowed to increase above this value so as to accommodate some runs of higher chloride concentration. An O_2-sensitive electrode was used to follow the reaction, and the pH maintained with a pH-stat.

The results show a half-order dependence on $[O_2]$, Figure 1, and lower than first-order dependence on total copper, $[Cu]_T$, Figure 2. There is a complicated dependence on total ascorbate, $[L]_T$, the rate becoming essentially independent of $[L]_T$ at high chloride concentration. This variation with $[Cl^-]$ is particularly inforative and is illustrated in Figures 3 and 4. The curves representing $k_{obs.}$ *versus* $[Cl^-]$, where $-d[O_2]/dt = k_{obs.}[O_2]^{\frac{1}{2}}$, were readily fitted to the general equation (1);

$$k_{obs.} = \frac{A + B[Cl^-]}{C + D[Cl^-] + E[Cl^-]^2} \qquad (1)$$

Table 1 shows the values obtained for a typical series of runs at constant $[Cu]_T$ and constant pH, but with varying $[L]_T$. Table 2 illustrates the effect of varying the pH whilst $[Cu]_T$ and $[L]_T$ are maintained constant. Note that this evaluation of *A*, *B*, *C*, *D*, and *E* was made possible by assuming that when $[Cl^-] = 0$, equation (1) should correspond to the results obtained in pure nitrate media (0.1 mol dm^{-3} $K[NO_3]$) and thus *A* and *C* were evaluated on this basis (the scheme of Jameson and Blackburn[2]) by putting $A = k_1'[Cu]_T[L]_T$ and $C = (1 + K_2^H[H^+] + K^M[L]_T + K^{MH}[L]_T[H^+])$ which is justified below.

It can be seen (from Table 1) that both *B* and *E* are directly proportional to $[L]_T^{\frac{1}{2}}$ and practically independent of $[H^+]$ (Table 2), whereas *D* is independent of $[L]_T$ but dependent on $[H^+]$ in a way that is readily explained (see below).

TABLE 1

Variation of coefficients with ascorbic acid concentration at pH 2.70; $[Cu]_T = 3.151 \times 10^{-5}$ mol dm^{-3}

$[L]_T$	$[L]_T^{\frac{1}{2}}$	10^2A *	B	C *	D	$10^{-4}E$	$D/(1 + K_2^H[H^+])$
0.010	0.100	4.06	4.30	28.0	58	1.19	2.03
0.017	0.130	6.91	5.80	30.0	53	1.61	1.85
0.030	0.173	12.22	7.59	33.9	60	2.11	2.10
0.040	0.200	16.34	8.92	36.7	54	2.48	1.89
0.050	0.224	20.33	10.10	38.3	56	2.81	1.96

* Calculated on the basis of nitrate-only studies.

TABLE 2

Variation of coefficients with $[H^+]$; $[L]_T = 0.017$ mol dm^{-3}, $[Cu]_T = 3.151 \times 10^{-5}$ mol dm^{-3}

pH	10^2A *	B	C *	D	$10^{-4}E$	$D/(1 + K_2^H[H^+])$
3.35	6.82	6.13	11.4	14	1.70	1.95
3.00	6.91	5.90	21.0	28	1.63	1.89
2.70	6.91	5.80	30.0	53	1.61	1.85
2.35	6.88	5.75	54.3	106	1.60	1.69
2.04	6.85	5.72	97.3	242	1.58	1.91

* Calculated on the basis of nitrate-only studies.

We explain these rate data as follows [note that we maintain the use of Cu^{2+} as the initial oxidation state of the copper in the reaction in agreement with our earlier work[1,2] but in contradistinction to the ideas put forward then, we now believe that these results cannot of themselves completely rule out the involvement of a copper(I) intermediate (in general agreement with Shtamm and Skurlatov[3]). This is discussed below].

The species that we propose to be responsible for the additional reactivity in chloride media is $[Cu_2LCl_2]$, where L^{2-} is the ascorbate anion, and that it remains in low concentration. The inhibiting effect of chloride at high chloride concentration we ascribe to the tetrachloro-species $[Cu_2LCl_4]^{2-}$. The equilibria (2)—(7) are required to explain the observations. At the pH's

$$L^{2-} + H^+ \rightleftharpoons HL^- \qquad K_1^H = [HL^-]/[L^{2-}][H^+] \qquad (2)$$

$$HL^- + H^+ \rightleftharpoons H_2L \qquad K_2^H = [H_2L]/[HL^-][H^+] \qquad (3)$$

$$Cu^{2+} + HL^- \rightleftharpoons [Cu(HL)]^+$$
$$K^M = [Cu(HL)^+]/[Cu^{2+}][HL^-] \qquad (4)$$

$$H^+ + Cu^{2+} + HL^- \rightleftharpoons [Cu(H_2L)]^{2+}$$
$$K^{MH} = [Cu(H_2L)^{2+}]/[H^+][Cu^{2+}][HL^-] \qquad (5)$$

$$Cu^{2+} + Cl^- \rightleftharpoons [CuCl]^+$$
$$K^{Cl} = [CuCl^+]/[Cu^{2+}][Cl^-] \quad (6)$$

$$2Cu^{2+} + L^{2-} + 4Cl^- \rightleftharpoons [Cu_2LCl_4]^{2-}$$
$$\beta = [Cu^{2+}]^2[L^{2-}][Cl^-]^4/[Cu_2LCl_4{}^{2-}] \quad (7)$$

concerned, the total ascorbate is given by equation (8) because the concentration of copper-containing species is negligible with respect to $[L]_T$. This leads to equation (9). The total copper present can be taken as shown in equations (10) and (11). This quadratic in $[Cu^{2+}]$ can be written to a high degree of accuracy as equation (12).

$$[L]_T = [HL^-] + [H_2L] \quad (8)$$

$$[HL^-] = [L]_T/(1 + K_2{}^H[H^+]) \quad (9)$$

$$[Cu]_T = [Cu^{2+}] + [Cu(HL)^+] + [Cu(H_2L)^{2+}] + [CuCl^+] + 2[Cu_2LCl_4{}^{2-}] \quad (10)$$

$$= [Cu^{2+}](1 + K^M[HL^-] + K^{MH}[HL^-][H^+] + K^{Cl}[Cl^-]) + [Cu^{2+}]^2(2[HL^-][Cl^-]^4/K_1{}^H[H^+]) \quad (11)$$

$$[Cu^{2+}] = [Cu]_T/\{1 + K^M[HL^-] + K^{MH}[HL^-][H^+] + K^{Cl}[Cl^-] + [Cu]_T{}^{\frac{1}{2}}\sqrt{2}[HL^-]^{\frac{1}{2}}[Cl^-]^2/(K_1{}^H[H^+])^{\frac{1}{2}}\} \quad (12)$$

We now, following the ideas expressed in Part 2,[2] and extended below, assume that the overall reaction can be expressed as shown in equations (13) and (14). Making use of equations (9) and (12), this leads to equation (15).

$$-d[O_2]/dt = k_1'[Cu^{2+}][HL^-][O_2]^{\frac{1}{2}} + k_2'[Cu^{2+}][L^{2-}]^{\frac{1}{2}}[Cl^-][O_2]^{\frac{1}{2}} \quad (13)$$

$$i.e.,\ k_{obs.} = k_1'[Cu^{2+}][HL^-] + k_2'[Cu^{2+}][L^{2-}]^{\frac{1}{2}}[Cl^-] \quad (14)$$

$$k_{obs.} = \frac{k_1'[Cu]_T[L]_T + \{k_2'[Cu]_T[L]_T{}^{\frac{1}{2}}(1 + K_2{}^H[H^+])^{\frac{1}{2}}/(K_1{}^H[H^+])^{\frac{1}{2}}\}[Cl^-]}{\{1 + K_2{}^H[H^+] + K^M[L]_T + K^{MH}[L]_T[H^+]\} + \{(1 + K_2{}^H[H^+])K^{Cl}\}[Cl^-] + \{2[Cu]_T[L]_T(1 + K_2{}^H[H^+])/K_1{}^H[H^+]\}^{\frac{1}{2}}[Cl^-]^2} \quad (15)$$

Now, equation (15) is clearly of the form required [equation (1)] and furthermore, *B* and *E* are virtually independent of $[H^+]$ as required and *A* and *C* are indeed the quantities calculable from the nitrate data. But the outstanding feature is the fact that *D* should be equal to $K^{Cl}(1 + K_2{}^H[H^+])$: in Tables 1 and 2, *D* has been divided by $(1 + K_2{}^H[H^+])$ to give a satisfyingly constant number. If this is taken to be the stability constant for the copper(II) monochloro-complex, then the eminently reasonable value of log K^{Cl} (below) is obtained

$$\log K^{Cl} = 0.28 \quad (16)$$

for the mean. Furthermore, $\beta_2{}^{Cl}$ for any dichloro-complex of copper would appear in the intercept of a plot of *E versus* $[L]_T{}^{\frac{1}{2}}$; this is in fact virtually zero. Finally, any such intercept would in any case have to be below 10^4 at pH = 2.7, leading to a $\beta_2{}^{Cl}$ value of below 10^3. This in itself rules out the presence of any more than trace concentrations of $[Cu^ICl_2]^-$ for which β_2 is reliably quoted [4] as 10^6.

Equation (15) finally allows a value for k_2' to be obtained (4.04×10^9 dm^6 mol^{-2} min^{-1}).

DISCUSSION

There is considerable evidence that under anaerobic conditions any intramolecular change transfer within the copper–ascorbate binary complex is only partial and that the complexes are best described as containing formally copper(II) ions. Thus, (*i*) the anaerobic copper(II)–ascorbate system titrates normally with base giving a series of parallel formation curves as the ratio of copper(II) to ascorbate is varied.[2] Invariant stability constants for the copper(II)–ascorbate complexes $[Cu(H_2L)]^{2+}$, $[Cu(HL)]^+$, $[Cu_2(HL)_2]^{2+}$, and $[Cu_2L_2]$ have been calculated from these data.[2] (*ii*) A green solution with a spectrum consistent with a copper(II) species is formed around pH 3.5. This spectrum exhibits *d–d* bands with $\lambda_{max.}$ = 410 and 680 nm reminiscent of those found in the copper(II) complexes of both substituted and unsubstituted catechols.[2] (*iii*) As the pH is raised above 3.5, copper metal is deposited indicating that any copper(I) complexes formed would be unstable in acid solution with respect to disproportionation. (*iv*) In chloride solutions the chloro-complexes present do not include $[Cu^ICl_2]^-$. (*v*) Below about pH 5, anaerobic copper(II)–ascorbic mixtures show e.s.r. signals [5] whose intensity reflects nearly 100% of total copper at pH 2.5, decreasing to 20% total copper at pH 5.0. Even at pH 3.0 (where oxidation is rapid) over 90% of the copper is e.s.r.-detectable. Note that the decrease in observed signal as the pH is increased is readily explicable on the basis of exchange coupling within the copper(II) dimers present. The e.s.r. spectrum of copper(II)–ascorbate

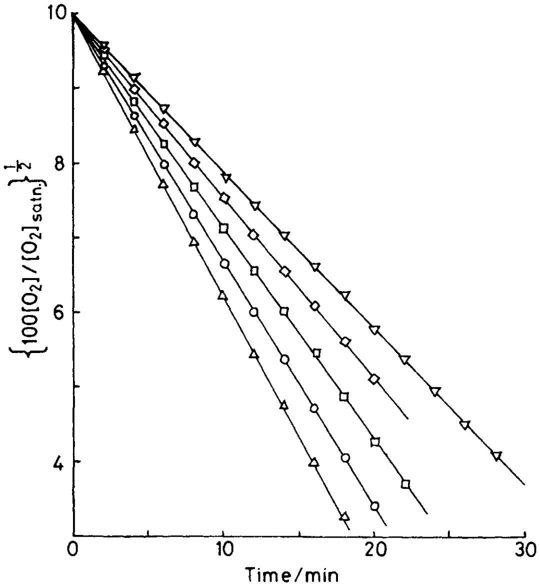


FIGURE 1 Typical half-order plots. $10^2[L]_T$ = 4.53—6.68 mol dm^{-3}; $10^5[Cu]_T$ = 1.71 mol dm^{-3}; at pH 2.20 (△), 2.60 (○), 2.89 (□), 3.15 (◇), 3.35 (▽); $[Cl^-]$ = 0.100 mol dm^{-3}

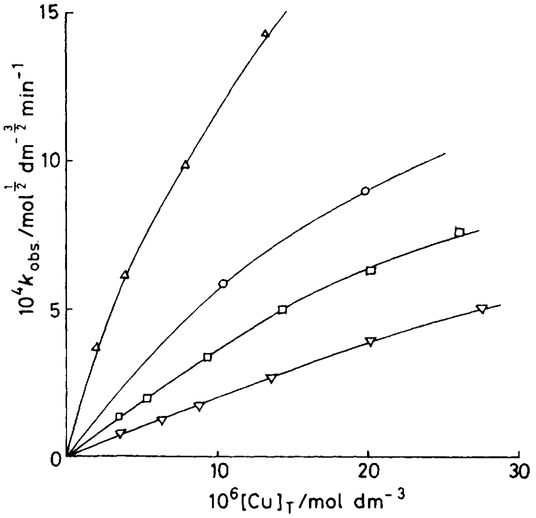


FIGURE 2 Illustration of dependence of $k_{obs.}$ on total copper concentration, $[Cu]_T$. At pH 2.42, $[Cl^-] = 0.05$ (△) 0.10 (○); pH 3.05, $[Cl^-] = 0.05$ (□) 0.10 mol dm^{-3} (▽)

complexes has also been reported by Hayakawa and Hayashi.[6]

These observations would furthermore seem to add credence to our previously published mechanism involving a formally higher oxidation state [*i.e.*, copper(III)] as the intermediate. Indeed this would seem to be reinforced by a recent reinvestigation of the iron(II)- and iron(III)-catalysed oxidation of cysteine.[7] The order with respect to oxygen is found to be unity and a half for iron(II) and iron(III) respectively. This is a direct analogy with the copper–ascorbate system in that the iron(III)-catalysed reaction can be ascribed to an initial two-electron transfer to O_2 *via* dimeric species involving formally iron(IV) species. Any build-up of iron(II) species would lead to kinetics identical to those found for the iron(II)-catalysed reaction. Furthermore, the reaction with iron(II) as the catalyst is considerably slower than when iron(III) is used and this is consistent with the idea that iron(II)–dioxygen complexes are thermodynamically more stable than those of iron(III).[8] Thus the transient nature of the oxygen adducts with the higher oxidation states of the metal ion, copper(II) and iron(III), would of itself lead to more rapid reaction in each case.

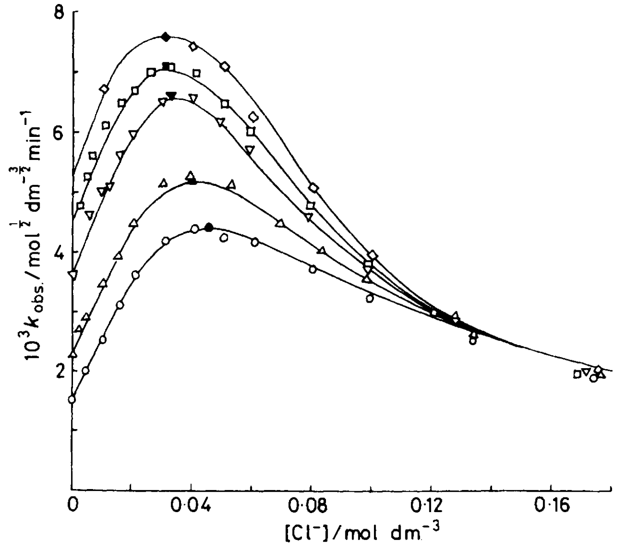


FIGURE 3 Dependence of rate of reaction on chloride concentration: variation with total initial ascorbate concentration, $[L]_T$. $[L]_T = 0.010$ (○), 0.017 (△), 0.030 (▽), 0.040 (□), 0.051 mol dm^{-3} (◇); at pH 2.70 and $[Cu]_T = 3.151 \times 10^{-5}$ mol dm^{-3}. Filled symbols are computed maxima using coefficients from Table 1. Curves are computed with experimental points

However, an intriguing feature of the copper(II)–ascorbate–chloride system is why the dichloro-species should be catalytically active (and decompose) whereas the tetrachloro-species should be inert (and build up) with increasing chloride concentration. A reasonable explanation can be inferred from the behaviour of the copper(II)–bipyridyl–ascorbate ternary system.[9] The complex $[Cu^{II}(bipy)(Asc)]$ (bipy = 2,2′-bipyridyl), which can be prepared under anaerobic conditions, is resistant to intermolecular electron transfer at low pH and exhibits a spectrum with $\lambda_{max.} = 420$ nm identical to that reported by us for the copper(II)–ascorbate system in nitrate medium.[2] In considerable contrast, the reaction of $[Cu^{II}(bipy)_2]$ with ascorbate leads to a rapid reduction of the bis(bipyridyl)copper complex (and a concomitant

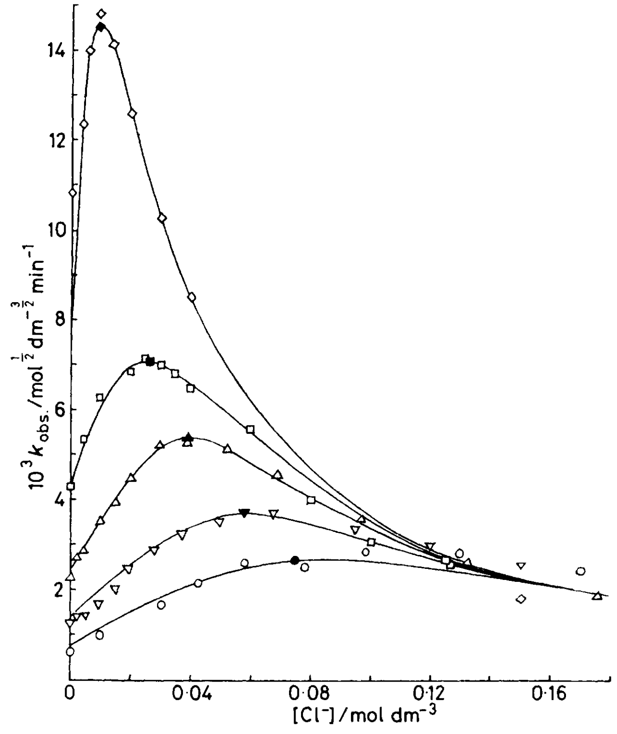
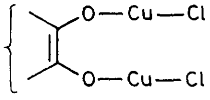


FIGURE 4 Dependence of reaction on chloride concentration: variation with pH. $[L]_T = 0.017$ mol dm^{-3}, $[Cu]_T = 3.151 \times 10^{-5}$ mol dm^{-3}; at pH 2.00 (○), 2.35 (▽), 2.70 (△), 3.00 (□), 3.35 (◇). Filled symbols are computed maxima using coefficients from Table 2. Curves are computed with experimental points

change of colour from green to brown) which is second-order in total copper concentration. This difference in behaviour is ascribed [9] to differences in the structure of the ternary complexes: the redox-stable $Cu^{II}(bipy)(Asc)$ system is mononuclear with bidentate ascorbate co-ordination whilst the reactive $Cu^{II}(bipy)_2(Asc)$ species is binuclear with an ascorbate bridge between two $Cu^{II}(bipy)_2$ units.

Similar considerations would go a long way to explaining the relative reactivity of the ternary copper(II)–ascorbate–chloride species, in that the reactive dimer might be expected to contain an ascorbate *bridge* whereas the unreactive tetrachloro-species would most likely not be bridged in this fashion.



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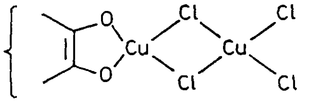
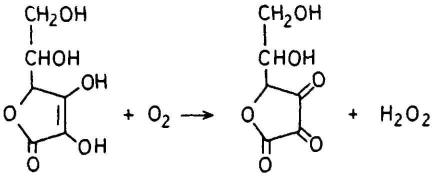


(I) (II)

The bridging ascorbate in (I) would, with or without an accompanying chloride bridge, confer an obvious advantage with respect to the accessibility of dioxygen binding.

Finally, it is almost certainly relevant to note that Skurlatov [10] found the postulation of copper(III) necessary as an intermediate in the oxidation of copper–ascorbate complexes with hydrogen peroxide.

Despite the strength of this evidence, it is only right to point out that the kinetic evidence does not *of itself* rule out the involvement of copper(I) as the intermediate in these reactions and we would like to put forward the following scheme, which although we feel still leaves questions unanswered, is nevertheless an important contribution to this work. The stoicheiometry of the reaction semains, of course, as it was (see below).



Ascorbic acid (H_2L) Dehydroascorbic acid (L′)

(III) (IV)

(*i*) *Absence of Chloride.*—We assume, as above, that the dimer is the reactive species, but do not consider it to react directly with molecular oxygen, but to break unsymmetrically, equation (17).

$$2[Cu(HL)]^+ \overset{K^{DH}}{\rightleftharpoons} [Cu_2(HL)_2]^{2+} \qquad (17)$$

Initiation

$$[Cu_2(HL)_2]^{2+} \xrightarrow{k_1} [Cu^I(HL)] + Cu^I + L' + H^+ \qquad (18)$$

$$Cu^I + HL^- \xrightarrow{K^{M'}} [Cu^I(HL)] \qquad (19)$$

Chain $$[Cu^I(HL)] + O_2 \xrightarrow{k_2} [Cu^I(HL)O_2] \qquad (20)$$

$$[Cu^I(HL)O_2] \xrightarrow{k_3} Cu^I + HO_2^- + L' \qquad (21)$$

Termination (resulting in re-formation of dimer)

$$[Cu^I(HL)O_2] + [Cu^I(HL)] \xrightarrow{k_4} [Cu_2(HL)_2]^{2+} + O_2^{2-}\ (O_2^{2-} + 2H^+ \xrightarrow{\text{fast}} H_2O_2) \qquad (22)$$

Steady-state treatment

$$(a)\ d[Cu^I(HL)O_2]/dt = k_2[Cu^I(HL)][O_2] - k_3[Cu^I(HL)O_2] - k_4[Cu^I(HL)O_2][Cu^I(HL)] = 0 \qquad (23)$$

$$(b)\ d[Cu^I(HL)]/dt = 2k_1[Cu_2(HL)_2{}^{2+}] - k_2[Cu^I(HL)][O_2] + k_3[Cu^I(HL)O_2] - k_4[Cu^I(HL)O_2][Cu^I(HL)] = 0 \qquad (24)$$

Since equation (22) is the termination step, we may legitimately assume that $k_3[Cu^I(HL)O_2] >> k_4[Cu^I(HL)O_2][Cu^I(HL)]$. Hence from equations (23) and (24) we obtain equations (25) and (26) respectively. However, since $-d[O_2]/dt = k_2[Cu^I(HL)][O_2]$, substituting for $[Cu^I(HL)]$ obtained from equations (25) and (26) yields the expressions (27)—(29) which is the required rate law arrived at by analysis of the data obtained in pure nitrate media.[2]

$$k_2[Cu^I(HL)][O_2] = k_3[Cu^I(HL)O_2] \qquad (25)$$

$$2k_1[Cu_2(HL)_2{}^{2+}] = k_4[Cu^I(HL)O_2][Cu^I(HL)] \qquad (26)$$

$$-d[O_2]/dt = (2k_1k_2k_3/k_4)^{\frac{1}{2}}[Cu_2(HL)_2{}^{2+}]^{\frac{1}{2}}[O_2]^{\frac{1}{2}} \qquad (27)$$

$$= (2k_1k_2k_3/k_4)^{\frac{1}{2}}(K^{MH}K^{D})^{\frac{1}{2}}[Cu^{2+}][HL^-][O_2]^{\frac{1}{2}} \qquad (28)$$

$$= k_1'[Cu^{2+}][HL^-][O_2]^{\frac{1}{2}} \qquad (29)$$

(*ii*) *Presence of Chloride.*—We start by requiring the reactant species (present in very low concentration) to be $[Cu_2LCl_2]$, reaction (30). *Propagation* and *Termination* steps as for the nitrate-only case. Since k_5, unlike k_1, now involves only one L^{2-} and two Cl^-, the final result for chloride species only is easily seen to be of the form of equation (31) which has been used in equation (14) above.

Initiation

$$[Cu_2LCl_2] \xrightarrow{k_5} Cu^+ + [Cu^ICl_2] + L' \qquad (30)$$

$$-d[O_2]/dt = k_2'[Cu^{2+}][L^{2-}]^{\frac{1}{2}}[Cl^-][O_2]^{\frac{1}{2}} \qquad (31)$$

The acceleration due to chloride is then due to the greater efficiency of reaction (30) over (18). The build-up of a tetrachloro-complex $[Cu_2LCl_4]$ inhibits the reaction either by releasing *two* $[Cu^ICl_2]^-$ species that are very slow to react or, more likely, by being very reluctant to break down in a redox step of the type (18) or (30).[1] The former suggestion is perhaps negated by the observation [1] that in the case of bromide the reaction becomes zero-order in ascorbate and first-order with respect to $[O_2]$ (strongly suggesting the re-oxidation of Cu^I by O_2 following stabilization of Cu^I by Br^-). Even at extremely high chloride concentrations the dependence on $[O_2]$ remains accurately half-order.

In the iron(III)-catalysed reaction between cysteine and oxygen [equation (32)] the kinetics involve half-order dependence [7] on $[O_2]$, whereas when iron(II) is used as the catalyst, the reaction displays first-order kinetics [7,8] with respect to $[O_2]$. This could also be explained on the basis of a scheme closely resembling that above if it is assumed that it is the kinetic step that is being followed

$$\text{iron(III)–cysteine–O}_2 \xrightarrow{k} \text{products} \qquad (32)$$

(spectrophotometrically) [8] and shown to persist over the pH range where the catalyst is effectively iron(III).

Thus, while by no means entirely ruling out our main contention that it is copper(III) that is involved in these reactions, the use of Occam's razor might well be invoked in favour of copper(I) involvement. But without further evidence this would be unwise as it means (*a*) too much reliance on purely kinetic arguments, (*b*) uncritical acceptance of the accidental equality (25), and (*c*) the rejection of our carefully presented non-kinetic arguments put forward in this and previous Papers in the series.

EXPERIMENTAL

Full experimental details are to be found in Parts 2 and 3. The temperature throughout was 25.00 ± 0.01 °C.

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REFERENCES

[1] Part 3, R. F. Jameson and N. J. Blackburn, *J. Chem. Soc., Dalton Trans.*, 1976, 1596.

[2] R. F. Jameson and N. J. Blackburn, *J. Chem. Soc., Dalton Trans.*, 1976, 534.

[3] E. B. Shtamm and Yu. Skurlatov, *Zh. Fiz. Khim.*, 1974, **48**, 1454.

[4] S. Ahrland and J. Rawsthorne, *Acta Chem. Scand.*, 1970, **24**, 157.

[5] N. J. Blackburn, unpublished work.

[6] K. Hayakawa and Y. Hayashi, *J. Nutr. Sci. Vitaminol.*, 1977, **23**, 395.

[7] R. F. Jameson, A. P. Masters, and J. M. Philp, *C.R. 21st Conf. Int. Chim. Coord., Toulouse*, 1980, 86.

[8] A. D. Gilmour and A. McAuley, *J. Chem. Soc. A*, 1970, 1006.

[9] Y. Yamamoto, K. Ishizu, and Y. Shimizu, *Chem. Lett.*, 1977, 735.

[10] Yu. Skurlatov, *Int. J. Chem. Kinet.*, 1980, **12**, 347.