Articles

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Redox Behavior of Blue Copper Model Complexes. Redox Potentials and Electron-Transfer Kinetics of Some Copper(II)–Copper(I) Complexes with Nitrogen and Thioether Donors

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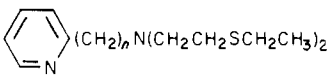
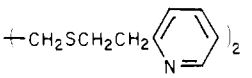
Copper(II)–copper(I) redox potentials, determined potentiometrically, of the copper complexes of



n = 1, pmas; n = 2, peas



pdto

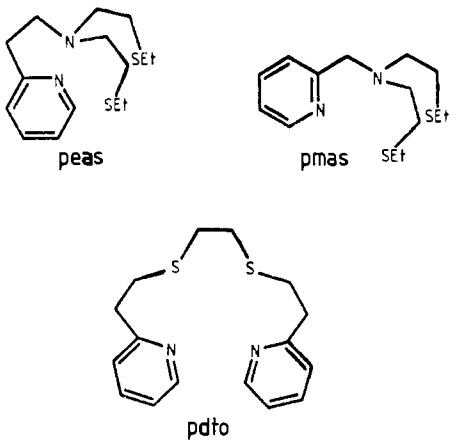
were found to be 398, 599, and 592 mV, respectively, in water at 25 °C and ionic strength 0.1 (KNO_3). Rate constants for one-electron reduction of the copper(II) complexes of these ligands by various reductants have been determined under the same conditions. Rate constants for reduction of the copper complexes are as follows: for Cu^{II}pmas, 1.66×10^3 M^{-1} s^{-1} by cytochrome c(II), 9.0×10^2 M^{-1} s^{-1} by $[Co(tpy)_2]^{2+}$, and 5.9×10^3 M^{-1} s^{-1} by $[Ru(NH_3)_5py]^{2+}$; for Cu^{II}peas, 6.9×10^3 M^{-1} s^{-1} by cytochrome c(II); for Cu^{II}pdto, 1.7×10^4 M^{-1} s^{-1} by cytochrome c(II) and 9.8×10^3 M^{-1} s^{-1} by $[Ru(NH_3)_4bpy]^{2+}$ (py = pyridine; bpy = bipyridine; tpy = terpyridine). These results are consistent with the Marcus cross relation. The structural barrier (internal rearrangement) to electron exchange is then estimated to be 32–36, 42, and 36–38 kJ mol^{-1}, respectively. It is concluded that the pseudotetrahedral geometry of the tripod ligands does not significantly increase the rate of electron transfer.

Introduction

Electron transfer between copper(II) and copper(I) is of central importance in a vast range of chemical and biochemical catalytic systems. Copper(II)–copper(I) electron transfer is also intrinsically interesting in comparison to the more well-studied iron, cobalt, and ruthenium electron-transfer couples. In these latter couples, both oxidation states are six-coordinate and usually one or both oxidation states are inert, whereas copper(II) and copper(I) prefer different coordination numbers and stereochemistry and both oxidation states are highly labile and stereochemically flexible.[1] In polar solvents copper(II) complexes exist predominately as five- or six-coordinate species, whereas copper(I) complexes are expected to prefer 4 or lower coordination numbers. Exceptions to this statement have been found only when steric constraints strongly favor another coordination number. Electron transfer between copper(II) and copper(I) will therefore normally be accompanied by major structural and stereochemical changes. These changes are expected to result in an appreciable activation energy barrier to electron transfer. In this and related papers[2] we examine the magnitude of this "structural barrier" to electron transfer and attempt to determine the important factors that contribute to it.

It has been shown previously that for bis(bipyridine)copper(II/I) and bis(phenanthroline)copper(II/I) complexes the structural barrier is around 30 kJ mol^{-1}.[3] It was also found that bis(2,9-dimethyl-1,10-phenanthroline)copper(II/I) electron transfer had a smaller structural barrier of 19 kJ mol^{-1}. The difference probably reflects the steric strain in the aquated bis(dimethylphenanthroline)copper(II) ion, which decreases the rearrangement energy required on reduction to the copper(I) complex.

Chart I



This paper reports on the redox properties of the copper complexes of some tetradentate N_2S_2 ligands (Chart I). These ligands were designed to examine two further aspects of copper(II/I) redox chemistry. First, kinetic data were obtained for sulfur donor ligands for the first time (apart from a preliminary communication[2]). Second, tetradentate tripodal ligands were chosen in order to ascertain whether the enforced pseudotetrahedral arrangement of the tripods would lead to

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(1) Jardine, F. H. *Adv. Inorg. Chem. Radiochem.* **1975**, *17*, 1975. Hathaway, B. J. *Struct. Bonding (Berlin)* **1973**, *14*, 49.
(2) Augustin, M. A.; Yandell, J. K.; Addison, A. W.; Karlin, K. D. *Inorg. Chim. Acta* **1981**, *55*, 235. Yandell, J. K. *Aust. J. Chem.* **1981**, *34*, 99.
(3) Augustin, M. A.; Yandell, J. K. *Inorg. Chem.* **1979**, *18*, 577.

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Table I. Ultraviolet-Visible Spectra of Copper(II) Complexes in Water[a]

	λ_{max}, nm (ϵ)
$Cu(pdto)^{2+}$	601 (5.7×10^2), 353 (4.4×10^3), 260 (1.07×10^4)
	600 (6.7×10^2), 354 (4.7×10^3)[b]
$Cu(pmas)^{2+}$	783 (4.5×10^2), 345 (2.7×10^3)
$Cu(peas)^{2+}$	675 (2.9×10^2), 352 (3.0×10^3), 262 (6.1×10^3)

[a] Conditions: pH 6; 10^{-2} MES buffer; ionic strength 0.1 (KNO_3). [b] Reference 17.

an appreciable decrease in the structural barrier and an enhancement of the rate of electron transfer.

The ligands also represent the best models[4] so far available for the N_2S_2[3] chromophore of the "blue" or type 1 copper enzymes[5] such as plastocyanin. In particular the tripodal ligands adopt a pseudotetrahedral arrangement about the copper atom in both copper(I) and copper(II) complexes similar to that in plastocyanin, though, unlike copper(II) plastocyanin, the copper(II) complexes probably coordinate a solvent molecule (in water). Nevertheless, the comparison with the enzymes was of considerable interest.

Experimental Section

Materials. [Cu(peas)]SO_4, [Cu(peas)]$NO_3 \cdot PF_6$, and [Cu(pmas)]SO_4 were synthesized as reported previously.[4] Solutions of these complexes were standardized by titration with ascorbic acid. Concentrations were as expected from mass except for [Cu(peas)]SO_4, which was found to be partly reduced (17%). pdto was prepared by the method of Goodwin and Lions.[6]

Copper(II) solutions of peas and pdto are partially dissociated under the conditions used in the kinetic and redox potential work, so it was necessary to add excess ligand or excess copper(II) to complex all of the ligand or all of the copper in solution. Visible spectra are reported in Table I. $[Ru(NH_3)_4bpy](ClO_4)_2$ was prepared by the method of Brown and Sutin.[7] The product, after purification as described by Brown and Sutin, was still contaminated with excess bipyridine. Solutions for kinetics were further purified by absorption of the $[Ru(NH_3)_4bpy]^{2+}$ on a short column of Sephadex C25 cation-exchange resin, washing of the column with 10^{-2} M phosphate buffer (pH 6) until it was free of bipyridine (UV absorption), and elution with 0.5 M KNO_3 or $NaClO_4$. Electronic spectra of these solutions had maxima at 522, 366, and 290 nm with $\epsilon(366)/\epsilon(522) = 1.63$ (cf. ref 7, 1.65) and $\epsilon(250)/\epsilon(366) = 5.30$ (cf. 5.50).

$[Ru(NH_3)_5py]Cl_3$ was prepared essentially by the method of Curtis and Meyer[8] and characterised by UV–visible spectra.[9] Solutions of $[Ru(NH_3)_5py]^{2+}$ were prepared by reduction with ascorbic acid followed by adsorption on Sephadex C25 ion-exchange resin and elution with 1 M KNO_3. $[Co(typ)_2]^{2+}$ solutions were prepared from cobalt nitrate solutions with 10% excess terpyridine. Solutions were prepared daily.

Equine cytochrome *c* (Sigma Type VI) was reduced in a buffered solution with excess ascorbic acid or sodium dithionite. The excess reagents and oxidized product were separated from the reduced cytochrome by gel filtration on a Sephadex G25 column. Solutions were prepared daily.

Analytical grade KNO_3, ascorbic acid (Sigma), and MES (morpholinoethanesulfonate, Sigma) were used without further purification. Ascorbic acid was shown to be better than 99% pure by titration with AR $K_3Fe(CN)_6$.

Methods. Redox Potentials. Potentials were measured between a saturated calomel electrode and a 1-cm platinum-wire electrode connected to a Metrohm pH/millivolt meter. Solutions of the copper complexes (approx 10^{-3} M) were titrated with ascorbic acid under an inert atmosphere (N_2 or Ar) in a glass cell maintained at 25 °C. Plots of cell potential against log ([Cu(II)]/[Cu(I)]) calculated from the stoichiometry yielded excellent straight lines except for [Cu(peas)]SO_4. In this case both the plot and the potential were consistent with some reduced complex (17%) being present initially. This was confirmed by titration of [Cu(peas)]$NO_3 \cdot PF_6$, which showed normal behavior. Errors quoted in Table VIII are the estimated random errors. Systematic errors (calomel electrode, liquid junction potentials) will probably add another ±5 mV error. The potential of the saturated calomel electrode was taken to be 0.242 V relative to the NHE.

Kinetics. Kinetics of oxidation of cytochrome *c*(II) to cytochrome *c*(III) were monitored spectrophotometrically at 420 nm by a conventional visible spectrophotometer (half-lives down to 5 s) or by a stopped-flow spectrophotometer. $Cu(pmas)^{2+}$ oxidation of $[Co(tpy)_2]^{2+}$ was studied with a stopped-flow spectrophotometer at 353 nm. $[Ru(NH_3)_4bpy]^{2+}$ and $[Ru(NH_3)_5py]^{2+}$ oxidations also required a stopped-flow spectrophotometer. These reactions were followed at 366 nm. All reactions, which were carried out with a greater than 10-fold excess of the copper complex, gave good first-order plots (log $(A_t - A_\infty)$, where A = absorbance, vs. time) over at least 2 half-lives. Small intercepts ($k_{obsd} < 0.1$ s^{-1}) were observed in some instances in plots of k_{obsd} vs. the copper(II) concentrations. Second-order rate constants were then obtained following subtraction of these intercepts.

Results

All the complexes are water soluble and in aqueous solution are stable for long periods. A strong near-ultraviolet absorption band is observed for all the complexes as well as weaker absorptions in the visible region (see Table I). No change in shape or position of these bands is observed in the pH range 5–6.5 though the intensity of the spectra decreases with decreasing pH because of dissociation of the complexes. Dissociation is not significant above pH 5 with $Cu(pmas)^{2+}$ and can be reduced with the remaining complexes by addition of excess ligand (or copper), so that under the conditions of the experiments described below no significant change in the spectrum of the copper(II) complexes occurs between pH 5 and 6.5.

Redox Potentials. Redox potentials (Table VIII) were measured by potentiometric titration with ascorbate ion. The stoichiometry of the reaction and the slope of the Nernst plot are consistent with one-electron reduction of the copper complexes. Potentials are therefore assigned to $Cu^{II}L/Cu^{I}L$ couples. Half-wave potentials for $Cu(pmas)^{2+/+}$ (0.24 V) and $Cu(peas)^{2+/+}$ (0.51 V) complexes in acetonitrile have been determined by Karlin and Sherman[4] and for $Cu(pdto)^{2+/+}$ (0.844 V) by Nikles, Powers, and Urbach.[10] A value of 0.577 V, determined by cyclic voltammetry, has also been reported[11] for $Cu(pdto)^{2+/+}$ in aqueous 0.1 M KNO_3. Considering the reported irreversibility of the cyclic voltammetry, the difference between this value and our value of 0.592 V is not significant.

Kinetics. All the copper(II) complexes rapidly oxidize cytochrome *c*(II) as judged by the characteristic difference in the visible absorption spectra of reduced and oxidized cytochrome *c*.[2] The reactions are therefore considered to be one-electron redox reactions (eq 1). The cyt *c*(III)/cyt *c*(II)

$$CuL^{2+} + \text{cyt } c(II) \rightarrow CuL^{+} + \text{cyt } c(III) \qquad (1)$$

redox potential of 0.26 V[12] is less than the copper(II)/copper(I) redox potentials, and thus the reactions go essentially to completion. Likewise the reactions with $[Co(bpy)_3]^{2+}$, $[Ru(NH_3)_4bpy]^{2+}$, and $[Ru(NH_3)_5py]^{2+}$ show the electronic spectral changes expected for one-electron oxidation of these

(4) Karlin, K. D.; Hayes, J. C.; Juen, S.; Hutchison, J. P.; Zubieta, J. *Inorg. Chem.* **1982**, *21*, 4108. Karlin, K. D.; Sherman, S. E. *Inorg. Chim. Acta* **1980**, *65*, 39. Karlin, K. D.; Dahlstrom, P. D.; Zubieta, J. *Transition Met. Chem. (Weinheim, Ger.)* **1981**, *6*, 314. Karlin, K. D.; Dahlstrom, P. D.; Hyde, J. R.; Zubieta, J. *J. Chem. Soc., Chem. Commun.* **1980**, 906. Karlin, K. D.; Dahlstrom, P. D.; Stanford, M. L.; Zubieta, J. *J. Chem. Soc., Chem. Commun.* **1979**, 465.

(5) Cass, A. E. G.; Hill, H. A. O. In "Biological Roles of Copper"; Evered, D., Lawrenson, G., Eds.; Excerpta Medica: Amsterdam, 1981; p 85. Colman, P. M., et al. *Nature (London)* **1978**, *272*, 319. Freeman, H. C. *Coord. Chem.* **1981**, *21*, 29. Lappin, G. A. *Met. Ions Biol. Syst.* **1981**, *13*, 15.

(6) Goodwin, H. A.; Lions, F. *J. Am. Chem. Soc.* **1960**, *82*, 5013.

(7) Brown, G. M.; Sutin, N. *J. Am. Chem. Soc.* **1979**, *101*, 883.

(8) Curtis, J. C.; Meyer, T. J. *Inorg. Chem.* **1982**, *21*, 1563.

(9) Gaunder, R.; Taube, H. *J. Am. Chem. Soc.* **1968**, *90*, 1187.

(10) Nikles, D. E.; Powers, M. J.; Urbach, F. L. *Inorg. Chim. Acta* **1979**, *37*, L499.

(11) Brubaker, G. R., et al. *Inorg. Chem.* **1979**, *18*, 299.

(12) Margalit, R.; Schejter, A. *Eur. J. Biochem.* **1973**, *32*, 492.

Table II. Rate Constants for Reduction of $Cu(pmas)^{2+}$ by Cytochrome *c*(II)

$10^5 \times$ [$Cu(pmas)^{2+}$], M	$10^{-3}k$,[a] $M^{-1}\ s^{-1}$	$10^5 \times$ [$Cu(pmas)^{2+}$], M	$10^{-3}k$,[a] $M^{-1}\ s^{-1}$
1.07	1.55	2.66	1.29[b]
2.13	1.55	2.66	1.51[c]
3.30	1.78	2.66	1.79[d]
4.22	1.57	5.40	0.56[e]
6.54	1.71	5.40	0.61[e]
11.29	1.68	5.40	0.26[f]
15.90	1.73	5.40	0.24[f]

[a] k = observed second-order rate constant; all data at 25 °C, pH 6.0, and ionic strength = 0.1 (KNO_3), except where indicated. [b] pH 7. [c] pH 5.5. [d] pH 5.0. [e] 14.0 °C. [f] 4.2 °C.

Table III. Rate Constants for Reduction of $Cu(pmas)^{2+}$ by $[Co(tpy)_2]^{2+}$

$10^4 \times$ [$Cu(pmas)^{2+}$], M	$10^{-2}k$,[a] $M^{-1}\ s^{-1}$	$10^4 \times$ [$Cu(pmas)^{2+}$], M	$10^{-2}k$,[a] $M^{-1}\ s^{-1}$
1.56	9.2	12.5	8.9
3.13	9.0	31.3	8.9
6.3	9.7	1.56	10.0[b]

[a] k = observed second-order rate constant; all data at pH 6.0, 25 °C, and ionic strength 0.1 (KNO_3), except where indicated. [b] pH 5.2.

Table IV. Rate Constants for Reduction of $Cu(pmas)^{2+}$ by $[Ru(NH_3)_5py]^{2+}$

$10^4 \times$ [$Cu(pmas)^{2+}$], M	$10^{-3}k$,[a] $M^{-1}\ s^{-1}$	$10^4 \times$ [$Cu(pmas)^{2+}$], M	$10^{-3}k$,[a] $M^{-1}\ s^{-1}$
0.61	5.85	9.85	5.90
1.23	5.96	0.61	6.24[b]
2.46	5.91	0.61	6.38[c]
4.93	5.92		

[a] k = observed second-order rate constant; all data at pH 6.0, 25 °C, and ionic strength 0.1 (KNO_3), except where indicated. [b] pH 5.5. [c] pH 6.5.

complexes. Again, this is consistent with the known redox potentials of the complexes.[13,14]

All of the reactions were found to be first order in the concentration of the reductant and first order in the concentration of the copper complex. Second-order rate constants are given in Tables II–VII.

Discussion

Nature of the Copper Complexes in Solution. Crystal structures of both copper(I) and copper(II) complexes of all the ligands have been published.[4,11] In all of these crystals the copper is coordinated by both nitrogen donors and both sulfur atoms of the tetradentate ligand. The copper(I) complexes are monomeric and four-coordinate, but with copper(II) five-coordinate species are formed by complexing with a counteranion.

In water there is little evidence to suggest that the Cu(I) complexes will not remain four-coordinate, though five-coordinate Cu(I) complexes are known[15] and particularly with $Cu(pmas)^+$ there would not appear to be any steric constraint to coordination of a solvent molecule.

The copper(II) complexes in water will no doubt coordinate at least one and possibly two water molecules and thus will be five- or six-coordinate. We expect that the ligands act as quadridentates in aqueous solutions of the copper(II) complexes as in the crystal structures though this is not entirely certain because thioether sulfur is a very weak donor for copper(II).[16] With the ligands under discussion, however, thioether coordination to copper(II) is evident in two ways. First, all the thioether complexes have a strong absorption around 350 nm, which is attributable to Cu(II) → S charge transfer.[17] Second, the stability constants for formation of the copper(II) complexes are much higher than would be

Table V. Rate Constants for Reduction of $Cu(pdto)^{2+}$ by Cytochrome *c*(II)

$10^5 \times$ [$Cu(pdto)^{2+}$], M	$10^4 \times$ [$Cu(aq)^{2+}$],[d] M	$10^{-3}k$,[a] $M^{-1}\ s^{-1}$
2.02	9.0	17.0
2.02	2.0	16.8
5.04	2.0	18.4
10.1	2.0	17.2
20.2	2.0	17.0
40.3	2.0	16.8
5.04	2.0	16.5[b]
4.73	2.0	5.9[c]
9.34	2.0	5.7[c]
18.3	2.0	5.6[c]
22.7	2.0	5.7[c]
42.5	2.0	5.7[c]

[a] k = observed second-order rate constant; all data at pH 6.0, 25 °C, and ionic strength 0.1 (KNO_3), except where indicated. [b] pH 5. [c] 5 °C. [d] Concentration of added aqueous Cu^{2+}.

Table VI. Rate Constants for Reduction of $Cu(pdto)^{2+}$ by $[Ru(NH_3)_4bpy]^{2+}$

$10^4 \times$ [$Cu(pdto)^{2+}$], M	$10^4 \times$ [pdto], M	$10^6 \times$ [$[Ru(NH_3)_4$-bpy$]^{2+}$], M	10^3k,[a] $M^{-1}\ s^{-1}$
0.5	5.0	3.5	9.8
1.0	5.0	3.5	9.7
1.25	5.0	5.7	10.0
2.50	0	5	10.0
2.50	5.0	3.5	10.8
5.0	0	12	9.6
5.0	5.0	12	9.8
5.0	5.0	5	10.4[b]
5.0	5.0	5	10.4[c]
10.0	10.0	10	9.8

[a] k = observed second-order rate constant; all data at pH 6.0, 25 °C, and ionic strength 0.1 (KNO_3), except where indicated. [b] pH 5.5. [c] pH 6.5.

Table VII. Rate Constants for Reduction of $Cu(peas)^{2+}$ by Cytochrome *c*(II)

$10^4 \times$ [$Cu^{2+}(aq)$],[d] M	$10^4 \times$ [pdto], M	10^3k,[a] $M^{-1}\ s^{-1}$
0.49	5.75	6.6
1.23	5.75	7.1
2.45	5.75	7.3
4.90	5.75	6.9
4.90	5.75	6.8[b]
4.90	5.75	7.0[c]
9.80	11.5	6.9
18.6	23.0	6.2

[a] k = observed second-order rate constant; all data at pH 6.5, 25 °C, and ionic strength 0.1 (KNO_3), except where indicated. [b] pH 6.0. [c] Cyt *c*(II) concentration = 0.5 μM. [d] Concentration of added aqueous Cu^{2+}.

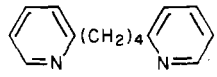
(13) Cummins, D.; Gray, H. B. *J. Am. Chem. Soc.* **1977**, *99*, 5158.
(14) Yee, E. L.; Weaver, M. J. *Inorg. Chem.* **1980**, *19*, 1077.
(15) Addison, A. W.; Carpenter, M.; Lau, L. K.-M.; Nicholas, M. *Inorg. Chem.* **1978**, *17*, 1545. Drew, M. G. B.; Cairns, C.; McFall, S. G.; Nelson, S. M. *J. Chem. Soc., Dalton Trans.* **1980**, 2020.
(16) Sigel, H.; Rheinberger, V. M.; Fischer, B. E. *Inorg. Chem.* **1979**, *18*, 3334.
(17) Amundsen, K. A.; Whelan, J.; Bosnich, B. *J. Am. Chem. Soc.* **1977**, *99*, 6730.

Table VIII. Copper(II)-Copper(I) Redox Potentials and Rate Constants for Electron Transfer and Exchange

oxidant	E°,[a,b] mV vs. NHE (slope)	reductant	k_{12},[a,c] M^{-1} s^{-1}	k_{11},[d] M^{-1} s^{-1}	$\Delta G^{\ddagger}$(int),[e] kJ mol^{-1}
$Cu(pdto)^{2+}$	592 ± 4 (59)	cyt *c*(II)	1.70×10^4	5	38
		$[Ru(NH_3)_4bpy]^{2+}$	9.8×10^3	11	36
$Cu(pmas)^{2+}$	398 ± 2 (59)	cyt *c*(II)	1.66×10^3	45	32
	394[f] (57)	$[Co(tpy)_2]^{2+}$	9.0×10^2	36	32
		$[Ru(NH_3)_5py]^{2+}$	5.9×10^3	6	36
$Cu(peas)^{2+}$	599 ± 2 (59)	cyt *c*(II)	6.9×10^3	0.7	42

[a] All data at 25 °C and ionic strength 0.1 (KNO_3). [b] From titration with ascorbate. [c] Observed second-order rate constants. [d] Cu(II)/Cu(I) electron exchange rate constant: r = 4.7 Å, peas, pmas; r = 5.0 Å, pdto. [e] Internal rearrangement energy. [f] From cyclic voltammetry.

expected for nitrogen coordination alone. For example, the stability constant for 1,4-dipyridylbutane (log K_1 = 2.6)



is less than the stability constant[18] for $Cu(pdto)^{2+}$ (log K_1 > 6). Likewise, the first stability constants for (aminomethyl)pyridine[19] (log K_1 = 10.0) and (aminoethyl)pyridine[19] (log K_1 = 7.8) are less than for pmas[18] (log K_1 > 11) and peas[18] (log K_1 > 9). It is particularly interesting that the terminal thioethers in the tripod ligands contribute substantially to the stability of the copper(II) complexes.

In the solid state[4] the tripodal ligands maintain a roughly tetrahedral arrangement of donor atoms in both copper(II) and copper(I) complexes, whereas the donor atoms of pdto are square planar in [Cu^{II}pdto] ([Cu(pdto)ClO_4]$^+$)[11] but are nearly tetrahedral in [Cu^{I}pdto]. This is probably true in solution also, because of the steric constraints in the tripodal ligands. Another interesting feature of the copper(II) tripods is the difference in their structures—[Cu^{II}peas] is pyramidal but [Cu^{II}pmas] is closer to a trigonal pyramid. This difference is found in solution as well as in the solid state. The visible spectra of these complexes in water (Table I) are consistent with this observation.

Hydroxy complexes do not appear to form appreciably below pH 7 since the visible spectra are independent of pH below 7 (except for dissociation). MES has been shown not to coordinate to copper(II).[3] The nitrate ion is not expected to coordinate in solution at the concentrations used in these experiments because nitrate has a low affinity for copper(II) in water.

Under the conditions of our experiments, therefore, it is likely that the copper(II) complexes are five-coordinate species with two nitrogen and two sulfur donor atoms from the quadridentate ligand and one water molecule. The copper(I) complexes are expected to have the same donor set except for the coordinated water molecule.

Redox Potentials. Redox potentials for $Cu^{II}L/Cu^{I}L$ show the expected dependence on the nature of the donor atoms. The presence of thioether donors raises Cu(II)/Cu(I) potentials by about 110 mV/donor[2] (relative to water) so the potentials are all fairly high. What is extraordinary, at first sight, is the 200-mV difference between the potentials of the two tripod ligand systems.

It now appears that an increase of 150–200 mV when the chelate ring size is increased from 5 to 6 is a feature of ligands containing aminopyridine chelate rings[2,4,10] and that smaller increases are found for other chelate systems. The higher potentials of the six-membered chelates have been attributed to an increase in stability of the Cu(I) complexes resulting from the larger chelate more closely adopting a tetrahedral configuration. This is not a sufficient explanation, however, since the effect of ring size on the stability of the copper(II) complexes cannot be neglected, particularly for aminopyridine chelates. The copper(II) complex of (aminomethyl)pyridine is more stable by a factor of $10^{2.2}$ than that of (aminomethyl)pyridine,[19] and a similar difference is found for the *N*'ethyl analogues of these ligands. $Cu(pmas)^{2+}$ is also more stable than $Cu(peas)^{2+}$ by a factor of $\sim 10^2$.[18] Such differences account for more than half the observed differences in redox potentials for the aminopyridine systems. Smaller differences are found between five- and six-membered chelate rings of other donor systems (e.g. $10^{0.8}$ for ethylenediamine relative to propylenediamine), which is reflected in smaller redox potential differences.

Kinetics. All the reactions are taken to be one-electron, outer-sphere electron-transfer reactions, principally because of the inertness of the reductants coupled with the lack of any suitable bridging ligands precludes inner-sphere mechanisms.[3] If the reactions are also adiabatic, then the Marcus cross relation may be used to determine the rate constants for self-exchange reactions between $Cu^{II}L$ and $Cu^{I}L$, i.e. for the reaction

$$Cu^{II}L + Cu^{I}L \rightleftharpoons Cu^{I}L + Cu^{II}L$$

The results are given in Table VIII. (These figures have been corrected for charge according to Wherland and Gray.[20] The corrections are small and only weakly dependent on the radii of the reactants.)

It is gratifying to find the excellent agreement between the k_{11} values derived from reaction with different reductants. This strongly supports the validity of this method of determining electron exchange rate constants.

Such rate constants are not expected to be reliable to better than a factor of 10.[3,21] The final column in Table VIII lists the internal free energy of activation, which is derived from k_{11} with a correction for outer-sphere effects.[7] These numbers are therefore uncertain to the same extent as k_{11} and, in addition, are significantly affected by the choice of radii for the complexes. Differences of less than 5 kJ mol^{-1} cannot therefore be regarded as significant.

Conclusions

The outstanding feature of the kinetic results both in this paper and in earlier work is the similarity of the Cu(II)–Cu(I) electron exchange rate constants found for copper complexes with a diverse range of donor atoms and stereochemistries.[2] With few exceptions rate constants for self-exchange fall in the range 1–10^2 M^{-1} s^{-1}. The associated internal free energies of activation (structural barriers) also fall in a narrow range around 30–35 kJ mol^{-1}. For the systems reported in this paper $Cu(pmas)^{2+}$ and $Cu(pdto)^{2+}$ fall in this "normal" range though $Cu(peas)^{2+}$ reacts a little more slowly and exhibits a slightly higher activation barrier. It appears therefore that the specific characteristics of these copper complexes do not strongly affect their rate of electron exchange or the magnitude of the structural barrier. In particular a decrease in the structural

(18) Yandell, J. K., unpublished results.
(19) Couturier, Y.; Petitfaux, C. *Bull. Soc. Chim. Fr.* **1978**, 435.
(20) Wherland, S.; Gray, H. *Proc. Natl. Acad. Sci. U.S.A.* **1976**, *73*, 2950.
(21) Chou, M.; Creutz, C.; Sutin, N. *J. Am. Chem. Soc.* **1977**, *99*, 5615.

barrier is not observed with the tripodal ligands even though the approximately tetrahedral arrangement of the tripod is expected to be maintained in both oxidation states. *We conclude that rearrangement from tetragonal to tetrahedral does not make a significant contribution to the structural barrier. We propose instead that the important factor is the change of coordination number on reduction*, which in the present case requires the loss of coordinated water molecules.

A second specific feature of the ligands in this paper is the presence of sulfur donor atoms. These too do not have any dramatic effect on the rate of electron transfer or exchange.

Comparison to Copper Proteins. The "blue" or type 1 copper proteins have intrigued chemists because of their unusual physical properties in the Cu(II) state. Crystal structure studies[5b,c] have now shown that the copper site in these proteins (plastocyanin and azurin) is four-coordinate, with two histidine nitrogen donor atoms and two sulfur donor atoms from cysteine and methionine side chains. Presumably this arrangement has some functional value to the proteins in their role in electron transport. It has been argued that the sulfur donor atoms are necessary to obtain the "high" redox potentials of the copper proteins (plastocyanin, 0.35 V; azurin, 0.33 V). The $E°$ values for $Cu(peas)^{2+/+}$ and $Cu(pdto)^{2+/+}$ are in fact considerably higher than those of the proteins, even though they have the same N_2S_2 donor set. Other features of the protein—the four-coordinate, tetrahedral coordination[2,22] and the protein envelope around the copper site[2,23] —are expected to further increase the potential of the proteins relative to those of small copper complexes, so that the protein redox potentials appear to be abnormally *low* rather than high compared to those of model compounds. In part this is because the thiol sulfur of cysteine will produce a lower potential than the thioether sulfur donors of the models.[24] Also, the considerable effect of chelate ring size on the redox potentials of copper complexes does suggest that strain within the coordination sphere of the copper proteins may be important in determining their redox potentials. Whatever the reasons for the low potentials, it is apparent that sulfur donors are not necessary to obtain potentials as high or higher than those of plastocyanin and azurin.

It might be argued instead that the sulfurs have a kinetic role. The results in this paper do not support this argument, however, at least not for thioether sulfur.

Electron exchange rate constants have been derived for plastocyanin and azurin from reactions with inorganic oxidants and reductants[5d,25] and cytochromes.[5d,26] For good reagents the exchange rate constants fall in the range 10^3–10^5 M^{-1} s^{-1}. This would suggest that these proteins have a kinetic advantage over model systems such as those in this paper of from 10 to 10^5 M^{-1} s^{-1} and that the structural barrier is lower in the copper proteins by 5–20 kJ mol^{-1}. As the protein shell will decrease the rate of electron transfer by reducing access to the copper center, the kinetic advantage of the copper structure in the protein may be even greater than suggested by these figures. The structural barrier in the blue copper proteins may therefore be quite small. This observation is consistent with the conclusion above—that the main factor contributing to the slow rate of electron transfer is the change in coordination number accompanying electron transfer—since the coordination number does not change on oxidation or reduction in the proteins.

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Registry No. $Cu(pdto)^{2+}$, 64685-82-1; $Cu(pdto)^{+}$, 68378-73-4; $Cu(pmas)^{2+}$, 88854-98-2; $Cu(pmas)^{+}$, 76682-77-4; $Cu(peas)^{2+}$, 72077-09-9; $Cu(peas)^{+}$, 72077-07-7; $Ru(NH_3)_4bpy^{2+}$, 54194-87-5; $Co(tpy)_2^{2+}$, 18308-16-2; $Ru(NH_3)_5py^{2+}$, 21360-09-8; cyt *c*(II), 78690-22-9.

(22) Yokoi, H.; Addison, A. W. *Inorg. Chem.* **1977**, *16*, 1341.
(23) Kassner, R. J. *Proc. Natl. Acad. Sci. U.S.A.* **1968**, *59*, 498.
(24) Addison, A. W.; Rao Nageswara, T.; Sinn, E. *Inorg. Chem.*, in press. Addison, A. W. In "Copper Coordination Chemistry: Biochemical and Inorganic Perspectives"; Karlin, K. D., Zubieta, J., Eds.; Adenine Press: New York, 1983; p 109.
(25) Mauk, A. G.; Bordignon, E.; Gray, H. B. *J. Am. Chem. Soc.* **1982**, *104*, 7654.
(26) Wherland, S.; Pecht, I. *Biochem. Biophys. Res. Commun.* **1973**, *50*, 853.

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New Multidentate Ligands. 22. *N,N'*-Dipyridoxylethylenediamine-*N,N'*-diacetic Acid: A New Chelating Ligand for Trivalent Metal Ions

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The synthesis and study of a new sexadentate ligand, *N,N'*-dipyridoxylethylenediamine-*N,N'*-diacetic acid (PLED), and of its proton and metal ion affinities are described. The ligand has two phenolate donors attached to a nitrogen heterocyclic ring to impart specificity for trivalent metal ions such as those of Ga(III), In(III), and Fe(III). The ligand has higher overall basicity than analogous phenolic sexadentate ligands containing hydroxybenzyl in place of pyridoxyl groups. Potentiometric studies of the protonation constants of the ligand and its metal chelates are reported. Stability constants of the metal chelates of Cu(II), Ni(II), Co(II), Zn(II), Fe(III), Ga(III), and In(III) ions are reported and are compared with those of analogous ligands. A comparison of the affinities of PLED and human serum protein transferrin for gallium(III) shows that PLED may compete successfully in vivo with transferrin for this metal ion.

Introduction

Frost et al.[1b] and Anderegg and L'Eplattenier[2] have reported the high stability of the Fe(III) chelate of *N,N'*-ethylenebis-[2-(*o*-hydroxyphenyl)glycine], EHPG (**1**), to be due to the high affinity of Fe(III) for the two phenolate groups present in the ionized ligand and to the orientation of these groups so as to permit their participation in chelate-ring formation. However, the structure of EHPG is such that steric hindrance may interfere somewhat with simultaneous participation in metal ion coordination by all six donor groups (two basic nitrogens, two carboxylate groups, and the two phenolate groups), with the result that the two (axial) carboxylate groups are displaced

(1) (a) Abstracted in part from a dissertation to be submitted by Christina H. Taliaferro to the faculty of Texas A&M University in partial fulfillment of the requirements for the degree of Doctor of Philosophy. (b) Frost, A. E.; Freedman, H. H.; Westerback, S. J.; Martell, A. E. *J. Am. Chem. Soc.* **1958**, *80*, 530.
(2) Anderegg, G.; L'Epllattenier, F. *Helv. Chim. Acta* **1964**, *47*, 1067.

0020-1669/84/1323-1188$01.50/0