Internal Reorganization Energies for Copper Redox Couples: The Slow Electron-Transfer Reactions of the $[Cu^{II/I}(bib)_2]^{2+/+}$ Couple

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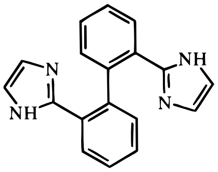
$[Cu^{II}(bib)_2]^{2+}$ and $[Cu^{I}(bib)_2]^{+}$ (bib = 2,2′-bis(2-imidazolyl)biphenyl) have similar geometries, with the Cu(II) complex best described as pseudo-square planar and the Cu(I) complex pseudo-tetrahedral. The cyclic voltammograms in acetonitrile are only quasi-reversible, with $\Delta E_{p/p} \sim 103$ mV, which leads to uncertainty in $E_{1/2}$. UV−vis measurements of the reaction $[Cu^{II}(bib)_2]^{2+} + [Co^{II}((nox)_3(BC_6H_5)_2)] \rightleftharpoons [Cu^{I}(bib)_2]^{+} + [Co^{III}((nox)_3(BC_6H_5)_2)]^{+}$ ($(nox)_3(BC_6H_5)_2$ = a clathrochelate ligand) yield $K = 2.6$ in acetonitrile at 25 °C with $\mu = 0.1$ M (Et_4NBF_4). Using the known $E°$ for the $Co^{III/II}$ couple yields $E° = 0.274$ V vs Ag/AgCl for $[Cu^{II/I}(bib)_2]^{2+/+}$. This, in turn, leads to $K = 72$ for reaction with $[Co^{II}((nox)_3(BC_4H_9)_2)]$. The reaction of $[Cu^{II}(bib)_2]^{2+}$ with $[Co^{II}((nox)_3(BC_4H_9)_2)]$ obeys simple bimolecular kinetics with $k = 87\ M^{-1}\ s^{-1}$. Similar behavior occurs with $[Co^{II}((nox)_3(BC_6H_5)_2)]$ and $[Co^{II}((dmg)_3(BC_4H_9)_2)]$ as reductants. A value of $k_{11} = 0.16\ M^{-1}\ s^{-1}$ is calculated for the $[Cu^{II}(bib)_2]^{2+}/[Cu^{I}(bib)_2]^{+}$ self-exchange reaction based on the Marcus cross-exchange relationship and the known electron-self-exchange rate constants for the Co(III)/Co(II) couples. This value of k_{11} is remarkably small for a Cu(II/I) system, especially in view of the small structural change at copper ($\Delta(Cu-N)_{ave} = 0.07$ Å) during redox turnover. However, molecular mechanics calculations with newly developed copper force fields demonstrate $\Delta G^{\ddagger}_{in}$ to be the major factor that accounts for the small value of k_{11}. Much of this large reorganization energy appears to arise from angular distortions around copper. Nevertheless, the structural changes for the $[Cu^{II/I}(bib)_2]^{2+/+}$ couple are small relative to those for systems that have much greater self-exchange rate constants. These considerations underscore the difficulties in applying the entatic hypothesis to cuproprotein active sites.

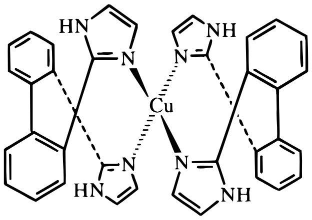
Introduction

Cu(II/I) electron-transfer reactions are currently of interest because the structural differences between the two oxidation states place unusual constraints on the electron-transfer kinetics. Insight into this behavior has application in understanding and manipulating the properties of cuproproteins.

Examples of small-molecule CNI (coordination-number-invariant) Cu(II/I) couples are rare. Even more rare are those that are conformationally invariant. It is this latter group, however, which is most directly evaluated by Marcus theory and which bears the closest relationship to the "entatic" properties of the CNI blue copper protein centers. A few years ago, Knapp et al. reported the synthesis of a bidentate chelating ligand 2,2′-bis(2-imidazolyl)biphenyl, bib.[1] This ligand has two imidazole groups appended to a biphenyl moiety, as shown directly below.



Knapp et al. also reported the synthesis of the Cu(II) and Cu(I) bis-chelates of bib and their X-ray crystal structures, which show that both oxidation states have the same basic structure as shown here.



The $[Cu^{II/I}(bib)_2]^{2+/+}$ couple has special importance because the structures and ligand conformations in the two oxidation states are essentially identical. Thus, this redox couple seems ideally suited as a test of the "entatic hypothesis"[2] for copper protein redox centers. In other words, should the small, restricted geometry differences between the two oxidation states in $[Cu(bib)_2]^{2+/+}$ lead to rapid or slow electron-transfer kinetics? In the original report on $[Cu^{II/I}(bib)_2]^{2+/+}$, NMR line-broadening measurements gave only a negative result (no line broadening for the di-*N*-methyl derivative), from which an upper limit of 100 $M^{-1}\ s^{-1}$ for k_{11} was inferred.[1] Here we report new results in the form of direct measurements of the electron-transfer kinetics for this system by use of the UV−vis spectrophotometry, as well as complementary molecular mechanics calculations using a force field approach for copper(I) and copper(II). The reactions studied are the reductions of $[Cu(bib)_2]^{2+}$ by a

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(1) Knapp, S.; Keenan, T. P.; Zhang, X.; Fikar, R.; Potenza, J. A.; Schugar, H. J. *J. Am. Chem. Soc.* **1990**, *112*, 3452−3464.

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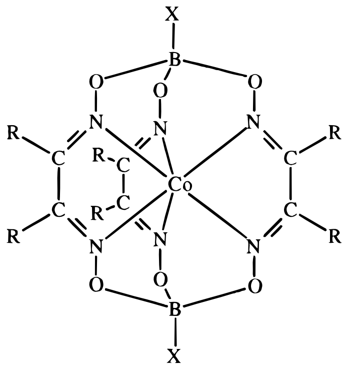
seriesof cobalt(II) clathrochelates of the general form as shown below.



$R = (C_4H_8)_{1/2}, CH_3; X = C_6H_5, C_4H_9$

These cobalt clathrochelates have been thoroughly studied and are well-known to act as outer-sphere reducing agents.[3−6] The electron-transfer kinetics yield an unusually low self-exchange rate constant, and the molecular mechanics calculations imply that the internal reorganization energy is the major component of the activation barrier.

Experimental Section

Materials. All solvents were reagent grade or better. Acetonitrile (Fisher) for electronic spectra, equilibrium constant, kinetic, and electrochemical measurements was distilled under N_2 from CaH_2 and stored in a glovebox under nitrogen.

Methanol, acetone, dichloromethane, benzene, ethyl acetate, dimethylformamide, 2-propanol, hexane, methylene chloride, barium chloride, and sodium hydroxide were reagent grade from Fisher. Pentane, hexamethylphosphoramide, tetrahydrofuran, thionyl chloride, diphenic acid, sulfamide, sulfolane, *p*-toluenesulfonic acid monohydrate, ethylenediamine, potassium hydride, tetraethylammonium tetrafluoroborate, 1-butaneboronic acid, 1,2-cyclohexanedione dioxime, phenylboronic acid, and nitrosonium tetrafluoroborate were purchased from Aldrich. Potassium permanganate and dimethylglyoxime were purchased from J. T. Baker. $Cu(ClO_4)_2{\cdot}6H_2O$ was from Alfa products.

The supporting electrolyte, n-Et_4NBF_4 (Aldrich), used in electronic spectra, equilibrium constant, kinetic, and electrochemical measurements was recrystallized three times from a mixture of methanol and hexane (4:1) instead of using a mixture of methanol and petroleum ether (9:1)[7] and dried under vacuum at 95 °C in an Abderhalden pistol.

Preparation of the Ligands. The synthesis of the ligand, 2,2′-bis(imidazolyl)biphenyl (bib), has been reported in the literature.[1] The ligand was purified by flash chromatography;[8] the literature method was modified by the use of a mixture of 3:7 methanol/ethyl acetate as eluant (54% yield).[1] The ligand was identified by melting point, FTIR (IBM FT-IR44 spectrometer), ^{1}H NMR (Bruker AC250 MHz NMR spectrometer), and MS (Finnigan 3300 spectrometer).

Preparation of the Copper Complexes. Bis[2,2′-bis(2-imidazolyl)biphenyl]copper(II) diperchlorate, $[Cu^{II}(bib)_2](ClO_4)_2$ was synthesized, purified, and isolated by the methods of Knapp et al.[1] The identity and purity of this compound were confirmed by comparison of its powder X-ray diffraction pattern with that generated from the reported single-crystal atomic coordinates, its cyclic voltammetry, and its UV−vis spectrum.[1] Bis[2,2′-bis(2-imidazolyl)biphenyl]copper(I) perchlorate, $[Cu^{I}(bib)_2](ClO_4)$ was likewise prepared according to the methods of Knapp et al.[1] and stored in a glovebox under nitrogen.

Preparation of the Cobalt Complexes. All cobalt compounds, $[Co^{II}((nox)_3(BC_6H_5)_2)]$, $[Co^{II}((nox)_3(BC_4H_9)_2)]$, $[Co^{II}((dmg)_3(BC_4H_9)_2)]$, $[Co^{III}((nox)_3(BC_6H_5)_2)]BF_4$, were prepared as previously described in the literature for the phenylborate-capped clathrochelates, $[Co^{II}((dpg)_3(BC_6H_5)_2)]$ and $[Co^{III}((dpg)_3(BC_6H_5)_2)]BF_4$,[4] but 1,2-cyclohexanedionedioxime (nox) and dimethylglyoxime (dmg) were substituted for diphenylglyoxime, and *n*-butaneboronic acid was substituted for benzeneboronic acid as required. All cobalt compounds were characterized by visible spectroscopy. Anal. Calcd for $[Co^{II}((nox)_3(BC_6H_5)_2)]{\cdot}CH_3OH$: C, 54.18; H, 5.57; N, 12.23. Found: C, 54.16; H, 5.24; N, 12.62. Anal. Calcd for $[Co^{II}((nox)_3(BC_4H_9)_2)]{\cdot}H_2O$: C, 49.30; H, 7.00; N, 13.27. Found: C, 49.23; H, 6.67; N, 13.27. Anal. Calcd for $[Co^{II}((dmg)_3(BC_4H_9)_2)]$: C, 44.67; H, 6.75; N, 15.64. Found: C, 44.64; H, 6.68; N, 15.69.

Electrochemistry. A BAS-100 electrochemical analyzer with Pt wire as an auxiliary electrode, Ag/AgCl as a reference electrode, and a Pt disk as a working electrode were used for the cyclic voltammetry measurements. The working electrode was preconditioned by electrolysis at −2 V in acetonitrile with 0.1 M n-Et_4NBF_4.[3] Scan rates ranged from 30 to 300 mV s^{-1}.

Spectrophotometric Equilibrium Constant. The equilibrium constant of reaction between $[Cu^{II}(bib)_2](ClO_4)_2$ and $[Co^{II}((nox)_3(BC_6H_5)_2)]$ was obtained by the standard method of spectrophotometry using a Hewlett-Packard 8452A spectrophotometer with quartz cells of 1.0-cm path length. The solutions were equilibrated in cuvettes thermostated at 25 °C. All solutions of $[Cu^{II}(bib)_2](ClO_4)_2$ (initial concentration: 0.25, 0.125, 0.05 mM, respectively), $[Co^{II}((nox)_3(BC_6H_5)_2)]$ (initial concentration: 0.025 mM), $[Co^{III}((nox)_3(BC_6H_5)_2)]BF_4$ (initial concentration: 0.25 mM) were made in acetonitrile with 0.1 M n-Et_4NBF_4 in a glovebox under nitrogen. The progress of the reaction was monitored at 472 nm, which corresponds to the absorption peak of the Co(II) complex. Three runs were made at the same conditions except various initial concentrations of Cu(II).

Cross-Exchange Kinetics. The kinetics studies were performed on a Hewlett-Packard 8452A spectrophotometer equipped with thermostated 1.0-cm quartz cuvettes. Reactions were initiated by mixing equal volumes of the two reactant solutions, both of which were maintained at 0.1 M ionic strength by use of n-Et_4NBF_4 as the background electrolyte. In all cases there was at least a 10-fold excess of $[Cu^{II}(bib)_2]^{2+}$ over the Co(II) complexes. Stock solutions of Cu(II) were prepared by weight. The decay of the absorbance of Co(II) was monitored at 472 nm for about 10 half-lives. The data were transferred to an OLIS 4300S system for data analysis; values of the pseudo-first-order rate constants, k_{obs}, were obtained by fitting the absorbance data with an exponential decay function. Values of k_{obs} were reproducible ±5% or better.

Temperature-dependent rate constants were obtained by use of a Hi-Tech SF-51 stopped-flow apparatus equipped with a SU-40 spectrophotometer and a C-400 circulating water bath that maintained and changed the temperature of the cell compartment from 7 to 45 °C. At least five shots were performed at each set of conditions, an OLIS 4300s system being used for data acquisition and analysis as described above.

A nonlinear-least-squares program was used to fit rate laws to the values of k_{obs}, with weighting as the inverse square of k_{obs}.[9] Activation parameters were obtained with k_{12} as the dependent variable and with weighting as the inverse square of k_{12}. Uncertainties are expressed as one standard deviation.

Molecular Mechanics. The SYBYL 6.3 program (Tripos, Inc.) was used in the molecular mechanics calculations with the TAFF force field[10] present in SYBYL. This force field was supplemented with the following preliminary parameters for copper: a van der Waals radius of 1.676 Å and a van der Waals constant of 0.100 kcal/mol were used for both Cu(I) and Cu(II); an out-of-plane force constant of 120 kcal mol^{-1} Å^{-2} was used for planes involving copper and three coordinated imidazole nitrogen atoms; a force constant of zero was used for torsional

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modes with copper as one of the central atoms; for torsions of the type *−C−N−Cu a force constant of 1.6 kcal mol^{-1} deg^{-2} and periodicity of −2 were used both for eclipsed and staggered minima; an N−Cu−N equilibrium angle of 109.5° and a force constant of 0.03 kcal mol^{-1} degree^{-2} were used for Cu(I), while these angular terms were omitted for Cu(II); the Cu(I)−N bond stretch was assigned an equilibrium length of 2.03 Å and a force constant of 400 kcal mol^{-1} Å^{-2}, while for Cu(II) the equilibrium bond length was 1.94 Å; an equilibrium angle of 120° and force constant of 0.02 kcal mol^{-1} deg^{-2} were used for the Cu−N−C angles. Structures were optimized starting from the crystallographic atomic coordinates.[1,11] Tables giving the equilibrium structures of $[\mathrm{Cu(bib)_2}]^{2+}$, $[\mathrm{Cu(bib)_2}]^{+}$, $[\mathrm{Cu(bimdpk)_2}]^{2+}$, and $[\mathrm{Cu(bimdpk)_2}]^{+}$ (bimdpk = bis(1-methyl-4,5-diphenylimidazol-2-yl)ketone) in the PDB format are shown in the Supporting Information.

Results

In this paper we report the electron-transfer reactions of $[\mathrm{Cu(bib)_2}]^{2+}$ with three different Co(II) clathrochelate complexes, as indicated in reactions 1−3.

$$[\mathrm{Cu^{II}(bib)_2}]^{2+} + [\mathrm{Co^{II}((nox)_3(BC_6H_5)_2)}] \rightleftharpoons [\mathrm{Cu^{I}(bib)_2}]^{+} + [\mathrm{Co^{III}((nox)_3(BC_6H_5)_2)}]^{+} \qquad k_{12}, k_{21} \quad (1)$$

$$[\mathrm{Cu^{II}(bib)_2}]^{2+} + [\mathrm{Co^{II}((nox)_3(BC_4H_9)_2)}] \rightleftharpoons [\mathrm{Cu^{I}(bib)_2}]^{+} + [\mathrm{Co^{III}((nox)_3(BC_4H_9)_2)}]^{+} \qquad k_{12}, k_{21} \quad (2)$$

$$[\mathrm{Cu^{II}(bib)_2}]^{2+} + [\mathrm{Co^{II}((dmg)_3(BC_4H_9)_2)}] \rightleftharpoons [\mathrm{Cu^{I}(bib)_2}]^{+} + [\mathrm{Co^{III}((dmg)_3(BC_4H_9)_2)}]^{+} \qquad k_{12}, k_{21} \quad (3)$$

For these clathrochelates the ligand moiety cyclohexanedione dioximate is abbreviated as nox, while dimethylglyoximate is abbreviated as dmg. UV−vis spectral data pertaining to the reactants and products are shown in Table S1 of the Supporting Information. Reactions 1−3 occur with a loss of absorbance at the 650 nm peak for $[\mathrm{Cu^{II}(bib)_2}]^{2+}$ and at 470 nm for the Co$^{\mathrm{II}}$ complexes, the products absorbing negligibly between 400 and 800 nm. UV−vis spectral analysis of reaction 1 confirms that under certain conditions it proceeds to an equilibrium position as indicated; reactions 2 and 3 proceed to completion, although they are represented as reversible processes in order to facilitate discussion in terms of Marcus theory. Blank spectrophotometric experiments have shown that there is no net reaction of $[\mathrm{Co^{III}((nox)_3(BC_6H_5)_2)}]^{+}$ with either $[\mathrm{Cu^{II}(bib)_2}]^{2+}$ or $[\mathrm{Co^{II}((nox)_3(BC_6H_5)_2)}]$.

Electrochemistry. The cyclic voltammograms of the compounds, $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$, $[\mathrm{Co^{III/II}((nox)_3(BC_4H_9)_2)}]^{+/0}$, $[\mathrm{Co^{III/II}((nox)_3(BC_6H_5)_2)}]^{+/0}$, and $[\mathrm{Co^{III/II}((dmg)_3(BC_4H_9)_2)}]^{+/0}$ show quasi-reversible waves with $\Delta E_{\mathrm{p/p}}$ values significantly greater than 59 mV, and are in agreement with reported results.[1,4] For the Co(III/II) couples, the anodic and cathodic waves are symmetric and the $\Delta E_{\mathrm{p/p}}$ values increase with scan rate as expected for slow electrode kinetics. The $E_{1/2}$ values are independent of scan rate and thus are good measures of E_{f}. On the other hand, for $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$ the anodic wave is not very well resolved, as shown in Figure 1, and the $\Delta E_{\mathrm{p/p}}$ values do not increase with scan rate. A summary of the results is given in Table S2 of the Supporting Information.

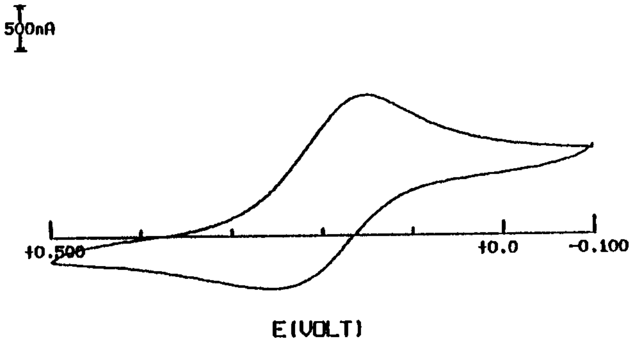

Figure 1. Cyclic voltammogram of 0.5 mM $[\mathrm{Cu^{II}(bib)_2}]^{2+}$ in acetonitrile with 0.1 M $\mathrm{Et_4NBF_4}$ at 298 K; scan range −0.1 to 0.5 V at a scan rate of 0.05 V s^{-1}.

Equilibrium Constants. Reaction 1 proceeds essentially to completion when conducted with no products present in the initial solution. On the other hand, when conducted in the presence of substantial amounts of the Co(III) product, it reaches a measurable equilibrium position. Accordingly, the value of the equilibrium quotient for reaction 1 at 25 °C in acetonitrile at $\mu = 0.1$ M was obtained by spectrophotometry at 472 nm in solutions that contained 0.25 mM Co(III) and no Cu(I) initially. For such solutions reacting as in eq 1 the absorbance is given by eq 4:

$$A/l = [\mathrm{Cu(II)}]\epsilon_{\mathrm{Cu(II)}} + [\mathrm{Co(II)}]\epsilon_{\mathrm{Co(II)}} + [\mathrm{Cu(I)}]\epsilon_{\mathrm{Cu(I)}} + [\mathrm{Co(III)}]\epsilon_{\mathrm{Co(III)}} \quad (4)$$

where A is the absorbance, l is the path length, and ϵ is the pertinent molar absorptivity. The absorbance change in the reaction is indicated as ΔA, such that

$$\Delta A = A_{\mathrm{o}} - A_{\infty} \quad (5)$$

where A_{o} and A_{∞} represent the absorbance of the reaction at the time of initiation and at equilibrium, respectively. Substitution of eq 4 into eq 5 leads to

$$\Delta A/l = \mathrm{X}(\epsilon_{\mathrm{Cu(II)}} + \epsilon_{\mathrm{Co(II)}} - \epsilon_{\mathrm{Cu(I)}} - \epsilon_{\mathrm{Co(III)}}) \quad (6)$$

where the X represents $[\mathrm{Cu(II)}]_{\mathrm{o}} - [\mathrm{Cu(II)}]_{\infty}$. Rearrangement of eq 6 and use of the spectral data in Table S1 at 472 nm with a 1.0 cm path length gives eq 7:

$$\mathrm{X} = \Delta A/(6027\ \mathrm{M}^{-1}) \quad (7)$$

In general, the equilibrium constant is given by

$$K_{12} = ([\mathrm{Cu(I)}]_{\mathrm{o}} + X)([\mathrm{Co(III)}]_{\mathrm{o}} + X)/\{([\mathrm{Cu(II)}]_{\mathrm{o}} - X)([\mathrm{Co(II)}]_{\mathrm{o}} - X)\} \quad (8)$$

Values of K_{12} were thus measured under three different sets of initial conditions as shown in Table 1. The average value of K_{12} for reaction 1 is 2.6, or Log $K = 0.41 \pm 0.16$. When the $E_{1/2}$ values given in Table S2 are used to calculate Log K, a value of −0.71 is obtained. The disagreement between these two results is attributed to the poor reversibility of the cyclic voltammetry of the $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$ redox couple. A value of 0.274 V for E_{f} for $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$ is derived from the measured equilibrium constant for reaction 1 and the $E_{1/2}$ value for the $[\mathrm{Co^{III/II}((nox)_3(BC_6H_5)_2)}]^{+/0}$ redox couple.

Equilibrium quotients for reactions 2 and 3 were calculated by the equation $-nf\Delta E° = -RT \ln K$ using the value of 0.274 V for E_{f} for $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$ and the measured $E_{1/2}$ values for the $[\mathrm{Co^{III/II}((nox)_3(BC_4H_9)_2)}]^{+/0}$ and $[\mathrm{Co^{III/II}((dmg)_3(BC_4H_9)_2)}]^{+/0}$ redox couples. These methods lead to values of K_{12} for the reactions 1, 2, and 3, listed in Table 2.

Cross-Exchange Kinetics. The electron-transfer kinetics of $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$ were studied by reducing the Cu(II) complex with a selection of suitable reducing agents. Three different

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Table 1. Spectrophotometric Equilibrium Constant for the Reduction of $[Cu^{II}(bib)_2]^{2+}$ by $[Co^{II}((nox)_3(BC_6H_5)_2)]$[a]

$[Cu^{II}(bib)_2{}^{2+}]_{init}$, mM	$[Co^{II}((nox)_3(BC_6H_5)_2)]_{init}$, mM	$[Co^{III}((nox)_3(BC_6H_5)_2)^+]_{init}$, mM	K_{12}
0.25	0.025	0.25	3.52
0.125	0.025	0.25	1.58
0.05	0.025	0.25	2.66
			K_{12}, avg = 2.56 ± 0.98

[a] At 25 °C in CH_3CN with $\mu = 0.1$ M (n-Et_4NBF_4) and $[Cu(I)]_{init} = 0.00$ mM. K_{12} calculated by use of eq 8.

Table 2. Cross Reactions between $[Cu^{II}(bib)_2](ClO_4)_2$ and Selected Counter-Reagents[a]

reducing agents	K_{12}	k_{12}, $M^{-1}\ s^{-1}$	k_{22},[d] $M^{-1}\ s^{-1}$	k_{11}, cal[e] $M^{-1}\ s^{-1}$
$[Co^{II}((nox)_3(BC_6H_5)_2)]$	2.56 ± 0.98[b]	21.7 ± 0.3	272 ± 57	0.36 ± 0.16
$[Co^{II}((nox)_3(BC_4H_9)_2)]$	72.4 ± 1.7[c]	87.4 ± 1.6	132 ± 16	0.49 ± 0.06
$[Co^{II}((dmg)_3(BC_4H_9)_2)]$	288 ± 6.6[c]	76.2 ± 1.8	546 ± 103	0.024 ± 0.004
			Log k_{11}, avg = −0.79 ± 0.72	

[a] In acetonitrile with 0.1 M Et_4NBF_4 at 25 °C. [b] Equilibrium constant by spectrophotometric measurements. [c] Equilibrium constants by the equation $-nF\Delta E° = -RT \ln K$ with $E_{1/2}$ for the Co(III/II) couples as in Table S2 and $E_f = 274$ mV for $[Cu^{II/I}(bib)_2]^{2+/+}$ as derived from the spectrophotometric equilibrium constant for reaction 1. [d] Reference 4. [e] Self-exchange rate constants by the Marcus Cross relationship as in eq 12.

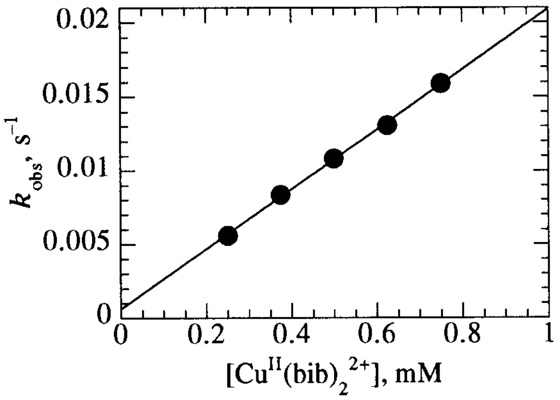

Figure 2. Plot of the pseudo-first-order rate constants, k_{obs}, as a function of $[Cu^{II}]$ for reaction 1 in acetonitrile with 0.1 M Et_4NBF_4 at 298 K.

cobalt compounds were used as reducing agents, as shown in reactions 1−3. Conditions were selected so as to drive the reactions to at least 95% completion, such that the reverse reactions could be safely neglected. The reactions were studied under pseudo-first-order conditions, with $[Cu^{II}(bib)_2]^{2+}$ in large excess, in acetonitrile with 0.1 M Et_4NBF_4 at 25 °C. Repetitive-scan studies showed the clean consumption of reactants, with no evidence for the buildup of intermediates. Under pseudo-first-order conditions the integrated rate equation is of the form

$$\ln([Co^{II}]_o/[Co^{II}]_t) = k_{obs}t \qquad (9)$$

where $[Co^{II}]_o$ and $[Co^{II}]_t$ represent the concentrations of the Co(II) complex at the time of reaction initiation and at any time, *t*, respectively. Five different concentrations of oxidizing agent were used in the kinetic determinations, with typical data as shown for reaction 1 in Figure 2. The linear dependence of k_{obs} on $[Cu(II)]_o$ as in eq 10

$$k_{obs} = k_{12}[Cu^{II}]_o \qquad (10)$$

indicates that, as expected, the reactions obey second-order kinetics conforming to the overall differential rate expression

$$-d[Co^{II}]/dt = k_{12}[Cu^{II}][Co^{II}] \qquad (11)$$

where $[Co^{II}]$ represents the concentration of the corresponding cobalt complex, and $[Cu^{II}]$ represents the concentration of the copper(II) complex. Values of k_{obs} are available in Tables S3−S5 of the Supporting Information, while the second-order rate constants are summarized in Table 2.

The temperature dependence of k_{obs} for reaction 2 was determined over the range from 7.9 to 44.3 °C with $[Cu^{II}(bib)_2{}^{2+}]_0 = 0.250$ mM and $[Co^{II}((nox)_3(BC_4H_9)_2)]_0 = 0.0250$ mM. The data are shown in terms of k_{12} as an Eyring plot in Figure S1 of the Supporting Information (data in Table S6), and the derived activation parameters are $\Delta H^\ddagger = 48.85 \pm 0.67$ kJ mol^{-1} and $\Delta S^\ddagger = -41.5 \pm 2.2$ J K^{-1} mol^{-1}.

Self-exchange rate constants for many metal complexes in acetonitrile have been summarized by Wherland.[6] These include the cobalt clathrochelates used in this work, although we have used the data as they appear in the original literature source.[4] These self-exchange rate constants are given as k_{22} in Table 2 and were used in the calculation of the copper(II/I) self-exchange rate constants, k_{11}, by iterative solution of eq 12:

$$k_{11} = (k_{12})^2/k_{22}K_{12}f_{12}(W_{12})^2 \qquad (12)$$

Here k_{12} is the second-order cross electron-transfer rate constant, K_{12} is the equilibrium constant, and k_{11} and k_{22} are the self-exchange rate constants for the component redox couples. Equation 12 is a rearranged form of the Marcus cross-relation, which is usually given as[12]

$$k_{12} = (k_{11}k_{22}K_{12}f_{12})^{1/2}W_{12} \qquad (13)$$

$$\ln f_{12} = \frac{[\ln K_{12} + (w_{12} - w_{21})/RT]^2}{4[\ln(k_{11}k_{22}/Z^2) + (w_{11} + w_{22})/RT]} \qquad (14)$$

$$W_{12} = \exp[(-w_{12} - w_{21} + w_{11} + w_{22})/2RT] \qquad (15)$$

$$w_{ij} = (Z_iZ_je^2)/(D_sa_{ij}(1 + \beta a_{ij}(\mu^{1/2}))) \qquad (16)$$

In these equations, *Z* is taken as 1×10^{11} M^{-1} s^{-1}, *a* is the center-to-center distance in angstroms when the reactants are in contact, and *R* is the gas constant (in kilocalories). β is the ion interaction parameter given by[12]

$$\beta = ((8\pi Ne^2)/(1000D_skT))^{1/2} \qquad (17)$$

where *N* is Avogadro's number 6.022×10^{23} mol^{-1}, *k* is the Boltzmann constant 1.38×10^{-23} J K^{-1}, and D_s is the static dielectric constant of the medium, 36.7 for acetonitrile at 25 °C.[13] These constants lead to the following expression for

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(13) Maryott, A. A.; Smith, E. R. *NBS Circular 514* **1951**, 1−7.

acetonitrile at 25 °C:

$$w_{\mathrm{ij}} = (9.05Z_{\mathrm{i}}Z_{\mathrm{j}})/(a_{\mathrm{ij}}(1 + 0.481a_{\mathrm{ij}}(\mu^{1/2}))) \qquad (18)$$

To calculate k_{11} for the $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$ couple by using eqs 12−17, the required value of *a* was measured by means of the CPK atomic models of these compounds. The average radius for the copper compound is 5.8 Å, and the average radii for the three cobalt compounds are approximately 7.5 Å. The calculated value W_{12} based on eqs 18 and 15 is 1.33 kcal/mol for three reactions 1, 2, and 3, while the calculated values of f_{12} based on eq 14 are 0.998, 0.920, and 0.863. The k_{11} values from the three cross-reactions 1, 2 and 3 are summarized in Table 2. They give a geometric mean of 0.16 $\mathrm{M^{-1}\ s^{-1}}$ for k_{11}.

Discussion

As indicated by the stoichiometry, the three reactions investigated here fall in the class of one-electron complementary redox reactions. A one-electron outer-sphere electron-transfer mechanism is virtually assured by the form of the rate law and the well-established substitution-inert properties of the cobalt counter-reagents. Further support for an outer-sphere mechanism is that the only potential bridging sites on the cobalt complexes are the oxygen atoms, and they are sterically protected. Given the evident lack of a bridging ligand, there is little possibility for the electronic coupling between the two reactants to be so large as to perturb the potential energy surfaces significantly. Thus, the activation enthalpy should be well represented by the intersection point of the parabolas in a traditional Marcus diagram.

Another potential concern is that the reactions might be significantly nonadiabatic. Indeed, nonadiabaticity is a feature of blue-copper protein electron transfer reactions, as has been elegantly discussed in section III.A.4 of a recent review by Holm et al.[14] The most direct probe of adiabaticity in the present system is its activation entropy, for which nonadiabaticity would make a negative contribution. Bimolecular adiabatic reactions have a characteristic activation entropy that depends on the solvent and reaction charge type, and so proper interpretation of the magnitude of $\Delta S^{\ddagger}$ is best performed by appropriate comparisons. To our knowledge, the only outer-sphere electron-transfer reactions of the 2+/0 charge type to have been studied in acetonitrile were reported by Anderson and Wherland in 1991.[5] These workers studied the reduction of $[\mathrm{Cr(CNdpp)_6}]^{2+}$ (CNdpp = 2,6-diisopropylphenyl isocyanide) by a series of cobalt(II) clathrochelates similar to those used in the present study, and their results are summarized in Table S7 of the Supporting Information. Our value of $\Delta S^{\ddagger}$ for the reaction of $[\mathrm{Cu(bib)_2}]^{2+}$ with $[\mathrm{Co((nox)_3(BC_4H_9)_2)}]$ is bracketed by the data of Anderson and Wherland. As it is believed that these reference reactions are not significantly nonadiabatic, we conclude that the present reactions of $[\mathrm{Cu(bib)_2}]^{2+}$ likewise are adiabatic. Note that the related system, $[\mathrm{Cu(bimdpk)_2}]^{2+/+}$ (bimdpk = bis(1-methyl-4,5-diphenylimidazol-2-yl)ketone, a chelating bisimidazole ligand), has a drastically different value of $\Delta S^{\ddagger}$ for its self-exchange reaction (−177 $\mathrm{J\ K^{-1}\ mol^{-1}}$);[11] this difference is in the expected direction in view of the difference in charge type (2+/+ vs 2+/0).

Given the above mechanistic considerations, reactions 1−3 conform to the model assumed in the Marcus cross relationship. The calculated values of k_{11} for $[\mathrm{Cu(bib)_2}]^{2+/+}$ shown in Table 2 display some scatter, as has been noted in other studies of reactions of the cobalt clathrochelates.[5] The problem seems to derive from uncertainties in the values of k_{22} used. Nevertheless, the data provide a reasonable estimate of k_{11} as 0.16 $\mathrm{M^{-1}\ s^{-1}}$. This result is entirely consistent with the prior inference of an upper limit of 100 $\mathrm{M^{-1}\ s^{-1}}$ for k_{11}, which was based on the lack of line broadening in an NMR study of $[\mathrm{Cu(bib)_2}]^{2+/+}$ mixtures.[1]

An appreciation of the significance of a self-exchange rate constant of only 0.16 $\mathrm{M^{-1}\ s^{-1}}$ can be gained by considering it in the context of self-exchange rate constants for other synthetic Cu(II/I) systems and also for pertinent cuproproteins. Table 3 contains a list of k_{11} values for a host of synthetic Cu(II/I) systems, from which it can be seen that the range of k_{11} values is enormous, spanning 14 orders of magnitude. The $[\mathrm{Cu(bib)_2}]^{2+/+}$ system is among the slowest, and in fact, it is the slowest when reactions involving probable changes in coordination number are excluded. A substantial number of the reactions in Table 3 proceed with significant structural changes, such as a change in coordination number, solvation number, or ligand configuration. Such reactions usually have two-step mechanisms consisting of a structural change and an electron-transfer step, a square scheme often being the outcome. Among the fastest reactions in this table are those for $[\mathrm{Cu^{II/I}(TAAB)}]^{2+/+}$, $[\mathrm{Cu^{II/I}([15]aneS_5)}]^{2+/+}$, and $[\mathrm{Cu^{II/I}(dmp)_2}]^{2+/+}$ for which such two-step mechanisms seem to apply. Thus, it is clear that the entatic hypothesis is not entirely correct in its implication that the fastest electron-transfer reactions should occur for systems having the least structural change. This conclusion is amplified by the present results for the $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$ system, which is among those systems shown in Table 3 undergoing the least structural change and yet is among the slowest!

The exact structural differences between $[\mathrm{Cu^{II}(bib)_2}]^{2+}$ and $[\mathrm{Cu^{I}(bib)_2}]^{+}$ are available from their crystal structures as reported by Knapp et al.[1] In both oxidation states the bib ligand is bidentate, conferring a quasi-square-planar geometry at Cu(II) and a pseudotetrahedral geometry at Cu(I). Both oxidation states have rigorous C_2 symmetry and approximate D_2 symmetry. The change in geometry is accompanied by significant torsional changes within the ligands and a 0.07 Å change in the Cu−N bond lengths. While these structural differences appear to be relatively modest, they are not accompanied by the compensating bond-lengthening/shortening effects seen in other systems such as $[\mathrm{Cu((py)_2DAP)}]^{2+/+}$.[15] On this basis we suggest that the low self-exchange rate constant for the $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$ system is due to a large Franck−Condon barrier rather than the lack of effective orbital coupling as suggested earlier.[1] This argument is strengthened by the adiabaticity inferred from the activation parameters as discussed above.

A fascinating comparison can be made with the results for the $[\mathrm{Cu^{II/I}(bimdpk)_2}]^{2+/+}$ system, which likewise is composed of a chelating biimidazole ligand and shows similar D_2 symmetry in both oxidation states.[11] Despite these similarities, the self-exchange rate constant for $[\mathrm{Cu^{II/I}(bimdpk)_2}]^{2+/+}$ is 5 orders of magnitude greater than for $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$. It is conceivable that some of this difference is due to greater adiabaticity in the $[\mathrm{Cu^{II/I}(bimdpk)_2}]^{2+/+}$ system arising from efficient metal−ligand superexchange pathways,[14] but the activation entropy data seem to indicate that both reactions are fully adiabatic. The change in the Cu−N bond lengths in the $[\mathrm{Cu^{II/I}(bimdpk)_2}]^{2+/+}$ system is 0.050 Å, i.e., somewhat less than for the $[\mathrm{Cu^{II/I}(bib)_2}]^{2+/+}$ system. Much more significant is the change in the intraligand bite angle, which is less than 1° for

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Table 3. Electron Self-Exchange Rate Constants of Some Synthetic $[Cu^{II/I}(L)]^{(n+1)/n}$ Couples

complex	k_{11}, $M^{-1}\,s^{-1}$	method	solvent	ref	temp, *K*
$[Cu^{II/I}(Cl_x)]^{(2-x)/(1-x)a}$	5×10^7	NMR	concd HCl_{aq}	24	–
$[Cu^{II/I}(TAAB)]^{2+/+}$	5.5×10^5	NMR	CD_3OD	25	295
$[Cu^{II/I}([15]aneS_5)]^{2+/+}$	2.0×10^5	NMR	D_2O	26	298
$[Cu^{II/I}(dmp)_2]^{2+/+\ a,b,c}$	1.8×10^5	NMR	D_2O	27	298
$[Cu^{II/I}(\textit{cis}\text{-cyhx-}[14]aneS_4)]^{2+/+\ c}$	5.0×10^4	NMR	H_2O	28	298
$[Cu^{II/I}((5\text{-}MeimidH)_2DAP)]^{2+/+}$	3.87×10^4	NMR	CD_3CN	29	293
$[Cu^{II/I}(bimdpk)_2]^{2+/+}$	1.9×10^4	NMR	CD_3CN	11	298
$[Cu^{II/I}((imidH)_2DAP)]^{2+/+}$	1.31×10^4	NMR	CD_3CN	30	298
$[Cu^{II/I}([14]aneS_4)]^{2+/+}$	7.8×10^3	Marcus	CD_3CN	31	298
$[Cu^{II/I}([14]aneS_4)]^{2+/+}$	7.6×10^3	NMR	D_2O	32	298
$[Cu^{II/I}(dmp)_2]^{2+/+\ b}$	6.2×10^3	NMR	CD_3CN	27	298
$[Cu^{II/I}([14]aneS_4)]^{2+/+}$	6.0×10^3	Marcus	H_2O	32	298
$[Cu^{II/I}(dpym)_2]^{0/-}$	5.9×10^3	NMR	acetone-d_6	33	298
$[Cu^{II/I}(bidhp)]^{2+/+}$	4×10^3	NMR	DMSO-d_6	34	301
$[Cu^{II/I}(dmp)_2]^{2+/+\ b}$	3.4×10^3	NMR	acetone-d_6	27	298
$[Cu^{II/I}((py)_2DAP)]^{2+/+}$	1.76×10^3	NMR	CD_3CN	15	298
$[Cu^{II/I}(me_2phenyl_2phen)_2]^{2+/+\ b}$	1.8×10^3	NMR	CD_3CN	35	298
$[Cu^{II/I}(bite)]^{2+/+}$	21.7^d	NMR	acetone-d_6	36	295
$[Cu^{II/I}(pdao)]^{2+/+}$	8.1	Marcus	H_2O	37	298
$[Cu^{II/I}W_{12}O_{40}]^{6-/7-}$	2.61	Marcus	H_2O	38	298
$[Cu^{II/I}(pdto)]^{2+/+\ c,e}$	0.63	Marcus	H_2O/CH_3OH	37	298
$[Cu^{II/I}(bib)_2]^{2+/+}$	0.16	Marcus	CD_3CN	f	298
$[Cu^{II/I}(\mathit{aq})]^{2+/+}$	5×10^{-7}	Marcus	H_2O	39	298
$[Cu^{II/I}(Im)_n]^{2+/+}$	1×10^{-7}	Marcus	H_2O	39	298

[a] May involve an inner sphere mechanism due to the presence of chloride. [b] The reported values in units of kg $mol^{-1}\,s^{-1}$ were converted to $M^{-1}\,s^{-1}$ by assuming solvent and solution densities to be equal. The actual values reported were for $[Cu^{I/II}(dmp)_2]^{+/2+}$ in D_2O, $k = 2.0\ (0.2) \times 10^5$ kg $mol^{-1}\,s^{-1}$, in CD_3CN, $k = 4.9\ (0.2) \times 10^3$ kg $mol^{-1}\,s^{-1}$, and in acetone-d_6, $k = 3.0\ (0.2) \times 10^3$ kg $mol^{-1}\,s^{-1}$; for $[Cu^{II/I}(me_2phenyl_2phen)_2]^{2+/+}$ $k = 1.4 \times 10^3$ kg $mol^{-1}\,s^{-1}$. [c] Loss of a water molecule coordinated to Cu(II) likely accompanies reduction.[40] [d] First-order rate constant in units of s^{-1}. [e] Reactions in acetonitrile with ferrocenes show rate saturation indicative of prior association.[41] [f] This work.

$[Cu^{II/I}(bimdpk)_2]^{2+/+}$ and 20° for $[Cu^{II/I}(bib)_2]^{2+/+}$. These considerations lend further support to the importance of Franck–Condon factors in these reactions.

According to the general concepts in Marcus theory, the origin of the unusually low self-exchange rate constant for the $[Cu(bib)_2]^{2+/+}$ system should lie in the solvent reorganization energy, $\Delta G^{\ddagger}_{solv}$, or in the internal reorganization energy, $\Delta G^{\ddagger}_{in}$. The theoretically predicted inverse dependence of $\Delta G^{\ddagger}_{solv}$ on molecular size is well-supported by experiments in aqueous media,[16] but in nonaqueous media the evidence is not as firm.[6,17] Nevertheless, since $[Cu(bib)_2]^{n+}$ is significantly smaller than $[Cu(bimdpk)_2]^{n+}$ it is reasonable to attribute some of the activation barrier to an increased contribution from $\Delta G^{\ddagger}_{solv}$. On the other hand, several of the systems in Table 3 that have very large self-exchange rate constants are comparable in size to $[Cu(bib)_2]^{n+}$. Clearly then, the contribution from the internal reorganization energy must also be considered.

A semiquantitative estimate of the internal reorganization energies can be made by use of molecular mechanics. In prior reports, molecular mechanics has been used to estimate $\Delta G^{\ddagger}_{in}$ for octahedral cobalt complexes where a change in the metal–ligand bond length corresponds to the reaction coordinate.[18–20] The approach taken has been to locate the transition state and calculate its energy. In one example this was done by defining the transition state as the point of lowest energy on a surface consisting of the sum of the relative strain energies of the two reactants distorted to a common geometry.[18] In the other cases the transition state was erroneously defined as the crossing point of the absolute strain energy curves for the reduced and oxidized reactants;[19,20] this definition can correctly yield a transition state in which electron transfer is isoenergetic, but it does not generally yield the lowest-energy point satisfying this criterion. Moreover, this reactant surface-crossing method fails entirely when the curves do not cross. In the present work it is clear that the reaction coordinate has a major component involving angular distortions, although the precise mode is difficult to specify. Because of these complications, we have taken a different approach to calculating $\Delta G^{\ddagger}_{in}$. First, we developed preliminary force fields for copper(I) and copper(II) that yield satisfactory equilibrium structures for $[Cu(bib)_2]^{n+}$ and $[Cu(bimdpk)_2]^{n+}$ in both oxidation states. The calculated equilibrium strain energies ($E_{str,eq}$) are given in Table S12 of the Supporting Information. Second, we calculated the (unoptimized) strain energies ($E_{str,vert}$) of these exact structures when the oxidation state of the copper atoms is reversed, as also given in Table S12 of the Supporting Information. Third, we calculated λ_{in}, the vertical internal reorganization energy, as $E_{str,vert}(Cu(I)) + E_{str,vert}(Cu(II)) - E_{str,eq}(Cu(I)) - E_{str,eq}(Cu(II))$. We then evaluated $\Delta G^{\ddagger}_{in}$ as $\lambda_{in}/4$, which is standard Marcus theory in the harmonic approximation. The result is a value of 12.3 kcal/mol for $\Delta G^{\ddagger}_{in}$ for the $[Cu(bib)_2]^{2+/+}$ system and a value of 7.4 kcal/mol for the $[Cu(bimdpk)_2]^{2+/+}$ system. If it is assumed that the ratio of the values of k_{11} for the two systems is due to a difference in $\Delta G^{\ddagger}$, the two rate constants yield a value of $\Delta\Delta G^{\ddagger}$ of 6.9 kcal mol^{-1} while the value of $\Delta\Delta G^{\ddagger}_{in}$ from molecular mechanics is 4.9 kcal mol^{-1}. It is to be expected that $\Delta G^{\ddagger}_{solv}$ is somewhat larger for the $[Cu(bib)_2]^{2+/+}$ system than for the $[Cu(bimdpk)_2]^{2+/+}$ system because of the smaller size of the former. The magnitude of this difference can be estimated from the equation $\Delta G^{\ddagger}_{solv} = \{(\Delta Z)^2 e^2/(8r^{\ddagger})\}(1/\epsilon_{op} - 1/\epsilon)$, which reduces to $\Delta G^{\ddagger}_{solv} = 21.5/r^{\ddagger}$ kcal/mol in acetonitrile at 25 °C. As mentioned above the value of $r^{\ddagger}$ is 5.8 Å for $[Cu(bib)_2]^{n+}$, and from the work of McMaster et al. we obtain a corresponding value of 9.16 Å for $[Cu(bimdpk)_2]^{n+}$. These data lead to a value

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Table 4. Electron Self-Exchange Rate Constants of Protein $[Cu^{II/I}(L)]^{(n+1)/n}$ Couples

complex	k_{11}, M^{-1} s^{-1}	method	solvent	ref	temp, *K*
[Cu$^{II/I}$(azurin(*P. a.*))]	2.4×10^6	EPR	H_2O (pH 5.0)	42	298
[Cu$^{II/I}$(azurin(*P. a.*))]	9.6×10^5	NMR	D_2O (pH 5.0)	43	298
[Cu$^{II/I}$(azurin(*P. a.*))]	7.0×10^5	NMR	D_2O (pH 9.0)	43	298
[Cu$^{II/I}$(plastocyanin(*A. v.*))]	5.9×10^5	Marcus	H_2O (pH 7.5)	44	298
[Cu$^{II/I}$(azurin(*A. d.*))]	4.0×10^5	NMR	D_2O (pH 6.7)	45	297
[Cu$^{II/I}$(plastocyanin(*A. v.*))]	3.2×10^5	NMR	D_2O (pH 7.5)	46	298
[Cu$^{II/I}$(amicyanin(*T. v.*))]	1.3×10^5	NMR	D_2O (pH 7.5)	47	298
[Cu$^{II/I}$(stellacyanin)]	1.2×10^5	EPR	H_2O (pH 7.0)	48	293
[Cu$^{II/I}$(plastocyanin(*French bean*)]	$\ll 2 \times 10^4$	NMR	D_2O (pH 7.4)	49	323
[Cu$^{II/I}$(rusticyanin(*T. f.*))]	$\sim 10^4$	NMR	D_2O (pH 5.7)	50	298
[Cu$^{II/I}$(M121Q azurin)]	8×10^3	NMR	D_2O (pH 7.0)	51	298
[Cu$^{II/I}$(umecyanin)]	6.1×10^3	NMR	D_2O (pH 7.5)	52	298
[Cu$^{II/I}$(plastocyanin(*Spinach*))]	$\sim 4 \times 10^3$	NMR	D_2O (pH 6.0)	53	298
[Cu$^{II/I}$(plastocyanin(*parsley*))]	3.3×10^3	Marcus	H_2O (pH 7.5)	44	298
[Cu$^{II/I}$(pseudoazurin(*A. c.*))]	2.9×10^3	NMR	D_2O (pH 7.5)	54	298
[Cu$^{II/I}$(pseudoazurin(*A. c.*))]	2.7×10^3	Marcus	H_2O (pH 7.5)	54	298

of 1.4 kcal/mol for $\Delta\Delta G^{\ddagger}_{\text{solv}}$, which leads to essentially perfect agreement between theory and the experimental kinetics results. This fine agreement between experiment and theory provides strong support for the concept that much of the difference in k_{11} for these two systems arises from a difference in the internal reorganization energy.

A listing of self-exchange rate constants for cuproproteins is given in Table 4. The rate constants range from 2.7×10^3 to 2.4×10^6 M^{-1} s^{-1} and hence are several orders of magnitude greater than for $[Cu^{II/I}(bib)_2]^{2+/+}$. Structural data for these cuproproteins indicate that they undergo electron transfer with no major structural changes at copper, and hence share this property with $[Cu^{II/I}(bib)_2]^{2+/+}$. One should be careful, however, not to overstate the lack of structural change as cuproproteins change oxidation state. X-ray crystallography shows that plastocyanin, for example, experiences a 0.21 Å change in the Cu−N(His37) bond length and L−Cu−L angular changes as large as 9° between the two oxidation states.[21,22] For these reasons, $[Cu^{II/I}(bib)_2]^{2+/+}$ at present is among the best dynamic small-molecule model compounds for the active site of blue cuproproteins. It shows that a wide range of internal reorganization energies are available at CNI copper redox sites, and it underscores the need to develop improved methods to evaluate these reorganization energies from structural data.

Overall, it appears that copper centers display the anticipated correlation between structural change and reorganization energy when the structural changes are quite limited, as in the $[Cu(bib)_2]^{2+/+}$ and $[Cu(bimdpk)_2]^{2+/+}$ systems. However, the structural flexibility of copper often leads to more drastic structural changes where such correlations are not found; examples include changes in coordination number or ligand configuration, which can lead to two-step mechanisms, associated square schemes, and rates that can appear anomalously rapid. Given the relatively large structural changes found for blue copper proteins,[21,22] it is not yet clear whether their low reorganization energies[23] arise from a combination of the "small" structural changes and low force constants or from low-energy conformational modes as could occur in a square scheme.

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Conclusions. (1) The low derived self-exchange rate constant for $[Cu^{II/I}(bib)_2]^{2+/+}$ is consistent with the failure to detect line broadening in the prior ^{1}H NMR study. (2) The derived self-exchange rate constant is the lowest yet recorded for Cu(II/I) systems, excluding those having complications such as changes in coordination number or a square scheme. (3) An explanation in terms of nonadiabaticity (weak orbital overlap) is not supported by the activation entropy, which is consistent with an adiabatic reaction. (4) Molecular mechanics calculations show that the geometry changes between Cu(II) (pseudo-square planar) and Cu(I) (pseudo tetrahedral) lead to an internal reorganization energy that is the major component of the Franck−Condon barrier and thus explain the slow kinetics. (5) The million-fold lower self-exchange rate constant for $[Cu^{II/I}(bib)_2]^{2+/+}$ relative to similar systems having qualitatively drastic structural changes raises questions as to the utility of the entatic hypothesis for blue copper proteins.

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Supporting Information Available: Tables of electronic spectra and electrochemical data for the compounds studied. Tables of pseudo-first-order rate constants for reactions 1−3 at 25 °C and for various temperatures for reaction 2. A table of activation parameters for various electron-transfer reactions of 2+/0 charge type. Molecular mechanics descriptions for $[Cu(bib)_2]^{2+/+}$ and $[Cu(bimdpk)_2]^{2+/+}$ (PDB format). Strain energies from molecular mechanics calculations for the four equilibrium states and the four vertical states. A figure showing the Eyring plot for reaction 2 (21 pages). Ordering information is given on any current masthead page.

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