$$A/A_0 = \alpha + \beta = e^{-k_1 t} + e^{-k_2 t} \quad (15)$$

Thus the activity does *not* decay by a simple first-order process, unless k_1 and k_2 are very different (which they seem to be), and this was the circumstance originally considered.[6,7] However, a *different* sum of two exponentials applies if the intermediate is fluxional. The exponents are complicated functions of both k_1 and k_2,[18] but for example, when $k_1 \gg k_2$ the expression is

$$A/A_0 = \alpha + \beta = (-2/13)e^{-4k_1 t} + (28/13)e^{-3k_1 t/4} \quad (16)$$

Here the loss of radiolabel is governed essentially by the dominant second term, and all five CO groups would be *observed* to exchange with the same specific rate. Provided $k_1 > k_2$,[13] then irrespective of the relative magnitudes of k_1 and k_2, the two exponents for the case of a fluxional intermediate never differ greatly.[13,18]

Clearly the kinetics are affected by the question of a fluxional intermediate, and the facts[6,7] indicate that the intermediate is rigid since two very different specific exchange rates are observed. This is inconsistent with the conclusions of Brown et al.,[11,12] drawn from the less clear-cut IR data, and it must be concluded that the problem of the lifetime of the intermediate remains unresolved. What is now clear, however, is that its lifetime does affect the rate laws for exchange, whether studied by the IR technique or by standard radiotracer methods, and it has been possible using the PMR to derive the exact rate laws for the two sets of circumstances.

Registry No. CO, 630-08-0.

(18) The exact expressions are as follows:[13]

$$\alpha = [k_1(k_1 - k_2)/(13k_1^2 + 6k_1k_2 + k_2^2)] \exp(-(4k_1 + k_2)t) + [(4k_1 + k_2)(3k_1 + k_2)/(13k_1^2 + 6k_1k_2 + k_2^2)] \exp(-(3k_1 + 2k_2)t/(4k_1 + k_2))$$

$$\beta = [(3k_1 + k_2)(k_2 - k_1)/(13k_1^2 + 6k_1k_2 + k_2^2)] \exp(-(4k_1 + k_2)t) + [4k_1(4k_1 + k_2)/(13k_1^2 + 6k_1k_2 + k_2^2)] \exp((-k_1(3k_1 + 2k_2)t/(4k_1 + k_2))$$

Contribution from the Department of Chemistry,
The University of Texas at Arlington, Arlington, Texas 76019

Acid-Induced Aquation of (Trifluoromethyl)cobaloximes with Aromatic Nitrogenous Axial Ligands

Kenneth L. Brown* and Te-Fang Yang

Received February 10, 1987

(Trifluoromethyl)(ligand)cobaloximes, $CF_3Co(D_2H_2)L$ (L = pyridine, 1-methylimidazole, and 1,5,6-trimethylbenzimidazole), have been prepared and characterized by ^1H and ^{19}F NMR spectroscopy. The kinetics of axial ligand dissociation from the pyridine complex have been studied as a function of acidity in sulfuric acid/water mixtures. The first equatorial oxime protonation is found to increase the rate of pyridine dissociation by 3 orders of magnitude while the second equatorial oxime protonation increased the rate of pyridine dissociation only 10-fold. A complete analysis of the relative free energies of all of the ground and transition states involved shows that the extraordinary effect of the first equatorial oxime protonation on the kinetics of pyridine dissociation is due to a significant stabilization of the transition state by this protonation. It is postulated that the transition states for ligand dissociation from the cationic complexes are substantially stabilized by intramolecular proton transfer from the protonated, equatorial oxime functionalities to the departing axial ligand. Kinetic measurements of the rate of axial ligand dissociation from the 1-methylimidazole complex in sulfuric acid/water mixtures suggest, but do not prove, that this ligand dissociation is specific-acid catalyzed via preequilibrium protonation of the axial ligand at N-1.

Introduction

Many studies of the kinetics and mechanism of axial ligand substitution in cobaloximes have appeared.[1-33] However, the effect of equatorial oxime protonation[34-38] on such processes has never been investigated. Organocobaloximes have been shown to un-

(1) Hague, D. N.; Halpern, J. *Inorg. Chem.* **1967**, *6*, 2059–2063.
(2) Tsiang, H. G.; Wilmarth, W. K. *Inorg. Chem.* **1968**, *7*, 2535–2542.
(3) Birk, J. P.; Chock, P. B.; Halpern, J. *J. Am. Chem. Soc.* **1968**, *90*, 6959–6963.
(4) Costa, G.; Tauzher, G.; Puxeddu, A. *Inorg. Chim. Acta* **1969**, *3*, 41–44.
(5) Ludwick, L. M.; Brown, T. L. *J. Am. Chem. Soc.* **1969**, *91*, 5188–5189.
(6) Crumbliss, A. L.; Wilmarth, W. K. *J. Am. Chem. Soc.* **1970**, *92*, 2593–2594.
(7) Sakurai, T.; Fox, J. P.; Ingraham, L. L. *Inorg. Chem.* **1971**, *10*, 1105–1106.
(8) Brown, T. L.; Ludwick, L. M.; Stewart, R. S. *J. Am. Chem. Soc.* **1972**, *94*, 384–388.
(9) Herlinger, A. W.; Brown, T. L. *J. Am. Chem. Soc.* **1972**, *94*, 388–393.
(10) Brown, K. L.; Kallen, R. G. *J. Am. Chem. Soc.* **1972**, *94*, 1894–1901.
(11) Brown, K. L.; Chernoff, D.; Keljo, D. J.; Kallen, R. G. *J. Am. Chem. Soc.* **1972**, *94*, 6697–6704.
(12) Earley, J. E.; Zimmerman, J. G. *Inorg. Nucl. Chem. Lett.* **1972**, *8*, 687–688.
(13) Guschl, R. J.; Brown, T. L. *Inorg. Chem.* **1973**, *12*, 2815–2819.
(14) Courtright, R. L.; Drago, R. S.; Nusz, J. A.; Nozari, M. S. *Inorg. Chem.* **1973**, *12*, 2809–2815.
(15) Trogler, W. C.; Stewart, R. C.; Marzilli, L. G. *J. Am. Chem. Soc.* **1974**, *96*, 3697–3699.
(16) Guschl, R. J.; Stewart, R. S.; Brown, T. L. *Inorg. Chem.* **1974**, *13*, 417–422.
(17) Espenson, J. H.; Russell, R. *Inorg. Chem.* **1974**, *13*, 7–11.
(18) Jensen, F. R.; Kiskis, R. C. *J. Am. Chem. Soc.* **1975**, *97*, 5820–5825. However, see ref 22.
(19) Brown, K. L.; Lyles, D.; Pencovici, M.; Kallen, R. G. *J. Am. Chem. Soc.* **1975**, *97*, 7338–7346.
(20) Seibles, L.; Deutsch, E. *Inorg. Chem.* **1977**, *16*, 2273–2278.
(21) Brown, K. L.; Autrey, A. W. *Inorg. Chem.* **1978**, *17*, 111–119.
(22) Stewart, R. C.; Marzilli, L. G. *J. Am. Chem. Soc.* **1978**, *100*, 817–822.
(23) Toscano, P. J.; Marzilli, L. G. *Inorg. Chem.* **1979**, *18*, 421–424.
(24) Lawrence, G. A.; Surachiltanout, S. *Inorg. Chim. Acta* **1980**, *44*, L61–L62.
(25) Bresciani-Pahor, N.; Randaccio, L.; Toscano, P. G.; Sandercock, A. C.; Marzilli, L. G. *J. Chem. Soc., Dalton Trans.* **1982**, 129–134.
(26) Dreos-Garlatti, R.; Tauzher, G.; Costa, G. *Inorg. Chim. Acta* **1983**, *70*, 83–86.
(27) Dreos-Garlatti, R.; Tauzher, G.; Costa, G. *Inorg. Chim. Acta* **1983**, *71*, 9–13.
(28) Dreos-Garlatti, R.; Tauzher, G.; Costa, G. *Inorg. Chim. Acta* **1984**, *82*, 197–200.
(29) Randaccio, L.; Bresciani-Pahor, N.; Orbell, J. D.; Calligaris, M.; Summers, M. F.; Snyder, B.; Toscano, P. J.; Marzilli, L. G. *Organometallics* **1985**, *4*, 469–478.
(30) Bresciani-Pahor, N.; Randaccio, L.; Zangrando, E.; Toscano, P. J. *Inorg. Chim. Acta* **1985**, *96*, 193–198.
(31) Bresciani-Pahor, N.; Randaccio, L.; Zangrando, E.; Summers, M. F.; Ramsden, J. H.; Marzilli, P. A.; Marzilli, L. G. *Organometallics* **1985**, *4*, 2086–2090.
(32) Bresciani-Pahor, N.; Forcolin, M.; Marzilli, L. G.; Randaccio, L.; Summers, M. F.; Toscano, P. J. *Coord. Chem. Rev.* **1985**, *63*, 1–125.
(33) Zangrando, E.; Bresciani-Pahor, N.; Randaccio, L.; Charland, J. P.; Marzilli, L. G. *Organometallics* **1986**, *5*, 1938–1944.
(34) Adin, A.; Espenson, J. H. *J. Chem. Soc., Chem. Commun.* **1971**, 653–654.
(35) Abley, P.; Dockal, E. R.; Halpern, J. *J. Am. Chem. Soc.* **1972**, *95*, 3166–3170.
(36) Crumbliss, A. L.; Bowman, J. T.; Gaus, P. L.; McPhail, A. T. *J. Chem. Soc., Chem. Commun.* **1973**, 415–416.
(37) Crumbliss, A. L.; Gaus, P. L. *Inorg. Chem.* **1975**, *14*, 486–490.
(38) Brown, K. L.; Lu, L.-Y. *Inorg. Chem.* **1981**, *20*, 4178–4183.

0020-1669/87/1326-3007$01.50/0

dergo such protonations with pK_a values of 0.67–0.00[34,35] for the first oxime protonation while only a single value for the pK_a for the second oxime protonation has been reported (−4.63, for the dicationic σ-bonded ethylcobaloxime carbonium ion intermediate in the acid-induced decomposition of (2-hydroxyethyl)- and (2-alkoxyethyl)cobaloximes[39].)

There are several reasons knowledge of the effect of equatorial oxime protonation on ligand-exchange processes of alkylcobaloximes is of interest. First, understanding of the nature of both the equatorial ligand–metal bonding and the bonding between such organocobalt centers and aromatic nitrogenous axial ligands (such as pyridine) is not sufficiently advanced to permit a priori prediction of either (a) the influence of equatorial oxime protonation on the electrophilicity of organocobaloxime cobalt centers or (b) the effect of changes in cobalt center electrophilicity on the energetics of both the ground state of organocobaloximes with such axial ligands and the transition state for ligand dissociation. It is hence not readily apparent how equatorial ligand protonation will affect the kinetics of the dissociation of pyridine, for instance, from organo(pyridine)cobaloximes. Second, such information is of practical significance. For various reasons, it is often more convenient to synthesize organocobaloximes with strongly donating axial ligands such as pyridine since for many interesting cases the organo aquo complexes are difficult to obtain. When such organo(pyridine)cobaloximes are used to study organocobalt decomposition mechanisms in acidic or basic aqueous environments, it is often necessary to assume that pyridine dissociation is rapid compared to carbon–cobalt bond-breaking processes[40,41] in order to simplify kinetic interpretations. Justification for such assumptions is usually based on known rate constants for the dissociation of pyridine from neutral organo(pyridine)cobaloximes.[19] However, while equatorial oxime proton dissociation in aqueous base is known to increase the rate of pyridine dissociation from such complexes,[19] nothing is known of the effect of equatorial oxime protonation in acid. Finally, Reenstra and Jencks[42] have studied the reaction of cyanide with aquocobalamin and cyanocobalamin. From distinct differences in the kinetic pattern of the displacement of the axial dimethylbenzimidazole ligand of cyanocobalamin by cyanide in basic and acidic solution, these authors suggested (with reservations based on contradictory results from others[43]) that dissociation of dimethylbenzimidazole from cyanocobalamin is specific-acid-catalyzed via preequilibrium protonation of the axial dimethylbenzimidazole nucleotide at N-1 (i.e., the atom involved in the N-glycosidic bond to ribose). However, in our studies of the ^{1}H-decoupled ^{31}P NMR spectra of cyanocobalamin[44] and aquocobalamin[45] in sulfuric acid/water mixtures, well-resolved independent resonances for the base-on and base-off species of both of these cobalamins were always observed at all acidities. Consequently the exchange between the base-on and base-off species is slow (on the NMR time scale) even at very high acidities. If this exchange were subject to specific-acid catalysis, it would be expected to become sufficiently fast, at some acidity, to collapse the ^{31}P resonances. It is thus not clear if specific-acid catalysis plays a role in dimethylbenzimidazole dissociation from cobalamins or not, although it would certainly be expected to do so at some acidity. It is thus of interest to model this process in the structurally more simple cobaloximes. However, in order to be able to understand the influence of acidity on the kinetics of dissociation of axial ligands with protonatable functionalities in addition to the liganding atom (such as 1,5,6-trimethylbenzimidazole (Me_3BzIm) or 1-methylimidazole (MeIm)), it is first necessary to understand the effect of equatorial oxime protonation on the dissociation of simpler ligands, such as pyridine (py).

Since preliminary measurements of the effect of acidity on the rate of pyridine dissociation from organo(pyridine)cobaloximes showed that equatorial protonation in fact greatly increases the rate of ligand dissociation, it was necessary to find an acid-stable organo(pyridine)cobaloxime with an organic group that would be both stable and nonprotonatable in very strong acid but would be sufficiently weakly electron donating[19,21] so that pyridine dissociation would remain slow enough to be measured at all acidities. (Trifluoromethyl)(pyridine)cobaloxime proved to be an excellent candidate. However, this material is not obtainable by standard reductive alkylation routes as it undergoes reductive defluorination in the presence of borohydride[46,47] (and probably other suitable reducing agents as well[48]). We consequently employed, with moderate success, the method of Widdowson,[49,50] which does not require reducing agents capable of generating cobalt(I) species. We now report what is, as far as we know, the first preparation of pure (trifluoromethyl)(ligand)cobaloximes and our study of the influence of acidity on the dissociation of axial pyridine and 1-methylimidazole from such species.

Experimental Section

Materials. Organic solvents and reagents and inorganic reagents were obtained in the highest purity commercially available and used without further purification. Zinc wool was freshened immediately before use (to remove surface oxide) by swirling with 4.0 NHCl, followed by rinsing with water and drying. Glass-distilled water was used throughout.

(Trifluoromethyl)(pyridine)cobaloxime, $CF_3Co(D_2H_2)py$, was synthesized at 0.025-mol scale (cobalt) by the method of Roussi and Widdowson[49,50] with benzene as the solvent and a reaction temperature of 70 °C. After introduction of the zinc wool, CF_3I was purged into the solution for 10 h. The reaction mixture was filtered and the filtrate evaporated to dryness. The solid was dissolved in chloroform, and the chloroform solution was washed with water and dried with $MgSO_4$. Evaporation provided 8.53 g of crude product, which was purified by silica gel chromatography with ethyl acetate as eluent. The purified product was recrystallized from methanol–water: yield 3.18 g (28%). Due to the difficulty of separating (trihalomethyl)(pyridine)cobaloximes from halo(pyridine)cobaloximes,[51] this material is not pure but is contaminated with $ICo(D_2H_2)py$ (see below). Pure $CF_3Co(D_2H_2)py$ is best obtained by conversion of this material to the mixture of aquo complexes and separation by chromatography, followed by addition of pyridine to pure $CF_3Co(D_2H_2)OH_2$.

$CF_3Co(D_2H_2)OH_2$ was prepared as follows. Purified $CF_3Co(D_2H_2)py$ (as above, 2.0 g, 4.6 mmol) was dissolved in 750 mL of methanol in the dark with stirring and warming. A 7.5-mL portion of settled AG50W-X8 cation-exchange resin (Bio-Rad), H^+ form, was then added and the mixture was stirred in the dark at room temperature for 11 days. The ion-exchange resin was removed by filtration and washed several times (until the washings were colorless) with methanol. The combined washings and supernatant were evaporated to dryness, and the solid was dissolved in 150 mL of methanol. Purification was effected by chromatography, in three 50-mL batches, on a 36 × 3 cm column of silica gel (in chloroform) with acetone as eluent. The purified material was recrystallized from acetone/water: yield 0.84 g (49%). ^{1}H NMR ($CDCl_3/CD_3OD$): δ_{Me_4Si} 2.326 (s). ^{19}F NMR ($CDCl_3/CD_3OD$): $\delta_{C_6H_5F}$ 86.656 (s).

Pure $CF_3Co(D_2H_2)py$ was then obtained from $CF_3Co(D_2H_2)OH_2$ as follows. $CF_3Co(D_2H_2)OH_2$ (0.2 g, 0.53 mmol) was dissolved in 50 mL of methanol in the dark, and pyridine (1.01 g, 12.3 mmol) was added. After the mixture was stirred 1 h in the dark, 50 mL of water was added and the solution was concentrated to 25 mL on a rotary evaporator. Another 25 mL of water was added, and the solution was again concentrated to 25 mL. After the solution was cooled in ice for 30 min,

(39) Brown, K. L.; Ramamurthy, S.; Maynick, D. S. *J. Organomet. Chem.* **1985**, *287*, 377–394.
(40) Espenson, J. H.; Wang, D. M. *Inorg. Chem.* **1979**, *18*, 2853–2859.
(41) Mock, W. L.; Bieniarz, C. *Organometallics* **1984**, *3*, 1279–1284.
(42) Reenstra, W. W.; Jencks, W. P. *J. Am. Chem. Soc.* **1979**, *101*, 5780–5791.
(43) Tacconi, N. R.; Lexa, D.; Saveant, J. M. *J. Am. Chem. Soc.* **1979**, *101*, 467–473.
(44) Brown, K. L.; Hakimi, J. M. *Inorg. Chem.* **1984**, *23*, 1756–1764.
(45) Brown, K. L.; Hakimi, J. M.; Jacobsen, D. W. *J. Am. Chem. Soc.* **1984**, *106*, 7894–7899.
(46) Ricroch, M.-N.; Bied-Charreton, C.; Gaudemer, A. *Tetrahedron Lett.* **1971**, 2859–2862.
(47) Ricroch, M.-N. Doctoral Dissertation, The University of Paris at Orsay, 1971.
(48) Brown, K. L.; Hakimi, J. M.; Nuss, D. M.; Montejano, Y. D.; Jacobsen, D. W. *Inorg. Chem.* **1984**, *23*, 1463–1471.
(49) Roussi, P. F.; Widdowson, D. A. *J. Chem. Soc., Chem. Commun.* **1979**, 810–812.
(50) Roussi, P. F.; Widdowson, D. A. *J. Chem. Soc., Perkin Trans. 1* **1982**, 1025–1028.
(51) Espenson, J. H.; McDowell, M. S. *Organometallics* **1982**, *1*, 1514–1518.
(52) Simonov, A. M.; Pozharskii, A. E.; Marianovskii, V. M. *Indian J. Chem.* **1967**, *5*, 81–82.

filtration provided 0.12 g (52%) of yellow $CF_3Co(D_2H_2)py$. ^1H NMR ($CDCl_3$): δ_{Me_4Si} 2.204 (S, 12 H), 7.387 (m, 2 H), 7.773 (m, 1 H), 8.539 (m, 2 H). ^{19}F NMR ($CDCl_3$): $\delta_{C_6H_5F}$ 81.281 (s).

The 1-methylimidazole (MeIm) and 1,5,6-trimethylbenzimadazole (Me_3BzIm, obtained by a method analogous to that for 1-methylbenzimidazole[50]) complexes were prepared analogously.

$CF_3Co(D_2H_2)MeIm$: yield 64%. ^1H NMR ($CDCl_3$): δ_{Me_4Si} 2.200 (s, 12 H), 3.641 (s, 3 H), 6.767 (s, 1 H), 6.947 (s, 1 H), 7.456 (s, 1 H). ^{19}F NMR ($CDCl_3$): $\delta_{C_6H_5F}$ 79.678 (s).

$CF_3Co(D_2H_2)Me_3Bz$: yield 65%. ^1H NMR ($CDCl_3$): δ_{Me_4Si} 2.173 (s, 12 H), 2.356 (s, 3 H), 2.382 (s, 3 H), 3.745 (s, 3 H), 7.080 (s, 1 H), 7.243 (s, 1 H), 7.978 (s, 1 H).

Methods. All work with the cobaloximes was performed in dim light, and solutions were covered with aluminum foil whenever possible. NMR spectra were obtained in a Nicolet NT 200 wide-bore superconducting NMR spectrometer operating at 200.067 MHz (^1H) or 188.238 MHz (^{19}F). UV–visible spectral measurements were made on a Cary 219 recording spectrophotometer whose cell block was maintained at 25.0 ± 0.1 °C (as measured directly in a dummy cell with a thermistor device (Yellow Springs Instruments)) by a circulating water bath. pH measurements were made on a Radiometer PHM 64 pH meter with samples and standards thermostated to 25.0 ± 0.1 °C.

Ligand dissociation rates were measured by diluting stock solutions of the appropriate $CF_3Co(D_2H_2)L$ complex into cuvettes containing sulfuric acid/water mixtures (which had incubated for 30 min at 25 °C) and monitoring the absorbance change at appropriate UV wavelengths (217 nm for L = py, 220 nm for L = MeIm). Rate constants were obtained from the slopes of semilogarithmic plots of $A_t - A_\infty$ vs. time or, in the case of reactions for which the half-time exceeded 24 h, from the slopes of Guggenheim plots.[53]

Molar acidities of kinetic samples were obtained by titration of duplicate aliquots (100–500 μL, by use of quantitative micro transfer pipets) with standard KOH to a phenolphthalein end point. All such duplicate titrations agreed to within 1%. Acidity function values were obtained from the determined molarities from literature data,[54] with interpolation between data points as necessary.

Ligand association rates for $CF_3Co(D_2H_2)OH_2$ and pyridine were measured by diluting stock solutions of $CF_3Co(D_2H_2)OH_2$ into solutions containing various concentrations of pyridine in water in 10 cm path length, water-jacketed cells and monitoring the decrease in absorbance at 410 nm. Pseudo-first-order rate constants were obtained from the slopes of semilogarithmic plots of $A_t - A_\infty$ vs. time, and the second-order rate constant, k_{on}, was obtained from the slope of a plot of the pseudo-first-order rate constants vs. pyridine concentration by the method of least squares. 1-Methylimidazole was titrated potentiometrically at ionic strength 1.0 M (KCl).

Results and Discussion

Synthesis and Characterization of (Trifluoromethyl)cobaloximes. (Trifluoromethyl)cobalt complexes are not obtainable by standard reductive alkylation procedures as they undergo reductive defluorination in the presence of reducing agents sufficiently strong to produce the needed cobalt(I) reagent. In the case of cobalamins, reductive alkylation with CF_3I produces a mixture of the CF_3 and CF_2H derivatives, which may be separated by chromatography on a reversed-phase absorbant.[48] Treatment of the purified CF_3 derivative with such reducing agents also leads to the CF_2H compound (as well as gives some C–Co bond cleavage). However, Gaudemer and co-workers[46,47] have shown that reductive alkylation of cobaloximes with CF_3I produces mixtures of CF_3, CF_2H, CFH_2, and even CH_3 cobaloximes. This problem is clearly avoided by using the Widdowson method,[49,50] since the $CF_3Co(D_2H_2)py$ obtained had only a singlet ^{19}F NMR resonance at 81.281 ppm (compared to 81.728 ppm for (trifluoromethyl)cobalamin[48]) and conspicuously lacked a doublet resonance near 18 ppm ((difluoromethyl)cobalamin, 18.261 ppm, J_{H-F} = 53.2 Hz[48]). However, owing to the difficulty of chromatographically separating (trihalomethyl)(pyridine)cobaloximes from halo(pyridine)cobaloximes,[51] $CF_3Co(D_2H_2)py$ obtained by the Widdowson method was about 13% contaminated with $ICo(D_2H_2)py$ (δ_{Me_4Si} 2.380 (s)). If the single silica gel chromatography band is cut into two roughly equal fractions, the first fraction contains only about 8% $ICo(D_2H_2)py$ while the second fraction contains 19%. Thus, purification by repeated chromatography is possible but would be wasteful and tedious. It is far more convenient to convert the mixture of materials to the aquo complexes by treatment with an acidic ion-exchange resin, during which process $ICo(D_2H_2)py$ is converted to aquohydroxocobaloxime. As the latter is readily separated from $CF_3Co(D_2H_2)OH_2$ chromatographically, pure $CF_3Co(D_2H_2)py$ is easily obtainable by addition of pyridine to the aquo complex. Unfortunately the 11-day reaction period for conversion of the pyridine to the aquo complex is mandated by the slow rate of pyridine dissociation from neutral $CF_3Co(D_2H_2)py$ (see below).

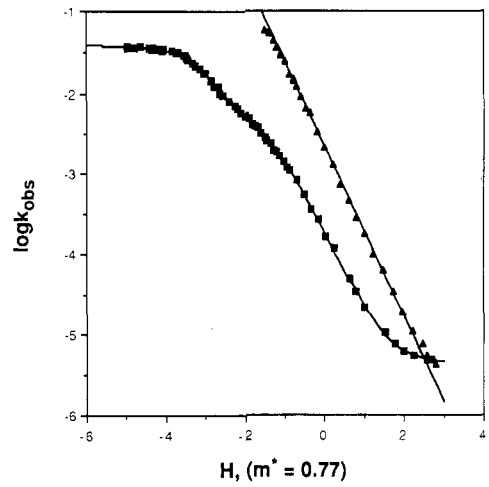
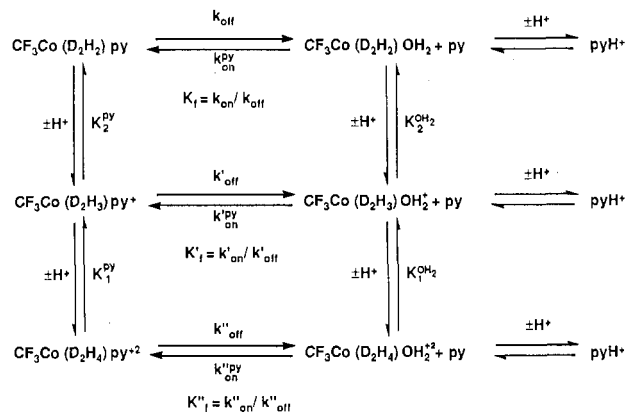


Figure 1. Plots of log k_{obsd} for ligand dissociation from $CF_3Co(D_2H_2)L$ in aqueous sulfuric acid vs. H, the generalized acidity function (eq 1[52]), at m^* = 0.77. L = py (■) the solid line was calculated from eq 2 by using the best-fit parameters listed in Table II. For L = MeIm (▲) the solid line is a least-squares fit for the data from H = 1.94 to H = −0.80, with slope 1.05 ± 0.008, intercept −2.685 ± 0.007, and r^2 = 0.999.

Scheme I



Kinetics of Pyridine Dissociation. Figure 1 shows the complete acidity function–rate profile for the dissociation of pyridine from $CF_3Co(D_2H_2)py$ in aqueous sulfuric acid. Acidity has been expressed by the generalized acidity function of Cox and Yates[54] (eq 1), where C_{H^+} is the concentration of hydrogen ion, X is the

$$-H = m^*X + \log C_{H^+} \quad (1)$$

so-called "excess acidity" of sulfuric acid, and m^* is an adjustable parameter characteristic of the weak base that is being protonated and indicative of the fact that all known acidity functions are linearly related to one another.[54-56] For the current work, the value of m^* has been set to 0.77 by a method to be described presently.

The H-rate profile (Figure 1) clearly shows the influence of the two anticipated equatorial oxime protonations. The complete reaction scheme for the dissociation of pyridine from CF_3Co-

(53) Guggenheim, E. A. *Philos. Mag.* **1926**, *2*, 538–543.
(54) Cox, R. A.; Yates, K. *J. Am. Chem. Soc.* **1978**, *100*, 3861–3867.
(55) Bunnett, J. F.; Olsen, F. P. *Can. J. Chem.* **1964**, *44*, 1899–1916.
(56) Morziano, N. C.; Cimino, G. M.; Passerini, R. C. *J. Chem. Soc., Perkin Trans. 2* **1973**, 1915–1922.

Table I. Best-Fit Parameters from the Fit of the Rate Data for Pyridine Dissociation from $CF_3Co(D_2H_2)py$ (25 ± 0.1 °C) to Eq 2, at $m^* = 0.77$ (Eq 1)

param	value	param	value
$G_1^{py}G_2^{py}k_{off}$	1.58×10^{-1} M^2 s^{-1}	$G_1^{py}G_2^{py}$	3.60×10^4 M^2
$G_1^{py}k'_{off}$	6.55 M s^{-1}	G_1^{py}	1.60×10^3 M
k''_{off}	3.79×10^{-2} s^{-1}		

Table II. Rate and Equilibrium Constants for the Dissociation of Pyridine from $CF_3Co(D_2H_2)py$ (25.0 ± 0.1 °C, Scheme I)

constant	value	constant	value
k_{off}	4.41×10^{-6} s^{-1}	K_f	3.06×10^3 M^{-1}
k'_{off}	4.09×10^{-3} s^{-1}	K_f'	1.57×10^3 M^{-1}
k''_{off}	3.79×10^{-2} s^{-1}	K_f''	3.13×10^3 M^{-1}
pK_1^{py}	−2.90	$pK_1^{OH_2}$	−3.20
pK_2^{py}	−1.65	$pK_2^{OH_2}$	−1.36
k_{on}	1.35×10^{-2} M^{-1} s^{-1}		
k'_{on}	6.42 M^{-1} s^{-1}		
k''_{on}	1.19×10^2 M^{-1} s^{-1}		

$(D_2H_2)py$ is thus as shown in Scheme I. Application of the law of mass action leads directly to the rate law of eq 2, where k_{obsd}

$$k_{obsd} = \frac{\frac{G_1^{py}G_2^{py}k_{off}}{a_{H^+}^2} + \frac{G_1^{py}k'_{off}}{a_{H^+}} + k''_{off}}{\frac{G_1^{py}G_2^{py}}{a_{H^+}^2} + \frac{G_1^{py}}{a_{H^+}} + 1} \quad (2)$$

is the rate constant for pyridine dissociation at a given acidity, a_{H^+} is hydrogen ion activity (calculated as 10^{-H}, eq 1), and the macroscopic proton dissociation constants, G_1^{py} and G_2^{py}, are related to the microscopic proton dissociation constants (Scheme I) as $G_1^{py} = 2K_1^{py}$ and $G_2^{py} = K_2^{py}/2$, where the factors of 2 are statistical factors due to the fact that the two ionizations are from identical ionizable groups. In order to determine the appropriate value of m^* (eq 1) for this system, the rate data (Figure 1) were fitted to eq 2 via an iterative, nonlinear least-squares program using a simplex minimization routine and an arbitrarily selected value of m^*. This procedure was then repeated with various values of m^* by using the sums of the squares of the deviations of the calculated rate constants from the observed rate constants (i.e., χ^2) as a "goodness of fit" criterion. χ^2 was found to go through a sharp minimum at $m^* = 0.77$. The best-fit parameters at this value of m^* are listed in Table I, and the solid line in Figure 1 for these data has been calculated from eq 1 by using these parameters. The kinetic and equilibrium constants that can be calculated from these results (i.e., k_{off}, k'_{off}, k''_{off}, pK_1^{py}, and pK_2^{py}) are listed in Table II (first five entries). It is readily apparent that the first equatorial oxime protonation increases the rate of pyridine dissociation by nearly 3 orders of magnitude while the second protonation increases the rate by only a factor of about 10. While such detailed measurements have not been made with other organo(pyridine)cobaloximes, it is very clear from our preliminary measurements as well as our recent work with (2-aryl-2-oxoethyl)(pyridine)cobaloximes[57] that large increases in the rate of pyridine dissociation upon equatorial protonation are characteristic of such complexes.

In order to obtain the remainder of the rate and equilibrium constants in Scheme I, it is necessary to have values of K_f, the formation constant for the neutral pyridine species, and $K_1^{OH_2}$ and $K_2^{OH_2}$, the microscopic proton dissociation constants for the aquo complex. The last two were determined by direct spectrophotometric titration of $CF_3Co(D_2H_2)OH_2$ at 4.00×10^{-5} M and 215 nm, the wavelength of maximal spectral change determined from preliminary scanning experiments at various acidities. The data were fitted to eq 3, where A_{obsd} is the observed absorbance at 215 nm of solutions of 4.00×10^{-5} M $CF_3Co(D_2H_2)OH_2$ at various acidities, A, A^+, and A^{2+} are the absorbances at 215 nm of 4.00

$$A_{obsd} = \frac{\frac{G_1^{OH_2}G_2^{OH_2}A}{a_{H^+}^2} + \frac{G_1^{OH_2}A^+}{a_{H^+}} + A^{2+}}{\frac{G_1^{OH_2}G_2^{OH_2}}{a_{H^+}^2} + \frac{G_1^{OH_2}}{a_{H^+}} + 1} \quad (3)$$

$\times 10^{-5}$ M neutral, monocationic, and dicationic (trifluoromethyl)aquocobaloxime, respectively, and the macroscopic proton dissociation constants ($G_1^{OH_2}$ and $G_2^{OH_2}$) are again related to the microscopic proton dissociation constants (Scheme I) as $G_2^{OH_2} = 2K_1^{OH_2}$ and $G_2^{OH_2} = K_2^{OH_2}/2$. From $m^* = 0.77$, i.e., the same acidity function previously found appropriate for the equatorial oxime protonations of the pyridine complex, the values of $pG_1^{OH_2}$ and $pG_2^{OH_2}$ were −3.50 and −1.06, respectively. The resulting values of $pK_1^{OH_2}$ and $pK_2^{OH_2}$ are listed in Table II. Surprisingly, the effect of substituting pyridine for water as the axial ligand is of the same magnitude on both ionizations but of opposite sign; i.e., this ligand substitution raises pK_1 by 0.3 unit but lowers pK_2 by essentially the same amount.

The value of $pK_2^{OH_2}$ (−1.36) is substantially below the range of values (0.67–0.0) previously reported for organoaquocobaloximes.[34,35] This is apparently indicative of a substantial electron inductive effect of the organic ligand on the upper axial ligand pK_a, as the previously reported values are for organocobaloximes with electron donating (propyl, ethyl, methyl) or slightly electron withdrawing (benzyl, phenyl), organic ligands while the CF_3 group is powerfully electron withdrawing. In fact, an excellent linear correlation was found (not shown) of $pK_2^{OH_2}$ with the Taft inductive substituent constant, σ^*[58] (slope −0.724 ± 0.019, intercept 0.533 ± 0.019, $r^2 = 0.996$), confirming the importance of the inductive properties of the organic ligand on the equatorial oxime pK_a's and providing an interesting example of the cis effect in such complexes.

Because of the very low rate constant for pyridine dissociation from the neutral species $CF_3Co(D_2H_2)py$ ($t_{1/2}$ = 43.7 h), direct measurement of K_f (Scheme I) by measurement of the amount of pyridine complex formed at equilibrium from the aquo complex in various concentrations of pyridine was not feasible. Consequently, k_{on} (Scheme I), the second-order rate constant for pyridine ligation to $CF_3Co(D_2H_2)OH_2$, was determined from the slope of a plot of the observed pseudo-first-order rate constants for pyridine ligation vs. the molar pyridine concentration. From the resulting value (Table II, $k_{on} = 1.35 \times 10^{-2}$ M^{-1} s^{-1}), K_f was calculated to be 3.06×10^3 M^{-1} from $K_f = k_{on}/k_{off}$. These values then permit calculation of the remaining equilibrium contents from eq 4 and 5 based on the cyclic nature of the equilibria, and the remaining

$$K_f' = K_2^{OH_2}K_f/K_2^{py} \quad (4)$$

$$K_f'' = K_1^{OH_2}K_f'/K_1^{py} \quad (5)$$

rate constants could be calculated from the equilibrium relationships $k'_{on} = k'_{off}K_f'$ and $k''_{on} = k''_{off}K_f''$. The values thus calculated are listed in Table II. Interestingly, equatorial oxime protonation has very little effect on the equilibrium constant for substitution of water by pyridine, there being no trend and little variation in the values of K_f, K_f', and K_f''.

Now that all of the rate and equilibrium constants for Scheme I have been obtained, it becomes possible, with use of transition-state theory, to calculate the relative free energies for all of the ground and transition states in Scheme I. For this purpose, the ground state of the neutral aquo complex was arbitrarily assigned an absolute free energy of 0.00 kcal mol^{-1} and all of the other free energies were calculated from this value. The results are given in Table III and shown graphically as reaction coordinate diagrams in Figure 2. The first equatorial oxime protonation of $CF_3Co(D_2H_2)py$ increases the ground-state free energy by 2.25 kcal mol^{-1} but actually stabilizes the transition state for pyridine dissociation by 1.80 kcal mol^{-1}. Consequently, the ΔG^* value for

(57) Wo, Y.-Y. P.; Brown, K. L. *J. Organomet. Chem.*, in press.

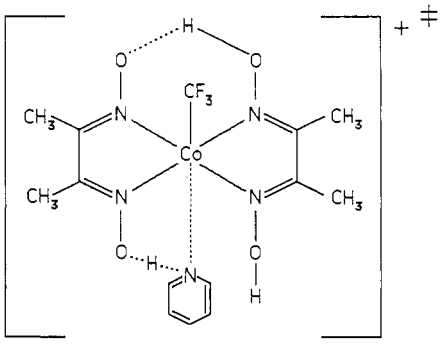
(58) Taft, R. W. In *Steric Effects in Organic Chemistry*; Newman, M. S., Ed.; Wiley: New York, 1956; Chapter 13.

Table III. Relative Gibbs Free Energies of the Ground and Transition States for the System Described in Scheme I[a]

complex	L = py $G°$, kcal mol^{-1}	transition state $G^‡$, kcal mol^{-1}	L = OH_2 $G°$, kcal mol^{-1}
$CF_3Co(D_2H_2)L$	−4.75	20.00	0.00
$CF_3Co(D_2H_3)L^+$	−2.50	18.20	1.85
$CF_3Co(D_2H_4)L^{2+}$	1.45	20.80	6.21

[a] Calculated from the rate and equilibrium constants in Table II, after arbitrary assignment of a value of 0.00 kcal mol^{-1} to the neutral $CF_3Co(D_2H_2)OH_2$ complex.

pyridine dissociation is decreased by 4.05 kcal mol^{-1} upon the first equatorial protonation and the dissociation rate of pyridine from $CF_3Co(D_2H_3)py^+$ is 927 times faster than from $CF_3Co(D_2H_2)py$. On the other hand, the second equatorial oxime protonation raises the ground state by an additional 3.95 kcal mol^{-1}, but now the transition state is destabilized by 2.60 kcal mol^{-1}. Thus the $\Delta G^‡$ value is decreased by 1.35 kcal mol^{-1}, and the reaction is only 9.3 times faster. Consequently, the surprisingly large increase in pyridine dissociation rate caused by the first equatorial oxime protonation is due to a combination of destabilization of the ground state and stabilization of the transition state, while the much more reasonable influence of the second equatorial oxime protonation on the kinetics of pyridine dissociation results from simultaneous destabilization of the ground and transition states, although, of course, the transition state is destabilized somewhat less than the ground state, as would be expected. If we consider the effect of the second equatorial protonation to be the "normal", expected effect, then it is clear that what is "abnormal" about the effect of the first protonation is the significant stabilization of the transition state that it causes. This may indicate that in the transition state for pyridine dissociation from the cationic species there is significant intramolecular proton transfer from the equatorial ligand to the departing pyridine ligand (as depicted in structure I) which cannot occur in the transition state of the neutral complex as the bridged hydrogens in the neutral complexes as many, many orders of magnitude less acidic.[3,19]



I

Kinetics of 1-Methylimidazole Dissociation. In order to attempt to gain direct evidence regarding the possibility of specific-acid catalysis of the dissociation of ligands with protontable functionalities other than the liganding atom, we have attempted measurement of the kinetics of acid-induced ligand dissociation from $CF_3Co(D_2H_2)Me_3BzIm$ and $CF_3Co(D_2H_2)MeIm$. Unfortunately, for the Me_3BzIm complex the spectral changes upon ligand dissociation were very small and preliminary measurements at low acidity showed that the rate of dissociation from the neutral complex was at least 1 order of magnitude faster than that of pyridine dissociation from $CF_3Co(D_2H_2)py$. However, a much wider range of measurements was possible with the MeIm complex, the results of which are shown in Figure 1, plotted against the generalized acidity function (eq 1) with $m^* = 0.77$, i.e., the acidity function found to be appropriate for cobaloxime equatorial oxime protonation from the kinetics of pyridine dissociation. The range of measurements possible was limited by the kinetics, from $H = 2.8$ ($t_{1/2} = 64$ h) to $H = -1.5$ ($t_{1/2} = 11.6$ s). Most of the data appear to lie on a straight line with a slight tendency to level

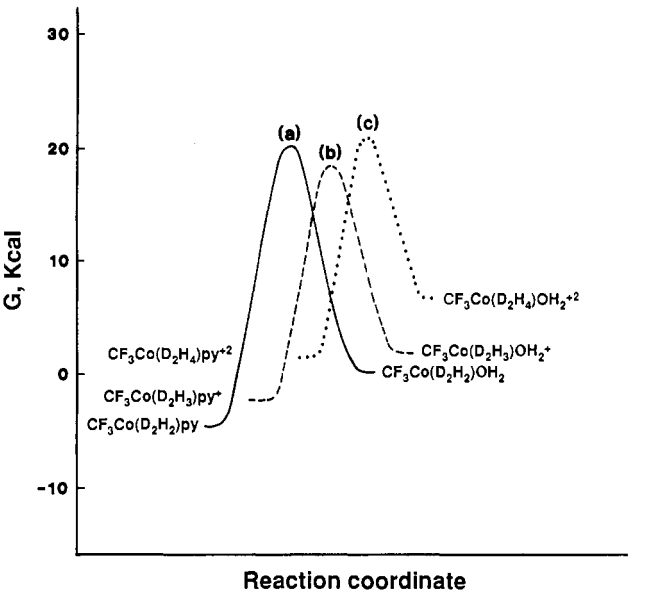
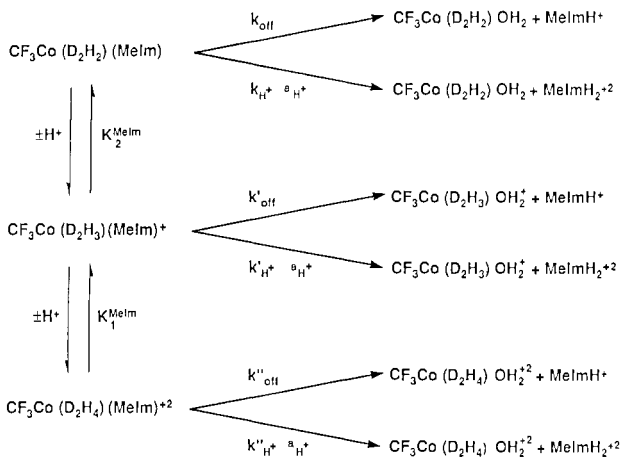


Figure 2. Free energy–reaction coordinate diagrams for dissociation of pyridine from (trifluoromethyl)(pyridine)cobaloxime in sulfuric acid/water mixtures (Scheme I): (a) neutral complexes; (b) monocationic complexes; (c) dicationic complexes. The relative free energies were calculated from the rate and equilibrium constants in Table II, after arbitrary assignment of a value of 0.00 kcal mol^{-1} to the neutral $CF_3Co(D_2H_2)OH_2$ complex.

Scheme II



off at low acidity. Unfortunately, the sluggishness of the reaction at low acidities prevented determination of k_{off} (for the neutral species), which is evidently lower than k_{off} for pyridine dissociation. At acidities higher than $H = -0.75$ there appears to be the beginning of a slight downward inflection. However, the linear portion of the data has a slope that exceeds 1.0, apparently statistically significantly ($N = 15$, slope 1.05 ± 0.008, $r^2 = 0.999$). This is clearly incompatible with the simple scheme for pyridine dissociation from $CF_3Co(D_2H_2)py$ (Scheme I). The alternative mechanism, involving specific acid-catalyzed 1-methylimidazole dissociation, is shown in Scheme II. The H–rate profile for such a mechanism would be expected to show an acidity-independent region at low acidity (i.e., k_{off}) followed by a first-order dependence on acidity (when plotted vs. an acidity function appropriate for MeIm protonation at N-1) punctuated by two inflections at pK_2^{MeIm} and pK_1^{MeIm}. As protonation of the equatorial ligand system must surely be expected to decrease the extent of protonation of the imidazole ligand at N-1, the inflections would be expected to be downward. Clearly, the data in Figure 1 for ligand dissociation from $CF_3Co(D_2H_2)MeIm$ resemble a portion of such an H–rate profile, i.e., up to the beginning of the first downward deflection (i.e., pK_2^{MeIm}). If this interpretation is correct, the value of pK_2^{MeIm} would be somewhat below −1.0, a perfectly reasonable

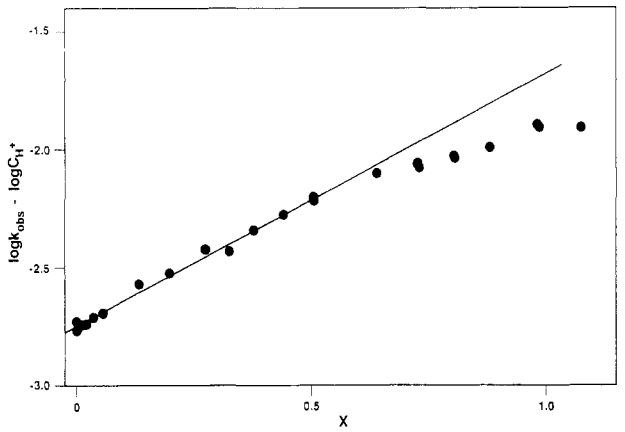


Figure 3. Plot of log k_{obsd} - log C_{H^+} vs. X (eq 8) for ligand dissociation from $CF_3Co(D_2H_2)MeIm$. The solid line is a least-squares fit to the data between $X = 0.0013$ and $X = 0.639$, with slope 1.07 ± 0.03, intercept -2.75 ± 0.01, and $r^2 = 0.992$.

value considering the fact that pK_2^{py} is -1.65 and 1-methylimidazole ($pK_a = 7.29$) is nearly 2 orders of magnitude more basic than pyridine ($pK_a = 5.51$[19]). In addition, if this interpretation (i.e., Scheme II) is correct, the region of the H-rate profile up to the inflection (i.e., $H >$ ca -0.75) would be governed by the simple rate law of eq 6. It is consequently possible to use the

$$k_{obsd} = k_{off} + k_{H^+}a_{H^+} \quad (6)$$

Cox and Yates generalized acidity function (eq 1) to determine what acidity function (i.e., what value of m^*) is appropriate for k_{H^+}. Expressing hydrogen ion activity via this acidity function leads to the rate law of eq 7, and for all acidities at which $k_{off} \ll k_{H^+}a_{H^+}$ eq 8 will hold. Hence, a plot of log k_{obsd} - log C_{H^+} vs.

$$k_{obsd} = k_{off} + k_{H^+}10^{m^*X+\log C_{H^+}} \quad (7)$$

$$\log k_{obsd} = \log k_{H^+} + m^*X + \log C_{H^+} \quad (8)$$

X should be linear with slope m^* and intercept log k_{H^+}. Such a plot is shown in Figure 3 and can be seen to be satisfactorily linear between $X = 0.0013$ ($H = 1.94$ at $m^* = 0.77$) and $X = 0.639$ ($H = -0.80$ at $m^* = 0.77$). At higher acidities the plot curves downward, confirming the downward inflection in the H-rate profile at $H < -0.75$. From the linear portion of the plot $k_{H^+} = 1.77 \times 10^{-3}$ M^{-1} s^{-1} and $m^* = 1.07 \pm 0.03$. Disappointingly, attempts to titrate the lower protonation of MeIm spectrophotometrically in order to determine the appropriate acidity function (i.e., m^* value) and hence to see if k_{H^+} follows the same acidity function as N-1 protonation of cationic $MeImH^+$ were unsuccessful due to the negligibly small spectral changes accompanying N-1 protonation of the $MeImH^+$ cation. However, the protonation of many nitrogen weak bases has been studied. The substituted-aniline indicator system originally used by Hammett and Dyrup to establish the H_0 acidity function and related compounds usually regarded as Hammett indicators follows the generalized acidity function with $m^* = 1.02 \pm 0.08$.[54] In addition, one of us and others[48] previously found the spectral changes accompanying N-1 protonation of α-ribazole (1-α-D-ribofuranosyl-5,6-dimethylbenzimidazole) to follow the H_0 acidity function. It thus seems reasonable to suggest that the kinetics of MeIm dissociation follow the Hammett acidity function, H_0, and that MeIm dissociation from $CF_3Co(D_2H_2)MeIm$ is indeed subject to specific-acid catalysis via protonation at N-1. However, given the limitations of the data collectible for this system, this must be considered to remain an open question until an appropriate experimental system can be devised to provide a definitive answer.

Acknowledgment. This research was supported by the Robert A. Welch Foundation, Houston, TX (Grant No. Y-749).

Contribution from the Departments of Chemistry, Wayne State University, Detroit, Michigan 48202, and University of Wisconsin at Eau Claire, Eau Claire, Wisconsin 54701

Structure-Reactivity Relationships in Copper(II)/Copper(I) Electron-Transfer Kinetics: Evaluation of Self-Exchange Rate Constants for Copper-Polythia Ether Complexes

Michael J. Martin,[1a] John F. Endicott,[1a] L. A. Ochrymowycz,[1b] and D. B. Rorabacher*[1a]

Received June 25, 1986

Kinetic parameters are reported for electron-transfer cross-reactions of eight Cu(II)/Cu(I) systems involving closely related open-chain and macrocyclic polythia ether ligands. Both the oxidation kinetics of the Cu(I) species and the reduction kinetics of the Cu(II) species are included by using tris(4,7-dimethyl-1,10-phenanthroline)iron(III) and diaquo(2,3,9,10-tetramethyl-1,4,8,11-tetraazacyclotetradeca-1,3,8,10-tetraene)cobalt(II) as the principal cross-reagents. Reactions with other cross-reagents are also reported for selected Cu(II)/Cu(I) systems. The applicability of the Marcus cross-relation to these reactions is examined and is shown to yield calculated Cu(II)/Cu(I) self-exchange rate constants that differ significantly for the corresponding oxidation and reduction reactions. To account for these apparent discrepancies, a dual-pathway mechanism is proposed in which a major part of the conformational reorganization at the copper center occurs sequentially, rather than concertedly, with the electron-transfer step. Differences in the relative kinetic behavior of the various copper-polythia ether complexes are discussed in terms of the influence of ligand constraints upon the bond-making and bond-breaking sequences that accompany the conversion of Cu(II) to Cu(I) (and vice versa).

Introduction

The prevalence of the Cu(II)/Cu(I) redox couple in enzymes involved in biological oxidation-reduction processes[2,3] has stimulated a high level of interest in the mechanistic details associated with electron transfer at a copper center. Specific attention has been focused on the kinetics of the blue electron carriers (azurin, plastocyanin, rusticyanin, stellacyanin), which contain a single copper atom and appear to exhibit relatively large self-exchange rate constants (eq 1) on the order of $k_{11} = 10^4$-10^6 M^{-1} s^{-1}.[4-19]

$$*Cu^{II} + Cu^{I} \underset{k_{11}}{\rightleftharpoons} *Cu^{I} + Cu^{II} \quad (1)$$

(1) (a) Wayne State University. (b) University of Wisconsin at Eau Claire.
(2) *Copper Coordination Chemistry: Biochemical and Inorganic Perspectives*; Karlin, K. D., Zubieta, J., Eds.; Adenine: Guilderland, NY, 1983.
(3) (a) *Met. Ions Biol. Syst.* **1981**, *13*. (b) Owen, C. A., Jr. *Biological Aspects of Copper*; Noyes: Park Ridge, NJ, 1982. (c) *Copper Proteins and Copper Enzymes*; Lontie, R., Ed.; CRC: Boca Raton, FL, 1984; Vols. I-III.
(4) (a) Holwerda, R. A.; Wherland, S.; Gray, H. B. *Annu. Rev. Biophys. Bioeng.* **1976**, *5*, 363-396 and references therein. (b) Cummins, D.; Gray, H. B. *J. Am. Chem. Soc.* **1977**, *99*, 5158-5167. (c) Holwerda, R. A.; Knaff, D. B.; Gray, H. B.; Clemmer, J. D.; Crowley, R. A.; Smith, J. M.; Mauk, A. G. *Ibid.* **1980**, *102*, 1142-1146.
(5) Farver, O.; Pecht, I. In *Copper Proteins*; Spiro, T. G., Ed.; Wiley-Interscience: New York, 1981; pp 151-192 and references therein.

0020-1669/87/1326-3012$01.50/0