Published in final edited form as:

Inorganica Chim Acta. 2008 March 3; 361(4): 1070–1078. doi:10.1016/j.ica.2007.07.038.

Comparison of Structurally–Related Alkoxide, Amine, and Thiolate–Ligated M^{II} (M= Fe, Co) Complexes: the Influence of Thiolates on the Properties of Biologically Relevant Metal Complexes

Lisa M. Brines, Gloria Villar-Acevedo, Terutaka Kitagawa, Rodney D. Swartz, Priscilla Lugo-Mas, Werner Kaminsky[§], Jason B. Benedict[§], and Julie A. Kovacs^{*}

The Department of Chemistry, University of Washington: Box 351700 Seattle, WA 98195-1700

Abstract

Mechanistic pathways of metalloenzymes are controlled by the metal ion's electronic and magnetic properties, which are tuned by the coordinated ligands. The functional advantage gained by incorporating cysteinates into the active site of non-heme iron enzymes such as superoxide reductase (SOR) is not entirely understood. Herein we compare the structural and redox properties of a series of structurally–related thiolate, alkoxide, and amine–ligated Fe(II) complexes in order to determine how the thiolate influences properties critical to function. Thiolates are shown to reduce metal ion Lewis acidity relative to alkoxides and amines, and have a strong *trans* influence thereby helping to maintain an open coordination site. Comparison of the redox potentials of the structurally analogous compounds described herein indicates that alkoxide ligands favor the higher-valent Fe³⁺ oxidation state, amine ligands favor the reduced Fe²⁺ oxidation state, and thiolates fall somewhere in between. These properties provide a functional advantage for substrate reducing enzymes in that they provide a site at the metal ion for substrate to bind, and a moderate potential that facilitates both substrate reduction, and regeneration of the catalytically active reduced state. Redox potentials for structurally–related Co(II) complexes are shown to be cathodically–shifted relative to their Fe(II) analogues, making them ineffective reducing agents for substrates such as superoxide.

Introduction

Cysteinate-ligated metalloenzymes and metalloproteins promote a number of critical biological processes.^{1–16} Mechanistic pathways are controlled by the metal ion's electronic and magnetic properties, which are tuned by the coordinated ligands, and subtly altered by H–bonding interactions within the protein. Supporting ligands play an important role in determining function by controlling a metal-ion's redox potential, as well as the basicity of bound substrates. Our group has put considerable effort into determining how coordinated cysteinate residues might control the chemistry taking place at the active–site of a metalloenzyme. Cysteinates are known to form highly covalent bonds to transition–metals^{17,18} and facilitate electron transfer reactions.^{19,20} Our work has shown that thiolate ligands make low–spin iron accessible even in a non–heme environment,^{21–23} stabilize iron

^{*} Author to whom correspondents should be addressed: kovacs@chem.washington.edu.

[§] UW staff crystallographers

Publisher's Disclaimer: This is a PDF file of an unedited manuscript that has been accepted for publication. As a service to our customers we are providing this early version of the manuscript. The manuscript will undergo copyediting, typesetting, and review of the resulting proof before it is published in its final citable form. Please note that during the production process errors may be discovered which could affect the content, and all legal disclaimers that apply to the journal pertain.

in the +3 oxidation state,^{21–23} and labilize sites *trans* to the thiolate thereby promoting reactivity by releasing product,²⁴ even with low-spin Co(III) which is typically inert.^{25,26} Recently, we showed that the thiolate ligand of a biomimetic superoxide reductase (SOR) model $[\text{Fe}(\text{II})(\text{cyclam-PrS})]^+$, and its *trans* positioning relative to the substrate binding site, contribute significantly to the catalyst's function by dramatically lowering the Fe–O stretching frequency ($\nu_{\text{Fe-O}} = 419 \text{ cm}^{-1}$) and force constant ($k = 1.20 \text{ mdynes/cm}$) well below that of any other reported iron-peroxo thereby favoring Fe–O bond cleavage and H_2O_2 release.²⁴ Superoxide reductase, peptide deformylase, and nitrile hydratase occupy a unique niche amongst known non-heme enzymes in that they contain cysteinate ligands, as opposed to the more frequently encountered nitrogen and oxygen based ligands. The functional advantage gained by incorporating cysteinates is not entirely understood. By systematically replacing a single atom or functional group within identical ligand frameworks and comparing the structural and redox properties of the resulting iron complexes, we hope to address this question. For this purpose, we synthesized alkoxide and amine analogues of our functional SOR models $[\text{Fe}^{\text{II}}(\text{S}^{\text{Me}_2}\text{N}_4(\text{tren}))]^+$ (**1**; *tren* = *tris*(2-aminoethyl)amine)) and $[\text{Fe}^{\text{II}}(\text{cyclam-PrS})]^+$ (**5**). Herein we examine and compare the structures, and electronic, magnetic, and redox properties of this series of structurally related complexes. We also compare primary versus secondary versus tertiary amines, and structurally analogous complexes containing Co(II) versus Fe(II) to observe how the metal ion, amine basicity, and steric bulk influences key properties critical to metalloenzyme function.

Experimental

General Methods

All reactions were performed using standard Schlenk techniques under an atmosphere of dinitrogen. The iron complexes $[\text{Fe}^{\text{II}}(\text{S}^{\text{Me}_2}\text{N}_4(\text{tren}))](\text{PF}_6)$ (**1**),²⁷ $[\text{Fe}^{\text{II}}(\text{O}^{\text{Me}_2}\text{N}_4(\text{tren})(\text{H}_2\text{O}))](\text{PF}_6)$ (**2**),²⁸ $[\text{Fe}^{\text{II}}(\text{cyclam-PrS})](\text{BPh}_4)$ (**5**),²⁴ and $[\text{Fe}^{\text{II}}(\text{S}^{\text{Me}_2}\text{N}_4(\text{tren-Et}_4))]^+$ (**8**),²⁹ as well as 3-Boc-cyclam and cyclam-PrSAc•4HCl,²⁴ were synthesized as previously described. All other reagents were obtained from Aldrich and used without further purification. Solvents were purified through standard procedures.³⁰ 3-methyl-3-mercapto-2-butanone was prepared according to a published procedure.²³ NMR spectra were recorded using Bruker AV-300, AV-301, or DPX 500 spectrometers and referenced to the residual protio solvent. ¹H NMR chemical shifts (δ) are reported in parts per million (ppm) and coupling constants (J) are reported in Hz. IR spectra were recorded on a Perkin-Elmer 1720 FT-IR. Electrospray mass spectra were recorded on a Bruker Esquire LC-Ion Trap Mass Spectrometer. EPR spectra were obtained using a Varian CW-EPR spectrometer at 4 K equipped with an Oxford helium cryostat. Cyclic voltammograms were recorded on a PAR 263A potentiostat with a glassy carbon working electrode, a platinum counter electrode, and a Ag/Ag⁺ (0.1 M) in acetonitrile as the reference electrode. Potentials were then converted to values referenced versus SCE. Electronic absorption spectra were recorded using a Hewlett-Packard 8453 diode array spectrometer. Magnetic moments were determined using a Quantum Design SQUID Magnetometer and by Evans' method as modified for a superconducting solenoid.^{31,32} Temperatures were obtained using Van Geet's method.³³ Elemental analyses were performed by Galbraith Laboratories (Knoxville, TN) and Atlantic Microlab, Inc (Norcross, GA).

Synthesis of $[\text{Fe}^{\text{II}}(\text{N}^{\text{Et}_2})\text{N}_4(\text{tren})\text{Cl}](\text{PF}_6)$ (**3**)

A solution of *tren* (0.584 g, 4 mmol) was dissolved in 10 ml of dry methanol, followed by the addition of diethylaminoacetone (0.532 g, 4.00 mmol). This solution was allowed to stir for 10 minutes. $\text{Fe}(\text{II})\text{Cl}_2$ (0.504 g, 4.00 mmol) was dissolved in 5 ml of dry methanol and added to the ligand mixture dropwise, immediately producing a yellow precipitate. NaPF_6

(1.34 g, 8.00 mmol) was added and the reaction mixture was allowed to stir overnight. The precipitate was filtered and washed with methanol to yield a pale yellow solid. Yield: 53 %. Yellow needle-like crystals were obtained by slow diffusion of diethyl ether into an acetonitrile solution. $E_{1/2}$ (MeCN, 298 K) = 303 mV vs SCE. Solution Magnetic Moment (303 K, MeCN) $\mu_{\text{eff}} = 5.37$ BM. Magnetic Moment in Solid State: $\mu_{\text{eff}} = 5.23$ BM. Elemental analysis calcd for $\text{FeN}_5\text{C}_{13}\text{H}_{31}\text{ClPF}_6$: C, 31.63; N, 14.19; H, 6.33. Found C, 31.68; N, 14.17; H, 6.40.

Synthesis of $[\text{Co}^{\text{II}}(\text{S}^{\text{Me}_2}\text{N}_4(\text{tren}))(\text{Cl})]$ (4**)**—Anhydrous cobaltous chloride (519 mg, 4.00 mmol) in methanol (20 mL) was added dropwise to a solution containing NaOCH_3 (216 mg, 4.00 mmol), 3-Methyl-3-mercapto-2-butanone (472 mg, 4 mmol), and tris-(2-aminoethyl)amine (585 mg, 4.00 mmol) and left to stir overnight. The resulting green solution was dried under reduced pressure, re-dissolved with a minimal amount of acetonitrile, and filtered through a fritted funnel. The solution was layered with diethyl ether and left to re-crystallize overnight giving pure **4** (700 mg, 52% yield) as a crystalline solid. Electronic absorption (CH_3CN): λ_{max} (ϵ) = 362 (1275) nm, 456 (243) nm, 504 (204) nm, 595 (179) nm. Redox potential (MeCN vs SCE): $E_{\text{pa}} = +270$ mV, $E_{\text{pc}} = -729$ mV. Elemental Analysis for $\text{C}_{11}\text{H}_{25}\text{ClCoN}_4\text{S}$: Calculated: C, 38.88; H, 7.42; N, 16.49. Found: C, 38.64; H, 7.34; N, 16.69.

Synthesis of CyclamPrOH•4HCl

Synthesis of 3-iodo-1-propanol—To a 50 mL round bottom flask was added 3-chloro-1-propanol (2.93 g, 316 mmol) and sodium iodide (9.47 g, 632 mmol) in 20 mL acetone. The solution was refluxed overnight. After cooling to room temperature, the solution was filtered and the solvent was removed by rotary evaporation. The residue was redissolved in methylene chloride, filtered again, and the solvent was removed. The red-orange oil was redissolved in 5 mL methylene chloride and passed through a small silica plug (100% CH_2Cl_2). The first several fractions were collected and evaporated to give a brown oil (3.06 g, 52% yield). ^1H NMR (CDCl_3): 3.73 (q, 2H, $J = 5.70$ Hz), 3.30 (t, 2H, $J = 6.60$ Hz), 2.13 (s, 1H), 2.06 (quintet, 2H, $J = 6.00$ Hz).

Synthesis of 3Boc-cyclamPrOH—In a 50 mL round bottom flask was added 3-iodo-1-propanol (1.79 g, 9.62 mmol), 3Boc-cyclam (2.41 g, 4.81 mmol), and K_2CO_3 (1.33 g, 9.62 mmol) in 20 mL DMF. The solution was stirred under nitrogen at room temperature for 9 days. 100 mL of DI water was added to the solution and the product was extracted with ethyl acetate (3×30 mL). The organic layer was then washed with brine (2×30 mL) and dried over MgSO_4 . Evaporation led to a sticky white residue of the product and excess 3-iodo-1-propanol. TLC revealed that the products would be difficult to separate, so the Boc deprotection step was carried out on the mixture. ESI-MS ($M + 1$) calcd for $M = [\text{N}_4\text{C}_{28}\text{H}_{54}\text{O}_7]$: 558.7, found: 559.9.

Synthesis of CyclamPrOH•4HCl—To the residue formed in the above step was added 15 mL of 4M HCl/dioxane. The solution was allowed to stir at room temperature overnight. The solvent was removed by rotary evaporation and the residue was taken up in a MeOH/ Et_2O solution (1:3). The solution was filtered and the precipitate was washed with additional ether (2×15 mL). The solid was further purified by triturating over 2 days in an $\text{Et}_2\text{O}/\text{CH}_2\text{Cl}_2$ mixture (1:1). The residue was then filtered, washed with additional ether, and pumped dry under vacuum. Yield: 1.92 g (~100%). ^1H NMR (D_2O): 3.65 (t, 2H, $J = 5.70$ Hz), 3.58-3.15 (m, 20H), 2.03 (m, 4H), 1.87 (m, 2H). ESI-MS ($M + 1$) calcd for $M = \text{N}_4\text{C}_{13}\text{H}_{30}\text{O}$: 258.4; found: 259.3.

Synthesis of [Fe^{II}(N₄(cyclam)PrO)(BPh₄) • MeOH (6)—Sodium methoxide (270 mg, 5.00 mmol) and CyclamPrOH•4HCl (404 mg, 1.00 mmol) were dissolved in 15 mL methanol and allowed to stir for 30 minutes. A solution of FeCl₂ (127 mg, 1.00 mmol) in methanol (10 mL) was added dropwise to the ligand solution at room temperature. Sodium tetraphenylborate (342 mg, 1.00 mmol) was added and the solution was allowed to stir overnight. The solution was then filtered through Celite and the methanol was evaporated. The residue was redissolved in 20 mL acetonitrile and filtered again. The filtrate was concentrated to 10 mL, layered with ether (35 mL), and cooled to −40 °C to yield white solid **6**. Yield: 368 mg (58.2%). Single crystals suitable for x-ray diffraction were grown at room temperature over several days by layering ether (20 mL) onto a methanolic solution (10 mL) of **6**. ESI-MS calcd for [FeN₄OC₁₃H₂₉]⁺: 313.2; found: 312.9. Magnetic Moment (solid state): 4.79 BM. Redox potential (MeCN vs SCE): E_{1/2} = + 35 mV. Elemental Analysis for FeC₃₈H₅₃N₄O₂B: Calculated: C, 68.68; H, 8.04; N, 8.43. Found: C, 68.63; H, 7.93; N, 8.40.

Synthesis of [Co^{II}(cyclam–PrS)][BPh₄] • MeCN (7)—Sodium methoxide (170 mg, 5.00 mmol) and cyclam–PrSAC•4HCl (200 mg, 1.00 mmol) were dissolved in 15 mL of methanol, allowed to stir for 30 minutes and cooled to −30 °C. CoCl₂•6H₂O (82 mg, 1.00 mmol) was dissolved in 5 mL of methanol, cooled to −30 °C and added dropwise to the cold ligand while stirring. A pink solution is formed instantly. Sodium tetraphenylborate (216 mg, 1.00 mmol) was then added in 10 mL of methanol and the solution was stirred overnight. The solution was filtered, the filtrate was concentrated to 5 mL, layered with ether (25 mL) and cooled to −35 °C to yield pink crystals. Yield: 46.5%. Electronic absorption (CH₃CN): λ_{max} (ε) = 324 (2029) nm, 366 (975) nm, 549 (116) nm; (MeOH) λ_{max} (ε) = 317 (2348) nm, 359 (1128) nm, 557 (134) nm. E_{pa} (MeCN) = +0.900 V vs SCE. Solution Magnetic Moment (310.2 K; MeOH) μ_{eff} = 3.97 BM. EPR (DCM/Tol glass (1:1), 10.3 K: g₁ = 4.99, g₂ = 3.93, g₃ = 2.61. ESI-MS calcd for [CoN₄SC₁₃H₂₉]⁺: 332.8; found 332.2. Elemental analysis calcd for CoC₃₇H₄₉N₄SB: C, 66.30; N, 8.60; H, 7.62. Found C, 66.07, N, 8.43, H, 7.72.

X-ray Crystallography—A bi-refrigent yellow crystal prism of **3** was cut down to 0.24 × 0.12 × 0.12 mm and mounted on a glass capillary with oil. Data was collected at 23 °C on a Nonius Kappa CCD diffractometer. The crystal-to-detector distance was set to 30 mm and exposure time was 15 seconds per degree for all data sets with a scan width of 1.8°. The data collection was 98.6% complete to 24.71° in θ . A total of 20518 partial and complete reflections were collected covering the indices, h = −15 to 16, k = −21 to 21, l = −10 to 10. 3472 reflections were symmetry independent and the R_{int} = 0.116 indicated that the data was of less than average quality (average quality = 0.07). Indexing and unit cell refinements indicated an orthorhombic P lattice in the space group Pna2₁ (No. 33). Data was collected at room temperature because a slow splitting of peaks was observed upon cooling the crystal. The splitting of peaks appeared to indicate that a low temperature crystallographic phase transition was occurring, though no structure of this phase could be obtained. This does, however, explain the large thermal parameters of several atoms in the structure. Furthermore, the BASF parameter indicates the crystal is comprised of approximately 46% of the inversion twin of this structure.

A red prism, size 0.44 × 0.38 × 0.27 mm of **4** was mounted on a glass capillary with oil. Data was collected at −143 °C. on a Nonius Kappa CCD diffractometer. The crystal-to-detector distance was set to 30 mm and exposure time was 15 seconds per degree for all data sets with a scan width of 2°. The data collection was 99.8% complete to 25° in θ . A total of 20162 partial and complete reflections were collected covering the indices, h = −13 to 13, k = −15 to 17, l = −15 to 15. 3700 reflections were symmetry independent and the R_{int} = 0.0366 indicated that the data was excellent (average quality = 0.07). Indexing and unit cell refinements indicated a monoclinic lattice in the space group P2₁/c (No. 14).

A yellow crystal prism size $0.54 \times 0.34 \times 0.19$ mm of **6** was mounted on a glass capillary with oil. Data was collected at -143 °C on a Nonius Kappa CCD diffractometer. The crystal-to-detector distance was set to 30 mm and exposure time was 20 seconds per degree for all data sets with a scan width of 1.1° . The data collection was 99.0% complete to 25° in ϑ . A total of 78675 partial and complete reflections were collected covering the indices, $h = -19$ to 18, $k = -12$ to 12, $l = -36$ to 43. 8278 reflections were symmetry independent and the $R_{\text{int}} = 0.0829$ indicated that the data was of average quality (average quality = 0.07). Indexing and unit cell refinements indicated a primitive monoclinic lattice in the space group $P2_1/c$ (No. 14).

A purple crystal plate of **7** was cut down to $0.29 \times 0.12 \times 0.12$ mm and mounted on a glass capillary with oil. Data was collected at -143 °C on a Nonius Kappa CCD diffractometer. The crystal-to-detector distance was set to 30 mm and exposure time was 40 seconds per degree for all data sets with a scan width of 1.6° . The data collection was 99.5% complete to 25° in ϑ . A total of 47402 partial and complete reflections were collected covering the indices, $h = -12$ to 12, $k = -23$ to 26, $l = -28$ to 28. 8679 reflections were symmetry independent and the $R_{\text{int}} = 0.0738$ indicated that the data was of average quality (average quality = 0.07). Indexing and unit cell refinements indicated a primitive monoclinic lattice in the space group $P2_1/c$ (No. 14).

The data for **3**, **4**, **6**, and **7** were all integrated and scaled using hkl-SCALEPACK, and an absorption correction was performed using SORTAV. Solution by direct methods (SIR97; default 4) produced a complete heavy atom phasing model consistent with the proposed structure. All non-hydrogen atoms were refined anisotropically by full-matrix least-squares methods, while all hydrogen atoms were initially located in a difference Fourier map and were refined with a riding model using a riding model. Crystal data for **3–4**, **6–7** is presented in Table 1. Selected bond distances and angles are assembled in Table 2.

Results and Discussion

Synthesis of Thiolate–, Alkoxide–, and Amine–Ligated 1–8

In order to correlate properties with metal ion structure and ligated atoms, one needs to compare structurally analogous systems in which the ligand backbone remains the same and only the heteroatoms or metal ion are varied. Multidentate ligands are generally required to maintain a somewhat rigid structure, since the biologically important first-row transition-metal ions tend to be substitution labile. Schematic drawing of the series of thiolate, alkoxide, and amine complexes described in this study are shown in Scheme I.

Thiolate–ligated $[\text{Fe}^{\text{II}}(\text{S}^{\text{Me}_2}\text{N}_4(\text{tren}))]^+$ (**1**),²⁷ $[\text{Fe}^{\text{II}}(\text{cyclam-PrS})]^+$ (**5**),²⁴ and $[\text{Fe}^{\text{II}}(\text{S}^{\text{Me}_2}\text{N}_4(\text{tren-Et}_4))]^+$ (**8**)²⁹ were synthesized as previously described. Thiolate–ligated $[\text{Co}^{\text{II}}(\text{S}^{\text{Me}_2}\text{N}_4(\text{tren}))(\text{Cl})]$ (**4**) and $[\text{Co}^{\text{II}}(\text{cyclam-PrS})]^+$ (**7**) were synthesized via a similar route using CoCl_2 or $[\text{Co}(\text{H}_2\text{O})_6]\text{Cl}_2$ in place of FeCl_2 . Thiolate–ligated **8**, which is a tertiary amine derivative of our functional SOR model **1**, was synthesized in order to determine if the amine N–H protons or H-atoms were involved in the proton–dependent mechanism of superoxide reduction. Alkoxide–ligated $[\text{Fe}^{\text{II}}(\text{O}^{\text{Me}_2}\text{N}_4(\text{tren})(\text{H}_2\text{O}))]^+$ (**2**) was synthesized as described elsewhere.²⁸ The cyclamPrOH·HCl ligand of alkoxide–ligated $[\text{Fe}^{\text{II}}(\text{cyclam-PrO})]^+$ (**6**) was pre-assembled via a four step synthesis (outlined in Scheme I), analogous to that previously reported for cyclam–PrSAc·4HCl.²⁴ For the sulfur-containing ligand, the thiolate must be protected (as a thioacetate group) throughout the synthesis to prevent oxidation of the thiol and also to inhibit its propensity to oligimerize. For the alcohol-substituted ligand, no protecting group was needed for the alcohol. The alkoxide complex $[\text{Fe}^{\text{II}}(\text{cyclam-PrO})]^+$ (**6**) was prepared via in situ deprotonation of the alcohol and amines using five equivalents of sodium methoxide, followed by the addition of FeCl_2 and

NaBPh₄ in MeOH. Recrystallization from MeOH/Et₂O afforded analytically pure **6** as a white solid. The ESI-MS spectrum of this solid redissolved in methanol shows the major peak at *m/z* = 312.9, consistent with a 5-coordinate geometry for the iron-containing cation. This structure was confirmed by x-ray crystallography. Amine-ligated [Fe^{II}(N^{Et}₂N₄(tren)(Cl))]⁺ (**3**) was synthesized via the FeCl₂-templated Schiff-base condensation between *tris*(2-aminoethyl)amine (tren) and diethylaminoacetone in MeOH.

Comparison of Structures 1–8

Single crystals of **3**, **4**, **6**, and **7** were obtained via layer diffusion of Et₂O into an MeCN (**3**, **4**, **7**) or MeOH (**6**) solution. The structures of **1**, **2**, **5**, and **8** were described elsewhere.^{24–27–29} Metrical parameters for **1** – **8** are compared in Table 2. ORTEP diagrams of **3**, **4**, **6**, and **7** are shown in Figures 1 and 2. In contrast to the other structures, X-ray data for **3** was collected at ambient temperature because a phase change occurred at lower temperatures, thus the significantly larger thermal ellipsoids (Figure 1). Alkoxide and amine-ligated **2** and **3** consistently crystallize as six-coordinate complexes with a water (**2**) or Cl[–] (**3**) occupying the sixth site (Figure 1; Scheme I). In contrast, all of the thiolate-ligated complexes **1**, **4**, **5**, **7**, and **8** remain five-coordinate, even in the presence of coordinating solvents or a Cl[–] counterion. This suggests that the metal ion Lewis acidity is lower for thiolate-ligated **1**, **4**, **5**, **7**, and **8**, and greatest for amine and alkoxide-ligated **2** and **3**. The noticeable *trans* influence of the thiolate^{25,34} could also be a factor partially responsible for the low coordination number. Amine-ligated **3** (Figure 1) would be a dication if the Cl[–] anion were not incorporated into its primary coordination sphere, explaining its higher affinity for Cl[–]. This is in contrast to all of the other molecules included in this study in which charge is conserved to afford monocationic metal complexes. The fact that the Cl[–] counterion of **4** (Figure 1) is not incorporated into its primary coordination sphere can be attributed to a decreased Lewis acidity (relative to **3**) caused by the replacement of a neutral amine ligand with a monoanionic thiolate. Charge is conserved however, by replacing the thiolate of **1** with an alkoxide in **2**, and yet the alkoxide complex reproducibly binds a sixth ligand, whereas the thiolate complex **1** does not. Again this implies that the thiolate decreases the metal ion Lewis acidity helping to maintain an open coordination site. Thiolate-ligated **7** (Figure 2) remains five-coordinate even though it cocrystallizes with a MeCN solvent. The acetonitrile nitrogen (N(5)) points directly towards the Co ion of **7**, is H-bonded to two of the cyclam secondary amine protons (N(5)•••H–N(3)= 2.245(9) Å; N(5)•••H–N(4)= 2.223(9) Å), and *trans* to the thiolate. These are not the first examples of coordinatively unsaturated thiolate complexes observed in our lab: our bis-thiolate-ligated iron complex [Fe^{III}(S₂^{Me}₂N₃(Pr,Pr))]⁺ remains five-coordinate even though it contains a more Lewis acidic Fe³⁺ metal ion, and the first-row series [M^{II}(tren)N₄S^{Me}₂]⁺ (M = Mn, Fe (**1**), Co, Ni, Cu, Zn)¹⁶ all maintain five-coordinate structures even in coordinating solvents such as MeCN. These findings are unusual in that the synthesis of complexes containing an open-coordination site usually requires the use of steric bulk, non-coordinating solvents, and strict avoidance of potentially coordinating counterions.^{35–38}

Why alkoxide-ligated **6** does not bind a sixth ligand, whereas **2** does, is unclear. The secondary versus primary amines may be responsible (by lowering the metal ion Lewis acidity), although the redox potentials (*vide infra*) imply that this is not the case. Ligand constraints favoring a distorted trigonal bipyramidal geometry may also be involved. This was the explanation given for the consistently observed five-coordinate structure of the first-row tren-ligated [M^{II}(tren)N₄S^{Me}₂]⁺ (M = Mn, Fe (**1**), Co, Ni, Cu, Zn) series recently reported by our group.¹⁶ The fact that alkoxide and amine-ligated tren complexes **2** and **3** afford six-coordinate structures would appear to rule this out. On the other hand, it may be that the cyclam ligand of **6** constrains the geometry to be five-coordinate when a metal ion

that is too large to fit inside the macrocyclic cavity is involved. The high spin Fe^{2+} ion of **6** (vide infra) would be larger than the preferred low-spin Fe^{2+} , or oxidized Fe^{3+} .

With the exception of **5**, the five-coordinate structures are all tetragonally distorted trigonal bipyramidal ($\tau = 0.47\text{--}0.67$) approximately half-way between square pyramidal ($\tau = 0.0$) and trigonal bipyramidal ($\tau = 1.0$). Thiolate-appended cyclam-ligated **5** is much closer to square pyramidal ($\tau = 0.13$). The X = S, O, N heteroatom occupies the apical position of the tren-derived complexes (**1**, **4**, and **8**), and the equatorial plane of the cyclam-derived complexes **6** and **7**. The Co^{II} -cyclam complex **7** crystallizes as the less frequently encountered trans-V (R, R, R, R) isomer, whereas **5** and **6** are both isolated as the more common trans-I (R, S, R, S) isomer. The trans-I configurational isomer is the most stable isomer for five-coordinate cyclam complexes.

For the five-coordinate structures, the alkoxide Fe–O bond (1.906(2) Å in **6**) is significantly shorter than the thiolate Fe–S bonds (range: 2.286(1) Å (in **5**) – 2.3281(9) Å (in **1**)), as one would expect based on differences between the covalent radii of O versus S. The alkoxide Fe–O bond (2.001(1) Å) for the six-coordinate structure **2** is longer than that of the five-coordinate structure **6**, as expected based on the size difference of 5- versus 6-coordinate metal ion radii.³⁹ The tertiary amine Fe–N(5) bond of six-coordinate **3** (2.293(5) Å) is longer than the tertiary amines of five-coordinate **5** (2.181(4) Å), **6** (2.225(2) Å), and **8** (mean: 2.222(5) Å), most likely because it is six-coordinate. Within the tren-ligated systems, there is less variability of these tertiary amine distances (range: 2.223(2) – 2.268(3) Å) most likely due to ligand constraints. The secondary amine Fe–N distances (2.138 Å in **5** to 2.206(2) in **6**) are noticeably shorter than the tertiary amine Fe–N bonds, but longer than the imine Fe–N bonds, and the only five coordinate primary amine Fe–N bonds (2.117 Å and 2.131(3) Å in **1**). The Co–S, Co–O, and Co–N distances in **4** and **7** are noticeably shorter than those of their iron analogue. This is expected based on periodic trends.¹⁶

Comparison of the Electronic and Magnetic Properties of 1–8—In contrast to the intensely colored, low-spin Fe(III)-thiolates, the Fe(II)-thiolates **1**, **5**, and **8** described herein are pale in color, and high-spin, due to the absence of charge transfer bands in the visible region, and the absence of pentet excited states (in the high-spin d_6 ion state splitting diagram). There is a poorer energy match between the reduced metal ion d- and sulfur-orbitals, this charge transfer bands fall outside the visible region. The alkoxide and amine complexes **2**, **3**, and **6** are pale yellow and high-spin. The pale color to these solutions is due to the weak intensity of Laporte-forbidden d-d transitions for the six-coordinate complexes, or poor energy match between the metal ion d- and alkoxide-orbitals. The thiolate-ligated cobalt complexes **4** and **7** are slightly more intensely colored due to the presence of lower energy charge transfer bands that tail into the visible (Figures 3 and supplemental Figure S-4), as well as the presence of two additional d-d transitions with moderate intensity due to the low symmetry and accessibility of a quartet 4P excited state. The charge transfer bands are lower in energy because the Co^{2+} ion orbitals lie lower than those of Fe^{2+} . The electronic absorption spectra of **1** and **4–8** are all solvent-independent implying that solvents do not bind to the metal ion's open coordination site, even in coordinating solvents.

The iron complexes **1**, **2**, **3** ($\mu_{\text{eff}} = 5.23$ BM; Figure 4), **5**, **6** ($\mu_{\text{eff}} = 4.79$ BM; Figure S-2), and **8** are all high-spin $S = 2$ both in the solid state and in solution. This is irrespective of whether the fifth ligand X (Table 2) is an alkoxide, thiolate, or amine (Figure 4). The cobalt complexes **4** ($\mu_{\text{eff}} = 4.05$ BM),¹⁶ and **7** ($\mu_{\text{eff}} = 3.97$ BM; supplemental Figure S-3) are both high-spin as well with an $S = 3/2$ ground state over the temperature range 4 – 300 K.

Comparison of Redox Potentials—Correlation with Apical Ligand X (S, O, N) and Metal Ion (Co²⁺, Fe²⁺)

Upon oxidation, the iron and cobalt complexes **1** and **4** both have been shown to convert to low-spin ($S = \frac{1}{2}$ or 0) six-coordinate structures.^{27,41} The iron complex, [Fe^{II}(S^{Me}₂N₄(tren))]⁺ (**1**), reacts with HO₂ to afford six-coordinate [Fe^{III}(S^{Me}₂N₄(tren))(MeCN)]²⁺ via a hydroperoxo intermediate [Fe^{III}(S^{Me}₂N₄(tren))(OOH)]⁺.^{27,42} Cyclic voltammograms of **3**, **6**, and **7**, are shown in Figure 5, and supplemental Figures S–6 and S–7, respectively. Redox potentials are assembled in Table 3. Comparison of the redox potential ($E_{1/2}$ and E_{pa} ; Table 3) of tren-ligated **1** versus that of tren-ligated **4** shows that replacement of the Fe²⁺ with Co²⁺ within an identical ligand environment causes the potential to shift cathodically by +470 mV. Similarly, comparison of $E_{1/2}$ for cyclam-ligated **5** versus $E_{1/2}$ for cyclam-ligated **7** shows that replacement of Fe²⁺ with Co²⁺, within an identical ligand environment, causes the potential to shift cathodically by +765 mV. These differences reflect differences in ionization potentials for the two metals, and follow expected periodic trends. As demonstrated by the lower energy charge transfer bands in their electronic absorption spectra (vide supra), the Co²⁺ d-orbitals lie lower in energy than the Fe²⁺ orbitals making Co²⁺ more difficult to oxidize. Given this, Co²⁺ would not be the metal ion of choice for metalloenzymes, such as SOR, that function as reducing agents.

Comparison of the redox potentials ($E_{1/2}$; Table 3) of **8** versus those of **5** and **1**, shows that a cathodic shift in the redox potential by +275 mV upon replacement of primary amines with secondary amines, and +610 mV upon replacement of primary amines with tertiary amines. Tertiary amines stabilize iron in the +2 oxidation state, whereas primary amines stabilize iron in the +3 oxidation state, and secondary amines fall somewhere in between.

Comparison of the redox potential ($E_{1/2}$; Table 3) of **1** versus that of **3**, shows that the replacement of a thiolate sulfur (X in Table 2) with an amine nitrogen causes the potential to shift cathodically, by +503 mV. In contrast, replacement of the thiolate sulfur in **1** with an alkoxide oxygen in **2** causes the potential to shift in the opposite direction (ie anodically). The redox potential of the water-ligated alkoxide complex **2** (Scheme I) is pH-dependent, shifting anodically by ~60 mV/pH unit, most likely because the coordinated water readily deprotonates.²⁸ At a higher pH, the potential is more anodic ($E_{1/2}$ (pH = 10.6) = –420 mV), while at a lower pH it is less anodically-shifted ($E_{1/2}$ (pH = 6.8) = –285 mV), consistent with increased charge build up at the metal center under more basic conditions, where a neutral hydroxide-bound derivative would form. Although the ligands are conserved with only minor changes to the apical X-atom, the coordination numbers differ for **1** (5-coordinate) versus **2** (6-coordinate) and **3** (6-coordinate), making these comparisons somewhat less valid. One would expect the redox potential to shift anodically in response to an increase in coordination number, implying that **3** would have an even more cathodically shifted potential in the absence of the sixth (Cl[–]) ligand. Alkoxide-ligated **2** would have a less anodically shifted potential in the absence of its sixth (H₂O) ligand. Another property influencing redox potentials is overall molecular charge. Molecular charge is conserved for **1**, **2**, and **3** — all are monocations. This would not be the case for **2** at higher pH where deprotonation of the water would afford a neutral hydroxide-bound species, or for **3** in the absence of the Cl[–] ligand where instead a dicationic (2+) species would form.

In summary, comparison of redox potentials indicates that the tertiary amine tren-ligands of **3** and **8** stabilize iron in the +2 oxidation state, whereas the alkoxide tren-ligand stabilizes iron in the +3 oxidation state. The primary amine thiolate tren-ligand falls somewhere in between, favoring the +3 oxidation state more than the amine, but less than an alkoxide. Thiolate-ligated [Fe^{II}(S^{Me}₂N₄(tren))]⁺ (**1**) is a better reducing agent than thiolate-ligated [Co^{II}(S^{Me}₂N₄(tren))]⁺ (**4**), and stabilizes iron in the +3 oxidation state, and cobalt in the +2 oxidation state. Within the cyclam-ligated series **5**–**7** similar trends are observed.

Replacement of the apical thiolate of $[\text{Fe}^{\text{II}}(\text{cyclam}-\text{PrS})]^+$ (**5**) with an apical alkoxide in $[\text{Fe}^{\text{II}}(\text{cyclam}-\text{PrO})]^+$ (**6**) causes the redox potential to shift anodically by -100 mV, and replacement of the Fe^{2+} ion of **5** with a Co^{2+} ion (in **7**) causes the redox potential to shift cathodically by $+765$ mV.

Summary and Conclusions

Comparison of the structurally-related derivatives described herein, and elsewhere,^{23,26,43} shows that thiolate ligands reduce metal ion Lewis acidity relative to alkoxides and amines, and have a strong *trans* influence thereby helping to maintain an open coordination site. With an open coordination site, the metal ion is accessible so that substrates can react via an inner sphere mechanism. Comparison of the redox potentials of the structurally analogous compounds described herein indicates that the higher-valent Fe^{3+} oxidation state is favored most by alkoxide ligands, and least by amine ligands. This fits with what one would expect based on the hard-soft acid-base classification of metal ions and ligands (Lewis acids and bases).⁴⁴ Thiolates favor Fe^{3+} more than amines but less than alkoxides. This is somewhat surprising given that sulfur is usually considered to be a soft Lewis base, and Fe^{3+} a hard Lewis acid. As shown previously however,⁴⁵ extensive π -interaction and delocalization stabilizes the $\text{Fe}(\text{III})$ -S bonds more than one would anticipate. For a redox active iron enzyme to function as a catalyst capable of reducing substrates, both the Fe^{3+} and the Fe^{2+} states must be accessible. Although a catalyst with a more anodic potential would be a more potent reducing agent, an excessively anodic potential would result in an inaccessible catalytically active reduced Fe^{2+} state.⁴⁶ Since an open coordination site, and a moderately low redox-potential are both necessary for efficient, inner sphere, substrate (e.g superoxide) reduction to occur, the combination of a thiolate (as opposed to an alkoxide or amine) coordinated to iron (as opposed to cobalt), with primary (as opposed to secondary or tertiary) amines creates both of these optimum conditions needed for this chemistry to occur. This implies that the cysteinyl ligand of the non-heme iron enzyme SOR plays an important role in promoting function.

Supplementary Material

Refer to Web version on PubMed Central for supplementary material.

Acknowledgments

This work was supported by the NIH (Grant Number: GM 45811-17).

Supporting Information. Supporting information contains crystallographic data for complexes **3**, **4**, **6**, and **7**, $1/\chi$ vs. T plot for **6** and **7**, EPR of **7**, UV/vis of **4**, and CV of **6** and **7**.

References

1. Solomon EI. Inorg Chem. 2006; 45:8012–8025. [PubMed: 16999398]
2. Kovacs JA. Chem Rev. 2004; 104:825–848. [PubMed: 14871143]
3. Rao PV, Holm RH. Chem Rev. 2004; 104:527–560. [PubMed: 14871134]
4. Gray HB, Malström B, Williams RJP. J Biol Inorg Chem. 2000; 5:551 – 559. [PubMed: 11085645]
5. Lindahl PA. Biochemistry. 2002; 41:2097–2105. [PubMed: 11841199]
6. Rao PV, Bhaduri S, Jiang J, Hong D, Holm RH. J Am Chem Soc. 2005; 127:1933–1945. [PubMed: 15701028]
7. Rodrigues JV, Abreu IA, Cabelli D, Teixeira M. Biochemistry. 2006; 45:9266–9278. [PubMed: 16866373]
8. Kurtz DM Jr. Acc Chem Res. 2004; 37:902–908. [PubMed: 15612680]

9. Fiedler AT, Bryngelson PA, Maroney MJ, Brunold TC. *J Am Chem Soc.* 2005; 127:5449–5462. [PubMed: 15826182]
10. Wuerges J, Lee JW, Yim YI, Yim HS, Kang SO, Carugo KD. *Proc Natl Acad Sci USA.* 2004; 101:8569–8574. [PubMed: 15173586]
11. Shearer J, Long LM. *Inorg Chem.* 2006; 45:2358–2360. [PubMed: 16529443]
12. Denisov IG, Makris TM, Sligar SG, Schlichting I. *Chem Rev.* 2005; 105:2253–2277. [PubMed: 15941214]
13. Ogata H, Hirota S, Nakahara A, Komori H, Shibata N, Kato T, Kano K, Higuchi Y. *Structure.* 2005; 13:1635–1642. [PubMed: 16271886]
14. Rajagopalan PTR, Yu XC, Pei D. *J Am Chem Soc.* 1997; 119:12418–12419.
15. Chang S, Karambelkar VV, diTargiani RC, Goldberg DP. *Inorg Chem.* 2001; 40:194–195. [PubMed: 11170522]
16. Brines LM, Shearer J, Fender JK, Schweitzer D, Shoner SC, Barnhart D, Kaminsky W, Lovell S, Kovacs JA. *Inorg Chem.* 2007; 46 in press.
17. Randall DW, George SD, Hedman B, Hodgson KO, Fujisawa K, Solomon EI. *J Am Chem Soc.* 2000; 122:11620–11631.
18. Glaser T, Hedman B, Hodgson KO, Solomon EI. *Acc Chem Res.* 2000; 33:859–868. [PubMed: 11123885]
19. George SD, Metz M, Szilagyi RK, Wang H, Cramer SP, Lu Y, Tolman WB, Hedman B, Hodgson KO, Solomon EI. *J Am Chem Soc.* 2001; 123:5757–5767. [PubMed: 11403610]
20. Lowery MD, Guckert JA, Gebhard MS, Solomon EI. *J Am Chem Soc.* 1993; 115:3012–3013.
21. Shoner S, Barnhart D, Kovacs JA. *Inorg Chem.* 1995; 34:4517–4518.
22. Jackson HL, Shoner SC, Rittenberg D, Cowen JA, Lovell S, Barnhart D, Kovacs JA. *Inorg Chem.* 2001; 40:1646–1653. [PubMed: 11261975]
23. Ellison JJ, Nienstedt A, Shoner SC, Barnhart D, Cowen JA, Kovacs JA. *J Am Chem Soc.* 1998; 120:5691–5700.
24. Kitagawa T, Dey A, Lugo-Mas P, Benedict J, Kaminsky W, Solomon E, Kovacs JA. *J Am Chem Soc.* 2006; 128:14448–14449. [PubMed: 17090014]
25. Shearer J, Kung IY, Lovell S, Kaminsky W, Kovacs JA. *J Am Chem Soc.* 2001; 123:463–468. [PubMed: 11456548]
26. Shearer J, Jackson HL, Schweitzer D, Rittenberg DK, Leavy TM, Kaminsky W, Scarrow RC, Kovacs JA. *J Am Chem Soc.* 2002; 124:11417–11428. [PubMed: 12236756]
27. Shearer J, Nehring J, Kaminsky W, Kovacs JA. *Inorg Chem.* 2001; 40:5483–5484. [PubMed: 11599942]
28. Brines LM, Kaminsky W, Kirk ML, Kovacs JA. 2007 manuscript in preparation.
29. Theisen RM, Shearer J, Kaminsky W, Kovacs JA. *Inorg Chem.* 2004; 43:7682–7690. [PubMed: 15554633]
30. Perrin, DD.; Armarego, WLF.; Perrin, DR. *Purification of Laboratory Chemicals.* 2. Pergamon Press; Elmsford, NY: 1980.
31. Live DH, Chan SI. *Anal Chem.* 1970; 42:791.
32. Evans DA. *J Chem Soc.* 1959:2005.
33. Van Geet AL. *Anal Chem.* 1968; 40:2227–2229.
34. Shearer J, Kung I, Lovell S, Kovacs JA. *Inorg Chem.* 2000; 39:4998–4999. [PubMed: 11233193]
35. Yandulov DV, Schrock RR. *J Am Chem Soc.* 2002; 124:6252–6253. [PubMed: 12033849]
36. Smith JM, Sadique AR, Cundari TR, Rodgers KR, Lukat-Rodgers G, Lachicotte RJ, Flaschenriem CJ, Vela J, Holland PL. *J Am Chem Soc.* 2006; 128:756–769. [PubMed: 16417365]
37. O’Keefe BJ, Breyfogle LE, Hillmyer MA, Tolman WB. *J Am Chem Soc.* 2002; 124:4384–4393. [PubMed: 11960467]
38. Brown SD, Betley TA, Peters JC. *J Am Chem Soc.* 2003; 125:322–323. [PubMed: 12517130]
39. Shannon RD. *Acta Cryst.* 1976; A32:751–767.
40. Theisen, RM. PhD thesis. University of Washington; 2005.
41. Swartz RD, Kovacs JA. manuscript in preparation.

42. Shearer J, Scarrow RC, Kovacs JA. J Am Chem Soc. 2002; 124:11709–11717. [PubMed: 12296737]
43. Shoner SC, Nienstedt A, Ellison JJ, Kung I, Barnhart D, Kovacs JA. Inorg Chem. 1998; 37:5721–5725.
44. Lippard, SJ.; Berg, JM. Principles of Bioinorganic Chemistry. University Science; Mill Valley: 1994.
45. Kennepohl P, Neese F, Schweitzer D, Jackson HL, Kovacs JA, Solomon EI. Inorg Chem. 2005; 44:1826–1836. [PubMed: 15762709]
46. Shearer J, Fitch SB, Kaminsky W, Benedict J, Scarrow RC, Kovacs JA. Proc Natl Acad Sci, USA. 2003; 100:3671–3676. [PubMed: 12655068]

Figure 1.

Figure 2.

Figure 3. Electronic absorption spectrum of $[\text{Co}^{\text{II}}(\text{cyclam-PrS})](\text{BPh}_4) \cdot \text{MeCN}$ (**7**) in MeCN at ambient temperature.

Figure 4.

Inverse molar magnetic susceptibility ($1/\chi_m$) vs. temperature (T) plot for $[\text{Fe}^{\text{II}}(\text{NEt}_2\text{N}_4(\text{tren})(\text{Cl}))](\text{PF}_6)$ (**3**) fit to an $S = 2$ spin-state.

Figure 5.

Cyclic voltammogram of $[\text{Fe}^{\text{II}}(\text{N}^{\text{Et}}_2\text{N}_4(\text{tren})(\text{Cl}))](\text{PF}_6)_3$ (**3**) in MeCN at 298 K (0.1 M $(\text{Bu}_4\text{N})\text{PF}_6$, glassy carbon electrode, 150 mV/sec scan rate). Peak potentials versus SCE indicated.

Scheme I.

Scheme II.

Table 1

Crystal Data, Intensity Collections^a and Structure Refinement Parameters for [Fe^{II}(N^{Et}₂N₄(tren)(Cl)](PF₆) (**3**), [Co^{II}(S^{Me}₂N₄(tren)](Cl) (**4**), [Fe^{II}(cyclam-PrO)](BPh₄) • MeOH (**6**), and [Co^{II}(cyclam-PrS)](BPh₄) • MeCN (**7**).

	3	4	6	7
formula	FeC ₁₃ H ₃₁ ClF ₆ N ₅ P	CoC ₁₁ H ₂₅ ClNS	FeC ₃₈ H ₅₃ BN ₄ O ₂	CoC ₃₉ H ₅₂ BN ₅ S
MW	493.70	339.79	664.50	692.66
T, K	296	130(2)	130(2)	130(2)
unit cell	orthorhombic	monoclinic	monoclinic	monoclinic
a, Å	13.6500(4)	10.3700(3)	8.1770(6)	9.2860(2)
b, Å	18.6480(6)	12.9290(4)	12.7180(6)	19.6960(5)
c, Å	8.575(1)	11.5130(4)	17.977(1)	21.3530(7)
α, deg	90	90	90	90
β, deg	90	94.654(1)	132.2251(8)	111.338(1)
γ, deg	90	90	90	90
V, Å ³	2182.7(3)	1538.50(8)	3535.3(1)	1863.4(2)
Z	4	4	4	4
d(calc), g/cm ³	1.502	1.467	1.248	1.265
space group	<i>P</i> na2 ₁	<i>P</i> 2 ₁ / <i>c</i>	<i>P</i> 2 ₁ / <i>c</i>	<i>P</i> 2 ₁ / <i>c</i>
R	0.0500 ^b	0.0443 ^b	0.0548 ^b	0.0553 ^b
R _w	0.1284 ^c	0.1180 ^c	0.1549 ^c	0.1238 ^d
GOF	0.998	1.094	0.982	0.997

^a Mo Kα(α, [−]) (λ = 0.71070 Å) radiation; graphite monochromator; −90 °C.

^b R = Σ||F_o| − |F_c||/Σ|F_o|.

^c R_w = [Σw(|F_o| − |F_c|)²/Σw (F_o²)²]^{1/2}, where w^{−1} = [σ²_{count} + (0.05 F²)²]/4F².

^d R_w = {Σ[w(F_o² − F_c²)²]/Σ[w(F_o²)²]^{1/2}; w = 1/[σ²(F_o²) + (0.1395P)² + 1.3703P], where P = [F_o² + 2F_c²]/3.

Table 2

Selected geometrical parameters for the cations of $[\text{Fe}^{\text{II}}(\text{S}^{\text{Me}2}\text{N}_4(\text{tren}))](\text{PF}_6)$ (1), $[\text{Fe}^{\text{II}}(\text{O}^{\text{Me}2}\text{N}_4(\text{tren})(\text{H}_2\text{O}))](\text{PF}_6)$ (2), $[\text{Fe}^{\text{II}}(\text{N}^{\text{Et}2}\text{N}_4(\text{tren})(\text{Cl}))](\text{PF}_6)$ (3), $[\text{Co}^{\text{II}}(\text{S}^{\text{Me}2}\text{N}_4(\text{tren}))](\text{Cl})$ (4), $[\text{Fe}^{\text{II}}(\text{cyclam-PrS})](\text{BPh}_4)$ (5), $[\text{Fe}^{\text{II}}(\text{cyclam-PrO})](\text{BPh}_4)\bullet\text{MeOH}$ (6), $[\text{Co}^{\text{II}}(\text{cyclam-PrS})](\text{BPh}_4)\bullet\text{MeCN}$ (7), and $[\text{Fe}^{\text{II}}(\text{S}^{\text{Me}2}\text{N}_4(\text{tren-Et}_4))]^{+}$ (8). Bond lengths and bond angles are given in Å and degrees, respectively.

	1	2	3	4	5	6	7	8
M-X(S,O,N)	2.328(9)	2.001(1)	2.293(5)	2.3006(7)	2.286(1)	1.906(2)	2.263(1)	2.317(9)
M-N(1)	2.091(3)	2.129(2)	2.131(5)	2.050(2)	2.181(4)	2.225(2)	2.221(2)	2.104(3)
M-N(2)	2.268(3)	2.257(2)	2.267(5)	2.223(2)	2.138(4)	2.206(2)	2.099(3)	2.227(3)
M-N(3)	2.131(3)	2.249(2)	2.215(15)	2.086(2)	2.174(4)	2.151(2)	2.175(2)	2.219(3)
M-N(4)	2.117(3)	2.222(2)	2.232(14)	2.107(2)	2.166(4)	2.142(2)	2.086(3)	2.220(3)
M-L	N/A	2.149(1)	2.424(2)	N/A	N/A	N/A	N/A	N/A
X-M-N(1)	84.02(8)	76.49(6)	75.1(2)	84.81(6)	101.2(1)	98.77(8)	99.07(7)	83.41(7)
X-M-N(2)	163.02(7)	153.14(6)	152.3(2)	163.96(6)	112.0(1)	95.70(8)	141.68(8)	161.53(7)
X-M-N(3)	108.8(8)	105.45(6)	107.0(6)	106.90(6)	104.9(1)	110.54(8)	91.29(8)	105.63(8)
X-M-N(4)	107.9(9)	108.00(6)	104.7(6)	112.4(2)	101.9(1)	121.16(8)	112.61(8)	109.64(7)
N(1)-M-N(2)	79.9(1)	76.71(7)	77.2(2)	79.49(8)	90.4(2)	165.45(8)	91.9(1)	78.6(1)
N(1)-M-N(3)	124.9(1)	94.77(6)	97.8(5)	125.14(8)	153.7(2)	92.37(8)	169.6(1)	121.4(1)
N(1)-M-N(4)	110.3(1)	103.51(7)	92.5(5)	122.54(8)	82.5(2)	82.21(8)	84.3(1)	111.4(1)
N(2)-M-N(3)	79.9(1)	78.57(7)	76.7(6)	79.93(8)	82.8(2)	81.09(9)	80.1(1)	81.3(1)
N(3)-M-N(4)	115.2(1)	144.73(7)	148.2(2)	102.54(8)	146.1(2)	128.26(8)	91.3(1)	118.49(9)
τ -value	0.64	N/A	N/A	0.65	0.13	0.62	0.47	0.67

X= S, N, or O

L= sixth ligand (O or Cl)

Table 3

Comparison of Redox potentials for thiolate, alkoxide, and amine–ligated M(II) Complexes

	$E_{1/2}$ (vs SCE) oxidation
$[\text{Fe}^{\text{II}}(\text{S}^{\text{Me}2}\text{N}_4(\text{tren}))]^+$ (1)	−200 mV
$[\text{Fe}^{\text{II}}(\text{S}^{\text{Me}2}\text{N}_4(\text{tren-Et}_4))]^+$	+410 mV (irrev)
$[\text{Fe}^{\text{II}}(\text{O}^{\text{Me}2}\text{N}_4(\text{tren})(\text{H}_2\text{O}))]^+$ (2)	−285 mV ($pH = 6.8$, irrev)
	−420 mV ($pH = 10.6$,
$[\text{Fe}^{\text{II}}(\text{N}^{\text{Et}2}\text{N}_4(\text{tren})(\text{Cl}))]^+$ (3)	+303 mV
$[\text{Co}^{\text{II}}(\text{S}^{\text{Me}2}\text{N}_4(\text{tren}))]^+$ (4)	+270 mV (irrev)
$[\text{Fe}^{\text{II}}(\text{cyclam-PrS})]^+$ (5)	+135 mV
$[\text{Fe}^{\text{II}}(\text{cyclam-PrO})]^+$ (6)	+35 mV
$[\text{Fe}^{\text{II}}(\text{cyclam})(\text{MeCN})_2]^{2+}$	+700 mV
$[\text{Co}^{\text{II}}(\text{cyclam-PrS})]^+$ (7)	+900 mV (irrev)