Mechanism for the Reduction of the Mixed-Valent $Mn^{III}Mn^{IV}$[2-OHsalpn]$_2^+$ Complex by Tertiary Amines

M. Tyler Caudle and Vincent L. Pecoraro*

Department of Chemistry, University of Michigan, Ann Arbor, Michigan, 48109-1055

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The mixed-valent dimanganese(III/IV) complex $Mn^{III}Mn^{IV}$(2-OHsalpn)$_2^+$, **1**, is cleanly reduced in acetonitrile by aliphatic tertiary amines to give the dimanganese(III) product Mn^{III}_2(2-OHsalpn)$_2$, **2**. Thorough characterization of the organic reaction products shows that tributylamine is converted to dibutylformamide and propionaldehyde. Kinetic studies and radical trapping experiments suggest that this occurs via initial single-electron transfer from the amine to **1** coupled with C−H$_\alpha$ proton transfer from the oxidized amine. EPR spectroscopy and base inhibition studies indicate that coordination of the amine to **1** is a critical step prior to the electron transfer step. Rate data and its dependence on the amine indicate that the ability of the amine to reduce **1** is correlated to its basicity rather than to its reduction potential. Weakly basic amines were unable to reduce **1** irrespective of their reduction potential. This was inferred to indicate that proton transfer from the amine radical cation is also important in the reduction of **1** by tertiary amines. Comparison of the activation energy with reaction thermodynamics indicates that proton transfer and electron transfer must be concerted to explain the rapidity of the reaction. The fate of the amine radical is dependent on the presence of oxygen, and labeling studies show that oxygen in the organic products arises from dioxygen, although incorporation from trace water was also observed. These data indicate that inhibition of the hydrolytic quenching of the amine radical in an aprotic solvent results in a different fate for the amine radical when compared to amine oxidation reactions in aqueous solution. The proposed mechanism gives new insight into the ability of amines with high reduction potential to reduce metal ions of lower potential. In particular, these data are consistent with the ability of small amines and certain amine-containing buffers to inhibit manganese-dependent oxygen evolution in photosynthesis, which arises in some cases as a result of manganese reduction and its concomitant loss from the PS II reaction center.

Introduction

Certain amines can function as reductants for metal complexes in solution, even in cases where the electrochemically measured redox potential of the amine is substantially higher than the metal complex. This property has been exploited for many years for synthetic purposes[1] using high-valent metal complexes without a clear understanding of the chemical mechanism. Our interest in the problem of manganese-promoted amine oxidation arises from observations concerning the interaction of amines and hydroxylamines with the oxygen-evolving manganese cluster in photosystem II, the site of water oxidation in photosynthesis.[2] It has been observed that the cluster is accessible to small amines,[3] and that loss of manganese from the cluster is one consequence.[4] These observations have been interpreted in terms of reduction of the metal ions by the amine.[4a] However, there have been few studies on the mechanism by which amines reduce manganese complexes, which would provide a chemical framework for understanding of synthetic and biological processes involving amine oxidation at manganese centers.

Previous work on the oxidation of amines by manganese-containing complexes has largely been restricted to manganoporphyrin systems.[5] However, there is a body of work on amine oxidation by simple non-heme iron[6] and copper[7] complexes. In

* Author to whom correspondence should be addressed.

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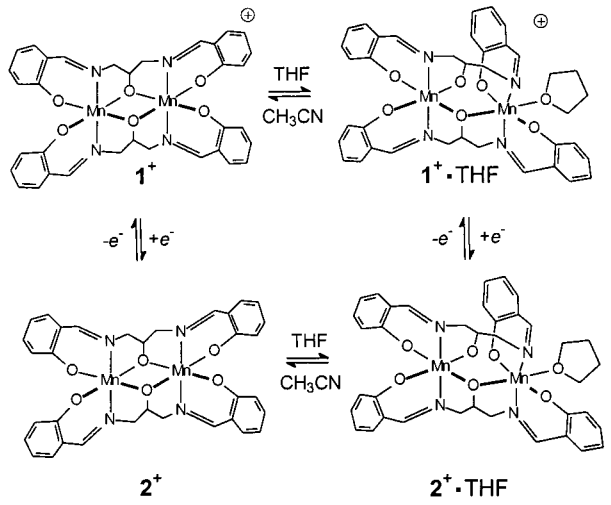
Scheme 1



aqueous media these reactions result in the oxidative dealkylation or dehydrogenation of amines. These reactions bear analogy to copper-containing amine oxidases, which also catalyze the oxidative dealkylation of small amines. These model and enzymatic studies indicate that single electron transfer from the amine to the oxidant is rate limiting. The resultant amine radical cation is probably deprotonated at the α-carbon,[8] although the precise mechanism resulting in net proton transfer has been questioned.[9] In aqueous solution, this initiates subsequent hydrolytic steps ultimately resulting in dealkylation of the oxidized amine to yield an aldehyde. Electrochemical oxidation of amines in aqueous solution yields similar products and is presumably the result of a similar mechanism.[10] It is therefore unclear whether coordination of the amine by the metal ion is required to affect the initial electron transfer using the metal-based oxidants. The fate of the radical in the absence of water has also not been explored in detail. We are reporting here on the tertiary amine-promoted reduction of a high-valent dimanganese complex in an effort to understand in more detail the chemical pathway for amine oxidation by metal complexes, and to develop a rudimentary analogy for the interaction of amines with the photosynthetic oxygen-evolving complex.

The redox and ligand substitution chemistry of the mixed-valent binuclear complex $Mn^{III}Mn^{IV}(2\text{-OHsalpn})_2^+$ (2-OHsalpn = 1,3-bis(salicylideneamino)-2-propanol), **1**, is summarized in Scheme 1.[11,12] The reduction potential of **1** is 0.52 V vs aqueous NHE, and the complex is readily reduced by one electron in the presence of tertiary amines. Complex **1** possesses a readily accessible solvent-labile substitution site in both oxidation states, offering the unique opportunity of probing the role of metal coordination in amine oxidation reactions. Because it is very soluble in organic solvents, it also gives us the opportunity of probing the initial steps in amine oxidation that may be obscured by hydrolytic chemistry in aqueous solution. This paper describes mechanistic studies on the one-electron reduction of **1** by tertiary amines, and the concomitant amine oxidation reactions. We present evidence that, in this case, coordination of the amine by the manganese complex is a critical step in the electron transfer from the amine to $Mn^{III}Mn^{IV}(2\text{-OHsalpn})_2^+$. Furthermore, this study illustrates the considerable differences in amine oxidation reactions in nonaqueous media when compared to aqueous solutions.

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Experimental Section

Materials and Equipment. The preparation of **1** as its perchlorate and hexafluorophosphate salts has been previously described, as has the preparation of the reduced analogue $Mn^{III}_2(2\text{-OHsalpn})_2$,[12] **2**. (***CAUTION:*** *Perchlorate salts of metal complexes with organic ligands are potentially shock sensitive. They should be handled only in small quantities behind a blast shield*.) Aliphatic amines were obtained from Aldrich, distilled by appropriate methods, and stored over 4 Å molecular sieves prior to use. Solid amines were recrystallized before use. Acetonitrile used in kinetic and mechanistic experiments was carefully purified to remove traces of amines by stirring over anhydrous cupric sulfate for 24 h followed by distillation over phosphorus pentoxide.

Static spectrophotometric measurements were made on a Perkin-Elmer Lambda 9 scanning instrument. Time-based absorption measurements were made using a Cary 14D spectrometer and digitized using Un-ScanIt for data workup. Fast kinetic measurements were made using an OLIS RSM-1000 rapid scanning stopped-flow instrument. Gas chromatograms were recorded on a Hewlett-Packard 5890. GC/MS data were collected on an HP 5890 interfaced to a Finnegan mass analyzer. EPR measurements were made at 110 K in the X-band region using a Bruker instrument with PC data acquisition. Proton NMR measurements were made on a Bruker 200 MHz instrument. Electrochemical data were measured using a BAS CV-27 potentiostat and recorded on an X-Y chart recorder.

Product Identification. The organic products of amine oxidation were identified by comparing gas chromatagrams of reaction mixtures with those of known standards. For highly volatile products which could not be identified by GC, the following procedure was used. The reaction was run on a 1 mL scale in deuterated solvent under aerobic conditions. The solvent and volatile reaction products were condensed into a small trap immersed in liquid nitrogen. The material in this trap was collected and its NMR spectrum measured. Product labeling was accomplished by running the amine oxidation reaction under an atmosphere of labeled $^{18}O_2$, followed by mass analysis using GC/MS. The identity of the reduced dimanganese(III) product **2** was established by comparing its infrared, UV−visible, and paramagnetic NMR spectra to authentic samples of **2**.

Kinetic Measurements. The electronic absorption at 480 nm was monitored with time. This was the position of largest absorption change between **1** and **2**. The dependence of decay time on the initial concentrations of the amine and **1** was measured under pseudo-first-order conditions of excess amine. The experiments were repeated for the various amines studied in this paper.

EPR Trapping Experiments. EPR trapping was accomplished by rapidly quenching reaction mixtures in liquid nitrogen at various times after initiation of the reaction. In this way, an EPR time course for the reaction was determined. For radical trapping reactions, the reactions were run in the presence of nitrosobenzene, which readily forms stable radical adducts. These reactions were also quenched at various times by freezing, and the EPR spectra measured.

Electrochemical Measurements. The reduction potential for **1** has been previously reported,[12] but was repeated here under our reaction conditions. Cyclic voltammograms for the amines were measured under identical conditions using a platinum disk working electrode, an aqueous Ag/AgCl reference electrode, and a Pt wire counter electrode. Because only the oxidative wave was observed in the cyclic voltammograms, the reduction potential of the amine was taken as the potential at half-

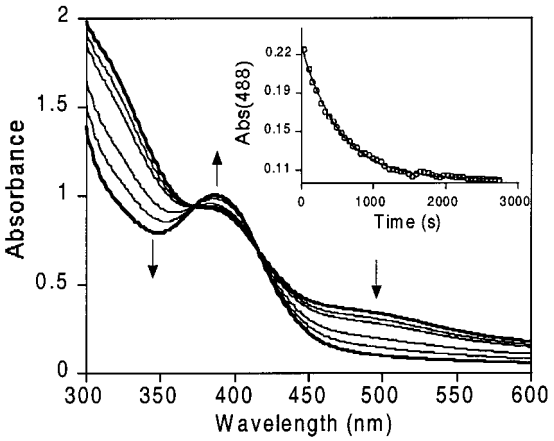


Figure 1. Time-dependent UV−visible spectra of **1** upon addition of excess tributylamine. Arrows indicate direction of spectrum changes with time. [**1**] = 0.10 mM, [bu_3N] = 1.0 mM, solvent = CH_3CN, T = 25 °C. Inset: Time-dependent absorbance change at 488 nm, same conditions. Solid line is the fit using the single-exponential decay equation $\Delta A = (A_0 - A_\infty)e^{-kt}$ to determine the observed pseudo-first-order rate constant k_{obs}.

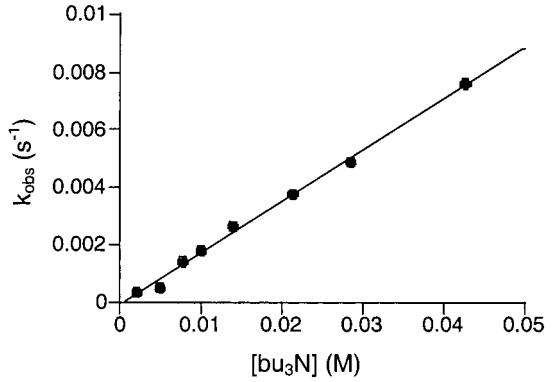


Figure 2. Pseudo-first-order rate constant, k_{obs}, as a function of tributylamine concentration. Slope = 0.178 M^{-1} s^{-1}, y-intercept = -4.1×10^{-5} s^{-1}. [**1**] = 0.10 mM, solvent = CH_3CN, T = 25 °C.

height for its positive-sweep voltammetry wave, and averaged at scan rates between 20 and 200 mV/s. The variation was never more than 30 mV.

Results

Figure 1 shows the time-dependent absorbance spectrum of **1** upon addition of excess tributylamine. The spectra are consistent with the one-electron reduction of **1** to give **2**, and the final spectrum is identical to that of **2** prepared independently. When the time-dependent absorption spectrum of **1** is monitored at 480 nm under pseudo-first-order conditions of excess tributylamine, a single-exponential decay is observed from which the pseudo-first-order rate constant for the conversion of **1** to **2** can be extracted, Figure 1 inset. Plotting the first-order rate constant as a function of amine concentration gave a linear plot whose slope is equal to the second-order rate constant for the reaction of **1** with tributylamine, Figure 2. When the initial concentration of **1** was varied under a constant excess of amine, no dependence on the concentration of **1** was observed. These data taken together show that the rate-determining step is first-order in amine and in **1**.

The rate for reduction of **1** by different amines was measured under identical conditions, and this data is summarized in Table 1 for tertiary amines of varying basicity and redox potential. The measured redox potentials and literature values for basicity[13] are also listed for each amine. In general, Table 1 shows that aliphatic amines are efficient reductants for **1**, but that aromatic amines such as aniline derivatives, pyridine, and imidazoles are ineffectual. The notable exception is 1,8-bis(dimethylamino)-naphthalene, DMAN, which has an unusually low redox potential and may proceed by a different mechanism (vide infra).

Table 1

amine	pK_a[a]	$E_{1/2}$ (V)[b]	rate constant (10^3 s^{-1})[c]
bu_3N	18.1	1.0	280
N-methylpiperidine	18.9	1.1	5.3
quinuclidine	19.5	1.3	28
DABCO[d]	18.2	0.90	29
N-methylaniline	11	0.96	0
DMAN[e]	20	0.57	7200
imidazole			0
methylimidazole			0

[a] In acetonitrile.[13] [b] Vs NHE. [c] Measured at 0.010 M amine and 1.0×10^{-4} M **1**. [d] 1,3-Diazabicyclononane. [e] 1,8-Bis(dimethylamino)naphthalene.

Under anaerobic conditions in the presence of nitrosylbenzene, a radical trapping reagent, reduction of **1** by tributylamine results in the formation of an organic radical that is persistent for several minutes and is clearly observable in the EPR spectrum of reaction mixtures. Under aerobic conditions, the radical intermediate is not observed and we instead observe the formation of dibutylformamide (DBF) by GC/MS, Figure S1A (Supporting Information). This product was formed in stoichiometric yield (1 DBF produced per **1** reduced) when measured in the GC against a decane internal standard. However, production of a quantitative yield required the presence of 1 atm of pure oxygen, and reactions run in air (20% O_2) gave lower yields. Reactions run in the dark gave identical final yields and proceeded at the same rate as determined by the appearance of dibutylformamide. Reactions run under nitrogen gave no dibutylformamide, Figure S1B. Reactions run under an atmosphere of $^{18}O_2$ showed incorporation of ^{18}O into dibutylformamide. Interestingly, the dibutylformamide product also showed oxygen uptake from ^{18}O-labeled water, which must occur through a reaction intermediate since control experiments showed no water exchange with dibutylformamide itself under our reaction conditions. The three-carbon cleavage product was shown to be propionaldehyde by NMR spectra measured on the volatile products of the reaction, and the formation of this product was also oxygen dependent. These data show the formation of an organic radical upon reaction of **1** with tributylamine which is quenched by oxygen to give dibutylformamide or by nitrosylbenzene to give a persistent radical adduct.

A change in the EPR spectrum of **1** is observed in the presence of a tertiary amine. The solid line in Figure 3 shows the EPR spectrum of **1** after mixing with tributylamine and rapid freezing to 77 K. The dotted line shows the EPR spectrum of **1** in the absence of amine for comparison. The high-field (g = 2) and low-field (g = 4−5) components have been interpreted to arise from ground state $S = {}^1/_2$ and excited state $S = {}^3/_2$ spin manifolds, consistent with the temperature dependence of the spectrum.[11] The largest change occurs at about g = 4.7, where a very weak signal in **1** is strongly enhanced in the presence of tributylamine, a change we associate with amine binding to **1**. This change is largely complete after addition of 2 equiv of tributylamine, suggesting a strong interaction between aliphatic

(13) Izutsu, K., Ed. *Acid-base dissociation constants in dipolar aprotic solvents*; International Union of Pure and Applied Chemistry: Oxford, 1990; pp 17−35.

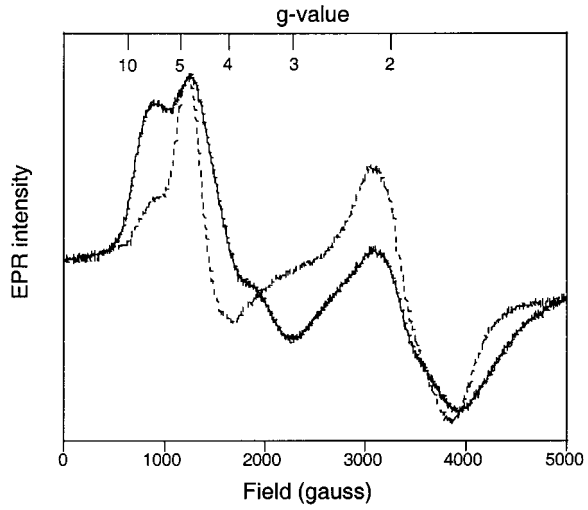


Figure 3. X-band EPR spectra of **1** in butyronitrile at 77 K, microwave power = 20 mW. Dotted line: **1** alone. Solid line: **1** + 5 equiv of tributylamine. This spectrum was measured by mixing **1** with tributylamine and immediately freezing the sample to 77 K. This ensured that negligible reduction of **1** took place during this experiment.

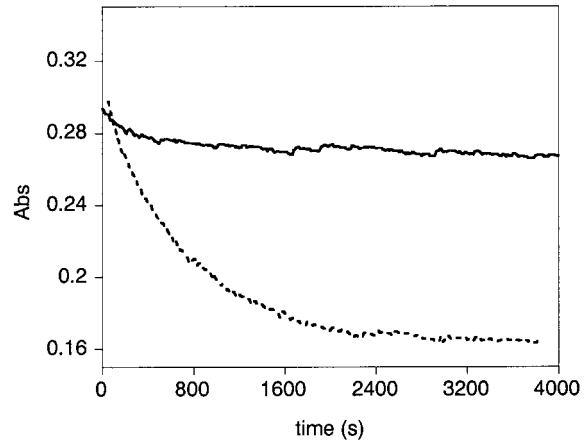


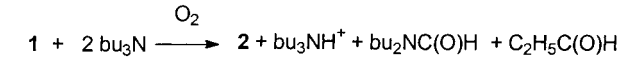
Figure 4. Absorbance decay curves of **1** with tributylamine at 488 nm. Dotted line: No hydroxide. Solid line: In the presence of 1 equiv of tetrabutylammonium hydroxide.

amines and **1** in acetonitrile. We have observed a similar EPR shift previously for this complex in the presence of other strong ligands such as hydroxide ion,[14a] which readily ligates in the solvent accessible site.[15] This is consistent with the observation in Figure 4 that reduction of **1** by amines is strongly inhibited by the presence of hydroxide ion, which binds strongly to **1**[14] and should prevent coordination of the amine. Only weak EPR changes were observed with the aromatic amines, which are not oxidized by **1**. The reaction with DMAN was too rapid to accurately assess any EPR shift associated with its interaction.

Discussion

The overall reaction of tributylamine with **1** proceeds according to Scheme 2. The reaction of other amines with **1** is similar, but tributylamine was chosen for extensive mechanistic analysis since its reaction products are more readily identified. In understanding the mechanism for the reaction in Scheme 2, it is useful to consider the reduction of the manganese complex and the quenching of the resultant amine radical separately.

Scheme 2



Scheme 3

1 + e^-	→	**2**	-12 kcal/mol
bu_3N:	→	$bu_3N^{+} + e^-$	+23 kcal/mol
1 + bu_3N:	→	**2** + bu_3N^{-+}	+11 kcal/mol

Scheme 4

1 + e^-	→	**2**	-12 kcal/mol
bu_3N	→	bu_3N^{-} + H^{-}	+84 kcal/mol
H^-	→	½ H_2	-52 kcal/mol
½ H_2	→	$H^+ + e^-$	0 kcal/mol
bu_3N + H^+	→	bu_3NH^+	-25 kcal/mol
1 + 2 bu_3N:	→	**2** + bu_3N^{-} + bu_3NH^+	-5 kcal/mol

Reduction of $Mn^{III}Mn^{IV}$(2-OHsalpn)$_2^+$. The reduction potential of **1** is 0.52 V vs NHE, which is substantially lower than the reduction potentials of all of the amines studied with the exception of the anomalous DMAN, which will be discussed separately. Any mechanism to account for reduction of **1** by tertiary amines must then be consistent with the very unfavorable thermodynamics for direct electron transfer from the amine to **1**. According to Table 1, there appears to be no clear correlation between amine reduction potential and ability to reduce **1**. In fact, direct oxidation of tributylamine by **1** to form the radical cation is disfavored by 11 kcal/mol, as shown in Scheme 3. However, a feature common to all of the amines that react with **1** is their strong basicity in acetonitrile solvent (see Table 1), which suggests that a proton transfer reaction may be critical in the reduction mechanism. The acidic properties of amine radical cations have been recognized,[8] and on the basis of available thermodynamic data in Scheme 3 and from literature,[16] we find that proton transfer from the tributylamine radical cation to tributylamine is favored by −16 kcal/mol. Coupling the electron transfer with subsequent proton transfer results in a net reaction that is thermodynamically favored by 5 kcal/mol, Scheme 4.

Scheme 4 indicates that the thermodynamic driving force for electron transfer could be provided by the subsequent proton transfer. The relationship between homolytic bond dissociation energy and acidity of organic molecules has been exploited extensively by Bordwell[16] and extended by Mayer[17] in the analysis of permanganate-promoted H-atom abstraction from aromatic hydrocarbons. By analogy, an alternative interpretation of Schemes 3 and 4 is to consider the tributylamine radical as arising by removing a hydrogen atom from tributylamine. However, the proton and the electron do not reside on the same species in the final products, and so our formulation is best considered as a proton-coupled electron transfer rather than as a hydrogen atom transfer reaction.

While Scheme 4 can be used to rationalize the thermodynamics of the system, its implication of a stepwise electron transfer

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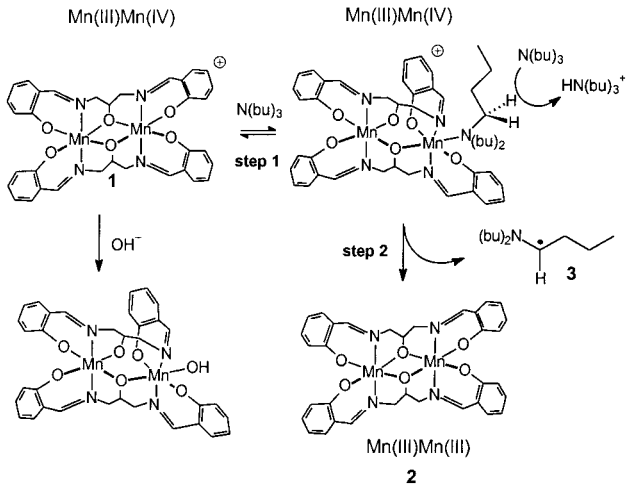
(15) The possibility that adventitious water was deprotonated by tributylamine to give the hydroxide-ligated complex was considered. However, under the conditions employed in this study, we do not obtain the characteristic change in the absorption spectrum of **1** that is observed upon hydroxide ligation (see ref 14).

(16) Homolytic bond dissociation energies were obtained from the following: *CRC Handbook of Chemistry and Physics, 1st Student Edition*; CRC Press: Boca Raton, 1988; pp F-124−127.

(17) (a) Bordwell, F. G.; Zhang, X. M. *Acc. Chem. Res.* **1993**, *26*, 510−517. (b) Bordwell, F. G.; Cheng, J. P. *J. Am. Chem. Soc.* **1988**, *110*, 1229−1231.

Scheme 5



followed by proton transfer cannot be interpreted as a literal mechanism. For such a reaction, the experimental second-order rate constant $k_{exp} = k_{ET}k_{PT}/k_{BET}$ (k_{ET} = electron transfer rate constant (M^{-1} s^{-1}), k_{BET} = back electron transfer rate constant (M^{-1} s^{-1}), k_{PT} = C−H_α proton transfer rate constant from the amine radical cation). The factor $k_{ET}/k_{BET} = K_{ET} = 1.0 \times 10^8$, which we derive from the thermodynamic driving force of 11 kcal/mol for direct electron transfer from tributylamine to **1**. For $k_{exp} = 0.178$ M^{-1} s^{-1}, a proton transfer rate constant of 1.8×10^7 M^{-1} s^{-1} is required. This is large when compared with the best present data on C−H_α proton transfer rates from tertiary amine cation radicals to amine bases, which range between 10^2 M^{-1} s^{-1} for aliphatic amines and about 10^6 for activated aromatic amines.[19] This would indicate that proton transfer is not likely to be sufficiently rapid in the present case to account for the observed rate constant for reduction of **1** by tertiary alkylamines via rate-limiting single electron transfer. We also deem it chemically unreasonable to expect proton transfer from the neutral amine to precede the electron transfer reaction in a stepwise reaction. We therefore conclude that a sequential reaction in which proton transfer and electron transfer reactions occur in separate steps is inconsistent with our rate data on amine oxidation by **1**.

We are therefore led to suggest a concerted pathway, in which both electron transfer and proton transfer reactions occur while the amine is still ligated to **1**. Our mechanistic proposition is illustrated schematically in Scheme 5. Step 1 in this scheme is consistent with our observed amine binding to **1** based on EPR spectroscopic changes. Coordination of the amine to **1** and interaction between an α-proton of the coordinated amine and the nitrogen of a free amine facilitates electron transfer to **1** and proton transfer to the free amine in step 2. This reaction is consistent with the proposition that metal coordination can stabilize the incipient amine radical cation,[6f] which is subjected to subsequent C^α−H proton transfer mediated by free amine to give the aminyl radical **3**. Step 2 is consistent with the 2:1 amine:**1** stoichiometry required to obtain quantitative yields and in principle circumvents the very unfavorable thermodynamics imposed on mechanistically separated proton transfer and electron transfer steps. However, to account for the first-order dependence on amine under our conditions, step 2 must be rate-limiting and step 1 must lie substantially to the right.

We interpret the hydroxide inhibition to result from coordination of OH^- to the solvent accessible site in **1**. This has been clearly observed in our previous work[14] showing that addition of OH^- to **1** results in EPR and visible spectroscopic shifts, as well as structural shifts observed by X-ray absorption spectroscopy, that are consistent with hydroxide ligation to this site. These studies show that the binding of neutral ligands such as tetrahydrofuran and water are weak when compared to hydroxide, and we expect that neutral amine binding would be weak in comparison to hydroxide as well. These data point to a role for hydroxide ion as a competitive inhibitor of amine oxidation by **1**, by blocking amine access to the metal cluster. However, it must be stressed that hydroxide ligation is also observed to decrease the reduction potential of **1** by 400 mV.[14b] As a result, we cannot entirely rule out the possibility that the inhibitory effect of hydroxide is due to depression of $E_{1/2}$ for **1** below that required for oxidation of the amine on the observed time scale.

Fate of the Amine Radical. Having established a viable mechanism whereby **1** is reduced by strongly basic alkylamines, we sought to develop a proposal for quenching of the produced radical to the observed products. Extensive characterization of the organic products from reaction of tributylamine with **1** showed that the only two products found in this case were *N*,*N*-dibutylformamide and propionaldehyde. Our products can only arise via oxidative cleavage of the C^α−C^β bond in the butyl group of tributylamine. This contrasts with most previously reported metal-mediated amine oxidation reactions which result in cleavage of the C^α−N bond to give amine dealkylation products, although small yields of dialkylformamide products are observed in the oxidation of tertiary amines by MnO_2.[20] The critical mechanistic data relating to the quenching of the amine radical are as follows: (1) no oxygenated products are formed in the absence of oxygen; (2) exclusively ^{18}O-labeled products are formed when the reaction is run under $^{18}O_2$ in anhydrous acetonitrile; (3) dibutylformamide shows some but not complete incorporation of oxygen from $H_2{}^{18}O$ when run under $^{16}O_2$; (4) formation of dibutylformamide lags behind reduction of **1**.

A mechanism based on literature precedent that we propose to account for these observations is shown in Scheme 6. In the first step, dioxygen attacks the amine α-radical to form an unstable peroxyl radical **4**. The formation of α-amino peroxides has been proposed in the aerobic oxidation of amines by metal complexes[5−7] and in the aerobic-oxygenation reactions of enamines that also result in C^α−C^β bond cleavage.[21] This radical is then reduced by a second dioxygen, followed by ring closure to form the 1,2-aminodioxetane, **5**, an intermediate in the singlet oxygenation of enamines that also results in C^α−C^β bond cleavage.[22] This intermediate is analogous to 1,2-dioxetanes which decompose thermally, sometimes explosively, with C^α−C^β bond cleavage to give aldehydes.[23] We suggest that the aminodioxetane decomposes via initial cleavage of the O−O bond to give the β-hydroxyamino intermediate **6**. This type of

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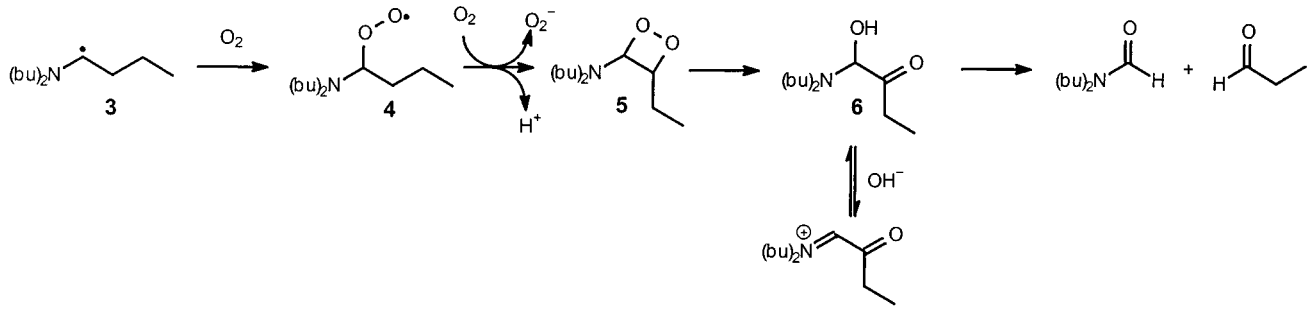
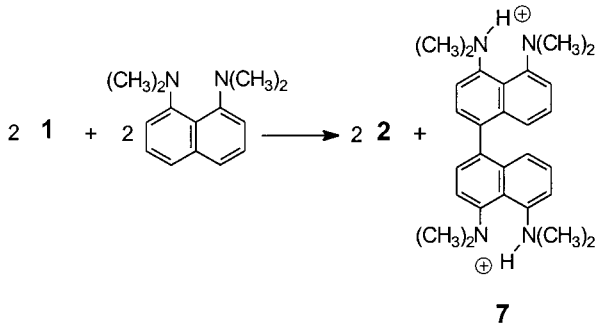
Scheme 6



Scheme 7



intermediate, suggested by Henbest and Stratford in the ozonolysis of tributylamine,[24] also provides a means for incorporation of oxygen derived from water into the dibutylformamide product via the ketoenamine species.

The results of these studies show that simple alkylamines are readily oxidized by redox-active metal centers of even moderate potential. This is consistent with the general observation that the lysine side chain is rarely employed as a ligand in metalloproteins, and never in redox-active metalloproteins. The preference in biology for the histidine imidazole group as a nitrogenous ligand is therefore consistent with the need for redox-stable ligands, as demonstrated by the observation that imidazoles do not reduce **1**, Table 1.

Reduction of 1 by DMAN. As illustrated in Table 1, DMAN is anomalous among aromatic amines in that it reduces **1** unusually rapidly. In fact, this occurs so rapidly that no radical intermediates could be observed by our techniques in this case. DMAN has a low redox potential, 570 mV, which is in a range where direct oxidation by **1** can occur on a kinetically competent time scale. DMAN furthermore has a high affinity for protons when compared to other aromatic amines as a result of its highly preorganized structure.[25] We therefore conclude that the reaction of **1** with DMAN proceeds via a direct electron transfer from DMAN to **1**. The manganese-containing products are identical to those obtained by reaction with tertiary amines. Analysis of the organic products showed them to consist of a mixture of $HDMAN^+$ and a radical coupling product that NMR analysis indicates is **7**. Therefore, this reaction is proposed to proceed as shown in Scheme 7. The dimerization via the para position to give **7** is consistent with experimental[26] and theoretical[27] studies on the oxidation of tertiary aromatic amines.

General Considerations

The experiments reported in this paper illustrate the complexity of amine oxidation by metal complexes. We have proposed a consistent mechanism for this reaction in nonaqueous media, but stress that these results may not be applicable to protic solvents. It is clear that the oxidation of amines takes a different path in nonaqueous media compared to aqueous solution since the resulting products are very different. Dealkylation of amines in aqueous solution is the result of a two-electron process that occurs via an iminium ion intermediate. In less polar solvents that do not stabilize ionic intermediates, radical pathways may be favored. This would be consistent with Scheme 5 where the oxidized amine is initially generated as its neutral radical. Because of the unique properties of **1** allowing controlled access to the metal ion by solvent and the amine, the results for this special case may also not be applicable to simpler metal-containing oxidants in general. However, it is because of the unique features of **1** in aprotic solvents that we suggest that this system may provide insight into the highly controlled oxidation of amine species in biological systems.

Amine inhibition of oxygen-evolving activity in PS II arises from apparently two effects. Small amines are apparently able to displace functionally necessary chloride from a site in the manganese cluster,[3] which precludes oxidation state advancement of the manganese cluster. This bears cursory analogy to the coordination of amines to **1** through the solvent labile site. In addition, it has been observed by several groups that the manganese cluster in photosystem II is irreversibly reduced by the presence of Tris buffer and small amines, resulting in the liberation of Mn(II).[4] Secondary and tertiary amines are apparently too bulky to enter the active site. This latter reaction is analogous to the reduction of **1** by the model amines. Of note, reduction of manganese in photosystem II is most facile when initiated from the more oxidizing S_2 state.[28] This is consistent with our observation in model chemistry that oxidation of the amine requires the oxidizing power of the Mn(IV) ion in **1**, and that the reduced complex **2** is unaffected by amines. Therefore, the reduction of **1** by amines is shunted into one-electron pathways by the relatively lower reduction potential of the dimanganese(III) complex **2**.

This highlights an important difference between the model complex **1** and the OEC in terms of the final oxidation state of the cluster. Reduction of **1** stops at the dimanganese(III) oxidation level, whereas Mn(II) is produced from PS II. This would indicate the presence of Mn(III) sites of high potential in PS II, which are capable of mediating further redox chemistry with amines. It has been suggested that oxidation of small amines by PS II proceeds via a two-electron process,[29] which would be consistent with coupled high-potential sites. However, in aqueous solution, differentiating one- vs two-electron path-

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ways is complicated by subsequent hydrolytic chemistry. A two-electron mechanism for manganese reduction by dihydroquinone or hydroxylamine is also postulated,[30] although these reductants appear to attack different manganese in the OEC as a result of differential access to the cluster or to the different redox potential requirements. In any case, it is likely that the precise mechanism for manganese reduction in the OEC depends on the particular reductant. Potent reductants with direct access to manganese ions can react via an inner-sphere mechanism, and probably by two-electron mechanisms. However, our data would indicate that amines and amine buffers such as Tris probably react via one-electron pathways, which may be obscured in biological systems by subsequent hydrolytic chemistry characteristic of the aqueous media.

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Supporting Information Available: Gas chromatograms of the reaction products (Figure S1). This material is available free of charge via the Internet at http://pubs.acs.org.

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