HYDROGEN PRODUCTION FROM MODEL COMPLEXES OF THE [FEFE]- AND [NIFE]-HYDROGENASE ACTIVE SITES

BY

BRYAN E. BARTON

DISSERTATION

Submitted in partial fulfillment of the requirements for the degree of Doctor of Philosophy in Chemistry in the Graduate College of the University of Illinois at Urbana-Champaign, 2010

Urbana, Illinois

Doctoral Committee:

Professor Thomas B. Rauchfuss, Chair
Professor Gregory S. Girolami
Professor Andrew Gewirth
Professor John F. Hartwig
Abstract

Facing a global energy crisis, many chemists have envisioned molecular hydrogen as an efficient and environmentally friendly fuel of the future. However, for the hydrogen economy to become viable, new hydrogen-processing catalysts are needed to replace platinum, which is costly and of limited supply. Nature offers direction toward this goal, as enzymes called hydrogenases evolved several billion years ago to utilize molecular hydrogen as a fuel. The [FeFe]-hydrogenases predominantly function to produce hydrogen from protons and electrons, while the [NiFe]-hydrogenases function to oxidize hydrogen. The active sites of both enzymes contain first-row transition metals and biologically exotic ligands. When assayed for their rates and efficiencies of hydrogen processing, the hydrogenase enzymes are directly comparable to platinum. Unfortunately, despite several crystal structures and a wealth of spectroscopic techniques, the mechanism of hydrogen processing remains speculative. Our goals as synthetic chemists have focused on the reactivity of models for the [FeFe]- or [NiFe]-hydrogenases in hopes of understanding of how Nature tunes these first-row transition metals into phenomenal catalysts.

Interestingly, active site models for the [FeFe]-hydrogenases were unknowingly present before the first crystal structure in 1999. In fact, the structurally similar diiron dithiolate hexacarboxyls had been investigated since the 1920s and had well-established chemistry. However, unique compared to all other diiron dithiolates, the active-site structure of [FeFe]-hydrogenase features a rotated diiron dithiolate core, exposing a vacant terminal position. Carbon monoxide binds to this terminal position
and inhibits catalysis. Thus, the mechanism of hydrogen processing is thought to occur by substrate (H₂, H⁺) binding in the terminal position. To properly model the biological mechanism of [FeFe]-hydrogenase, we sought terminal hydrides of diiron dithiolates. After the first terminal hydride complex, [HFe₂(edt)(CO)₂(PMe₃)₄]⁺, was published from our group others quickly followed. This new class of diiron dithiolate terminal hydrides was derived by the biologically relevant pathway of protonation of a Fe(I)Fe(I) precursor. Unfortunately, the terminal hydrides derived in this fashion were unstable, and when warmed above –80 °C quickly isomerized to isomeric bridging hydrides.

To understand and control the selective formation of terminal hydride species and subsequent isomerization pathway, a series of diiron dithiolates were investigated. All diiron dithiolates explored showed the kinetic formation of a terminal hydride species that subsequently isomerized via a series of turnstile rotations to bridging hydrides. We learned that these turnstile rotations were controlled by a combination of electronic and steric effects, as the 1,3-propanedithiolate derivatives were vastly more stable than their corresponding 1,2-ethanedithiolate derivatives. Additionally, more phosphine ligands generally resulted in a more stable terminal hydride. Thus, protonation of Fe₂(pdt)(CO)₂(dppv)₂ provided the terminal hydride complex [(t-H)Fe₂(pdt)(CO)₂(dppv)₂]⁺, which isomerized at room temperature (t₁/₂ ~ 10 min) to the bridging hydride [((µ-H)Fe₂(pdt)(CO)₂(dppv)₂]⁺. With a pseudo-stable terminal hydride complex in hand, we sought to explore the catalytic mechanism of proton reduction via the terminal hydride. To our surprise, although the mechanism was very similar to that proposed in biology, the catalytic efficiency suffered greatly when compared to the bridging hydride complex. We continued focusing our research efforts on proton relay.
The active site of [FeFe]-hydrogenase is speculated to contain a 2-azapropane-1,3-dithiolate as the bridging dithiolate ligand, although the exact identity of the bridgehead atom could be either carbon, nitrogen, or oxygen. Recent work on mononuclear nickel phosphines led to the impression that an azadithiolate (adt) could function as a proton relay lowering the kinetic barrier of proton transfers to and from the terminal hydride position. Due to the significant amount of steric congestion in [(H)Fe₂(pdt)(CO)₂(dppv)₂]⁺, the iron hydride does not deprotonate with tetramethylguanidine (pKₐ = 26) and requires the strong acid HBF₄·Et₂O (pKₐ = -2) for its formation. Upon incorporation of the proposed azadithiolate cofactor, [(H)Fe₂(adt)(CO)₂(dppv)₂]⁺ was observed to have a significantly smaller barrier for proton transfer to and from the terminal position. In addition, [(H)Fe₂(adt)(CO)₂(dppv)₂]⁺ was observed to be a remarkably fast and efficient catalyst for proton reduction, with turnover frequencies approaching that of the enzyme.

Unlike the [FeFe]-hydrogenases, model complexes for the [NiFe]-hydrogenases were unknown prior to the crystal structure in 1996. However, most synthetic efforts focused on structural models for the active site, and neglected the catalytically imperative hydride ligand. Thus, we sought a nickel-iron hydride complex to explore the relevant reactivity of the first (µ-H)Ni(µ-SR)₂Fe complex. We found that the previously reported (dppe)Ni(µ-pdt)Fe(CO)₃, a Ni(I)Fe(I) complex, reacted with acid to provide [(dppe)Ni(µ-H)(µ-pdt)Fe(CO)₃]⁺, the first nickel-iron hydride. After protonation, the hydride complex is amenable to substitution chemistry at the Fe(CO)₃ subunit. Further derivatives altering the Ni(diphosphine)(SR)₂ subunit have been achieved through an alternative synthetic procedure to the Ni(I)Fe(I) complex. All nickel-iron hydrides
investigated are active catalysts for the reduction of protons. As the catalytic mechanism of [NiFe]-hydrogenase is widely speculative, the reactivity of this new class of nickel-iron hydrides offers powerful insights into Nature's catalytic mechanism.
To my wife Ashley, and daughter Layla—without whom my life would be meaningless
Acknowledgements

Once someone has decided to pursue a Ph.D. in Chemistry they are met with little pay, long hours, and a complete loss of self-confidence. Perhaps the design is intentional, as those who continue this pursuit are absolutely committed to a career in research. However, research is a love-hate relationship, some days the results are so invigorating that you feel like your chemistry can change the world. This feeling is typically shortly followed with the horrible realization that you have either done something incorrectly, or that you don’t understand the chemistry at all. During these dramatic emotional changes a stable support system is a requirement. Thankfully, I have had an encouraging advisor, good friends, and above all a supportive family.

I would first like to first thank my advisor, Tom Rauchfuss. You have always given me the right amount of positive and negative criticism to keep me learning and working, while keeping my optimism in check. I would also like to thank the remaining members of my committee: Professors Greg Girolami, Andy Gewirth, and John Hartwig. You have provided very helpful discussions on my research and guided me to explore areas I would have otherwise not have considered.

I would also like to thank past and present members of the Rauchfuss group. The Rauchfuss group has always maintained a positive and humorous workplace environment, without which I would have been very depressed. I would also like to specifically thank Aaron Justice and Chrissy Boyke for training me to understand that I knew very little about chemistry and that I needed to talk to others and read, read, read. This was a pivotal step toward my Ph.D. and only after this has my education progressed.
From the beginning of teaching freshman chemistry, I have had a strong network of friends. Without the football, basketball, and frequent lunch outings life in Champaign-Urbana would have been much less tolerable. Because of you my time here has been filled with wonderful memories of playing snow tackle football, tailgating, my first hockey game, and late-nights.

Lastly I have to thank my family for their continued support and encouragement. My parents, being just an hour away, have always encouraged me to enjoy sunny days outside and to take vacations. Above all friendships and parental support, I thank my wife Ashley for showing me that although I am a chemist, chemistry is not my life. Beyond all my accomplishments in school I am most proud of our marriage, and now with our now 8 month old daughter, Layla, our family is the best part of my life.
# Table of Contents

Chapter 1: Mechanistic Insights for the [FeFe]-Hydrogenases: Nature’s Fastest Hydrogen Catalyst.................................................................1

Chapter 2: Terminal Hydrides of Diiron Dithiolates as Kinetic Products of Protonation.................................................................................14

Chapter 3: Reactivity of Terminal Hydride Complexes and Functional Azadithiolate Proton Relays.................................................................60

Chapter 4: Proton Reduction Catalysis by Azadithiolate Terminal Hydrides........106

Chapter 5: The Very Electron-Rich Disubferrous Dithiolate: Fe\textsubscript{2}(pdt)(CNMe)\textsubscript{6}........143

Chapter 6: The Development of Functional Models for [NiFe]-Hydrogenase.......................................................................................159

Chapter 7: Hydride-Containing Models for the Active Site of the Nickel-Iron Hydrogenases.................................................................169

Chapter 8: New Routes to Nickel-Iron Dithiolate Hydrides.................................237
Chapter 1

Mechanistic Insights for the [FeFe]-Hydrogenases: Nature’s Fastest Hydrogen Catalyst

During the development of the earliest lifeforms on earth, the atmosphere was highly reducing and hydrogen was abundant. Although the exact process is not understood, iron-sulfur minerals were incorporated into enzymes, which utilized the reductive atmosphere to perform redox reactions.\(^1\) Hydrogen processing enzymes, known as hydrogenases, mediated both hydrogen oxidation and hydrogen production (Eq. 1.1). Today, the hydrogenase enzymes remain prevalent in anaerobic environments and typically function to expel additional reducing equivalents when typical oxidants such as \( \text{O}_2 \), sulfate, \( \text{CO}_2 \), and nitrate are not present.\(^2\) There are three evolutionarily distinct hydrogenase enzymes, the [FeFe]-, [NiFe]-, and [Fe]-hydrogenase, each named for the metal(s) of the active site. As [FeFe]- and [NiFe]-hydrogenase are the most prevalent in biology, our work has focused on these two enzymes.\(^3\)

\[
\begin{align*}
\text{H}_2 & \quad \text{Hydrogen Oxidation} \quad \text{2H}^+ + 2\text{e}^- \quad (\text{Eq. 1.1}) \\
\text{H}_2 & \quad \text{Hydrogen Production} \\
\end{align*}
\]

Industrially, hydrogen gas is produced from the steam reforming of hydrocarbons or high temperature water electrolysis.\(^4\) However, for numerous environmental and economic reasons chemists aspire to produce hydrogen from water, a carbon-free source. In the reaction protons are reduced at one electrode to give hydrogen and water is oxidized at the other to provide molecular oxygen. For this water-splitting reaction to
be feasible the individual catalysts for each half-reaction must be thermodynamically efficient. The standard catalyst for the processing of hydrogen is platinum, and although very efficient, its cost and scarcity prevent its utilization in mass-production.\textsuperscript{5} Recently catalysts have been designed that perform efficiently without platinum,\textsuperscript{6} and even without a metal.\textsuperscript{7} The hydrogenase enzymes display efficiencies and rates of hydrogen processing that are very similar to platinum,\textsuperscript{8} and use cheap and abundant first-row metals (Table 1).\textsuperscript{9}

\textbf{Table 1.1.} Comparison of catalytic rates for hydrogenase enzymes (turnovers per second per active site).\textsuperscript{8}

<table>
<thead>
<tr>
<th></th>
<th>H\textsubscript{2} Oxidation</th>
<th>H\textsubscript{2} Production</th>
</tr>
</thead>
<tbody>
<tr>
<td>[FeFe]-Hydrogenase</td>
<td>28,000</td>
<td>6,000-9,000</td>
</tr>
<tr>
<td>[NiFe]-Hydrogenase</td>
<td>700</td>
<td>700</td>
</tr>
</tbody>
</table>

The [FeFe]-hydrogenase enzyme contains several [4Fe4S] clusters that serve as a kind of wire to deliver electrons to the center of the protein where the [2Fe2S] active site performs the hydrogen processing (Figure 1.1).\textsuperscript{3,10} Hydrogen processing in this context refers to both hydrogen oxidation and proton reduction. The [2Fe2S] active site contains several biologically unique ligands and cofactors. The active site contains both cyanide and carbon monoxide ligands, both of which, as their free molecules, are typically toxic to organisms.\textsuperscript{3} In addition, a [4Fe4S] cluster is directly attached to the [2Fe2S] cluster through a cysteinate residue forming the so-called “H-cluster”. The exact identity of the central atom “X” in the dithiolate cofactor in Figure 1.1 is highly debated but is either carbon, nitrogen (NH, azadithiolate), or oxygen based on X-ray crystallography.\textsuperscript{11} In the fully reduced state, “H\textsubscript{red}”, the diiron center is in the Fe(I)Fe(I) state, and contains a “rotated” geometry exposing a vacant terminal site.
Although the detailed mechanism of the enzymatic reaction is unknown, substrate turnover is thought to occur only on the single iron atom furthest from the [4Fe4S] cluster and undergo single electron steps.\textsuperscript{12} A consistent mechanism is presented in Figure 1.2. Here, catalysis involves terminal hydrides on the distal iron center and the dithiolate cofactor is presented as an azadithiolate and serves to relay protons for the heterolysis of H\textsubscript{2}.

Since the structure of the active site of [FeFe]-hydrogenase was published in 1999,\textsuperscript{13} synthetic chemists have aimed to mimic its behavior with diiron dithiolate complexes.\textsuperscript{14} Interestingly, diiron dithiolate carbonyls had been known since the 1920s,\textsuperscript{15-17} so the transition to functional models that catalytically reduced protons to hydrogen was short. Since the report of the first catalytically active diiron dithiolate,\textsuperscript{18} several different mechanisms have been published.
Catalytic Fe\textsubscript{2}(SR)\textsubscript{2}(CO)\textsubscript{6} “models” are pervasive,\textsuperscript{19} however more interesting from the perspective of biochemistry are complexes that incorporate azadithiolate’s,\textsuperscript{20,21} hydrides,\textsuperscript{22} and cyanide ligands\textsuperscript{18,23}. These components are rarely realized in one model and models involving terminal hydrides are unknown (Figure 1.3). After the discovery that cyanide-containing models protonate at the cyanide ligand and are slow to perform proton reduction catalysis,\textsuperscript{18,23} our modeling efforts migrated towards the utility of phosphine ligands to simulate the electronic environment of the active site. Although both the fully reduced state of the enzyme, “H\textsubscript{red}”, and model complexes are in the Fe(I)Fe(I) oxidation state, the active site features a “rotated geometry”, exposing a vacant terminal position. Diiron dithiolato carbonyls with several (3 or 4) phosphine ligands are accurate models in terms of their CO stretching frequencies and the potential of the Fe\textsubscript{2}\textsuperscript{I,I/II} redox event (Table 1.2). When, however, Fe\textsubscript{2}(SR)\textsubscript{2}(CO)\textsubscript{x}(L)\textsubscript{6-x}
models approach properly modeling the νCO frequencies, a ~400 mV gap is observed between the model’s Fe$_2^{I/II}$ couple to that of the enzyme (Figure 1.4). This Fe$_2^{I/II}$ couple was observed to be highly dependent on the size of the dithiolate ligand$^{24}$ and larger dithiolates are thought to stabilize a rotated Fe(I)Fe(I) center.$^{25}$ Thus, the origin of the 400 mV gap in $E_{1/2}$’s is likely due to Nature’s ability to stabilize the rotated structure. Despite these differences and initial indications that phosphine-substituted models were worse catalysts, we studied the catalytic mechanism of these substituted derivatives with the goal of gaining a better understanding of the natural system.

![Figure 1.3. Diiron dithiolate model complexes that incorporate adt, CN$^-$, phosphines, and hydrides.](image-url)
Table 1.2 Comparison of vCO frequencies, average vCO frequency, and the potential for the I,II,I couple of diiron dithiolate model complexes with that of the fully-reduced enzyme.

<table>
<thead>
<tr>
<th>FT-IR vCO (cm⁻¹)</th>
<th>Average vCO (cm⁻¹)</th>
<th>( E_{1/2}^{(I),(II)}/E_{1/2}^{(I)} ) vs Fc⁰+/⁺ (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>( D. \text{ desulfuricans, &quot;H}_\text{red}&quot; )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>( \text{Fe}<em>2(pdt)(CO)</em>{6} )</td>
<td>1965, 1916, 1895</td>
<td>1925</td>
</tr>
<tr>
<td>( \text{Fe}<em>2(pdt)(CO)</em>{4}(dppv) )</td>
<td>2076, 2035, 1997 ( ^a )</td>
<td>2036</td>
</tr>
<tr>
<td>( \text{Fe}<em>2(pdt)(CO)</em>{4}(PMe_3)_2 )</td>
<td>2021, 1950, 1912 ( ^a )</td>
<td>1961</td>
</tr>
<tr>
<td>( [\text{Fe}<em>2(pdt)(CN)</em>{2}(CO)_{4}]^{2-} )</td>
<td>1982, 1943, 1899 ( ^a )</td>
<td>1941</td>
</tr>
<tr>
<td>( \text{Fe}<em>2(pdt)(CO)</em>{3}(dppv)(PMe_3) )</td>
<td>1961, 1917, 1880 ( ^a )</td>
<td>1919</td>
</tr>
<tr>
<td>( \text{Fe}<em>2(pdt)(CO)</em>{2}(dppv)_2 )</td>
<td>1943, 1892 ( ^a )</td>
<td>1917</td>
</tr>
<tr>
<td>( 1888, 1868 ( ^a )</td>
<td>1878</td>
<td>-0.88 ( ^d )</td>
</tr>
</tbody>
</table>

a) Recorded in CH₂Cl₂ solution.
b) Ref # 12
c) Ref # 18
d) Ref # 24

Figure 1.4. Correlation of 1,3-propanedithiolate diiron dithiolate model complexes with that of the fully-reduced enzyme “\( H_{\text{red}} \)”. Data from Table 1.2.
Catalysis by hexacarbonyl complexes, e.g. Fe₂(pdt)(CO)₆, proceeds via initial reduction to give the anion \([\text{Fe}_2(SR)_2(CO)_6]^-,\) which then is protonated. In contrast, the phosphine-substituted derivatives \((\text{Fe}_2(SR)_2(CO)_{6-x}(PR_3)_x)\) enable protonation as the initial step to form \([(\mu-H)\text{Fe}_2(SR)_2(CO)_{6-x}(PR_3)_x]^+\). This first step in the catalytic mechanism is reminiscent of that proposed for the enzyme, although the stereochemistry of the hydride should be in the terminal position. Since 1976 diiron dithiolato carbonyls have been known to protonate at the metal-metal bond (Figure 1.5). A terminal hydride complex has recently been synthesized, albeit by the treatment of a diferrous complex with a hydride reagent. The terminal hydride \([(t-H)\text{Fe}_2(edt)(CO)_2(PMe_3)_4]^+\) (where \(t-H\) denotes terminal hydride) was observed to isomerize to the bridging position at room temperature, establishing that thermodynamically stable terminal hydrides would be difficult to prepare. This kinetically unstable terminal hydride complex displayed uniquely liberates \(H_2\) upon treatment with \(\text{HBF}_4\cdot\text{Et}_2\text{O}\), whereas the bridging hydride complex does not (Figure 1.6).

Figure 1.5. Depiction of Fe₂ “banana-bond” of Fe₂(SR)₂(CO)₄(PMe₃)₂ and its subsequent protonation.
Shortly after this initial communication, terminal hydrides were observed as kinetic products of protonation for $\text{Fe}_2(pdt)(\text{CO})_4(dppe)$, suggesting that the more reactive terminal hydrides could be incorporated into catalysts that are mechanistically relevant to Nature (Figure 1.7).\textsuperscript{27} However, $[\{(t\text{-}H)\text{Fe}_2(pdt)(\text{CO})_4(dppe)]^+$ is only stable at temperatures below $-70^\circ\text{C}$, and isomerizes to $[\{(\mu\text{-}H)\text{Fe}_2(pdt)(\text{CO})_4(dppe)]^+$. The mechanism of protonation is thought to involve a rotated Fe(I)Fe(I) center. This proposal is supported as a similar complex, $\text{Fe}_2(pdt)(\text{CO})_4(dppv)$, adopts a rotated structure during substitution reactions.\textsuperscript{28} Rotated Fe(I)Fe(I) states had been calculated to be thermally-accessible conformers that were stabilized by a highly-asymmetric electronic environment.\textsuperscript{29}
Several features of Nature’s terminal hydride fit in context with a wealth of information obtained from transition-metal hydride complexes. In an octahedral ligand framework, the reactivity of the hydride is highly sensitive to the \textit{trans} ligand.\textsuperscript{30} For example, \textit{trans}-[HFe(dppe)$_2$(CO)]$^+$, protonates only by the strong acids HOTf or HBF$_4$•Et$_2$O to give [(H$_2$)Fe(dppe)$_2$(CO)]$^+$.\textsuperscript{31} However, replacement of the CO ligand with a more donating ligand, such as CH$_3$CN or F$^-$ results in dramatic shifts in the Fe-(H$_2$) p$K_a$.\textsuperscript{32} In the active site of [FeFe]-hydrogenase the proposed terminal hydride is \textit{trans} from a CO ligand, which forces the Fe-(H$_2$) species to be very acidic, and susceptible to heterolytic cleavage. Additionally, the [FeFe]-hydrogenases incorporate redox-active [4Fe4S] ligands. Although the effect of changing a ligand oxidation state on the reactivity of a metal hydride is not well understood, changing the metal’s oxidation state results in a change of the hydride p$K_a$ by up to 20 units.\textsuperscript{33}

In addition to these direct comparisons, thermodynamic analysis of transition-metal hydrides has proven useful in predicting reactivity. Such calculations combine metal hydride p$K_a$’s and electrochemical data to describe the ability of the metal hydride to donate H$^-$ ($\Delta G_{\text{H}^-}$) or H$^+$ ($\Delta G_{\text{H}^+}$).\textsuperscript{34} This understanding has been used to tune mononuclear Ni phosphine complexes from proton reducers to H$_2$ oxidation catalysts.\textsuperscript{35} Although a powerful predictive tool, the basis is not always useful as kinetic barriers occasionally prevent the reaction from proceeding under relevant conditions.\textsuperscript{36} This gap in predictive utility is particularly relevant to diiron dithiolates as model complexes for the [FeFe]-hydrogenases as they often employ phosphine ligands, where steric bulk can hinder reactions at the metal center.
The [FeFe]-hydrogenases utilize iron, a poor hydrogen processing catalyst, and transform it into being one of the fastest and most efficient hydrogen processing catalysts known. Our goal for the inorganic modeling of the [FeFe]-hydrogenases is to understand how Nature performs this remarkable transformation of reactivity. To address this, we synthesize and investigate the reactivity of diiron dithiolato carbonyls that closely resemble the electronic and/or structural environment of the active site. The methods by which the [FeFe]-hydrogenases use to transform the reactivity of iron to that of platinum is essential to the development of iron-based catalysts, and potentially broadly applicable to other catalytic processes.

References:


Chapter 2

Terminal Hydrides of Diiron Dithiolates as Kinetic Products of Protonation

Introduction

Protonation of di- and polynuclear compounds almost invariably results in products containing µ-hydride ligands. Kinetic barriers for protonation arise from reorganizational energies, which are expected to be lower for protonation at main group vs transition metal centers. Consistent with this view, low temperature studies on the protonation of anionic metal carbonyls indicate that the carbonyl oxygen is often the kinetic site of protonation followed by proton transfer to the metal. Polynuclear complexes, in principle, contain multiple possible sites of protonation, each with their own kinetic and their thermodynamic preferences. The interchange of bridging and terminal hydrides in clusters is well established (e.g. for \( \text{H}_2\text{Os}_3(\text{CO})_{10}(\text{PPh}_3) \)). Additionally, the (kinetic) reactivity of hydride ligands is well known to be affected by their geometry: bridging hydrides are more acidic than terminal hydrides. A parallel pattern has long been noted for the boron hydrides. Some hydrides are known to exist as isomers (Scheme 2.1). The present results suggest that protonation of more complex metal clusters, which invariably give µ-hydrides, could occur at a single metal followed by rapid transfer of the hydride to a bridging position.
Scheme 2.1. Two isomers of HFe(CO)$_3$(µ-PR$_2$)Pt(PR$_3$)$_2$.

The structure of dimetallic hydrides is relevant to the [FeFe]-hydrogenases, which feature a pre-rotated Fe(I)Fe(I) center that facilitates binding and turnover localized at one vacant terminal site (Scheme 2.2).$^7$ As even highly-developed diiron dithiolate model complexes lack the rotated entatic structure we proposed that a highly-assymmetric electronic environment may be required to achieve the entatic state.$^8$ A DFT investigation has supported this theory finding that chelating ligands on a single iron center modestly stabilized the rotated state and thus were more accessible as a excited conformer at room temperature.$^9$ Despite the lack of direct evidence for a rotated Fe(I)Fe(I) conformer, recent protonation studies provide an indirect argument for their existence.

Scheme 2.2. Protonation of the rotated Fe(I)Fe(I) state of [FeFe]-hydrogenase providing a terminal hydride complex.
The recently described diiron(I) dithiolates $\text{Fe}_2(S_2C_nH_{2n})(\text{CO})_2(\text{dppv})_2$ [$n = 2$ (1), 3 (2); dppv = cis-1,2-bis(diphenylphosphino)ethene] have recently been shown to react with Lewis acids to provide rotated complexes.\textsuperscript{10} It is however unclear if the Lewis acid binds first and then the complex isomerizes to the rotated state or if the rotate state is thermally accessed and then binds the Lewis acid. Perhaps indicating the former, DFT calculations indicate that the rotated state lies $\sim$10 kcal/mol higher in energy however the reaction with Lewis acids occurs immediately at $\sim$80 °C.\textsuperscript{10}

The diiron(I) dithiolates are fluxional species wherein the various rotamers could exhibit differing basicities.\textsuperscript{11,12} However, it has long been understood that diiron(I) dithiolates protonate to give $\mu$-hydride derivatives. Early studies by Poilblanc led to the phosphine-substituted complexes such as $[\text{Fe}_2(\mu-H)(\text{SMe})_2(\text{CO})_4(\text{PPhMe}_2)_2]^+$.\textsuperscript{13} After these compounds were found to resemble the active sites of the [FeFe]-hydrogenases,\textsuperscript{14} analogous protonations were described for the related ethane- and propanedithiolato diiron complexes, all of which provided bridging hydrides.\textsuperscript{14-16}

The first terminal hydride complex, $[(t-H)\text{Fe}_2(\text{edt})(\text{CO})_2(\text{PMe}_3)_4]\text{BF}_4$ (where edt = 1,2-ethanedithiolate), was not generated by protonation, but by the treatment of $[(\text{CH}_3\text{CN})\text{Fe}_2(\text{edt})(\text{CO})_2(\text{PMe}_3)_4]\text{BF}_4$ with LiAlH$_4$.\textsuperscript{4} This complex is not thermodynamically stable and was found to isomerize unimolecularly to the $\mu$-hydride isomer at room temperature with a halflife of minutes. The relatively long lifetime permitted crystallographic characterization. The regiochemistry of the hydride ligand defines its reactivity, as the terminal hydride isomer is more susceptible to protonation than the $\mu$-hydride isomer.
Recently, the low temperature protonation of several diiron(I) dithiolates has revealed the intermediacy of terminal hydrides, which subsequently form the more stable µ-hydrides. The asymmetrically substituted Fe$_2$(pdt)(dppe)(CO)$_4$ (where pdt = 1,3-propanedithiolate) has been shown to protonate at − 75 °C with the strong acid HBF$_4$•Et$_2$O at a single Fe site to afford a terminal hydrides.$^{17}$ The high-field $^1$H NMR spectrum displays a singlet at δ - 4.33, which has a similar chemical shift to that of [(t-$^3$H)Fe$_2$(edt)(CO)$_2$(PMe$_3$)$_4$]BF$_4$.

In this chapter, we describe and analyze protonations of a series of substituted diiron(I) dithiolates to better understand the factors leading to the preferential formation of terminal hydride species and the pathway for conversion of terminal- to µ-H species. Reactions were monitored at low temperatures by both $^1$H and $^{31}$P NMR spectroscopies, techniques that usually provide definitive structural assignments, allowing us to deduce reaction sequences. Our results point to a general mechanism for the protonation of phosphine-substituted diiron(I) dithiolates, one that begins with terminal protonation.
Results and Discussion

**Isomers of [HFe₂(edt)(CO)₄(dppv)]BF₄ and [HFe₂(pdt)(CO)₄(dppv)]BF₄.** To provide reference spectra for relatively simple systems, we first examined the protonation of dichloromethane solutions of Fe₂(edt)(CO)₄(dppv) (1) and Fe₂(pdt)(CO)₄(dppv) (2) with HBF₄·Et₂O. The structures of the resulting hydrides can be deduced from the coupling patterns since in diiron dithiolato hydrides (and other mononuclear iron hydrides), trans $J_{P\text{-hydride}}$ coupling constants are smaller ($J_{PH} \sim 5$ Hz) than the corresponding cis coupling constants ($J_{PH} > 25$ Hz).¹⁸,¹⁹ When protonations of 1 and 2 were monitored at lower temperatures, we observed intermediates. In the case for the edt derivative (1), the low temperature $^{31}$P NMR spectrum featured two signals, and the corresponding $^1$H NMR signal at $\delta$ -16.1 ($dd$, $J_{PH} = 26$, 6 Hz). This pattern is consistent with dppv spanning apical and basal sites (Scheme 2.3). Upon warming to room temperature, this hydride quantitatively converted to a new isomer with a single $^{31}$P NMR signal at $\delta$ 85 and $^1$H NMR signal at $\delta$ -14.7 ($t$, $J_{PH} = 21$ Hz) (Figure 2.1).

![Scheme 2.3. Products formed upon protonation of Fe₂(pdt)(CO)₄(dppv).](image)
Figure 2.1. $^1$H NMR (500 MHz) spectra of a CD$_2$Cl$_2$ solution of 1, Fe$_2$(edt)(CO)$_4$(dppv), after protonation with 3 equiv HBF$_4$•Et$_2$O at -80 °C, then warmed to recorded temperatures. Spectrum recorded at 20 °C is after 48 h at room temperature.
Figure 2.2. $^1$H NMR (500 MHz) spectra of a CD$_2$Cl$_2$ solution of 2, Fe$_2$(pdt)(CO)$_4$(dppv), after protonation with 3 equiv HBF$_4$•Et$_2$O at -80 °C, then warmed to recorded temperatures. Spectrum recorded at 20 °C is after 16 h at room temperature.
Figure 2.3. Protonation of 2, Fe$_2$(pdt)(CO)$_4$(dppv), with 3 equiv HBF$_4$•Et$_2$O at -80 °C, then warmed to recorded temperatures. $^{31}$P{$^1$H} NMR (202 MHz) acquired with -50 °C $^1$H NMR above shows the growth of signals at δ 97 and 72 ([2H-Term]$^+$, for terminal hydride derivative wherein dppv is apical, basal) and a signal at δ 89 ([2H-Term$^*$]$^+$, terminal hydride with hydride on dibasal-dppv side). When sample is warmed to – 30 °C, new signals are observed at δ 92 and 95 ([2H-A]$^+$ - is bridging hydride of apical,basal-dppv). When sample is warmed to room temperature, only one signal is observed at δ 86 ([2H-B]$^+$) - thermodynamic product with dppv as dibasal (triplet in high-field $^1$H NMR). At low temperature, signals at δ 96 and 82 are due to remaining starting material.
Low temperature protonation of the propanedithiolate (2) resulted in an identical isomerization sequence as seen for 1; However, they appear to arise from kinetic terminal hydrides. The formation of a pair of isomeric hydrides were indicated by characteristically\(^4,20,21\) low field signals at \(\delta -4.4\) (s) and \(\delta -3.3\) (t, \(J_{PH} = 70\) Hz) (Figure 2.2). The predominant (~92%) product, [2H-Term]+ is characterized by \(^{31}\)P NMR signals at \(\delta 97\) and 71, consistent with a terminal hydride on the Fe(CO)\(_3\) subunit and an adjacent Fe(CO)(dppv) subunit with an apical-basal dppv (Figure 2.3). The minor (~8%) product, [2H-Term*]+ terminal hydride isomer features a single \(^{31}\)P NMR signal at \(\delta 89\), and is consistent with a terminal-apical hydride on the Fe(CO)(dppv) subunit wherein the dppv is dibasal (Scheme 2.3). Isomerization of the major isomer commences at a convenient rate around ~30 °C, and upon warming the sample to room temperature, a single \(\mu\)-hydride, [2H-A]+, is formed with \(^{31}\)P NMR signals at \(\delta 95\) and 92 and \(^1\)H NMR signal at \(\delta -14.4\) (dd \(J_{PH} = 25\), 5 Hz). The spectrum is consistent with an apical-basal dppv (see Scheme 2.3), identical to the initially observed \(\mu\)-hydride for the edt example. As in the edt case, the sample further isomerized at room temperature to give exclusively [2H-B]+, a dibasal dppv isomer with a hydride signal at \(\delta -14.8\) (t, \(J_{PH} = 21\) Hz). Ezzaher et al. had observed a similar sequence of reactions for the protonation of Fe\(_2\)(pdt)(CO)\(_4\)(chel) (chel = dmpe\(^17\) (1,2-C\(_2\)H\(_4\)(PMe\(_2\))\(_2\) and dppe (1,2-C\(_2\)H\(_4\)(PPh\(_2\))\(_2\)) – these terminal hydrides convert to the apical-basal then the dibasal isomer of the \(\mu\)-hydride. The isomerization pathway for the minor isomer of the terminal hydride could not be studied because of inadequate signal/noise ratio.
Protonation of Fe$_2$(edt)(CO)$_3$(PMe$_3$)(dppv) and Isomerization of the Resulting Hydrides. Protonation of room temperature solutions of Fe$_2$(edt)(CO)$_3$(PMe$_3$)(dppv) (3)\textsuperscript{22} by HBF$_4$Et$_2$O was observed to yield a mixture of three isomeric µ-hydrides, as indicated by three multiplets in the high field region of the $^1$H NMR spectrum.\textsuperscript{15,23} When this protonation was conducted at –90 °C, only a single µ-hydride was obtained, [3H-A]BF$_4$. In this species, the dppv ligand occupies one apical and one basal site as indicated by $^{31}$P NMR signals at δ 21, 91, 96.\textsuperscript{22} The $^1$H NMR spectrum displayed a doublet of triplets at δ -16.9 ($J_{PH1} = 23$, $J_{PH2} \sim 3$ Hz) (Figure 2.4, 2.5). This pattern is consistent with the dppv ligand spanning apical ($J_{PH} \sim$3 Hz) and basal positions ($J_{PH} \sim$23 Hz). Upon warming the sample to 20 °C, [3H-A]$^+$ was found to convert to [3H-B]$^+$. For this isomer, $^1$H and $^{31}$P NMR data indicated that the dppv remains apical-basal, but the PMe$_3$ has shifted to a basal site. Over the course of ~20 h at room temperature, two further isomerizations occur giving rise to [3H-C]$^+$ and [3H-D]$^+$. Patterns for $^{31}$P and high-field $^1$H NMR spectra also allowed us to uniquely identify the stereochemistry of these species. In [3H-C]$^+$, the PMe$_3$ was found to remain basal, but the dppv shifted to a dibasal orientation. In [3H-D]$^+$, the PMe$_3$ returns to an apical position to give [3H-D]$^+$. The equilibrated (room temperature) solution was found to contain 75% [3H-B]$^+$, 20% [3H-C]$^+$, and 5% [3H-D]$^+$ (Figure 2.9, Scheme 2.4).
**Figure 2.4.** $^1$H NMR (500 MHz) spectra of CD$_2$Cl$_2$ solution of 3, Fe$_2$(edt)(CO)$_3$(PMe$_3$)(dppv), after protonation with [H(Et$_2$O)$_2$]BAr$_4$ at -90 °C and recorded at various temperatures. At -75 °C, a doublet is the sole species present in the $^1$H NMR spectrum, J$_{PH}$ ~30 Hz (indicating a single P atom cis to the hydride). Upon warming to -30 °C, two new triplets are observed in the $^1$H NMR, J$_{PH}$ ~25 Hz (indicating two P atoms cis to the hydride). Upon equilibration of [3H]$^+$ at room temperature for several months, the $^1$H NMR displays a triplet and a triplet of doublets, and possibly one more signal with unresolved coupling under the major resonance at δ -16.8. As seen for [4H]$^+$, the unstable rotamer [3H-B$^*$]$^+$ is observed at δ -16.8.
Figure 2.5. $^{31}\text{P}^{1\text{H}}$ NMR spectra of CD$_2$Cl$_2$ solution of 3, Fe$_2$(edt)(CO)$_3$(PMe$_3$)(dppv), after protonation with [H(Et$_2$O)$_2$]BAR$_4^-$ at -90 °C and recorded at various temperatures. At -75 °C, the $^{31}\text{P}^{1\text{H}}$ spectrum shows mostly one isomer, with two signals in the dppv region (apical-basal) and one signal in the PMe$_3$ region. At -30 °C, the same peaks remain, smaller broad peaks are due to slight remainder of starting material. Upon equilibration of [3H]$^+$ at room temperature for several months, the $^{31}\text{P}^{1\text{H}}$ NMR spectrum displays in the dppv region a set of equally intense peaks at δ 92, 93 (apical-basal, major isomer), a single resonance at δ 89 (dibasal, minor isomer), and in the PMe$_3$ region three peaks. The missing dppv signal for the third isomer wherein the dppv is possibly dibasal is not observed.
Scheme 2.4. Isomerization pathway for $[\text{HFe}_2\text{(pdt)}(\text{CO})_3(\text{dppv})(\text{PMe}_3)]^+$, $[\text{4H}]^+$, and $[\text{HFe}_2(\text{edt})(\text{CO})_3(\text{dppv})(\text{PMe}_3)]^+$, $[\text{3H}]^+$, and their equilibrium isomer distribution.

In a preparative-scale study, 3 was protonated with HCl followed by ion exchange with PF$_6^-$ to give the thermally and air-stable salt. From an equilibrated solution of [3H]PF$_6$, we grew crystals of exclusively [3H-B]PF$_6$. Crystallographic analysis confirmed the assignment of the dppv orientation as apical-basal and trans-basal PMe$_3$ (Figure 2.6). Fresh solutions of these crystals contained only [3H-B]$^+$. Upon allowing its solutions to stand, this species isomerized to the equilibrium mixture of [3H-B]$^+$, [3H-C]$^+$, and [3H-D]$^+$, as discussed above.
Figure 2.6. Structure of the cation in [Fe₂(edt)(m-H)(CO)₃(dppv)(PMe₃)]PF₆ ([3H-B]⁺) with thermal ellipsoids drawn at the 50% probability level, phenyl hydrogen atoms omitted for clarity. Key distances (Å): Fe-Fe, 2.5744 (5); Fe(1)-S(1), 2.2581 (7); Fe(1)-S(2), 2.2794 (6); Fe(2)-S(1), 2.2668 (7); Fe(2)-S(2), 2.2760 (6); Fe(1)-P(1), 2.2137 (6); Fe(1)-P(2), 2.2282 (6); Fe(1)-H(1), 1.69 (2); Fe(2)-H(1), 1.64 (2).
Table 2.1. Crystallographic data for [3H-B]^+, [HFe_2(edt)(CO)_3(dppv)(PMe_3)]PF_6, including selected lengths (Å) and angles (°).

<table>
<thead>
<tr>
<th>Bond</th>
<th>Length (Å)</th>
<th>Bond</th>
<th>Angle (°)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe(1) – Fe(2)</td>
<td>2.5744 (5)</td>
<td>Fe(1) – S(1) – Fe(2)</td>
<td>69.355 (17)</td>
</tr>
<tr>
<td>Fe(1) – S(1)</td>
<td>2.2581 (7)</td>
<td>Fe(1) – Fe(2) – S(1)</td>
<td>55.164 (17)</td>
</tr>
<tr>
<td>Fe(1) – S(2)</td>
<td>2.2794 (6)</td>
<td>Fe(2) – Fe(1) – P(1)</td>
<td>148.911 (18)</td>
</tr>
<tr>
<td>Fe(2) – S(1)</td>
<td>2.2668 (7)</td>
<td>Fe(2) – Fe(1) – P(2)</td>
<td>113.20 (2)</td>
</tr>
<tr>
<td>Fe(2) – S(2)</td>
<td>2.2760 (6)</td>
<td>Fe(2) – Fe(1) – C(1)</td>
<td>111.37 (6)</td>
</tr>
<tr>
<td>Fe(1) – P(1)</td>
<td>2.2137 (6)</td>
<td>P(1) – Fe(1) – P(2)</td>
<td>86.80 (2)</td>
</tr>
<tr>
<td>Fe(1) – P(2)</td>
<td>2.2282 (6)</td>
<td>P(1) – Fe(1) – C(1)</td>
<td>90.78 (6)</td>
</tr>
<tr>
<td>Fe(1) – C(1)</td>
<td>1.765 (2)</td>
<td>P(2) – Fe(1) – C(1)</td>
<td>91.10 (6)</td>
</tr>
<tr>
<td>Fe(2) – P(3)</td>
<td>2.2549 (7)</td>
<td>Fe(1) – Fe(2) – C(2)</td>
<td>112.44 (7)</td>
</tr>
<tr>
<td>Fe(2) – C(2)</td>
<td>1.790 (2)</td>
<td>Fe(1) – Fe(2) – C(3)</td>
<td>141.78 (7)</td>
</tr>
<tr>
<td>Fe(2) – C(3)</td>
<td>1.771 (2)</td>
<td>Fe(1) – Fe(2) – P(3)</td>
<td>109.54 (2)</td>
</tr>
<tr>
<td>Fe(1) – H(1)</td>
<td>1.69 (2)</td>
<td>C(3) – Fe(2) – C(2)</td>
<td>95.45 (10)</td>
</tr>
<tr>
<td>Fe(2) – H(1)</td>
<td>1.64 (2)</td>
<td>C(3) – Fe(2) – P(3)</td>
<td>94.15 (7)</td>
</tr>
<tr>
<td>C(1) – O(1)</td>
<td>1.142 (2)</td>
<td>C(2) – Fe(2) – P(3)</td>
<td>92.44 (7)</td>
</tr>
<tr>
<td>C(2) – O(2)</td>
<td>1.140 (3)</td>
<td>Fe(1) – H(1) – Fe(2)</td>
<td>99.99</td>
</tr>
<tr>
<td>C(3) – O(3)</td>
<td>1.148 (3)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Protonation of Fe$_2$(pdt)(CO)$_3$(PMe$_3$)(dppv) and Isomerization of the Resulting Hydrides. As in the protonation of 2, low temperature protonation of Fe$_2$(pdt)(CO)$_3$(PMe$_3$)(dppv) $^{22}$ (4) with HBF$_4$•Et$_2$O allowed the observation of kinetic terminal hydrides. At -90 °C, the $^1$H NMR spectrum displayed signals at both δ -3.2 and -4.0, a chemical shift range that is again diagnostic of terminal hydrides (Figure 2.7).$^{4,20}$ Both signals were doublets ($J_{PH} \sim 70$ Hz), consistent with protonation at the Fe(CO)$_2$(PMe$_3$) center, and cis orientation of the terminal-apical hydride and PMe$_3$. The presence of two doublets is attributed to the relative orientation of the propanedithiolate strap, which is known to refold slowly at low temperatures (Scheme 2.5).$^{12}$

![Scheme 2.5. Possible terminal hydride isomers observed during the low-temperature protonation of 4.](image-url)
Figure 2.7. $^1$H NMR (500 MHz) spectrum of a CD$_2$Cl$_2$ solution of 4, Fe$_2$(pdt)(CO)$_3$(PMe$_3$)(dppv), after protonation with HBF$_4$•Et$_2$O at -90 °C. Signals at δ -3, -4 ppm are assigned to isomers of the terminal hydride and appear as doublets with $J_{PH} \sim 75$ Hz. The intense signal at δ -14.3 ppm with $J_{PH} \sim 30$ Hz is assigned to [4H-A]$^+$. For spectra at higher temperature, see Figure S5.
Alternatively, the data do not exclude the possibility that the two doublets indicate the presence of two stereoisomers arising from either cis or trans basal phosphine ligands (as in [3H-B]⁺ and [3H-B*]⁺, see below). ¹H NMR spectra indicate that [4H-Term]⁺ isomerizes quickly at -90 °C (it isomerizes faster than the tetracarbonyl [2H-Term]⁺). The initial μ-hydride, [4H-A]⁺, is stereochemically analogous to the edt compound [3H-A]⁺, having apical-basal dppv and apical PMe₃. ¹H NMR spectra indicated that [4H-A]⁺ isomerizes at -30 °C to afford approximately equal amounts of two isomeric μ-hydrides, labeled [4H-B]⁺ and [4H-B*]⁺ (Figure 2.8). ¹H NMR spectra indicate that both isomers feature two basal and one apical phosphine and that both also feature apical-basal dppv and basal PMe₃ ligands (the isomer with dibasal-dppv and apical PMe₃ can be eliminated since this isomer corresponds to [4H-D]⁺, to be described, see Scheme 2.4). The presence of the two species ([4H-B]⁺ and [4H-B*]⁺) is attributed to the two directions for rotation of the HFe(CO)₉(PMe₃) center, placing the PMe₃ either trans-basal ([4H-B]⁺) or cis-basal ([4H-B*]⁺). Upon standing at -30 °C, the [4H-B]⁺/[4H-B*]⁺ mixture was found to convert entirely to [4H-B]⁺. We also observed low concentrations of an analogous pair of isomers for the ethanedithiolate (i.e., [3H-B]⁺ and [3H-B*]⁺). At room temperature, [4H-B]⁺ was found to isomerize to [4H-C]⁺ and [4H-D]⁺. The equilibrium ratio for [4H]BF₄ was 71% [4H-C]⁺, 18% [4H-B]⁺ (0% [4H-B*]⁺), 11% [4H-D]⁺ (Figure 2.9). The propanedithiolate favors the isomer wherein all phosphines are basal ([4H-C]⁺), whereas for the edt compound, the major isomer features basal PMe₃ but apical-basal dppv ([3H-B]⁺).
Figure 2.8. $^1$H NMR (500 MHz) spectra of CD$_2$Cl$_2$ solution of 4 Fe$_2$(pdt)(CO)$_3$(PMe$_3$)(dppv) after warming to various temperatures. Labeling is as follows (for peaks observed at -30 °C): $\delta$ -14.4 (dt) is $[4\text{H-A}]^+$, $\delta$ -14.7 (td) is $[4\text{H-B}^*]^+$, $\delta$ -15.3 (td) is $[4\text{H-B}]^+$ (for peaks observed at 23 °C): $\delta$ -14.3 (q) is $[4\text{H-C}]^+$, $\delta$ -14.8 (td) is $[4\text{H-D}]^+$. Peak at $\delta$ -16.4 is unknown.
The kinetics of the transformation of [4H-B]+ to [4H-C]+ and [4H-D]+ were measured by \textsuperscript{1}H NMR spectroscopy. The linear dependence of ln([4H-B]) vs time (Figure 2.10, Figure 2.11) indicated that the isomerization is first-order. The rate constant of isomerization of [4H-B]+ to [4H-C]+ was quantified as $1.1 \times 10^{-5}$ s\(^{-1}\) at 23 °C ($t_{1/2} \approx 17$ h). The rate was unaffected by 1 atm. H\(_2\). When the isomerization of [4H-B]+ was conducted under an atmosphere of D\(_2\), the hydride ligand did not exchange.

**Figure 2.9.** Difference in energy for isomers for [3H]+ and [4H]+, [HFe\(_2\)(xdt)(CO)\(_3\)(PMe\(_3\))(dppv)]BF\(_4\), as calculated experimentally.
Figure 2.10. Plot of the isomerization of \([4H-B]^+\) to \([4H-C]^+\), as monitored by \(^1\)H NMR spectroscopy.
Figure 2.11. Plot monitoring the concentration of $[4\text{H-B}]^\circ$, $[4\text{H-C}]^\circ$ and $[4\text{H-D}]^\circ$ as a function of time. Data was collected using $^1\text{H}$ NMR spectroscopy.
The equilibration of $\textbf{[4H]}^+$ was examined by high temperature NMR measurements. Between 80 and 120 °C, the signals for two isomers ($\textbf{[4H-C]}^+$ and $\textbf{[4H-D]}^+$) were observed to broaden, but coalescence was not observed (Figure 2.12). Consistent with the proposed dynamics, the two isomers that undergo exchange are related by a single turnstile rotation. After three days in dimethylformamide, samples of $\textbf{[4H]}^+$ were observed to degrade to a mixture of $\textbf{[Fe}_2(pdt)(\mu\text{-H})(\text{CO})_2(dppv)_2]^+$ and unidentified poly-phosphine hydrido complexes.
Protonation of Fe$_2$(xdt)(CO)$_2$(dppv)$_2$ and Isomerization of the Resulting Hydrides (where xdt = edt, pdt). The complexes Fe$_2$(S$_2$C$_n$H$_{2n}$)(CO)$_2$(dppv)$_2$ (n = 2, 5; n = 3, 6) are protonated by HBF$_4$•Et$_2$O at room temperature and isolated as two bridging hydride isomers [Fe$_2$(S$_2$C$_n$H$_{2n}$)(µ-H)(CO)$_2$(dppv)$_2$]BF$_4$ ([5H-A,B]BF$_4$ and [6H-A,B]BF$_4$). The $^1$H NMR spectrum of [6H]$^+$ displays a triplet-of-triplets ($\delta - 14.6, J_{PH} = 24, 6$ Hz) and a doublet-of-doublet-of-doublet-of-doublets ($\delta - 15.6, J_{PH} = 24, 19, 19, 10$ Hz) (Figure 2.12).
2.13). One isomer of the propanedithiolato derivative ([6H-B]BF₄) was characterized crystallographically. The structure of [6H-B]BF₄ identifies the bridging hydride approximately equidistant between the two iron centers (Figure 2.14). The Fe-Fe bond distance has slightly increased from that of the starting complex 6 (Table 2.2).

![Figure 2.13](image)

**Figure 2.13.** High-field $^1$H NMR (500 MHz, CD₂Cl₂) spectrum of the two isomers of [Fe₂(S₂C₆H₆)(μ-H)(CO)₂(dppv)₂]BF₄, [6H-A,B]BF₄
Figure 2.14. Structure of [6H-B]BF₄. Thermal ellipsoids are drawn at 50%, phenyl groups are drawn as lines with hydrogen atoms omitted and counter-anion has been omitted for clarity.

Table 2.2. Selected bond distances for the molecular structure of [Fe₂(pdt)(µ-H)(CO)₂(dppv)₂]BF₄, [6H-B]BF₄.

<table>
<thead>
<tr>
<th>Table of Selected Bond Distances</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe(1) - Fe(2)</td>
</tr>
<tr>
<td>Fe(1) - H(1)</td>
</tr>
<tr>
<td>Fe(2) - H(1)</td>
</tr>
<tr>
<td>Fe(1) - S(1)</td>
</tr>
<tr>
<td>Fe(2) - S(1)</td>
</tr>
<tr>
<td>Fe(1) - C(5)</td>
</tr>
<tr>
<td>Fe(2) - C(4)</td>
</tr>
<tr>
<td>Fe(1) - P(1)</td>
</tr>
<tr>
<td>Fe(1) - P(3)</td>
</tr>
<tr>
<td>Fe(2) - P(2)</td>
</tr>
<tr>
<td>Fe(2) - P(4)</td>
</tr>
</tbody>
</table>
When 5 and 6 were protonated at low temperatures, we selected [H(Et₂O)₂]BAr₄⁻ for to ensure precise stoichiometry; However, replicate experiments with HBF₄•Et₂O provided qualitatively similar results. Upon protonation, we observed high-field ¹H NMR signals characteristic of terminal hydrides. The hydride signals are both triplets ([5-tH]ₕ: δ -6.1, t, JₚH = 74 Hz; [6-tH]BF₄, δ -3.5, t, JₚH = 78 Hz) indicating coupling to chemically equivalent phosphorus atoms (Figure 2.15). The corresponding ³¹P{¹H} NMR spectrum displays four signals (δ 99, 91, 85, 68) two of which (δ 91, 85) show additional coupling to the hydride resonance (Figure 2.16, Figure 2.17). The spectroscopy is consistent with one apical-basal diphosphine ligand and the other dibasal (Scheme 2.5). The FT-IR of [6-tH]BF₄ displays one strong band at 1964 cm⁻¹, a weaker bridging CO vibration at 1905 cm⁻¹, and a weak shoulder at 1888 cm⁻¹ (Figure 2.18). The weak signal at 1888 cm⁻¹ is not present for the deuteride [6-tD]⁺ (generated from HBF₄•Et₂O and D₂O) and thus is assigned to the Fe-H vibration.

Figure 2.15. High-field ¹H NMR Spectra of [5-tH]BAr₄⁻ (bottom) and [6-tH]BAr₄⁻ (top).
Figure 2.16. $^{31}P\{^1H\}$ NMR (500 MHz, CD$_2$Cl$_2$, -35°C) spectrum of [(t-H)Fe$_2$(pdt)(CO)$_2$(dppv)$_2$]BAr$_4$, [6-tH]BAr$_4$.

Figure 2.17. $^{31}P\{^1H\}$ NMR (600 MHz, CD$_2$Cl$_2$) spectrum of [(t-H)Fe$_2$(edt)(CO)$_2$(dppv)$_2$]BAr$_4$, recorded at -35°C. Peaks corresponding to [5H-Term]$^+$ are at δ 99, 91, 84, 74 ppm. Remaining signals are due to [5H-A]$^+$ (97, 87 ppm) and 5 (92, 93 ppm).
Figure 2.18. FT-IR spectrum (CH₂Cl₂, ~25 °C) of Top: [6-tH]BF₄ displaying a terminal (1964 cm⁻¹) and bridging (1901 cm⁻¹) $\nu_{CO}$ vibration, and possibly $\nu$FeH (1888 cm⁻¹). Bottom: FT-IR spectrum of [6-tD]BF₄ generated by addition of D₂O to HBF₄•Et₂O, showing no vibration of FeH.
Unlike the other terminal hydrides reported in this chapter which rapidly isomerizes to bridging hydrides at less than –60 °C, the terminal hydrides \([5-tH]^+\) and \([6-tH]^+\) are more stable. The ethanedithiolate terminal hydride, \([5-tH]^+\), isomerized in minutes at -20 °C to \([5-\mu H]^+\). However, the propanedithiolate (\([6-tH]^+\)) derivative proved even more stable and isomerized at \(+20^\circ \text{C with a } t_{1/2} \sim 10 \text{ min.}\) We attribute the slower isomerization of the propanedithiolate (pdt) derivative to the steric clash between the dppv and the middle methylene group of pdt inhibiting the rotation of the FeH(CO)(dppv) site (Scheme 2.6). A similar steric argument has recently been made to describe the lifetime of the \([6(CO)]^+\), a mixed valence species that undergoes a similar isomerization process.\(^{24}\)

**Scheme 2.6.** Depiction of the steric clash between propanedithiolate and diphosphine that hinders the isomerization of the terminal hydride.

The isomerization of the terminal \([6-tH]^+\) to the bridging \([6-\mu H]^+\) hydride proceeds via a first-order pathway \((k = 1 \times 10^{-3} \text{ s}^{-1}, 25 ^\circ \text{C})\) and was unaffected by solvent polarity or the presence of CO (Figure 2.19). These findings are indicative of an intramolecular process, consistent with isomerization kinetics previously reported for \([\text{Fe}_2(\text{pdt})(t-H)(\text{CO})_2(\text{PMe}_3)_4]\text{BF}_4\). An Eyring plot indicates \(\Delta H^\ddagger = 80 \text{ kJ/mol and } \Delta S^\ddagger = -23 \text{ J/mol K,}\)
also consistent with an intramolecular process that proceeds without ligand dissociation (Figure 2.20).
Further investigating the mechanism of isomerization, both a deprotonation-reprotonation mechanism and a intramolecular turnstile rotation mechanism would display first-order kinetics. To address this issue we monitored the isomerization under conditions that might change this isomerization rate. A reversible deprotonation would depend on the concentration of acid. However, in the presence of excess [H(Et₂O)₂]BAR₄, the isomerization rate of [6-tH]⁺ was unaffected. Additionally, because such a mechanism would transiently produce free acid, we would expect deuteride incorporation if the isomerization is conducted in the presence of a deuterium source. However, no H/D exchange was observed when a solution of [6-tH]⁺ was allowed to isomerize in the presence of D₂O at 10 °C. Confirming the intramolecular turnstile rotation mechanism, the rate of isomerization was indistinguishable for CD₃CN and CD₂Cl₂ supporting the absence of a solvent role in the isomerization process.

The isomerization process converting the terminal hydride to the bridging hydride
displays a series of rotations consistent with the other hydrides, [1-4-tH]+. The kinetic bridging hydride, [5,6H-A]+ contains two dppv ligands that are both apical-basal (Scheme 2.7). Upon warming the solution a second bridging hydride is observed, [5,6H-B]+. Both isomers are the same as that obtained from room temperature protonation, and [6H-B]+ has been characterized crystallographically (see above).

Scheme 2.7. Protonation scheme for Fe₂(xdt)(CO)₂(dppv)₂, here depicted as pdt, although the results for edt are consistent.
Discussion

The first suggestion of isomeric hydrides in the diiron dithiolates came from the characterization of \([\text{Fe}_2(\text{edt})(\mu-\text{H})(\text{CO})_4(\text{CN})_2]\).\(^{15}\) The \(^1\text{H}\) NMR spectrum showed two hydride signals, which are assigned to isomers of \([((\text{CO})_2(\text{CN})\text{Fe}(\mu-\text{edt})(\mu-\text{H})\text{Fe}(\text{CO})_2(\text{CN}))]^{25}\) Related derivatives of the type \([\text{Fe}_2(\text{edt})(\mu-\text{H})(\text{CO})_3\text{L}(\text{CN})_2]\) also exist as isomers (\(\text{L} = \text{PPh}_3, \text{CN}^{-}\)).\(^{25}\) Additionally, Schollhammer and Hogarth \textit{et al.} have demonstrated isomerism in \([\text{HFe}_2(\text{pdt})(\text{CO})_4(\text{chel})]^+\) (\text{chel} = \text{dmpe}^{17}, \text{dppe}^{20}, \text{phen}^{21}).\(^{26}\)

**Terminal Protonation.** Diiron complexes with terminal hydrides are rarely observed.\(^{27,28}\) Poilblanc proposed that protonation of the diiron dithiolates occurs directly at the Fe-Fe bond.\(^{29}\) However, when these protonations are monitored at low temperatures by NMR spectroscopy, terminal hydrides are often observed, especially for complexes containing chelating ligands.\(^{17,20,21,30}\) The chelating phosphine ligands elevate the barrier for the conversion of the terminal hydrides to the bridging hydrides. Even though terminal hydrides are not observed consistently, results suggest that terminal hydrides are the sole kinetic product of protonation. In cases where the terminal hydrides have not been observed spectroscopically - \(1, 3,\) and \(\text{Fe}_2(\text{xdt})(\text{CO})_4(\text{PMe}_3)_2\) (\text{xdt} = \text{adt, edt, pdt}) - the initially observed \(\mu\)-hydrides are explicable as arising via intramolecular rearrangement from the analogous terminal hydrides.

Potential exceptions to the universality of terminal protonation come from studies on the protonation of diiron dithiolates that are constrained by special chelating ligands. The complex \(\text{Fe}_2(\text{pdt})(\text{Cy}_2\text{PCH}_2\text{PCy}_2)(\text{CO})_4,\) where the diphosphine links the two Fe
centers, protonates rapidly to give the $\mu$-hydride.\textsuperscript{31} The two isomers of $\text{Fe}_2(\text{SCH}_2\text{CH}_2\text{PPh}_2)_2(\text{CO})_4$, protonate with retention of geometry (Eq. 2.1).\textsuperscript{16} These protonations are slow at room temperature; However, possibly reflecting the high barrier to protonation of the Fe-Fe bond. A recent report suggests that protonation at a Fe-Fe bond can be accelerated by ligands that can serve as proton relays.\textsuperscript{32} In such cases however, it is difficult to exclude relay to a single Fe center followed by rapid rearrangement to the $\mu$-hydride isomer. In contrast, the protonation of the complexes described in this paper occur within the time of mixing at room temperature. Although fast at room temperature, the protonation of $\text{Fe}_2(\text{edt})(\text{CO})_4(\text{dppv})$ is noticeably slower than the protonation of $\text{Fe}_2(\text{pdt})(\text{CO})_4(\text{dppv})$ at $-80\,^\circ\text{C}$.

\begin{equation}
\text{Fe-S-S-Fe} + \text{CF}_3\text{SO}_3\text{H} \rightarrow \text{Fe} \equiv \text{Fe} + + \text{C} = \text{O}
\end{equation}

**Mechanism of Terminal to $\mu$-Hydride Isomerization.** In those cases where they can be observed, terminal hydrides do not readily deprotonate. For example, separate NMR signals for hydrides and their conjugate bases can be resolved readily, consistent with a relatively large barrier ($> 12\,\text{kcal/mol}$) to intermolecular proton transfer. Indicative of their significant kinetic robustness is the high barrier for deprotonation. Thus, the terminal hydride $[(t\text{-H})\text{Fe}_2(\text{pdt})(\text{CO})_4(\text{dppv})]^+$ resists deprotonation by large excess of $\text{NEt}_3$, which is far more basic than the diiron center ($\Delta pK_a \sim 6$).\textsuperscript{33-36} Thus, the isomerization of the terminal hydrides to the $\mu$-hydrides does not occur via a
deprotonation-reprotonation pathway. Consistent with this view, conversion of $[\text{4H-Term}]^+$ to $[\text{4H-A}]^+$ occurs without H-D exchange with D$_2$O.

The terminal hydrides are however labile with respect to conversion to the $\mu$-hydrides. The rate of this isomerization is strongly affected by the dithiolate, the pdt derivatives always being more stable than the analogous edt compounds.$^{37}$ Steric effects alone do not explain the stability trends for the terminal hydrides as indicated by the finding that $[\text{HFe}_2(\text{pdt})(\text{CO})_3(\text{PMe}_3)(\text{dppv})]\text{BF}_4$ is significantly less stable than $[\text{HFe}_2(\text{pdt})(\text{CO})_4(\text{dppv})]\text{BF}_4$ ($[\text{2H-Term}]\text{BF}_4$). The terminal hydride $[\text{4H-Term}]^+$ is observed only fleetingly at $-90$ °C, whereas $[\text{2H-Term}]\text{BF}_4$ is stable until ca. $-40$ °C. We suggest the first step in the isomerization is a rate-determining 120° “turnstile” rotation that shifts the terminal apical hydride to a terminal basal site. DFT calculations predict that the basal hydrides are significantly less stable than the apical hydrides, consistent with the fact that basal hydrides species have so far never been observed (Scheme 2.8).$^{26}$ The higher barrier for the isomerization of $[\text{2H-Term}]^+$, in comparison to $[\text{4H-Term}]^+$, is attributed to the instability of $[\text{2H-Basal}]^+$ where CO is trans to $\mu$-CO. The second step in the isomerization involves the interchange of the basal hydride ligand with the $\mu$-CO ligand. Precedent for the conversion of the terminal to the $\mu$-hydride comes from the pairwise exchange of terminal and $\mu$-hydrides in species such as $\text{HOS}_3(\mu$-$\text{H})(\text{CO})_{10}\text{L}$. $^{38}$
Interconversions within the \(\mu\)-Hydride Manifold. Two or more isomers of the \(\mu\)-hydrides are observed in all cases reported here. In general, these isomers exist in dynamic equilibrium, and the overall isomer distribution is dictated by the energetic advantage associated with CO ligands \(\textit{trans}\) to the hydride.\(^{39}\) In addition to this \(\textit{trans}\)-influence of dinuclear \(\mu\)-hydrides, steric interactions between the organic ligands influence the stability of the various isomers. The \(\mu\)-hydrides are proposed to interconvert via 120° turnstile rotations of a single FeL\(_3\) subunit. The isomerization occurs at \textit{alternating} Fe sites. Particularly telling mechanistically is the observation of pairs of \(\mu\)-H isomers in the case of the pdt-(PMe\(_3\))(dppv) system ([4H-B]\(^+\)/[4H-B\(^+\)]\(^+\)). These rotamers contain basal PMe\(_3\) ligands that are mutually \(\textit{cis}\) and \(\textit{trans}\) with respect to the apical-basal dppv ligand. These apparent rotamers arise from the clockwise and anticlockwise rotations of the Fe(CO)\(_2\)(PMe\(_3\)) center. A high temperature \(^1\)H NMR spectrum of [4H]\(^+\) showed that a pair of isomers related by a single 120° rotation of the
Fe(CO)$_2$(PMe$_3$) subunit ([4H-C]$^+$ and [4 H-D]$^+$) approached coalescence. An unanswered question is whether isomerizations within the μ-hydride manifold occur via scission of a Fe-H bond. The scission of individual Fe-μ-H bonds has been proposed to occur upon photolysis of [Fe$_2$(pdt)(μ-H)(CO)$_4$(PMe$_3$)$_2$]$^+$.\textsuperscript{40,41}

**Concluding Remarks.** The protonation of the substituted diiron dithiolato carbonyls provides new insights into protonation of polynuclear complexes, a broadly important theme.\textsuperscript{42} The main pathway entails protonation at a single metal followed by isomerization to series of μ-hydrides. As described in chapters 3 and 4, the terminal hydrides are mechanistically related to the [FeFe]-hydrogenases whereas the μ-hydrides are more closely related to the [NiFe]-hydrogenases.
Experimental

Manipulations were conducted using standard Schlenk techniques. Solvents were filtered through activated alumina and subsequently degassed. $^1$H and $^{31}$P NMR spectra were acquired on a Unity Varian 500 or a Unity Varian 600 spectrometer. IR spectra were collected on a Mattson Infinity Gold FTIR spectrometer. Fe$_2$(edt)(CO)$_4$(dppv)$_2$ [22] Fe$_2$(edt)(CO)$_4$(dppv)$_2$ [22] Fe$_2$(edt)(CO)$_4$(PMe$_3$)(dppv)$_2$ [22] Fe$_2$(pdt)(CO)$_3$(PF$_6$)(ppv)$_2$ [22] Fe$_2$(pdt)(CO)$_3$(PMe$_3$)(dppv)$_2$ [22] Fe$_2$(dppv)$_2$(CO)$_2$ [22,30] [H(Et$_2$O)$_2$]BAR$_4^{43}$ were prepared according to literature procedures. HBF$_4$•Et$_2$O and dppv (cis-1,2-bis(diphenylphosphino)ethylene) were purchased from Aldrich and used as received.

$\text{[Fe}_2\text{(edt)(µ-H)(CO)}_4\text{(dppv)]PF}_6$ ([1H]PF$_6$). A solution of 0.110 g (0.155 mmol) of Fe$_2$(edt)(CO)$_4$(dppv) in 10 mL of CH$_2$Cl$_2$ was treated with 0.20 mL (0.40 mmol) of a 2.0 M solution HCl in Et$_2$O, the reaction flask was immediately stoppered. After the reaction solution stirred for 3 h, solvent was removed in vacuum. The residue, a red powder, was then re-dissolved in 5 mL of MeOH. The red solution was then treated with 0.025 g of NH$_4$PF$_6$ (0.160 mmol) in 5 mL of MeOH, precipitating the product. The red solid was collected via filter cannula and washed with 2 x 10 mL Et$_2$O. Anal. Calcd for C$_{32}$H$_{27}$Fe$_2$O$_4$F$_6$P$_3$S$_2$ (found): C, 44.78 (44.70); H, 3.17 (3.02). $^1$H NMR (CD$_2$Cl$_2$): δ 7.2 – 8.3 (m, 22H, P(C$_6$H$_5$)$_2$; C$_2$H$_2$(PPh$_2$)$_2$); 0.7 - 2.9 (m, 4H, S$_2$C$_2$H$_4$); δ -15 (t, 1H, Fe-H). $^{31}$P NMR (CD$_2$Cl$_2$): δ 85. FT-IR (CH$_2$Cl$_2$): 2098 (s), 2052 (m), 2038 (m), 1980 (m) cm$^{-1}$.

$\text{[Fe}_2\text{(pdt)(µ-H)(CO)}_4\text{(dppv)]PF}_6$ ([2H]PF$_6$). A solution of 0.066 g (0.091 mmol) of Fe$_2$(pdt)(CO)$_4$(dppv) in 10 mL of CH$_2$Cl$_2$ was treated with 0.2 mL (0.40 mmol) of a 2.0 M solution HCl in Et$_2$O, the reaction flask was then stoppered. After the reaction solution stirred for 3 h, solvent was removed in vacuum. The residue, a red powder, was then
re-dissolved in 5 mL of MeOH. The red solution was then treated with 0.015 g of NH₄PF₆ (0.092 mmol) in 5 mL of MeOH, precipitating the product. The red solid was collected via filter cannula and washed with 2 x 10 mL Et₂O. ¹H NMR (CD₂Cl₂): δ 7.2 – 8.5 (m, 22H, P(C₆H₅)₂; C₂H₂(PPh₂)₂); 2.3–3.6 (m, 6H, S₂C₃H₆); δ -14.7 (t, 1H, Fe-H, Jₓᵧ = 23 Hz). ³¹P NMR (CD₂Cl₂): δ 85.6. FT-IR (CH₂Cl₂): 2098 (s), 2052 (m), 2038 (m), 1980 (m) cm⁻¹.

[Fe₂(edt)(μ-H)(CO)₃(PMe₃)(dppv)]PF₆ ([3H]PF₆). A solution of 0.125 g (0.164 mmol) of 3 in 25 mL of MeCN was treated with 8.2 mL (0.820 mmol) of a 0.10 M solution of 12M HCl in MeCN. After the reaction solution stirred for 3 h, solvent was removed in vacuum. The residue, an orange powder, was rinsed with 30 mL of Et₂O and then re-dissolved in 5 mL of MeOH. This extract was treated with a solution of 0.026 g of NH₄PF₆ (0.164 mmol) in 5 mL of MeOH, and the product was precipitated by the addition of 30 mL of H₂O. The orange solid was collected and rinsed with 15 mL each of H₂O and Et₂O. Crystals were grown via slow diffusion of hexanes into a CH₂Cl₂ solution. Anal. Calcd for C₃₄H₃₆Fe₂O₃F₆P₄S₂ (found): C, 45.06 (44.89); H, 4.00 (3.93). ¹H NMR (CD₃CN): δ 7.2 – 8.5 (m, 22H, P(C₆H₅)₂; C₂H₂(PPh₂)₂); 0.4–2.8 (m, 4H, S₂C₂H₄); 1.6 (d, 9H, PMe₃); δ -17 (td, 1H, Fe-H). ³¹P NMR (CD₃CN): δ 92.3, 90.7 (d, t, dppv); δ 22.8 (d, PMe₃). FT-IR (MeCN): 2031, 1978 cm⁻¹.

[Fe₆(pdtd)(μ-H)(CO)₃(PMe₃)(dppv)]PF₆ ([4H]PF₆). The sample was prepared following the procedure for [3H]PF₆. Anal. Calcd for C₃₅H₃₈Fe₂O₃F₆P₄S₂ (found): C, 45.67 (46.20); H, 4.16 (4.22). ¹H NMR (CD₃CN): δ 7.2 – 8.5 (m, 22H, P(C₆H₅)₂; C₂H₂(PPh₂)₂); 0.4–2.8 (m, 6H, S₂C₃H₆); 1.6 (d, 9H, PMe₃); δ -15 (td, 1H, Fe-H). ³¹P NMR (CD₃CN): δ 91.6, 90.6 (t, d, dppv); δ 21.6 (d, PMe₃). FT-IR (MeCN): 2028, 1973
Procedure for Protonation of 1, 2, 3, 4 at Low Temperature. Into a J-Young NMR tube dried CD$_2$Cl$_2$ was distilled onto 5 mg diiron dithiolate. To the frozen solution ~5 mL HBF$_4$•Et$_2$O (6.91 M) (0.035 mmol, 1 to 10 equiv.) was added was freeze-pump-thawed to ensure a vacuum atmosphere. The frozen J-Young tube was then placed directly into a dry ice/Acetone bath (-78 °C), thawed slowly, and characterized by low temperature NMR studies with a pre-cooled NMR spectrometer probe.

Isomerization of [HFe$_2$(pdt)(CO)$_3$(PMe$_3$)(dppv)]PF$_6$, [4H]PF$_6$. In a J-Young NMR tube dried CD$_2$Cl$_2$ was distilled onto a solid mixture of ~5 mg (0.005 mmol) of [HFe$_2$(pdt)(CO)$_3$(PMe$_3$)(dppv)]PF$_6$ ([4H]PF$_6$) and ~1 mg (0.006 mmol) hexamethylbenzene (as internal standard). The frozen J-Young tube was thawed carefully and immediately inserted into the NMR spectrometer. The signals of interest (the PCH$_3$ signals and the hydride signals) were integrated relative to the internal standard over the course of several days, weeks, and months.

H/D Exchange from D$_2$ + [HFe$_2$(pdt)(CO)$_3$(PMe$_3$)(dppv)]PF$_6$. In a J. Young tube and following the conditions for the isomerization of [4H]PF$_6$, D$_2$ (~ 1.0 atmosphere) was added to the freshly prepared and freshly frozen solution of [4H]PF$_6$ in what solvent. The sample was then thawed carefully, shaken (to ensure complete and immediate saturation of solution), and immediately inserted into the NMR spectrometer for kinetic characterization. Experiments were conducted in both ambient light and dark, and no measurable differences were observed.

H/D Exchange from D$_2$O + [HFe$_2$(pdt)(CO)$_2$(dppv)$_2$]BAr$_4^-$F. In a J-Young NMR tube dried CD$_2$Cl$_2$ was distilled onto a solid mixture of 10 mg (0.009 mmol)
Fe$_2$(pdt)(CO)$_2$(dppv)$_2$ and $\sim$10 mg (0.010 mmol) [H(Et$_2$O)$_2$]BAr$_4^\text{F}$. The frozen J-Young tube was then thawed carefully at – 40 °C (CH$_3$CN: dry ice) and then after complete thawing (~1 min) was frozen in liquid N$_2$ and 1 drop D$_2$O was added. The sample was then evacuated on the high-vac line and carefully thawed at -40 °C and immediately inserted into the NMR spectrometer for characterization.

**X-ray Crystallography.** Structure was phased by direct methods. Systematic conditions suggested the ambiguous space group. The space group choice was confirmed by successful convergence of the full-matrix least-squares refinement on F$^2$. The highest peaks in the final difference Fourier map were in the vicinity of atoms F1, C9, and F4; the final map had no other significant features. A final analysis of variance between observed and calculated structure factors showed no dependence on amplitude or resolution.
References:


29. Arabi, M. S.; Mathieu, R.; Poilblanc, R. "Protonation of the Metal-Metal Bond in Fe₂(μ₄-A)(μ₄-A')(CO)₄L₂ Complexes (A = A' = SC₆H₅, A' = P(C₆H₅)₂; L = P(C₆H₅)₃-n(CH₃)n)." *J. Organomet. Chem.* **1979**, 177, 199-209.


41. Zhao, X.; Chiang, C.-Y.; Miller, M. L.; Rampersad, M. V.; Daresbourg, M. Y. "Activation of Alkenes and \(\text{H}_2\) by [Fe]-\(\text{H}_2\)ase Model Complexes" \textit{J. Am. Chem. Soc.} \textbf{2003,} \textit{125,} 518-524.


43. Brookhart, M.; Grant, B.; Volpe, A. F. "\([\{3,5-(\text{CF}_3)\text{C}_6\text{H}_3\}\text{B}]\text{H(OEt}_2\text{)}\text{2}\)\textsuperscript{+}: a convenient reagent for generation and stabilization of cationic, highly electrophilic organometallic complexes" \textit{Organometallics} \textbf{1992,} \textit{11,} 3920-3922.
Chapter 3

Reactivity of Terminal Hydride Complexes and Functional Azadithiolate Proton Relays

Introduction

Though the crystal structure of fully-reduced form of [FeFe]-hydrogenase, $\text{H}_{\text{red}}$, has been known since 1999, the inclusion of a terminal hydride ligand on the distal iron ("Fe\text{d}") in this structure is still debated.\(^1\) However, the observation that the enzyme is inhibited by carbon monoxide at a single iron center conforms to a mechanism whereby proton reduction and hydrogen oxidation occur via substrate binding to the same coordination site on the distal iron.\(^2\) The identity of the central atom of the dithiolate is also debated (Figure 3.1).\(^3\) This atom was originally proposed to be a propanedithiolate.\(^1\) However, after it was found to be within hydrogen-bonding distance to a sulfur atom of a nearby cysteine residue, it was proposed to be nitrogen.\(^4\) Recently, DFT calculations have suggested that this atom is instead oxygen, while $^{14}\text{N}$ HYSCORE and ESEEM spectroscopy has supported the assignment as nitrogen.\(^3,5\)

![Figure 3.1.](image.png)

**Figure 3.1.** Representation of the active site for [FeFe]-Hydrogenase. The identity of the atom in the dithiolate denoted as “X” has been a matter of debate.
Inspired by the combination of the terminal hydride and amine-proton relay combination, DuBois et al. have synthesized nickel diphosphine complexes that incorporate a pendant amine base as a proton relay (Figure 3.2).\textsuperscript{6-9} From this work it has become apparent that for pendant bases to act as functional proton relays, the pK\textsubscript{a} of the ammonium must closely match that of the metal hydride.\textsuperscript{6} When the pK\textsubscript{a}'s are properly tuned, these hydride complexes exist in dynamic equilibrium with their ammonium tautomer and display time-averaged signals by \textsuperscript{1}H NMR spectroscopy. The oxidation of hydrogen by complexes of the type [Ni(diphosphine)\textsubscript{2}]\textsuperscript{2+} is usually slow (t\textsubscript{1/2} \sim hours), but when complemented with pendant amine bases, hydrogen oxidation is much faster and catalytic.

![Figure 3.2. Mononuclear nickel diphosphine complexes incorporating pendant amine groups that serve as proton relays.](image)

Models for [FeFe]-hydrogenase have also incorporated pendant bases with the hopes of facilitating proton transfer (Figure 3.3). The azadithiolate Fe\textsubscript{2}(adtN-Bn)(CO)\textsubscript{4}(PMe\textsubscript{3})\textsubscript{2} (where adt = 2-azapropane-1,3-dithiolate; Bn = benzyl) has been observed to undergo protonation at both the amine and the diiron center, and their pK\textsubscript{a}'s in MeCN solution have been individually determined as 12 and 15, respectively.\textsuperscript{10} The investigation of proton reduction catalysis with amine-containing dithiolates has been
reported to shift electrocatalysis by ~ 400 mV.\textsuperscript{11-14} However, this effect is purely coulombic, as rates are not enhanced. Not surprisingly, the reactivity of the oxadithiolate complexes Fe\textsubscript{2}(odt)(CO)\textsubscript{4}(P\textsubscript{2}) (where P\textsubscript{2} = diphosphine) and Fe\textsubscript{2}(odt)(CO)\textsubscript{6} (where odt = 2-oxapropane-1,3-dithiolate) are virtually indistinguishable from that of the propanedithiolate derivative. For this reason, we conclude that oxadithiolate is not basic enough to serve as a pendant proton relay.\textsuperscript{15} For weakly basic diiron dithiolates, incorporation of the amine-containing azadithiolate offers no advantage to proton relay because the amine base is far more basic than the iron center. In the case of the oxadithiolate, the ether is far less basic than the iron center (Scheme 3.1).\textsuperscript{16}

\begin{figure}
\centering
\includegraphics[width=\textwidth]{diagram.png}
\caption{Flawed diiron dithiolate model complexes incorporating aza- and oxadithiolates.}
\end{figure}
Prior to this work, protonations of diiron(I) dithiolates were observed to occur at the metal-metal bond to give only bridging hydrides. The discovery that diiron(I) dithiolates initially protonate to give terminal, not bridging, hydrides opens a new and potentially significant phase in elucidating the role of the dithiolate cofactor in the catalytic cycle (Chapter 2). We demonstrated that protonations of the electronically symmetrical Fe$_2$(edt)(CO)$_2$(dppv)$_2$ and Fe$_2$(pdt)(CO)$_2$(dppv)$_2$ (1) yield relatively stable ($t_{1/2}$ = minutes at 25 °C), terminal hydride derivatives (dppv = cis-1,2-bis(diphenylphosphino)ethylene; pdt = 1,3-propanedithiolate). These terminal hydrides are sufficiently robust to study in detail. This chapter focuses on the characterization and reactivity of these terminal hydrides. In addition, complexes incorporating the biologically relevant dithiolates (Fe$_2$(adt)(CO)$_2$(dppv)$_2$, 2, and Fe$_2$(odt)(CO)$_2$(dppv)$_2$, 3 have been synthesized to address whether the heteroatom in the dithiolate participates in proton transfer to and from the terminal hydride.

**Scheme 3.1.** Relative basicities of the dithiolate cofactor and the iron center for diiron dithiolate model complexes. Effective proton relays have pK$_a$’s that closely match the metal center.
Results

Reactivity of the Terminal Hydride $[(\tau{H})Fe_2(pdt)(CO)_2(dppv)]^+$, $([1-\rho H])^+$. As the terminal hydride ligand in $[(\tau{H})Fe_2(edt)(CO)_2(PMe_3)_4]^+$ reacts with excess HBF$_4$$\cdot$Et$_2$O in CH$_3$CN to provide H$_2$ and the acetonitrile complex $[Fe_2(edt)(CO)_2(MeCN)(PMe_3)_4]^{2+}$, we sought to probe the basicity of the terminal hydride complex $[1-\rho H]^+$.\(^{21}\) However, CD$_2$Cl$_2$ solutions of $[1-\rho H]^+$ are unreactive toward triflic acid ($pK_a \sim -12$)\(^{22,23}\), showing no tendency to evolve H$_2$. Additionally, the hydride ligand does not exchange with CD$_3$OD over the course of several hours, confirming the low basicity of the hydride ligand.\(^{24,25}\) Given the dissimilarity of $\nu_{CO}$ stretching frequencies of $[(\tau{H})Fe_2(edt)(CO)_2(PMe_3)_4]^+$ (1940 (s), 1874 (w) cm$^{-1}$) and $[1-\rho H]^+$ (1964, 1901 cm$^{-1}$) we were not surprised the strong acid HOTf did not protonate the hydride ligand. However, this differing reactivity might be attributed to the presence of CH$_3$CN, which has demonstrated a profound impact on the protonation chemistry of other terminal hydrides (see the section below on $[(\tau{H})Fe_2(adt,NH)(CO)_2(dppv)]^+$ and CH$_3$CN). In terms of models of the enzyme, $[1-\rho H]BF_4$ is spectroscopically a better active site model for the H$_{red}$ state of the enzyme ($[1-\rho H]BF_4$: 1964, 1901 cm$^{-1}$ vs H$_{red}$ for D.d.: (1965, 1916, and 1894 cm$^{-1}$)).\(^{26}\)

Fe$_2$(adt)(CO)$_2$(dppv)$_2$, 2. The azadithiolate 2 was synthesized from the hexacarbonyl Fe$_2$(adt,NH)(CO)$_6^{27,28}$ in a similar manner to that of Fe$_2(pdt)(CO)_2$(dppv)$_2$. Recrystalization from CH$_2$Cl$_2$ and MeOH provided analytically pure samples of 2. The FT-IR spectrum of 2 is very similar to that of 1 showing two equally intense $\nu_{CO}$ bands (1888, 1868 cm$^{-1}$) (Figure 3.4). The azadithiolate Fe$_2$(adt)(CO)$_2$(dppv)$_2$ (2), like the
structurally related 1 and Fe$_2$(edt)(CO)$_2$(dppv)$_2$, is stereochemically nonrigid displaying both rapid “turnstile rotation” of the Fe(dppv)(CO) subunits and dithiolate “flipping”. At -80 °C, the $^{31}$P{$^1$H} NMR spectrum displays four equally intense signals indicating that the two dppv ligands are chemical inequivalent and the azadithiolate points toward a single iron center (Figure 3.22). In contrast, the related odt, edt, pdt complexes show only a pair of signals at low temperatures. We propose the additional signals arise from the dithiolate flipping process, wherein the azadithiolate has a higher barrier compared to the structurally related pdt, due to the anomeric affect.$^{29,30}$ (Scheme 3.2) In a previous report, the observation of an additional peak in the $^{31}$P NMR spectrum led to the speculation that an additional apical-basal, dibasal dppv isomer existed.$^{30}$ However, in cleaner samples this signal is not present and such asymmetric isomer would likely have four signals.

![FT-IR spectrum of Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$ (2) in CH$_2$Cl$_2$.](image)

**Figure 3.4.** FT-IR spectrum of Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$ (2) in CH$_2$Cl$_2$. 
**Scheme 3.2.** Dynamic “flipping” process proposed to be slow for the azadithiolate in 2.

**Azaditiholate μ-Hydrides** $\left[(\mu-H)\text{Fe}_2(\text{adt,NH})(\text{CO})_2(\text{dppv})_2\right]^+$ and $\left[(\mu-H)\text{Fe}_2(\text{adt,NH}_2)(\text{CO})_2(\text{dppv})_2\right]^{2+}$. Protonation of 2 at room temperature with HBF₄•Et₂O followed by equilibration overnight at room temperature provided two μ-hydride isomers of $\left[(\mu-H)\text{Fe}_2(\text{adt})(\text{CO})_2(\text{dppv})_2\right]\text{BF}_4$. The isomers are distinguishable at room temperature by $^{31}\text{P}\{^1\text{H}\}$ and $^1\text{H}$ NMR spectroscopy, and differ with respect to the positions of the dppv ligands, identical to those seen for the related propanedithiolate, $[1-\mu\text{H}]^+$ (Chapter 2). The bridging hydride isomers are named from their kinetic origins, $[2-\mu\text{H}_A]^+$ being formed first and then isomerizes to $[2-\mu\text{H}_B]^+$. The major isomer in solution, $[2-\mu\text{H}_B]^+$, is $C_s$-symmetric with one diphosphine chelating in the apical and basal orientation and the other dppv is dibasal. In the minor $C_2$-symmetric isomer, $[2-\mu\text{H}_A]^+$, both dppv ligands coordinate in the apical basal orientation. Upon precipitation of a thermally equilibrated solution, the most stable isomer, $[2-\mu\text{H}_B]^+$, is isolated as a brownish-red solid. Upon addition of HBF₄•Et₂O to a CH₂Cl₂ solution of $\left[(\mu-H)\text{Fe}_2(\text{adt})(\text{CO})_2(\text{dppv})_2\right]\text{BARF}_4$, the $\nu_{\text{CO}}$ pattern shifts about 15 cm⁻¹ to higher energy indicative of N-protonation (Scheme 3.3). Solutions of the ammonium μ-hydride, $\left[(\mu-H)\text{Fe}_2(\text{adt,NH}_2)(\text{CO})_2(\text{dppv})_2\right]^{2+}$ ([2-μH-
adt,NH$_2^{2+}$), degrade over a couple days in sealed NMR tubes, providing unknown products observed by $^{31}$P and $^1$H NMR spectroscopy.

**Scheme 3.3.** Protonation and isomerization of 2, showing both terminal and bridging hydride isomers and amine and ammonium forms thereof.

**The Terminal Hydride:** $[(t$-$H$)Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$]$^+$. Addition of [H(OEt)$_2$]$_2$BF$_4$ to a cold CH$_2$Cl$_2$ solution of Fe$_2$(adt)(CO)$_2$(dppv)$_2$ was observed to initially give the terminal hydride $[(t$-$H$)Fe$_2$(adt)(CO)$_2$(dppv)$_2$]$^+$ ($[1$-$t$H]$^+$). Its $^1$H NMR spectrum ($\delta$ -4.2, $t$, $J_{PH} = 73$ Hz) indicates the presence of a single hydride isomer (Figure 3.5). The $^{31}$P NMR spectrum, with proton decoupler off, indicates that the diphosphine on the FeH center is dibasal, since both values of $J_{PH}$ are appropriate for cis coupling. The diphosphine on the other Fe center spans apical and basal sites (Figure 3.6). Solutions of the terminal hydride isomer are stable for ca. one hour at 0
°C, but only for minutes room temperature. As expected from comparison to [1-tH]^+, we found that the isomerization to the \( \mu \)-hydride follows first order kinetics (Figure 3.7).

![Figure 3.5](image)

**Figure 3.5.** High-field \(^1\)H NMR (500 MHz, CD\(_2\)Cl\(_2\)) spectra of
(a) \([\text{(t-H)Fe}_2\text{(pdt)(CO)}_2\text{(dppv)}_2]\text{BF}_4\) (\([1\text{-tH}]\text{BF}_4\)),
(b) \([\text{(t-H)Fe}_2\text{(adt,NH)(CO)}_2\text{(dppv)}_2]\text{BArF}_4\) (\([2\text{-tH}]\text{BArF}_4\)),
(c) \([\text{(t-H)Fe}_2\text{(adt,NH}_2\text{(CO)}_2\text{(dppv)}_2]\text{BArF}_4\)\(_2\) (\([2\text{-tH-adt,NH}_2]\text{(BArF}_4\)\(_2\)),
(d) \([\text{(t-H)Fe}_2\text{(edt)(CO)}_2\text{(dppv)}_2]\text{BF}_4\) (See Chapter 2)
Figure 3.6. $^{31}$P NMR (600 MHz, CD$_2$Cl$_2$, -35 °C) spectrum of [(t-H)Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$]BF$_4$ ([2-tH]BF$_4$).

Figure 3.7. Plot of decay of terminal hydride [2-tH]BF$_4$ at -10 °C, CH$_2$Cl$_2$ solution, as assayed by $^1$H NMR.
The Terminal Hydride Ammonium: \([{(t-H)Fe_2(adt,NH_2)(CO)_2(dppv)}_2]^{2+}\) Upon addition of excess \([H(Et_2O)_2]BAr^F_4\) or \(HBF_4\cdot Et_2O\) to a \(CH_2Cl_2\) solution of \([2-tH]^+\), the mixture converts to the doubly protonated species \([2-tH-adt,NH_2]^2+\). The FT-IR spectrum displays resembles that for \([2-tH]\) but is shifted by \(\sim 15\) cm\(^{-1}\) (Figure 3.12). The \(^1H\) and \(^{31}P\) NMR spectra confirm that \([2-tH-adt,NH_2]^2+\) is a single isomer with both dibasal and apical-basal diphosphine ligands (Figures 3.8, 3.9). Recording the \(^{31}P\) NMR spectrum with the \(^1H\) decoupler off allows for quick location of the \(HFe(dppv)\) \(^{31}P\) NMR resonances (δ 74, 88). A –10 °C \(CD_2Cl_2\) solution containing equimolar concentrations of \([2-tH]^+\) and \([2-tH-adt,NH_2]^2+\) displays individual \(^1H\) and \(^{31}P\) NMR resonances indicating slow intermolecular proton exchange on the NMR time scale (Figure 3.8).

As indicated by both \(^{31}P\{^1H\}\) NMR and FT-IR spectroscopy, exposure of \(CH_2Cl_2\) solutions of \([2-tH-adtNH_2](BAr^F_4)_2\) to air or an equivalent of MeOH results in rapid deprotonation of the ammonium and regeneration of \([2-tH]BAr^F_4\). This deprotonation by MeOH (MeOH\(^+\), \(pK_a(aq) = -2\)\(^{31}\)) demonstrates the very acidic nature of the ammonium in \([2-tH-adt,NH_2]^2+\). Importantly, \([2-tH-adtNH_2](BAr^F_4)_2\) does not spontaneously eliminate \(H_2\), confirming the low hydridicity of the hydride ligand. This result is consistent with that previously observed for \([1-tH]^+\), which also does not eliminate \(H_2\) in the presence of strong acids.
Figure 3.8. $^1$H NMR spectrum (CD$_2$Cl$_2$, 500 MHz, -10 °C) of an equimolar solution of [2-t-H]BAr$_4^-$ and [2-tH-adt,NH$_2$]BAr$_4^-$ F. The signal at δ = -3 is unknown, but increases with time.

Figure 3.9. $^{31}$P NMR (600 MHz, CD$_2$Cl$_2$, -35 °C) spectrum of [(t-H)Fe$_2$(adt,NH$_2$)(CO)$_2$(dppv)$_2$](BAr$_4^-$)$_2$ ([2-tH-adt,NH$_2$](BAr$_4^-$)$_2$). Signals at δ 88 and 74 show additional coupling to the hydride resonance.
Reactivity of [(t-H)Fe₂(adt)(CO)₂(dppv)]⁺ in the Presence of CH₃CN. As stated previously, in CD₂Cl₂ solutions, [2-tH]⁺ and [2-tH-adtNH₂]⁺ do not spontaneously eliminate H₂, confirming the low hydridicity of the hydride ligand. However, the addition of excess or a single equivalent of CH₃CN either [2-tH]⁺ or [2-tH-adtNH₂]⁺ eliminate H₂ and provide several unidentifiable products by ³¹P{¹H} NMR and FT-IR spectroscopies. DuBois has shown that substituting CO for CH₃CN in [HFe(P₂)₂L]⁺ (where L = CO or CH₃CN and P₂ = diphosphine) changes the pKₐ of the dihydrogen ligand by several units. Similarly, we propose that CH₃CN coordinates at the Fe-Fe bond trans to the hydride ligand thereby increasing the hydridicity and allowing subsequent protonation reactions. A logical product would be the CH₃CN-bound diferrous species, [Fe₂(adt)(CO)₂(CH₃CN)(dppv)₂]²⁺; however, the ³¹P{¹H} NMR and FT-IR spectroscopies are inconclusive. The propanedithiolate [1-tH]+ does not eliminate H₂ in the presence of CH₃CN, even in the presence of excess acid.

Hydride-Ammonium Tautomerization in [(t-H)Fe₂(adt)(CO)₂(dppv)]⁺. Protonation of a CH₂Cl₂ solution of 2 with one equiv of HBF₄•Et₂O (vs [H(OEt)₂]BARF₄ used above) afforded a tautomeric mixture of the terminal hydride [2-tH]⁺ and the ammonium [2-adt,NH₂]⁺ compounds, as indicated by FT-IR (Figure 3.10). The formation of the tautomeric mixture is attributed to the ability of the BF₄⁻ to participate in hydrogen-bonding. Consistent with this view, addition of 10 equivalents of [NBu₄]BF₄ increased the [2-tH]⁺/[2-adt,NH₂]⁺ ratio from 2:1 to approximately 1:1 mixture. The dependence of the tautomeric ratio on Log[BF₄] is linear consistent with the stoichiometry (Figure 3.11, Equation 3.1). The [2-tH]⁺/[2-adt,NH₂]⁺ ratio is also strongly affected by solvent. Thus, the ammonium [1H]BF₄ is the exclusive tautomer observed in
solutions of acetone and methanol (Figure 3.12). The IR spectrum for this tautomer has the same $\nu_{CO}$ pattern as 2, but is shifted to higher energy by $\sim$20 cm$^{-1}$. As expected, the tautomerization does not apply to the bridging hydrides as the NMR and IR spectra of $[(\mu-H)Fe_2(adt)(CO)_2(dppv)]BArF_4$ is not noticeably affected by changing counter-anions ($BArF_4^-$ vs $BF_4^-$).

Figure 3.10. React-IR spectra of a solution of [2-tH]BF₄ with 0, 1, 2, 3, 10 equivalents of added [NBu₄]BF₄.

Figure 3.11. Graph of the equilibrium ratio of [2-tH]⁺/[2-adt,NH₂]⁺ (K) vs. Log(excess equivalents [BF₄⁺]) for a CH₂Cl₂ solution of [2-tH]BF₄.
Figure 3.12. Selected FT-IR spectra for the protonation of 2 in CH₂Cl₂ and MeOH solutions to give the terminal hydride and ammonium tautomers. [2-adt,NH₂]⁺ was generated from protonation of 2 with [H(Et₂O)₂][BArF₄] in 100% MeOH at room temperature, the region above 2000 cm⁻¹ was removed due to a MeOH absorption band. The terminal hydride, [2-tH]⁺, was generated from protonation of 2 with 1 equivalent of [H(Et₂O)₂][BArF₄] in CH₂Cl₂ solution at −40 °C. The terminal hydride ammonium [2-tH-adt,NH₂]²⁺, was generated from protonation of 2 with 2 equivalents of [H(Et₂O)₂][BArF₄] in CH₂Cl₂.
The Oxadithiolate Fe$_2$(odt)(CO)$_2$(dppv)$_2$ and Its Terminal and Bridging Hydrides. We prepared Fe$_2$(odt)(CO)$_2$(dppv)$_2$ from Fe$_2$(odt)(CO)$_6^{27,28}$ and confirmed spectroscopically (odt = 2-oxopropane-1,3-dithiolate) (Figure 3.13). Protonation of Fe$_2$(odt)(CO)$_2$(dppv)$_2$ (3) at -78 °C with the strong acid [H(Et$_2$O)$_2$]BAR$_4^+$ afforded the terminal hydride [3-tH]BAR$_4^+$. $^1$H and $^{31}$P NMR analysis confirmed that protonation occurred at a single Fe center, similar to related derivatives (Figure 3.14, Figure 3.15). This terminal hydride was found to isomerize upon warming to give the µ-hydride complex, [3-µH]BAR$_4^+$ (k = 2.6 x 10$^{-4}$ s$^{-1}$, -10 °C), a process following unimolecular kinetics (Figure 3.16). The isomerization rate is similar to that for [2-tH]BAR$_4^+$ (k = 1.4 x 10$^{-4}$ s$^{-1}$, -10 °C) but is faster than [1-tH]BAR$_4^+$ (k = 2.5 x 10$^{-5}$ s$^{-1}$).

Figure 3.13. FT-IR spectra (-40 °C, CH$_2$Cl$_2$) of Fe$_2$(odt)(CO)$_2$(dppv)$_2$ (blue, 3), and [HFe$_2$(odt)(CO)$_2$(dppv)$_2$]BF$_4$ (red, [3-tH]BF$_4$). For [3-tH]BF$_4$, the band for the semi-bridging CO is at ~1910 cm$^{-1}$. 
Figure 3.14. $^{31}\text{P}^{1\text{H}}$ NMR (242 MHz, CD$_2$Cl$_2$) spectra of solution of 3 before (a, -90 °C) and after (b, -90 °C) treated with 1 equiv [H(Et$_2$O)$_2$]BAr$_4^-$ showing complete conversion to [3-$tH$]BAr$_4^-$. 
**Figure 3.15.** $^1H$ NMR spectra before and after the protonation of Fe$_2$(odt)(CO)$_2$(dppv)$_2$ (3) using [H(Et$_2$O)$_2$]BAr$_4$F$_4$. The kinetic terminal hydride (a) (-75 °C, 600 MHz, CD$_2$Cl$_2$) [($t$-H)Fe$_2$(odt)(CO)$_2$(dppv)$_2$]BAr$_4$F$_4$, [3-$t$H]BAr$_4$F$_4$ was found to isomerize upon warming to the bridging hydride isomers (b) (25 °C, 500 MHz, CD$_2$Cl$_2$) [($µ$-H)Fe$_2$(odt)(CO)$_2$(dppv)$_2$]$^+$ [3-$µ$H]BAr$_4$F$_4$. Bridging hydride isomer (b) at $δ$ - 14.5 (qd, Fe-H, $J_{PH1,2,3}$ $\sim$ 20 Hz, $J_{PH4}$ $\sim$ 7 Hz) is assigned to the asymmetric C$_3$ isomer, [3-$t$H$_B$]$^+$ and $δ$ - 15.4 (tt, Fe-H, $J_{PH1,2}$ $\sim$ 20 Hz, $J_{PH3,4}$ $\sim$ 7 Hz) is assigned to the C$_2$ symmetric isomer, [3-$µ$H$_A$]$^+$. 
Figure 3.16. Plot of decay of terminal hydride $[3\text{-tH}]\text{BAr}_4$ at -10 °C, CH$_2$Cl$_2$ solution, as assayed by $^1$H NMR.

Influence of Dithiolate on Equilibrium Position of Bridging Hydride Isomers.

After days in CD$_2$Cl$_2$ solution at room temperature, complexes of the type [HFe$_2$(xdt)(CO)$_2$(dppv)$_2$]BAr$_4$ (xdt = odt, pdt, adtNH, adtNH$_2$) reach an equilibrium of both the $C_S$ ($\#_{-\mu}H_B$) and $C_2$ ($\#_{-\mu}H_A$) symmetric isomers. As indicated by the $^{31}$P and $^1$H NMR spectra, equilibria favor the $C_S$-symmetric isomer (Table 3.1, Figure 3.17). In chapter 2 it was observed that the steric profile of the dithiolate had a profound influence on the half-life of the terminal hydride. Here, we show that the identity of the dithiolate affects the position of this $\mu$-hydride $C_S/C_2$ equilibrium. The equilibrium ratio $C_S/C_2$ follows the trend pdt $>$ adtNH$_2$ $>$ adtNH $>$ odt. Thus, the larger dithiolates stabilize the $C_S$-symmetric isomer ($[\#_{-\mu}H_B]^+$). The propanedithiolate derivative, being the largest in the series, shows the greatest preference for the $C_S$-symmetric isomer ($[1-\mu H_B]^+$). Protonation of the amine to give the ammonium hydride $[2-\mu H\text{-adt,NH}_2]^{2+}$ alters the
equilibrium to give a $C_5/C_2$ isomer ratio closely matching that for $[(\mu\text{-}H)\text{Fe}_2\text{pdt}(\text{CO})_2\text{dppv}_2]^+$. 

<table>
<thead>
<tr>
<th>Bridging Hydrides:</th>
<th>$C_2$ [(\mu\text{H}_A)]$^+$</th>
<th>$C_5$ [(\mu\text{H}_B)]$^+$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$[(\mu\text{-}H)\text{Fe}_2\text{pdt}(\text{CO})_2\text{dppv}_2]^+$, $[1-\mu\text{H}]^+$</td>
<td>6%</td>
<td>94%</td>
</tr>
<tr>
<td>$[(\mu\text{-}H)\text{Fe}_2\text{adt}(\text{CO})_2\text{dppv}_2]^+$, $[2-\mu\text{H}]^+$</td>
<td>32%</td>
<td>68%</td>
</tr>
<tr>
<td>$[(\mu\text{-}H)\text{Fe}_2(\text{adt,NH})(\text{CO})_2\text{dppv}_2]^+$, $[2-\mu\text{H}]^+$</td>
<td>19%</td>
<td>81%</td>
</tr>
</tbody>
</table>

Table 3.1. Thermodynamic ratios of bridging hydride isomers after equilibration in CD$_2$Cl$_2$ solution at room temperature for several weeks. The ratios emphasize the effect of the dithiolate on equilibrium position.

Figure 3.17. Free energy difference for the isomer ratio $C_2$ vs. $C_5$ calculated from the equilibrium mixture of isomers in a CD$_2$Cl$_2$ solution at room temperature (298 K).
To investigate the effect of the bridging hydride equilibrium ratio ($C_s/C_2$) on the corresponding the terminal hydride isomer, we plotted the observed rate constant at $-10$ °C vs the equilibrium ratio (Figure 3.18). Not surprisingly, such plot gives a nearly linear regression indicating that the bridging hydride isomer ratio is a good indicator of the steric profile for the dithiolate, which is directly related to the isomerization rate of the terminal to bridging hydride. The development of such predictive tools could be useful for assaying terminal hydride stability indirectly.

**Figure 3.18.** Graph of the isomerization rate constant at $-10$ °C (CD$_2$Cl$_2$) of the terminal hydride isomer vs the % C$_2$-symmetric isomer.
Large Kinetic Barrier of Proton Transfer for \([\text{([t-H]}\text{Fe}_2\text{(pdt)(CO)}_2\text{(dppv)}_2]^+\). Solutions of \([1-\text{tH}]^+\) in \(\text{CD}_2\text{Cl}_2\) are unreactive toward \(\text{NET}_3\) \((pK_a^{\text{CH}_3\text{CN}} = 18)\) and \(\text{TMG}\) \((pK_a^{\text{CH}_3\text{CN}} = 26)\) (Figure 3.20). Based on the acidity of \([\text{[(H]}\text{Fe}_2\text{(pdt)(CO)}_4\text{(PMe}_3\text{)}_2]^+\) \((pK_a^{\text{CH}_3\text{CN}} \sim 15)\) \(\text{TMG}\) should be sufficiently basic to deprotonate the hydride of \([1-\text{tH}]^+\).\(^\text{10}\) We therefore conclude that the lack of reactivity is a kinetic phenomenon and not a thermodynamic one. Norton et al. have found that ligands with large steric profiles can hinder proton transfer.\(^\text{32,33}\) When large ligands are employed they can prevent access to the transition state of protonation, which is thought to involve a \(\text{BH}^+\cdots\cdots\text{M}\) interaction.\(^\text{34,35}\) Therefore, it is not surprising given the large steric profile of the diphosphine ligand, \(\text{dppv}\), that \([1-\text{tH}]^+\) is incapable of proton transfer. In fact, the formation of the hydride itself is only facile when employing the very acidic and sterically small \(\text{HBF}_4\cdot\text{Et}_2\text{O}\) or \([\text{H(Et}_2\text{O)}_2]\text{Bar}^\text{F}_4\) (although \(\text{Bar}^\text{F}_4\) is large, the acid itself as protonated \(\text{Et}_2\text{O}\) is quite small) (Figure 3.19). Otherwise, reaction of \(\text{1}\) with weak acids, such as \([\text{HPCy}_3]\text{BF}_4\) do eventually provide the expected bridging hydride \([1-\mu\text{H}]^+\) although several days is required for the protonation to occur.
Figure 3.19. $^{31}$P($^1$H) NMR (242 MHz, CD$_2$Cl$_2$) spectrum of 1 after treatment with ~1 equiv of [H(Et$_2$O)$_2$]BAr$_4^+$ showing immediate and complete conversion to [1-tH]BAr$_4^+$ at -80 °C.
Figure 3.20. $^1$H NMR (500 MHz, CD$_2$Cl$_2$) spectra of solutions of [1-tH]BAR$_4$ (a, 20 °C) treated with large excess of NEt$_3$ showing no deprotonation. Scales of intensity are included to show the concentration of NEt$_3$ (δ 2.5, 1.0 ppm) compared to terminal hydride species (triplet δ -3.5 ppm) and growth of bridging hydride species (multiplets δ -14.5, -15.7 ppm).
Odt and Adt Accelerate the Deprotonation of \([{(t\text{-}H)\text{Fe}_2(xdt)(CO)_2(dppv)}]^+\).

To probe the effect of the aza and oxadithiolate on proton relay, we observed the facility with which a CD$_2$Cl$_2$ solution of [3-tH]BAR$_4^F$ deprotonates. At -78 °C, [3-tH]BAR$_4^F$ is unreactive toward base (Figure 3.21), but upon warming to 0 °C, two products form, [3-µH]BAR$_4^F$ and 3, as assayed by $^1$H and $^{31}$P NMR spectroscopy. The ratio of these two products was unaffected by the concentration of the base as well as its pK$_a$, as indicated by deprotonations with both strong and weak bases, respectively, tetramethylguanidine (TMGH$^+$, pK$_a = 23$) and PPh$_3$ ([HPPh$_3$]BF$_4$, pK$_{CD2Cl2} = 1.6$) (Figure 3.22).$^{16}$ This result is in contrast to [1-tH]BAR$_4^F$, which cannot be deprotonated even at room temperature, when isomerization to [1-µH]$^+$ eventually occurs. By comparison, the azadithiolate derivative [2-tH]BAR$_4^F$ is immediately deprotonated with PBu$_3$ ([HPBu$_3$]BF$_4$, pK$_{CD2Cl2} = 8.2$) even at -90 °C, exclusively providing 2 (Figure 3.23). The close similarity of the FT-IR spectra in the ν$_{CO}$ region for [1-tH]BAR$_4^F$, [2-tH]BAR$_4^F$, and [3-tH]BAR$_4^F$ suggest that these terminal hydrides should have similar thermodynamic acidities.$^{36}$ The similar thermodynamic acidities of these three hydrides indicate that the differing chemistry is a kinetic phenomenon and related to identity of the heteroatom in the dithiolate.
Figure 3.21. $^{31}$P{$^1$H} NMR (242 MHz, CD$_2$Cl$_2$, -75 °C) spectra before (left) and after (right) treatment of solutions (CD$_2$Cl$_2$, -75 °C) of [2-tH]BAR$_4^F$ and [3-tH]BAR$_4^F$ with ~100 equiv of Et$_3$N. Upon addition of NEt$_3$ to [2-tH]BAR$_4^F$, resulting $^{31}$P NMR (upper left) shows complete conversion to 2, whereas for [3-tH]BAR$_4^F$, no change (lower left) is seen until warming near 0 °C.
Figure 3.22. $^{31}$P{$^1$H} NMR (242 MHz, CD$_2$Cl$_2$) spectra of solutions of [3-tH]BAr$_4^F$ treated with $\sim$1 equiv of tetramethylguanidine (a, 20 °C), and repeated with $\sim$ 1 equiv of PPh$_3$, (b, 20 °C) and large excess ($>$100 equiv) of NEt$_3$ (c, 20 °C) showing no dependence on concentration or p$K_a$. Upon addition of either base to [3-tH]BAr$_4^F$ the resulting $^{31}$P{$^1$H} NMR shows ca. 50% conversion to 3 (δ 90) and ca. 50% conversion to [3–µH]BAr$_4^F$ (δ 89, 88).
Figure 3.23. $^3{\text{P}}^1{\text{H}}$ NMR (242 MHz, CD$_2$Cl$_2$) spectra of a solution of [2-tH]BAR$_4^F$ treated with ~1 equiv of PMe$_2$Ph ([HPMe$_2$Ph]BF$_4$, pK$_{CD_2Cl_2}=$ 5.7) (a, -60 °C) showing no reaction and repeated with ~2 equiv of PBu$_3$ ([HPBu$_3$]BF$_4$, pK$_{CD_2Cl_2}=$ 8.2) (b, -60 °C) showing complete conversion to 2 and some generation of [HPBu$_3$]BAR$_4^F$ ([HPBu$_3$]$^+$ δ 11, PBu$_3$ δ -33).
The oxadithiolate provides further insight into the mechanism of proton relay given that a 50/50 mixture of [3-μH]^+ and 3 is generated regardless of the base and Only when the solution is warmed to 0 °C. A consistent mechanism involves the terminal hydride in equilibrium with an oxonium tautomer. This equilibrium is defined by the relative first-order rate constants $k_{XH}$ and $k_{FeH}$ (Scheme 3.4). The iron hydride is much more abundant than to oxonium hydride, but deprotonates far more slowly. Thus, deprotonation only occurs when the oxygen atom relays a proton to the base. The rate constant of the internal proton relay ($k_{OH}$) is coincidentally the same as that for the terminal to bridging hydride, $k_{ISO}$, which explains the observed 50/50 mixture upon warming to 0 °C. The regiochemistry of the hydride also has a profound effect as the three bridging hydrides, [1-μH]^+, [2-μH]^+, and [3-μH]^+, are not deprotonated by NEt₃ at room temperature (Figure 3.24).
Figure 3.24. $^1$H NMR (500 MHz, CD$_2$Cl$_2$) spectra of solutions of [1-$\mu$H]Bar$_4$ (a, 20 °C), [2-$\mu$H]Bar$_4$ (b, 20 °C), and [3-$\mu$H]Bar$_4$ (c, 20 °C) treated with large excess of NEt$_3$ showing no deprotonation after 1 day. High-field region was magnified by 500x.
Azadithiolate Strongly Affects the Protonation of Fe$_2$(xdt)(CO)$_2$(dppv)$_2$. To confirm the proposed proton relay mechanism, we studied the protonation of 1, 2, and 3 at various temperatures. The presence of the azadithiolate was found to strongly affect the rate of protonation. The strong acid [H(Et$_2$O)$_2$]BAR$_4^+$ protonated 1, 2, and 3 quickly at -90 °C, but the significantly weaker acid [HPMe$_2$Ph]BF$_4$ (pK$_{CD2Cl2}^{+}$ = 5.7) protonated only 2 (-90°C) (Figure 3.25), and not 1 or 3. The pK$_a$ of [2-tH]$^+$ is bracketed by the finding that 2 is not protonated by [HPBu$_3$]BF$_4$ (pK$_{CD2Cl2}^{+}$ = 8.2). The implication that the acidity of the ammonium and terminal hydride tautomers of [2-tH]$^+$ are comparable is supported by the previously reported finding that the ratio of the ammonium and terminal hydride tautomers can be shifted by the solvent and concentration of BF$_4^-$.

These results are consistent with a mechanism whereby hydride formation is
regulated by the basicity of the heteroatom in the dithiolate: the ammonium center in \( \text{2} \) is easily protonated and then quickly relays protons to Fe. In contrast, for complexes with weakly basic oxadithiolate \( (pK\text{CD}_{2}Cl_{2}(R_{2}OH^{+}) = -4.7 \text{ to } 1.6) \) or nonbasic propanedithiolate, the Fe site can only be protonated by strong acids, even though the basicities of these diiron centers are very similar. In efforts to observe the kinetic product of protonation for \( \text{2} \) (the ammonium \( [\text{2-adt,NH}_{2}^{+}] \)) we recorded FT-IR spectra within seconds after the addition of 1 equivalent \( [\text{H(Et}_{2}\text{O)}_{2}]\text{BAR}_{4}^{\text{F}} \) at \(-97 \degree \text{C}\) (Figure 3.26). The spectrum obtained 30 seconds after addition of acid was, however, inconclusive.
Figure 3.25. $^{31}$P($^1$H) NMR (242 MHz, CD$_2$Cl$_2$) spectra of solutions of 2 (a, -80 °C) treated with ~1 equiv of [HPPh$_3$]BAR$_4^-$ (pK$_{CD_2Cl_2}$ = 1.6) (b, -80 °C) showing complete conversion to [2-tH]BAR$_4^-$.
Figure 3.26. *In situ* RIR spectra of a CH$_2$Cl$_2$ solution of 2 at –97 °C before (a), and after (b) treatment with 1 equiv [H(EEt$_2$O)$_2$]BAr$_4^F$ showing initial formation of the N-protonated tautomer [2-adtNH$_2$]BAr$_4^F$. Spectrum c was recorded 30 seconds later showing complete conversion to [2-fH]BAr$_4^F$. Spectrum d is a MeOH solution of [2-adtNH$_2$]BAr$_4^F$ recorded at 20 °C.
Conclusions.

In our modeling of the active-site of [FeFe]-hydrogenase we work solely with diphosphine-substituted derivatives. The highly-substituted diiron complexes of the type Fe$_2$(xdt)(CO)$_2$(dppv)$_2$ represent promising models for the enzyme active site (Table 3.2). As such, the properties of the terminal hydride derivatives, [(t-H)Fe$_2$(xdt)(CO)$_2$(dppv)$_2$]$^+$, may be expected to likely display hydridicity and acidity similar to that of the enzyme. If these assumptions are valid, the resistance of [1-tH]$^+$ toward protonolysis by very strong acids is highly relevant to a key step in the enzymatic mechanism: diferrous terminal hydrides must first be reduced prior to protonation.

Table 3.2  Selected properties of Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$ and the H$_{\text{red}}$ form of D. desulfuricans [FeFe]-hydrogenase.

<table>
<thead>
<tr>
<th></th>
<th>$\nu_{\text{CO}}$ cm$^{-1}$</th>
<th>$E_{1/2}(\text{Fe}_2^{\text{I,I}},\text{H}^+,\text{H}^-)$ vs Fc$^{0/+}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>H$_{\text{red}}$</td>
<td>1925</td>
<td>1.07 V$^a$</td>
</tr>
<tr>
<td>[2-tH]$^+$</td>
<td>1878</td>
<td>0.88 V</td>
</tr>
</tbody>
</table>

Incorporation of the biologically relevant azadithiolate strongly facilitates proton transfer to and from the apical site on Fe, but only to the extent that the acid can protonate the bridgehead atom. The oxadithiolate is not as effective as a proton relay, owing to the weak basicity of the ether. Thus, only the azadithiolate 2 enables hydride formation from weak acids. Consistent with the highly-active proton relay present in 2, the pK$_a$’s of the ammonium tautomer and the terminal hydride tautomer appear to be closely matched, as changing the solvent and even counterion can influence the position of this equilibrium.
Although the electronic parameters of Fe₂(xdt)(CO)₂(dppv)₂ approach those of the enzyme, the structural characteristics do not. In the fully-reduced form of the enzyme, H₉red, one iron coordination site is “rotated”, exposing a vacant terminal position primed for protonation. Such rotated structures require no rearrangement to undergo protonation. The diiron dithiolates in this work contain the sterically bulky diphosphine ligand, dppv. It has been previously demonstrated that such bulky ligands often inhibit the first step of protonation reactions, the formation of a hydrogen bond. The results present in this report may not be consistent with the reactivity of the enzyme. Independent of these concerns, it is apparent that azadithiolates can function as a proton relay from the terminal hydride position. Azadithiolate may therefore be required to facilitate deprotonation of a H₂ species, which likely is very unstable with respect to loss of H₂ (Figure 3.27).

**Figure 3.27.** Possible utility of the azadithiolate to aid in deprotonation of a transient dihydrogen complex.
Experimental Procedures

All manipulations were conducted using standard Schlenk techniques. Solvents were filtered through activated alumina and subsequently degassed. $^1$H and $^{31}$P NMR spectra were acquired on a Unity Varian 500 or a Unity Varian 600 spectrometer. IR spectra were collected on a Mattson Infinity Gold FTIR spectrometer. Cis-1,2-bis(diphenylphosphino)ethylene (dppv) and HBF$_4$•Et$_2$O solution were purchased from Aldrich. Fe$_2$(edt)(CO)$_2$(dppv)$_2$, Fe$_2$(pdt)(CO)$_2$(dppv)$_2$, and [H(Et$_2$O)$_2$]BAR$_4$F ($\text{BAR}_4^2$ = [B(C$_6$H$_3$-3,5-(CF$_3$)$_2$)$_4$]) were prepared according to literature procedures.

Preparation of [HPPPh$_3$]BAR$_4^F$. A solution of [H(Et$_2$O)$_2$]BAR$_4^F$ (0.385 g, 0.377 mmol) in Et$_2$O (10 mL) at -40 °C was transferred via cannula into a solution of PPh$_3$ (0.097 g, 0.370 mmol) in Et$_2$O (10 mL) at -40 °C. Solvent was removed under vacuum, leaving a white solid. Yield: 0.350 g (83%). $^{31}$P{$^1$H} NMR (202 MHz, CD$_2$Cl$_2$, 20 °C): $\delta$ 7.0 (s).

Fe$_2$(adt,NH)(CO)$_4$(dppv). A solution of Fe$_2$(adt,NH)(CO)$_6$ (806 mg, 2.08 mmol) and dppv (820 mg, 2.06 mmol) in 30 mL degassed of toluene was treated with a solution of anhydrous Me$_3$NO (156 mg, 2.06 mmol) in ca. 5 mL of MeCN, resulting in bubbling and darkening of the reaction mixture. After 30 min, FT-IR spectroscopy indicated complete conversion to product. The solvent was removed under vacuum and the solid was extracted into 5 mL of CH$_2$Cl$_2$. This green solid was washed with 50 mL of hexane to yield the green product 1.03 g (1.41 mmol, 68% yield). All spectroscopic analyses matched those of the previously reported synthesis.$^{30}$

Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$, (2). A solution of Fe$_2$(adt,NH)(CO)$_4$(dppv) (300 mg, 0.412 mmol) in 50 mL toluene was treated with dppv (165 mg, 0.416 mmol) in 10 mL of
toluene. The mixture was photolyzed with a 100 W UV immersion lamp, $\lambda_{\text{max}} = 356$ nm (Spectroline) until the conversion was complete (~12 h) as indicated by IR spectroscopy. The solvent was removed in vacuum, and product was extracted into 10 mL of CH$_2$Cl$_2$. The product precipitated as an olive-green powder upon the addition of 50 mL of hexanes. Yield: 190 mg (43%). $^1$H NMR (500 MHz, CD$_2$Cl$_2$): $\delta$ 8.1 – 7.0 (m, 40H, P(C$_6$H$_5$)$_2$), 2.3 (s, 4H, (SCH$_2$)$_2$NH). $^{31}$P NMR (CD$_2$Cl$_2$): $\delta$ 91.2. FT-IR (CH$_2$Cl$_2$, cm$^{-1}$): $\nu$CO = 1888, 1868. 

\[
\text{[(t-H)Fe}_2\text{(pdt)(CO)}_2\text{(dppv)}_2]\text{BF}_4, \ [1\text{-tH}]\text{BF}_4. \ A \text{ dark green solution of Fe}_2\text{(pdt)(CO)}_2\text{(dppv)}_2 \ (205 \text{ mg, 0.192 mmol) in 5 mL of CH}_2\text{Cl}_2 \text{ was treated at -40 °C with HBF}_4 \text{•Et}_2\text{O (5 mL 0.04 M, 0.2 mmol). The resulting darker green solution was then treated with 100 mL cold (-78 °C) Et}_2\text{O producing a green precipitate. The precipitate was collected by filtration at -78 °C, washed with 2 x 10 mL of Et}_2\text{O, and pumped dry prior to storage in nitrogen glove-box freezer (140 mg, 63% yield).} \]

$^1$H NMR (500 MHz, CD$_2$Cl$_2$): $\delta$ -3.5 (t, Fe-H, $J_{\text{PH}} = 76$ Hz). $^{31}$P{$^1$H} NMR (202 MHz, CD$_2$Cl$_2$): $\delta$ 99 (s), 91 (s), 86 (s), 68 (s). Selective $^{31}$P decoupled $^1$H NMR verified that only the signals at $\delta$ 91 and 86 coupled to the hydride ligand, each yielding doublets with $J_{\text{PH}} = 76$ Hz. FT-IR (CH$_2$Cl$_2$, cm$^{-1}$) $\nu$CO: 1965, 1905 $\nu$Fe-H: 1888.

\[
\text{[(t-H)Fe}_2\text{(adt,NH)(CO)}_2\text{(dppv)}_2]\text{BAr}_4, \ [2\text{-tH}]\text{BAr}_4. \ In \ a \ J. \ Young \ NMR \ tube,} \text{ CD}_2\text{Cl}_2 \text{ was distilled and frozen above solid Fe}_2\text{(adt,NH)(CO)}_2\text{(dppv)}_2 \ (5 \text{ mg, 0.005 mmol) and [H(Et}_2\text{O)}_2]}\text{BAr}_4 \ (5 \text{ mg, 0.005 mmol). The J. Young tube was then placed directly into a -78 °C bath, and analyzed by low temperature NMR spectroscopy. High field $^1$H NMR (600 MHz, CD$_2$Cl$_2$, -40 °C): $\delta$ -4.2 (t, Fe-H, $J_{\text{PH}} = 73$ Hz). $^{31}$P{$^1$H} NMR (242 MHz, CD$_2$Cl$_2$, -40 °C): $\delta$ 100 (s), 99 (s), 82 (s), 72 (s). $^{31}$P NMR (decoupler off)
verified that only the signals at δ 99 and 82 were coupled to the hydride.

\((t\text{-H})\text{Fe}_2(\text{adt},\text{NH}_2)(\text{CO})_2(\text{dppv})_2[\text{Bar}_4^\text{F}_2]_2\) \text{[2-adt,NH}_2-t\text{H}(\text{Bar}_4^\text{F}_2)_2).\) In a J. Young NMR tube CD$_2$Cl$_2$ was distilled and frozen above solid Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$ (5 mg, 0.005 mmol) and [H(Et$_2$O)$_2$]Bar$_4^\text{F}_4$ (10 mg, 0.010 mmol). The J. Young tube was then placed directly into a -78 °C bath and analyzed with low temperature NMR spectroscopy. High field $^1$H NMR (600 MHz, CD$_2$Cl$_2$, -40 °C): δ - 4.95 (t, Fe-H, J$_{PH}$ = 72 Hz). $^{31}$P{$^1$H} NMR (242 MHz, CD$_2$Cl$_2$, -40 °C): δ 98 (s), 89 (s), 76 (s), 74 (s). $^{31}$P NMR (decoupler off) verified that only the signals at δ 89 and 74 coupled to the hydride.

Treatment of Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$ (2) with 1.5 equiv of [H(Et$_2$O)$_2$]Bar$_4^\text{F}_4$. In a J. Young NMR tube CD$_2$Cl$_2$ was distilled and frozen above solid Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$ (5 mg, 0.005 mmol) and [H(Et$_2$O)$_2$]Bar$_4^\text{F}_4$ (7.5 mg, 0.0075 mmol). The frozen J. Young tube was then placed directly into a -78 °C bath, and analyzed with low temperature NMR spectroscopy. High field $^1$H NMR (600 MHz, CD$_2$Cl$_2$, -40 °C): δ - 4.95 (t, Fe-H, J$_{PH}$ = 72 Hz), δ - 4.2 (t, Fe-H, J$_{PH}$ = 73 Hz).

Treatment of Fe$_2$(adtNH)(CO)$_2$(dppv)$_2$ (2) with 1 equiv of [H(Et$_2$O)$_2$]Bar$_4^\text{F}_4$ at -97 °C monitored by insitu IR-spectroscopy. In a 50-mL 3 neck flask Fe$_2$(adtNH)(CO)$_2$(dppv)$_2$ (35 mg, 0.033 mmol) was connected to the ReactIR instrument, dissolved in 3 mL of CH$_2$Cl$_2$, and cooled to -97 °C with a MeOH/N$_2$(l) bath. The instrument was then programmed to collect spectra at 30 second intervals. Solid [H(Et$_2$O)$_2$]Bar$_4^\text{F}_4$ (37 mg, 0.032 mmol) was then added. The temperature of the reaction solution was monitored by a type-K braided thermocouple inserted through a septum into the reaction solution.

[Fe$_2$(pdt)(µ-H)(CO)$_2$(dppv)$_2$]PF$_6$. [1-µH]PF$_6$. In a 250-mL Schlenk flask, a dark
green solution of Fe₂(pdt)(CO)₂(dppv)₂ (200 mg, 0.187 mmol) in 5 mL CH₂Cl₂ was treated at room temperature with HCl•Et₂O solution (0.2 mL, 2.0 M, 0.2 mmol). Solvent was removed under vacuum, and the solid was redissolved in MeOH. The product was precipitated as its hexafluorophosphate salt by addition of ca. 5 mL of an aqueous saturated solution of NH₄PF₆. The brownish precipitate was then transferred onto a Celite plug in an “air sensitive column” where it was washed with 6 x 30 mL H₂O and 6 x 30 mL Et₂O. The product was then eluted through the Celite with CH₂Cl₂, and solvent removed under vacuum. The product was dissolved with a mixture of 5 ml of CH₂Cl₂ and 5 mL of MeOH; treatment of this solution with 50 mL of hexane yielded a brown powder, which was collected by filtration, dried, and stored in the glove box (170 mg, 75% yield). High-field ¹H NMR (500 MHz, CD₂Cl₂): δ -14.5 (dtd, µ-H, JₚH₁ = 24 Hz, JₚH₂ = 19 Hz, JₚH₃ = 10 Hz), -15.6 (tt, µ-H, JₚH₁ = 24 Hz, JₚH₂ = 6 Hz). ³¹P{¹H} NMR (202 MHz, CD₂Cl₂): δ 89.6 (s), 89.6 (s); δ 76.7 (s), 82.8 (s), 84.2 (s), 87.8 (s). FT-IR (CH₂Cl₂, cm⁻¹) νCO: 1958. Anal. Calcd for C₅₇H₅₁F₆Fe₂O₂P₅S₂: C, 56.45 (56.28); H, 4.24 (4.28); Fe 9.21 (8.79).

**Isomerization of [1-µH]⁺ to [1-µµH]**. In a J. Young NMR tube CD₃CN (0.7 mL) was distilled and frozen above solid Fe₂(pdt)(CO)₂(dppv)₂ (5 mg, 0.005 mmol), H(OEt₂)₂BArF₄ (5 mg, 0.005 mmol), and hexamethylbenzene (0.5 mg, 0.005 mmol). The J. Young tube was then placed into a -40 °C bath and analyzed with low temperature NMR spectroscopy. Data was collected as an array over 2 h showing complete consumption of [1-µH]⁺ and growth of two isomers of [1-µµH]⁺, first [1-µµH₃]⁺ then [1-µµH₄]⁺. The terminal hydride triplet at δ -3.5 was integrated from each FID against the internal standard hexamethylbenzene. The decay of terminal hydride species followed first order
kinetics at all temperatures. An Eyring plot was constructed from these data indicating
\[ \Delta H^\ddagger + 81 \text{ kJ}, \Delta S^\ddagger - 23 \text{ J/K\cdot mol}. \]

\[
[(t-H)\text{Fe}_2(\text{odt})(\text{CO})_2(\text{dppv})_2]\text{BAr}_4, [3\text{-}tH]\text{BAr}_4.
\]
In a J. Young NMR tube CD$_2$Cl$_2$
was distilled and frozen above solid Fe$_2$(odt)(CO)$_2$(dppv)$_2$ (7 mg, 0.007 mmol) and
[H(Et$_2$O)$_2$][BArF$_4$] (7 mg, 0.007 mmol). The J. Young tube was then placed directly
into a -78 °C bath and analyzed by low temperature NMR spectroscopy. High field
$^1$H NMR (600 MHz, CD$_2$Cl$_2$, -40 °C): \( \delta = 2.7 \) (t, Fe-H, \( J_{PH} = 72 \text{ Hz} \)). $^{31}$P{$^1$H} NMR (242 MHz, CD$_2$Cl$_2$, -40 °C): \( \delta = 99 \) (s), 94 (s), 89 (s), 69 (s). After isomerization: $^1$H NMR (600 MHz, CD$_2$Cl$_2$, 25 °C): \( \delta = 14.5 \) (qd, Fe-H, \( J_{PH1,2,3} \sim 20 \text{ Hz}, J_{PH4} \sim 7 \text{ Hz} \)), \( \delta = 15.4 \) (tt, Fe-H, \( J_{PH1,2} \sim 20 \text{ Hz}, J_{PH3,4} \sim 7 \text{ Hz} \)). $^{31}$P{$^1$H} NMR (242 MHz, CD$_2$Cl$_2$, 25 °C): \( \delta = 89, 88; 86, 84, 83, 78 \).

Isomerization of [3-tH]$^+$ to [3-\( \mu \)H]$^+$. In a J. Young NMR tube CD$_2$Cl$_2$ (0.7 mL)
was distilled and frozen above solid 3 (7 mg, 0.007 mmol), [H(Et$_2$O)$_2$]BArF$_4$ (7 mg, 0.007
mmol), and hexamethylbenzene (0.5 mg, 0.005 mmol). The J. Young tube was then
placed into a -40 °C bath (MeCN, CO$_2$(s)) and analyzed by low temperature NMR
spectroscopy. Data was collected as an array over 2 h showing nearly complete
consumption of [3-tH]$^+$ and growth of two isomers of [3-\( \mu \)H]$^+$. The terminal hydride
triplet at \( \delta = 2.7 \) was integrated from each FID against the internal standard
hexamethylbenzene. The decay of terminal hydride species followed first order kinetics.

References:


28. Lawrence, J. D.; Li, H.; Rauchfuss, T. B. "Beyond Fe-only Hydrogenases: N-Functionalized 2-Aza-1,3-dithiolates Fe₂[(SCH₂)₂NR](CO)ₓ (x = 5, 6)" Chem. Commun. 2001, 1482-3.


Chapter 4

Proton Reduction Catalysis by Azadithiolate Terminal Hydrides

Introduction

The [FeFe]-hydrogenase’s are Nature’s fastest catalyst for both hydrogen oxidation and hydrogen production.\textsuperscript{1-3} To compare the catalytic properties of [FeFe]-hydrogenase with other heterogeneous catalysts two key parameters are usually considered: overpotential and turnover frequency (molecules H\textsubscript{2} produced per active site per second).\textsuperscript{4-7} Overpotential is defined as an additional electromotive potential required for a reaction to occur at a particular rate, it is a kinetic parameter.\textsuperscript{8} By examining active enzymes supported on a pyrolytic graphite edge (PGE) electrode, the turnover frequency of the [FeFe]-hydrogenase of \textit{C. acetobutylicum} has been estimated to be 4,000-6,000 turnovers per second.\textsuperscript{9-11} These enzyme-modified electrodes achieve \textasciitilde40\% of the cathodic current densities of a platinum electrode. When taking into account that the amount of absorbed hydrogenase (3.3 pmol/cm\textsuperscript{2}) is much less than that of the active sites per platinum (1.1 nmol/cm\textsuperscript{2}),\textsuperscript{12} it appears that hydrogenase enzymes are at least similar if not faster than platinum (Table 4.1).\textsuperscript{9,13}
Table 4.1 Catalytic parameters of a [FeFe]-hydrogenase and platinum.

<table>
<thead>
<tr>
<th></th>
<th>Turnover frequency (molecules H₂ s⁻¹)</th>
<th>Overpotential (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>[FeFe]-Hydrogenase</td>
<td>4,000-6,000</td>
<td>0.080</td>
</tr>
<tr>
<td>Polished Pt-electrode</td>
<td>~100</td>
<td>0.080</td>
</tr>
</tbody>
</table>

In contrast to the [FeFe]-hydrogenase enzyme, typical catalysts derived from diiron dithiolates require high overpotentials and still operate at low turnover rates. Following the discovery that H₂-evolution is catalyzed by [Fe₂(pdt)(CN)(CO)₄(PMe₃)]⁻¹⁴ many diiron dithiolato carbonyl complexes have been found to exhibit similar behavior. Two general mechanisms have been identified that differ in the sequence of reduction vs protonation of the [Fe(I)]₂ center. From the highly negative reduction potentials of diiron dithiolato cyanides, e.g. [Fe₂(pdt)(CN)₂(CO)₄]²⁻ \( (E_{\text{red}} = -2.74 \text{ V vs Fc}^{0/+}) \), we concluded that reduction prior to protonation was unlikely in a biological setting.¹⁴,¹⁵ Instead, we propose that the diiron(I) center protonates prior to reduction. Thus, functional models for the reduced state of the active site must protonate initially. Indeed many complexes of the type \([\text{Fe}_2(\text{pdt})(\text{CO})_{4-x}L_{2+x}]^z\) protonate readily and the products, \(\mu\)-hydrido complexes,¹⁶ are catalytically active. Such \(\mu\)-hydrides are probably not relevant to the catalytic mechanism of the [FeFe]-hydrogenases, which are assumed¹⁷ to operate via a terminal hydride on the distal Fe center.¹⁸

As discussed in Chapters 2 and 3, diiron(I) dithiolates protonate initially to provide terminal hydrides, which tend to isomerize to the \(\mu\)-hydrides.¹⁹ Many of the early examples of terminal hydrides were restricted by very low temperatures,¹⁹ but we
subsequently reported that the stabilities of the hydrides was improved when incorporated into sterically crowded diiron dithiolato carbonyls. Thus, \( \text{Fe}_2(\text{pdt})(\text{CO})_2(\text{dppv})_2 \) undergoes protonation to give terminal hydrides, with a half-life of several minutes at room temperature. Poorly-basic diiron azadithiolates, such as \( \text{Fe}_2(\text{adt})(\text{CO})_4(\text{dppe}) \), do not provide hydrides upon protonation but instead protonate at the amine.\(^{20}\) However, with the more basic “tetra-phosphine” \( \text{Fe}_2(\text{adt,NH})(\text{CO})_2(\text{dppv})_2 \) protonation occurs at the amine and is quickly relayed to the terminal position, providing a model complex suitably poised for biomimetic catalysis. This chapter describes the culmination of thermally robust terminal hydride complexes incorporating the azadithiolate cofactor leading to functional catalysts. These synthetic catalysts display turnover frequencies around 2,000 s\(^{-1}\) at 0 °C, approaching those of the enzyme.
Reduction of the Terminal Hydride \([((t\text{-}H)\text{Fe}_2(pdt)(CO))_2(dppv)]^+\). Reductions of metal hydrides have been reported to substantially increase the basicity of the hydride ligand. According to Tilset et al., one-electron redox changes can change the acidity of hydrides by more than 20 $pK_a$ units. However, these changes have only been observed for monometallic complexes.\(^{21}\)

Cyclic voltammetric (CV) studies (CH$_2$Cl$_2$ solution, 0 °C, vs Ag|AgCl) show that the terminal hydride [1-tH]BF$_4$ reduces at a potential ~200 mV milder than the isomeric bridging hydride [1-µH]BF$_4$ (Figure 4.1). This difference in energy corresponds to ~4.6 kcal/mol (23.06 kcal/V x 0.2 V), and represents a significant difference when compared to reducing a bridging hydride.\(^{22}\)

**Figure 4.1.** Cyclic voltammogram (CH$_2$Cl$_2$, 0°C) of a 0.1 M [NBu$_4$]PF$_6$ solution of both [1-tH]$^+$ and [1-µH]$^+$ (a), and again after isomerized fully to [1-µH]$^+$. Although clean solutions of [1-tH]$^+$ have been obtained, the slightly isomerized (a) unequivocally shows two discrete couples, $\Delta E_{1/2} = 200$ mV. The working electrode was a 0.3 mm diameter glassy carbon electrode, the reference was a Ag wire in saturated KCl solution, and the counter electrode was a platinum wire.
We assigned both the terminal hydride and bridging hydride reduction events of [1-H]BF₄ to 1e⁻ processes as indicated by the similarity of the dependence of \( i_p \) vs \( \nu^{1/2} \) for both [1-tH]BF₄ and its conjugate base 1 (Figure 4.2), which we have established undergoes a 1e⁻ oxidation to 1⁺ (the recently described model of Hox).²³,²⁴ Nineteen electron metal hydrides are extremely rare. There is only one example to our
knowledge, which involves a quasi-reversible redox couple of CpCo(dppe)H$^{0/-}$ at 20 V/s.$^{25}$ We therefore propose that for the $[1\text{-tH}]^{+/-0}$ couple the “Fe$(\mu$-pdt)(dppv)(CO)$(\mu$-CO)” subunit undergoes reduction forming a mixed-valence 35-electron terminal hydride, $[1\text{-tH}]^{0}$. Such reduced terminal hydride could isomerize to the bridging hydride, however, at 0 °C, the $[1\text{-tH}]^{+/-0}$ couple is reversible ($i_{pa}/i_{pc} = 0.64$) at scan rates as slow as 25 mV/s, which implies that the reduced hydride has a half-life of at least several seconds at this temperature.

Figure 4.3. Cyclic voltammogram (CH$_2$Cl$_2$, 0 °C) of 1 mM $[(t\text{-H})Fe_2(pdt)(CO)_2(dppv)_2]BF_4$ ([1-tH]BF$_4$) (a) and with one equivalent of [HPPPh$_3$]BF$_4$ (b).
Once reduced, $[1\text{-}t\text{H}]^0$ is susceptible to protonation by weak acids. We can estimate the $pK_a$ of $[1\text{H}]^0$ by the strength of the acid that renders the $[1\text{H}]^{+/0}$ couple irreversible. After protonation of $[1\text{H}]^0$ and elimination of $H_2$, $1^+$, would be immediately reduced at the electrode following an ECE type mechanism, thus not only is the wave irreversible but the current is doubled. For these CV titrations, we used phosphonium acids, for which Morris has established a $pK_{\text{CD2Cl2}}$ scale.\textsuperscript{26} In the presence of $[\text{HPPPh}_2\text{Me}]\text{BF}_4$ ($pK_{\text{CD2Cl2}} = 3.3$), the $[1\text{H}]^{+/0}$ couple remained reversible; however, the couple became irreversible with the slightly stronger acid $[\text{HPPPh}_3]\text{BF}_4$ ($pK_{\text{CD2Cl2}} = 1.6$) (Figure 4.3). These experiments show that reduction of $[1\text{H}]^+$ increases the basicity of the iron hydride by at least eleven $pK_{\text{CD2Cl2}}$ units ($[1\text{H}]^+$ is unreactive toward HOTf, $pK_{\text{CH3CN}} \sim -12$).\textsuperscript{27,28} As the 35-electron mixed valence hydride $[1\text{-}t\text{H}]^0$ contains two metals by which to distribute a redox change a $\Delta pK_a < 20$ would be expected.\textsuperscript{29}
Figure 4.4. Cyclic voltammograms of 1 mM [(t-H)Fe$_2$(S$_2$C$_3$H$_6$)(μ-CO)(CO)(dppv)$_2$]BF$_4$ ([1-tH]$^+$) with increasing equivalents of HBF$_4$$\cdot$Et$_2$O collected at 100 mV/s at 0 °C. The growth of the peak at –1.3 V are assigned to catalysis by trace amounts of [1-μH]$^+$.

Proton Reduction Catalysis from [(t-H)Fe$_2$(pdt)(CO)$_2$(dppv)]$^+$. To probe the catalytic proton reduction by terminal hydride derivatives, cyclic voltammetry was conducted at 0 °C. At this temperature the isomerization of [t-2H]$^+$ to [μ-2H]$^+$ is slow. The strong acid HBF$_4$$\cdot$Et$_2$O (pK$_a$ ~ -2 in MeCN) readily converts the diiron(I) complex to its protonated state [1-tH]$^+$. Addition of increasing equiv of HBF$_4$$\cdot$Et$_2$O results in an increase current for the −1.67 V couple (vs Fe$^{0+/}$, -1.25 V vs Ag|AgCl) indicative of catalysis (Figure 4.4). Plots of $i_d/i_p$ vs [H$^+$] are linear up to 17 equiv of acid, indicating a
catalytic pathway that is second order with respect to [H$^+$].$^{30-34}$ The pathway for hydrogen evolution therefore entails protonation of the mixed-valence hydride [1-$t$H]$^0$ to release H$_2$, affording 1$^+$ (Scheme 4.1). As the 1$^{+/-0}$ couple occurs at $\sim$0 V vs Ag|AgCl, at the applied potential 1$^+$ would be immediately reduced to 1, completing the cycle. Interestingly, the small amount of [1-$\mu$H]$^+$ present in solution also responds to increasing equiv of HBF$_4$•Et$_2$O but at potentials more negative by 200 mV. It is clear that [1-$\mu$H]$^+$ is a faster catalyst than is [1-$t$H]$^+$ because $i_c$ is more responsive to [HBF$_4$•Et$_2$O], and from NMR experiments we expect that the concentration of [1-$\mu$H]$^+$ is much less than that of [1-$t$H]$^+$. However, bridging hydride catalysis was not further explored here as a recent paper describes their mechanistic details (which is proposed to not regenerate a Fe(I)Fe(I) complex).$^{35-39}$ The overpotential associated with proton reduction with [1-$t$H]$^+$ can be estimated from $E^\circ$(HBF$_4$/H$_2$) = -0.02 V (in CH$_3$CN) as $\sim$1.6 – 1.7 V. The cyclic voltammogram of [1-$t$H]$^+$ is unaffected by the weak acid chloroacetic acid ($pK_a^{CH3CN} = 15.3$), consistent with the large kinetic barrier to protonation by weak acids (Chapter 3).
Scheme 4.1. Proposed mechanism of proton reduction catalysis by $[1$-$tH]^+$. 

**General Redox Properties of Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$ (2) and Protonated Derivatives.** Compound 2 contains a functional proton relay and is capable of protonation by weak acids (see Chapter 3). Thus, we sought to explore the effects of proton relay on proton reduction catalysis. The cyclic voltammograms of 2 exhibits two well-separated one-electron oxidation waves centered near -500 mV vs Fc$^{0/+}$. The reversibility and separation of these two waves is highly dependent on electrolyte and solvent.\textsuperscript{40,41} For CH$_2$Cl$_2$ solutions with [NBu$_4$]BAR$_4^-$ as electrolyte, 2 displays two reversible oxidation couples at $-0.63$ V and $-0.40$ V vs Fc$^{0/+}$ corresponding to the previously studied 2$^{0/+}$ and 2$^{+/-+}$ couples (Figure 4.5).

In CH$_2$Cl$_2$ solutions at 0 °C, the amino-hydride [2-$tH]^+$ displays an *irreversible* reduction at $-1.64$ V vs Fc$^{0/+}$. The pdt derivative [1-$tH]^+$ displays a *reversible* reduction
(\(i_{pa}/i_{pc} = 0.64\)) at nearly the same potential (−1.67 V vs Fc\(^{0/+}\)). As for [1-H]\(^+\), the [2-tH]\(^+0\) couple is milder than the [2-\(\mu\)H]\(^+0\) couple. As invoked for related amino-hydrides\(^{31,42}\) the irreversibility of the [2-tH]\(^+0\) couple is rationalized by an EC mechanism whereby the mixed valence hydride ([2-tH]\(^0\)) undergoes protonation by [2-tH]\(^+\).

**Figure 4.5.** Cyclic voltammogram of a 0.0028 mM CH\(_2\)Cl\(_2\) solution of 1 using 0.1 M [NBu\(_4\)]BAr\(_F^4\) displaying two reversible oxidation events. Reference electrode: Ag|AgCl, but here shown vs Fc\(^{0/+}\). Working electrode: 3 mm diameter glassy carbon. Counter electrode: Pt wire. Graph presented in IUPAC convention.

**Proton Reduction Catalysis by Fe\(_2\)(adt)(CO)\(_2\)(dppv)\(_2\), 2.** In contrast to the propanedithiolate, the voltammetry of [2-tH]\(^+\) is strongly affected by weak acids by virtue of the amine relay site in the dithiolate. Experiments employed [NBu\(_4\)]BAr\(_F^4\) as the electrolyte to maximize the equilibrium concentration of [2-tH]\(^+\) vs the tautomer [2-adtNH\(_2\)]\(^+\) (Chapter 3). Unlike the propanedithiolate derivative 1, addition of the weak acid chloroacetic acid (\(pK_a^{CH_3CN} = 15.3, E^\circ(AN) = -1.05\)) sharply increased the current at
−1.64 V, indicative of catalytic proton reduction (Figure 4.7, 4.8). The azadithiolate also greatly enhances the rate of proton reduction catalysis. A plot of $i_{cl}/i_p$ vs. $[H^+]$ is linear up to $i_{cl}/i_p \sim 95$ (~700 equiv chloroacetic acid). The linearity of the $i_{cl}/i_p$ vs $[H^+]$ indicates that catalysis is second-order with respect to $[H^+]$ (Figure 4.8). Doubling the concentration of 2 doubled the catalytic current response $i_c$ vs $[H^+]$, in agreement with our proposed mechanism confirming that catalysis is first order with respect to $[Fe_2]$. On return scans after catalytic waves, the shape of the oxidation couples for 2 are merged and no-longer independent processes (inset Figure 4.6). This observation is consistent with the previous understanding that these oxidation couples are very sensitive to electrolyte and the catalytic proton reduction wave generates several equivalents ClCH$_2$CO$_2^-$ within the double layer thus altering the appearance of the 2$^{0/+}$ and 2$^{+/-}$ couples.
Figure 4.6. Cyclic voltamagrams of a 0.0028 M solution of 2 (0 °C, 0.1M [NBu$_4$]BAR$_4^-$, CH$_2$Cl$_2$) recorded with increasing (labeled) equivalents of ClCH$_2$CO$_2$H. The inset shows the effect of [ClCH$_2$CO$_2$]$^-$ on the shape of the 2$^{0/+}$ and 2$^{+/-}$ couples. Graph presented in IUPAC convention. The reduction event appears to shift in the negative direction with increasing acid equivalents; however, this is a consequence of a lack of iR compensation, and has been corrected in Figure 4.7.
Figure 4.7. Cyclic voltammogram of $[2\text{-tH}]^+$ in the presence of increasing equivalents of ClCH$_2$CO$_2$H. The current for the one-electron $2^{0/+}$ couple in the absence of acid corresponds to 3.8 $\mu$A (inset). In this CV the shape of the return scan with 1000 equivalents of acid is peculiar, however not reproducible.

*iR Compensation.* As observed in Figure 4.6, the position of the peak maximum ($E_p$) with increasing equivalents of acid appears to drift in the more reducing direction; however the onset is the same. This effect is due to the poor conductance of a CH$_2$Cl$_2$ solution of [NBu$_4$]BAR$_4^-$ or [NBu$_4$]PF$_6$ which leads to unusually high resistance.$^8$ Ohm's law is the relationship of potential with current and resistance (equation 4.1). The potentiostat however is normally not automatically set to adjust for this added solution resistance. Thus, as current increases with increasing equivalents of acid the applied
voltage from the potentiostat is actually much less after adjusting for increasing solution resistance. After turning on the potentiostat’s iR compensation feature CV’s were obtained without the observed drift at high currents (Figure 4.7).

![Graph](image)

**Figure 4.8.** Graph of $i_c/i_p$ vs equivalents of added ClCH₂CO₂H. Catalysis becomes acid-independent at $i_c/i_p = 95$.

$$E = iR \quad \text{(Eq. 4.1)}$$

$E = $ potential  
$i = $ current  
$R = $ resistance

**Estimation of Overpotential.** Neither the $pK_a$’s of relevant acids nor the standard reduction potential of $H^+$ are known in CH₂Cl₂ solution, thus the overpotential can only be estimated.⁵ It is, however, well established that proton reduction is efficiently catalyzed by Pt with low overpotentials. Under conditions used for catalysis by [2-tH]+, we employed a Pt electrode to directly measure the reversible hydrogen couple (Figure 4.9). As the solubility of H₂ in CH₂Cl₂ is low, the reverse scan (oxidation) was
used as a more reliable measure of $E^\circ$, as the concentration of $H_2$ at the electrode would more closely approximate 1 atm. Under our reaction conditions, the Pt couple for $E^\circ_{(HA/H_2)}$ from chloroacetic acid is -1.0 V vs $Fc^{0/+}$, which agrees well with the value in MeCN ($E^\circ_{(HA/H_2)} = -1.05$ V vs $Fc^{0/+}$). The potential at half-height for an acid-independent voltammogram ($E_{cat}$) for [2-tH]$^+$ occurs at -1.45 V vs $Fc^{0/+}$, thus indicating an overpotential of ~ 450 mV.

Figure 4.9. Cyclic voltammogram of a CH$_2$Cl$_2$ solution of 0.07 M ClCH$_2$CO$_2$H with 0.1 M [NBu$_4$]BARF$_4$ electrolyte at 0.1 V/s with platinum working electrode (d = 0.15 mm), Ag|AgCl reference, Pt counter electrode. Graph presented in IUPAC convention.
**Estimation of rate constant.** Conventionally, rate constants for proton reduction are estimated from the plateau region of plots of \(i_c/i_p\) vs [H⁺] (Equation 4.2).\(^{31}\) In this region, the rate-determining step is no longer dependent on [H⁺]. As catalytic runs investigating [2-tH]\(^+\) were performed by the *in-situ* generation from 2, we used the oxidation current from the 2\(^{0/+}\) couple as a measure for the [2-tH]\(^{+/0}\) couple. As both events are one-electron couples, and the concentration of the diiron complex is the same before and after protonation, this method for determining \(i_p\) (peak current in the absence of excess acid) should be accurate. In the presence of ~1,000 equiv of CICH\(_2\)CO\(_2\)H with a scan rate of 100 mV/s the \(i_c/i_p\) vs [H⁺] is no longer affected by [H⁺] and this \(i_c/i_p\) ratio is ~ 95. Using the relationship in equation 4.2,\(^{31}\) the rate constant is approximately **2,000 s\(^{-1}\)**.

\[
\frac{i_c}{i_p} = \frac{n}{0.4463} \sqrt{\frac{RTk}{Fv}}
\]  
**EQ. 4.2**

where:
- \(n = \# \text{ electrons (2)}\)
- \(v = \text{the scan rate (0.1 V/s)}\)
- \(F = \text{Faraday's constant (96485 C/s)}\)
- \(R = \text{Gas constant (8.314 J/mol}\cdot\text{K)}\)
- \(T = \text{Temperature (278 K)}\)
- \(k = \text{rate constant}\)
Figure 4.10. Cyclic Voltammograms of a CH₂Cl₂ solution of 0.0017 M 2 in the presence of 0, 1, 5, 9, 12, 18 equivalents of HBF₄•Et₂O with [NBu₄]BAR₄ as supporting electrolyte. Reference electrode: Ag|AgCl. Working electrode: glassy carbon (d = 3.0 mm). Counter electrode: Pt wire. Graph presented in IUPAC convention.

**Catalytic Reduction of HBF₄•Et₂O by the Azadithiolato Complex.** It was known that complex 2 reacts with multiple equiv of the strong acid HBF₄•Et₂O to undergo double protonation to give the ammonium hydride. Addition of more than two equivalents of HBF₄•Et₂O to CH₂Cl₂ solutions of 2 gave complicated CV’s consisting of multiple catalytic reduction currents (Figure 4.10). In addition to the expected catalytic wave at −1.64 V for the [2-tH]⁺/₀ couple, a new catalytic wave is observed at −1.2 V vs Fc⁰/⁺, ca. 400 mV less negative than for [2-tH]⁺. This process is proposed to be catalysis by the ammonium hydride [2-adtNH₂-tH]²⁺. Other diiron systems incorporating a protonated ligand exhibit similar anodic shifts of around ~400 mV. The origin of this effect is proposed to be coulombic, because rate enhancements are typically not observed.
Catalytic Reduction of HBF₄•Et₂O by [(t-H)Fe₂(odt)(CO)₂(dppv)]⁺, [3-tH]⁺. In Chapter 3, the oxadithiolate ligand in 3 was shown to function poorly as a proton relay. We sought therefore to explore the effect of the oxadithiolate ligand on proton reduction catalysis. We found that, unlike the azadithiolate 2, cyclic voltammograms of 3 were unaffected by addition of weak acids. Upon addition of HBF₄•Et₂O, however, the [3-tH]⁺/0 reduction event increased proportional to the equivalents of added acid (Figure 4.11). A plot of $i_c/i_p$ vs [HBF₄•Et₂O] is linear indicating a catalytic mechanism second-order with respect to [H⁺] (Figure 4.12). Saturation kinetics, where $i_c/i_p$ is no longer dependent on [H⁺], was not achieved with [3-tH]⁺ because of the large catalytic response to the presence of small amounts of [3-µH]⁺ (Figure 4.12). As with the propanedithiolate 1, the expected concentration of [3-µH]⁺ at −40 °C is small; thus the large catalytic wave associated with [3-µH]⁺ strongly indicates that [3-µH]⁺ is a faster catalyst than [3-tH]⁺. The overpotential associated with [3-tH]⁺ is very similar to that of [1-tH]⁺, being ca. 1.6 – 1.7 V. Under identical experimental conditions, the direct comparison of [1-tH]⁺ and [3-tH]⁺ shows that [3-tH]⁺ is slightly more responsive within the acid-dependent region (Figure 4.12). As the rate-limiting step in this acid-dependent regime involves [H⁺] we conclude that 3 is slightly faster to protonate than 1. However, both 1 and 3 suffer from high overpotentials and slow catalytic rates.
Figure 4.11. Cyclic voltammogram (-40 °C, 1 mM catalyst, ~1 mM ferrocene) for [3-tH]BF₄ with increasing equivalents of HBF₄•Et₂O recorded at 50 mV/s. During return scans, the presence of trace amounts of unprotonated 3 is seen at ~600 mV. The event at ~1.8 V is attributed to catalysis by [3-tH]BF₄, although the concentration of this species at –40 °C is expected to be very small.
Figure 4.12. Plot of $i_{C}/i_{D}$ vs $[\text{H}^+]$ for both $[3-\text{tH}]^+$ and $[1-\text{tH}]^+$ under identical experimental conditions. (Conditions: see Figure 4.10).

Thermodynamics and Mechanism of Proton Reduction – Implications for Hydrogen Oxidation Catalysis. DuBois, Norton, and others have demonstrated that thermodynamic parameters for transition-metal hydrides have been powerful tools for predicting their reactivity such as hydride insertion, alkene hydrogenation, and hydrogen oxidation. The free energy relationships for catalysis by $[2-\text{tH}]^+$ are depicted in Scheme 4.2. The free energy relationships have been calculated in reference to the standard reduction potential of chloroacetic acid ($E^0$), and as such the overall $\Delta G$ is zero. The first step in hydrogen evolution catalysis rotation of the Fe(I)Fe(I) complex. Protonation then occurs forming the terminal hydride $[2-\text{tH}]^+$. The subsequent step is the reduction of the terminal hydride species, forming a reduced terminal hydride. The $\Delta G$
associated with this step is determined by an electrochemical measurement ($\Delta G = 23.06x E_{1/2} - E^\circ$). This step is the largest thermodynamic step of the catalytic reaction, but is not however the rate-limiting step as the potential is provided by the electrode. Following this initial electron-transfer step, the proposed subsequent chemical steps are protonation of the reduced species to provide a $\eta^2$-H$_2$ complex, and loss of this H$_2$ ligand to provide the rotated mixed-valence Fe(II)Fe(I) complex. Reduction to regenerate the unrotated Fe(I)Fe(I) complex completes the catalytic cycle. Interestingly, the free-energy of formation of the rotated Fe(I)Fe(I) species must be taken into account to achieve an energy-neutral diagram. Additionally, the return “unrotation” energy is accounted for in the $[2]^{+/0}$ redox couple. As the overall energy of the reaction is known, we can determine the $\Delta G$ associated with the two unknown steps (protonation of the reduced $[2-tH]^0$, elimination of H$_2$ reforming $[2]^+$) as $\sim -8$ kcal/mol. This thermodynamic analysis provides a detailed picture of the complete catalytic process, and can be used to interpret the reverse reaction (Scheme 4.2). Although hydrogen oxidation catalysis has not yet been achieved experimentally, a reaction with hydrogen has been achieved from the mixed-valence $2^+$, albeit only at high pressures.
Scheme 4.2. Free-energy relationships along the steps for the proposed mechanism of proton reduction from ClCH₂CO₂H by [2-tH]⁺.
Discussion

Role of Azadithiolate in Proton Reduction Catalysis. Incorporation of an amine into the dithiolate, e.g. in [2-tH]⁺, enables reduction of weak acids. Isostructural complexes incorporating propanedithiolate and oxadithiolate ([1-tH]⁺ and [3-tH]⁺, respectively) are poor catalysts because of the significant kinetic barrier to proton transfer (chapter 3). Consequently, proton reduction catalysis by [1-tH]⁺ and [3-tH]⁺ is only possible with very strong acids, resulting in very large overpotentials. The azadithiolate [2-tH]⁺ however, operates at remarkably lower overpotentials, because protons from weak acids are efficiently relayed to the terminal position. Direct comparison of acid independent rate constants highlights the effective combination of terminal hydrides with azadithiolates (Table 4.2).

Table 4.2. Comparison of catalytic properties of terminal and bridging hydrides of [1-H]⁺ and [2-H]⁺.

<table>
<thead>
<tr>
<th>Complex</th>
<th>$E_{\text{overpot. (acid)}}$</th>
<th>rate (s⁻¹)</th>
</tr>
</thead>
<tbody>
<tr>
<td>[([t-H]Fe₂(pdt)(CO)₂(dppv)₂]⁺</td>
<td>~ 1.6 V (HBF₄)</td>
<td>&lt; 20</td>
</tr>
<tr>
<td>([µ-H]Fe₂(pdt)(CO)₂(dppv)₂]⁺</td>
<td>~ 0.65 V (ClCH₂CO₂H)</td>
<td>~ 20</td>
</tr>
<tr>
<td>([t-H]Fe₂(adt)(CO)₂(dppv)₂]⁺</td>
<td>~ 0.45 V (ClCH₂CO₂H)</td>
<td>~ 2,000</td>
</tr>
<tr>
<td>([µ-H]Fe₂(adt)(CO)₂(dppv)₂]⁺</td>
<td>~ 0.65 V (ClCH₂CO₂H)</td>
<td>~ 20</td>
</tr>
</tbody>
</table>
Without the incorporation of an effective proton relay, catalysis from bridging hydride complexes have significantly lower overpotentials than terminal hydride complexes. However, they function in a non-biomimetic manner. To expand our understanding of Nature, we have focused on the differences between the aza-, oxa-, and propanedithiolate with the biologically relevant terminal hydride. The differing catalytic properties are largely associated with the high kinetic barrier to protonation from weak acids for 1 and 3. Typically graphs of \( \frac{I_c}{I_p} \) vs [H\(^+\)] contain an acid-independent regime wherein the rate-limiting step is not dependent on [H\(^+\)]. In this regime, the rate-limiting step is typically proposed to be H\(_2\)-elimination or electron transfer at the electrode surface. For [2-tH\(^+\)] we favor a rate-limiting step that involves H\(_2\) dissociation from the mixed-valence [2-tH\(_2\)]\(^+\), although we have no spectroscopic support for this conclusion. As both 1 and 2 have similar electronic profiles (\( E_{1/2} \)'s, pK\(_a\)'s) it would be logical to assume they share a common mechanism. However, if H\(_2\) dissociation is the rate-limiting step for the propanedithiolate [1-tH\(^+\)], this implies that H\(_2\) binds much stronger to [1]\(^+\) than to [2]\(^+\), which is not consistent with the similar electronic profiles. Thus, we propose that the rate-limiting step for [1-tH\(^+\)] is the protonation of 1 to form [1-tH\(^+\)], and that the mechanism for 1 involves a pre-equilibrium with a rotated excited state (Scheme 4.3).
Scheme 4.3. The proposed mechanism for protonation of 1, at high acid concentration the equilibrium between 1 and 1* is rate-limiting.

Table 4.3. Various hydrogen processing catalysts and their overpotential and estimated turnover frequencies.

<table>
<thead>
<tr>
<th>Catalyst</th>
<th>Acid Used (\text{pK}_a^{\text{CH}_3\text{CN}})</th>
<th>Overpotential (V)</th>
<th>Est. Rate (\text{s}^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>(\text{Fe}_2\text{(pdt)(CO)}_4\text{(PMe}_3)_2)</td>
<td>22.3 (HOAc)</td>
<td>0.79</td>
<td>(&lt; 10^{(14)})</td>
</tr>
<tr>
<td>([\text{Fe}_2\text{(pdt)(CN)(CO)}_4\text{(PMe}_3)_2]^+)</td>
<td>8.7 (TsOH)</td>
<td>0.45</td>
<td>(&lt; 10^{(14)})</td>
</tr>
<tr>
<td>(\text{Fe}_2\text{(bdt)(CO)}_6)</td>
<td>8.7 (TsOH)</td>
<td>0.6</td>
<td>(&lt; 10^{(55)})</td>
</tr>
<tr>
<td>([\text{Fe}_2\text{(pdt)(NO)(CO)}_3\text{(dppv)}]^+)</td>
<td>12.5 (TFA)(^a)</td>
<td>0.3</td>
<td>(&lt; 5^{(56)})</td>
</tr>
<tr>
<td>([\text{Fe}_4\text{S}_6]^{2-})</td>
<td>14.1 (LutH(^+))</td>
<td>0.38</td>
<td>900 (^{(58)})</td>
</tr>
<tr>
<td>([((t\text{-H})\text{Fe}_2\text{(adt)(CO)}_2\text{(dppv)}_2]^+)</td>
<td>15.3 (ClAA)(^a)</td>
<td>0.45</td>
<td>2,000 (^{(b)})</td>
</tr>
<tr>
<td>([\text{Ni}(\text{P}^\text{Ph}_2\text{N}^\text{Ph}_2)_2(\text{CH}_3\text{CN})]^{2+})</td>
<td>-4; (HOTf)</td>
<td>0.41</td>
<td>130 (^{(31)})</td>
</tr>
<tr>
<td>(\text{Co}(\text{dmgBF}_2)_2(\text{CH}_3\text{CN})_2)</td>
<td>8.7 (TsOH)(^a)</td>
<td>0.04</td>
<td>(~ 1 \times 10^4 \text{M}^{-1} \text{s}^{-1}) (^{(59)})</td>
</tr>
</tbody>
</table>

\(^a\) TFA is CF\(_3\)CO\(_2\)H, ClAA is ClCH\(_2\)CO\(_2\)H,
\(^b\) This work
Comparison to Other Synthetic H₂-Production Catalysts. The [FeFe]-
hydrogenases operate at low overpotentials (< 50 mV) and are fast (~6,000 turnovers
per second) and as such rival the activity of platinum. For this reason, diiron dithiolates
have sparked immense interest as synthetic catalysts. However, to date only poor
catalysts have been designed. To compare the terminal hydride azadithiolate, [1-tH]⁺,
we have selected “good” catalysts (Table 4.3). For diiron dithiolates, the maximum
turnover frequencies are all less than 20 s⁻¹, while overpotentials are typically 0.5 V or
greater. A rare example of a hydrogenase mimic that is thermodynamically efficient (~
0.38 V overpotential, ~ 900 s⁻¹) is Pickett et al.’s 4Fe6S species. This 4Fe6S species is
rotated and suspected to involve a terminal hydride.⁵⁷,⁵⁸ DuBois et al. have incorporated
proton relays with mononuclear nickel hydrides to achieve some of the best-reported
homogeneous catalysts to date (~ 0.4 V overpotential, 130 s⁻¹).³¹ Some cobaloximes
have been reported to have remarkably low overpotentials (~40 mV) and high rates,
although the exact mechanism is thought to involve bimolecular pathways.⁵⁹

Pre-Rotation in Diiron Dithiolates. In the enzyme, the H₉red state is either a
rotated Fe(I)Fe(I) species or a stable terminal hydride. These possibilities cannot be
distinguished crystallographically. However, DFT studies have indicated that a terminal
hydrides are invariably unstable with respect to bridging hydrides.⁶⁰ Thus, we favor H₉red
being a rotated Fe(I)Fe(I) species, containing an ammonium azadithiolate. We have
shown in Chapter 3 that non-rotated diiron(I) ammonium species exists in dynamic
equilibrium with the hydrido diiron(II) amine. In the model compound, [2-tH]⁺, a
Fe(CO)(dppv) subunit must rotate prior to the relay from the ammonium form to the
terminal hydride position (Scheme 4.4). This rotational energy has previously been
estimated as \(~9\text{ kcal/mol (400 mV)}.\) If this rotational barrier were removed we suspect that the \([\text{Fe}_2(\text{xdt})(\text{CO})_2(dppv)_2]^{0/+}\) couple would shift anodically by \(~400\text{ mV}.\) Indeed, the difference of \(E_{1/2}\) between our model complex and the enzyme is \(~400\text{ mV (Chapter 1).}\) As the catalytic mechanism for model complexes proceeds through this barrier we propose it could be related to the overpotential for catalysis (\(~450\text{ mV).}\) Future models will need to address this overpotential by designing complexes that alleviate the need for rotation. Interestingly, if such complexes were “pre-rotated” little reorganization would occur during proton transfer, and as such there would be no need for a proton relay to aide in protonation reactions. However, the amine relay could still play a vital role in \(\text{H}_2\) oxidation- where \(\text{H}_2\) must transiently bind and then be quickly deprotonated by a base.\(^{61}\)

\[
\text{Scheme 4.4. The rotational barrier within the catalytic cycle of synthetic diiron dithiolates.}
\]

In conclusion, the incorporation of azadithiolate into diiron dithiolate terminal hydride complexes has resulted in an effective proton relay. When this system is examined catalytically, turnover frequencies are remarkably similar to that of the enzyme (\(~2,000 \text{ s}^{-1}\)). However, the diiron dithiolate model complexes employed retain a significant flaw when compared to Nature as they do not have rotated \(\text{Fe(I)Fe(I)}\)
structures. This rotational barrier is proposed to be the source of the observed ~400 mV overpotential. Future efforts will need to eliminate or reduce this rotational barrier to achieve a fully functional synthetic hydrogenase.

**Experimental Procedures**

All manipulations were conducted using standard Schlenk techniques. Solvents were filtered through activated alumina and subsequently degassed. $^1$H and $^{31}$P NMR spectra were acquired on a Unity Varian 500 or a Unity Varian 600 spectrometer. IR spectra were collected on a Mattson Infinity Gold FTIR spectrometer. *Cis*-1,2-bis(diphenylphosphino)ethylene (dppv) and HBF$_4$$\cdot$Et$_2$O solution were purchased from Aldrich. Fe$_2$(S$_2$C$_3$H$_6$)(CO)$_2$(dppv)$_2$ (1) and [H(Et$_2$O)$_2$]BAr$_4^F$ were prepared according to literature procedures ($\text{BAr}_4^F- = [\text{B(C}_6\text{H}_3-3,5-(\text{CF}_3)_2)_4]^{-}$).

**Preparation of [HPPPh$_3$]BAr$_4^F$.** A solution of [H(Et$_2$O)$_2$][BAr$_4^F$] (0.385 g, 0.377 mmol) in Et$_2$O (10 mL) at -40 °C was transferred via cannula into a solution of PPh$_3$ (0.097 g, 0.370 mmol) in Et$_2$O (10 mL) at -40 °C. Solvent was removed under vacuum, leaving a white solid. Yield: 0.350 g (83%). $^{31}$P($^1$H) NMR (202 MHz, CD$_2$Cl$_2$, 20 °C): $\delta$ 7.0 (s).

**Electrochemistry.** Cyclic voltammetry experiments were carried out in either a 9-mL or 20-mL one-compartment glass cell. The working electrode was a glassy carbon disk (3.0 mm in diameter). The reference electrode for experiments conducted less than 0 °C was a pseudo-reference silver wire, for experiments greater than 0 °C a Ag|AgCl electrode (ca. -0.420 V vs Fc/Fc$^+$) was employed. The counter electrode was a Pt wire. The electrolyte was 0.1 M [NBu$_4$]PF$_6$ in CH$_2$Cl$_2$ or [NBu$_4$]BAr$_4^F$ in CH$_2$Cl$_2$ when
specified. The concentration of the organometallic complex was 1 mM. For low
temperature cyclic voltammetry, a thermocouple was inserted into the CV cell
periodically to record the temperature. We routinely observed ~10 °C differences
between the solution temperature and the surrounding cold bath.

Cyclic Voltammetry for [(t-H)Fe$_2$(pdt)(CO)$_2$(dppv)$_2$]BF$_4$, [2-tH]BF$_4$. A solution
containing 1:1 ratio of 1 and [1-tH]BF$_4$ was generated by addition of 0.5 equiv of
[H(ET$_2$O)$_2$]BAR$_4$ to 1 at 0 °C. The resulting solution was analyzed with cyclic
voltammetry at various scan rates and a Cottrell graph of $i_p$ versus square-root of scan
rate for 1 and [1-tH]$^+$ was constructed. Both 1 and [1-tH]$^+$ share similar diffusion
coefficients indicated by the similar their slopes.

[(t-H)Fe$_2$(pdt)(CO)$_2$(dppv)$_2$]BF$_4$, [1-tH]BF$_4$ and Proton Reduction by Cyclic
Voltammetry. A solution of Fe$_2$(pdt)(CO)$_2$(dppv)$_2$ (1) (7.5 mg, 0.007 mmol) in 6 mL
CH$_2$Cl$_2$ was treated with aliquots of a 0.0691 M HBF$_4$$^•$2Et$_2$O CH$_2$Cl$_2$ solution (100 µL,
0.07 mmol) at 0 °C. Cyclic voltammograms were collected at 500 mV/s.

Protonation of Reduced [(t-H)Fe$_2$(pdt)(CO)$_2$(dppv)$_2$] ([1-tH]$^0$) by Cyclic
Voltammetry. A solution of Fe$_2$(pdt)(CO)$_2$(dppv)$_2$ (1) (7.5 mg, 0.007 mmol) in 6 mL
CH$_2$Cl$_2$ was treated with less than one equiv 0.0691 M HBF$_4$$^•$2Et$_2$O CH$_2$Cl$_2$ solution
(~75 µL, 0.005 mmol) at -40 °C. This solution was then treated with 0.07 M [HPPh$_3$]BF$_4$
CH$_2$Cl$_2$ solution (100 µL, 0.007 mmol). Cyclic voltammograms were collected at 200
mV/s using a silver wire as the pseudo-reference electrode.

Proton Reduction Catalysis Cyclic Voltammetry for [(t-
H)Fe$_2$(odt)(CO)$_2$(dppv)$_2$]BF$_4$, [3-tH]BF$_4$. A solution of Fe$_2$(odt)(CO)$_2$(dppv)$_2$ (3) (7.5 mg,
0.007 mmol) in 6 mL CH$_2$Cl$_2$ was treated with aliquots of a 0.691 M HBF$_4$$^•$2Et$_2$O CH$_2$Cl$_2$
solution (10 µL, 0.07 mmol) at -40 °C. Cyclic voltammograms were collected at 50 mV/s.

Proton Reduction Catalysis Cyclic Voltammetry for [HFe$_2$(adt,NH)(CO)$_2$(dppv)$_2$]BF$_4$, [2-tH]BF$_4$. A solution of Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$ (2) (7.5 mg, 0.007 mmol) in 3 mL CH$_2$Cl$_2$ was treated at 0 °C with aliquots of a 0.0691 M HBF$_4$•Et$_2$O solution (100 µL, 0.07 mmol). Cyclic voltammograms were collected at 100 mV/s.

Proton Reduction Catalysis Cyclic Voltammetry for [HFe$_2$(adt,NH)(CO)$_2$(dppv)$_2$]BF$_4$, [2-tH]BF$_4$. A solution of Fe$_2$(adt,NH)(CO)$_2$(dppv)$_2$ (2) (3.1 mg, 0.047 mmol) in 4 mL CH$_2$Cl$_2$ was treated at 0 °C with aliquots of a 0.0691 M ClCH$_2$CO$_2$H CH$_2$Cl$_2$ solution (100 µL, 0.07 mmol). Cyclic voltammograms were collected at 100 mV/s. The iR compensate feature was turned on and the resistance of the solution was recorded and automatically adjusted prior to each scan.

References:


38. Adam, F. I.; Hogarth, G.; Kabir, S. E.; Richards, I. "Models of the iron-only hydrogenase: Synthesis and protonation of bridge and chelate complexes [Fe$_2$(CO)$_4$(Ph$_2$P(CH$_2$)$_n$PPh$_2$)(µ-pdt)] (n =2-4) - evidence for a terminal hydride intermediate" C. R. Chim. 2008, 11, 890-905.


Chapter 5

The Very Electron-Rich Disubferrous Dithiolate: Fe₂(pdt)(CNMe)₆

Introduction

Although there are several hundred model complexes for the active site of [FeFe]-hydrogenase,¹ most are hexacarbonyls and poorly mimic the electronic environment of the “H_{red}” state (Table 1.1, Figure 1.1).² The closest match by IR spectroscopy is the tetra-substituted Fe₂(adt)(CO)₂(dppv)₂.³,⁴ However, by comparison to the H_{red}/H_{ox} couple (a Fe₂^{I,1}/Fe₂^{I,II} couple), the [Fe₂(adt)(CO)₂(dppv)₂]^{0/+} couple is easier to oxidize by 200 mV.

To minimize this redox disparity, we have explored the utility of isocyanides, which are isoelectronic to CO yet are much weaker π-acceptors leading to dramatic changes in oxidation and protonation chemistry.⁵ For example, the homoleptic Fe(CN₆Bu)₅, unlike Fe(CO)₅, is susceptible to protonation with weak acids and binds diphenylacetylene.⁶-⁹ Isocyanides are potentially relevant to the [FeFe]-hydrogenases, because the cyanide ligands employed in the H-cluster are hydrogen-bonded to amino acid residues,¹⁰ which in turn makes them similar to M-CNR groups (Figure 5.1).¹¹ Thus, isocyanide ligands are likely more biomimetic than phosphine ligands, which typically have large steric profiles that can hinder proton transfer reactions.¹²
Isocyanide ligands exhibit a wealth of reactivity. When bonded to very electron-rich metal centers, the isocyanide ligand is typically bent. In Ru(CN\text{Bu})\textsubscript{5} the C-N-C angles are 130°.\textsuperscript{8} Protonation of such electron-rich isocyanides often affords an aminocarbene (M-CN(H)R).\textsuperscript{13,14} However, protonation can also occur at the metal, as is the case for the reaction of Fe(CN\text{Bu})\textsubscript{5} with HBF\textsubscript{4}•Et\textsubscript{2}O providing [HFe(CN\text{Bu})\textsubscript{5}]\textsuperscript{+} (Scheme 5.1).\textsuperscript{15,16} Some hydrides have been observed to insert into the metal isocyanide bond providing the corresponding iminoformyl (M-C(H)=NR).\textsuperscript{17} Some isocyanide complexes can undergo multiple protonation to the corresponding aminoformyl (M\textsuperscript{n}C(H)N(H)R). Recently, the aminoformyl complex [Ni(triphos)(C(H)N(H)xylyl)]\textsuperscript{2+} has been demonstrated to be perform the hydrogenation of ketones.\textsuperscript{16}
Scheme 5.1. Typical protonation reactions of metal isocyanide complexes, and their nomenclature.

Previous work in our laboratory has shown that the tetra-substituted $\text{Fe}_2(\text{pdt})(\text{CO})_2(\text{CNMe})_4$ reacts with $\text{CF}_3\text{CO}_2\text{H}$ at $-80 \, ^\circ\text{C}$ to form what is proposed to be $[\text{Fe}_2(\text{pdt})(\text{CO})_2(\text{CH=NMMe})(\text{CNMe})_3]^+$. This iminoformyl later isomerizes to provide several isomers of a bridging hydride $[(\mu\text{-H})\text{Fe}_2(\text{pdt})(\text{CO})_2(\text{CNMe})_4]^+$. The iminoformyl ligand is likely generated by hydride insertion into the $\alpha$-carbon of an isocyanide ligand, thus supporting previous results of kinetic terminal protonation for diiron dithiolates.
Results:

Treatment of Fe₂(pdt)(CO)₆ with excess methyl isocyanide in refluxing toluene for two days was found to result in the formation of the very air-sensitive red compound, Fe₂(pdt)(CNMe)₆, 1. Recrystallization of 1 from benzene provided red X-ray quality crystals of Fe₂(pdt)(CNMe)₆•2(C₆H₆) (Figure 5.2). The Fe(1)-Fe(2) distance of 2.540 Å compares well to other Fe₂(pdt)(PR₃)ₓ(CO)₆₋ₓ species.²⁰ Interestingly, the isocyanides show different degrees of bending, four of which are ~175-180° while C(6)-N(3)-C(12) and C(5)-N(2)-C(11) are bent and have C-N-C angles of 166°. The nature of this bending suggests strong π-backbonding from the metal center into the carbon-nitrogen σ* orbital (Scheme 5.1).

Figure 5.2. Structure of Fe₂(pdt)(CNMe)₆, 1, thermal ellipsoids drawn at 35 %, hydrogen atoms and co-crystallized benzene are emitted for clarity.
Consistent with rapid dithiolate “flipping” as observed for the starting Fe$_2$(pdt)(CO)$_6$,$^{21}$ the $^1$H NMR spectrum of 1 shows just two signals for the pdt ligand. Additionally, only one signal for the methyl resonance of the isocyanide ligands was observed, indicating rapid turnstile rotation of the Fe(CNMe)$_3$ subunits (Figure 5.3).$^{21}$

The $^{13}$C{$^1$H} NMR (CD$_2$Cl$_2$, 150 MHz) spectrum is also consistent with six rapidly converting coordinated CNMe ligands ($\delta$ 182, Fe-CNMe; $\delta$ 30.2 Fe-CNMe) (Figure 5.4).

The FT-IR spectrum (CH$_2$Cl$_2$) of 1 displays several metal isocyanide bands (2000-2150 cm$^{-1}$), and no $\nu$CO bands (Figure 5.5).

Figure 5.3. $^1$H NMR (600 MHz, CD$_2$Cl$_2$) spectrum of 1•2(C$_6$H$_6$) at ambient temperature showing the CNMe resonance ($\delta$ 3.35) and two pdt resonances ($\delta$ 1.72, 1.57).
**Figure 5.4.** $^{13}$C NMR (150.86 MHz, CD$_2$Cl$_2$) spectrum of 1 recorded at ambient temperature displaying a Fe-CNMe signal ($\delta$ 182) and two pdt signals ($\delta$ 30, 24). The signal at $\delta$ 127 is benzene, while the signal at $\delta$ 54 is CD$_2$Cl$_2$. 
Figure 5.5. FT-IR spectrum (CH$_2$Cl$_2$, room temperature) of 1 (bottom), and after addition of [HNEt$_3$]BF$_4$ providing 2 (top).

As CNMe is a weaker $\pi$-acceptor than CO,$^5$ the reactivity of 1 is greatly modified in comparison to Fe$_2$(pdt)(CO)$_6$. The all-carbonyl complex is stable in solution in air indefinitely, whereas Fe$_2$(pdt)(CNMe)$_6$ (1) is very air-sensitive oxidized by trace amounts of air to a green complex, presumably [Fe$_2$(pdt)(CNMe)$_7$]$^{2+}$.$^18$

The Iminoformyl [Fe$_2$(pdt)(CH=NMe)(CNMe)$_3$]$^+$. Addition of a single equivalent or excess [HNEt$_3$]BF$_4$ to a CD$_2$Cl$_2$ solution of 1 resulted in the immediate formation of an extremely air-sensitive blue solution. This species, 2, displays a new resonance in the $^1$H NMR spectrum at $\delta$ 9.5, as well as four CNMe resonances (Figure 5.6). The $^{13}$C($^1$H} NMR displays three CNMe signals ($\delta$ 169, 162.7, 157; 2:2:1 ratio) and four CNMe signals ($\delta$ 31.6, 31.3, 31.0, 30.7; 2:1:2:1 ratio) (Figure 5.7). The $\nu$CN stretching frequencies in the FT-IR spectrum are shifted to higher frequencies than those for 1
(Figure 5.5). Literature precedents suggests that the $^1$H NMR signal at $\delta$ 9.5 arises from an iminoformyl species, and as such we propose 2 is $[\text{Fe}_2(pdt)(\text{CHNMe})(\text{CNMe})_3]^+$, which is consistent with the $^{13}$C NMR and FT-IR.

Complex 2 is a diferrous species and electronically unsaturated as it has only five 2e$^-$ donor ligands, and one “X” ligand (CH=NMe). Other iminoformyl species, such as Tp*(CO)$_2$Mo($\eta^2$-C(NBu$t^1$)Me) (where Tp* = hydridotris(3,5-dimethylpyrazolyl)borate) have been shown to coordinate through the nitrogen atom via an $\eta^2$-\alpha-iminoacyl.$^{22}$ An $\eta^2$-\alpha-iminoacyl ligand is also consistent with the observed spectroscopy for 2 and would satisfy the 18 e$^-$ rule (Scheme 5.1).$^{13}$ Despite repeated attempts to isolate 2, we were unsuccessful due to its extreme air-sensitivity.
Figure 5.6. $^1$H NMR (600 MHz, CD$_2$Cl$_2$, -70 °C) spectrum of 2 showing the iminoformyl resonance ($\delta$ 9.27), four CNMe signals ($\delta$ 3.7, 3.5, 3.3, 3.3), and two pdt signals ($\delta$ 1.7, 2.1). Other signals are assigned to benzene ($\delta$ 7.3), NEt$_3$, ($\delta$ 2.4, 0.9), and CDHCl$_2$ ($\delta$ 5.3).
Figure 5.7. $^{13}$C NMR spectrum (150 MHz, CD$_2$Cl$_2$, -20 °C) of 2 showing three separate 2:1:2 CNMe signals, and four CNMe signals at δ 31. The imino carbon (g) was not observed.

Double protonation eliminating H$_2$ and generating [Fe$_2$(pdt)(CNMe)$_7$]$^{2+}$:

Treatment of CD$_2$Cl$_2$ solutions of 1 with two equivalents [H(Et$_2$O)$_2$]BAR$_4^-$ or HBF$_4$•Et$_2$O liberated H$_2$, as detected by $^1$H NMR spectroscopy. When conducted at low temperature, the $^1$H NMR spectrum shows the formation of H$_2$ at −75 °C, suggesting a facile process (Figure 5.8). After treatment with [H(Et$_2$O)$_2$]BAR$_4^-$, the only $^1$H NMR signals observed were that of [H(Et$_2$O)$_2$]BAR$_4^-$, benzene, and CDHCl$_2$. Thus, we could not identify the organometallic product of the reaction; however, from this solution we crystallized the previously reported [Fe$_2$(pdt)(CNMe)$_7$]$^{2+}$ (Figure 5.9). As samples are free from excess CNMe, we propose that the immediate product of the reaction is not stable and allows the formation of [Fe$_2$(pdt)(CNMe)$_7$]$^{2+}$. 
Figure 5.8. $^1$H NMR (600 MHz, CD$_2$Cl$_2$, -40 °C) of 1 + 2 equiv [H(Et$_2$O)$_2$]BAR$_4^F$, inset shows expanded region of H$_2$ signal ($\delta$ 4.6). All other peaks are assigned to CDHCl$_2$ ($\delta$ 5.3), [H(Et$_2$O)$_2$]BAR$_4^F$ (BAR$_4^F$; $\delta$ 7.8, 7.6, Et$_2$O; $\delta$ 3.5, 1.2), or benzene ($\delta$ 7.3).
Figure 5.9. Structure of the dication \([\text{Fe}_2(\text{pdt})(\text{CNMe})_7]^2^+\), formed by treatment of 1 with excess \([\text{H(Et}_2\text{O})_2]\text{BAr}^\text{F}_4\). Thermal ellipsoids drawn at 50%. Reprinted from ref # 18.

Conclusions:

The complete substitution of \(\text{Fe}_2(\text{pdt})(\text{CO})_6\) with \(\text{CNMe}\) has provided the homoleptic diiron dithiolate \(\text{Fe}_2(\text{pdt})(\text{CNMe})_6\), 1. Complex 1 is significantly more basic than \(\text{Fe}_2(\text{pdt})(\text{CO})_6\) and exposure to trace oxygen results in immediate oxidation. Similar to \(\text{Fe}_2(\text{edt})(\text{CO})_2(\text{CNMe})_4\), we propose protonation of 1 forms the iminoformyl \([\text{Fe}_2(\text{pdt})(\text{CH=NMe})(\text{CNMe})_3]^+\) (2).\(^\text{19}\) Further equivalents of acid result in the elimination of \(\text{H}_2\) and generation of a diferrous species which degrades to the known \([\text{Fe}_2(\text{pdt})(\text{CNMe})_7]^2^+\). This work represents the first model complex that protonates twice to liberate \(\text{H}_2\). Complex 1 is likely significantly more basic than the active site of \(\text{[FeFe]}\)-hydrogenase. As such, this reactivity may not represent a biological pathway.
Isocyanides are possibly related to the [FeFe]-hydrogenases, which contain hydrogen-bonding residues near the cyanide ligands of the active site. Whereas \( \text{Fe}_2(\text{pdt})(\text{CO})_6 \) can only accommodate two mono-dentate phosphines (i.e. \( \text{Fe}_2(\text{xdt})(\text{CO})_4(\text{PR}_3)_2 \)), the sterically small \( \text{CNMe} \) ligand allows further substitution. The results present in this chapter suggest that complexes of the type \( \text{Fe}_2(\text{xdt})(\text{CO})_3(\text{L})(\text{CNMe})_2 \) (where \( \text{L} = \text{PR}_3 \) or other \( 2e^- \) donor) could be synthesized to more accurately model the coordination environment of the active site.

**Experimental**

**Preparation of \( \text{Fe}_2(\text{pdt})(\text{CNMe})_6 \).** To a 250-mL Schlenk flask 2.00 g (5.18 mmol) \( \text{Fe}_2(\text{pdt})(\text{CO})_6 \) was dissolved in 50 mL of degassed toluene. The solution was then frozen and 2.4 mL of CNMe (1.8 g, 44 mmol, 8.5 equivs) was distilled onto the frozen solution. The flask was then placed under argon, fitted with a reflux condenser, and heated to reflux for 2 days, upon which the FT-IR spectrum showed no \( \nu_{\text{CO}} \) stretching bands. Solvent was then removed, and the red solid was redissolved in 130 mL of benzene, heated to reflux, and then cooled to \( \sim 5^\circ \text{C} \) for 30 min. The red crystalline product was isolated by cannula filtration and washed with 2 x 50 mL of hexane. Yield: 2.0 g (4.31 mmol, 83 % yield). \(^1\text{H} \) NMR (600 MHz, CD\(_2\)Cl\(_2\)): \( \delta \) 3.35 (18H, CNMe), 1.72, 1.57 (4H, 2H, pdt). \(^{13}\text{C}\{^1\text{H}\} \) NMR (150.86 MHz, CD\(_2\)Cl\(_2\)): \( \delta \) 182.55 (CNMe), 30.22 (CNMe), 24.1 (pdt) FT-IR (CH\(_2\)Cl\(_2\)): \( \nu_{\text{CNMe}} \) 2200 (w), 2142 (m), 2092 (s), 2065 (s), 1953 (w) cm\(^{-1}\). Anal. Calcd for \( \text{C}_{27}\text{H}_{24}\text{N}_8\text{Fe}_2\text{S}_2\cdot1.4(\text{C}_6\text{H}_6) \) (found): \( \text{C} \), 49.00 (49.16); \( \text{H} \), 5.69 (5.48); \( \text{N} \), 14.65 (13.70).
Protonation of Fe₂(pdt)(CNMe)_6 with [HNEt₃]BF₄. In a 50-mL Schlenk flask specifically designed for in-situ IR measurements, 0.044 g Fe₂(pdt)(CNMe)_6 (0.090 mmol) and 0.017 g [HNEt₃]BF₄ (0.090 mmol) were combined in the glovebox. The flask was then connected to the React-IR and 4-5 mL dry degassed CH₂Cl₂ was added. A similar reaction was preformed in a J. Young NMR tube for NMR analysis. ¹H NMR (600 MHz, CD₂Cl₂): δ 9.27 (1H, C(H)NMe), 3.72, 3.51, 3.31 (3H, 6H, 6H, CNMe), 2.12, 1.70 (2H, 4H, pdt). ¹³C{¹H} NMR (150.86 MHz, CD₂Cl₂): δ 169, 162.7, 157 (2:1:1 CNMe), 31.6, 31.26, 31.0 30.75 (2:1:2:1 CNMe). FT-IR (CH₂Cl₂): υ_CNMe 2231 (w, sh), 2196 (s), 2165 (m), 2138 (m), 2095 (w) cm⁻¹.

Protonation of Fe₂(pdt)(CNMe)_6 with 2 equivalents [H(Et₂O)]BAr⁴⁺. In a J. Young NMR tube 0.006 g Fe₂(pdt)(CNMe)_6 (0.0147 mmol) and 41 mg [H(Et₂O)]BAr⁴⁺ (0.040 mmol) were combined and frozen CD₂Cl₂ was then distilled above the sample and then melted onto sample by placing the J. Young tube into a −78 °C (acetone/dry ice) bath. The resulting solution was then monitored by low-temperature NMR spectroscopies. ¹H NMR (600 MHz, CD₂Cl₂): δ 3.35 (18H, CNMe), 4.6 (H₂) all other peaks to small to resolve near large Et₂O signals of acid.

References:


The [NiFe]-hydrogenases are Nature’s most prevalent hydrogen processing enzymes.¹ Unlike the [FeFe]-hydrogenases, which are highly sensitive to O₂, the [NiFe]-hydrogenases are more tolerant to aerobic environments.² The structure of the active site contains a heterometallic core, consisting of a Ni(SCys)₄ center coordinated to a Fe(CN)₂(CO) center.³ The Ni(SCys)₄ center is redox active, and switches between the Ni(III)/Ni(II) oxidation states, whereas the Fe center is always in the Fe(II) oxidation state (Figure 6.1).⁴ Unlike the [FeFe]-hydrogenases, there is no direct attachment of the active site to a redox-active [4Fe4S] cluster, although a highly conserved [4Fe4S] cluster is present in close proximity. As isolated, the enzyme is not active, and contains additional µ-OH or µ-OOH ligands. Upon treatment with H₂, these deactivated states are reduced to the Ni-SI state with concomitant release of the bridging ligand. Further reaction with H₂ provides the Ni-C and Ni-R states, which exist in rapid equilibrium with H₂ (Figure 6.2). When the [4Fe4S] clusters are fully reduced, treatment of Ni-SI with H₂ provides the Ni-R state directly, whereas otherwise the Ni-C state would be obtained.⁵⁷
Figure 6.1. Structure of the fully-reduced [NiFe]-hydrogenase active site. (PDB: 1WUL)

The most well-characterized active state of the enzyme, Ni-C, contains a paramagnetic (µ-H)Ni(III)Fe(II) center. The hydride ligand has been confirmed by hyperfine coupling in EPR spectra, followed by ENDOR, and HYSCORE spectroscopies.\textsuperscript{8,9} In addition, the hydride ligand is observed to be photo-labile, where it undergoes proton transfer to a cysteinate ligand providing (CysSH)Ni(I)(Fe(II)), the Ni-L state.\textsuperscript{10-12} Oxidation of Ni-C provides Ni-SI, which is proposed to be a Ni(II)Fe(II) center wherein the Fe center is a 16 e\textsuperscript{-} fragment, and the proton has been lost to the enzyme. Electrochemical titrations have demonstrated that Ni-R is one-electron and one-proton separated from Ni-C. Thus, Ni-R is most likely a (CysSH)(µ-H)Ni(II)Fe(II) center. At high pH, Ni-R is converted to Ni-R’, which is speculated as containing a (CysS)(µ-H)Ni(II)Fe(II) center, although the exact identity is not determined.\textsuperscript{13}
Figure 6.2. The proposed active states of [NiFe]-hydrogenase.

The catalytic cycle of [NiFe]-hydrogenase is currently quite speculative. The most agreed upon mechanism involves all three reduced forms within the catalytic cycle (Figure 6.3). This simplified mechanism is poorly understood as observed chemical transformations often include concerted proton and electron transfer steps. The observation of several different protonated states of Ni-R has led to several different mechanisms, some involving spectator hydride ligands, mixed-valence intermediates (Figure 6.4), and tri-hydrides.
Figure 6.3. Simplified catalytic mechanism of [NiFe]-hydrogenase.

Figure 6.4. A proposed catalytic cycle of [NiFe]-hydrogenase including a transient Ni-L state.
Unlike [FeFe]-hydrogenase, wherein diiron dithiolates had been known 70 years prior to the structure of the enzyme, heterobimetallic complexes containing Ni and Fe were rare.\textsuperscript{17} Prior to the structural characterization of [NiFe]-hydrogenase in 1995,\textsuperscript{3} modeling efforts were benchmarked on reactivity, and monomeric Ni complexes were developed that reacted with CO, H\textsuperscript{+}, H\textsuperscript{-}, and even H\textsubscript{2}.\textsuperscript{18-20} However, after the structure synthetic chemists quickly altered their focus and developed structural models incorporating both Ni and Fe (Figure 6.5).\textsuperscript{21-23} These structural model complexes had little to no functional similarities with the enzyme. In addition, the oxidation state of the metal centers were either Ni(II)Fe(II) and contained 18e\textsuperscript{-} Fe fragments (compared to Nature’s 16e\textsuperscript{-}) or a very abiological Ni(II)Fe(0) center. Despite the definitive spectroscopic evidence that the catalytic Ni-C state contained a $\mu$-hydride ligand, synthetic chemists had failed to include a hydride ligand in any NiFe heterobimetallic center.\textsuperscript{24}

\begin{figure}[h]
\centering
\includegraphics[width=\textwidth]{figure6_5.png}
\caption{First-generation [NiFe]-hydrogenase model complexes.}
\end{figure}

After this series of structural models, synthetic chemists split into two directions. The first continued to pursue highly developed structural model complexes and focused
on combining the Ni(SR)$_4$ fragments with a Fe(CN)$_2$(CO) center (Figure 6.6). However, these models were plagued by an additional ligands at Fe, and despite being a very good structural models, there was no functionality.$^{25,26}$ The second direction focused on achieving functional model complexes, and several nickel iron complexes were developed that were capable of catalytically reducing protons to H$_2$, although none isolated proposed hydride intermediates.$^{27}$

![Figure 6.6. Well-developed structural models of the [NiFe]-hydrogenases.](image)

A significant advance was made with a ruthenium nickel complex that incorporated a hydride ligand, was capable of oxidizing H$_2$ catalytically, and even soluble in water (Figure 6.7).$^{28-31}$ Despite the wealth of known dihydrogen chemistry with ruthenium,$^{32-37}$ and previous ruthenium nickel complexes,$^{38}$ the work was praised as a functional [NiFe]-hydrogenase model complex. A review of the work reenergized synthetic chemists to focus on hydride ligands, and reactivity of the Ni(µ-SR)$_2$Fe core.$^{39}$ Thus, future generations of NiFe model complexes moved away from the idea of
achieving functionality through structural models, and toward investigations of Ni(µ-SR)₂Fe centers with hydride-delivering substrates (H⁻, H⁺, H⁺).⁴⁰

![Figure 6.7](image.png)

**Figure 6.7.** Functional hydrogen oxidation catalyst based on a Ni(µ-SR)₂(µ-H)Ru center.

**References**


18. Baidya, N.; Olmstead, M. M.; Mascharak, P. K. "Mononuclear nickel(II) complex with [NiN₃S₂] chromophore that readily affords the nickel(I) and nickel(III)


Chapter 7

Hydride-Containing Models for the Active Site of the Nickel-Iron Hydrogenases

Introduction

Two families of hydrogenases, the [FeFe]-hydrogenases and the [NiFe]-hydrogenases,\textsuperscript{1-3} have stimulated intense work on the development of bioinspired catalysts for hydrogen processing.\textsuperscript{5,6} In the case of the [FeFe]-hydrogenases, the transition from structural to functional model complexes was rapid owing to foundational work conducted, albeit without awareness of the biological connection,\textsuperscript{7-11} years before the structural characterization of the proteins. Modeling the [NiFe]-hydrogenases has proven more challenging, despite the fact that the relevant proteins had been structurally characterized already in the 1990’s. Relative to the [FeFe]-enzymes, the [NiFe]-hydrogenases are well suited for modeling since these protein are diverse,\textsuperscript{12} sometimes oxygen-tolerant,\textsuperscript{13} and are catalytic biased for H\textsubscript{2} oxidation,\textsuperscript{14} which is the more challenging reaction for model complexes.

Diverse structural models for the [NiFe]-hydrogenases have been described,\textsuperscript{15,16} including many that feature Ni(SR)\textsubscript{2}Fe cores complemented by diatomic ligands on Fe.\textsuperscript{17-19} Unlike model complexes for the active site of the [FeFe]-hydrogenases models for the [NiFe]-enzymes lack hydride ligands. As the hydride ligand is an definitive structural feature of the active site, and required for catalysis, we have focused our strategies on installing hydride ligands.\textsuperscript{20} For example, one could attach Ni modules to substitutionally labile iron hydrides and, complementarily, attach Fe species to Ni hydrides. We have investigated the former route with partial success. The species
[HFe(CN)₂(CO)₃]⁻, which contains the biomimetic ferrous HFe(CO)(CN)₂⁻ module, has been prepared efficiently.²¹ Although this anion undergoes well-behaved substitution reactions, thus far we have been unable to couple it to nickel thiolates. Alternatively, we envisioned reactions of preformed Ni(SR)₂Fe ensembles with the equivalent of H⁻ or H⁺. Hydrides can be installed on diferrous dithiolates using BH₄⁻ salts.²² We have generated impure samples of [(dppe)Ni(μ-pdt)(μ-H)Fe(CO)(dppe)]⁺ from [(dppe)Ni(μ-pdt)Fe(CO)₂(dppe)]²⁺ via this method (dppe = 1,2-bis(diphenylphosphino)ethane, pdt = 1,3-propanedithiolate).²³

The final approach to suitable nickel-iron hydrides calls for protonation of reduced Ni(SR)₂Fe species, an analogous route is well developed in the modeling of the [FeFe]-hydrogenases, leading to catalysts for proton reduction.²⁴ Most structural models feature Ni(II)(SR)₂Fe(II) cores, but a few consist of L₂(RS)Ni(II)(SR)Fe(0)(CO)₄ centers, which in principle could be protonated at iron.²⁵,²⁶ Despite its reputed instability, the Ni(I)(SR)₂Fe(I) species (dppe)Ni(μ-pdt)Fe(CO)₃ (1) was attractive to us.²⁷ This chapter describes a synthetic procedure for the clean preparation of 1 and its protonation affording the first hydride-containing nickel-iron model complex, [1H]⁺. Complex [1H]⁺ is substitutionally labile allowing the synthesis of several phosphine-substituted derivatives. All nickel-iron hydride complexes investigated were shown to be active catalysts for the reduction of protons.
Results and Discussion

**Synthesis and characterization of NiFe(pdt)(dppe)(CO)$_3$, 1.** Complex 1 has been recently synthesized and characterized crystallographically by Schröder et al. Following a modified synthetic procedure employing Fe$_2$CO$_9$ as the “Fe(CO)$_3$” source, 1 can be cleanly synthesized up to 35% yield following an air-sensitive silica column. As isolated, 1 is a stable crystalline green solid which decomposes slowly upon exposure to air, or when heated to 90 °C. The FT-IR spectrum displayed a simple pattern expected for C$_{3v}$ geometry containing a single strong A$_1$ band at 2028 cm$^{-1}$, and a slightly weaker and broader E$_1$ band at 1952 cm$^{-1}$ (Figure 7.1). In addition, the $^{31}$P{$^1$H} NMR spectrum of our samples differ from that previously reported, consisting of a broadened singlet at room temperature that decoaleses into two doublets at -68 °C, consistent with rotation of the trigonal bipyramidal Ni(dppe) site (Scheme 7.1). Cyclic voltammetry of CH$_2$Cl$_2$ solution of 1 displayed one reversible oxidation event at -0.5 V vs Fc$^{0/+}$ and one irreversible couple at ~+0.3 V vs Fc$^{0/+}$ (Figure 7.2). Indicating a one-electron oxidation, bulk chemical oxidation of 1 with FcBF$_4$ in CH$_2$Cl$_2$ shifted the ν$_{CO}$ bands by ~ 29 cm$^{-1}$ (Figure 7.1). An EPR spectrum of a frozen 8:2 CH$_2$Cl$_2$:THF solution of [1]BF$_4$ exhibits an intense nearly axial signal which lacks phosphine hyperfine coupling (Figure 7.3). This spectrum is best described as a square-pyramidal Fe(I) ion where the SOMO resides near the Ni-Fe bond (Figure 7.4). This radical cation exists in the same oxidation state assigned to Ni-L state of the enzyme. 4
Figure 7.1. FT-IR spectra of (CO)$_3$Fe(pdt)Ni(dppe) (1, top), [(CO)$_3$Fe(pdt)Ni(dppe)]BF$_4$ ([1]BF$_4$, middle) [(CO)$_3$Fe(pdt)(μ-H)Ni(dppe)]BF$_4$ ([H]BF$_4$, bottom) in CH$_2$Cl$_2$.

Scheme 7.1. Proposed Ni-centered dynamics for 1 (L = CO) and 2, 3, 4 (L = PR$_3$).
Figure 7.2. Cyclic voltammetry of (CO)$_3$Fe(pdt)Ni(dppe) (1) in CH$_2$Cl$_2$ solution (0.1M NBu$_4$PF$_6$, scan rate of 100 mV/s). The electrochemical events at –0.4 V and –0.8 V result from the $1^{+/-2+}$ couple, as they are not present initially. To convert Ag|AgCl to Fc$^{0/+}$ add 0.520 V.²⁹
Figure 7.3. X-band EPR spectrum of a 8:2 CH₂Cl₂:THF solution of [(CO)₃Fe(S₂C₃H₆)Ni(dppe)]BF₄ ([1]+). The high and low field signals correspond to \( g_\perp = 2.005 \) and \( g_\parallel = 2.05 \), respectively. The signal simulated as subspectrum B is of unknown origin. Fitting parameters for Simulated A: \( g_x = 2.0561 \), \( LW = 11.40 \) G; \( g_y = 2.0536 \), \( LW = 10.42 \) G; \( g_z = 2.0080 \), \( LW = 7.61 \) G. Fitting parameters for Simulated B: \( g_x = 2.0561 \), \( LW = 15.33 \) G; \( g_y = 2.0401 \), \( LW = 11.82 \) G; \( g_z = 2.0110 \), \( LW = 10.05 \) G.

Figure 7.4. The mixed-valence radical cation [1H]BF₄. The depicted SOMO is supported by the EPR spectrum, which indicated a square pyramidal iron-localized SOMO lacking \( ^{31}P \) hyperfine coupling.
The first nickel-iron hydride \([\text{HNiFe(pdt)(dppe)(CO)}_3]^+\). Treatment of a \(\text{CH}_2\text{Cl}_2\) solution 1 with \(\text{HBF}_4\cdot\text{Et}_2\text{O}\) or \(\text{CF}_3\text{COOH}\) resulted in immediate protonation to yield the respective salts of the cationic hydride. The tetrafluoroborate salt ([1H]BF₄) was isolated as a stable red microcrystalline solid that is soluble in \(\text{CH}_2\text{Cl}_2\), THF, MeCN, and MeOH. Consistent with protonation at the Ni-Fe bond, the \(v_{\text{CO}}\) pattern in the FT-IR spectrum displays an identical pattern to that of 1, shifted by 54 cm\(^{-1}\) (Figure 7.1). The \(^{31}\text{P}\{^1\text{H}\}\) NMR spectrum displays a sharp singlet at \(\delta 71\) (Figure 7.5), whereas the \(^1\text{H}\) NMR spectrum displays a high-field hydride signal at \(\delta -3.5\) (\(tt, J = 6, 0.6\) Hz, Figure 7.7). The 6 Hz coupling is a typical \(J_{\text{PH}}\) for nickel phosphine hydrides, e.g. \(J_{\text{PH}} = 6\) Hz for \([\text{HNi(dppe)}_2]\text{AlCl}_4\).\(^{30}\) The smaller coupling, confirmed by the \({^1\text{H}}_{-}{^1\text{H}}\) COSY spectrum (Figure 7.6), arises from coupling to protons on the dithiolate. Excess acid is not deleterious, i.e. [1H]\(^+\) does not protonate further, even with large excess HBF\(_4\) Et\(_2\)O.
Figure 7.5. $^{31}\text{P}^1\text{H}$-NMR spectrum of $[(\text{CO})_3\text{Fe}^{}(\text{pdt})(\mu\text{-H})\text{Ni}^{}(\text{dppe})]\text{BF}_4$ in CD$_2$Cl$_2$. 
Figure 7.6. Top: $^1$H NMR spectrum of [H]BF$_4$ in CD$_2$Cl$_2$ with labeled hydrogen atoms, hydrogens contributing the largest coupling constants to the labeled atom are given in parentheses. For example, H$_B$ appears as a quartet of triplets, the quartet is from coupling to H$_B$ and two H$_D$’s. Bottom: {$^1$H-$^1$H} COSY spectrum of the same solution highlighting the observed cross-peak between the hydride (H$_E$) and pdt proton (H$_C$).
Solutions of $[{\text{H}}]\text{BF}_4$ are stable in air for days. Representing the first example of a nickel-iron hydride, $^{31}$ $[{\text{H}}]\text{BF}_4$ was characterized crystallographically (Figure 7.7). Including the bridging hydride ligand, the Fe center is quasi-octahedral and Ni is approximately square-pyramidal. The hydride connectivity, the metal coordination numbers, and the presence of three diatomic terminal ligands on Fe match features of the active site of the enzyme in the Ni-R state. $^{3}$ The complex $[{\text{H}}]^+$ resembles the 34e diiron dithiolato hydrides $[\text{Fe}_2(\text{pdt})(\mu-\text{H})(\text{CO})_{6-x}(\text{PR}_3)_x]^+$, which are considered to be useful models for the [FeFe]-hydrogenases. $^{6}$ In contrast to the diiron hydrides, $[{\text{H}}]^+$ deprotonates rapidly. Upon treatment with NEt$_3$ ($[\text{HNEt}_3]\text{BF}_4$, $pK_a = 18$), $[{\text{H}}]^+$ was found to rapidly give 1.

Figure 7.7. Structure of the cation in $[(\text{dppe})\text{NiFe(pdt)(µ-H)(CO)}_3]\text{BF}_4$ ($[{\text{H}}]\text{BF}_4$). Key distances (Å): Ni-Fe, 2.6131(14); Ni-S(1), 2.210(2); Ni-S(2), 2.219(2); Fe-S(1), 2.321(2); Fe-S(2), 2.322(2); Ni-H, 1.64(6); Fe-H, 1.46(6).
Owing to the poor Lewis acidity of 1, no reaction occurs in the presence of weak Lewis acids such as BPh$_3$. However, these Lewis acids have been shown to form frustrated Lewis acid-base pairs and are capable of the activation of H$_2$. The bimetallic nickel-iron complexes are capable of similar activation, $^{31}$P{$^1$H} and $^1$H NMR spectra of CD$_2$Cl$_2$ solutions of 1 and BPh$_3$ show trace conversions to [1H]$^+$ (Figure 7.8). This protonation is presumably due to trace H$_2$O, as after several hours no more [1H]$^+$ is formed. Upon addition of H$_2$, $^{31}$P{$^1$H} and $^1$H NMR spectra show complete and immediate conversion to [1H]$^+$, and some trace impurities, which are similar to the products from the reaction of [NBu$_4$]BH$_4$ and 1 (Figure 7.9). This reaction can be thought of as an analogue to the [Fe]-hydrogenases, which are incapable of H$_2$ activation without the Lewis acidic methenyl-H$_4$MPT cofactor.$^{33,34}$

Figure 7.8. $^{31}$P{$^1$H} NMR spectrum of an equimolar solution of 1 and BPh$_3$. Inset shows the corresponding $^1$H NMR spectrum, displaying [1H]BF$_4$. 

179
Figure 7.9. $^{31}$P($^1$H) NMR spectrum of 1, BPh$_3$, and H$_2$, inset shows the corresponding $^1$H NMR spectrum, displaying the HNiFe signal for [1H]$^+$. 
Phosphine substitution of [(dppe)NiFe(pdt)(µ-H)(CO)₃]BF₄ providing [(dppe)NiFe(pdt)(µ-H)(CO)₂(PR₃)]BF₄. Mono-substituted complexes [HNiFe(pdt)(dppe)(PR₃)(CO)₂]BF₄ were prepared via thermal and photochemical substitution of [1H]BF₄ (Eq. 7.2). FT-IR spectra of these adducts feature a pair of approximately equally intense νₐ₈ bands at about 2025 and 1970 cm⁻¹ (Figure 7.11). The positions of these bands indicate the expected sequence of basicity, i.e. P(OPh)₃ < PPh₂Py < PPh₃. The ¹H NMR spectra exhibit doublet of triplets (Jₚₜ ~ 35, 4 Hz) in the hydride region, the coupling to the dppe ligands being slightly less than that of the tricarbonyl hydride, [1H]BF₄ (4 vs 6 Hz, Figure 7.10). In addition, all three monosubstituted hydrides display an additional doublet of triplets (Jₚₜ ~ 40, 5 Hz).
accounting for ~1% of the sample. The identity of this second species is not known, but we suggest that it is an isomer as it occurs in all three derivatives, but its NMR shift varies with the identity of the monodentate ligand L.

\[
\text{[HNiFe(pdt)(dppe)(CO)₃]BF₄ + L \rightarrow [HNiFe(pdt)(dppe)(L)(CO)₂]BF₄ + CO (Eq. 7.2)}
\]

\[
\begin{align*}
2\text{H} & \text{BF₄, } L = \text{P(OPh)}₃ \\
3\text{H} & \text{BF₄, } L = \text{PPh₃} \\
4\text{H} & \text{BF₄, } L = \text{PPh₂Py}
\end{align*}
\]
Figure 7.11. FT-IR spectra in the $\nu_{\text{CO}}$ region for CH$_2$Cl$_2$ solutions of the nickel-iron hydride complexes described in this work (top to bottom): [1H]BF$_4$; [2H]BF$_4$; [3H]BF$_4$; [4H]BF$_4$. The $\nu_{\text{CO}}$ band for the Ni-R state occurs at 1936-1948 cm$^{-1}$, depending on the organism.$^4$
Variable temperature $^{31}$P{$^1$H} NMR spectra provide insights into the dynamics that cannot be readily obtained for the more symmetrical [1H]BF$_4$. The signal assigned to dppe (δ 65) was broad at room temperature but decoupled at -30 °C into two doublets. The singlet for PPh$_3$ remained unchanged throughout this experiment (Figure 7.12). At room temperature, the $^{13}$C{$^1$H} spectrum for [3H]BF$_4$ displays three inequivalent methylene signals for the μ-pdt ligand, while just one for the dppe backbone (Figure 7.13). Upon cooling, the Ph$_2$PCH$_2$CH$_2$PPh$_2$ signal decouples into two separate peaks (Figure 7.14). This DNNMR pattern is consistent with a dynamic square-pyramidal to trigonal-bipyramidal interconversion of the HNi(dppe)(pdt) subunit.
(Scheme 7.2). The Fe(CO)$_2$(PPh$_3$) subunit remains unchanged throughout this process, resulting in inequivalent pdt signals. This process is also present in [1H$^+$], which displays a broadened $^1$H NMR signal for the dppe methylene’s, and two inequivalent Fe-CO signals by $^{13}$C{$^1$H} NMR (Figure 7.15). We have previously shown that related diirondithiolates, e.g. Fe$_2$(pdt)(CO)$_3$(dpv)(PMe$_3$) are subject to a turnstile rotations of each Fe(L)$_3$ subunit, although with much higher barriers.$^{35,36}$

Scheme 7.2. Representation of the Ni-centered dynamic process observed for [2H]BF$_4$, [3H]BF$_4$, and [4H]BF$_4$, involving a square-pyramidal to trigonal bipyramidal rearrangement of the HNi(dppe)(pdt) subunit.
Figure 7.13. $^{13}$C$^{1}$H NMR spectrum (CD$_2$Cl$_2$, +19 °C) of [3H]BF$_4$, showing three inequivalent $^{13}$C signals for the pdt ligand (25, 26, 36 ppm) and one broad PPh$_2$CH$_2$CH$_2$PPh$_2$ signal (28 ppm).
Figure 7.14. $^{13}$C$^1$H NMR spectrum (CD$_2$Cl$_2$, -60 °C) of [3H]BF$_4$ showing three inequivalent $^{13}$C signals for the pdt ligand (25, 26, 37 ppm), two signals for PPh$_2$CH$_2$CH$_2$PPh$_2$ (28, 29 ppm), and two signals for Fe(CO)$_2$(PPh$_3$), each appearing with some $^2$J$_{PC}$ coupling from PPh$_3$. 
Figure 7.15. $^{13}\text{C}^{1\text{H}}$ NMR spectrum (CD$_2$Cl$_2$, +19 °C) of [1H]BF$_4$ showing two pdt signals (26, 36 ppm), one PPh$_2$CH$_2$CH$_2$PPh$_2$ signal ($t$, 30 ppm, $^1J_{PC}$~$^2J_{PC}$), and two Fe(CO)$_3$ signals (204, 205 ppm).
Complex [3H]BF₄ was further characterized by X-ray crystallography, which confirmed that PPh₃ is cis to the hydride (Figure 7.16). The Fe(1)-Ni(1) distance of 2.6432(7) Å is only slightly longer than that of [1H]BF₄ (2.6131(14) Å). The iron nickel distance in the *D. vulgaris* enzyme for the Ni-C/Ni-R state is 2.55 Å.¹ The bridging hydride position for [3H]BF₄ is unsymmetrical, closer to iron than nickel by 0.40 Å, whereas in [1H]BF₄ the hydride ligand is closer to iron by only 0.18 Å. The unsymmetrical character of the hydride ligand is also indicated by the diminished value of $J_{\text{dppheH}}$. The Fe(1)-S(1)-Ni(1) angle in [3H]BF₄ (71.42(3)°) is almost identical to that for [1H]BF₄ (70.39(6)°). The OC-Fe-CO angle is ~ 99°, vs the value of 96° calculated from the difference in the intensities of the two $\nu_{\text{CO}}$ bands.³⁷
Deprotonation to provide \( \text{NiFe(pdt)(dppe)(PPh}_3\text{)(CO)}_2 \) (3).

NiFe(pdt)(dppe)(PPh\(_3\))(CO)\(_2\) (3) can be prepared in analytical purity by the deprotonation of [3H]BF\(_4\) with NaOMe. The rate of deprotonation varies with base, but is generally slow, requiring \( \sim 1 \) h for NaOMe in MeOH/CH\(_2\)Cl\(_2\) and 3-5 hours for NEt\(_3\) and pyridine-derived bases. The \( \nu_{\text{CO}} \) bands of 3 (CH\(_2\)Cl\(_2\): 1971, 1916 cm\(^{-1}\)) shift by about 45 cm\(^{-1}\) toward lower energy (Figure 7.17). The \( ^{31}\text{P}\{^{1}\text{H}\} \) NMR spectrum (Figure 7.18) of a CD\(_2\)Cl\(_2\) solution at room temperature displays a triplet (\( \delta \) 55) assigned to the PPh\(_3\) ligand and two broad signals for the dppe (\( \delta \) 77, 45). Upon cooling the sample to -20 °C, the dppe signals sharpen to the expected AB quartet and the PPh\(_3\) signal appears as a doublet-of-doublets. This dynamic behavior is analogous to that proposed for 1.
The Pyridylphosphine $\text{[NiFe}(\mu$-$\text{H})(\text{pdt})(\text{dppe})(\text{PPh}_2\text{py})(\text{CO})_2]$. In efforts to incorporate a proton relay within the coordination sphere of $(\mu$-$\text{H})\text{NiFe}$, diphenylpyridylphosphine was employed. $^1\text{H}$, $^{31}\text{P}$, and FT-IR spectroscopies confirm that $[4\text{H}]^+$ is structurally similar to $[3\text{H}]^+$. However unlike $[3\text{H}]^+$, CH$_2$Cl$_2$ solutions of $[4\text{H}]\text{BF}_4$ were immediately deprotonated by NEt$_3$, (Scheme 7.3). Addition of D$_2$O to a $d^6$-acetone solution of $[4\text{H}]\text{BF}_4$ resulted in complete conversion to $[4\text{D}]\text{BF}_4$ within seconds. Under the same conditions the corresponding PPh$_3$ derivative $[3\text{H}]\text{BF}_4$ was found to exchange only slowly ($t_{1/2} = 20$ min) with D$_2$O. Further supporting this facile intramolecular proton relay, CH$_2$Cl$_2$ solutions of $[4\text{H}]^+$ exposed to air provide FT-IR spectra shifted negative by ~ 20 cm$^{-1}$ (2000, 1940 cm$^{-1}$). Although speculative, we
assign this species to a pyridinium mixed-valence Ni(II)Fe(I), [4H]$^{2+}$. Such species would be formed by \textit{intramolecular} proton transfer to the pyridine ligand, followed by oxidation of the Ni(II)Fe(I) center. The FT-IR spectrum agrees with this assignment, $\sim 15$ cm$^{-1}$ higher in energy than the mixed valence [3]$^+$. This process is related to [NiFe]-hydrogenase, wherein oxidation of Ni-C, ($\mu$-H)Ni(III)Fe(II), provides (CysSH)Ni(II)Fe(II); presumably via (CysSH)Ni(I)Fe(II).

**Scheme 7.3.** Intramolecular proton relay for [4H]$^+$, subsequent oxidation to provide [4H]$^{2+}$, and protonation of [4H]$^+$ to provide [4H$_2$]$^{2+}$.

Confirming the Brønsted basic nature of the pyridyl group of [4H]$^+$, addition of excess CF$_3$CO$_2$H to a CH$_2$Cl$_2$ solution of [4H]$^+$ gave a new hydride with a concomitant increase of $\nu$$_{CO}$ by $\sim 10$ cm$^{-1}$, consistent with protonation of the pyridine ligand providing
[4H₂]²⁺ (Figure 7.20). The ¹H NMR spectrum of [4H₂]²⁺ features a doublet of triplets centered at δ = -3.11 (J₁H = 40, 3.5 Hz). The ³¹P{¹H} NMR spectrum resembles that for [3H]BF₄⁻, suggesting that the overall structure and stereodynamics of the hydride remain unchanged. Consistent with the weak donor ability of PPh₂pyH⁺, [4H₂]²⁺ is not stable for prolonged periods of time or in the presence of excess acid and degrades to [1H]⁺, and free phosphine.

In addition to facile intramolecular proton transfer, solutions of [4H]⁺ also indicate facile intermolecular proton transfer as FT-IR spectra display the expected frequencies for [4H]⁺, but also for the neutral 4 and the N-protonated hydride [4H₂]²⁺ (Figure 7.19, Scheme 7.3). The relative amounts of the organometallic components in the solution were estimated by simulation of the IR spectrum, giving Kₚʳᵒᵗ (Eq. 7.5).

\[
K_{prot} = ([4])([4H₂]²⁺)/([4H]⁺)² \approx 0.03 \quad \text{(Eq. 7.5)}
\]
Figure 7.20. FT-IR spectrum of $[4\text{H}_2]^2+$ in CH$_2$Cl$_2$ solution.
The Tetra-phosphine $[\text{NiFe(μ-H)(pdt)(dppe)}_2(\text{CO})]^{\text{+}}$. Highly substituted derivatives incorporating four phosphine ligands can be synthesized by ultraviolet irradiation of CH$_2$Cl$_2$ or THF solutions of [1H]$^+$ with appropriate ligand. However, this method is only applicable with weakly basic phosphine ligands, as more basic ligands result in the deprotonation of [1H]$^+$ to provide 1, and small isocyanide or cyanide ligands result in complicated reaction mixtures, often with elimination of the Ni(dppe) ligand. Tetra-phosphines made by this method from P(OPh)$_3$, P(OMe)$_3$, and PTA have been characterized by FT-IR and $^1$H NMR spectroscopies. Treatment of [1H]$^+$ with dppe provides $[\text{NiFe(μ-H)(pdt)(dppe)}_2(\text{CO})]^{\text{+}}$, [5H]$^+$. The FT-IR spectrum of [5H]$^+$ in CH$_2$Cl$_2$ displays two overlapping ν$_{\text{CO}}$ peaks at 1932, 1928 cm$^{-1}$ (Figure 7.21). The $^{31}$P{$^1$H} displays two large signals at 82 and 60 ppm consistent with a C$_{2v}$ structure, and a three smaller signals at 64, 62, and 48 ppm. The $^1$H NMR displays two hydride signals at $-2.8$ and $-6.2$ ppm (Figure 7.22), both triplets with J$_{\text{PH}} \sim 25$ Hz. The $^{31}$P{$^1$H} and $^1$H NMR spectra are consistent with two isomers, one C$_{2v}$ and the other C$_6$ (Figure 7.22).
Figure 7.21. FT-IR spectrum of [5H]BF₄ in CH₂Cl₂ solution.
Figure 7.22. Top: High-field $^1$H NMR spectrum (CH$_2$Cl$_2$, 500 MHz) of [5H]BF$_4$ showing two hydride signals, both triplets. Bottom: proposed isomers of [5H]BF$_4$. 
It was recently demonstrated that P(OPh)₃ can act as a labile ligand for HFe(CO)(L)ₓ centers, providing clean routes to more highly-substituted complexes.²¹ For this reason [2H]BF₄ was synthesized, and phosphite-displacement reactions were attempted. Unfortunately, thermal substitution of THF solutions of [2H]BF₄ with simple ligands such as PPh₃ or dppe provided no reaction. However, upon treatment of THF solutions of [2H]BF₄ with 1,2-bis-dimethylphosphinoethane (dmpe) we isolated low yields of [NiFe(µ-H)(pdt)(dmpe)(dppe)(CO)]⁺, [6H]⁺. The crystallographic structure firmly establishes that dmpe has displaced a the nickel diphosphine, and the dppe ligand has migrated to iron (Figure 7.23). When monitoring the reaction by FT-IR spectroscopy it is apparent that upon mixing [2H]⁺ with dmpe, deprotonation occurs. After stirring overnight, slow growth of [6H]⁺ is observed, with no intermediates detected.

Figure 7.23. Crystallographic structure of [NiFe(µ-H)(pdt)(dmpe)(dppe)(CO)]⁺, [6H]BF₄.
Although with low-conversion, [5H]$^+$ can be synthesized by the treatment of [NiFe(pdt)(dppe)$_2$(CO)$_2$]$^{2+}$ ([(5(CO)$_2$)$_2$$]^{2+}$) with [NBu$_4$]BH$_4$. As highly substituted diiron dithiolates have been successfully synthesized by bulk oxidations in the presence of ligands, we attempted similar synthesis with [(5(CO)$_2$)$_2$$]^{2+}$. Indeed, one-pot synthesis of [(5(CO)$_2$)$_2$$]^{2+}$ was achieved by treatment of 1 with 2 equivalents of FcBF$_4$ in the presence of dppe (Scheme 7.4). The resulting dicarbonyl dication was isolated and characterized by FT-IR and $^{31}$P($^1$H) spectroscopies (Figure 7.24, Figure 7.25). Similar to results obtained with [(5(CO)$_2$)(OTf)$_2$, [(5(CO)$_2$)]BF$_4$ reacted with [NBu$_4$]BH$_4$, providing a very small conversion to [5H]BF$_4$. In efforts to utilize these easily prepared NiFe complexes, we sought to reduce [(5(CO)$_2$)$_2$$]^{2+}$ to mixed-valence Ni(II)Fe(I) or fully-reduced Ni(I)Fe(I) derivatives. Upon addition of CoCp$_2$ FT-IR spectroscopy shows a clear shift of the $\nu_{CO}$ peak to lower frequency, $\sim$1935 cm$^{-1}$ (Figure 7.26). However, the resulting complex could not be isolated, and is not consistent with the expected $\nu_{CO}$ frequency for a mixed-valence [5]$^+$, nor fully-reduced 5. Further work in this direction is needed to fully understand the resulting reactions.

**Scheme 7.4.** One-pot oxidation of 1 to provide [(5(CO)$_2$)](BF$_4$)$_2$. 
Figure 7.24. FT-IR spectrum of [5(CO)$_2$](BF$_4$)$_2$ in CH$_2$Cl$_2$ showing one major trans-isomer at 1986 cm$^{-1}$, and a small cis-isomer at ~2053 cm$^{-1}$.

Figure 7.25. $^{31}$P{$^1$H} NMR spectrum of [5(CO)$_2$](BF$_4$)$_2$ in CD$_2$Cl$_2$. 

200
Figure 7.26. In-situ reduction of a CH$_2$Cl$_2$ solution of [5(CO)$_2$](BF$_4$)$_2$ with CoCp$_2$. 
Redox Properties of Nickel-Iron Hydrides. The complexes \([1\text{H}]\text{BF}_4\), \([2\text{H}]\text{BF}_4\), \([3\text{H}]\text{BF}_4\), undergo quasi-reversible reductions at relatively mild potentials, as judged by cyclic voltammetry (Table 7.1). The corresponding reduction for \([4\text{H}]\text{BF}_4\) was irreversible. For complex \([2\text{H}]\text{BF}_4\), this couple was quite reversible. The peak-to-peak separation for the \([2\text{H}]^+\) reduction couple (\(\Delta E_p \sim 65\) mV) indicated a one-electron process. Additionally, the scan rate dependence was very similar to that of the \(2^{0/+}\) couple, which we have independently established undergoes a one-electron oxidation providing the mixed-valence Ni\(^{II}\)Fe\(^{I}\) complex (Figure 7.27, Figure 7.28).\(^\text{40}\) Following the Randles-Sevick equation, the diffusion constants for \([2\text{H}]^+\) and 2 can be calculated as \(2.9 \times 10^{-6}\) and \(3.5 \times 10^{-6}\), respectively. Additionally, \([2\text{H}]\text{BF}_4\), \([3\text{H}]\text{BF}_4\), and \([4\text{H}]\text{BF}_4\) irreversibly oxidize at \(\sim 0.5\) V vs Fc\(^{0/+}\), for \([1\text{H}]\text{BF}_4\) this oxidation process is \(\sim 0.75\) V vs Fc\(^{0/+}\).

**Table 7.1.** Selected electrochemical properties of the hydrides. Data was collected from a \(\sim 1\) mM freshly-prepared CH\(_3\)CN solution of nickel-iron hydride and 0.1 M \([\text{NBu}_4][\text{PF}_6]\) as electrolyte. **Cell conditions:** see Figure 7.18.

<table>
<thead>
<tr>
<th>Complex</th>
<th>(E_{1/2}) (V) vs Fc(^+/0)</th>
<th>(i_{pal}/i_{pc}) (at 0.1 V/s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>([1\text{H}]^+)</td>
<td>-1.29</td>
<td>0.26</td>
</tr>
<tr>
<td>([2\text{H}]^+)</td>
<td>-1.44</td>
<td>0.93</td>
</tr>
<tr>
<td>([3\text{H}]^+)</td>
<td>-1.49</td>
<td>0.06</td>
</tr>
<tr>
<td>([4\text{H}]^+)</td>
<td>-1.49</td>
<td>0.00</td>
</tr>
<tr>
<td>([4\text{H}_2]^{2+})</td>
<td>-1.28</td>
<td>---</td>
</tr>
</tbody>
</table>
Figure 7.27. Cyclic voltamograms of a 1.852 mM CH$_3$CN solution of [2H]$^+$ (left) and a 1.683 mM 90/10 CH$_3$CN/CH$_2$Cl$_2$ solution of 2 (right) at various scan rates denoted in mV/s.

\[ i_p = (2.687E5) \times n^{3/2} \nu^{1/2} \nu^{1/2} D^{1/2} A C \]  
\text{(Eq. 7.6)}

where:  
2.687E5 (C mol$^{-1}$ V$^{-1/2}$)  
\( n \) = number or electrons  
\( \nu \) = scan rate (V s$^{-1}$)  
\( D \) = diffusion constant (cm$^2$ s$^{-1}$)  
\( A \) = area of electrode (cm$^2$)  
\( C \) = concentration of species (mol/cm$^3$)
Figure 7.28. Scan rate dependence of the [2H]⁺/0 and 2⁺/0 couples, concentrations are ~1.8 mM (see Figure 7.27).

Catalytic Proton Reduction from Nickel-Iron Hydrides. Upon addition of CF₃CO₂H (CH₃CN pKₐ = 12.9, E° = -0.89 V) to CH₂Cl₂ solutions of [1H]⁺, [2H]⁺ and [3H]⁺, cyclic voltammograms displayed increased cathodic current coinciding with the hydride reduction event indicative of proton reduction catalysis (Figure 7.29, Figure 7.30). At the relevant potentials (~ -1.5 V) proton reduction by the glassy carbon working electrode was confirmed to be negligible.⁴¹ In the case of [4H]⁺, addition of CF₃CO₂H resulted in the appearance of a new catalytic current ~200 mV milder than the [4H]⁺/0 couple (Figure 7.31). We attribute this new feature to catalysis by the N-protonated pyridine [4H₂]²⁺ complex, consistent with the previous spectroscopic results. As [4H₂]²⁺ degrades into [1H]⁺ at very high concentrations of CF₃CO₂H (10-100 equiv), the voltammograms display an additional catalytic wave for [1H]⁺.
Figure 7.29. Cyclic voltammograms of a CH₂Cl₂ solution of [1H]BF₄ with increasing equiv of CF₃COOH. Conditions: ~0.5 mM in CH₂Cl₂ (see experimental), 0.1M [NBu₄]PF₆, scan rate 0.1 V/s, glassy carbon working electrode (d = 3.0 mm); Ag wire pseudoreference with internal Fc standard at 0 V; Pt counter electrode).
Figure 7.30. Cyclic voltammograms of CH₂Cl₂ solutions of [2H]BF₄ (left) and [3H]BF₄ (right) with increasing equiv of CF₃COOH (denoted on right). Overpotentials were estimated by the standard potential for hydrogen evolution from CF₃COOH in CH₃CN (see text). Conditions: see caption for Figure 7.29.
Plots of \( i_c/i_p \) (\( i_c = \) catalytic current, \( i_p = \) current in the absence of acid) vs [CF\(_3\)CO\(_2\)H] are linear up to about \( i_c/i_p \sim 16-20 \) (Figure 7.32). This linear dependence indicates a rate law for H\(_2\) evolution that is second order with respect to [H\(^+\)], i.e. H\(_2\) evolution follows protonation of the reduced hydride.\(^{42}\) The initial slope of this plot is a measure of the acid-dependent rate-determining step. For the hydrides investigated, this step is the protonation to reform [HNiFe]\(^+\). Thus, the kinetically more accessible
hydride ([1H]^+) displays a steeper initial slope than the substituted hydrides [2H]^+ and [3H]^+. The relative catalytic activity can be estimated from the acid-independent regions of the graph of $i_c/i_p$ vs [CF$_3$CO$_2$H]. When the rate of catalysis in no longer dependent on [H$^+$], the values for $i_c/i_p = 16$-20 yield an acid-independent rate constant between 50-75 s$^{-1}$ (see SI). For the tricarbonyl [1H]^+, an $i_c/i_p = 10$ represents an acid-independent rate constant of just 10 s$^{-1}$. Thus, phosphine-substitution has a dramatic effect on the maximum turnover rate for nickel-iron hydride complexes.

![Graph](image)

**Figure 7.32.** Influence of [acid]:catalyst ratio on catalytic current ($i_c/i_p$) for [1H]BF$_4$ (black circles), [2H]BF$_4$ (red squares), [3H]BF$_4$ (blue triangles), and [4H]BF$_4$ (green diamonds). **Conditions:** see figure 7.29.

An important way of evaluating hydrogen evolution catalysts is overpotential. To properly determine overpotential, catalysis should be performed in a solvent where the standard reduction potential of the acid, $E^{\circ}_{HA/H2}$, has been determined. However we
faced the practical problem that the hydride complexes are poorly stable in MeCN solution. Nonetheless, cyclic voltammetry of a freshly prepared CH$_3$CN solution of [3H]$^+$ displayed the [3H]$^{+0}$ couple at $-1.45$ V, and the cathodic current dramatically increased upon addition of CF$_3$CO$_2$H. Thus, we conclude that the $E^\circ_{(HA/H_2)}$ values in MeCN solution are applicable to our catalysts. To calculate overpotential, we used the potential at half-height from an acid-independent cyclic voltammogram and subtracted the $E^\circ_{(HA/H_2)}$ value for the acid investigated (CF$_3$CO$_2$H, $E^\circ = -0.89$ V).$^{41,44}$ Complexes [2H]$^+$ and [3H]$^+$ operate at an overpotential of $\sim 430$ mV (Figure 7.30). The pyridinium complex [4H$_2$]$^{2+}$ operates at an overpotential of only $\sim 260$ mV (Figure 7.31).

To address the exact mechanism of H$_2$ production, the reversibility of the [2H]$^{+0}$ couple is diminished in the presence of [HNEt$_3$]BF$_4$. This result is consistent with the reduced derivative [2H]$^0$ being sufficiently basic to undergo protonation by [HNEt$_3$]$^+$. Since [HNEt$_3$]BF$_4$ is unable to protonate 2, the system 2/[HNEt$_3$]$^+$ is catalytically inactive. To further clarify the mechanism of hydrogen production, the catalytic current with excess acid displays a linear dependence on the concentration of added [3H]$^+$, indicating a catalytic mechanism first-order in [3H]$^+$.

**Quantification of Hydrogen Production by Bulk Electrocatalysis of CF$_3$CO$_2$H by [1H]$^+$.** To verify the electrocatalytic production of H$_2$ from CF$_3$CO$_2$H by [1H]BF$_4$, bulk electrolysis was performed in a $\sim 100$ mL glass electrolysis cell. The electrolysis cell was to connected to an online gas chromatograph equipped with a thermal conductivity detector via a gas-inlet and outlet. The electrolysis proceeded with continuous purge of argon through the cell, at a rate of 50 mL/min. A 1 mL sample was
automatically injected into the gas chromatography instrument every 4 min. The hydrogen peak area was then related to concentration by a calibration curve.

Because of the continuous purge of argon through the sample, relatively large quantities of hydrogen were required to be produced continuously. To accomplish this, a Hg pool electrode was used as the working electrode. However, because of the high-resistance of CH₂Cl₂ and the size of the working electrode, the potentiostat could not provide the adequate potential. To overcome this challenge, we explored the use of CH₃CN as solvent during electrolysis. Previously, it had been found that solutions of [1H]⁺ in CH₃CN are not stable and form products with no v_CO bands after ~2 hours, this process is accelerated by the presence of light.

Electrolysis of a 1.3 mM solution (0.04 mmol in 30 mL) of [1H]BF₄⁻ in CH₃CN at -800 mV (vs Ag|AgCl) by the Hg pool electrode resulted in the continuous evolution of hydrogen, as detected by the gas chromatogram. After 30 min of electrolysis, 31 C had passed through the cell. An integration of the hydrogen concentration detected every 4 minutes was used to conclude that ca. 0.13 mmol of H₂ (3 mL) was produced within the 30 min, accounting for 19 C (61 % faradaic yield) (Figure 7.33). A background electrolysis reaction without [1H]BF₄⁻ was performed to quantify the amount of H₂ produced from the Hg pool alone, after 30 min 5 C had passed through the cell. Thus, after 30 min of electrolysis, hydrogen production by [1H]⁺ accounted for 26 C, or 0.13 mmol H₂, ~ 3.2 turnovers. After 30 min the hydrogen production slowed, and the CH₃CN solution was colorless, indicating loss of the dissolved catalyst. As the catalytic cycle proceeds through 1, which is not soluble in CH₃CN, we propose that its precipitation leads to the loss of catalysis.
Figure 7.33. Bulk electrolysis of a 1.3 mM solution of [1H]BF₄ with 87 mM CF₃CO₂H, 0.1 M [NBu₄]PF₆. Conditions: Hg pool working electrode, Ag|AgCl reference electrode, Pt counter electrode. The total coulombs passed through the cell arising from [1H]BF₄ is given by the black solid line. The red dotted line is the total coulombs passed under identical conditions without [1H]BF₄. The solid black dashes represent the coulombs required to account for the H₂ produced from [1H]BF₄ (hydrogen produced from the Hg pool has been subtracted) as determined by gas chromatography.
For future bulk electrolysis reactions CH₂Cl₂ should be used to avoid the loss of catalytic activity, with a smaller working electrode. As the continuous flow of Ar might not be feasible in future cell designs, sealed electrocatalytic reactions should be performed to quantify H₂.

**Thermodynamic Properties of Nickel-Iron Hydrides.** To better understand the properties controlling catalysis, we sought to determine both the pKₐ of the nickel-iron hydrides investigated and the electrochemical properties of the resulting neutral NiFe compounds. As most neutral nickel-iron compounds investigated are not soluble in CH₃CN, a solvent with a well-established pKₐ scale, we employed PhCN for all pKₐ determinations. Although PhCN does not have an established pKₐ scale, it has been used by others and shown to provide similar results as from CH₃CN solutions.⁴⁵ A dilute solution of [1H]⁺ in PhCN with one equivalent of aniline ([PhNH₃]BF₄ CH₃CN pKₐ = 10.7) provided a 1:1 equilibrium mixture of [1H]⁺ and 1, several hours later the [1H]⁺:1 ratio was unchanged thus indicating a pKₐ of ~10.7 for [1H]⁺. Similarly, four equivalents of 4-methoxypyridine ([H(4-methoxypyridine)]BF₄ CH₃CN pKₐ = 14.23)⁴⁶ and [3H]⁺ provided an 2:1 equilibrium ratio of 3 and [3H]⁺, as determined by ³¹P NMR spectroscopy. Although decomposition of the organometallic species occurred over several hours, we confidently assign the pKₐ as ~ 14.9 ± 0.1 as the 3:[3H]⁺ ratio remained unchanged throughout the decomposition. For comparison, the μ-hydrido diiron complex [HFe₂(pdt)(CO)₄(PMe₃)₂]BF₄ has a MeCN pKₐ of 12.⁴⁷

Using cyclic voltammetry, we determined the oxidation potentials 1 and 3 referenced vs Fc⁰/⁺. Similar to the pKₐ determination, PhCN was used to approximate
$E_{1/2}$ values on the MeCN scale. A ~ 1 mM solution of 1 in PhCN displayed two oxidation events, one reversible oxidation at -0.543 V and a second irreversible oxidation at -0.124 V (Table 7.2). The cyclic voltammogram for 3 was similar, displaying a reversible oxidation event at -0.722 V and an irreversible oxidation event at -0.191 V. We assigned all of these couples as one-electron processes as the oxidation of 1 performed chemically provides mixed-valence species previously characterized by EPR spectroscopy.\textsuperscript{23}

The $pK_a$ and electrochemical data were combined in a thermodynamic cycle to calculate the free energy of hydrogen atom donation. Although the oxidation couples 1$^{+/2+}$ and 3$^{+/2+}$ were irreversible, we used their values as rough approximations to determine a hydride donor strength for [1H]$^+$ and [3H]$^+$ (Table 7.2). Following thermodynamic equations 7.7 and 7.8 the free energies of H• and H$^-$ donation, $\Delta G_{H^•}$ and $\Delta G_{H^-}$, were calculated to be ~ 57 and ~79 kcal/mol, respectively. The free energies of hydrogen atom donation represent relatively weak M-H bond strengths\textsuperscript{48,49}, and are relevant to their catalytic efficiency\textsuperscript{14}. The negligible difference between these thermodynamic parameters indicates that carbonyl substitution with PPh$_3$ affects the NiFe$^{I,II/III}$ and NiFe$^{I,II/III}$ couples the same magnitude as the $\Delta pK_a$. Comparing [1H]$^+$ to [3H]$^+$, the energy associated with the $\Delta pK_a$ is 5.7 kcal/mol; However, this energy is not equally offset by the $\Delta E_{1/2}$ of the [HNiFe]$^{+/0}$ couples (4.6 kcal/mol). The result is a more basic hydride, which for its $pK_a$, is slightly easier to reduce.
\[ \Delta G_{\text{H}^+} = 1.37pK_a + 23.06E_{1/2}^{\text{I,II,II}} + 54.9 \]  
(Eq. 7.7)

\[ \Delta G_{\text{H}^-} = 1.37pK_a + 23.06E_{1/2}^{\text{I,II,II}} + 23.06E_{1/2}^{\text{I,II,II}} + 79.6 \]  
(Eq. 7.8)

**Table 7.2.** Thermodynamic data regarding the strength of the \( \mu \)-H bond (E's vs Fc\(^{0+/+}\), \( \Delta G \)'s in kcal/mol)

<table>
<thead>
<tr>
<th></th>
<th>pK(_a)</th>
<th>NiFe(^{0+/+})</th>
<th>NiFe(^{+/+})</th>
<th>HNiFe(^{+/0})</th>
<th>( \Delta G_{\text{H}^+} )</th>
<th>( \Delta G_{\text{H}^-} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>([1H]^+)</td>
<td>~10.7</td>
<td>-0.543 V</td>
<td>-0.124 V</td>
<td>-1.29 V</td>
<td>57</td>
<td>79</td>
</tr>
<tr>
<td>([3H]^+)</td>
<td>~14.9</td>
<td>-0.722 V</td>
<td>-0.191 V</td>
<td>-1.49 V</td>
<td>58</td>
<td>79</td>
</tr>
</tbody>
</table>

**Discussion**

The new hydride complexes are confirmed to be functional models for the [NiFe]-hydrogenases, at least with respect to certain structural features and their ability to catalyze hydrogen evolution. Specifically, the hydrides represent structural mimics of the Ni-R form of these enzymes,\(^4\) an \( S = 0 \) state that is thought to feature an Fe(\( \mu \)-SR)\(_2\)(\( \mu \)-H)Ni core. In the model complexes, the coordination sphere at Ni is square pyramidal, having rearranged from the tetrahedral geometry of the Fe(I)Ni(I) precursor. Such a rearrangement does not occur in the protein, which orients the terminal ligands to sites intermediate between square-planar and tetrahedral.\(^3\) Our DNMR studies also reveal a oscillatory motion that interchanges equivalent structures of the Fe(CO)\(_2\)(PR\(_3\)) site. This dynamic equilibrium does *not* operate in the protein, as the interchange process would place a cyanide ligand in an apical site.
One of the more striking results is the asymmetry of the Fe-H-Ni linkage, which is probably relevant to the mechanism by which these complexes reduce protons. In the parent hydride $[(\text{CO})_3\text{Fe(pdt)(H)}\text{Ni(dppe)}]\text{BF}_4$, $[(\text{H})\text{BF}_4]$ the difference of the iron-hydride and nickel-hydride bond distances ($\Delta d(M-H)$) is 0.15 Å, whereas in $[\text{Fe}(\text{PPh}_3)(\text{CO})_2(\mu-H)(\mu-pdt)\text{Ni(dppe)}]\text{BF}_4$, the Ni-H bond is 0.4 Å longer than the Fe-H bond (Table 7.3). Although terminal hydrides are invoked for the catalytic mechanism of [FeFe]-hydrogenases, catalytically active $\mu$-hydrido diiron dithiolates have been reported.\textsuperscript{24} Such $\mu$-hydride catalysts are mechanistically and structurally more closely related to the [NiFe]-hydrogenases.\textsuperscript{20} The high asymmetry of the Ni-H-Fe linkage in the present cases suggests that even in the [NiFe] enzymes, the Fe-H bond may have significant character as a terminal (vs bridging) hydride.

### Table 7.3. Fe-H and Ni-H Bond Distances in Diiron and Nickel-Iron Dithiolato Hydrides.

<table>
<thead>
<tr>
<th>Hydride</th>
<th>M-H Distances (Å)</th>
<th>$\Delta d(M-H)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$[\text{HFe}_2(\text{pdt})(\text{CO})_4(\text{dppe})]\text{BF}_4$\textsuperscript{50}</td>
<td>(dppe)$_{\text{dibasal}}$(CO)Fe-H: 1.627(3) (CO)$_3$Fe-H: 1.640(4)</td>
<td>0.013</td>
</tr>
<tr>
<td>$[\text{HFe}_2(\text{pdt})(\text{CO})_4(\text{NHC-chelate})]\text{BF}_4$\textsuperscript{51}</td>
<td>(NHC)$_2$(CO)Fe-H: 1.710 (CO)$_3$Fe-H: 1.562 Å</td>
<td>0.15</td>
</tr>
<tr>
<td>$\text{unsym-}[\text{HFe}_2(\text{SC}_2\text{H}_4\text{PM}_2)_2(\text{CO})_4]\text{BF}_4$\textsuperscript{52}</td>
<td>(Palkyl)$_{\text{apical}}$(CO)$<em>2$Fe-H: 1.59(1) (Palkyl)$</em>{\text{basal}}$(CO)$_2$Fe-H: 1.74(1)</td>
<td>0.15</td>
</tr>
<tr>
<td>$\text{cis-}[\text{HFe}_2(\text{pdt})(\text{CN})(\text{CO})_4(\text{PMe}_2)]\text{BF}_4$\textsuperscript{53}</td>
<td>(PMe)$_{\text{basal}}$(CO)$<em>2$Fe-H: 1.63(1) (NC)$</em>{\text{basal}}$(CO)$_2$Fe-H: 1.70(1)</td>
<td>0.07</td>
</tr>
<tr>
<td>$[\text{HFe}(\text{CO})_3(\text{pdt})\text{Ni(dppe)}]\text{BF}_4$\textsuperscript{23}</td>
<td>Fe-H: 1.46(6) Ni-H: 1.64(6)</td>
<td>0.18</td>
</tr>
<tr>
<td>$[\text{HFe}(\text{CO})_2(\text{PPh}_3)(\text{pdt})\text{Ni(dppe)}]\text{BF}_4$</td>
<td>Fe-H: 1.49(3) Ni-H: 1.89(3)</td>
<td>0.4</td>
</tr>
</tbody>
</table>
Scheme 7.5. Catalytic cycle proposed for proton reduction by [HFe(CO)$_2$(PPh$_3$)(pdt)Ni(dppe)]BF$_4$.

Catalysis by these complexes operate via the sequence of reactions in Scheme 7.5. The voltammetric responses indicate that the reaction is second order in protons and first order in the bimetallic complex. Unlike catalysis by phosphine-substituted diiron dithiolates as models for the [FeFe]-hydrogenases, the pK$_a$ of acid required appears to match well with the pK$_a$ of the complex, suggesting that a more complicated mechanism is not operating. In the regime where the catalytic rate is affected by [H$^+$], protonation is probably rate determining, and the facility of this process
depends on the basicity of the metal center and steric effects. Such acid-dependent rate constants are not effective benchmarks for catalytic efficiency, and are dependent on the proton source. More useful are the relative rates for catalysis in the acid-independent regime obtained at high $[H^+]$, and the overpotential (Table 7.4). This high $[H^+]$ regime resembles enzymatic conditions where protons are efficiently provided to the NiFe center. The rate-limiting step is assumed to be the dissociation of $H_2$ from the doubly protonated, one-electron reduced catalyst. The overpotential with the present phosphine-substituted nickel-iron hydrides is $\sim 0.4$ V; However, slightly weaker acids were not explored catalytically.

Table 7.4. Selected Catalytic Properties of the Hydrides. *Conditions*: see experimental.

<table>
<thead>
<tr>
<th>Catalyst</th>
<th>$E_{\text{cat}}^a$ (V, vs Fc$^{0+/+}$)</th>
<th>Rate$^b$ (s$^{-1}$)</th>
<th>Overpotential$^c$ (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$[1H]^+$</td>
<td>-1.20</td>
<td>20</td>
<td>0.31</td>
</tr>
<tr>
<td>$[2H]^+$</td>
<td>-1.32</td>
<td>50</td>
<td>0.43</td>
</tr>
<tr>
<td>$[3H]^+$</td>
<td>-1.30</td>
<td>50</td>
<td>0.41</td>
</tr>
<tr>
<td>$[4H]^+$</td>
<td>$\sim$-1.3</td>
<td>50</td>
<td>$\sim$0.4</td>
</tr>
<tr>
<td>$[4H_2]^{2+}$</td>
<td>-1.15</td>
<td>50</td>
<td>0.26</td>
</tr>
<tr>
<td>[NiFe]-hydrogenase (A. vinosum)</td>
<td>-0.345 (pH 7.4)$^d$</td>
<td>$\sim$500$^e$</td>
<td>$\sim$0$^59$</td>
</tr>
</tbody>
</table>

$^a$For $[1H]^+$ - $[4H]^+$, potential at half-height for an acid-independent voltammogram.$^{41}$ Potentials can be corrected from Fc$^{+/0}$ (MeCN) to NHE using the relation $E_{\text{NHE}} = E_{\text{Fc}} + 0.717$ V.$^{60}$

$^b$Estimated from acid-independent region of $i_\text{c}/i_\text{p}$ vs [acid] plot.

$^c$Calculated using the relation Overpotential = $E_{\text{cat}}^c + 0.89$ V.$^{41}$

$^d$Onset potential, 40 °C, vs NHE.$^61$

$^e$Lower limit estimate, 30 °C, pH 6.
Pyridylphosphines confer fascinating properties to certain catalysts, and offered the possibility of facilitating proton transfer. When installed on diiron dithiolates, pyridyl phosphines undergo $N$-protonation, which leads to milder $E_{\text{cat}}$, an effect attributable to an electrostatic influence. We see similar effects in this report: the overpotential decreases by $\sim 140$ mV for the dication $[4\text{H}_2]^{2+}$ vs $[3\text{H}]^+$ (Table 7.4). The amine also accelerates the deprotonation of the $\mu$-hydride, which is otherwise slow. The pathway by which this deprotonation occurs is suggested by the observation that the hydride exists in equilibrium with the pyridinium salt. In the protein, proton relay is proposed to occur via the terminal thiolate ligands.

The current nickel-iron hydrides employ phosphine ligands to modulate the electronic environment of the HNiFe(SR)$_2$ core, which is in contrast to Nature’s selection of cyanide and cysteinate ligands. However, our goal is to correctly understand the mechanism of [NiFe]-hydrogenase by focusing on the hydride ligand reactivity. In this report we show that replacement of a single CO ligand in $[\text{HFe(CO)}_3(\text{pdt})\text{Ni(dppe)}]\text{BF}_4$ ($[1\text{H}]\text{BF}_4$) with the modest donor ligand PPh$_3$ has a mild affect on the acidity of the hydride ligand, shifting the $pK_a$ by about 4 units (6.7 kcal/mol). However, the reduction potential of the hydride complex is shifted by only 200 mV (4.6 kcal/mol). As the $pK_a$ of the [HNiFe]$^+$ species is that required for catalysis, a correctly matched acid for $[3\text{H}]^+$ ($pK_a \sim 14.5 \ \text{E}^o = -1.02 \ \text{V vs Fc}^{0/+}$) would yield a reduced overpotential of 300 mV; however, such acids are slow to protonate 3. Thermodynamic cycles provide that the metal hydride bond free energies ($\Delta G_{\text{H\text{-}}}$) are quite weak, and are properly poised for efficient catalysis.
The present paper describes examples of hydride complexes that sustain at least quasi-reversible redox events. Reversible redox has been observed in other bimetallic hydrides but is rare for monometallic hydrides where redox changes the $pK_a$ by many orders of magnitude, which precludes reversibility. Nature’s selection of bimetallic active sites in the hydrogenases thus provides a way to soften the effect of redox on the acid-base properties of the hydride. One-electron oxidation of these $\text{HN}^{\text{III}}\text{Fe}^{\text{II}}$ “Ni-R” models would provide a $\text{HN}^{\text{III}}\text{Fe}^{\text{II}}$ “Ni-C” model, which is a well-characterized catalytically active state, proposed to bind $\text{H}_2$. However, in this generation of complexes the potentials are severe, and the oxidation couples irreversible. In Nature, $\text{[NiFe]}$-hydrogenase predominantly functions as a hydrogen oxidation catalyst, and proton reduction is inhibited by trace quantities of $\text{H}_2$. The present complexes only catalyze proton reduction, are not inhibited by $\text{H}_2$, and do not have a $\text{HN}^{\text{III}}\text{Fe}^{\text{II}}$ center in the catalytic cycle, thus appear to operate by a different mechanism than Nature. However, the exact mechanism of Nature is unclear, and may use a $\text{HN}^{\text{III}}\text{Fe}^{\text{II}}$ center solely for $\text{H}_2$ oxidation.
Experimental

Unless otherwise indicated, reactions were conducted using Schlenk and cannula-filtration techniques at reduced temperatures. Solvents for syntheses were HPLC-grade and further purified using an alumina filtration system (Glasscontour Co., Irvine, CA), NMR solvents were either dried with CaH₂ and stored under nitrogen over activated 3 Å molecular sieves or purchased as ampoules from Cambridge Isotope Laboratories. Diiron nonacarbonyl, tetrafluoroboric acid in diethyl ether, triphenylphosphite, triphenylphosphine, and trifluoroacetic acid were purchased from Aldrich and used as received. Tetrabutylammonium hexafluorophosphate (Aldrich) was recrystallized from methylene chloride and hexane. NMR spectra were recorded at room temperature on a Varian Mercury spectrometer. NMR chemical shifts are quoted in ppm; spectra are referenced to TMS for ¹H and 85% H₃PO₄ for ³¹P{¹H} spectra. EPR simulations performed by Mark Nilges: (Nilges, M. J.; Matteson, K. L. SIMPOW6: A Software Package for the Simulations of ESR Powder-type Spectra in ESR Spectroscopy in Membrane Biophysics).

NiFe(pdt)(CO)₃(dppe) (1). To a 500-mL round bottomed Schlenk flask with stir bar was added 2.25 g (4.01 mmol) of Ni(pdt)(dppe), 1.52 g (4.19 mmol) of Fe₂(CO)₉, and 40 mL of CH₂Cl₂. After stirring the red slurry for 6 h, solvent was removed under vacuum, and the red residue was washed with four 30 mL portions of MeCN to remove Fe₂(pdt)(CO)₆ and Fe(CO)₅. The remaining red-green solid was extracted into ca. 5 mL of CH₂Cl₂, and this extract was filtered through 4 x 12 cm plug of silica gel, rinsing with CH₂Cl₂. A mobile green product eluted, leaving unreacted Ni(pdt)(dppe). The green
solution was then concentrated and then diluted with 100 mL of hexane to precipitate green microcrystals. Yield: 0.745 g (1.06 mmol, 27%). $^1$H NMR (500 MHz, CD$_2$Cl$_2$, 20 °C): δ 1.3 (1H, qt, axial proton - (SCH$_2$)$_2$CH$_2$), 1.85 (1H, dt, equatorial proton- (SCH$_2$)$_2$CH$_2$), 1.9 (2H, t, axial protons- (SCH$_2$)$_2$CH$_2$), 2.5 (2H, dt, equatorial protons- (SCH$_2$)$_2$CH$_2$), 2.2 (4H, m, dppe), 7.4 - 7.7 (20H, m, dppe). $^{31}$P($^1$H) NMR (202 MHz, CD$_2$Cl$_2$): δ 63.6. FT-IR (CH$_2$Cl$_2$): ν$_{CO}$ = 2028, 1952 cm$^{-1}$.

$$[(CO)_3Fe(S_2C_3H_6)Ni(dppe)]BF_4 ([1]BF_4).$$ To a 10 mL Schlenk flask 1 (7 mg, 0.01 mmol) and FcBF$_4$ (2.5 mg, 0.01 mmol) were dissolved in a mixture of 8 mL of CH$_2$Cl$_2$ and 2 mL of THF. This solution was then transferred via air-tight syringe to a sealable quartz EPR tube. The solution was then frozen and the tube was flame-sealed under vacuum.

$$[(CO)_3Fe(S_2C_3H_6)(\mu-H)Ni(dppe)]BF_4 ([1H]BF_4).$$ To a 100-mL round bottom Schlenk flask with magnetic stir bar was added 1.25 g (1.78 mmol) of 1 and 10 mL of CH$_2$Cl$_2$. To this green solution was added 0.30 mL (2.078 mmol) of HBF$_4$ Et$_2$O, immediately producing a red solution. The solution was then concentrated under vacuum and the product was precipitated by the addition of ~20 mL of Et$_2$O. Recrystallization from CH$_2$Cl$_2$/Et$_2$O afforded red microcrystals. Yield: 1.35 g (1.71 mmol, 96%). $^1$H NMR (500 MHz, CD$_2$Cl$_2$, 20 °C): δ -3.53 (1H, tt: J$_{PH}$ = 6, J$_{HH}$ = 0.6 Hz correlates with signal at δ2.5, HNiFe), 1.57 (1H, qt, axial proton – (SCH$_2$)$_2$CH$_2$), 2.0 (2H, t, axial protons- (SCH$_2$)$_2$CH$_2$), 2.5 (2H, d, equatorial protons- (SCH$_2$)$_2$CH$_2$), 2.65 (1H, dt, equatorial proton- (SCH$_2$)$_2$CH$_2$), 2.78 (4H, m, dppe), 7.5 - 8.0 (20H, m, dppe- (C$_6$H$_5$)$_4$).

$^{31}$P NMR (202 MHz, CD$_2$Cl$_2$): δ 71. FT-IR (CH$_2$Cl$_2$): ν$_{CO}$ = 2082, 2024 cm$^{-1}$. Anal. Calcd for C$_{32}$H$_{31}$BF$_4$FeNiO$_3$P$_2$S$_2$ (found): C, 50.10 (50.16); H, 4.55 (4.75). Single crystals of
[1H]BF₄CH₂Cl₂ were obtained from CH₂Cl₂-ether.

**Reaction of 1 with B(C₆F₅)₃ and H₂.** Under an inert atmosphere 4.0 mg B(C₆F₅)₃ (0.0078 mmol) and 6.6 mg (0.0094 mmol) of 1 was dissolved with ~0.5 mL CD₂Cl₂ in a J. Young NMR tube. The ¹H and ³¹P{¹H} NMR spectra were recorded initially showing ~16% conversion to the hydride [1H]+, which we attribute to the action of (H₂O)B(C₆F₅)₃. Spectra recorded after 1 h verified that no change occurred. The sample was then frozen and put under an H₂ atmosphere. ¹H and ³¹P{¹H} NMR spectra showed nearly complete conversion to [1H]+.

[(PhO)₃P(CO)₂Fe(S₂C₃H₆)(μ-H)Ni(dppe)]BF₄, ([H]BF₄). To a 250-mL round bottomed Schlenk flask was dissolved 1.245 g (1.58 mmol) of [2H]BF₄ in 40 mL of CH₂Cl₂. To this solution, 414 µL (1.58 mmol) of P(OPh)₃ was added and the mixture was stirred for 6 h at 35 °C. Solvent was then removed under vacuum, and the product was extracted into a small amount of warm EtOH. Cooling of this extract to ~78 °C precipitated the red product. This process was repeated 3x followed by recrystallization of the material from an EtOH solution by the addition of hexane. Yield: 1.06 g (1.0 mmol, 62%). ¹H NMR (400 MHz, CD₂Cl₂): δ 6.6-8.0 (35H, m, Ph's), 2.88-1.1 (10H, m, dppe PPh₂CH₂CH₂PPh₂; pdt SCH₂CH₂CH₂S), -3.45 (1H, dt, Ni(m-H)Fe). ³¹P{¹H} NMR (161 MHz, CD₂Cl₂): δ 161 (s, P(OPh)₃), 65.8 (br, dppe). FT-IR (CH₂Cl₂): νCO = 2031, 1981 cm⁻¹.

[Ph₃P(CO)₂Fe(S₂C₃H₆)(μ-H)Ni(dppe)]BF₄, ([3H]BF₄). **Method A.** To a 100-mL round bottomed Schlenk flask fitted with magnetic stir bar was added 0.126 g (0.180 mmol) of NiFe(pdt)(CO)₃(dppe) from the glove-box and dissolved in 25 mL of CH₂Cl₂. To this green solution 0.077 g of (0.220 mmol) [HPPh₃]BF₄ and 0.105 g (0.400 mmol) of
PPh₃ were added. After 3.5 h photolysis with a Spectroline black light lamp (365 nm), the FT-IR spectrum showed complete consumption of [1H]BF₄. The solution was then concentrated under vacuum and addition of 40 mL of Et₂O provided a red precipitate. The product was washed with 3 x 10 mL of Et₂O, and dried under vacuum. Yield: 0.113 g (0.121 mmol, 67%).

**Method B.** To a 250-mL round bottomed flask fitted with magnetic stir bar was prepared a solution of 0.262 g (0.333 mmol) of [1H]BF₄ in 50 mL of THF. To this solution 0.98 g (3.74 mmol, ~10x) of PPh₃ was added. After stirring the solution for 2 h at 40 °C, the solvent was removed in vacuum yielding an red colored oil, which was washed with four 20 mL portions of hexane. The remaining oil was redissolved in 30 mL of CH₂Cl₂, and the microcrystalline product was precipitated by addition of 100 mL of hexane. Yield: 0.235 g (0.230 mmol, 70%).

**1H NMR** (400 MHz, CD₂Cl₂): δ 6.8-7.9 (35H, m, Ph’s), 2.7 (4H, m, dppe PPh₂CH₂CH₂PPh₂), 2.7-1.4 (6H, m, SCH₂CH₂CH₂S), d -3.08 (1H, dt, Ni(m-H)Fe).

**31P{¹H} NMR** (161 MHz, CD₂Cl₂): δ 69.5 (s, PPh₃), 65.8 (br, dppe). FT-IR (CH₂Cl₂): νCO = 2016, 1964 cm⁻¹. Anal. Calcd for C₄₉H₄₆BF₄FeNiO₂P₃S₂ (found): C, 57.40 (57.48); H, 4.52 (4.36).

**[(Ph₃P)(CO)₂Fe(S₂C₃H₈)Ni(dppe)] (3).** In a 100-mL round-bottomed Schlenk flask was dissolved 0.110 g (0.107 mmol) of [3H]BF₄ in 5 mL of CH₂Cl₂ and 2 mL of MeOH. To this red solution 5.8 mg (0.107 mmol) of NaOMe was added. After stirring for 3 h, the reaction mixture was evaporated under vacuum. The residue was washed with H₂O (3 x 5 mL) and MeOH (3 x 5 mL), and the green powder was dried under vacuum. Yield: 74 mg (0.079 mmol, 74%).

**1H NMR** (400 MHz, CD₂Cl₂): δ 7.2-8.0 (35H, m, Ph’s), 2.1 (4H, m, dppe PPh₂CH₂CH₂PPh₂), 1.8-0.8 (6H, m, SCH₂CH₂CH₂S).
NMR (161 MHz, CD$_2$Cl$_2$): δ 55 (t, PPh$_3$), 45, 77 (br, dppe). FT-IR (CH$_2$Cl$_2$): ν$_{CO}$ 1971, 1916 cm$^{-1}$. Anal. Calcd for C$_{49}$H$_{45}$FeNiO$_2$P$_3$S$_2$ (found): C, 62.10 (62.78); H, 4.78 (4.45).

[(Ph$_2$PyP)(CO)$_2$Fe(S$_2$C$_3$H$_6$)(μ-H)Ni(dppe)]BF$_4$ ([4H]BF$_4$) and its Protonation.
To a 250-mL round bottom Schlenk flask was added 0.400 g [1H]BF$_4$ (0.508 mmol) and dissolved in 30 mL THF. To this solution 0.150 g (0.570 mmol, ~1.1x) of Ph$_2$PyP was added. After stirring the solution for 3 h at 40 °C, solvent was concentrated and the product was precipitated by addition of Et$_2$O. The red solid was recrystallized from 15 mL of acetone by the addition of 60 mL of EtOAc. The red microcrystalline material was dried under vacuum and stored in the glovebox. Yield: 0.309 g (0.302 mmol, 59%). $^1$H NMR (500 MHz, CD$_2$Cl$_2$): δ 6.8-7.9, 8.8 (35H, m, PPh$_2$Py, dppe Ph$_2$'s), δ 2.62 (4H, m, dppe PPh$_2$CH$_2$CH$_2$PPh$_2$), δ 2.53-1.49 (6H, m, SCH$_2$CH$_2$CH$_2$S), δ -3.19 (1H, dt, Ni(m-H)Fe). $^{31}$P{$^1$H} NMR (202 MHz, CD$_2$Cl$_2$): δ 73.7 (s, PPh$_2$Py), 65.7 (br, dppe). FT-IR (CH$_2$Cl$_2$): ν$_{CO}$ = 2022, 1971 cm$^{-1}$. Samples of [4H$_2$]$^{2+}$ were generated by protonation of degassed CH$_2$Cl$_2$ solutions with ~5 equiv of CF$_3$CO$_2$H, however the resulting dication was observed to decompose over the course of minutes. The decomposition mixture consisted of [1H]BF$_4$, Ni(pdt)(dppe), and [HPPh$_2$Py]$^{+}$ as indicated by $^{31}$P{$^1$H} and $^1$H NMR spectra. $^1$H NMR (400 MHz, CD$_2$Cl$_2$): δ 7.2-7.8 (m), 8.0 (t), 8.3 (t), 8.8 (d) (35H, PPh$_2$Py, dppe Ph$_2$'s), δ 2.6 (4H, m, dppe PPh$_2$CH$_2$CH$_2$PPh$_2$), δ 2.7-1.5 (6H, m, SCH$_2$CH$_2$CH$_2$S), δ -3.14 (1H, dt, Ni(m-H)Fe). $^{31}$P{$^1$H} NMR (161 MHz, CD$_2$Cl$_2$): δ 79 (s, PPh$_2$Py), 66 (br, dppe). FT-IR (CH$_2$Cl$_2$): ν$_{CO}$ = 2032, 1982 cm$^{-1}$.

Preparation of [H(dppe)(CO)Fe(S$_2$C$_3$H$_6$)Ni(dmpe)]BF$_4$. To a 100 mL round bottom Schlenk flask 50 mg [2H]BF$_4$ was dissolved in 10 mL of THF. To this solution 8 μL dmpe was added. The solution was allowed to sit at room temperature for ca. 2 hr,
then heated with a warm water bath. The reaction progress was monitored by FT-IR spectroscopy. Initially the loss of ν\textsubscript{CO} peaks for [2H]BF\textsubscript{4} are observed with concomitant growth of 2. Then, 2 converts slowly to a single ν\textsubscript{CO} peak at 1934 cm\textsuperscript{-1}. Upon completion, the solution is concentrated and addition of 20 mL hexane produces a brown-red solid. \textsuperscript{1}H NMR (500 MHz, CD\textsubscript{2}Cl\textsubscript{2}): δ – 4.2 (J\textsubscript{PH} = ~75 Hz). \textsuperscript{31}P{\textsuperscript{1}H} NMR (161 MHz, CD\textsubscript{2}Cl\textsubscript{2}): δ 86 (s, Fe-dppe), 51 (s, Ni-dmpe). FT-IR (CH\textsubscript{2}Cl\textsubscript{2}): ν\textsubscript{CO} = 1934 cm\textsuperscript{-1}.

Preparation of [(dppe)(CO)\textsubscript{2}Fe(S\textsubscript{2}C\textsubscript{3}H\textsubscript{6})Ni(dppe)](BF\textsubscript{4})\textsubscript{2}, [5(CO)\textsubscript{2}](BF\textsubscript{4})\textsubscript{2}. To a 100 mL round bottom Schlenk flask was added 126 mg 1 (0.18 mmol), 98 mg FcBF\textsubscript{4} (0.38 mmol), and 77 mg (0.198 mmol) dppe. The solids were then dissolved in 15 mL CH\textsubscript{2}Cl\textsubscript{2} and let sit for 30 min, upon which the IR spectrum indicated a complete reaction. The solution was then concentrated and addition of ca 30 mL Et\textsubscript{2}O resulted in the precipitation of an orange microcrystalline solid. Yield: 141 mg (65%). \textsuperscript{31}P{\textsuperscript{1}H} NMR (161 MHz, CD\textsubscript{2}Cl\textsubscript{2}): δ 59.5, 58.5. FT-IR (CH\textsubscript{2}Cl\textsubscript{2}): ν\textsubscript{CO} 1986 (s), 2051 (w) cm\textsuperscript{-1}.

Attempted Reduction of [5(CO)\textsubscript{2}](BF\textsubscript{4})\textsubscript{2}. To a 10 mL round bottom Schlenk flask was added 6.3 mg [5(CO)\textsubscript{2}](BF\textsubscript{4})\textsubscript{2} (0.005 mmol) and dissolved in 2 mL CH\textsubscript{2}Cl\textsubscript{2}. To this solution was added via cannula transfer 4 mg CoCp\textsubscript{2} (0.01 mmol) dissolved in 2 mL CH\textsubscript{2}Cl\textsubscript{2}. The reaction was monitored by FT-IR spectroscopy.

pK\textsubscript{a} Determination of [3H]BF\textsubscript{4}. In a J. Young NMR tube, ~0.8 mL dry degassed PhCN was added to 10.0 mg (0.0098 mmol) of [3H]BF\textsubscript{4}. To this solution 197 µL of a freshly prepared solution of 0.197 M of 4-methoxypyridine in PhCN (pK\textsubscript{a}\textsubscript{MeCN} = 14.23) was added. The \textsuperscript{31}P NMR spectrum was then recorded after 1, 2, 3, and 5 h, each time showing a 2:1 ratio of 3:[3H]\textsuperscript{+} with increasing decomposition. The integrations was determined from integration of the respective PPh\textsubscript{3} signals (pK\textsubscript{a} ~14.9).
**pKₐ Determination of [1H]BF₄.** In a 25 mL Schlenk flask, ~4 mL dry degassed PhCN was added to 5.8 mg (0.0073 mmol) of [1H]BF₄. To this solution 27.6 µL of a freshly prepared solution of 0.5 M of aniline in PhCN (pKₐMeCN = 10.7) was added. The FT-IR spectrum was then recorded after 3, 8 and 18 h, each time showing a 1:1 ratio of 1:[1H]⁺.

**H/D Exchange of [3H]BF₄ with D₂O.** Under an inert atmosphere 4.3 mg (0.0042 mmol) of [3H]BF₄ was dissolved with ~0.5 mL d⁶-acetone (ampoule, Cambridge) in a J. Young NMR tube. A ¹H NMR spectrum was recorded for t = 0, then 10 mL (0.56 mmol) D₂O was added (in air) and subsequent scans were collected by at ~2 min intervals. The first-order decay plot was generated by integration of the µ-H signal (δ –3.08) against a normalized phenyl region. After the complete disappearance of the hydride signal for [3H]BF₄, a ³¹P{¹H} spectrum was recorded verifying the presence of the deuteride complex ([3D]BF₄) with no decomposition.

**H/D Exchange of [4H]BF₄ with D₂O.** In a J. Young NMR tube was prepared a solution of 7.2 mg (0.00705 mmol) of [4H]BF₄ in ~0.5 mL d⁶-acetone. The ¹H NMR spectrum was recorded for t = 0. The sample was then frozen, 10 mL (0.56 mmol) D₂O was added, and the sample tube evacuated and then thawed. A ¹H NMR spectrum recorded ~ 5 min after thawing showed nearly complete consumption of the hydride signal (δ –3.19). A ³¹P{¹H} spectrum was recorded verifying the presence of the deuteride complex ([4D]BF₄) with no decomposition.

**Kinetics of Deprotonation of [3H]BF₄ with NEt₃.** In a J. Young NMR tube, a solution of 5.4 mg (0.0052 mmol) of [3H]BF₄ in ~0.5 mL CD₂Cl₂ was treated with 12 µL (0.086 mmol) of NEt₃ added by syringe. The tube was then sealed, and ¹H and ³¹P{¹H}
NMR spectra were recorded. The first-order decay plot was constructed from $^{31}$P($^1$H) NMR spectra, as the ratio of $3/[^3$H]$\text{BF}_4$ could be readily determined by integration of the respective PPh$_3$ signals. The $^1$H NMR data confirms the pseudo-first order behavior.

**Electrochemistry, General Considerations.** As the nickel-iron hydrides presented in this paper degrade over the course of minutes in MeCN solution, electrochemistry was mainly performed on CH$_2$Cl$_2$ solutions. Cyclic voltammetry experiments were carried out in a 20-mL one compartment glass cell with tight-fitting Teflon lid with tight-fitting three electrodes and nitrogen gas inlet, interfaced with a BAS-100 Electrochemical Analyzer. The working electrode was a glassy carbon disk (0.3 cm in diameter). A silver wire was used as a pseudo-reference electrode, and the counter electrode was a Pt wire. The electrolyte was 0.1 M Bu$_4$NPF$_6$ in CH$_2$Cl$_2$. Ferrocene was added as an internal reference and cyclic voltammograms were each referenced to this Fc$^{0/+}$ couple (0.00 V). iR compensation was applied: solutions were pulsed prior to each scan to determine the cell resistance, this compensation was applied to the subsequent voltammagram. Electrodes were cleaned between scans by polishing with alumina.

**Cyclic Voltammetry of [2H]$\text{BF}_4$.** A solution of 2.6 mg (0.00243 mmol) of [2H]$\text{BF}_4$ in 5 mL of CH$_2$Cl$_2$ was prepared in the CV cell and was treated with successive aliquots (18 µL, 2 equiv) of a freshly prepared solution of 0.268 M CF$_3$CO$_2$H in CH$_2$Cl$_2$ solution. Cyclic voltammograms were recorded at 100 mV/s. Studies on [3H]$\text{BF}_4$ (3.3 mg in 5 mL CH$_2$Cl$_2$) and [3H]$\text{BF}_4$ (4.9 mg in 5 mL) were done similarly.

**Cyclic Voltammetry of [2H]$\text{BF}_4$ in MeCN.** A solution of 5.9 mg (0.0055 mmol) of 2 in 3.0 mL of MeCN (1.7 mM) with 0.1 M [NBu$_4$]PF$_6$ was prepared in the CV cell. Cyclic
voltammograms were then recorded between 100 – 1000 mV/s. From these data the diffusion constant (Do) was calculated to be 2.9 \times 10^{-6} \text{ cm}^2\text{s}^{-1}.

**Cyclic Voltammetry of 2 in MeCN.** A solution of 5.1 mg (0.0052 mmol) of 2 in 2.7 mL of MeCN and 0.3 mL CH₂Cl₂ (for solubility) (1.683 mM) with 0.1 M [NBU₄]PF₆ was prepared in the CV cell. Cyclic voltammograms were then recorded between 100 – 1000 mV/s. From these data the diffusion constant (Do) was calculated to be 3.5 \times 10^{-6} \text{ cm}^2\text{s}^{-1}.

**Cyclic Voltammetry of [3H]BF₄ in MeCN.** A solution of 6.8 mg (0.0066 mmol) of [3H]BF₄ in 5 mL of MeCN was prepared in the CV cell and was treated with successive aliquots (23 \mu L, 3 equiv) of a freshly prepared solution of 0.857 M CF₃CO₂H in MeCN. Cyclic voltammograms were recorded at 100 mV/s.

**Cyclic Voltammetry of [5(CO)₂](BF₄)₂ in MeCN.** A solution of 6.9 mg (0.0056 mmol) of [5(CO)₂](BF₄)₂ in 5 mL of MeCN with 0.1 M [NBU₄]PF₆ was prepared in the CV cell. A Cyclic voltammogram was recorded at 100 mV/s, and displayed a single irreversible reduction event at E_{1/2} = -0.963 V vs Fe^{0/+}.

**Determination of order of with respect to [4H]BF₄.** A solution of 145 \mu L CF₃CO₂H in 5 mL CH₂Cl₂ was prepared in the CV cell and treated with successive amounts of solid [4H]BF₄. The result of this titration assigns catalysis as first-order with respect to [4H]BF₄.

**Details of H₂ separation by Gas Chromatography.** An Agilent GC-6890 gas chromatograph (GC) equipped with a 30 m, 0.53 \mu m OD, Agilent HP-Molesieve column and a G1532-60720 thermal conductivity detector (TCD) was chosen for separating and analyzing the evolved H₂. GC-TCD systems with this configuration have been shown
previously in the literature to be effective for the separation and accurate measurement of the concentration of H₂.\textsuperscript{73,74} Additionally, the GC-TCD system had a pneumatically actuated, computer controlled sampling valve with a 1 mL sample loop. The 1 mL gas samples were then directed into the split/splitless column inlet before finally entering the column for separation and detection. The reaction flask was purged with Ar, which was continued into the GC as the carrier gas. The thermal conductivity ratio for H₂:Ar was 187:18.

**Bulk Electrolysis of [1H]BF\textsubscript{4} in MeCN.** In a \textasciitilde100 mL bulk electrolysis cell 31 mg (0.039 mmol) was dissolved in 30 mL CH\textsubscript{3}CN with 0.1 M [NBu\textsubscript{4}]PF\textsubscript{6} and 200 \textmu L CF\textsubscript{3}CO\textsubscript{2}H (87 mM). The working electrode was a Hg pool; reference a Ag|AgCl electrode; counter a platinum wire. A cyclic voltammogram was recorded with Hg pool as the working electrode prior to electrolysis, displaying an onset of reduction \textasciitilde – 700 mV vs Ag|AgCl. During electrolysis the potentiostat was maintained at – 800 mV vs the Ag|AgCl reference. A flow of argon at 50 mL/min was purged through the cell and then into the GC as described above.

**References:**


Chapter 8

New Routes to Nickel-Iron Dithiolate Hydrides

Introduction

Recent inorganic modeling efforts of the [NiFe]-hydrogenases have provided (µ-H)Ni(µ-SR)₂Fe compounds that are both structural and functional mimics of Nature’s most-prevalent hydrogen processing enzyme. Originally obtained in low-yields from (BDA)Fe(CO)₃, (CO)₃Fe(pdt)Ni(dppe) (1) was described as an unstable brown solid although the compound was characterized crystallographically (pdt = 1,3-propanedithiolate, dppe = Ph₂PCH₂CH₂PPh₂, BDA = benzylideneacetone).¹² After optimization, 1 can be prepared in up to 35% yield from Fe₂(CO)₉ and is a robust, easily recrystallized oxygen-sensitive green solid.³ Upon protonation, [1H]⁺ is isolated as a stable red microcrystalline solid, and is a catalyst for the reduction of protons to H₂ at mild overpotentials (~ 400 mV).

In view of the newly recognized robustness of its precursors,¹ and profound catalytic properties of [1H]⁺, we desired to modify the (CO)(L)₂Fe(µ-SR)₂Ni(L₂) framework and observe the resulting thermodynamic (acidity, E½’s, ∆G₁⁺, ∆G₂⁺) and catalytic consequences. We wished to retain the flexible coordination sphere at the Ni center, which is required for catalysis,³ thus we avoid rigid tetratentate scaffolds.⁴ The first modification was first achieved by ligand substitution at the Fe(CO)₃ center, and the resulting (CO)₂(PR₃)Fe(µ-pdt)₂(µ-H)Ni(dppe) complexes displayed a profound effect on both the acidity of the hydride and E½’s (Chapter 7). The catalytic and thermodynamic properties of monomeric nickel phosphines have been shown to be highly dependent on
bite-angle and ligand basicity.\textsuperscript{5,6} Thus, we were especially interested in the consequences of (L\textsubscript{2})Ni(SR)\textsubscript{2}-based modifications on the structural and catalytic properties of the hydride. It is widely accepted that the Ni-R and the Ni-C states of the enzyme feature hydride ligands spanning the Fe-Ni vector, and the presence of the hydride has established in the case of the Ni-C state. Because this species is described by an \( S = 1/2 \) state, it has been well examined by electron spin resonance techniques. The ENDOR and HYSCORE data are consistent with Ni\textsuperscript{III}-H and Fe\textsuperscript{II}-H distances of 1.61 and 1.72 Å, respectively.\textsuperscript{7} In contrast, the Fe-H bonds are 0.2-0.4 Å shorter than the Ni-H bonds in the \( S = 0 \) Ni\textsuperscript{III}-Fe\textsuperscript{II} models [(CO)\textsubscript{2}LFe(pdt)(\( \mu \)-H)Ni(dppe)]BF\textsubscript{4} (L = CO, PPh\textsubscript{3}).\textsuperscript{1}

To achieve these (L\textsubscript{2})Ni(SR)\textsubscript{2}-based modifications, new Ni(I)Fe(I) model complexes would need to be synthesized. In addition to Fe\textsubscript{2}(CO)\textsubscript{9} as a “Fe(CO)\textsubscript{3}” source, we revisited the “Fe(CO)\textsubscript{n}\textsuperscript{2+}” source FeI\textsubscript{2}(CO)\textsubscript{4}. This reagent has enjoyed renewed popularity since its value was demonstrated in the preparation of the hydrogenase-relevant ferrous carboxyls.\textsuperscript{8-12} Typical substitution reactions of FeI\textsubscript{2}(CO)\textsubscript{4} replace one or more CO ligands with \( 2e^- \) donors, retaining the anionic iodide ligands.\textsuperscript{13} In this chapter, we show that although Fe\textsubscript{2}(CO)\textsubscript{9} routes to provide new Ni(I)Fe(I) compounds have proven ineffective, the ferrous carbonyl I\textsubscript{2}Fe(CO)\textsubscript{4} efficiently condenses with nickel dithiolato diphosphines to give intermediates that can be converted to the targeted hydrides. Modification of the nickel center in (CO)(L)\textsubscript{2}Fe(\( \mu \)-SR)\textsubscript{2}Ni(L\textsubscript{2}) has proven a valuable target, as new model compounds demonstrate that the (SR)\textsubscript{2}Ni(L\textsubscript{2}) fragment has a powerful influence on redox potentials, and only a mild effect on hydride acidity.
Results and Discussion

Ni(SR)$_2$(diphosphine). We examined the preparation of several new examples of the general type of Ni(SR)$_2$(diphosphine), which could be used to prepare Fe-Ni derivatives. The procedures entailed the initial formation of NiCl$_2$(diphosphine), which were subsequently converted to the dithiolates using salt-forming reactions. The alkyl thiolato derivatives of nickel(II) phosphines are known to be somewhat labile,$^{14}$ but propane- and ethanedithiolates are more stable. Although some of the dichlorides exist as an equilibrium mixture of the square planar and tetrahedral geometries,$^{15}$ their dithiolato derivatives proved to be uniformly diamagnetic and hence are assumed to be square planar. We were unable to prepare propanedithiolate derivatives with wide bite-angle diphosphines: attempted syntheses of Ni(pdt)(dppf), Ni(pdt)(dppp), and Ni(pdt)(Ph$_2$Si(CH$_2$PPh$_2$)$_2$) afforded the polymer [Ni(pdt)$_n$ indicating the apparent weaker binding of ligands with larger bite angles. The diphosphine dppn, which is known to provide small chelate bite-angles ($\Theta \sim 80^\circ$) and be quite rigid,$^{16,17}$ allowed the formation of the square planar Ni(pdt)(dppn) complex. The following new derivatives were prepared in analytical purity: Ni(pdt)(dppbz), Ni(pdt)(dppn), Ni(pdt)(dppv), Ni(pdt)(dmpe), Ni(edt)(dmpe), and Ni(pdt)(dcpe).

The Dimeric [(CO)Ni(SPh)$_2$Fe(CO)$_6$]$_2$(µ-dppe)$_2$. Employing the utility of Fe$_2$(CO)$_9$ as a “Fe(CO)$_3$” source, reaction of the Ni(SR)$_2$(diphosphine)’s listed in Table 8.1 provided mixtures of several organometallic products by IR spectroscopy, and purification attempts proved ineffective. However, the reaction of Ni(SPh)$_2$(dppe) with Fe$_2$(CO)$_9$ appeared to be stable and well-behaved, but did not resemble the C$_{3v}$ IR
Crystallographic analysis indicated that the product is a dimeric [(CO)Ni(SPh)_2Fe(CO)]_2(µ-dpone) (Figure 8.1). The structure is corroborated by the $^{31}$P{^1}H NMR spectrum, which displays doublets of doublets at δ 33.9 and at 49.9 ($J_{PP} = 30, 60$ Hz) (Figure 8.2), which also indicates that the SPh ligands are equivalent and hence all-equatorial. The FT-IR ($\nu_{CO} = 1993$ (s), 1938 (m) cm$^{-1}$) resembled that reported for (CO)Ni(pdt)Fe(CO)]_2(µ-dpone). The reaction of Ni(SPh)_2(dpdbz) appeared to give a similar product to that obtained from the Ni(SPh)_2(dppe) reaction ($\nu_{CO} = 2006$(s), 1945(m) cm$^{-1}$), although is likely (CO)Ni(pdt)Fe(CO)]_2(µ-dppe). The quest for non-chelating thiolates was inspired by Nature, which uses cysteinate residues as the µ-thiolate linkage. Similar to the active site of [NiFe]-hydrogenase, the SPh$^-$ groups are equatorial.

**Figure 8.1.** Structure of Ni$_2$Fe$_2$(SPh)$_4$(CO)$_6$(dppe)$_2$, obtained by the reaction of Ni(SPh)$_2$(dppe) with Fe$_2$(CO)$_3$. As is the case with the enzyme, the thiolate groups are equatorial. Key distances (Å): Ni-Fe 2.4567(3). Fe(1)-P(2) 2.2326(5). Ni(1)-P(1) 2.1905(5). Ni(1)-S(1) 2.2633(5). Ni(1)-S(2) 2.2784(5). Fe(1)-S(1) 2.2833 (5). Fe(1)-S(2) 2.3167(6).
Figure 8.2. $^{31}$P\{$^1$H\} NMR (20 °C, 161 MHz, CD$_2$Cl$_2$) spectrum of Ni$_2$Fe$_2$(SPh)$_4$(CO)$_6$(dppe)$_2$. The doublet of doublets at δ 50 and δ 34 are assigned to the phosphorus atom coordinated to iron and nickel, respectively. The small signal at δ 57 is Ni(SPh)$_2$(dppe).
Scheme 8.1. Illustrative preparation of reduced NiFe model complexes from I$_2$Fe(CO)$_4$.

The preparation of (CO)$_3$Fe(µ-SR)$_2$Ni(diphosphine)’s via I$_2$Fe(CO)$_4$. As the Fe$_2$(CO)$_9$ method proved ineffective for the preparation of new nickel-iron models, we sought the utility of the “Fe(CO)$_n^{2+}$” fragment. Using this new approach, the one-pot reaction of FeI$_2$(CO)$_4$ and Ni(pdt)(dppe) followed by reduction with two equivalents of Cp$_2$Co provided (CO)$_3$Fe(pdt)Ni(dppe) in ~ 35% yield (Scheme 8.1). The reactions of the Ni(SR)$_2$(diphosphine) with FeI$_2$(CO)$_4$ are proposed to proceed via the intermediacy of the µ-I cations [(CO)$_3$Fe(SR)$_2$(µ-I)Ni(diphosphine)]$^+$. IR spectra of these intermediates resemble those for the corresponding µ-hydrides but are shifted to higher energies by about 20 cm$^{-1}$. The spectra are also usually more complex, probably due to some sample decomposition, reflecting the instability of these halide complexes. The action of I$_2$ with isolated Fe(CO)$_3$(pdt)Ni(dppe) in CH$_2$Cl$_2$ provide a cleaner FT-IR spectrum, and supports the (µ-I)Fe(CO)$_3$ assignment (Figure 8.3).
Table 8.1. Selected IR Data for \((\text{CO})_3\text{Fe(SR)}_2\text{Ni(diphosphine)}\) and Derivatives.

<table>
<thead>
<tr>
<th>diphosphine, thiolate</th>
<th>(v_{\text{CO}} \text{ (cm}^{-1}, \text{CH}_2\text{Cl}_2) for ([\text{CO}])_3\text{Fe(SR)}_2(\mu-\text{l})\text{Ni(P}_2))^* )</th>
<th>(v_{\text{CO}} \text{ (cm}^{-1}, \text{CH}_2\text{Cl}_2) for ([\text{CO}])_3\text{Fe(SR)}_2\text{Ni(P}_2) )</th>
<th>(v_{\text{CO}} \text{ (cm}^{-1}, \text{CH}_2\text{Cl}_2) for ([\text{CO}])_3\text{Fe(SR)}_2(\mu-H)\text{Ni(P}<em>2))^* \text{ (NMR } \delta</em>{\text{hydride}} \text{)} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>dppe, (\text{S}_2\text{C}_3\text{H}_6)</td>
<td>2096, 2054, 2023</td>
<td>monomer 2028, 1952</td>
<td>2082, 2024 ((\delta -3.53))$^3$</td>
</tr>
<tr>
<td>(SPh)$_2$</td>
<td>2082, 2032 (unstable)</td>
<td>2035, 1970 dimer$^a$: 2001, 1942</td>
<td></td>
</tr>
<tr>
<td>dppe, (\text{S}_2\text{C}_3\text{H}_6)</td>
<td>2094, 2075, 2048, 2040 (unstable)</td>
<td>2031, 1963 dimer: 1997, 1936</td>
<td></td>
</tr>
<tr>
<td>dmpe, (\text{S}_2\text{C}_3\text{H}_6)</td>
<td>2095, 2052, 2024</td>
<td>2016, 1941 (unstable)</td>
<td>2078, 2018 ((\delta -3.63))</td>
</tr>
<tr>
<td>dmpe, (\text{S}_2\text{C}_2\text{H}_4)</td>
<td>2094, 2049, 2026</td>
<td>2018, 1943 (unstable)</td>
<td>2079, 2018 ((\delta -3.71))</td>
</tr>
<tr>
<td>dppbz, (\text{S}_2\text{C}_3\text{H}_6)</td>
<td>2096, 2053, 2027</td>
<td>2030, 1954</td>
<td>2082, 2023 ((\delta -3.42))</td>
</tr>
<tr>
<td>dppe, (\text{S}_2\text{C}_3\text{H}_6)</td>
<td>2090, 2075, 2054, 2024</td>
<td>dimer$^a$: 1996, 1936</td>
<td></td>
</tr>
<tr>
<td>dcpe, (\text{S}_2\text{C}_3\text{H}_6)</td>
<td>2095, 2053, 2021</td>
<td>2014, 1940</td>
<td>2078, 2017 ((\delta -3.00))</td>
</tr>
</tbody>
</table>

$^a$Dimer" refers to derivatives assumed to be \([\text{CO}])_3\text{Fe(SR)}_2\text{Ni(})(\text{CO})_2(P_2)\). The monomeric NiFe complex was not observed for dppe.
Figure 8.3: IR spectra (CH$_2$Cl$_2$): (CO)$_3$Fe(pdt)Ni(dppe) + I$_2$ (top), intermediate [(CO)$_3$Fe(pdt)(I)Ni(dppe)]I (second), (CO)$_3$Fe(pdt)Ni(dppe) (third), and [(CO)$_3$Fe(pdt)(H)Ni(dppe)]BF$_4$ (bottom).

Addition of cobaltocene to red solutions of [(CO)$_3$Fe(SR)$_2$($\mu$-I)Ni(diphosphine)]$^+$ provided green solutions of (CO)$_3$Fe(pdt)Ni(diphosphine), identified by IR and $^{31}$P NMR spectroscopies. Unfortunately, several nickel dithiolates were not well behaved in their reactions with I$_2$Fe(CO)$_4$. The reaction of Ni(pdt)(dmpe) with FeI$_2$(CO)$_4$ produced a thermally sensitive $\mu$-I intermediate, which reduced with 2 equiv of cobaltocene to provide a greenish colored product. Low-temperature protonation of this species with HCl$\cdot$Et$_2$O produced a hydride derivative ($\delta$ -3.63, broad). However, above -78 °C the sample decomposed. Similar stability problems were encountered with Ni(edt)(dmpe).

Stable $\mu$-iodo intermediates were observed with the reactions of Ni(pdt)(dppbz), Ni(pdt)(dcpe), and Ni(edt)(dppe) with FeI$_2$(CO)$_4$, and were reduced to the corresponding
Fe(I)Ni(I) compounds with CoCp₂. (CO)₃Fe(pdt)Ni(dcpe) (2) and (CO)₃Fe(edt)Ni(dppe) (3) were isolated and characterized as their neutral Ni(I)Fe(I) compounds. Their $^{31}$P{$^1$H} NMR spectra displayed a single broad resonance, consistent with the Ni-based dynamics observed for 1. Protonation of 2 or 3 with HBF₄•Et₂O allowed the isolation of the targeted hydrides [(CO)₃Fe(pdt)(μ-H)Ni(dcpe)]BF₄ ([2H]⁺), and [(CO)₃Fe(edt)(μ-H)Ni(dppe)]BF₄ ([3H]⁺). The $^{31}$P{$^1$H} spectrum of [2H]BF₄ in CD₂Cl₂ displayed a single resonance at δ 90 (Figure 8.4), while the $^1$H NMR spectrum displayed a broad singlet at δ −3.0 (Figure 8.5). Presumably the hydride signal of [2H]BF₄ is a triplet with some degree of $J_{PH}$ coupling from the Ni(dcpe), although the best NMR spectrum obtained had a peak-width at half height of ~3 Hz, implying little to no coupling.

![Figure 8.4. $^{31}$P{$^1$H} NMR spectrum (CD₂Cl₂) of [(CO)₃Fe(H)(pdt)Ni(dcpe)]BF₄, [2H]BF₄.](image-url)
Figure 8.5. $^1$H NMR spectrum (CD$_2$Cl$_2$ solution) of [(CO)$_3$Fe(µ-H)(pdt)Ni(dcpe)][BF$_4$] ([3H]BF$_4$), showing the hydride species at δ -3.00.

The well-behaved salt [2H]BF$_4$ was also characterized crystallographically (Figure 8.6). Overall, the structure closely resembles that for [(CO)$_3$Fe(pdt)(µ-H)Ni(dppe)]BF$_4$. During the crystallographic analysis, the hydride was located and refined. The Fe-H-Ni linkage is unsymmetrical in both the dppe and dcpe cases. The Ni-Fe distance is 2.68 Å, whereas this distance is 2.61 Å in [(CO)$_3$Fe(pdt)(µ-H)Ni(dppe)]BF$_4$. 
**Figure 8.6.** Molecular structure of the cation in [(CO)$_3$Fe(pdt)(H)Ni(dcpe)]BF$_4$. Hydrogen (except the hydride ligand) atoms and BF$_4^-$ have been omitted for clarity. Key distances (Å): Ni-Fe, 2.6843(5); Ni-S(1), 2.2274(7); Ni-S(2), 2.2307(7); Fe-S(1), 2.3208(8); Fe-S(2), 2.3169(8); Ni-H, 1.90(2); Fe-H, 1.53(2).

**Redox Properties of (CO)$_3$Fe(pdt)Ni(dcpe), 2 and [2H]$^+$.** In benzonitrile solution with [NBu$_4$]PF$_6$ as electrolyte, complex 2 displays a reversible oxidation event at −0.840 V vs Fc$^{0/+}$ characterized by $\Delta E_p$ similar to the internal standard Fc$^{0/+}$ couple. At significantly more positive potentials, a second irreversible oxidation is observed (Table 8.2). As the previously reported 1 behaves similarly, and chemical oxidation of 1 by FcBF$_4$ has provided the fully-characterized mixed-valence [1]$^+$, we assign both couples to one-electron events. By comparison to 1, 2 is more readily oxidized by 300 mV.
In CH$_3$CN with [NBu$_4$]PF$_6$ as electrolyte, the protonated complex [2H]$^+$ displays a single reduction event at -1.36 V vs Fc$^{0/+}$. This couple is only shifted by 70 mV vs. the reduction of [1H]$^+$, which has previously been assigned as a one-electron process.

Table 8.2. Selected Electrochemical Properties (V vs Fc$^{0/+}$) of NiFe(pdt)(dcpe)(CO)$_3$ and NiFe(pdt)(dppe)(CO)$_3$ and the Derived Hydrides. Conditions: see Caption of Figure 8.7

<table>
<thead>
<tr>
<th></th>
<th>NiFe$^{0/+} E_{1/2}$</th>
<th>NiFe$^{+/2+} E_{1/2}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>NiFe(pdt)(dppe)(CO)$_3$, 1</td>
<td>-0.543 V</td>
<td>-0.124 V</td>
</tr>
<tr>
<td>NiFe(pdt)(dcpe)(CO)$_3$, 2</td>
<td>-0.840 V</td>
<td>+0.220 V</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th></th>
<th>[HNiFe]$^{+/0} E_{1/2}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>[HNiFe(pdt)(dppe)(CO)$_3$]$^+$, [1H]$^+$</td>
<td>-1.29 V</td>
</tr>
<tr>
<td>[HNiFe(pdt)(dcpe)(CO)$_3$]$^+$, [2H]$^+$</td>
<td>-1.36 V</td>
</tr>
</tbody>
</table>

**Electrocatalytic Properties of [HNiFe(pdt)(dcpe)(CO)$_3$]BF$_4$, [2H]$^+$.** In dichloromethane solution, the electrochemistry of [2H]$^+$ is highly responsive to the addition of the weak acid, CH$_2$ClCO$_2$H ($pK_a^\text{MeCN} = 15.3$, $E^\circ_{(HA/H_2)} = -1.05$ V) indicating the catalytic production of H$_2$ (Figure 8.7). The [2H]$^{+/0}$ reduction potential is apparently highly dependent on solvent, as in CH$_2$Cl$_2$ it appears at −1.5 V whereas in CH$_3$CN this couple was observed at −1.36 V. Insight into the catalytic mechanism is provided by the effect of [H$^+$] on $i_c/i_p$, the catalytic current ($i_c$) normalized relative to the noncatalytic current ($i_p$). This dependence is initially linear, indicating a mechanism involving protonation of a reduced reduced hydride (Figure 8.8). The catalytic activity of the Ni-Fe hydrides can be estimated by the value of $i_c/i_p$ where catalysis is independent of [H$^+$], which occurs at an $i_c/i_p$ of ~15. Under the conditions of the experiment, this value
corresponds to an acid-independent turnover rate of $\sim 50$ s$^{-1}$.\textsuperscript{1} By comparison, amine-complemented nickel phosphine complexes operate at rates up to $\sim 400$ s$^{-1}$.\textsuperscript{5,6}

**Figure 8.7.** Cyclic voltammogram of $[2H]^+$ before and after addition of varying equiv (denoted on right) of CH$_2$ClCO$_2$H. A control voltammogram in the absence of $[2H]^+$ at 100 equivalents of acid is shown with dotted red line. **Conditions:** $\sim 0.5$ mM $[2H]^+$ in CH$_2$Cl$_2$, 0.2M [NBu$_4$]PF$_6$, scan rate 0.1 V/s, glassy carbon working electrode ($d = 3.0$ mm); Ag wire pseudoreference with internal Fc standard at 0 V; Pt counter electrode. 
Figure 8.8 Dependence of $i_c/i_p$ vs equiv CH$_2$ClCO$_2$H. Conditions: See Figure 8.7.

Table 8.3. Parameters Relevant to the Electrocatalysis by [2H]$^+$ and Related Ni-Fe Hydrides. Potentials in V vs Fe$^{0+}$.

<table>
<thead>
<tr>
<th>Complex</th>
<th>$E_{cat}^a$</th>
<th>CH$_3$CN $pK_a$</th>
<th>Overpotential$^b$</th>
<th>[H$^+$]-independent rate, s$^{-1}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>[HNiFe(pdt)(dppe)(CO)$_3$]$^+$</td>
<td>-1.20</td>
<td>10.7</td>
<td>$\sim 0.31$ V</td>
<td>20</td>
</tr>
<tr>
<td>[HNiFe(pdt)(dppe)(PPh$_3$)(CO)$_2$]$^+$</td>
<td>-1.30</td>
<td>14.9</td>
<td>$\sim 0.41$ V</td>
<td></td>
</tr>
<tr>
<td>[HNiFe(pdt)(dcpe)(CO)$_3$]$^+$</td>
<td>-1.46</td>
<td>$\sim 12.7$</td>
<td>$\sim 0.41$ V</td>
<td>50</td>
</tr>
</tbody>
</table>

a) Potentials taken from potential at half-height for an acid-independent voltammogram from CH$_2$Cl$_2$ solutions.$^{18}$

b) Overpotentials calculated from $E^\circ$-$E_{cat}$, assuming negligible difference from $E^\circ$(CH$_3$CN) vs $E^\circ$(CH$_2$Cl$_2$).

**Thermodynamic Analyses.** Benzonitrile was used to estimate the $pK_a^{MeCN}$ of several complexes. The $pK_a^{MeCN}$ for [2H]$^+$ was bracketed to be between 12.3 and 13.2 based on the observation that a benzonitrile solution of 2 and one equivalent of dichloroacetic acid (Cl$_2$CHCO$_2$H $pK_a = 13.2$)$^{18,19}$ provided only trace protonation, whereas one equivalent of pyridinium tetrafluoroborate ([pyH]BF$_4$ $pK_a = 12.3$)$^{20}$ provided
complete protonation. The $pK_a^{MeCN}$ of $[1H]^+$ was previously measured as 10.7. Thus, substitution of the dppe by dcpe enhances the basicity of the NiFe center by ca. 2 $pK_a$ units. Following the Bordwell-like calculations utilized by DuBois,\textsuperscript{5,21} we can estimate the hydricity and bond dissociation free energies of the NiFe hydrides. The hydricity calculation requires the potential for the two electron oxidation of the Ni(II)Fe(II) species. Although the irreversibility of the second oxidation of the NiFe complexes is a serious source of error, even liberal errors in this potential suggest that the dicationic Ni(II)Fe(II) centers have high affinities for hydride. For example, $[\text{Ni(dppe)}_2\text{H}]^+$ has a hydricity of about 65 kcal/mol\textsuperscript{21} whereas for the bimetallic cation $[2\text{H}]^+$, the hydricity is >80 kcal/mol.

Table 8.4. Thermodynamic Parameters Relevant to NiFe(pdt)(dcpe)(CO)$_3]^+$ and Related Ni-Fe Hydrides. Potentials are in mV vs Fc$^{0+/+}$, $\Delta G_{H-}$ and $\Delta G_{H•}$ in kcal/mol.

<table>
<thead>
<tr>
<th>Complex</th>
<th>$pK_a$</th>
<th>$E_{1/2}$ $[\text{NiFe}]^{0+}$, $[\text{NiFe}]^{+2+}$</th>
<th>$\Delta G_{H-}$</th>
<th>$\Delta G_{H•}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$[\text{HNiFe(pdt)(dppe)(CO)}_3]^+$</td>
<td>10.7</td>
<td>-543, -124 (irrev)</td>
<td>79\textsuperscript{a}</td>
<td>57</td>
</tr>
<tr>
<td>$[\text{HNiFe(pdt)(dppe)(PPh}_3(\text{CO)}_2]^+$</td>
<td>14.9</td>
<td>-722, -191 (irrev)</td>
<td>79\textsuperscript{a}</td>
<td>58</td>
</tr>
<tr>
<td>$[\text{HNiFe(pdt)(dcpe)(CO)}_3]^+$</td>
<td>~12.7</td>
<td>-840, +220 (irrev)</td>
<td>83\textsuperscript{a}</td>
<td>53</td>
</tr>
</tbody>
</table>

\textsuperscript{a) Estimated assuming that $E_{1/2}$ for the irreversible $[\text{NiFe(pdt)(dxpe)(CO)}_2\text{L}]^{+2+}$ couples are accurate.}
Conclusions

The new method for preparing nickel-iron dithiolato hydrides entails addition of the nickel dithiolates to an electrophilic ferrous carbonyl center. The initially observed product is proposed to be \([(\text{CO})_3\text{Fe(SR)}_2(\mu-\text{I})\text{Ni(diphosphine)})]^+\). We independently generated this unstable intermediate by iodination of Ni(I)Fe(I) compounds. The iodo tricarbonyl would be a close relative of \([(\text{dppe})(\text{CO})\text{Fe(pdt)}(\mu-\text{Cl})\text{Ni(dppe)})]\text{BF}_4,\) which enjoys enhanced stabilization owing to the presence of the two diphosphine ligands. Consistent with its high reactivity, and in contrast to \([(\text{dppe})(\text{CO})\text{Fe(pdt)}(\mu-\text{Cl})\text{Ni(dppe)})]\text{BF}_4,\) the \(\mu\)-iodide is readily reduced by cobaltocene to give the targeted Ni(I)Fe(I) derivatives. The Ni(I)Fe(I) derivatives do not necessarily exhibit a higher affinity for \(\text{H}^+\) vs \(\text{I}^+\), but the hydrides are stable whereas the iodides are not.

This work allows comparison of \([(\text{CO})_3\text{Fe(pdt)}_2(\mu-\text{H})\text{Ni(dxpe)})]\text{BF}_4,\) for diphosphines of disparate basicities. The diphosphine dcpe is known to substantially enhance the basicity of its complexes relative to dppe. The heat of protonation of Fe(CO)\(_3\)(dcpe) (28.4 kcal/mol) is 5.2 kcal/mol greater than that for Fe(CO)\(_3\)(dppe).\(^{22}\) In the absence of entropy effects, this difference corresponds to a difference in the basicity of Fe(CO)\(_3\)(dxpe) of about 4 \(pK_a\)\(^{C2H4Cl2}\) units. We observe a difference of only 2 \(pK_a\) units (2.7 kcal/mol) in the bimetallic complex, which suggests that the dcpe vs dppe effect is attenuated in the bimetallic system. Electrochemically, benzonitrile solutions of the hydride complexes show only a minor effect (\(\Delta E_{1/2} = 70\) mV, 1.6 kcal/mol) in the \([\text{HNiFe(pdt)(dxpe)(CO)}_3]^{+}\) couple (where dxpe = dcpe and dppe). This minor effect in reduction potentials is relevant to catalysis, as the reduction potential per a specific \(pK_a\)
determines the overpotential. In theory, replacing dppe for dcpe allows for a slightly weaker acid (2 $pK_a$ units, 2.7 kcal/mol) to be reduced catalytically, whereas the change in reduction potential is thermodynamically not as great (70 mV, 1.6 kcal). However, catalytic runs were only performed in $\text{CH}_2\text{Cl}_2$, and the reduction potential therein is greatly altered ($E_{\text{cat}}$: -1.46 V (dcpe) vs -1.29 V (dppe)). Thus, it would appear that the $\Delta pK_a$ change of 2 is more than accounted for by this $\Delta E$ (0.17 V = 4 kcal/mol).

In a thermodynamic sense, the redox properties of the Ni(I)Fe(I) center are more strongly affected by the basicity of the nickel center as the $E_{1/2}$ for the [NiFe(pdt)(dcpe)(CO)$_3$]$^{+/0}$ couple is 297 mV (6.8 kcal/mol) more negative than the [NiFe(pdt)(dppe)(CO)$_3$]$^{+/0}$ couple. The finding that the basicity at Ni more strongly affects the redox potential (6.88 kcal/mol at 298K) than the basicity (2.7 kcal/mol at 298K) is consistent with redox being localized at Ni whereas protonation involves greater cooperativity.

The dcpe catalysts are marginally faster than the dppe-based systems but operate at higher overpotentials, which in turn lead to enhanced rates. The higher rates associated with the more basic Ni site are offset by the corresponding penalty in overpotential. In future studies, we hope to introduce terminal ligands on nickel and to replace some of the CO ligands on Fe with stronger donors as well. These studies will benefit from the new synthetic methods defined in this work.
Experimental Section

Typically, all reactions were conducted using Schlenk techniques at room temperature using cannula filtration techniques. Reagents were purchased from Aldrich and Strem. Solvents were HPLC-grade and dried by passing through an alumina filtration system or distilled under nitrogen over an appropriate drying agent. Ni(pdt)(dppe),\(^{24}\) Ni(SPh)\(_2\)(dppe),\(^{25}\) Ni(SPh)\(_2\)(dmpe),\(^{16}\) NiCl\(_2\)(dppbz),\(^{26}\) NiCl\(_2\)(dcpe),\(^{27}\) and (CO)\(_3\)Fe(pdt)Ni(dppe)\(^{12}\) were prepared according to modified literature procedures. Two other nickel complexes were also examined, NiCl\(_2\)(dpff)\(^{28}\) and NiCl\(_2\)(dppp).\(^{29}\) For silica gel, we used Siliaflash® P60 from Silicycle (230-400 mesh). HBF\(_4\)·Et\(_2\)O was purchased from Sigma-Aldrich and supplied as either a 1:1 molar ratio of HBF\(_4\)·Et\(_2\)O, or 51-57% HBF\(_4\) in Et\(_2\)O, or 6.91 – 7.71 M. A 2.0 M solution of HCl in Et\(_2\)O was purchased from Sigma-Aldrich. [Bu\(_4\)N]PF\(_6\) was purchased from GFS Chemicals and recrystallized multiple times from CH\(_2\)Cl\(_2\) and hexane. NMR spectra were recorded with Varian 500 MHz and Varian 400 MHz spectrometers. \(^1\)H NMR spectra were referenced to TMS from residual solvent. \(^{31}\)P\(^{\{1\}H}\) spectra were referenced to external 85% H\(_3\)PO\(_4\). FT-IR spectra were recorded on a Perkin Elmer Spectrum 100 FTIR spectrometer.

**Ni(pdt)(dmpe).** A suspension of 0.875 g (3.13 mmol) of NiCl\(_2\)(dmpe) and 0.34 g (3.13 mmol) of 1,3-propanedithiol in 10 mL of CH\(_2\)Cl\(_2\) was slowly (15 min) treated with a solution of 0.34 g (6.25 mmol) of NaOMe in 8 mL of MeOH. After the addition of the NaOMe/MeOH solution, the reaction mixture was stirred for 30 min. Solvent was removed under reduced pressure, and the residue was extracted into about 15 mL of CH\(_2\)Cl\(_2\). The reaction mixture was filtered through a pad of Celite to remove NaCl.
Evaporation of solvent afforded a bright orange powder. Yield: 0.80 g (81%). $^{31}$P($^1$H) NMR (CD$_2$Cl$_2$, 23 °C): δ 40.2 (s). Anal. Calcd for C$_9$H$_{22}$NiP$_2$S$_2$ (found): C, 34.31 (34.67); H, 7.04 (7.11).

**Ni(edt)(dmpe).** Ni(edt)(dmpe) was prepared analogously to the method for Ni(pdt)(dmpe). $^{31}$P($^1$H) NMR: (CD$_2$Cl$_2$, 23 °C): δ 42.8 (s). Anal. Calcd for C$_8$H$_{20}$NiP$_2$S$_2$ (found): C, 31.92 (32.12); H, 6.70 (6.81).

**Ni(pdt)(dppn).** A suspension of 0.78 g (1.25 mmol) NiCl$_2$(dppn) in 20 mL of CH$_2$Cl$_2$ was treated with 0.13 mL (1.25 mmol) of 1,3 propanedithiol followed by 0.69 mL (5 mmol) of triethylamine. The reaction turned a brown/green color. The solution was concentrated under reduced pressure, and 30 mL of hexane was added to the concentrate, causing precipitation of a red/orange powder. This solid was washed with hexanes, EtOH, and then finally Et$_2$O. Yield: 0.63 g (76%). Unlike related compounds, this species required storage in an N$_2$-filled glovebox. $^{31}$P($^1$H) NMR (CD$_2$Cl$_2$, 23 °C): δ 23.9 (s). Anal. Calcd for C$_{37}$H$_{32}$NiP$_2$S$_2$ (found): C, 67.19 (66.91); H, 4.88 (4.81). Single crystals suitable for X-ray diffraction were obtained by layering a CH$_2$Cl$_2$ solution with Et$_2$O.

**Ni(pdt)(dppbz).** A solution of 1.59 g (2.76 mmol) of NiCl$_2$(dppbz) in 20 mL of CH$_2$Cl$_2$ was treated with 0.28 mL (2.76 mmol) of 1,3 propanedithiol followed by 1.53 mL (11.04 mmol) of Et$_3$N, which caused the solution color to immediately turn a darker brown/red. The solution was stirred for 20 min., and then most of the solvent was removed under reduced pressure. When 3 mL of CH$_2$Cl$_2$ remained, 50 mL of Et$_2$O was added to the solution to precipitate an orange fluffy powder. This powder was dissolved in 5 mL of CH$_2$Cl$_2$, and this solution was diluted with 100 mL of ethanol. The reaction
mixture was placed in the freezer at -30 °C for 30 min. to afford orange crystals, which were washed with ethanol. Yield 1.00 g (59%). $^{31}$P{$^1$H} (CD$_2$Cl$_2$, 23 °C): δ 61.1. Anal. Calcd for C$_{33}$H$_{33}$NiP$_2$S$_2$ (found): C, 64.83 (64.52); H, 4.95 (4.74).

**NiCl$_2$(dcpe).** The literature procedure proceeded as described, although no spectroscopic data has been reported. $^{31}$P{$^1$H} NMR (CD$_2$Cl$_2$): δ 81.2. $^1$H NMR (CD$_2$Cl$_2$): δ 1.2-2.6 (broad m).

**Ni(pdt)(dcpe).** A solution of 3.5 g (6.34 mmol) of NiCl$_2$(dcpe) in 30 mL of CH$_2$Cl$_2$ was treated with 0.685 g (6.34 mmol) of 1,3-propanedithiol followed by a solution of 1.37 g (25.3 mmol) of NaOMe in 10 mL of MeOH. The mixture was stirred for 15 min., during which time the color of the reaction solution deepened. Solvent was removed under reduced pressure, and the residue was extracted into CH$_2$Cl$_2$ (20 mL). The slurry was filtered through a pad of Celite to remove NaCl. The dark orange filtrate was concentrated to dryness. Yield: 3.61 g (97%). $^{31}$P{$^1$H} (CD$_2$Cl$_2$): δ 72.31. Anal. Calcd for C$_{29}$H$_{54}$NiP$_2$S$_2$ (found): C 59.29 (59.57); H, 9.26 (9.33).

**Ni(SPh)$_2$(dcpe).** NiCl$_2$(dcpe) (0.820 g, 1.48 mmol, 1 equiv) was dissolved in CH$_2$Cl$_2$ (15 ml). NaSPh (0.392 g, 2.97 mmol, 2 equiv) was dissolved in MeOH (15 mL). The solution of NaSPh in MeOH was transferred into the solution of NiCl$_2$dcpe in CH$_2$Cl$_2$. The color of the reaction mixture immediately changed from bright orange to dark purple. The mixture was stirred for 30 minutes, and then solvent was removed under reduced pressure. The dark purple residue was extracted into CH$_2$Cl$_2$ (30 mL) and the slurry was filtered through a pad of Celite to remove NaCl. The dark purple filtrate was concentrated to dryness. Yield: 0.49 g, (47 %). $^{31}$P{$^1$H} (CD$_2$Cl$_2$): δ 75.71.
(CO)$_3$Fe(pdt)Ni(dpbz). A 100-mL Schlenk flask was charged with a Teflon-coated stir bar, 0.53 g (1.27 mmol) of FeI$_2$(CO)$_4$, and 0.776 g (1.27 mmol) of Ni(pdt)(dpbz). The flask was cooled to −78 °C. To the cooled mixture of solids was added 10 mL of CH$_2$Cl$_2$ (unchilled), resulting in a dark brown homogeneous solution. After 5 min, a 0.5-mL sample was removed (into an ambient temperature syringe) and checked by FT-IR spectroscopy to confirm the formation of the proposed µ-iodo intermediate (Table 8.1). The reaction solution, still at −78 °C, was then treated with a pre-cooled (−78 °C) solution of 0.48 g (2.54 mmol) of Cp$_2$Co in 5 mL of CH$_2$Cl$_2$, which was transferred via cannula. The reaction solution became brown-green and was allowed to warm to room temperature. After stirring for 1 h at room temperature, the reaction solution was concentrated under reduced pressure to a volume of about 5 mL. The solution was filtered through a 30 x 3 cm column of silica gel, eluting with about 100 mL of CH$_2$Cl$_2$. Unreacted Ni(pdt)(dpbz) remained on the column. The green eluate was concentrated under reduced pressure and then diluted with 50 mL of hexanes to precipitate a light green powder. Yield: 0.60 g (63%). IR (CH$_2$Cl$_2$, cm$^{-1}$): ν$_{CO}$ 2030, 1954. $^{31}$P($^1$H) (CD$_2$Cl$_2$, 23 °C): δ 60.1. ESI-MS: m/z 750 ([M$^+$]). Anal. Calcd for C$_{36}$H$_{30}$FeNiO$_3$P$_2$S$_2$ (found): C, 57.56 (54.89); H, 4.03 (4.01).

[(CO)$_2$Fe(SPh)$_2$Ni(CO)]$_2$(dppe)$_2$. A 100-mL Schlenk flask was charged with Teflon stir-bar, 0.264 g Ni(SPh)$_2$(dppe) (0.3916 mmol) and 0.142 g Fe$_2$(CO)$_9$ and dissolved in 8 mL degassed and dried THF. The flask was then stirred for 1 h and solvent removed under vacuum. The residue was then dissolved in Et$_2$O and filtered through an air-sensitive Celite column to remove insolubles. The Et$_2$O soluble portion was then dried under vacuum and washed via filter cannula with 2 x 10 mL MeOH, and
2 x 10 mL hexane. The resulting black/brown solid was dried under vacuum yielding 0.100 g (0.122 mmol, 31%). IR (CH$_2$Cl$_2$, cm$^{-1}$): $\nu$$_{CO}$ 1993 (s), 1938 (m) cm$^{-1}$. $^{31}$P{$^1$H} NMR (CD$_2$Cl$_2$): 33.9 (d), 49.9 (d).

[Fe(CO)$_3$(pdt)(H)Ni(dppbz)]BF$_4$. To a solution of 0.030 g (0.039 mmol) (CO)$_3$Fe(pdt)Ni(dppbz) in 10 mL of CH$_2$Cl$_2$ was added 0.050 mL (0.345 mmol) of HBF$_4$•Et$_2$O. The reaction mixture immediately changed from dark green to light orange. The solution was concentrated, and 20 mL of Et$_2$O was added to precipitate a light orange powder. Yield: 0.083 g (32%). IR (CH$_2$Cl$_2$, cm$^{-1}$): $\nu$$_{CO}$ 2082, 2023. $^{31}$P{$^1$H} NMR (CD$_2$Cl$_2$, 23 °C): $\delta$ 69.6. $^1$H-NMR (CD$_2$Cl$_2$, 23 °C): $\delta$ –3.42 (s, 1H, hydride). Anal. Calcd for C$_{36}$H$_{31}$BF$_4$FeNiO$_3$P$_2$S$_2$ (found): C, 51.53 (48.02); H, 3.72 (3.60).

Fe(CO)$_3$(pdt)Ni(dcpe). A 100-mL Schlenk flask was charged with a Teflon-coated stir bar, 0.350 g (0.85 mmol) of FeI$_2$(CO)$_4$, and 0.500 g (0.85 mmol) of Ni(pdt)(dcpe). The flask was cooled to –78 °C. To the cooled mixture of solids was added 10 mL of CH$_2$Cl$_2$, resulting in a dark brown homogeneous solution. A 0.5-mL sample was removed (using an ambient temperature syringe) and checked by FT-IR spectroscopy to confirm the formation of the proposed $\mu$-iodo intermediate (see Table 1). The remaining cold reaction solution was then treated with a pre-cooled (–78 °C) solution of 0.320 g (1.70 mmol) of Cp$_2$Co in 5 mL of CH$_2$Cl$_2$, which was transferred via cannula. The reaction solution became brown-green in color and was allowed to warm to room temperature. After stirring for 1 h at room temperature, the reaction mixture was evaporated under reduced pressure, and the dark brown residue was washed with 3 x 15 mL of MeCN (until the washings were colorless). The brown residue was extracted into 5 mL of CH$_2$Cl$_2$, and the dark brown product was precipitated by addition
of 30 mL of hexanes. Yield: 0.21 g (34 %). IR (CH₂Cl₂, cm⁻¹): νCO 2014, 1940. ³¹P{¹H} (CD₂Cl₂): δ 79.3.

Fe(CO)₃(pdt)Ni(dppe) and [Fe(CO)₃(pdt)(I)Ni(dppe)]I. The procedure developed for Fe(CO)₃(pdt)Ni(dcpe) was applied to the synthesis of Fe(CO)₃(pdt)Ni(dppe). Yield: 0.22 g (38 %). A solution of 30 mg (0.0426 mmol) of Fe(CO)₃(pdt)Ni(dppe) in 5 mL of CH₂Cl₂ was cooled in a –78 ºC bath and then treated with 10.8 mg (0.0426 mmol) of solid I₂. A 0.5 mL sample was removed (into an ambient temperature syringe) and checked by FT-IR spectroscopy: νCO 2095, 2055, 2025 (see Figure 3).

[Fe(CO)₃(pdt)(H)Ni(dcpe)]BF₄. To a solution of 0.150 g (0.207 mmol) of (CO)₃Fe(pdt)Ni(dcpe) in 10 mL of CH₂Cl₂, was injected 100 µL (0.691 mmol) of HBF₄·Et₂O. An immediate color change from green-brown to dark orange was observed. The reaction solution was concentrated under reduced pressure to about 5 mL, and about 30 mL of hexanes were added to precipitate a dark orange powder. Yield: 0.090 g (53%). ³¹P{¹H} NMR (CD₂Cl₂): δ 89.75. ¹H NMR (500 MHz, CD₂Cl₂): δ -3.00 (br. s, 1H, hydride). Anal. Calcd for C₃₂H₅₅BF₄FeNiO₃P₂S₂·CH₂Cl₂ (found): C, 44.03 (43.72); H, 6.38 (6.41). ESI-MS: m/z 727 ([M]⁺). IR (CH₂Cl₂, cm⁻¹): νCO 2078, 2017.

Fe(CO)₃(pdt)Ni(dmpe) and [Fe(CO)₃(pdt)(H)Ni(dmpe)]Cl. A 100-mL Schlenk flask was charged with a Teflon-coated stir bar, 0.053 g (0.127 mmol) of FeI₂(CO)₄, and 0.040 g (0.127 mmol) of Ni(pdt)(dmpe). The flask was cooled to –78 ºC. To the cooled mixture of solids was added 10 mL of CH₂Cl₂, resulting in a dark brown homogeneous solution. A 0.5-mL sample was removed (using an ambient temperature syringe) and
checked by FT-IR spectroscopy to confirm the formation of the proposed μ-I intermediate (Table 1). The cold reaction solution was then treated with a pre-cooled (−78 °C) solution of 0.320 g (1.70 mmol) of Cp₂Co in 5 mL of CH₂Cl₂, which was transferred via cannula. The addition of the Cp₂Co induced a color change from bright red to bright green. A sample of the cold (−78 °C) reaction solution was withdrawn and injected into an ambient temperature IR cell, and the spectrum was recorded within a few minutes (Table 1.2). If, however, the reaction solution was allowed to warm to room temperature, the resulting IR spectrum indicated decomposition to a more complex mixture.  

31P{¹H} NMR (CD₂Cl₂): δ 46.6. A cold (−78 °C) solution of Fe(CO)₃(pdt)Ni(dmpe) was treated with 0.33 mL (0.64 mmol, 5 equiv) of a 1.90 M solution of HCl in Et₂O. The product was also thermally unstable and could only be observed by IR spectroscopy if the solution was maintained near −78 °C. IR (CH₂Cl₂, cm⁻¹): ν₉ 2078, 2018. ¹H NMR (CD₂Cl₂, 25 °C): δ -3.63 (s, 1H, hydride). 31P{¹H} (CD₂Cl₂): δ 61.3. To obtain the NMR spectra, all of the solvent was removed under reduced pressure and the isolate orange residue was brought into a glovebox. 5 mg of the sample was dissolved in CD₂Cl₂, placed in a J Young tube and a ¹H NMR spectrum was acquired at room temperature.

Fe(CO)₃(edt)Ni(dmpe) and [Fe(CO)₃(H)(edt)Ni(dmpe)]Cl. This compound was prepared in the same manner and behaved in a similar fashion to Fe(CO)₃(pdt)Ni(dmpe). 31P{¹H} NMR (CD₂Cl₂): δ 48.2. Protonation with HCl as in the preceding experiment afforded a solution assumed to contain [Fe(CO)₃(H)(pdt)Ni(dmpe)]Cl. IR (CH₂Cl₂, cm⁻¹): ν₉ 2079, 2018. ¹H NMR (CD₂Cl₂): δ −3.71 (s, 1H, hydride). 31P{¹H} (CD₂Cl₂): δ 59.8. To obtain the NMR spectra, all of the
solvent was removed under reduced pressure and the isolate orange residue was brought into a glovebox. 5 mg of the sample was dissolved in CD$_2$Cl$_2$, placed in a J Young tube and a $^1$H NMR spectrum was acquired at room temperature.

**Electrochemistry.** As the nickel-iron hydrides presented in this paper all degrade over time to uncharacterized products in MeCN solution ($t_{1/2} \sim 60$ min), electrochemical measurements were performed on CH$_2$Cl$_2$ solutions. Cyclic voltammetry experiments were conducted with 20-mL one-compartment glass cell with a tight-fitting Teflon top using a BAS-100 Electrochemical Analyzer. The working electrode was a glassy carbon disk (3.00 mm in diameter). A silver wire was used as a pseudo-reference electrode, and the counter electrode was a Pt wire. The electrolyte was 0.1 M Bu$_4$NPF$_6$ in CH$_2$Cl$_2$. Ferrocene (~ 1 mM) was added as an internal reference, and cyclic voltamgramms were each referenced to this Fc$^{0+/+}$ couple = 0.00 V. $iR$ compensation was applied to all measurements using the BAS software. Cell resistance was determined prior to each scan and the correction automatically applied to the subsequently collected cyclic-voltammagram.

**Cyclic Voltammetry for [2H]BF$_4$.** A solution of 3.4 mg (0.0046 mmol) in 5 mL CH$_2$Cl$_2$ was prepared in the CV cell and was treated with successive aliquots (22 µL, 2 equivs) of a freshly prepared solution of 0.427 M ClCH$_2$CO$_2$H in CH$_2$Cl$_2$ solution. Cyclic voltammograms were recorded at 100 mV/s.
References


