PROTON-COUPLED ELECTRON TRANSFER REACTIONS
IN BIO-INSPIRED CATALYSIS

BY

MIOY T. HUYNH

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, 2017

Urbana, Illinois

Doctoral Committee:

Professor Sharon Hammes-Schiffer, Chair
Professor Thomas B. Rauchfuss
Professor Dallas R. Trinkle
Assistant Professor Josh Vura-Weis
ABSTRACT

Nature often serves as a seminal source of inspiration for the design of catalysts, which is likely due to the efficiency of many biological systems. Bio-inspired design is especially applicable in many energy conversion processes, such as artificial photosynthesis, and in the development of alternative renewable energy technologies. Often proton-coupled electron transfer (PCET) is central to the interconversion of energy. Consequently, understanding the role of PCET in the mechanisms of bio-inspired catalysts can aid in the design of next generation catalysts and the elucidation of guiding design principles. In nature, hydrogenases are the most active catalysts for hydrogen production, with rates and overpotentials comparable to leading, synthetic Pt catalysts while utilizing earth abundant metal centers in their active sites. Density functional theory calculations, in conjunction with complementary experiments, were conducted to understand the underlying physical principles of hydrogenase-based models. Our collective studies on hydrogenase-based models have revealed two emerging themes: (1) one metal center is acid-base active while the adjacent metal center is redox active, and (2) the importance of considering thermodynamically less stable but active intermediates.

Nature also often relies on mediators to couple electrons and protons between various cycles in catalysis. A prime example of the importance of electron-proton transfer mediators is photosynthesis, where the slow water splitting reaction and the fast photoinitiation of a reaction center are mediated by a Tyr-His redox pair in Photosystem II. Understanding the fundamental PCET processes in Tyr-His models is helpful for the design of photoelectrochemical water splitting cells, where water oxidation provides protons and electrons for hydrogen production via a hydrogen-evolving catalyst, such as those based on hydrogenases. To understand the kinetics of this reaction, a nonadiabatic PCET rate theory is applied to predict and interpret kinetic isotope
effects for Tyr-His model systems. Importantly, theory predicted new mediators capable of two proton transfers coupled to an electron transfer, and these predictions were later validated experimentally. In addition to the Tyr-His redox mediator, Photosystem II also employs a quinone derivative to mediate the transfer of electrons and protons across the membrane. This plastoquinone cycles between the oxidized quinone state and the doubly-reduced, doubly-protonated hydroquinone state (i.e., a $2\,e^-/2\,H^+$ PCET process). Our analysis of fundamental redox behaviors of over one hundred quinones using linear scaling relationships, such as Hammett correlations, revealed linear correlations between $1\,e^-$ reduction potentials, $pK_a$ values, and $2\,e^-/2\,H^+$ reduction potentials with an effective Hammett constant. More importantly, key deviations resulting from hydrogen-bonding, halogenated, charged, and sterically-bulky substituents were identified and analyzed. In principle, these types of deviations can be leveraged to further develop and tune quinone-based catalysts, mediators, or devices beyond the redox properties predicted by standard linear scaling relationships.
ACKNOWLEDGEMENTS

I would like to dedicate this thesis to my mom and dad. They often say that it takes a village to raise a child. If that is true, then I am not sure what it must have taken to raise the nine of us. Thank you for giving up everything so that we could have something. Thank you for being a part of each of our lives. And thank you for always asking if I am happy.

To my advisor, Professor Sharon Hammes-Schiffer, I am grateful for your endless support, mentorship, invaluable guidance and advice, and welcoming me into the group. I would also like to extend special thanks to Professor Tom Rauchfuss, for being a pseudo-coadvisor to me and providing me with the opportunity for such in-depth collaborations. To my undergraduate advisor, Professor Anastassia Alexandrova, thank you for setting me on a path I did not know existed. Of course, I am also indebted to the many collaborators I have worked with over the years: Dave Schilter, Geoffrey Chambers, Cass Richers, Olbelina Ulloa, and Wenguang Wang from UIUC; Colin Anson and Prof. Shannon Stahl from UW-Madison; S. Jimena Mora, Professor Tom Moore, and Professor Ana Moore from ASU.

I am thankful for the amazing and supportive group members I have been able to work with. To Alexander Soudackov, Soumya Ghosh, and Puja Goyal, thank you for the countless discussions—I could not have learned so much without you! I am also grateful to the recipients of my endless rants: fellow crafter Nihan Ucisik, perpetual lunch-buddy Tanner Culpitt, and the constituents of “the better office” for providing me with endless laughs. Of course, thank you for putting up with my regularly-scheduled Tuesday “I quit!” exclamations. And, lastly, to my partner, Alex Dunn, thank you for your never-ending support, patience, advice, and humor—it means the world to me.
# TABLE OF CONTENTS

- **Chapter 1.** Introduction.................................................................................................................. 1

- **Chapter 2.** [FeFe]-Hydrogenase Models: Characterization and Mechanistic Insights......................... 21

- **Chapter 3.** Interplay between Terminal and Bridging Diiron Hydrides........................................... 52

- **Chapter 4.** Protonation of [NiFe]-Hydrogenase Models ................................................................ 79

- **Chapter 5.** H$_2$ Evolution Mechanism for [NiFe]-Hydrogenase Models........................................... 112

- **Chapter 6.** Models of the Ni-L and Ni-SI$_2$ States of the [NiFe]-Hydrogenase ............................... 146

- **Chapter 7.** Quinone 1 e$^-$ and 2 e$^{-}$/2 H$^+$ Reduction Potentials .................................................. 184

- **Chapter 8.** Concerted Electron-Proton Transfer in Tyr-His Models ............................................. 208
CHAPTER 1. Introduction*

1.1. Energy Consumption and Production and Renewable Sources

The production and demand for energy is an ever-growing concern for society. In 2015, the United States consumed 97.7 quadrillion BTU of energy (Figure 1.1).\(^1\) And of the amount of energy consumed, an overwhelming 80% was derived from the three major fossil fuels: petroleum/oil, natural gas, and coal—all considered nonrenewable energy sources. Though there is substantial uncertainty in the levels of future supply and demand, current estimates of the global supply of fossil fuels are expected to meet energy demands for a fair amount of time (tens to hundreds of years depending on the energy source). However, continued reliance on nonrenewable fossil fuels and hydrocarbon fuels depends on the recurrent discovery of new energy reserves, as well as the development of new technologies for access and harvesting.

![Figure 1.1. United States energy consumption by energy source in 2015. The numerical data is obtained from the U. S. Energy Information Administration.](image)

---

In stark contrast to fossil fuels, renewable energy sources only accounted for 10% of the total energy consumed in 2015.\textsuperscript{1} While much of the history of renewable energy has been dominated by biomass and hydroelectric power, the contributions from wind and solar energies have made remarkable gains in the past few decades. For instance, in the early 1990’s, solar energy contributed 5% to total renewable energy produced; today, solar energy contributes 15% of total renewable energy produced.\textsuperscript{1} Of course, despite this progress, the ever-present critique of renewable energy sources is their intrinsic intermittency. Consequently, the ability to capture, convert, store, and transport solar energy has garnered increased attention in recent years.

1.2. Energy Storage

Batteries are, perhaps, one of the most discernable types of energy storage technologies. As batteries are low energy density (energy per unit weight) storage devices, advances in battery technology have largely focused on increasing their power density (rate of energy extraction) and cycle lifetime.\textsuperscript{2} Their low energy densities are generally constrained by the material’s physical properties (Figure 1.2).\textsuperscript{3} However, since the electricity can be stored directly from solar energy, batteries benefit from typically high conversion efficiencies.\textsuperscript{4}

\textbf{Figure 1.2.} Ragone plot of power density versus energy density of various energy storage technologies. Figure adapted from Ref. 3.
Chemical fuels, on the other hand, offer similar power densities to batteries, but possess far greater energy densities through the storage of electrons in the bonds of light elements (e.g. C–H, N–H, and H–H bonds), as shown in Figure 1.2.\textsuperscript{3} This energy density is critical for the large-scale energy storage demands of society. In fact, the storage of energy in a chemical bond is the basis for hydrocarbons (fossil) fuels, since C–H bonds provide the optimum supply of energy in terms of volumetric energy densities. Hydrogen, meanwhile, possesses the largest energy density by mass, but its gaseous nature limits its volumetric energy density.

1.3 Photosynthesis: Providing Solar Fuels in Biology

Nature also stores solar energy in the form of chemical bonds. In photosynthesis, solar energy storage is achieved by splitting water. The initial steps of water splitting involve breaking the four O–H bonds of two water molecules, concomitant with the formation of an O=O bond and release of four protons (H\textsuperscript{+}) and four electrons (e\textsuperscript{–}).

\[
2 \text{H}_2\text{O} \rightleftharpoons \text{O}_2 + 4 \text{H}^+ + 4 \text{e}^- \quad E^\circ = -1.23 \text{ V vs. NHE}
\]

These protons and electrons are then combined with CO\textsubscript{2} to produce hydrocarbon fuels. Conversely, the protons and electrons can be combined to form hydrogen via a hydrogen evolution reaction (HER) catalyst.

\[
2 \text{H}^+ + 2 \text{e}^- \rightleftharpoons \text{H}_2 \quad E^\circ = 0 \text{ V vs. NHE}
\]

In both half reactions, $E^\circ$ is the formal reduction potential relative to the normal hydrogen electrode (NHE) at pH = 0 and 25 °C. The pH dependence for both reactions can be derived for mixed electron-proton equilibrium: Ox $+ m$ H\textsuperscript{+} $+ n$ e\textsuperscript{–} $\rightleftharpoons$ Red. The results are summarized as a Pourbaix ($E$-pH stability) diagram shown in Figure 1.3.

\[
E^\circ = E^\circ' - \frac{m}{n} \cdot 0.0591 \cdot \text{pH} \quad (1.1)
\]
Overall, the water splitting reaction catalyzed by an oxygen evolving complex (OEC) found in Photosystem II (PS II) is

\[ 2 \text{H}_2\text{O} \rightleftharpoons 2 \text{H}_2 + \text{O}_2 \quad E^\circ_{\text{cell}} = -1.23 \text{ V} \]

where \( E^\circ_{\text{cell}} \) is the cell potential (voltage), indicating that \(-470 \text{ kJ/mol}\) of energy can be stored via this process. In PS II, the energy from sunlight is propagated to a reaction center, P680, through a network of antenna complexes. An electron in P680 is then excited, and the photoexcited electron induces a cascade of subsequent electron transfers, resulting in the transport of one electron from Photosystem II and through Photosystem I (i.e., the electron transport chain). A “hole” (vacancy upon electron loss) is passed to a Mn₄O₄Ca cluster, known as the oxygen evolving complex (OEC), which splits water into oxygen and protons. The net result of these processes is the migration of an electron and a hole to opposite sides of the biological membrane. These steps are depicted schematically in Figure 1.4, where only the reactions and formal potentials (V vs. NHE) within the vicinity of PS II are shown.
Figure 1.4. Schematic diagram of Photosystem II showing the absorption of light, photoexcitation of an electron in P680, the cascade of electron transfers composing the electron transport chain, and the water splitting reaction catalyzed by the oxygen evolving complex (OEC). The approximate redox potentials for the P680/P680$^{+•}$ and the water splitting reactions are given in units of V vs. NHE.

An important feature of photosynthesis is the interface between the photochemical redox of P680 and the multi-electron, multi-proton water oxidation reaction (or the reduction of CO$_2$). In essence, these types of interfaces would involve the coupled transfer of electrons and protons—in other words, proton-coupled electron transfer (PCET).$^5$ This is one of the most widely-cited examples of PCET reactions in biology.$^6$-$^12$ The interface between the fast photoinitiated steps of photosynthesis and the slow catalytic process of water oxidation is a Tyr-His redox pair (Figure 1.5). Functionally, the Tyr-His pair acts as a mediator between the oxidized electron donor (P680$^{+•}$) and the OEC (a distance of ~10 Å), coupling proton and electron transfer between the two sites.
Figure 1.5. The Tyr-His redox pair in PS II, showing the transfer of an electron from the phenol to P680$^{+*}$ and the transfer of the phenolic proton from Tyr to a hydrogen-bonded His nitrogen, leading to the formation of a phenoxyl radical.

For perspective, the estimated potential for the P680$^{+*}$/P680 couple is $\sim$1.3 V vs. NHE,$^{13}$ while the redox potential for the water oxidation reaction at pH $\sim$5 (the approximate pH of the lumen) is $\sim$0.8 V vs. NHE.$^{14}$ This leaves a fairly narrow potential window for water oxidation by tyrosine. Studies on aqueous solutions of tyrosine indicate that between pH $\sim$2 and pH $\sim$10, the tyrosine redox couple is pH-dependent (see Pourbaix diagram in Figure 1.6).$^{15}$

$$\text{TyrO}' + H^+ + e^- \rightleftharpoons \text{TyrOH} \quad E^\circ = (1.34 - 0.0591 \cdot \text{pH}) \text{ V vs. NHE}$$

Figure 1.6. Pourbaix diagram for water oxidation, along with the pH-independent P680$^{+*}$/P680 couple and the pH-dependent tyrosine reduction in aqueous solution (white dashed line). The shaded window is relevant to the pH conditions of the photosynthetic membrane.
From a thermodynamic standpoint, the oxidation of Tyr by P680$^{+•}$ likely occurs with the transfer of the phenolic proton to the hydrogen-bonded His (Figure 1.5), with a potential that is appropriately between the P680 and water oxidation potentials at near-neutral to acidic pH conditions (shaded window in Figure 1.6). In other words, the redox potential leveling provided by a PCET reaction would tune the Tyr redox potential to lie between that needed for P680$^{+•}$ reduction and water oxidation at the OEC. It is the thermodynamic activity of this proton that preserves the high redox potential necessary for water oxidation.

In addition to redox leveling, this PCET reaction is also proposed to play a large role in the generation of proton-motive force. Tommos and Babcock proposed that an extensive hydrogen bond network originating at the OEC and extending to the lumen (~20 Å) could act as a proton channel, and hypothesized that such a mechanism would arise through a series of proton transfers commencing with the proton-coupled oxidation of the Tyr-His relay. A possible proton channel of this sort, extending from the OEC through Tyr-His and finally to the lumen, was identified in a crystal structure obtained by Umena et al. with 1.9 Å resolution. Nevertheless, the subject of proton channels within PS II remains largely controversial, and many channels have been proposed thus far, operating via a Grotthuss-like mechanism. Unequivocal, though, is the fact that the protons from water oxidation are pumped into the lumen and contribute to proton-motive force.

As an intermediary, the Tyr-His pair also averts charge recombination, which is vital for the accumulation of reducing equivalents on the OEC required to oxidize water. Understanding this fundamental PCET reaction in PS II may aid in the adaptation of these processes to artificial photosynthetic energy conversion devices, where a similar architecture could prevent back electron transfer after photoinduced charge separation. In conjunction,
understanding proton activity in PS II could also provide design principles for controlling the proton activity of water oxidation catalysts and the design of bio-inspired proton wires capable of translocating protons over large distances.

1.4. Bio-Inspired Tyr-His Models and Artificial Photosynthesis

Experimental and theoretical interest in the role of PCET in photosynthesis, specifically the Tyr-His redox pair, has led to the synthesis of a number of bio-inspired molecular constructs featuring intramolecularly hydrogen-bonded phenols\textsuperscript{22} and photochemically-active synthetic PS II mimics bearing chromophores.\textsuperscript{23,24} Moore et al. have designed a porphyrin (PF\textsubscript{10}) construct bearing a covalently attached benzimidazole-phenol (BIP) moiety, which is adsorbed to the surface of a colloidal TiO\textsubscript{2} nanoparticle (Figure 1.7).\textsuperscript{25-27}

![Figure 1.7](image)

Figure 1.7. Bio-inspired molecular construct, with the red benzimidazole-phenol (BIP) mimicking the Tyr-His redox pair, the blue porphyrin moiety (PF\textsubscript{10}) mimicking the chlorophyll of P680, and the gray TiO\textsubscript{2} nanoparticle mimicking the pheophytin (electron acceptor) of PS II. The colored arrows drawn on the initial state (left structure) indicate the electron and proton transfers occurring upon irradiation of the porphyrin moiety (step 1), which leads to the final charge-separated state shown on the right side. Note that step 3 involves the concerted transfer of an electron from the BIP to the porphyrin moiety and a proton from the phenol to the benzimidazole nitrogen.

In these molecule-nanoparticle constructs (BIP-PF\textsubscript{10}–TiO\textsubscript{2}), irradiation of the porphyrin moiety leads to a photoinduced electron transfer from the excited porphyrin to the conduction band of the TiO\textsubscript{2}, forming the initial charge-separated state BIP-PF\textsubscript{10}–•–TiO\textsubscript{2}–•. A PCET reaction then facilitates the transfer of a hole from the photo-oxidized porphyrin to the BIP moiety,
generating the final charge-separated state (BIP$^{••}$-PF$_{10}$$^{−}$-TiO$_2$$^{−}$) in which a phenoxy radical is stabilized through intramolecular hydrogen bonding. These processes are shown in Figure 1.7 and purposely resemble those occurring in PS II.

An artificial photosynthetic water splitting cell could also use sunlight to drive the splitting of water in a similar fashion to biology, providing equivalents of electrons and protons that can be recombined to form hydrogen, which can be stored as a fuel. A general schematic of an artificial photosynthetic water splitting cell is shown in Figure 1.8 based on the work of Mallouk and coworkers.$^{28,29}$ A pervasive problem in synthetic molecular energy conversion devices, however, is the fast rate of back electron transfer relative to the slow rate of water oxidation. The back electron transfer leads to thermalization of photoinduced charge-separated states and results in low quantum yields for water splitting.

**Figure 1.8.** Schematic diagram of an artificial photosynthetic water splitting cell, which uses sunlight to generate molecular oxygen and hydrogen. This diagram is inspired by the device reported by Mallouk and coworkers.$^{28,29}$
In these devices, the initiation step involves irradiation, which is followed by the ejection of an electron from the sensitizer to TiO\textsubscript{2} nanoparticles on the photoanode. The injected electrons are then conducted to the cathode by a wire. Meanwhile, the oxidized sensitizer is reduced by a water oxidation catalyst, which generates oxygen and protons. A permeable membrane separating the two sides of the cell allows for the diffusion of protons across the cell, where they combine with electrons to form hydrogen on the cathode. Like the anode, the cathode can also be decorated with molecular proton reduction electrocatalysts to generate hydrogen. One of the first molecular electrocatalysts to be covalently attached to an electrode was the hydrogenase-based Ni(P\textsubscript{2}N\textsubscript{2})\textsubscript{2} catalyst.\textsuperscript{30}

![Dye-sensitized TiO\textsubscript{2} photoanode of a water splitting solar cell showing the BIP mediator-decorated iridium oxide water oxidation catalyst. The electron transfer and intramolecular proton transfer reactions are indicated with arrows. Figure adapted from Ref. 31.](image)

Returning to the BIP construct discussed previously, the advantages of an electron-proton mediator become apparent. In PS II, the Tyr-His redox mediator not only provides the high redox potential required for water oxidation, but also circumvents charge recombination. Interestingly, the experimentally measured potential corresponding to the oxidation of BIP is
1.24 V vs. NHE (or 1.00 V vs. SCE), indicating that the bio-inspired molecular constructs would be thermodynamically poised for water oxidation. Indeed, a subsequent study by Zhao et al. showed that covalently attaching a related BIP construct to a water oxidation catalyst (a colloidal iridium oxide nanoparticle), which is then coadsorbed onto a dye-sensitized TiO₂ photoanode, results in photoelectrochemical water splitting with enhanced quantum efficiency relative to the mediator-free device (Figure 1.9). These devices relied on the increased spatial separation of the eventual electron donor-acceptor pair introduced by the BIP redox mediator to dampen the rate of back electron transfer and increase the lifetime of the charge-separated state.

1.5. Hydrogenases and Hydrogen in Nature

Certain prokaryotes live off energy liberated from the metabolism of simple oxidants and hydrogen. In a landmark paper by Stephenson and Stickland, they described bacterial cultures from river mud that reduced NO₃⁻, O₂, and fumarate, writing that “hydrogen is in some way activated, and this activation can be conveniently expressed H₂ → 2H without implying anything about the nature of the reaction” and coining the catalyst “hydrogenase.” These H₂-processing metalloenzymes are the basis for many anaerobic bacteria, protozoa, fungi, and algae, collectively producing and consuming an estimated 0.3 Gt of H₂ each year.

Three major classes of hydrogenase have been identified, featuring either an FeFe, a NiFe, or an Fe core. The [FeFe]- and [NiFe]-hydrogenases have been shown to mediate the bidirectional interconversion of hydrogen (H₂) with protons and electrons at high rates and potentials very close to those bounded by thermodynamics. It is worth emphasizing that the hydrogenase families share several mechanistic motifs—principal among them H₂ heterolysis—that are in no small part borne of the structural similarities revealed by their active sites.
Despite the three enzyme classes being phylogenically unrelated,\textsuperscript{39} “convergent evolution” causes expression of these common features that are apparently expedient for H\textsubscript{2} processing.\textsuperscript{40}

\textbf{Figure 1.10.} The [FeFe]- and [NiFe]-hydrogenase enzymes and their active sites with white arrows indicating the open coordination site for the binding of protons or H\textsubscript{2}.

The pyramidalized bimetallic active sites contain low spin Fe-CO-thiolate centers with biologically unusual CO and CN\textsuperscript{−} cofactors. Additionally, hydrogenases also utilize Fe\textsubscript{4}S\textsubscript{4} clusters to mediate the transfer of electrons, and function with high rates and low overpotentials. In addition to the identity of the metal centers that comprise their active sites, [FeFe]- and [NiFe]-hydrogenases also differ in the location of the open coordination site for proton or H\textsubscript{2} binding. In [FeFe]-hydrogenase, this is a terminal site, whereas in [NiFe]-hydrogenases, this is a bridging site between the two metal centers. Additionally, the [FeFe]-hydrogenase employs an azadithiolate (adt) cofactor to mediate the transfer of protons to and from the metal site;\textsuperscript{41} meanwhile, the identity of the proton relay remains an open question for [NiFe]-hydrogenases. While most hydrogenases are considerably oxygen sensitive, [NiFe]-hydrogenases have been shown to exhibit tolerance for oxygen.\textsuperscript{42-44}
1.6. Bio-Inspired Hydrogenase Models

Without a doubt, the growing interest in renewable, carbon-neutral energy technologies has motivated considerable research into the expression, structure, and function of hydrogenases.\cite{40} Given that H\textsubscript{2} processing is currently championed by the Pt group metals, more cost-effective alternatives would involve the use of purely organic catalysts,\cite{45} or those based on earth abundant metals, like the hydrogenases.\cite{46-54} Better yet, their synthetic models may be potentially more robust and mass-producible, with smaller spatial footprints that allow densely decorated, highly active electrodes to be prepared and incorporated into devices,\cite{55-57} such as fuel cells or the water splitting cell shown in Figure 1.8.

![Figure 1.11. Structure of the doubly protonated [tH-Fe\textsubscript{2}(adt-H)(dppv)\textsubscript{2}(CO)\textsubscript{2}]\textsuperscript{2+} (left) and bridging hydride [(dppe)Ni(pdt)(\mu-H)Fe(CO)\textsubscript{3}]\textsuperscript{+} (right) model complexes.](image)

Of the two hydrogenases, models of [FeFe]-hydrogenase are more advanced. A notable example is the Fe\textsubscript{2}(adt)(dppv)\textsubscript{2}(CO)\textsubscript{2} model complexes designed by Rauchfuss and coworkers (left structure in Figure 1.11).\cite{58} In the presence of the strong acid CF\textsubscript{3}CO\textsubscript{2}H (TFA, pK\textsubscript{a}\textsubscript{MeCN} = 12.7), these models are very fast electrocatalysts for hydrogen evolution, catalyzing HER at an incredible rate of 58 000 s\textsuperscript{–1} at –1.11 V vs. Fe\textsuperscript{0/+} in dichloromethane, but at an overpotential of 0.51 V. The overpotential is defined as the difference between the catalytic potential (\(E_{\text{cat}} = –1.11\) V) and the standard reduction potential of the acid (\(E_{\text{a, HA/H}_2}^\circ \approx –0.6\) V vs. Fe\textsuperscript{0/+} for TFA).\cite{59}

Impressively, a crystal structure was obtained revealing a short distance (1.89 Å) between the ammonium of the protonated adt and the terminal hydride on the iron center, a very probable
intermediate in the hydrogen evolution mechanism. The remarkable catalytic activity of this model is proposed to result from the adt proton relay, which increases the local concentration of protons at the terminal iron hydride site.

The most widely studied [NiFe]-hydrogenase models are derived from the (dppe)Ni(pdt)Fe(CO)₃ models designed by Schröder and coworkers.⁶⁰ Rauchfuss and coworkers later demonstrated that a stable bridging hydride complex, [(dppe)Ni(pdt)(μH)Fe(CO)₃]⁺, was formed upon treatment with acid (right structure in Figure 1.1).⁶⁰⁻⁶² Moreover, the models proved to be competent proton reduction catalyst in the presence of TFA acid, with HER rates of 20 s⁻¹ at −1.20 V vs. Fe⁰/⁺ in acetonitrile and an overpotential of 0.50 V. The rate of HER for the modified (dppe)Ni(edt)Fe(CO)₃ derivative is ~300 s⁻¹ under similar conditions, indicating that the catalytic activity can be further tuned.

An important feature of the hydrogenases is the employment of organic bases in the outer-coordination sphere to mediate or shuttle protons to and from the metal centers. Generally, the protonation of metal centers is slower than the protonation of main-group elements, such as nitrogen and oxygen.⁶³ Consequently, the incorporation of pendant bases (e.g., the adt cofactor) enhances the kinetics of metal-site protonation, as seen in many enzymes.⁶⁴ The role of the pendant base in the hydrogenase-based Ni(P₂N₂)₂ catalyst, originally reported by DuBois and coworkers, has been particularly well-studied experimentally and theoretically, correlating increased HER rates with the presence of a proton relay.⁵⁴,⁶⁵ Notably, derivatives of these Ni(P₂N₂)₂ catalyst have been covalently attached to carbon nanotubes and deposited onto indium tin oxides,⁶⁰ glassy carbon electrodes,⁶⁶ and gallium phosphide and silicon photocathodes.⁶⁷

Because of the nature of hydrogen evolution, the mechanisms by which these hydrogenase-based models operate must involve electron and proton transfers. Shared motifs
between the hydrogenases and their models are the mixed-valency of the bimetallic cores and the role of reduced metal-hydrides. To this end, understanding the electrochemical (redox) and acid-base properties, as well as elucidating possible hydrogen evolution mechanisms, would aid in further tuning the catalytic activity of molecular electrocatalysts.

1.7. Outline of Thesis

Chapters 2–6 are centered around synthetic models of [FeFe]- and [NiFe]-hydrogenases and their proposed hydrogen evolution mechanisms. Chapter 2 deals with a biomimetic model of the [FeFe]-hydrogenase enzyme for which a \( \text{H}_2 \) evolution mechanism is proposed that proceeds through a terminal hydride intermediate. An attempt to stabilize the terminal hydride isomer relative to the bridging hydride isomer, which is typically more thermodynamically stable, is presented in Chapter 3. The first generation [NiFe]-hydrogenase model is discussed in Chapter 4, with experimental and theoretical results indicating that protonation of NiFe models proceeds upon isomerization at the Ni center. Subsequently, a proposed \( \text{H}_2 \) evolution mechanism, which proceeds through a reduced bridging hydride intermediate, is proposed for these dithiolate NiFe models in Chapter 5. Chapter 6 details a newer generation NiFe model in which Ni-site isomerization is circumvented in the hydrogen evolution pathway.

Chapters 7 and 8 are focused on studies of electron-proton transfer mediators, which facilitate the transfer of electrons and/or protons between coupled redox reactions, such as in PS II. In particular, Chapter 7 explores the fundamental redox properties of substituted quinones, analyzing the relationships between electron transfer (redox), proton transfer (protonation), and proton-coupled electron transfer reactions. For perspective, plastoquinone (PQ) mediates the transfer of electrons and protons across the photosynthetic membrane in PS II (Figure 1.4). Chapter 8 focuses on a set of biomimetic molecular constructs of the Tyr-His redox pair in PS II,
which mediates the transfer of electrons from the OEC to P680 and protons from the OEC to the lumen (inside of the membrane). Thermodynamic and kinetic analysis is presented to showcase the theory-led design of a construct featuring two proton transfers coupled to the oxidation of a phenol.

1.8. References


(66) Das, A. K.; Engelhard, M. H.; Bullock, R. M.; Roberts, J. A. A hydrogen-evolving Ni (P₂N₂)₂ electrocatalyst covalently attached to a glassy carbon electrode: Preparation,

Chapter 2. [FeFe]-Hydrogenase Models: Characterization and Mechanistic Insights*†

2.1. Abstract

The [FeFe]-hydrogenase enzymes catalyze hydrogen oxidation and production efficiently with binuclear iron metal centers. Recently the bio-inspired H₂-producing model system, Fe₂(adt)(CO)₂(dppv)₂ (adt = azadithiolate), was synthesized and studied experimentally. In this system, the azadithiolate bridge facilitates the formation of a doubly protonated, ammonium-hydride species through a proton relay. Herein computational methods are utilized to examine this system in the various oxidation states and protonation states along proposed mechanistic pathways for H₂ production. The calculated results agree well with the experimental data for the geometries, CO vibrational stretching frequencies, and reduction potentials. The calculations illustrate that the NH···HFe₄ dihydrogen bonding distance in the doubly protonated species is highly sensitive to the effects of ion-pairing between the ammonium and BF₄⁻ counterions, which are present in the crystal structure, in that the inclusion of BF₄⁻ counterions leads to a significantly longer dihydrogen bond. The non-hydride Fe₃ center was found to be the site of reduction for terminal hydride species and unsymmetric bridging hydride species, whereas the reduced symmetric bridging hydride species exhibited spin delocalization between the Fe centers. According to both experimental measurements and theoretical calculations of the relative pKₐ values, the Fe₄ center of the neutral species is more basic than the amine, and the bridging hydride species is more thermodynamically stable than the terminal hydride species. The calculations implicate a possible pathway for H₂ evolution that involves an intermediate with H₂ weakly bonded to one Fe, a short H₂ distance

† Author contributions: Dr. Wenguang Wang and Prof. Thomas Rauchfuss were responsible for all experiments and associated analyses in this chapter.
similar to the molecular bond length, the spin density delocalized over the two Fe centers, and a nearly symmetrically bridged CO ligand. Overall, this study illustrates the mechanistic roles of the ammonium-hydride interaction, flexibility of the bridging CO ligand, and intramolecular electron transfer between the Fe centers in the catalytic cycle. Such insights will assist in the design of more effective bioinspired catalysts for H₂ production.

2.2. Introduction

The design of H₂ oxidation and production electrocatalysts is critical for the development of alternative renewable energy technologies. The [FeFe]-hydrogenase enzymes catalyze these reactions effectively with binuclear iron metal centers, as depicted in Figure 2.1. The detailed catalytic mechanism of these enzymes is still not completely understood.

![Figure 2.1. Oxidized (H₉ox) and reduced (H₉red) intermediates of the [FeFe]-hydrogenase enzyme’s active site with the azadithiolate (adt) cofactor.](image)

Fourier Transform infrared spectroscopy (FTIR) and electron paramagnetic resonance (EPR) studies on the native enzyme have identified a mixed-valence oxidized state, H₉ox, and a diamagnetic reduced state, H₉red. The H₉ox state is well-characterized and features an obvious site for binding the H₂ substrate on the distal Fe center, denoted Fe₉d in Figure 2.1. The structure of the H₉red state is less clear because X-ray crystallography of biological systems cannot easily resolve hydrogen atoms. Although hypothesized for more than a decade, an azadithiolate bridgehead with an amine group pendant to the Fe metal center has only recently been confirmed experimentally. This recent insight is guiding current modeling and computational efforts.
Relative to biophysical studies, an advantage of model systems is that hydride and dihydrogen species are readily observed by $^1$H NMR and, in some cases, X-ray crystallography.$^{11,12}$ Moreover, model systems are more computationally tractable, thereby providing additional structural and mechanistic insights.$^{8,13-33}$

A variety of H$_2$-producing model systems inspired by [FeFe]-hydrogenases have been synthesized.$^{34-38}$ Many of these models suffer from high overpotentials (i.e., negative reduction potentials) for hydrogen production. The recent H$_2$-producing models, Fe$_2$(xdt)(CO)$_2$(dppv)$_2$ (xdt = pdt, adt), designed by Rauchfuss and coworkers rely on an electrode in place of the [Fe$_4$S$_4$] clusters that normally support electron transfer in the enzyme. Furthermore, in these models the CN and one of the CO ligands are replaced by diphosphine (dppv) ligands.$^{38}$ The presence of sterically crowded, electron-rich ligands on diiron carbonyl has been shown to stabilize the desired terminal hydride species.$^{35,39}$ Through this approach, key hydride intermediates have been crystallographically characterized.$^{35,38,40,41}$ The most recent advance has revealed a key intermediate wherein the hydride is dihydrogen bonded to the protonated azadithiolate cofactor.$^{38}$

This chapter analyzes computationally the recent advances in the models and biological chemistry of the [FeFe]-hydrogenases, and is composed of two components: (i) benchmarking the computational methodology for these model systems and (ii) providing insights into the mechanism of H$_2$ production. The benchmarking provides the validation that is necessary to support the more hypothetical mechanistic predictions. In the first part of this paper, the computational methods are benchmarked by comparing experimental and theoretical structures, infrared (IR) vibrational frequencies of the CO groups, reduction potentials, and relative pK$_a$ values. The second part of this paper focuses on the mechanism for H$_2$ production. In particular, we examine these model systems in the various oxidation states and protonation states along
proposed mechanistic pathways for H₂ production. Of particular interest, we aim to clarify the nature of the doubly protonated ammonium-hydride intermediate, which has been characterized crystallographically for this bimetallic system. Such doubly protonated intermediates have not been observed in bioinspired monometallic Ni-based catalysts, although a mono-Fe catalyst in the doubly protonated state has very recently been characterized by DuBois and coworkers. We also present additional experimental data to validate the theoretical studies and to test predictions generated by the calculations.

2.3. Defining the Nomenclature

The diiron model systems can adopt numerous isomeric and geometric forms. Borrowing the nomenclature generally employed in enzymatic studies, the Fe center with the hydride is denoted distal (Fe₃), and the opposite Fe center is denoted proximal (Fe₄). Herein, we will abbreviate the (Fe₃Fe₄)(xdt)(CO)₂(dppv)₂ catalysts as [(FeFe)xdt] for notational simplicity. Figure 2.2 depicts the electrocatalyst in the terminal hydride (tH) form with xdt representing either a propanedithiolate (pdt), [tH-(Fe²⁺Fe²⁺)pdt]⁺, or an azadithiolate (adt), [tH-(Fe²⁺Fe²⁺)adt]⁺. In this notation, the terminal hydride is always assumed to reside on the Fe₃ center.

Figure 2.2. Model of the H_red state of the enzyme, where the role of the [Fe₄S₄]²⁺ cluster is replaced by an electrode and the other ligands are replaced by 1,2-bis(diphenylphosphino)ethylene (dppv). The x represents the heteroatom in the dithiolate bridge (xdt=pdt, adt). The configuration shown is the terminal hydride, [tH-(Fe²⁺Fe²⁺)xdt]⁺.
Figure 2.3. Schematic depiction of the various [(FeFe)adt] species, where each row represents a different protonation state of the catalyst. Only the isomers with the lowest free energies are shown for each species in this figure. The dashed lines (P1, P2, P3, P4, and P5) correspond to pathways for which structural and electronic changes can be monitored by $\Delta \nu_{\text{CO}}$ values; they do not necessarily correspond to reaction pathways.

Figure 2.3 depicts all of the protonation states of the [(FeFe)adt] species studied in this work. A similar scheme for the [(FeFe)pdt] species is available in the Supporting Information (Figure S17). Although the terminal hydride ($tH$) species are thermodynamically less stable than their associated bridging hydride ($\mu H$) isomers in model systems, experiments indicate that the terminal hydride species are much more catalytically active: they generate $H_2$ faster and at less negative potentials than their bridging hydride counterparts.$^{35,38}$ The dependence of the reduction potential on the regiochemistry of the hydride remains unexplained but appears to be catalytically relevant. Whereas the [(Fe$^{1}$Fe$^{1}$)pdt] complex requires strong acids to form the terminal hydride
species, \([tH-(Fe^{II}Fe^{II})pdt]^+\), the nearly isostructural \([(Fe^{I}Fe^{I})adt]\) complex can be protonated with weak acids due to the proposed relay activity of the azadithiolate.\(^{34}\) Specifically, the presence of the azadithiolate bridge permits the formation of the ammonium tautomer, \([(Fe^{I}Fe^{I})adt-H]^+\), and subsequent intramolecular proton transfer to the Fe\(_d\) leads to the terminal hydride species, \([tH-(Fe^{II}Fe^{II})adt]^+\) (Figure 2.3). Both \([tH-(Fe^{II}Fe^{II})adt]^+\) and \([tH-(Fe^{II}Fe^{II})pdt]^+\) convert to their respective bridging hydride isomers irreversibly at room temperature.\(^{38,44}\) Upon equilibration, two bridging hydride species form, \(sym-[\mu H-(Fe^{II}Fe^{II})adt]^+\) and \(unsym-[\mu H-(Fe^{II}Fe^{II})adt]^+\), which differ in the orientation of the dppv ring on the Fe\(_d\) center relative to the dppv ring on the Fe\(_p\) center (Figure 2.3).

The adt-containing hydride species can undergo a second protonation with stronger acids, leading to the doubly protonated species, \([tH-(Fe^{II}Fe^{II})adt-H]^2+\) (Figure 2.3). Again, \([H-(Fe^{II}Fe^{II})adt-H]^2+\) isomerizes into two bridging hydrides, \(sym-[\mu H-(Fe^{II}Fe^{II})adt-H]^2+\) and \(unsym-[\mu H-(Fe^{II}Fe^{II})adt-H]^2+\), albeit at a slower rate than the singly protonated analog.\(^{38}\) Although less thermodynamically stable than the bridging hydride species, the terminal hydride species, \([tH-(Fe^{II}Fe^{II})adt-H]^2+\), is poised for \(H_2\) production due to the close proximity of the ammonium and Fe-hydride centers in the doubly protonated species.

2.4. Electronic Structure Methods

Density functional theory (DFT) calculations were performed to characterize the structural and energetic properties of the pertinent species in the catalytic scheme. All calculations were performed using the Gaussian 09 electronic structure program.\(^{45}\) For geometry optimizations, the starting geometries were obtained from the corresponding crystal structures, which are available for the \([(Fe^{I}Fe^{I})pdt]^0,\) \([(Fe^{I}Fe^{II})pdt]^+,\) \(unsym-[\mu H-(Fe^{II}Fe^{II})pdt]^+,\) \(sym-[\mu H-(Fe^{II}Fe^{II})pdt]^+\) and \(sym-[\mu H-(Fe^{II}Fe^{II})pdt]^0\) species,\(^{41}\) as well as the \([(Fe^{I}Fe^{I})adt]^0\) and \([tH-(Fe^{II}Fe^{II})adt-H]^2+\) species.\(^{38}\) The
starting geometries for species without crystal structures were obtained by manually altering the most closely related crystal structure. Optimizations in the gas phase and solution phase resulted in similar structures for several representative species, so gas phase optimizations were used herein for computational efficiency. A comparison of the structures optimized in the gas phase and solution phase for several representative species is available in the Supporting Information (Table S2).

The DFT calculations were performed using the B3P86 density functional\textsuperscript{47,48} with the Stuttgart pseudopotential and associated basis set of Preuss and coworkers (SDD) for the Fe centers,\textsuperscript{49} the 6-31G** basis set for the active hydrogens,\textsuperscript{50} the 6-31+G* basis set for the BF\textsubscript{4}\textsuperscript{-} counterions,\textsuperscript{51} and the 6-31G* basis set for all other atoms.\textsuperscript{52,53} To obtain the solvation free energies, we used the conductor-like polarizable continuum model (C-PCM)\textsuperscript{54,55} with the Bondi atomic radii\textsuperscript{56} and included nonelectrostatic interactions, namely the dispersion,\textsuperscript{57,58} repulsion,\textsuperscript{57,58} and cavitation energies.\textsuperscript{59} The solvation free energies were calculated in dichloromethane (CH\textsubscript{2}Cl\textsubscript{2}) to coincide with the experiments unless otherwise stated.

\textbf{2.5. Calculation of Vibrational Frequencies, Reduction Potentials, and }ΔpK\textsubscript{a} \textbf{Values}

The CO vibrational frequencies were calculated within the harmonic model and were scaled by a factor of 0.9850, as indicated by previous studies.\textsuperscript{60} The vibrational frequency shifts, Δν\textsubscript{CO}, were calculated for the dashed pathways shown in Figure 2.3. For the $[t\text{H}\text{-(Fe}^{II}\text{Fe}^{II})\text{adt}]^+$ and $[t\text{H}\text{-(Fe}^{II}\text{Fe}^{II})\text{adt-H}]^{2+}$ species, the presence of the semibridging CO ligand allows for the assignment of two distinguishable CO stretching frequencies: the higher ν\textsubscript{CO} corresponds to the terminal CO attached to the Fe\textsubscript{p} center, and the lower ν\textsubscript{CO} corresponds to the semibridging CO. For all other species, both CO ligands are terminal, so their stretching frequencies are less separated.
In general, DFT does not consistently provide reliable absolute $pK_a$'s and reduction potentials because of limitations in the functionals, basis sets, and solvation models, as well as uncertainties in the free energies of the electron, proton, and electrode. As discussed previously,$^{61}$ the calculation of $pK_a$'s and reduction potentials relative to related reference systems with known experimental values has been shown to be more reliable because many of the systematic errors cancel. For this reason, we discuss only relative reduction potentials and $pK_a$'s in this paper. The structures and energies of all complexes discussed herein are provided in the SI.

The reaction free energies associated with reduction and deprotonation were calculated using the Born-Haber thermodynamic cycle. The reaction free energy for reduction, $\Delta G_r^o(e^-)$, was calculated using the expression

$$\Delta G_r^o(e^-) = \Delta G_{\text{gas}}^o + \Delta G_s^o(\text{Red}) - \Delta G_s^o(\text{Ox}).$$

(1)

In this expression, $\Delta G_s^o(\text{Red})$ and $\Delta G_s^o(\text{Ox})$ are the solvation free energies of the reduced and oxidized species, respectively, and $\Delta G_{\text{gas}}^o$ is the gas phase reaction free energy calculated using the standard relation $\Delta G_{\text{gas}}^o = \Delta H_{\text{gas}}^o - T\Delta S_{\text{gas}}^o$. A similar expression was used to calculate the reaction free energy for deprotonation, $\Delta G_r^o(H^+)$.

For reasons mentioned above, the reduction potentials were calculated relative to a related reference reaction for which the experimental value is known. This procedure accounts for systematic errors associated with the selected theoretical method and basis set and avoids the determination of the free energies of the electron and the electrode because these terms cancel. Detailed descriptions of this methodology can be found elsewhere.$^{27,61-63}$ For our specific application, we used the following expression to calculate the reduction potentials:

$$E^o = E_{\text{ref}}^o - \Delta \Delta G_r^o(e^-) / F.$$

(2)
Here, $E^o$ is the reduction potential of the [(FeFe)adt] species of interest; $E^o_{\text{ref}}$ is the experimentally measured reduction potential of the reference species, which is the [(FeFe)pdt] species in this work; $\Delta G^o_r(e^-)$ is the calculated difference in reaction free energies (Eq. (1)) for reduction of the species of interest and the reference; and $F$ is the Faraday constant. All calculated and experimental reduction potentials are given in volts (V) relative to the ferrocene/ferrocenium (Fc$^0$/Fc$^+$) couple in CH$_2$Cl$_2$. The irreversible reduction potentials of the doubly protonated species are calculated relative to the reduction potentials of the corresponding singly protonated species and are labeled $\Delta E^o_{\text{prot}}$. The cyclic voltammetry (CV) experiments were performed at 273 K, and the reduction potentials were also calculated at this temperature.$^{38}$

We used the standard relationship, $pK_a = \Delta G^o_r(H^+)/ (RT \ln 10)$, to calculate the $pK_a$ values. In this expression, $\Delta G^o_r(H^+)$ corresponds to the calculated reaction free energy for deprotonation, $R$ is the gas constant, and $T = 233.15$ K to coincide with experimental conditions. Due to the limited experimental $pK_a$ data for the [(FeFe)pdt] systems, and its inherent inability to doubly protonate, these complexes were not used as references. Instead, we calculated $\Delta pK_a$, which is defined to be the difference between the $pK_a$ values for two specified states of the [(FeFe)adt] complex. The reported $\Delta pK_a$ values were calculated in CH$_2$Cl$_2$ to coincide with experimental conditions and in acetonitrile (CH$_3$CN) to coincide with the reference $pK_a$ values used in the interpretation of the experimental values because a $pK_a$ scale is not well established for CH$_2$Cl$_2$.

2.6. Geometries, Free Energies, and Vibrational Frequencies

We optimized the geometries of the various species discussed in the Introduction and compared the structures and key vibrational frequencies to available experimental data. The optimized geometries are in reasonable agreement with the X-ray crystal structures. Table 2.1
provides a comparison of the most relevant bond lengths for the various species. Structural and energetic information of all systems studied herein are provided in the Supporting Information. The relative free energies of the species depicted in Figure 2.3 are provided in Table 2.2. Moreover, the relative free energies of these species and other isomers of higher free energy are provided in Figures S15 and S16.

The optimized geometry for the doubly protonated species, \([\mu H-(Fe^{II}Fe^{II})adt-H]^{2+}\), exhibited one key difference from the crystal structure, namely the dihydrogen distance \(NH\cdots HFe\). This dihydrogen distance was \(d_{H\cdots H} = 1.88\ \text{Å}\) in the crystal structure and \(d_{H\cdots H} = 1.40\ \text{Å}\) in the optimized geometry. Thus, DFT leads to a significantly shorter dihydrogen distance that corresponds to a stronger dihydrogen bond. A previous DFT study on a model of the \([FeFe]\)-hydrogenase active site also showed an underestimation of the dihydrogen distance for a related

<table>
<thead>
<tr>
<th>species</th>
<th>(\text{Fe}^{2+}\cdots \text{Fe}^{2+})</th>
<th>(\text{Fe}^{2+}\cdots \text{CO})</th>
<th>(\text{Fe}^{0}\cdots \text{CO})</th>
<th>(\text{Fe}^{2+}\cdots H)</th>
<th>(\text{Fe}^{2+}\cdots H)</th>
</tr>
</thead>
<tbody>
<tr>
<td>([\text{Fe}^{2+}\text{Fe}^{2+}]^{0})</td>
<td>2.60</td>
<td>1.73</td>
<td>1.73</td>
<td>g</td>
<td>g</td>
</tr>
<tr>
<td>([\text{Fe}^{2+}\text{Fe}^{2+}]^{1})</td>
<td>2.59</td>
<td>1.75</td>
<td>1.74</td>
<td>g</td>
<td>g</td>
</tr>
<tr>
<td>([\text{Fe}^{2+}\text{Fe}^{2+}]^{2})</td>
<td>2.66</td>
<td>1.78</td>
<td>1.74</td>
<td>1.69</td>
<td>1.68</td>
</tr>
<tr>
<td>([\text{Fe}^{2+}\text{Fe}^{2+}]^{3})</td>
<td>2.80</td>
<td>1.75</td>
<td>1.74</td>
<td>1.78</td>
<td>1.76</td>
</tr>
<tr>
<td>([\text{Fe}^{2+}\text{Fe}^{2+}]^{4})</td>
<td>2.60</td>
<td>1.74</td>
<td>1.73</td>
<td>1.75</td>
<td>1.73</td>
</tr>
<tr>
<td>([\text{Fe}^{2+}\text{Fe}^{2+}]^{5})</td>
<td>2.62</td>
<td>2.52</td>
<td>2.50</td>
<td>1.79</td>
<td>1.78</td>
</tr>
</tbody>
</table>

\(^\text{a}\) Values given in Å. \(^\text{b}\) Experimental data from ref. 40. \(^\text{c}\) Experimental data from ref. 46. \(^\text{d}\) Experimental data from ref. 35. \(^\text{e}\) Experimental data from ref. 41. \(^\text{f}\) Not applicable because the hydrogen is not present or the distance is not relevant for the doubly protonated species.

The optimized geometry for the doubly protonated species, \([\mu H-(Fe^{II}Fe^{II})adt-H]^{2+}\), exhibited one key difference from the crystal structure, namely the dihydrogen distance \(NH\cdots HFe\). This dihydrogen distance was \(d_{H\cdots H} = 1.88\ \text{Å}\) in the crystal structure and \(d_{H\cdots H} = 1.40\ \text{Å}\) in the optimized geometry. Thus, DFT leads to a significantly shorter dihydrogen distance that corresponds to a stronger dihydrogen bond. A previous DFT study on a model of the \([FeFe]\)-hydrogenase active site also showed an underestimation of the dihydrogen distance for a related
doubly protonated species. This smaller dihydrogen distance in the $[tH-(Fe^{II}Fe^{II})adt-H]^{2+}$ species was found for a wide range of density functionals, as shown in Table S3. However, dihydrogen bond distances have also been shown to be extremely sensitive to weak intra- and intermolecular interactions, such as ion-pairing.\textsuperscript{66} Moreover, neutron diffraction experiments are expected to provide more reliable positions for hydrogen atoms than X-ray diffraction, which most likely overestimates this distance.\textsuperscript{67} A recent neutron diffraction study on a monometallic Fe hydrogenase model using the noncoordinating counterion, BAr$^{+}$, found the NH···HFe dihydrogen bond to be 1.49Å.\textsuperscript{43} In contrast, the doubly protonated structure, $[tH-(Fe^{II}Fe^{II})adt-H]^{2+}$, was determined with X-ray crystallography and included BF$_4^-$ ions that could exhibit significant ion-pairing effects.\textsuperscript{68}

Given the difference in the crystallographic and calculated dihydrogen bond length in $[tH-(Fe^{II}Fe^{II})adt-H]^{2+}$, we hypothesized that this discrepancy was due to the omission of the BF$_4^-$ ions in the geometry optimizations. To test this hypothesis, we also optimized the structure including one or two of the BF$_4^-$ ions closest to the azadithiolate bridge in the crystal structure. Figure 2.4 illustrates that the geometry optimizations including either one or two BF$_4^-$ ions improved the agreement between DFT and experiment by shifting the amine bridgehead and elongating the dihydrogen distance. The bond lengths given in Table 2.3 show that including the ion-pairing interactions preserves the Fe$\_p$···Fe$\_d$ and Fe$\_d$···H distances while elongating the dihydrogen distance. Although the inclusion of counterions improves the agreement with the doubly protonated crystal structure, we omit the counterions in the calculations of other quantities, such as reduction potentials and spin densities, to maintain consistency among the various species studied under a range of experimental conditions.
Figure 2.4. Superimposed structures of the doubly protonated species, \([tH-(Fe^{II}Fe^{II})adt-H]^2+\), illustrating the better agreement between the crystal structure and the optimized geometries when BF\(_4^-\) ions are included in the calculations. The phenyl groups have been removed and the Fe\(_d\)H···HN interaction is circled for clarity. Color Scheme: crystal structure (gray) and optimized geometry with no BF\(_4^-\) (red), one BF\(_4^-\) (blue), and two BF\(_4^-\) (green) counterions.

Table 2.3. Selected Bond Lengths from DFT Calculations and X-Ray Structures for Doubly Protonated Species Optimized with and without BF\(_4^-\) Anions\(^{a,b}\)

<table>
<thead>
<tr>
<th></th>
<th>Fe(_d)···Fe(_d)</th>
<th>Fe(_d)···H</th>
<th>N···H</th>
<th>H···H</th>
</tr>
</thead>
<tbody>
<tr>
<td>Expt(^c)</td>
<td>2.62</td>
<td>1.44</td>
<td>1.00</td>
<td>1.88</td>
</tr>
<tr>
<td>No BF(_4^-)</td>
<td>2.65</td>
<td>1.53</td>
<td>1.07</td>
<td>1.40</td>
</tr>
<tr>
<td>Solution(^d)</td>
<td>2.61</td>
<td>1.51</td>
<td>1.04</td>
<td>1.55</td>
</tr>
<tr>
<td>1·BF(_4^-)</td>
<td>2.62</td>
<td>1.51</td>
<td>1.04</td>
<td>1.65</td>
</tr>
<tr>
<td>2·BF(_4^-)</td>
<td>2.59</td>
<td>1.50</td>
<td>1.04</td>
<td>2.04</td>
</tr>
</tbody>
</table>

\(^a\)Values given in Å. \(^b\)All optimizations were conducted in the gas phase unless stated otherwise. \(^c\)Experimental data from ref 38. \(^d\)This structure was optimized in solution (CH\(_2\)Cl\(_2\)) using C-PCM without any BF\(_4^-\) counterions.

The strong, distinctive absorptions at ~2000 cm\(^{-1}\) associated with CO stretching can be used to probe changes in the electronic structure due to protonation, isomerization, reduction, and oxidation.\(^{13,23,69,70}\) We calculated the change in the CO frequencies, \(\Delta \nu_{CO}\), for the five protonation reactions depicted by dashed lines in Figure 2.3. As shown in Table 2.4, the calculated \(\Delta \nu_{CO}\) values due to protonation are within ~10 cm\(^{-1}\) of the experimentally measured values. This level of agreement provides validation for the structures of species for which crystal structures are not
Typically DFT calculations are more reliable for changes in frequencies than for absolute frequencies; a comparison between the calculated and experimental absolute frequencies is provided in Table S4. For the present study, the previously reported\textsuperscript{38} spectroscopic studies were repeated using the acid [H(OEt$_2$)$_2$]BARF$_4$ to eliminate the possibility of ion pairing arising from the coordination of BF$_4^-$ when using HBF$_4$·Et$_2$O, which has been shown to participate in hydrogen-bonding interactions,\textsuperscript{71} particularly in complexes with protonated adt ligands.\textsuperscript{72} The agreement between calculated and experimental frequencies is better for the BARF$_4^-$ salts; these spectra are provided in the Supporting Information (Figures S6-S7).

The $\Delta\nu_{\text{CO}}$ values reflect the protonation sites, as indicated by both experimental measurements and theoretical calculations. The $\Delta\nu_{\text{CO}}$ values associated with $N$-protonation are typically characterized by CO frequency increases of $\sim$20 cm$^{-1}$.\textsuperscript{73-75} The experimental and calculated $\Delta\nu_{\text{CO}}$ values associated with $N$-protonation to form either the singly or doubly protonated species (P1, P4, and P5 protonations) exhibit this shift. While $N$-protonation is associated with a relatively small $\nu_{\text{CO}}$ shift, formation of the terminal and bridging hydrides (P2 and P3 protonations, respectively) are associated with much higher shifts of $\sim$70 cm$^{-1}$. In both cases, two well-resolved bands are observed in the spectra, but the bands are much more energetically separated for the terminal hydride species than for the bridging hydride species. The band at lower frequencies in the terminal hydride species is assigned to the semibridging CO, which exhibits a smaller $\nu_{\text{CO}}$ shift of $\sim$40 cm$^{-1}$. Our calculated $\Delta\nu_{\text{CO}}$ values for the protonations leading to the terminal and bridging hydride species (P2 and P3 protonations, respectively) demonstrate this trend and agree well with the experimental values (Table 2.4). An analogous figure and table for the [(FeFe)pdt] complexes is provided in the Supporting Information (Figure S17 and Table S1).
Table 2.4. Experimental and Calculated $\Delta \nu_{\mathrm{CO}}$ Values upon Protonation

<table>
<thead>
<tr>
<th></th>
<th>$\text{expt}^c$</th>
<th>$\text{calc}^d$</th>
</tr>
</thead>
<tbody>
<tr>
<td>P1</td>
<td>22, 22</td>
<td>31, 30</td>
</tr>
<tr>
<td>P2$^e$</td>
<td>78, 46</td>
<td>86, 37</td>
</tr>
<tr>
<td>P3</td>
<td>82, 79</td>
<td>89, 77</td>
</tr>
<tr>
<td>P4$^e$</td>
<td>24, 21</td>
<td>27, 34</td>
</tr>
<tr>
<td>P5</td>
<td>19, 17</td>
<td>29, 29</td>
</tr>
</tbody>
</table>

$^a$Values given in cm$^{-1}$. $^b$The P1, P2, P3, P4, and P5 protonation reactions are defined in Figure 2.3. $^c$Experimental data from ref 38. $^d$Frequencies scaled by a factor of 0.9850. $^e$The second $\Delta \nu_{\mathrm{CO}}$ value is assigned to the semibridging CO.

2.7. Reduction Potentials

The calculated reduction potentials for the species in CH$_2$Cl$_2$ are presented in Table 2.5. Because the [(FeFe)pdt] species are used as the reference systems, these reduction potentials agree exactly with experiment by construction, and all other reduction potentials are calculated relative to these values. The agreement between theory and experiment for the [(FeFe)adt] species is reasonable for the singly protonated species with errors of $\sim$0.05 V vs Fe$^{0/+}$ in CH$_2$Cl$_2$. Additional calculated reduction potentials are provided in the Supporting Information (Table S5). A detailed analysis of the metal oxidation states of the reduced species in terms of the spin densities is provided in the next subsection.

Table 2.5. Experimental and Calculated Reduction Potentials

<table>
<thead>
<tr>
<th>species</th>
<th>$E^o$ [(FeFe)$^{1+}$]$^g$</th>
<th>$E^o$ [H$^{-}$[(FeFe)$^{1+}$]$^h$</th>
<th>$E^o$ unsym-[µH-[(FeFe)$^{1+}$]$^i$</th>
</tr>
</thead>
<tbody>
<tr>
<td>[(FeFe)pdt]$^y$</td>
<td>expt$^f$ -0.94</td>
<td>-1.70$^i$</td>
<td>-1.80</td>
</tr>
<tr>
<td></td>
<td>calc$^f$ -0.94</td>
<td>-1.70$^i$</td>
<td>-1.80$^i$</td>
</tr>
<tr>
<td>[(FeFe)adt]$^y$</td>
<td>expt$^f$ -0.76</td>
<td>-1.66$^g$</td>
<td>-1.79</td>
</tr>
<tr>
<td></td>
<td>calc$^f$ -0.84</td>
<td>-1.69</td>
<td>-1.77$^h$</td>
</tr>
</tbody>
</table>

$^a$Values given in volts vs Fe$^{0/+}$ in dichloromethane (CH$_2$Cl$_2$). Experimental $E_{1/2}$ values given unless otherwise indicated. $^b$[(FeFe)pdt] reactions were used as the references reactions, so the calculated and experimental values agree by construction. $^c$Experimental data from ref 76. $^d$Quasi-reversible reaction so $E_p$ is reported. $^eE^o = -1.82$ V vs Fe$^{0/+}$ in CH$_2$Cl$_2$ if calculated using the symmetric bridging form. $^f$Experimental data from ref 38. $^g$Irreversible reaction so $E_p$ is reported. $^hE^o = -1.83$ V vs Fe$^{0/+}$ in CH$_2$Cl$_2$ if calculated using the symmetric bridging form.

Due to difficulties in determining a suitable reference system for the doubly protonated species, we calculated the difference between the reduction potentials for the doubly and singly
protonated species ($\Delta E^\circ_{\text{prot}}$). Using this formulation, we calculated $\Delta E^\circ_{\text{prot}} = 0.59$ V vs Fe$^{0+/+}$ in CH$_2$Cl$_2$. As expected from basic electrostatic arguments, the doubly protonated species has a less negative reduction potential than the singly protonated species. The experimentally measured difference in these reduction potentials was 0.39 V vs Fe$^{0+/+}$ in CH$_2$Cl$_2$. As discussed above, this error could be due to the neglect of BF$_4^-$ in the calculations. Note that BF$_4^-$ ions were present in the electrochemical measurements for the doubly protonated species but not for the singly protonated species. Furthermore, the effects of ion pairing on reduction potentials are known to be pronounced in solvents of low dielectric constants, such as CH$_2$Cl$_2$. Thus, the difference in the counterions present for the electrochemical experiments of the singly and doubly protonated species could lead to discrepancies in the comparison between calculated and experimental results. Note that the experimentally measured reduction potentials for the terminal hydride species are for quasireversible couples, which could introduce additional discrepancies.

2.8. Analysis of Metal Oxidation States

For species with an unpaired electron, we analyzed the spin densities to determine the localization of the unpaired electron. The spin densities for the reduced species are presented in Table 2.6. For the reduced terminal hydride species [tH-(Fe$^\text{I}$Fe$^\text{II}$)adt]$^0$, the unpaired spin density is localized entirely on the Fe$_p$ center, suggesting that the Fe$_p$ center is reduced. This localization on the Fe$_p$ center is similar for the reduced terminal hydride species [tH-(Fe$^\text{I}$Fe$^\text{II}$)pdt]$^0$, which is produced at a similar potential as the [tH-(Fe$^\text{I}$Fe$^\text{II}$)adt]$^0$ species, indicating that the dithiolate is not a major influence on the redox behavior for the terminal hydride species. The spin densities pertaining to the [(FeFe)pdt] complexes are provided in the Supporting Information (Table S6).
Table 2.6. Spin Densities of the Reduced Mixed-Valence Species

<table>
<thead>
<tr>
<th>species</th>
<th>ρ(Feₚ)</th>
<th>ρ(Feₚ)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$[\mu\text{H}-(\text{Fe}^{II}\text{Fe}^{II})\text{adt}]^+$</td>
<td>1.03</td>
<td>0.05</td>
</tr>
<tr>
<td>sym-$[\mu\text{H}-(\text{Fe}^{1.5}\text{Fe}^{1.5})\text{adt}]^0$</td>
<td>0.76</td>
<td>0.37</td>
</tr>
<tr>
<td>sym-$[\mu\text{H}-(\text{Fe}^{1.5}\text{Fe}^{1.5})\text{adt-H}]^+$</td>
<td>0.81</td>
<td>0.29</td>
</tr>
<tr>
<td>unsym-$[\mu\text{H}-(\text{Fe}^{II}\text{Fe}^{II})\text{adt}]^0$</td>
<td>0.98</td>
<td>0.15</td>
</tr>
<tr>
<td>unsym-$[\mu\text{H}-(\text{Fe}^{II}\text{Fe}^{II})\text{adt-H}]^+$</td>
<td>0.96</td>
<td>0.14</td>
</tr>
</tbody>
</table>

Previous studies also examined the metal oxidation states upon reduction of $[\mu\text{H}-(\text{Fe}^{II}\text{Fe}^{II})\text{pdt}]^+$ by analyzing the spin density with the BP86 functional and EPR spectra of the mixed-valence species.⁴¹ These studies showed that reduction of sym-$[\mu\text{H}-(\text{Fe}^{II}\text{Fe}^{II})\text{pdt}]^+$ resulted in a species with the unpaired electron delocalized between the Feₚ and Feₚ centers, sym-$[\mu\text{H}-(\text{Fe}^{1.5}\text{Fe}^{1.5})\text{pdt}]^0$. However, reduction of unsym-$[\mu\text{H}-(\text{Fe}^{II}\text{Fe}^{II})\text{pdt}]^+$ resulted in a species with the unpaired electron localized on the Feₚ center, unsym-$[\mu\text{H}-(\text{Fe}^{II}\text{Fe}^{II})\text{pdt}]^0$. Our calculations revealed a similar trend with the bridging hydride species of the singly and doubly protonated [(FeFe)adt] (Table 2.6), although several important differences were observed.

In the singly protonated symmetric and unsymmetric [(FeFe)adt] bridging hydride species, sym-$[\mu\text{H}-(\text{Fe}^{II}\text{Fe}^{II})\text{adt}]^+$ and unsym-$[\mu\text{H}-(\text{Fe}^{II}\text{Fe}^{II})\text{adt}]^+$, the orientations of the dppv ligands on the two Fe centers dictate the overall symmetry of the system, as depicted in Figure 2.3. Furthermore, the hydride is located symmetrically between the two Fe centers in both isomers for the oxidized state. These distances are reported in Figure 2.5.
Figure 2.5. Bond lengths (Å) of symmetric and unsymmetric bridging hydrides. The bond lengths given without parentheses correspond to the oxidized species as drawn while those given in parentheses correspond to the reduced mixed-valence species (not drawn). The spin densities of the reduced mixed-valence species are reported in Table 2.6.

Upon reduction, the bridging hydride is perturbed from this symmetric position. In the reduced unsymmetric bridging hydride species, the hydride is more closely bound to Fe\textsubscript{d} ($d_\text{Fe-d···H}=1.58$ Å) than to Fe\textsubscript{p} ($d_\text{Fe-p···H}=2.05$ Å) and resembles a terminal hydride. In the reduced symmetric bridging hydride species, the hydride is more symmetrically bound to the two iron centers ($d_\text{Fe-d···H}=1.60$ Å for Fe\textsubscript{d} and $d_\text{Fe-p···H}=1.77$ Å for Fe\textsubscript{p}). The observation that the reduction of $\text{sym-}[\mu\text{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{adt}]^+$ maintains a more symmetric bridging hydride is consistent with the observation that the spin density is somewhat delocalized between the two Fe centers. In contrast, reduction of $\text{unsym-}[\mu\text{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{adt}]^+$ leads to a much more asymmetric bridged hydride, and the spin density is localized on Fe\textsubscript{p}, which does not have a bound hydride. Thus, $\text{unsym-}[\mu\text{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{adt}]^+$ is similar to $[\mu\text{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{adt}]^+$ in that reduction occurs at the Fe\textsubscript{p} site, but reduction
of unsym-$[\mu \mathbf{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{adt}]^+$ also involves motion of the bridged hydride toward the Fe$_d$ site to resemble a terminal hydride. These observations are consistent with the observation that the terminal hydride species is reduced more easily (i.e., at a slightly less negative potential) than the unsymmetric bridging hydride species (Table 2.5).

The presence of the amine group in the azadithiolate bridgehead affects the degree of asymmetry observed upon reduction of the symmetric bridging hydride species. Although the bridging hydride is more symmetric in the $\text{sym}-[\mu \mathbf{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{adt}]^+$ species than in the $\text{unsym}-[\mu \mathbf{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{adt}]^+$ species, more asymmetry is exhibited in the $\text{sym}-[\mu \mathbf{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{adt}]^+$ than in the analogous $\text{sym}-[\mu \mathbf{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{pdt}]^+$ species. In the previous study of [(FeFe)pdt] bridging hydride species, reduction of $\text{sym}-[\mu \mathbf{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{pdt}]^+$ preserves the symmetry of the hydride position to a higher degree, with $d_{\text{Fe} \cdots \text{H}}=1.68$ Å and $d_{\text{Fe} \cdots \text{H}}=1.70$ Å for the Fe$_p$ and Fe$_d$ centers, respectively.\(^\text{41}\)

We have confirmed these observations using the level of theory described within the present paper and various other basis sets. These results are given in the Supporting Information (Table S7). Moreover, these nearly symmetric Fe···H bond lengths are associated with nearly equal spin densities on the two Fe centers. As shown in Figure 2.5 and Table 2.6, the reduced symmetric protonated [(FeFe)adt] system, $\text{sym}-[\mu \mathbf{H}-(\text{Fe}^{1.5}\text{Fe}^{1.5})\text{adt}]^0$, exhibits more asymmetry in these Fe···H bond lengths and unequal delocalized spin densities on the two Fe centers.

The reduction of the doubly protonated bridging hydride species, $\text{unsym}-[\mu \mathbf{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{adt-H}]^2+$ and $\text{sym}-[\mu \mathbf{H}-(\text{Fe}^{\text{II}}\text{Fe}^{\text{II}})\text{adt-H}]^2+$, follows a similar trend but displays a slightly smaller degree of delocalization of the spin density, as shown in Table 2.6. According to the distances given in Figure 2.5, the asymmetry of the bridging hydride is slightly more pronounced in both the symmetric and unsymmetric forms for the doubly protonated species. Thus, reduction is expected to occur primarily at the non-hydride Fe$_p$ center, as supported by the unpaired spin densities given
in Table 2.6. This pattern for the electron localization is also seen in doubly protonated terminal hydride species, which will be discussed below.

2.9. Analysis of the Protonation Scheme and ΔpKₐ

Previous experiments on the [(FeFe)adt] system estimated a small difference in the pKₐ of the first and second protonation steps (ΔpKₐ ~ −2), however, our calculated values indicated much larger differences, as seen from ΔpKₐ (P4 − P2) in Table 2.7. This discrepancy between experiment and theory prompted a reinvestigation of the acid-base chemistry of the [(FeFe)adt] species. The pKₐ values can be estimated by NMR measurements because each protonated species displays a unique ³¹P NMR spectrum. The experimental details and spectra are provided in the Supporting Information (Figures S12-S13). According to these experiments, the protonation reactions leading to the formation of [(Fe²⁺Fe⁺)adt-H]⁺, [tH-(Fe⁺⁺Fe⁺⁺)adt]⁺, and [tH-(Fe⁺⁺Fe⁺⁺)adt-H]²⁺, defined as P1, P2, and P4 in Figure 2.3, correspond to pKₐ values of >14.2, ~15.5, and ~6, respectively.

<table>
<thead>
<tr>
<th>ΔpKₐ (P2 − P1)</th>
<th>ΔpKₐ (P4 − P2)</th>
<th>ΔpKₐ (P3 − P2)</th>
</tr>
</thead>
<tbody>
<tr>
<td>expt</td>
<td></td>
<td></td>
</tr>
<tr>
<td>CH₃Cl₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>&gt;1.3</td>
<td>6</td>
<td>5</td>
</tr>
<tr>
<td>233.15 K</td>
<td>298.15 K</td>
<td>233.15 K</td>
</tr>
<tr>
<td>CH₃CN</td>
<td></td>
<td></td>
</tr>
<tr>
<td>−9.5</td>
<td>−18</td>
<td>−14</td>
</tr>
<tr>
<td>233.15 K</td>
<td>298.15 K</td>
<td>233.15 K</td>
</tr>
<tr>
<td>&gt;2.0</td>
<td>8</td>
<td>6</td>
</tr>
<tr>
<td>233.15 K</td>
<td>298.15 K</td>
<td>233.15 K</td>
</tr>
</tbody>
</table>

The protonation pathways P1, P2, P3, and P4 are defined in Figure 2.3, and the pKₐ is determined for the protonated species in each case. These experiments were performed in CH₃Cl₂ solvent, but the reference pKₐ values pertain to CH₃CN solvent.

For reasons discussed in the Methods section, we calculated only the relative pKₐ values for these species. This procedure accounts for systematic errors associated with the selected theoretical method and avoids the determination of the free energy of the solvated proton because this term cancels. Note that the experimentally measured values are only qualitatively meaningful because the experiments were performed in CH₃Cl₂, which does not have a well-defined pKₐ scale, and the reference pKₐ values of the acids pertain to CH₃CN. As a result, we calculated the relative
pK\textsubscript{a} values in both CH\textsubscript{2}Cl\textsubscript{2} and CH\textsubscript{3}CN. Additionally, the effects of homoconjugation in these experiments may distort the apparent pK\textsubscript{a} measured. As discussed above in the context of comparing the crystal structure and optimized geometries, the impact of ion pairing is significant for the doubly protonated species. Furthermore, it has been shown experimentally that ion pairing between the BF\textsubscript{4}\textsuperscript{-} counterion and the ammonium center can shift the equilibrium of the singly protonated species toward [(Fe\textsuperscript{II}Fe\textsuperscript{II})adt-H]\textsuperscript{+}.\textsuperscript{38}

The experimental and calculated ΔpK\textsubscript{a} values for these species are given in Table 2.7. The experimental ΔpK\textsubscript{a} (P2–P1) value indicates that the Fe\textsubscript{d} site is more basic than the amine by at least ~1.3 pK\textsubscript{a} units; however, the amine is more readily protonated than the stereochemically crowded Fe site. The greater basicity of the Fe\textsubscript{d} site than the amine is corroborated by our calculated ΔpK\textsubscript{a} (P2–P1) value of ~6 in CH\textsubscript{2}Cl\textsubscript{2}. The large, negative ΔpK\textsubscript{a} (P4–P2) value measured experimentally indicates that N-protonation of [tH-(Fe\textsuperscript{II}Fe\textsuperscript{II})adt]\textsuperscript{+} to form [tH-(Fe\textsuperscript{II}Fe\textsuperscript{II})adt-H]\textsuperscript{2+} requires much stronger acids, as supported by our calculated large, negative ΔpK\textsubscript{a} (P4–P2). Note that this experimentally measured value of ΔpK\textsubscript{a} (P4 – P2) is significantly more negative than the previously measured value.\textsuperscript{38} Again, it is worth noting that the presence of the BF\textsubscript{4}\textsuperscript{-} ions in the experiments may contribute to the remaining discrepancy between theory and experiment. Moreover, the theoretical calculation of pK\textsubscript{a} differences for species with different charges (i.e., +2 and +1 for the doubly and singly protonated species) is more challenging. Turning our attention to the bridging hydride isomers, previous experiments have suggested that the conversion to the bridging hydride isomers is irreversible, and the bridging hydride is estimated to be at least ~2 pK\textsubscript{a} units less acidic than the terminal hydride.\textsuperscript{35,38} This observation is supported by the large, positive calculated ΔpK\textsubscript{a} (P3–P2) value and the thermodynamic stability of the bridging hydride isomers.
These ΔpKₐ trends are reproducible using different functionals, as indicated by additional data provided in the Supporting Information (Table S8).

2.10. Analysis of the H₂ Evolution Mechanism

The proposed mechanisms for hydrogen evolution catalyzed by the [(FeFe)adt] system for weak and strong acids are depicted in Figure 2.6. Initially, the amine group on [(Fe¹Fe¹)adt]₀ is protonated to form the ammonium tautomer, [(Fe¹Fe¹)adt-H]⁺. An intramolecular proton transfer occurs from the N to the Feₖ to form the terminal hydride, [tH-(Fe²Fe²)adt]⁺. For weak acids, the terminal hydride species, [tH-(Fe²Fe²)adt]⁺, is reduced prior to subsequent protonation and reduction to evolve H₂. This pathway is shown in dotted lines in Figure 2.6. For stronger acids, the intramolecular proton transfer is followed by another protonation at the amine group to generate [tH-(Fe²Fe²)adt-H]⁺. Reduction of this doubly protonated species then evolves H₂. A possible mechanism for this process is indicated with solid lines on the right side of Figure 2.6.

The values for the spin densities and bond distances for the species associated with this pathway are given in Table 2.8.

<table>
<thead>
<tr>
<th>species</th>
<th>ρ(Feₚ)</th>
<th>ρ(Feₖ)</th>
<th>Feₚ···Feₖ</th>
<th>Feₖ···H</th>
<th>H···H</th>
<th>Feₚ···CO</th>
<th>Feₖ···CO</th>
</tr>
</thead>
<tbody>
<tr>
<td>[tH-(Fe¹Fe¹)adt-H]⁺</td>
<td>0.00</td>
<td>0.00</td>
<td>2.65</td>
<td>1.53</td>
<td>1.40</td>
<td>2.50</td>
<td>1.78</td>
</tr>
<tr>
<td>[tH-(Fe²Fe²)adt-H]⁺</td>
<td>1.02</td>
<td>0.05</td>
<td>3.03</td>
<td>1.56</td>
<td>1.30</td>
<td>2.84</td>
<td>1.78</td>
</tr>
<tr>
<td>[H₂-(Fe¹Fe¹)adt]⁺</td>
<td>0.43</td>
<td>0.45</td>
<td>2.69</td>
<td>1.75ᵇ</td>
<td>0.79</td>
<td>2.06</td>
<td>1.89</td>
</tr>
<tr>
<td>[(Fe²Fe²)adt]⁺</td>
<td>0.03</td>
<td>1.11</td>
<td>2.59</td>
<td>n/a</td>
<td>n/a</td>
<td>2.51</td>
<td>1.78</td>
</tr>
</tbody>
</table>

ᵇ Values given in Å. This bond length is the average of the two Feₖ···H distances.
Figure 2.6. Possible mechanisms for H$_2$ production with strong or weak acids. The steps indicated with solid lines along the perimeter represent the proposed pathway discussed in the main text. The steps indicated with dotted lines represent the pathway with a weak acid.

Our calculations indicate that reduction of [tH-(Fe$^{II}$Fe$^{II}$)adt-H]$^{2+}$ leads to a mixed-valence complex, [tH-(Fe$^{I}$Fe$^{II}$)adt-H]$^+$. The spin densities and bond distances of this species are given in Table 2.8. This mixed-valence species is characterized by the unpaired electron localized on the Fe$_p$ center, which is in agreement with the reduction scheme exhibited by the [tH-(Fe$^{II}$Fe$^{II}$)adt]$^+$ and [tH-(Fe$^{II}$Fe$^{II}$)pdt]$^+$ species discussed previously. The NH···HFe$_d$ distance is shorter in [tH-(Fe$^{I}$Fe$^{II}$)adt-H]$^+$ ($d_{NH···H} = 1.30 \text{ Å}$) than in [tH-(Fe$^{II}$Fe$^{II}$)adt-H]$^{2+}$ ($d_{NH···H} = 1.40 \text{ Å}$) but is still larger than the distance in molecular H$_2$ ($d_{H···H} = 0.74 \text{ Å}$). Although we were unable to locate a minimum-energy species with H$_2$ weakly bound to an Fe$_d$(I) center, we found another type of H$_2$ adduct intermediate with the unpaired electron delocalized between the Fe$_p$ and Fe$_d$ metal centers. This H$_2$ adduct, [H$_2$-(Fe$^{1.5}$Fe$^{1.5}$)adt]$^+$, is higher in free energy than the mixed-valence doubly protonated
species, $[tH-(Fe^{I}Fe^{II})adt-H]^+$, by 5.82 kcal/mol but has several features that are relevant to its potential catalytic activity, as will be discussed below. Note that previous theoretical studies on small models of the [FeFe]-hydrogenase active site found that reduction of a doubly protonated $Fe_p(II)Fe_d(II)$ species could lead to the formation of a stable mixed-valence $H_2$ adduct intermediate, $Fe_p(I)Fe_d(II)$, where the $Fe_p$ center is the site of reduction and the $H_2$ ligand is loosely bound to the $Fe_d(II)$ center.\textsuperscript{8,19} These mixed-valence intermediates are similar to the $[tH-(Fe^{I}Fe^{II})adt-H]^+$ species found herein; however, the $[H_2-(Fe^{1.5}Fe^{1.5})adt]^+$ species was not found in these previous studies.

In contrast to the $[tH-(Fe^{I}Fe^{II})adt-H]^+$ species, the $[H_2-(Fe^{1.5}Fe^{1.5})adt]^+$ species features a much shorter $NH\cdots HFe_d$ distance ($d_{H\cdots H} = 0.79 \text{ Å}$), suggesting an activated $H_2$ molecule. In conjunction with the activated $H_2$ bond distance, the average distance between each hydrogen and the $Fe_d$ center is elongated ($d_{Fe\cdots H} = 1.75 \text{ Å}$), suggesting that the $H_2$ ligand is weakly bound to the $Fe_d$ center and poised for release. Another important structural feature is that unlike all terminal hydride species reported in the present paper, the semibridging CO ligand in this $H_2$ adduct is nearly symmetrically bridged between the $Fe_p$ and $Fe_d$ centers, where $d_{Fe\cdots CO} = 2.06$ and $1.89 \text{ Å}$, respectively. The $Fe_p\cdots Fe_d$ distance contracts ($d_{Fe\cdots Fe} = 2.69 \text{ Å}$) to accommodate this bridging interaction. This structural property is reminiscent of the enzyme’s $H_{ox}$ state, which features a nearly symmetrically bridged CO ligand (Figure 2.1). Moreover, the spin density is now delocalized between the $Fe_p$ and $Fe_d$ center, 0.43 and 0.45, respectively, warranting the $[H_2-(Fe^{1.5}Fe^{1.5})adt]^+$ assignment. This delocalization of the spin density suggests the possibility for intramolecular electron transfer to occur between the $Fe_p$ and $Fe_d$ centers concomitant with formation of a shorter $H_2$ bond. Note that the proposed mechanism in Figure 2.6 is similar to mechanisms proposed previously for [FeFe]-hydrogenases in terms of the reduction and
Our objective is to characterize the proposed intermediates and to analyze aspects such as dihydrogen bonding and electron delocalization or localization with respect to the metal centers.

Release of H\textsubscript{2} from [(H\textsubscript{2}-(Fe\textsuperscript{1.5}Fe\textsuperscript{1.5})adt\textsuperscript{+}] would generate another mixed-valence state, the [(Fe\textsuperscript{II}Fe\textsuperscript{I})adt\textsuperscript{+}] species shown in Figure 2.6. This mixed-valence species is closely related to the H\textsubscript{ox} state of the enzyme shown in Figure 2.1. Interestingly, the spin densities for this species, which are given in Table 2.8, indicate Fe oxidation states that are opposite to the various other mixed-valence species discussed in this paper. Whereas the spin density was localized on the Fe\textsubscript{p} center for the other purely mixed-valence species, such as [(H-(Fe\textsuperscript{I}Fe\textsuperscript{II})adt\textsuperscript{0}] and [(H-(Fe\textsuperscript{I}Fe\textsuperscript{II})adt-H]\textsuperscript{+}], this H\textsubscript{ox}-like species has the spin density entirely localized on the Fe\textsubscript{d} center, suggesting that the Fe\textsubscript{d} center is now more reduced than the Fe\textsubscript{p} center. Upon formation of this [(Fe\textsuperscript{II}Fe\textsuperscript{I})adt\textsuperscript{+}] species, the bridging CO ligand returns to the original asymmetric semibridging position. This analysis of the mechanism points to the importance of the bridging CO ligand, which seemingly behaves as a lever and indicator for intramolecular electron transfer between the Fe centers. Moreover, there is 0.16 spin density localized on the symmetric bridging CO ligand in the [(H\textsubscript{2}-(Fe\textsuperscript{1.5}Fe\textsuperscript{1.5})adt\textsuperscript{+}] species, suggesting that it may be even more involved in the electron transfer process.

2.11. Concluding Remarks

In this paper, we investigated a proton reduction catalyst, [(FeFe)adt], which is modeled after the [FeFe]-hydrogenase enzyme. This model is characterized by incorporation of the azadithiolate bridge, which facilitates the formation of a doubly protonated species through a proton relay. The calculated results agree well with the experimental data for the geometries, CO vibrational stretching frequencies, and reduction potentials for the systems studied. This
agreement provides validation for the computational methods. The calculations illustrate that the 
NH···HFe \textsubscript{d} dihydrogen bonding distance in the doubly protonated species is highly sensitive to 
the effects of ion-pairing between the ammonium and BF\textsubscript{4}⁻ counterions, which are present in the 
crystal structure. Specifically, the inclusion of BF\textsubscript{4}⁻ counterions leads to a significantly longer 
dihydrogen bond. Comparable weak interactions with the protein backbone might be expected to 
influence the H···H bonding in the enzyme. Analyses of spin densities provide insight into the 
oxidation states of the metal centers for the reduced singly and doubly protonated species. The 
non-hydride Fe center (Fe\textsubscript{p}) was found to be the site of reduction for terminal hydride species and 
for unsymmetric bridging hydride species. In contrast, the reduced symmetric bridging hydride 
species exhibited a significant degree of spin delocalization between the Fe centers, although these 
species exhibited more asymmetry than the analogous [(FeFe)pdt] species. These differences in 
the degree of spin localization, as well as the movement of the hydride upon reduction of the 
unsymmetric bridging species, may be related to the experimental observation that the terminal 
hydride species are reduced at less negative potentials than their bridging hydride isomers.\textsuperscript{35,41,79} 
This analysis suggests that the bridging hydrides may be considered to be quasi-terminal hydrides 
for the unsymmetric species but not for the symmetric species.

A combined experimental and theoretical study of the relative pK\textsubscript{a} values for the various 
species in the catalytic cycle provides further insights. The Fe\textsubscript{d} site was found to be more basic 
than the amine, although the amine is more readily protonated due to steric crowding at the Fe\textsubscript{d} 
site. Moreover, the bridging hydride species was found to be more thermodynamically stable than 
the terminal hydride species. These observations are consistent with the initial protonation of the 
amine, followed by intramolecular proton transfer to the Fe\textsubscript{d} site to produce a terminal hydride 
species, followed by isomerization to the bridging hydride species at room temperature but not at
low temperatures. In terms of forming the doubly protonated species, the amine site is significantly less basic for the terminal hydride species than for the unprotonated species. Both the experimental and theoretical data indicate that a stronger acid is required for this second protonation.

In addition, these calculations implicate a possible pathway for H₂ evolution that involves an intermediate with H₂ weakly bonded to one Fe, a short H₂ distance similar to the molecular bond length, the spin density delocalized over the two Fe centers, and a nearly symmetrically bridged CO ligand. The short H₂ distance implies that the Fe-H₂ bonding is relatively weak and that the product H₂ is poised for release. Upon H₂ release, the mixed-valence species produced is closely related to the Hₐox state of the enzyme. The key features of this proposed mechanism highlight the roles of the ammonium-hydride interaction, flexibility of the bridging CO ligand, and intramolecular electron transfer between the Fe centers in the catalytic cycle. While these bioinspired models are effective catalysts for H₂ production, the challenge of designing molecular electrocatalysts that can perform with weaker acids and lower overpotentials remains.¹ Further investigation of the effects of ligand substitution at the Fe centers will help guide the design of more effective H₂-producing catalysts.

2.12 References


Cao, Z.; Hall, M. B. Modeling the active sites in metalloenzymes. 3. Density functional calculations on models for [Fe]-hydrogenase: Structures and vibrational frequencies of the observed redox forms and the reaction mechanism at the diiron active center. J. Am. Chem. Soc. 2001, 123, 3734-3742.


(35) Barton, B. E.; Rauchfuss, T. B. Terminal hydride in [FeFe]-hydrogenase model has lower potential for H₂ production than the isomeric bridging hydride. *Inorg. Chem.* **2008**, *47*, 2261-2263.


Chapter 3. Interplay between Terminal and Bridging Diiron Hydrides\textsuperscript{*}*\textsuperscript{†}

3.1. Abstract

This study describes structural and redox properties of electronically unsymmetrical diiron hydrides. The terminal hydride \( \text{Cp}^*\text{Fe(pdt)Fe(dppe)(CO)H} \) (\([1(t-H)]^0\), \( \text{Cp}^* = \text{Me}_5\text{C}_{5}^- \), \( \text{pdt}^2^- = \text{CH}_2(\text{CH}_2\text{S}^-)_{2} \), \( \text{dppe} = \text{Ph}_2\text{PC}_{2}\text{H}_{4}\text{PPh}_2 \)) was prepared by hydride reduction of \([\text{Cp}^*\text{Fe(pdt)Fe(dppe)(CO)(NCMe)}]^+\). As established by X-ray crystallography, \([1(t-H)]^0\) features a terminal hydride ligand. Unlike previous examples of terminal diiron hydrides, \([1(t-H)]^0\) does not isomerize to the bridging hydride \([1(\mu-H)]^0\). Oxidation of \([1(t-H)]^0\) gives \([1(t-H)]^+\), which was also characterized crystallographically as its \(\text{BF}_4^-\) salt. Density functional theory (DFT) calculations indicate that \([1(t-H)]^+\) is best described as containing an \(\text{Cp}^*\text{Fe(III)}\) center. In solution, \([1(t-H)]^+\) isomerizes to \([1(\mu-H)]^+\) as anticipated by DFT. Reduction of \([1(\mu-H)]^+\) by \(\text{Cp}_2\text{Co}\) affords the diferrous bridging hydride \([1(\mu-H)]^0\). Electrochemical measurements and DFT calculations indicate that the couples \([1(t-H)]^{+/0}\) and \([1(\mu-H)]^{+/0}\) differ by 210 mV. Qualitative measurements indicate that \([1(t-H)]^0\) and \([1(\mu-H)]^0\) are close in free energy. Protonation of \([1(t-H)]^0\) in MeCN solution affords \(\text{H}_2\) even with weak acids via hydride transfer. In contrast, protonation of \([1(\mu-H)]^0\) yields 0.5 equiv of \(\text{H}_2\) by a proposed protonation-induced electron transfer process. Isotopic labeling indicates that \(\mu\)-H/D ligands are inert.


\textsuperscript{†} Author contributions: Xin Yu, Chen-Ho Tung, Dr. Danielle L. Gray, Prof. Wenguang Wang, and Prof. Thomas Rauchfuss were responsible for all experiments and associated analyses in this chapter.
3.2. Introduction

The interaction of the most abundant metal, iron, with the most abundant element, hydrogen, is of obvious fundamental importance. Indeed, many iron hydride (Fe–H) compounds are known. Iron hydrides play important roles in biology, in industry, and even in geosciences.¹ Not only has the Fe–H bond been invoked in all three hydrogenases²³ but also in nitrogenase.⁴⁵ Molecular iron hydrides are the basis of emerging "earth-abundant" catalytic technologies.⁶⁷ Iron hydrides are almost certainly involved in the Haber-Bosch process, one of mankind's most energy intensive industrial operations.⁸

Intense studies have focused on the design and characterization of molecular iron complexes with hydride ligands.⁹ Ferrous compounds of the type H₂FeL₄ and HFeXL₄ and related cyclopentadienyl derivatives are particularly well-studied.⁹¹⁰ With respect to redox, the system [Cp*Fe(dppe)H]ⁿ⁺ (dppe = 1,2-C₂H₄(PPh₂)₂) is renowned because it has been characterized in both its ferrous (n = 0) and ferric (n = 1) states.¹¹ Such work has inspired the development of CpFe(diphosphine-amine)-based catalysts for the oxidation of H₂.¹²⁻¹⁵

In the context of biological and heterogeneous catalysis, multi-iron hydrides are of special relevance. Of general interest are the factors that influence the regiochemistry of hydride ligands, terminal (attached to a single Fe center) vs bridging (attached to two or more Fe centers). With the exception of a few metastable models for the [FeFe]-hydrogenases,¹⁶⁻²¹ hydride ligands in multi-iron complexes occupy bridging sites.²²⁻²⁴ This pattern applies to [Fe(0)]₂H systems (e.g., [HFe₂(CO)₈]⁻),²⁵ [Fe(II)H]ⁿ systems,²³,²⁶,²⁷ mixed-valence Fe(III)Fe(II) systems (e.g., [Fe₂(NH)(H)(PR₃)₆]²⁺),²₈ and [Fe(III)]₂Hₓ systems (e.g., [Cp*FeH₂]₂).²⁶,²⁹ A possible exception is the minor isomer of [HFe₂(S₂C₆H₄)(CO)₆]⁻, but this assignment is tentative.³⁰ In this paper, we report results demonstrating the isolability of a thermodynamically stable diiron complex with a
terminal hydride ligand. This advance is disclosed in the context of new transformations involving all members of the series \([\text{Fe–Fe–H}]^z\) and \([\text{Fe–H–Fe}]^z\) \((z = 0, 1; \text{Scheme 3.1})\).

![Scheme 3.1](image)

**Scheme 3.1.** Diiron hydrides described in this work.

### 3.3. Fe(II)–Fe(II)–H ([1(κ-H)]\(^0\))

The key diiron hydride was prepared in three steps as described in Scheme 3.2. The electrophilic reagent \([\text{Cp}*\text{Fe(NCMe)}_3]\text{PF}_6\) condenses rapidly at room temperature with \(\text{Fe(pdt)(CO)}_2(\text{dppe})\)\(^{31}\) to give the cation \([\text{Cp}*\text{Fe(pdt)Fe(dppe)(CO)}_2]^+\) \(([1(\text{CO})]^+)\). The salt \([1(\text{CO})]\text{PF}_6\) was obtained in high yield (83%) as red-brown crystals. The IR spectrum of \([1(\text{CO})]\text{PF}_6\) exhibits \(v_{\text{CO}}\) bands at 1950 and 1856 cm\(^{-1}\), assigned to terminal and bridging CO ligands, respectively.
Scheme 3.2. The synthesis of the thermally stable terminal hydride [I(t-H)]\(^0\). Reaction conditions: i) 1:2 MeCN/CH\(_2\)Cl\(_2\), –30 °C → RT; ii) MeCN, 365 nm; iii) MeCN, Bu\(_4\)NBH\(_4\).

To activate the diiron center for installation of the hydride ligand, [I(CO)]\(^+\) was converted to [I(MeCN)]\(^+\). This conversion was achieved by photosubstitution in MeCN solution. The lability of the MeCN ligand in [I(MeCN)]\(^+\) is indicated by its quantitative carbonylation to regenerate [I(CO)]\(^+\). The \(^{31}\)P NMR spectrum of [I(MeCN)]\(^+\) displays a singlet at \(\delta\) 65, indicating that the \(C_s\) symmetry of the precursor is retained in the product. Crystallographic analysis confirms that the overall structure of the Cp*Fe(pdt)Fe(dppe)L\(_2\) core is retained (Figure S1). Upon replacement of the terminal CO by MeCN, the bridging CO shifts from an unsymmetrical (\(\Delta(\text{Fe–}\mu-\text{CO}) = 0.5 \text{ Å}\)) to a symmetrical (\(\Delta(\text{Fe–}\mu-\text{CO}) = 0.05 \text{ Å}\)) position. This change is reflected in an 89 cm\(^{-1}\) shift for \(\nu_{\text{CO}}\) to 1769 cm\(^{-1}\). The repositioning of the \(\mu\)-CO ligand indicates that the Fe sites in the [Fe(dppe)(NCMe)]\(^2+\) and [Cp*Fe]\(^+\) fragments are more equivalent electronically.

The salt [I(MeCN)]PF\(_6\) reacts with Bu\(_4\)NBH\(_4\) to give the targeted hydride [I(t-H)]\(^0\). Thus, upon addition of Bu\(_4\)NBH\(_4\) to an MeCN solution of [I(MeCN)]\(^+\), a color change from dark brown to black was observed, and microcrystals of the neutral product precipitate. The charge-neutral
hydride exhibits good solubility in toluene and was purified by crystallization from cold pentane. In its IR spectrum, the \( \nu_{\text{CO}} \) band of \([1(t-H)]^0\) shifted to 1752 cm\(^{-1}\), in contrast to 1769 cm\(^{-1}\) for \([1(\text{MeCN})]^+\). Its \(^{1}\text{H}\) NMR spectrum features a triplet at \( \delta -4.95 \) (t, \( J_{P,H} = 76 \) Hz), a chemical shift that is diagnostic of terminal iron hydrides.

**Figure 3.1.** The structure of \([1(t-H)]^0\) with H atoms (except for the hydride ligand) omitted, and thermal ellipsoids shown at the 50% probability level.

In contrast to other diiron complexes with terminal hydride ligands, \([1(t-H)]^0\) is robust. Heating to the point of its decomposition (hours at 50 °C in THF solution) gave none of the isomeric bridging hydride \([1(\mu-H)]^0\), which we have prepared as described below. The molecular structure of \([1(t-H)]^0\) was confirmed crystallographically (Figure 3.1). It has idealized \(C_s\) symmetry. As in \([1(\text{NCMe})]^+\), the bridging CO is symmetrically bound (\(\Delta(\text{Fe–}$$\mu$$-CO) = 0.02 Å).

### 3.4. Fe(III)–Fe(II)–H ([1(t-H)]\(^{+}\))

The redox properties of \([1(t-H)]^0\) were evaluated by cyclic voltammetry (CV, Figure 3.2). It is a fairly strong reducing agent, exhibiting a reversible oxidation at \(-0.87 \) V vs Fe\(^{+}/0\). A second, but irreversible, oxidation was observed at 0.38 V (Figure S16). These oxidations are assigned to
the \([1(t-H)]^0/+\) and \([1(t-H)]^{+2/+}\) couples, respectively. The irreversibility of the \([1(t-H)]^{+2/+}\) couple is expected owing to the strongly enhanced acidity upon oxidation.\(^3\)

**Figure 3.2.** Cyclic voltammogram for \([1(t-H)]^0\). _Conditions:_ 1 mM sample in THF, 0.1 M Bu4NPF6; scan rate, 100 mV/s; potentials vs Fe(+/0). Analysis of this CV indicates that \(E_{1/2}\text{[1}(t-H)\text{]}^+/0 = -0.87 \text{ V}\), \(i_{pa}/i_{pc} = 0.97\); \(E_{pc}\text{[1}(t-H)\text{]}^{2+/+} = 0.38 \text{ V}\), irreversible.

The oxidation of \([1(t-H)]^0\) was conducted on a preparative scale with FcBF4 in CH2Cl2. At \(-40^\circ\text{C}\), the solution color changed from black to olive-green immediately upon addition of the oxidant. The IR spectrum of \([1(t-H)]^+\) exhibits a broad \(\nu_{\text{CO}}\) band at 1875 cm\(^{-1}\), a \(\sim 130 \text{ cm}^{-1}\) shift to higher energy relative to \([1(t-H)]^0\). This redox process is chemically reversible, as addition of CpCo2 to the reaction mixture gave back \([1(t-H)]^0\). Notably, \([1(t-H)]^+\) is unreactive toward base.

X-ray crystallographic refinement of the structure of \([1(t-H)]\text{BF}_4\) was sufficient to confirm the overall connectivity. The diiron centers in \([1(t-H)]^+\) and \([1(t-H)]^0\) are similar, each with a \(\eta^1\) - CO ligand (Figure S20). The 1 e\(^-\) oxidation also causes the Cp*-Fe(1) distance to increase from 1.715 in \([1(t-H)]^0\) to 1.742 Å in \([1(t-H)]^+\). A similar change was also observed for the oxidation of Cp*-Fe(dcpe)H (dcpe = 1,2-C2H4(P(C6H11)2)2).\(^33,34\) The EPR spectrum of \([1(t-H)]^+\) is rhombic, but approaching axial in symmetry (Figure 3.3). The spectrum was simulated as an 82:18 mixture of two rhombic subspectra with \(g = 1.957, 2.074, 2.412\) and \(g = 1.957, 2.062, 2.413\). These subspectra result from two "flippamers", isomers that differ with respect to the conformation of
the propanedithiolate ligand. Similar effects have been observed previously in dimetallic complexes of pdt$^{2-}$, e.g., in the case of [Fe$_2$(pdt)(CO)$_3$(PMe$_3$)(dpv)]BF$_4$ and [NiFe(pdt)(CO)$_3$(dppe)]BF$_4$. The broad ν$_{CO}$ band for [I(t-H)]$^+$ is also consistent with conformers. The calculated free energy difference between these two flippamers is 2.16 kcal/mol in THF, with the flippamer in which the pdt is oriented toward the Fe(dppe)CO center being thermodynamically more stable, which is in agreement with the crystal structure. This calculated free energy difference implicates a 97:3 ratio, which is also in qualitative agreement with the ratio from the EPR simulation. Additionally, the calculated free energy barrier for the interconversion between the two flippamers is 10.54 kcal/mol in THF, which is consistent with previous studies (Figure S39).

**Figure 3.3.** X-band EPR spectra (CH$_2$Cl$_2$/PhMe, 110 K) of [I(t-H)]BF$_4$. The experimental spectrum (Exp) could be simulated as the sum of two components, denoted A+B, where the A and B subspectra represent the two flippamers differing in the conformation of the pdt ligand.

**3.5. DFT Analysis of [I(t-H)]$^0$ and [I(t-H)]$^+$**

The DFT-optimized structures of [I(t-H)]$^0$ and [I(t-H)]$^+$ are in good agreement with the crystallographic results (Table 3.1 and Tables S2-S3). Since metal-hydride distances are often
underestimated by X-ray crystallography, discussion of these bonds focuses on the distances calculated by DFT. Consistent with an unchanged Fe–H distance is the finding that the oxidation is localized on the Cp*Fe center, as revealed by the highly localized nature of the spin density (Figure 3.4). The greatest differences between the structures of [1(t-H)]\(^0\) and [1(t-H)]\(^+\) are associated with the Fe–\(\mu\)-CO distances, which are also the most variable structural elements of the [FeFe]-hydrogenases.\(^2,42\) Upon oxidation, the CO ligand shifts from a nearly symmetric position between the two Fe centers to a position closer to the Fe(II) hydride center and farther from the Cp*Fe(III) center. Such a shift is also consistent with the anticipated tendencies of neutral Fe(II) vs cationic Fe(III) centers to engage in \(\pi\)-backbonding. The \(\mu\)-CO ligand "points" to the electron-rich Fe center.

**Table 3.1.** Crystallographic and DFT-Calculated Bond Distances (Å) for [1(t-H)]\(^0\) and [1(t-H)]\(^+\)

<table>
<thead>
<tr>
<th></th>
<th>[1(t-H)](^0)</th>
<th>[1(t-H)](<del>BF_4</del>)(^a)</th>
<th>[1(t-H)](<del>BF_4</del>)(^a)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Exptl</td>
<td>Calcd</td>
<td>Exptl</td>
</tr>
<tr>
<td>Fe(1)–Fe(2)</td>
<td>2.4615(3)</td>
<td>2.478</td>
<td>2.513(1)</td>
</tr>
<tr>
<td>Fe(2)–H(1)</td>
<td>1.46(2)</td>
<td>1.522</td>
<td>n.d.(^c)</td>
</tr>
<tr>
<td>Fe(1)–C(11)</td>
<td>1.948(2)</td>
<td>1.923</td>
<td>2.359(6)</td>
</tr>
<tr>
<td>Fe(2)–C(11)</td>
<td>1.927(2)</td>
<td>1.933</td>
<td>1.796(7)</td>
</tr>
<tr>
<td>Fe(1)–Cp*</td>
<td>1.715</td>
<td>1.738</td>
<td>1.742</td>
</tr>
<tr>
<td>Fe(1)–S</td>
<td>2.2453(5)</td>
<td>2.273</td>
<td>2.223(2)</td>
</tr>
<tr>
<td></td>
<td>2.2447(5)</td>
<td>2.269</td>
<td>2.227(2)</td>
</tr>
<tr>
<td>Fe(2)–S</td>
<td>2.2560(5)</td>
<td>2.282</td>
<td>2.264(2)</td>
</tr>
<tr>
<td></td>
<td>2.2716(5)</td>
<td>2.298</td>
<td>2.283(2)</td>
</tr>
<tr>
<td>Fe(2)–P</td>
<td>2.1758(5)</td>
<td>2.203</td>
<td>2.202(2)</td>
</tr>
<tr>
<td></td>
<td>2.1807(5)</td>
<td>2.203</td>
<td>2.208(2)</td>
</tr>
</tbody>
</table>

\(^a\)BF\(_4\) was not included in the DFT calculations. \(^c\)The distances are reported for the flippamer in which the pdt is oriented toward the Fe(dppe)CO center. The bond distances for the flippamer with the pdt oriented toward the Cp*Fe center are given in Table S13. \(^c\)Not observed.
Figure 3.4. Spin density plot for $[1(t-H)]^+$ on the left and $[1(\mu-H)]^+$ on the right showing predominant localization of unpaired electron on the Cp*Fe(1) center. Spin density is shown in blue wireframe. Numerical values for $[1(t-H)]^+$: $\rho$(Fe1) = 1.17, $\rho$(Fe2) = −0.05. Numerical values for $[1(\mu-H)]^+$: $\rho$(Fe1) = 1.10, $\rho$(Fe2) = −0.03. The hydrogen atoms, except the $t$-H and $\mu$-H, have been removed for clarity.

Other DFT results are in agreement with experimental measurements. In particular, the calculated and observed $\nu_{CO}$ bands agree well for both the neutral and cationic complexes (Table 3.2 and Table S7). The experimentally measured and calculated redox potentials are also in good agreement (Table S6).

Table 3.2. Crystallographic and DFT-Calculated Bond Distances (Å) for $[1(t-H)]^0$ and $[1(t-H)]^+$

<table>
<thead>
<tr>
<th></th>
<th>$\nu_{CO}$ (cm$^{-1}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Exptl</td>
</tr>
<tr>
<td>$[1(t-H)]^0$</td>
<td>1752</td>
</tr>
<tr>
<td>$[1(t-H)]BF_4$</td>
<td>1875</td>
</tr>
<tr>
<td>$[1(\mu-H)]^0$</td>
<td>1914</td>
</tr>
<tr>
<td>$[1(\mu-H)]BF_4$</td>
<td>1953</td>
</tr>
</tbody>
</table>

3.6. Fe(III)–H–Fe(II) ([1(μ-H)]$^+$)

In contrast to the inertness of its neutral parent, the cation $[1(t-H)]^+$ is labile. After a few hours at room temperature it isomerizes to $[1(\mu-H)]^+$ (Figure 3.5). This conversion, which is quantitative, was monitored by IR spectroscopy. The appearance of a $\nu_{CO}$ band at 1953 cm$^{-1}$ indicates a terminal CO ligand, which is also consistent with the site exchange between CO and
H\(^-\) in this mixed-valence cation. The isomerization follows unimolecular kinetics with a rate constant of \(3.78 \times 10^{-5}\) s\(^{-1}\) at 24 \(^\circ\)C (Figure 3.4). Interestingly, isomerism of [I(t-H)]\(^+\) was strongly affected by the counterion, while remaining first order. For [I(t-H)]BAr\(^5\)\(_4\), \(t_{1/2}\) increased to 98.7 h from 5.1 h (Figure S22).

**Figure 3.5.** IR spectra of [I(t-H)]BF\(_4\) (1\(\times\)10\(^{-3}\) M in CH\(_2\)Cl\(_2\)) (left) and the integrated absorbance of the peak at 1953 cm\(^{-1}\) vs time (right).

In solution, [I(\(\mu\)-H)]\(^+\), which is NMR-silent, exhibits a rhombic EPR spectrum (Figure S38). With values of \(g = 1.9827, 2.0732,\) and 2.4419, the spectrum is approaching axial symmetry consistent with a nearly \(C_s\)-symmetric environment. The absence of hyperfine coupling suggests that the unpaired electron is localized on the Cp*Fe site, consistent with the spin density calculated with DFT (Figure 3.4). Related to [I(\(\mu\)-H)]\(^+\) is the symmetrical mixed-valence (Fe(II)Fe(III)) bridging hydride \([\text{Cp}\^*\_2\text{Fe}_2(\text{S}_2\text{C}_6\text{H}_4)(\text{\(\mu\)-H})_0.43]\).

X-ray crystallographic analysis of [I(\(\mu\)-H)]BF\(_4\) confirmed that this cation retains the structure of its precursor (Figure 3.6 and Figure S24). The two cations have crystallographically imposed \(C_s\) symmetry, but they differ with respect to the conformation of the pdt\(^2\)\(^-\) ligand. Some disorder is observed in the dppe groups. The structural details are discussed below in the context of the DFT calculations.
According to the usual oxidation state and electron counting rules, $\text{[1(\mu-H)]}^+$ is a 33 e$^-$ species, described as Fe(III)--Fe(II) or Fe(II)--Fe(III). The magnetic moment for $\text{[1(\mu-H)]BF}_4$ in solution was determined to be 2.53 $\mu_B$ by the Evans method. This value is comparable to that reported for $[\text{Cp}^*\text{Fe(dppe)}X]^+$ salts (2.16–2.79 $\mu_B$). Notably, isomerization of $\text{[1(t-H)]}^+$ to $\text{[1(\mu-H)]}^+$ leads to shortening of the $\text{Cp}^*$--Fe distances, from 1.742 Å to 1.715 and 1.727 Å, depending on the flippamer.

3.7. Fe(II)--H--Fe(II) ($\text{[1(\mu-H)]}^0$)

The final member of the [1H]$^+$ series is $\text{[1(\mu-H)]}^0$, diferrous with a bridging hydride. The accessibility of this species is indicated by the CV of $\text{[1(\mu-H)]}^+$, which exhibits a reversible reduction at $-1.06$ V vs Fc$^{+/0}$ (Figure 3.7). Relative to $\text{[1(t-H)]}^0$, the isomeric bridging hydride $\text{[1(\mu-H)]}^0$ is about 200 mV more reducing. A similar trend was observed for the isomeric hydrides of $[\text{HFe}_2(\text{xdt})(\text{CO})_2(\text{dppv})_2]^+/0$ (xdt$^{2-}$ = azadithiolate$^{2-}$ or pdt$^{2-}$). On a preparative scale, treatment of a $\text{CH}_2\text{Cl}_2$ solution of $\text{[1(\mu-H)]BF}_4$ with one equivalent of $\text{Cp}_2\text{Co}$ ($E_{1/2} = -1.33$ V vs Fc$^{+/0}$) gave a bright green product. The FT-IR spectrum features a $v_{\text{CO}}$ band at 1908 cm$^{-1}$, which is $\sim 45$ cm$^{-1}$ lower in energy vs $\text{[1(\mu-H)]}^+$. The diamagnetism of $\text{[1(\mu-H)]}^0$ is indicated by its well-behaved NMR properties. The $^{31}\text{P}$ NMR spectrum features a singlet at $\delta 80$, and the $^1\text{H}$ NMR
spectrum displays a hydride signal as a triplet at $\delta -12.66$ with $J_{P-H} = 28$ Hz. Efforts to obtain high-quality single crystals suitable for X-ray crystallographic analysis were unsuccessful. Fortunately, this $S = 0$, diferrous, $\mu$-hydride species is the most routine and predictable of the series of four compounds.

![Cyclic voltammogram for [1(μ-H)]⁺](image)

**Figure 3.7.** Cyclic voltammogram for [1(μ-H)]⁺. Conditions: 1 mM sample in CH₂Cl₂, 0.1 M Bu₄NPF₆; scan rate, 100 mV/s; potentials vs Fe⁷⁺⁰. Analysis of this CV indicates that $E_{1/2} = -1.06$ V, $i_{pa}/i_{pc} = 0.99$; $E_{1/2} = 0.36$ V, $i_{pa}/i_{pc} = 1.00$.

### 3.8. DFT Analysis of [1(μ-H)]⁰ and [1(μ-H)]⁺

In good agreement with the crystal structure, the DFT-optimized structures of the two flippamers of [1(μ-H)]⁺ reveal $C_s$-symmetric species (Table 3.3). The two computed Fe–H bond lengths are very similar for the neutral and cationic states. As is typical, Fe–(μ-H) distances are longer than terminal Fe–H bonds. According to the DFT calculations, the Fe–H distances in the two flippamers change by < 0.02 Å. It is difficult to determine Fe–H distances reliably using X-ray crystallography because of the large electron density of the Fe atoms dominating the diffraction pattern. Disorder in the dppe ligands limits the signal to noise in the high resolution data preventing accurate refinement of the Fe–H distances because of Fourier truncation ripples located close to the Fe atom positions.
Table 3.3. Crystallographic and DFT-Calculated Bond Distances (Å) for [1(µ-H)]0 and [1(µ-H)]+

<table>
<thead>
<tr>
<th></th>
<th>[1(µ-H)]0</th>
<th>[1(µ-H)]BF4</th>
<th>Cacld</th>
<th>Exptl</th>
<th>Cacld a</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>(pdt CO)</td>
<td>(pdt CO)</td>
<td>(pdt CO)</td>
<td>(pdt CO)</td>
<td>(pdt CO)</td>
</tr>
<tr>
<td>Fe–Fe</td>
<td>2.600</td>
<td>2.5868(9)</td>
<td>2.572(1)</td>
<td>2.592</td>
<td>2.607</td>
</tr>
<tr>
<td>Cp*Fe–H</td>
<td>1.650</td>
<td>1.64(7)</td>
<td>1.56(5)</td>
<td>1.661</td>
<td>1.684</td>
</tr>
<tr>
<td>(dppe)Fe–H</td>
<td>1.719</td>
<td>1.43(6)</td>
<td>1.78(5)</td>
<td>1.710</td>
<td>1.697</td>
</tr>
<tr>
<td>Fe–CO</td>
<td>1.742</td>
<td>1.754(4)</td>
<td>1.741(5)</td>
<td>1.748</td>
<td>1.746</td>
</tr>
<tr>
<td>Fe–Cp*</td>
<td>1.680</td>
<td>1.727</td>
<td>1.715</td>
<td>1.748</td>
<td>1.742</td>
</tr>
<tr>
<td>Cp*Fe–S</td>
<td>2.233</td>
<td>2.2101(9)</td>
<td>2.223(1)</td>
<td>2.243</td>
<td>2.255</td>
</tr>
<tr>
<td>(dppe)Fe–S</td>
<td>2.315</td>
<td>2.2657(9)</td>
<td>2.264(1)</td>
<td>2.303</td>
<td>2.292</td>
</tr>
<tr>
<td>Fe–P</td>
<td>2.245</td>
<td>2.205(5)</td>
<td>2.235(4)</td>
<td>2.271</td>
<td>2.272,</td>
</tr>
<tr>
<td></td>
<td></td>
<td>2.221(5)</td>
<td>2.217(5)</td>
<td>2.275</td>
<td>2.274</td>
</tr>
</tbody>
</table>

aBF4 was not included in the DFT calculations. bThese bond distances correspond to the flippamer in which the central CH2 of the pdt− is oriented toward the Fe(dppe)CO center. These bond distances correspond to the flippamer in which the central CH2 of the pdt− is oriented toward the Cp*Fe center.

The calculated free energy difference is 1.08 kcal/mol in THF between these two flippamers, with the flippamer with the pdt− oriented toward the Fe(dppe)dpv center being thermodynamically favored. The calculated free energy barrier for the interconversion between the two flippamers is ΔG‡ = 9.37 kcal/mol in THF, consistent with previous calculations.36-40

The greatest change induced by oxidation of [1(µ-H)]0 is the ~0.06 Å elongation of the Fe–Cp* distance, as seen by DFT. Similar differences are seen for [Cp*Fe(dppe)H]+/0.48,49 The elongation of the Fe–Cp* distances upon oxidation is attributed to diminution of π-backbonding from Fe(III) to Cp*. This difference is not observed in [1(t-H)]0/+ because of the buffering effect of the strongly π-acidic µ-CO ligand. These results are consistent with the EPR data and calculations indicating that the spin density is localized on the Cp*Fe center (Figure 3.4).

3.9. Isomerization of Neutral and Cationic Hydrides

The isolation of [1(µ-H)]0 provided an opportunity to evaluate the terminal-bridging isomerization of hydride ligands. Although [1(t-H)]0 shows no tendency to convert to [1(µ-H)]0, equilibration may be subject to high kinetic barriers. In toluene solution, [1(µ-H)]0 slowly converts to [1(t-H)]0 but only partially. About ~10% conversion is observed after 36 h, competitive with decomposition. These results are consistent with DFT calculations suggesting that the reaction free energy for the isomerization of [1(µ-H)]0 to [1(t-H)]0 is endergonic by ~4 kcal/mol. Furthermore, according to DFT, isomerization of [1(t-H)]+ to [1(µ-H)]+ is exergonic by ca. −7
kcal/mol in THF, consistent with our observation of a full conversion in solution. As shown in Tables S8-S9, however, the isomerization reaction free energies vary for different functionals and should be viewed as only qualitative estimates of relative thermodynamic stabilities. For example, all functionals studied are consistent in that the bridging hydride is more stabilized relative to the terminal hydride for the cationic species than for the neutral species. Because of this limitation of DFT for the isomerization process, the free energy barriers that would be required for a complete kinetic analysis were not calculated. However, previous computational studies on related models predict that this isomerization likely proceeds through a series of Ray-Dutt (rhombic) and Bailar (trigonal) twists.50

3.10. Protonation of Hydrides

In addition to isomerization and redox studies, the reactivity of diferrous hydrides toward protons was examined. Solutions of [1(t-H)]0 in MeCN react with protic acids such as H(OEt)2BARF4 producing [1(NCMe)]+ (eq 1).

\[
\text{Cp*Fe(pdt)Fe(dppe)(CO)H} + H^+ + \text{MeCN} \rightarrow [\text{Cp*Fe(pdt)Fe(dppe)(CO)(NCMe)}]^+ + H_2 \quad (1)
\]

The yield of H2 is close to stoichiometric (quantified by GC analysis, 89±6% yields in three experiments). The hydridic character of [1(t-H)]0 is indicated by the finding that H2 evolution is induced with even the weakly acidic pyrrolidinium tetrafluoroborate ([HPyrr]BF4, pK_a MeCN = 19.651). Solutions of [1(t-H)]+ are unreactive toward acids, indicating its diminished hydricity.

Similar to [1(t-H)]0, the isomeric [1(†-H)]0 also reacts with strong acids, but by a different pathway. The yield of H2 was 43±5% using H(OEt)2BARF4 and [HPPh3]BF4 (pK_a MeCN = 7.6). The diiron product is exclusively [1(µ-H)]+, as indicated by IR spectral analysis. The ~50% yield of H2 indicates that the µ-H ligand serves as a spectator, not as a source of H2. This unusual behavior, which has been reported for [(µ-H)Fe2(pdt)(CO)2(dppv)]0 and [Cp*2Fe2(S2C6H4)(µ-H)]0,43 is consistent with the reactions in Scheme 3.3.
Protonation gives a diferric dihydride:

$$[\text{I}(\mu-\text{H})]^0 + \text{H}^+ \rightarrow [\text{HI}(\mu-\text{H})]^+ \quad (2)$$

Electron transfer from the parent hydride to its conjugate acid gives a mixed valence dihydride:

$$[\text{HI}(\mu-\text{H})]^+ + [\text{I}(\mu-\text{H})]^0 \rightarrow [\text{HI}(\mu-\text{H})]^0 + [\text{I}(\mu-\text{H})]^+ \quad (3)$$

Protonation of the dihydride results in hydrogen evolution:

$$[\text{HI}(\mu-\text{H})]^0 + \text{H}^+ \rightarrow [\text{I}(\mu-\text{H})]^+ + \text{H}_2 \quad (4)$$

Net reaction:

$$2 [\text{I}(\mu-\text{H})]^0 + 2 \text{H}^+ \rightarrow 2 [\text{I}(\mu-\text{H})]^+ + \text{H}_2 \quad (5)$$

Scheme 3.3. H$_2$ Evolution Mechanism

This pathway is analogous to protonation-induced electron-transfer (ET) reactions implicated in the proposed H$_2$ evolution mechanisms for two other cases of electron-rich ($E_{\text{MHM}}^{\text{+/0}} \sim -500$ mV vs Fe$^{+/0}$) bridging hydride complexes.$^{52,53}$ The role of the bridging hydride as a spectator in $[\text{I}(\mu-\text{H})]^0$ was confirmed by conducting the same experiments starting with $[\text{I}(\mu-\text{D})]^0$.

Protolysis of this deuteride with H(OEt)$_2$BAr$_F$$_4$ gave exclusively $[\text{I}(\mu-\text{D})]^+$, as confirmed by IR and $^2$H NMR spectroscopies.

3.11. Conclusions

This paper describes the characterization of all four members of a square scheme for diiron hydrides (Scheme 3.1). The total free energy change for the cycle should be 0 kcal/mol, see eq 6.

$$\Delta G\{[\text{I}(t-\text{H})]^{+/0} \} + \Delta G\{[\text{I}(\mu-\text{H})]^+ / [\text{I}(t-\text{H})]^+\} - \Delta G\{[\text{I}(\mu-\text{H})]^{+/0} \} - \Delta G\{[\text{I}(t-\text{H})]^0 / [\text{I}(\mu-\text{H})]^0\}\} = 0 \quad (6)$$

The uniquely well-behaved nature of this system is established by the reversibility of the redox reactions and the crystallographic analysis of several members of the cycle. The ideality of this system is attributable to a number of factors: the steric protection afforded by the dppe and Cp* ligands, the high barriers to isomerization because the CO and hydride ligands are mutually trans, and the nonequivalent environments of the two Fe centers.
While much work has been reported on monometallic hydrides, the following conclusions and hypotheses emerge from this work pertaining to bimetallic hydrides:

1) **Relative to the more prevalent bridging hydrides, terminal hydrides of bimetallic complexes are more stabilized in reduced states.** This effect is observed in the experimentally measured equilibrium ratios for $[1(t-\text{H})]^+/[1(\mu-\text{H})]^+ (K_{298} \ll 1)$ and $[1(t-\text{H})]^0/[1(\mu-\text{H})]^0 (K_{298} > 0.15)$ at room temperature, and affirmed by DFT, which indicates that the bridging hydride is more stabilized relative to the terminal hydride for the cationic species than for the neutral species.

2) **Terminal hydrides exhibit greater hydridic character.** This effect is illustrated by the contrasting behavior of $[1(t-\text{H})]^0$ vs $[1(\mu-\text{H})]^0$ toward protic reagents: the terminal hydride serves as a hydride donor (eq 1) whereas its bridging isomer serves as an electron donor (Scheme 3.3). The contrasting hydridic nature of terminal vs bridging hydrides parallels the behavior of the polyhedral boranes.\(^5^4\) Although the bridging hydride ligand does not serve as a $\text{H}_2$ source, the bridging hydride isomer does evolve hydrogen, but with the $\mu$-hydride ligand serving as a spectator after protonation produces a diferric dihydride. In the present case, the bridging hydride isomer is a superior electron donor. This effect is illustrated by the 220 mV difference in the $[1(t-\text{H})]^{+/0}$ vs $[1(\mu-\text{H})]^{+/0}$ couples.

3) **Electronically unsymmetrical Fe\(_2\) species stabilize terminal hydrides.** Previous work has consistently shown that HFe\(_2\)L\(_x\) compounds form symmetrical structures with bridging hydrides.\(^5^0,^5^5\) In this work, the introduction of an anionic C\(_5\)Me\(_5\)\(^-\) ligand strongly desymmetrizes the two iron centers such that the terminal hydride is stabilized.

Future work on the new complexes will focus on the reactivity of the unusual paramagnetic hydrides.\(^5^6\)
3.12. Experimental Methods

Materials and Methods. Unless otherwise noted, the experimental methods are the same as those described recently.\textsuperscript{16} Single-crystal X-ray diffraction data were collected using a Bruker SMART APEX II diffractometer with a CCD area detector (graphite monochromatic Mo K\textalpha radiation) at 173 K. Microanalytical data were acquired using Exeter Analytical CE-440 elemental analyzer EPR spectra were recorded on a Varian E-line 12” Century series X-band CW spectrometer. Infrared spectra, recorded on a Perkin Elmer FT-IR Spectrometer Spectrum Two, are reported for the $\nu_{\text{CO}}$ region only. Cyclic voltammetry was performed using a CHI 760e electrochemical workstation (Shanghai Chen Hua Instrument Co., Ltd) with a glassy carbon working electrode, Pt wire counter electrode, and the pseudo-reference electrode Ag wire.

Fe(pdt)(dppe)(CO)$_2$\textsuperscript{14} and [Cp*Fe(NCMe)$_3$]PF$_6$\textsuperscript{16} were prepared according to published procedures. [Cp*Fe(pdt)(dppe)(CO)$_2$]PF$_6$ (504 mg, 1.10 mmol) in 30 mL MeCN was added a solution of Fe(pdt)(dppe)(CO)$_2$ (600 mg, 1.10 mmol) in 20 mL of CH$_2$Cl$_2$ at −40 °C. The reaction solution was allowed to slowly warm to room temperature. After stirring for 30 min, the solution was concentrated under vacuum. Dilution of the concentrate with 50 mL of Et$_2$O resulted in precipitation of a brown powder. Yield: 870 mg (83%). Single crystals suitable for X-ray diffraction were obtained by slow diffusion of hexane into a CH$_2$Cl$_2$ solution of [1(CO)]PF$_6$ at −30 °C. $^1$H NMR (300 MHz, CD$_2$Cl$_2$): $\delta$ 7.68−7.28 (m, 20 \times \text{ArH}), 3.66−2.93 (m, 4H, Ph$_2$PCH$_2$), 2.76−2.32 (m, 6H, SCH$_2$ and CH$_2$CH$_2$CH$_2$) 1.38 (s, 15H, C$_5$Me$_5$). $^{31}$P{$^1$H} NMR (300 MHz, CD$_2$Cl$_2$): $\delta$ 62.7. FT-IR (MeCN): $\nu_{\text{CO}}$ = 1953, 1856 cm$^{-1}$. ESI-MS: Calcd for [1(CO)]$^+$, 807.1035; found, 807.1024. Anal. Calcd for C$_{41}$H$_{45}$F$_6$Fe$_2$O$_2$P$_3$S$_2$: C, 51.70 (51.93); H, 4.76 (4.97).
In a 50-mL Pyrex Schlenk tube, 100 mg of \([\text{I}(\text{CO})]\)PF$_6$ (0.10 mmol) was dissolved in 15 mL of MeCN. The solution was irradiated with a 365 nm LED array (15 W) until the conversion was complete (~ 4 h) as monitored by IR spectroscopy. The solvent was removed under vacuum, and product was extracted into 3 mL of MeCN. The extract was filtered through Celite and diluted with 30 mL of Et$_2$O to precipitate a brown powder. The resulting powder was recrystallized by layering a CH$_2$Cl$_2$ solution with hexane at –30 ºC. Yield: 78 mg (81%). $^1$H NMR (300 MHz, CD$_2$Cl$_2$): $\delta$ 7.58–7.08 (m, 20 × ArH), 3.54–3.04 (m, 4H, SCH$_2$CH$_2$SCH$_2$), 2.64 (s, 1H, SCH$_2$CH$_2$SCH$_2$), 2.54–2.08 (m, 4H, PPh$_2$CH$_2$CH$_2$PPh$_2$), 1.90 (m, 1H, SCH$_2$CH$_2$SCH$_2$) 1.37 (s, 15H, C$_5$Me$_5$), 1.27 (s, 3H, CH$_3$CN). $^3$P{$^1$H} NMR (300 MHz, CD$_2$Cl$_2$): $\delta$ 65.1. FT-IR (MeCN): $\nu$$_{CO}$ = 1769 cm$^{-1}$. ESI-MS: Calcd for [M-(MeCN)]$^+$, 779.1086; found, 779.1074. Anal. Calcd for C$_{42}$H$_{48}$NF$_6$Fe$_2$OP$_3$S$_2$ (found): C, 52.24 (52.56); H, 5.01 (5.47); N, 1.45 (1.62).

To a solution of [I(MeCN)]PF$_6$ (100 mg, 0.1 mmol) in 15 mL of MeCN was added 27 mg of Bu$_4$NBH$_4$ (0.105 mmol) solids. Over the course of 20 min at room temperature, the stirred solution deposited a black solid precipitate. Yield: 59 mg (75%). Black microcrystals were obtained by layering the toluene solution with pentane at –30 ºC. $^1$H NMR (400 MHz, CD$_2$Cl$_2$): $\delta$ 7.93–7.20 (m, 20 × ArH), 2.90–2.14 (m, 5H, SCH$_2$CH$_2$SCH$_2$), 2.09–1.71 (m, 5H, PPh$_2$CH$_2$CH$_2$PPh$_2$, SCH$_2$CH$_2$SCH$_2$) 1.38 (s, 15H, C$_5$Me$_5$), -4.76 (t, 1H, $J_{P-H}$ = 76 Hz, Fe-H). $^{31}$P{$^1$H} NMR (400 MHz, CD$_2$Cl$_2$): $\delta$ 87.7. FT-IR (CH$_2$Cl$_2$): $\nu$$_{CO}$ = 1752 cm$^{-1}$. ESI-MS: Calcd for [I(t-H)]$^0$: 780.1164; found, 780.1156. Anal. Calcd for C$_{40}$H$_{46}$Fe$_2$OP$_2$S$_2$ (found): C, 61.55 (61.79); H, 5.94 (6.28).

To a solution of [I(MeCN)]PF$_6$ (200 mg, 0.1 mmol) in 20 mL of THF was added 10 mg of NaBD$_4$ (0.24 mmol) solids. Over the course of 30
min at room temperature, solvent was removed under vacuum. The residue was extracted with toluene (2 mL × 2) and filtered through Celite, then laid with n-hexane, resulting in precipitation of a black solid. Yield: 72 mg (46%). 1H NMR (500 MHz, CD2Cl2): δ 7.93–7.24 (m, 20 × ArH), 2.90–2.19 (m, 5H, SCH2CH2SCH2), 2.10–1.78 (m, 5H, PPh2CH2CH2PPh2, SCH2CH2SCH2) 1.38 (s, 15H, C5Me5). 31P{1H} NMR (500 MHz, CD2Cl2): δ 88.2. 2H NMR (500 MHz, toluene): δ -4.66 (br). FT-IR (CH2Cl2): νCO = 1752 cm⁻¹.

[Cp*Fe(pdt)Fe(dppe)(CO)H]BF4 ([1(t-H)]BF4). A solution of [1(t-H)]0 (100 mg, 0.026 mmol) in 2 mL of CH2Cl2 was treated with one equiv of FcBF4 (35 mg) at –30 ºC. The color of the solution changed from brown to olive-green immediately. The reaction mixture was diluted with 20 mL of cold pentane (–30 ºC). The product precipitated as olive-green powder. Yield: 105 mg (94%). Single crystals suitable for X-ray diffraction were obtained by slow diffusion of hexane into a CH2Cl2 solution of [1(t-H)]BF4 at –30 ºC. FT-IR (CH2Cl2): νCO = 1875 cm⁻¹. ESI-MS: Calcd for [1(t-H)]⁺, 780.1164; found, 780.1153.

[(μ-D)Cp*Fe(pdt)Fe(dppe)(CO)]BF4 ([1(μ-D)]BF4). Compound [1(μ-D)]BF4 was prepared by oxidation of [1(t-D)]0 following the same procedure as [1(μ-H)]BF4. Yield: 80%. FT-IR (toluene): νCO = 1952 cm⁻¹.

[(μ-H)Cp*Fe(pdt)Fe(dppe)(CO)]BF4 ([1(μ-H)]BF4). In a 50-mL Schlenk flask, a solution of 60 mg (0.077 mmol) of [1(t-H)]0 in 10 mL of CH2Cl2 was maintained at room temperature. After ~ 24 h, the IR spectrum indicated complete isomerization. The solution was evaporated to dryness under reduced pressure. The residue was extracted into MeCN, filtering through Celite. The filtrate was concentrated, the concentrate diluted with 20 mL of Et2O, and this mixture was kept at –30 ºC overnight. Dark-green crystals were collected by filtration. Yield: 56
mg (80%). FT-IR (CH$_2$Cl$_2$): $\nu_{\text{CO}}$ = 1953 cm$^{-1}$. ESI-MS: Calcd for [I($\mu$-H)][$^{+}$, 780.1164; found, 780.1145. Anal. Calcd for C$_{40}$H$_{46}$Fe$_2$OP$_2$S$_2$ (found): C, 55.39 (55.73); H, 5.35 (5.68).

$[\mu$-$\text{H}]\text{Cp}^*\text{Fe(pdt)}\text{Fe(dppe)}\text{(CO)}\text{]}$ ([$\mu$-$\text{H}$])$^0$. To a dark-green solution of [I($\mu$-H)]BF$_4$ (100 mg, 0.115 mmol) in 20 mL of THF was added 22 mg of Cp$_2$Co (0.116 mmol). The color of the solution changed immediately. After 5 min, solvent was removed under vacuum. The residue was extracted into 6 mL of toluene, and the combined extracts were filtered through Celite. The green filtrate was layered with 50 mL of hexane, and this mixture was cooled at –30 ºC to afford green crystals. Yield: 75 mg (83%). $^1$H NMR (500 MHz, toluene-$d_8$): $\delta$ 7.71–7.55 (m, 8H, ArH), $\delta$ 7.14–6.99 (m, 12H, ArH), 2.48–2.46 (m, 2H, SCH$_2$CH$_2$SCH$_2$), 2.23–2.20 (m, 4H, PPh$_2$CH$_2$CH$_2$PPh$_2$), 2.13–2.03 (m, 4H, SCH$_2$CH$_2$SCH$_2$), 1.82 (s, 15H, C$_5$Me$_5$), –12.67 (t, 1H, $J_{\text{P-H}}$ = 28 Hz, Fe-$\text{H}$-Fe). $^{31}$P{ $^1$H} NMR (500 MHz, toluene-$d_8$): $\delta$ 80.2. FT-IR (THF): $\nu_{\text{CO}}$ = 1914 cm$^{-1}$. ESI-MS: Calcd for [I($\mu$-H)]$^0$, 780.1164; found, 780.1159. Anal. Calcd for C$_{40}$H$_{46}$Fe$_2$OP$_2$S$_2$ (found): C, 61.55 (61.87); H, 5.94 (6.16).

$[\mu$-$\text{D}]\text{Cp}^*\text{Fe(pdt)}\text{Fe(dppe)}\text{(CO)}\text{]}$ ([$\mu$-$\text{D}$])$^0$. Compound [I($\mu$-D)]$^0$ was prepared by reduction of [I($\mu$-D)]BF$_4$ following the same procedure as [I($\mu$-H)]$^0$. Yield: 86%. $^1$H NMR (500 MHz, toluene-$d_8$): $\delta$ 7.69–7.53 (m, 8H, ArH), $\delta$ 7.13–6.97 (m, 12H, ArH), 2.44–2.46 (m, 2H, SCH$_2$CH$_2$SCH$_2$), 2.22–2.18 (m, 4H, PPh$_2$CH$_2$CH$_2$PPh$_2$), 2.09–1.96 (m, 4H, SCH$_2$CH$_2$SCH$_2$), 1.80 (s, 15H, C$_5$Me$_5$). $^{31}$P{ $^1$H} NMR (500 MHz, toluene-$d_8$): $\delta$ 79.6. $^2$H NMR (500 MHz, toluene): $\delta$ -12.84 (br). FT-IR (toluene): $\nu_{\text{CO}}$ = 1914 cm$^{-1}$.

### 3.13. Computational Methods

The DFT calculations were performed with the BP86 functional$^{57,58}$ with the SDD pseudopotential and associated basis set$^{59}$ for the Fe centers, the 6-31G** basis set$^{60}$ for the $\mu$-H
and $t$-H ligands, and the 6-31G* basis set\textsuperscript{61,62} for all other atoms. Additional benchmarking was performed using the B3P86,\textsuperscript{57,63} M06-L,\textsuperscript{64} B3LYP,\textsuperscript{63,65} BLYP,\textsuperscript{58,65} TPSSH,\textsuperscript{66} and \omega B97XD\textsuperscript{67-70} functionals and is presented in Tables S6-S12. The starting coordinates for all geometry optimizations were obtained from the available crystal structures. The conductor-like polarizable continuum model\textsuperscript{71,72} was used to calculate the solvation free energies using the Bondi atomic radii\textsuperscript{73} and including non-electrostatic contributions from dispersion,\textsuperscript{74,75} repulsion,\textsuperscript{74,75} and cavitation formation.\textsuperscript{76} The solvent was THF unless stated otherwise. Structural and energetic information for all studied compounds are provided in the Supporting Information (Tables S13-S18).

Herein, the geometry optimizations were performed in the gas phase; however, a comparison of the structures optimized in the gas and solution phases is provided in the SI (Tables S2-S5). All minimum energy structures were verified to have no imaginary frequencies, and transition states (TSs) were confirmed to have only a single imaginary frequency. The TS for isomerization of the pdt\textsuperscript{2-} ligand was located utilizing the synchronous transit-guided quasi-Newton method,\textsuperscript{77,78} and the intrinsic reaction coordinate (IRC) was followed to verify that the TS leads to the correct reactant and product states (flippamers). Thermochemical data were calculated at $T = 298.15$ K. Zero-point energy, entropic contributions, and solvation effects were included in the calculated reaction free energies ($\Delta G$) and free energy barriers ($\Delta G^\ddagger$) in solution. For the $\Delta G$ value of the isomerization between the terminal and bridging hydride, an MP2\textsuperscript{79-81} single-point energy calculation was performed using the DFT-optimized geometry to obtain an MP2-corrected electronic energy. Thus, the electronic energy was obtained from MP2 while the zero-point energy, entropic, and solvation free energy contributions were obtained from DFT. This treatment of the electronic energy generally improves consistency among the functionals (Table S9).
procedure was applied to the neutral species but not to the oxidized species due to spin contamination issues, which were least pervasive for the BP86 functional (Table S10). The relative redox potentials were calculated using methodology described elsewhere. All calculations were performed with the Gaussian 09 electronic structure program.

3.14. References


(19) Adam, F. I.; Hogarth, G.; Kabir, S. E.; Richards, I. Models of the iron-only hydrogenase: Synthesis and protonation of bridge and chelate complexes [Fe₂(CO)₄{Ph₂P(CH₂)ₙPPh₂}(μ-pdt)] (n = 2-4) – Evidence for a terminal hydride intermediate. *C. R. Chim.* **2008**, *11*, 890-905.

(20) Barton, B. E.; Rauchfuss, T. B. Terminal hydride in [FeFe]-hydrogenase model has lower potential for H₂ production than the isomeric bridging hydride. *Inorg. Chem.* **2008**, *47*, 2261-2263.


(49) Argouarch, G.; Hamon, P.; Toupet, L.; Hamon, J.-R.; Lapinte, C. [(η⁵-C₅Me₅)Fe(Ph₂PCH₂CH₂CH₂PPh₂)][SO₃CF₃], a stable 16-electron complex with a coordinating counteranion and without agostic interaction: The dramatic role of a trivial methylene group. *Organometallics* 2002, 21, 1341-1348.


Chapter 4. Protonation of [NiFe]-Hydrogenase Models*†

4.1. Abstract

Theory and experiment indicate that the protonation of reduced NiFe dithiolates proceeds via a previously undetected isomer with enhanced basicity. In particular, it is proposed that protonation of (OC)$_3$Fe(pdt)Ni(dppe) (1; pdt$^{2−}$ = $^{−}$S(CH$_2$)$_3$S$^{−}$; dppe = Ph$_2$P(CH$_2$)$_2$PPh$_2$) occurs at the Fe site of the two-electron mixed-valence Fe(0)Ni(II) species, not the Fe(I)-Ni(I) bond for the homovalence isomer of 1. The new pathway, which may have implications for protonation of other complexes and clusters, was uncovered through studies on the homologous series L(OC)$_2$Fe(pdt)M(dppe) where M = Ni, Pd (2), and Pt (3) and L = CO, PCy$_3$. Similar to 1, complexes 2 and 3 undergo both protonation and 1 $\text{e}^{−}$ oxidation to afford well-characterized hydrides ([2H]$^{+}$ and [3H]$^{+}$) and mixed-valence derivatives ([2]$^{+}$ and [3]$^{+}$), respectively. Whereas the Pd site is tetrahedral in 2, the Pt site is square planar in 3, indicating that this complex is best described as Fe(0)Pt(II). In view of the results on 2 and 3, the potential energy surface of 1 was reinvestigated with density functional theory. These calculations revealed the existence of an energetically accessible and more basic Fe(0)Ni(II) isomer with a square planar Ni site.

4.2. Introduction

The protonation of metal complexes and clusters continues to attract much attention because of its relevance to catalysis involving proton and hydride transfer reactions. In organometallic and heterogeneous catalysis, protonation is a step in substrate hydrogenation and hydrogen evolution.$^{1-4}$ In biology, metal protonation is implicated in the reduction of CO$_2$ and N$_2$, not to mention hydrogen evolution.$^{2,5,6}$ For all of these reactions, a central question is the

† Author contributions: Dr. Dave Schilter and Prof. Thomas Rauchfuss were responsible for all experiments and associated analyses in this chapter.
regiochemistry of the protonation. For compounds with metal-metal bonds, protonation can in principle occur at a single metal or at the metal-metal bond.

The regiochemistry of protonation is relevant to a mechanistic understanding of the [NiFe]- and [FeFe]-hydrogenases, which also feature Ni-Fe and Fe-Fe bonding interactions.\(^7^,^8\) By proton transfer, these complexes equilibrate protons and dihydrogen, i.e., \(2 \text{H}^+ + 2 \text{e}^- \rightleftharpoons \text{H}_2\) in accordance with the local redox environment and activities of \(\text{H}^+\) and \(\text{H}_2\). Functional models for the active sites of these enzymes present opportunities for translating biological insights into practical catalytic systems.\(^9^,^11\) In addition to being fast and operating at low overpotentials, the hydrogenases are attractive targets for functional modeling because they utilize earth abundant metals.\(^12^,^13\) For the above reasons, much effort has been directed, both computationally\(^14^,^15\) and experimentally,\(^16^,^18\) at understanding the details of proton transfer to and from synthetic NiFe and FeFe complexes.

While many complexes engage separately in acid-base and redox reactions, few exhibit both kinds of reactivity, providing a possible explanation for the paucity of functional hydrogenase (H\(_2\)ase) models. With respect to the [NiFe]-H\(_2\)ases (Figure 4.1, left), the premier example of such a model is \((\text{OC})_3\text{Fe(pdt)Ni(dppe)}\) (1).\(^19\) Tricarbonyl 1 and analogous phosphine-substituted derivatives undergo: (i) protonation to afford rare examples of nickel-iron hydrides, and (ii) \(1 \text{e}^-\) oxidation to give Fe(I)Ni(II) derivatives (Figure 4.1, right). Several examples of charge-neutral Fe(I)Ni(I) species, as well as their related Fe(II)Ni(II) hydrides\(^20^,^22\) and Fe(I)Ni(II) cations,\(^23^,^24\) have been characterized.
The active site of [NiFe]-H$_2$ase, a bidirectional catalyst, participates in redox and acid-base chemistry (left). Similar behavior is observed for the prototypical model complex 1 (right).

The [NiFe]-H$_2$ase active site features Ni bound to four Cys ligands in a distorted tetrahedral (SF$_4$-like) arrangement. Two of the thiolates link to an Fe center whose coordination sphere includes the unusual CO and CN$^-$ cofactors, the latter H-bonding to neighboring Ser and Arg residues.$^7$ In the active states of the enzyme, the FeNi core appears to be relatively conformationally rigid, with the Fe···Ni distance and coordination geometry being rather insensitive both to redox state and the presence/absence of H$^-$ substrate.$^{25,26}$ The rigidity of the active site is undoubtedly a factor in the high rates and low overpotentials at which the interconversion of H$^+$ and H$_2$ is mediated.

In contrast to the rigid active site, small molecule mimics are conformationally dynamic. For example, 1 undergoes a tetrahedral → square planar twist at Ni upon oxidation or protonation (Figure 4.1). While the structures of 1, [1]$^+$, and [1H]$^+$ are established,$^{19,20-22,23,24}$ those of any intermediates are not. Such transformations may, for example, involve ‘pre-twisting’ of the Ni coordination in 1 (Scheme 4.1, top).
Distortion of the ground state structure of 1 could well affect oxidation states by inducing $2e^-$ mixed-valency.\textsuperscript{27} Furthermore, the degree of Fe–Ni bonding and, most importantly, reactivity, would also be perturbed. While ligand-centered isomerism is a common theme in the chemistry of multimetallic compounds, drastic differences in coordination geometry and metal-metal bonding between two (or more) interconverting species are extremely rare. They are, however, not without precedent; for example, Jones and co-workers reported X-ray structures for two complexes of formula (OC)$_2$Rh(µ-PR$_2$)$_2$Rh(CO)$_2$ (Scheme 4.1, bottom).\textsuperscript{28} One species takes the form of a red compound featuring Rh–Rh bonding (2.761 Å), in which one Rh center is square planar and the other tetrahedral. A second (orange) isomer was also observed, in which non-interacting Rh centers (3.717 Å) adopt square planar geometries.

The present study involves parallel experimental and density functional theory (DFT) investigations to obtain a complete structural, energetic, and mechanistic picture of complexes derived from FeNi complex 1. This work was extended to FePd and FePt species such that a homologous series can be interrogated for information regarding conformational motions and metal-metal bonding.
4.3. Synthesis of Neutral Compounds

The major routes to low-valent FeNi complexes of type (OC)_3Fe(pdt)Ni(diphosphine) involve treatment of (pdt)Ni(diphosphine) with either (i) Fe_2(CO)_9 or (ii) FeL_2(CO)_4/CoCp. As applied to the Pd and Pt derivatives, the second route proved more attractive, giving the diamagnetic heterobimetallics (OC)_3Fe(pdt)Pd(dppe) (2) and (OC)_3Fe(pdt)Pt(dppe) (3) (Scheme 4.2).

![Scheme 4.2.](image)

Isolated as dark green needles, 2 is sensitive to O_2 (decomposing to CO-free paramagnetic species) but is thermally stable in the solid state under N_2. The energies of the two ν_CO bands (A + E modes) compare well with those observed for 1 (Table 4.1).

Table 4.1. IR Data (ν_CO / cm\(^{-1}\)) for FeM (M = Ni, Pd, Pt) Complexes in CH_2Cl_2 Solution.

<table>
<thead>
<tr>
<th>Compound</th>
<th>FeNi</th>
<th>FePd</th>
<th>FePt</th>
</tr>
</thead>
<tbody>
<tr>
<td>(OC)_3Fe(pdt)M(dppe)</td>
<td>2028, 1952</td>
<td>2024, 1955</td>
<td>1962, 1883, 1870</td>
</tr>
<tr>
<td>[(OC)_3FeH(pdt)M(dppe)]BF_4</td>
<td>2082, 2024</td>
<td>2081, 2020</td>
<td>2081, 2022</td>
</tr>
<tr>
<td>[(OC)_3Fe(pdt)M(dppe)]BF_4</td>
<td>2057, 1986</td>
<td>2058, 1986</td>
<td>2058, 1987</td>
</tr>
<tr>
<td>[Cy_3P(OC)_2Fe(pdt)M(dppe)]BF_4</td>
<td>1966, 1899</td>
<td>1964, 1900</td>
<td>1966, 1900</td>
</tr>
</tbody>
</table>
The FePd species 2 is similar to 1 in that its room temperature $^{31}$P{^1}H NMR spectrum features a broad singlet resonance ($\delta$ 51.3) for the dppe ligand. At lower temperature, the resonance for 1 decoalesces into two peaks (well-resolved at −68 °C), consistent with a dynamic process that interconverts the $^{31}$P sites. The previously reported coalescence temperature (−30 °C)$^{21}$ corresponds to a free energy barrier of 9.8 kcal/mol for the FeNi species 1. This process presumably occurs via an intermediate or transition state with square planar Ni. No decoalescence was observed for 2 even at −85°C, behavior suggestive of either: (i) a tetrahedral Pd ground state with a particularly low rotational barrier or (ii) a square planar Pd ground state. The second possibility is ruled out based on the similarity of the IR data for 1 and 2. The structure of 2 in the solid state was confirmed crystallographically (Figure 4.2).

**Figure 4.2.** ORTEP of 2 with ellipsoids drawn at the 50% probability level and H atoms omitted for clarity. Selected distances /Å (mean DFT values in parentheses): Pd1-Fe1, 2.561 (2.56); Pd1-P1, 2.249; Pd1-P2, 2.271 (2.26); Pd1-S1, 2.573; Pd1-S2, 2.402 (2.49); Fe1-S1, 2.299; Fe1-S2, 2.275 (2.30); Fe1-C30, 1.795; Fe1-C31, 1.793; Fe1-C32, 1.798 (1.77). Selected angles /° (mean DFT values in parentheses): P1-Pd1-S1, 135.2; P1-Pd1-S2, 118.83; P2-Pd1-S1, 115.7; P2-Pd1-S2, 122.5 (135.1); S1-Fe1-C30, 162.2; S2-Fe1-C32, 163.1 (161.1).
Given the covalent radii of Pd (1.39 Å) and low-spin Fe (1.32 Å), the Fe-Pd distance of 2.561 Å suggests the presence of bonding between the metal centers in 2. The intermetallic separation is somewhat shorter in 1 (2.467 Å). Metal-metal interactions aside, the Pd and Fe centers in 2 exist in roughly tetrahedral and square-pyramidal (τ = 0.01 indicates a low degree of trigonality) coordination geometries, respectively. The complex is isostructural to its Fe(I)Ni(I) congener, and it is expected that an Fe(I)Pd(I) description is appropriate, particularly considering the tetrahedral Pd coordination, as such a geometry is unlikely for Pd(II).

The structures of several complexes reported here were calculated using DFT, and the optimized structures were in excellent agreement with available crystal structures (selected metrics are presented in Tables S3 and S4). In the case of 2, calculations accurately predicted the geometry of Pd, as well as its proximity to Fe. Additional isomers characterized as minima with DFT will be discussed below. NBO (natural bond orbital) analysis revealed the Fe-Pd bond to arise from overlap of an Fe-centered $d(z^2)$ orbital with a Pd-centered orbital with $d$ character (Figure S56b). NBO analysis of FeNi congener 1 also revealed metal-metal bonding (Figure S56a) consistent with previous DFT analyses of this system. NBO analysis of other Fe-M complexes discussed herein did not indicate mixing of metal $d$ orbitals; hence, no Fe-M bonding is expected for these species.

4.4. Mixed-Valence Derivatives

The cyclic voltammogram of 2 features reversible and quasi-reversible oxidations at $E_{1/2} = -0.87$ and $+0.26$ V, respectively (vs ferrocene/ferrocenium, Fe$^{0/+}$; Figure 4.3). The reversible wave, assigned to the $[2]^{0/+}$ couple, is 300 mV more negative than the wave for $[1]^{0/+}$, a result reflective of the more reducing nature of the heavier group 10 metal.
DFT was used to calculate the potentials associated with the two couples for each 
\[ [L(CO)_2Fe(pdt)M(dppe)]^{0/+2+} \] (L = CO, PCy\textsubscript{3}; M = Ni, Pd, Pt) system (Table 4.2). The calculated potentials are in excellent agreement with the experimental values for the neutral/monocation couples, while the values for the monocation/dication couples are in qualitative but not quantitative agreement. The latter discrepancies are mainly a result of these more anodic couples being only partially reversible, the associated potentials being less thermodynamically reliable.

Chemical oxidation of \( \text{2} \) with \( \text{FcBF}_4 \) afforded \([\text{(OC)}_3\text{Fe(pdt)Pd(dppe)}]\text{BF}_4 \) ([\( \text{2} \)]\text{BF}_4), isolated as an olive-brown powder. Positive ion ESI-MS analysis (\( m/z \) 749.8) was consistent with the formula, and the lability of \([\text{(OC)}_3\text{Fe(pdt)Pd(dppe)}]\text{\textsuperscript{+}} \) was evidenced by observation of CO dissociation products (\( m/z \) 721.8, 693.8). The \( \nu_{\text{CO}} \) bands for \([\text{2}]^+ \) (2058, 1986 cm\(^{-1}\)) are only modestly shifted from those of \( \text{2} \), implicating oxidation at the Pd center. EPR measurements also indicated an Fe(I)Pd(II) description for \([\text{2}]^+ \), the assignment being based on the magnitude of the g shifts and the absence of \( ^{31}\text{P} \) hyperfine coupling (Figure 4.4, top; for simulated EPR spectra and parameters see Figure S23 and Table S1, respectively). The data are similar to those for
Fe(I)Ni(II) complexes, for which theory indicated an Fe-centered $d(z^2)$ doublet ground state singly occupied molecular orbital (SOMO). These computational results, as well the implied absence of spin-orbit coupling expected for a $d(z^2)$ ground state, are consistent with a resonance in the vicinity of $g_c$. Owing to the inversion of the pdt$^{2-}$ ring, the [(OC)$_3$Fe(pdt)M(dppe)]$^+$ radicals exist as two conformers, one giving rise to an axial pattern and the other to a rhombic pattern.

Table 4.2. Redox Potentials (V vs Fe$^{0+/+}$, Experimental and DFT-Calculated) and Associated Electrochemical Data (Recorded at 0.1 Vs$^{-1}$) for FeNi, FePd, and FePt Complexes (1 mM) in NBu$_4$PF$_6$ Electrolyte (100 mM CH$_2$Cl$_2$ Solution).

| Complex | $E_{1/2}$ (Fe$^{1+}$/Fe$^{1+}$) (|$i_{pc}$/|$i_{pa}$|, |$\Delta E_{pl}$|) | $E^{\circ}$ (Fe$^{1+}$/Fe$^{1+}$) (|$i_{pc}$/|$i_{pa}$|, |$\Delta E_{pl}$|) | $E_{1/2}$ (Fe$^{1+}$/Fe$^{1+}$) (|$i_{pc}$/|$i_{pa}$|, |$\Delta E_{pl}$|) | $E^{\circ}$ (Fe$^{1+}$/Fe$^{1+}$) |
|---------|---------------------------------|---------------------------------|---------------------------------|---------------------------------|
| 1       | +0.34$^b$                       | +0.52                           | -0.59 (0.90, 0.059)             | -0.62                           |
| [1$'$]BF$_4$ | -0.14 (0.68, 0.076)$^b$           | -0.14$^d$                       | -0.92 (1.00, 0.072)$^c$         | -0.92$^d$                       |
| 2       | +0.26 (0.29, 0.115)             | +0.39                           | -0.87 (0.98, 0.061)             | -0.92                           |
| [2$'$]BF$_4$ | -0.20 (0.67, 0.080)             | -0.20$^e$                       | -1.14 (0.83, 0.122)             | -1.14$^e$                       |
| 3       | +0.33 (0.49, 0.168)             | N.D.$^f$                        | -0.93 (0.97, 0.105)$^g$         | -0.91$^g$                       |
| [3$'$]BF$_4$ | -0.19 (0.50, 0.076)             | -0.24                           | -1.37 (0.66, 0.067)             | -1.38                           |

$^a$$\Delta E_{pl}$(Fe$^{0+/+}$) = 0.064 V under these conditions. $^b$Anodic potential for irreversible oxidation. $^c$Data from reference.$^{24}$ $^d$Calculated and experimental values for [1$'$]BF$_4$ agree by construction and were used as references for the corresponding [1] waves. $^e$Calculated and experimental values for [2$'$]BF$_4$ agree by construction and were used as references for the corresponding [2], [3] and [3$'$] waves. $^f$This value was not determined as difficulties arising from local minima led to inconclusive results. $^g$This wave is assigned to a Fe$^{0}$/Pt$^{II}$/Fe$^{1+}$/Pt$^{II}$ couple (vide infra).

Figure 4.4. X-band EPR spectra (CH$_2$Cl$_2$/PhMe, 110 K) of [2]BF$_4$ and [2$'$]BF$_4$. 

87
DFT analysis of spin densities was used to determine the localization of the unpaired electron in [2]⁺ as well as all other mixed-valence bimetals (Table 4.3). In each case, the spin is entirely Fe-centered, consistent with the Fe(I)M(II) assignments from EPR analyses. The spin density distribution is unaffected by inversion of the pdt²⁻ ring, as observed for related FeNi complexes.²³,²⁴

<table>
<thead>
<tr>
<th>Complex</th>
<th>(\rho(M))</th>
<th>(\rho(Fe))</th>
</tr>
</thead>
<tbody>
<tr>
<td>[(OC)₃Fe(pdt)Ni(dppe)]⁺ ([1]⁺)</td>
<td>0.04</td>
<td>1.03</td>
</tr>
<tr>
<td>[Cy₃P(OC)₂Fe(pdt)Ni(dppe)]⁺ ([1']⁺)</td>
<td>0.01</td>
<td>1.14</td>
</tr>
<tr>
<td>[(OC)₃Fe(pdt)Pd(dppe)]⁺ ([2]⁺)</td>
<td>0.03</td>
<td>1.04</td>
</tr>
<tr>
<td>[Cy₃P(OC)₃Fe(pdt)Pd(dppe)]⁺ ([2']⁺)</td>
<td>0.03</td>
<td>1.06</td>
</tr>
<tr>
<td>[(OC)₃Fe(pdt)Pt(dppe)]⁺ ([3]⁺)</td>
<td>0.03</td>
<td>1.03</td>
</tr>
<tr>
<td>[Cy₃P(OC)₃Fe(pdt)Pt(dppe)]⁺ ([3']⁺)</td>
<td>0.04</td>
<td>1.05</td>
</tr>
</tbody>
</table>

Table 4.3. Calculated Spin Densities on Metal Centers in Mixed-Valence Fe(I)M(II) Species.

Cation [2]⁺ undergoes substitution with PCy₃ to afford [Cy₃P(OC)₂Fe(pdt)Pd(dppe)]BF₄ ([2']BF₄) as a yellow-brown solid. Similar to its tricarbonyl precursor, [2']⁺ is also described as an Fe(I)Pd(II) compound, the rhombic signals split by interaction of the SOMO with \(^{31}\)PCy₃ (\(\Delta^{31}P = 54, 61, 57\) MHz). The data are distinct from those of the green Fe(I)Ni(II) analog [1']⁺, which show \(^{31}\)P hyperfine to be absent due to structural distortions at Fe.

As with 2, the cyclic voltammogram of [2']BF₄ features two waves, the events at \(E_{1/2} = -1.14\) and \(-0.20\) V being assigned to [2']⁰⁺ and [2']⁺⁺ couples, respectively. The waves are cathodically shifted relative to the tricarbonyl parent, with a smaller shift in the first wave (\(\Delta E_{1/2} = -0.27\) V) than in the second (\(\Delta E_{1/2} = -0.46\) V). These trends were also reproduced by DFT calculations (Table 4.2). The fact that ligand substitution at Fe perturbs the anodic wave to a greater degree points to this redox event being Fe-centered. These data are consistent with the sequence: Fe(I)Pd(I) \(\leftrightarrow\) Fe(I)Pd(II) \(\leftrightarrow\) Fe(II)Pd(II). The reversibility of the
Fe(I)Pd(II)/Fe(II)Pd(II) ([2']^{+/2+}) couple is greater than that of [2]^{+/2+}, the ligands in the former case being better suited to stabilize the ferrous state.

4.5. Palladium-Iron Hydride

The protonation of 2 with HBF₄ affords the hydride salt [2H]BF₄, the formation of which is signaled by the appearance of high energy νCO bands (2081, 2020 cm⁻¹) suggestive of an Fe(II)HPd(II) product. A high-field ¹H NMR signal (δ ~4.3) and a single downfield ³¹P NMR resonance (δ 62.8) are consistent with [2H]⁺ adopting a Cₛ-symmetric structure in solution. When the acidification of 2 was performed in the presence of CD₃OD, ESI-MS and ²H NMR spectroscopy confirmed formation of the deuteride isotopologue [2D]⁺. The evidence for the solution structure of [2H]⁺ is complemented by crystallographic analysis (Figure 4.5).

Figure 4.5. ORTEP of one of two independent complexes in [2H]BF₄·THF·0.5Et₂O with ellipsoids drawn at the 50% probability level. The solvate molecules, BF₄⁻ anion and H atoms (except the H⁻ ligand) are omitted for clarity. Mean distances in the complexes /Å (mean DFT values in parentheses): Pd2-Fe2, 2.882 (2.92); Pd2-H2, 2.173 (2.19); Pd2-P3, 2.253; Pd2-P4, 2.270 (2.30); Pd2-S3, 2.367; Pd2-S4, 2.363 (2.39); Fe2-S3, 2.330; Fe2-S4, 2.338 (2.34); Fe2-H2, 1.483 (1.54). Mean angles in the complexes ° (mean DFT values in parentheses): Pd2-H2-Fe2, 102.4 (101.2); P3-Pd2-S4, 171.5; P4-Pd2-S3, 177.7 (177.6); S3-Fe2-C35, 165.4; S4-Fe2-C33, 167.1 (165.8).

By single crystal X-ray diffraction, two structurally similar FePd hydrides were found in the asymmetric unit of [2H]BF₄, in each case with the Fe-Pd distance (2.845, 2.918 Å) being
greater than the sum of the covalent radii (2.71 Å).\textsuperscript{31} Protonation at Fe is accompanied by rotation of the Pd(dppe) center, affording square planar Pd linked by the pdt\textsuperscript{2-} group to octahedral Fe. Indeed, the arrangement of metal centers in [2H]\textsuperscript{+} is analogous to that in [(OC)\textsubscript{3}FeH(pdt)Ni(dppe)]\textsuperscript{+} ([1H]\textsuperscript{+}). The H\textsuperscript{-} ligand in [2H]\textsuperscript{+} was located in the difference map and allowed to refine; its final location was significantly closer to the Fe center. Considering the covalent radius of H (0.31 Å),\textsuperscript{31} the H\textsuperscript{-} ligand in [2H]\textsuperscript{+} is formally bonded to Fe (1.469 Å), whereas its interaction with Pd is weak (2.168 Å). DFT calculations also predict unsymmetrical binding of the hydride ligand, which, despite being poised between the metals, can be viewed as a terminal ligand.

4.6. Platinum-Iron Complexes

The first indication of the unusual structure of (CO)\textsubscript{3}Fe(pdt)Pt(dppe) (3), which is red-brown, is that its UV-vis spectrum features two absorptions (480, 360 nm), whereas single bands were observed for the green complexes 1 (398 nm) and 2 (376 nm). The IR spectrum of 3 is also distinctive, with the ν\textsubscript{CO} bands (1962, 1883, 1870 cm\textsuperscript{-1}) observed at energies ~70 cm\textsuperscript{-1} lower than those for Fe(I)Ni(I) and Fe(I)Pd(I) analogs. The \textsuperscript{31}P\{\textsuperscript{1}H\} NMR spectrum for 3 consists of a singlet (and \textsuperscript{195}Pt-coupled doublet) and is temperature invariant down to −85 °C. The chemical shift (δ 44.1) and \textsuperscript{1}J\textsubscript{PPt} (3277 Hz) are close to those of the (pdt)Pt\textsuperscript{II}(dppe) precursor (δ 44.5, \textsuperscript{1}J\textsubscript{PPt} = 2717 Hz). While NMR data are in line with Pt coordination being either (i) highly fluxional tetrahedral or (ii) square planar, the latter is most consistent with low energy ν\textsubscript{CO} bands. Square planarity at Pt is corroborated by DFT geometry optimization (Figure 4.6).
Figure 4.6. Energy-minimized structure of 3 obtained with DFT. Selected distances /Å: Fe-Pt, 2.87; Fe-C1, 1.75; Fe-C2, 1.75; Fe-C3, 1.76; Fe-S1, 2.34; Fe-S2, 2.37; Pt-S1, 2.37; Pt-S2, 2.38; Pt-P1, 2.28; Pt-P2, 2.28. Selected angles /°: C1-Fe-S2, 169.4; C2-Fe-S1, 136.8; S1-Pt-P2, 176.1; S2-Pt-P1, 177.7.

Indeed, rather than being isostructurual to 1 and 2, 3 is predicted to feature a square planar Pt site, with the angle between the PtP₂ and PtS₂ planes being 4°. The calculated Fe-Pt distance (2.87 Å) exceeds the sum of the covalent radii (2.68 Å), suggesting an absence of Fe-Pt bonding and a purely pentacoordinate Fe. This site, in contrast to the square-pyramidal Fe(I) in 2, is distorted trigonal bipyramidal ($\tau = 0.54$). As mentioned previously, NBO analysis did not reveal $d$ orbital mixing, which was observed in 1 and 2. Overall, the coordination geometries are consistent with an Fe(0)Pt(II) description for 3.

Quasi-reversible oxidations for (CO)$_3$Fe(pdt)Pt(dppe) were observed at −0.93 and +0.33 V (Figure S29), potentials similar to those for the FePd analog 2. The proposed sequence of 1e$^-$ events: Fe(0)Pt(II) ↔ Fe(I)Pt(II) ↔ Fe(II)Pd(II) is supported by the IR spectrum for [(OC)$_3$Fe(pdt)Pt(dppe)]BF₄ ([3]BF₄, prepared from 3 and FcBF₄), in which the $\nu_{CO}$ bands (2058, 1987 cm$^{-1}$) are at virtually identical energies to the bona fide Fe(I)M(II) species [1]$^+$ and [2]$^+$. While the EPR spectrum of [3]$^+$ indicates some unpaired spin resides on Pt ($A^{(195Pt)} \approx 97$ MHz,
\( A(31P) \approx 11 \text{ MHz} \), the \( g \)-shifts suggest an Fe(I)Pt(II) description to be appropriate, as mirrored in the spin density calculations (Table 4.3). The paramagnetic salt \([3]\text{BF}_4\) undergoes ligand substitution with PCy\(_3\) to afford dicarbonyl \([\text{Cy}_3\text{P(OC)}_2\text{Fe(pdt)}\text{Pt(dppe)}]\text{BF}_4\) (\([3']\text{BF}_4\)), the characterization for which mirrors that for \([2']\text{BF}_4\).

As expected, reaction of low-valent \(3\) with HBF\(_4\) furnished the hydride salt \([\text{(OC)}_3\text{FeH(pdt)}\text{Pt(dppe)}]\text{BF}_4\) (\([3\text{H}]\text{BF}_4\)), for which a \( ^1\text{H} \) NMR resonance at \( \delta -3.57 \) is particularly diagnostic. The small coupling observed for the satellites (\( ^1J_{\text{HPt}} = 173 \text{ Hz} \)) indicates that the \( \text{H} \cdots \text{Pt} \) interactions are rather weak. For comparison, square planar platinum hydrides exhibit \( ^1J_{\text{HPt}} \) from \( \sim 650 \) to \( 1700 \text{ Hz} \),\(^{34}\) with \( ^1J_{\text{HPt}} \) being at the low end of this range for five coordinate complexes of the type \([\text{Pt(PR}_3)_4\text{H}]^+\).\(^{35,36}\) Thus, \([3\text{H}]^+\) is similar to its FePd analog in that the \( \text{H}^- \) ligand can be considered terminally Fe-bound, as also supported by the \( 31P \) NMR resonances for \(3\) being insensitive to protonation at Fe (for \([3\text{H}]^+ \delta 46.4, \ ^1J_{\text{PPt}} = 3220 \text{ Hz} \)). In contrast, the energies of the \( \nu_{\text{CO}} \) bands for the FePt conjugate acid and base species are very different – those for \([3\text{H}]^+ \) (2081, 2022 cm\(^{-1}\)) are greatly shifted relative to those for \(3\) (the average shift, taking into account the A + E patterns was \( \Delta \nu_{\text{CO}} = 137 \text{ cm}^{-1}\)). A similar shift (139 cm\(^{-1}\)) was observed by Angelici and co-workers upon conversion of the Fe(0) carbonyl \( \text{Fe(dppe)(CO)}_3 \) to the Fe(II) hydride \([\text{FeH(dppe)(CO)}_3]^+\).\(^{37}\) Taken together, these data are consistent with Fe(0)Pt(II), Fe(II)Pt(II), and Fe(II)HPt(II) assignments for \(3\), \([3]\text{BF}_4\), and \([3\text{H}]\text{BF}_4\), respectively.

While only a handful of FePt complexes bearing \( \text{H}^- \) and CO ligands exist, the compounds \((\text{CO})_3\text{FeH(µ-PR}_2\text{)}\text{Pt(PR}_3\text{)}_2\) are particularly pertinent.\(^{38}\) These complexes interconvert between isomers with terminal ((H)Fe\(^0\)Pt\(^0\)) and bridging (Fe\(^0\)(µ-H)Pt\(^0\)) hydride ligands. When these FePt species feature the substituents \( R = \text{Ph}, R' = \text{OPh} \), a non-fluxional terminal hydride results (\(^1J_{\text{HPt}}\)
= 28 Hz), whereas when R = Cy, R’ = Et the hydride is evenly shared ($J_{\text{H}^1 \text{Pt}} = 520$ Hz). One would expect values double the latter (~1000 Hz) for purely terminal Pt–H moieties.$^{39}$

\[ \text{[1H]}^+ \text{Fe(II)Ni(II)} \xrightarrow{\text{H}^+} \text{Fe(II)Ni(I)} \xrightarrow{\text{Pt}^2-} \text{Fe(II)Pt(II)} \]

\[ \text{[2H]}^+ \text{Fe(II)Pd(II)} \xrightarrow{\text{H}^+} \text{Fe(I)Pd(II)} \xrightarrow{\text{Pt}^2-} \text{Fe(II)Pd(II)} \]

\[ \text{[3H]}^+ \text{Fe(II)Pt(II)} \xrightarrow{\text{H}^+} \text{Fe(0)Pt(II)} \xrightarrow{\text{Pt}^2-} \text{Fe(II)Pt(II)} \]

\[ \Delta\nu_{\text{CO}} = 66 \text{ cm}^{-1} \text{ (exp.)} \quad 56 \text{ cm}^{-1} \text{ (DFT)} \]

\[ \Delta\nu_{\text{CO}} = 32 \text{ cm}^{-1} \text{ (exp.)} \quad 32 \text{ cm}^{-1} \text{ (DFT)} \]

\[ \Delta\nu_{\text{CO}} = 137 \text{ cm}^{-1} \text{ (exp.)} \quad 112 \text{ cm}^{-1} \text{ (DFT)} \]

\[ \Delta\nu_{\text{CO}} = 32 \text{ cm}^{-1} \text{ (exp.)} \quad 27 \text{ cm}^{-1} \text{ (DFT)} \]

\[ \Delta\nu_{\text{CO}} = 72 \text{ cm}^{-1} \text{ (exp.)} \quad 90 \text{ cm}^{-1} \text{ (DFT)} \]

\[ [1]^{+} \text{Fe(II)Ni(II)} \xrightarrow{\text{Pt}^2-} \text{Fe(II)Ni(II)} \]

\[ [2]^{+} \text{Fe(I)Pd(II)} \xrightarrow{\text{Pt}^2-} \text{Fe(II)Pd(II)} \]

\[ [3]^{+} \text{Fe(I)Pt(II)} \xrightarrow{\text{Pt}^2-} \text{Fe(II)Pt(II)} \]

\[ \Delta\nu_{\text{CO}} \approx 50 \text{ cm}^{-1} \]

\[ \Delta\nu_{\text{CO}} \approx 110 \text{ cm}^{-1} \]

\[ \Delta\nu_{\text{CO}} \text{ values correlate with the formal changes in Fe oxidation state: from } +1 \text{ to } +2 \text{ upon protonation of 1 and 2 and from 0 to } +2 \text{ upon protonation of 3. For the 1 e}^{-} \text{ oxidations, trends in } \Delta\nu_{\text{CO}} \text{ for the couples } [1]^{+}/[0], [2]^{+}/[0], \text{ and } [3]^{+}/[0] \text{ are similar to those seen for protonation, the effects again being greater for the Pt case. The } \Delta\nu_{\text{CO}} \text{ values for these oxidations are smaller,} \]

\[ \text{Scheme 4.3.} \]

4.7. Comparison of CO Frequencies

Good agreement was found between experimental and calculated shifts in weighted average CO vibrational frequency ($\Delta\nu_{\text{CO}}$) induced by protonation and oxidation of bimetals 1-3 (Scheme 4.3). Protonation of the FeNi and FePd complexes is predicted to shift $\nu_{\text{CO}}$ by ~50 cm$^{-1}$. In contrast, the shift induced by protonation of the FePt analog is much greater, with $\Delta\nu_{\text{CO}}$ ~110 cm$^{-1}$. These $\Delta\nu_{\text{CO}}$ values correlate with the formal changes in Fe oxidation state: from +1 to +2 upon protonation of 1 and 2 and from 0 to +2 upon protonation of 3. For the 1 e$^{-}$ oxidations, trends in $\Delta\nu_{\text{CO}}$ for the couples [1]$^{+}/[0]$, [2]$^{+}/[0]$, and [3]$^{+}/[0]$ are similar to those seen for protonation, the effects again being greater for the Pt case. The $\Delta\nu_{\text{CO}}$ values for these oxidations are smaller,
consistent with their occurring remote from the Fe(CO)₃ center for [1]⁺/0 and [2]⁺/0 or entailing only a one-electron change for [3]⁺/0.

4.8. Stereochemical Non-Rigidity

As discussed above, the fluxionality of 1 and 2, evidenced by the broad singlet in their ³¹P{¹¹H} NMR spectra, implicate the presence of a transient or intermediate species featuring square planar Ni/Pd centers (Scheme 4.1, top). This conformational motion is likely to perturb both the oxidation states and intermetallic bonding of the Fe and M centers. Two optimized DFT structures characterized as minima were calculated for each bimetallic 1-3: one in which the (pdt)M(dppe) fragment is tetrahedral, and one in which it is square planar. Selected bond lengths and angles are provided in Table S5. In each case, the Fe center is within bonding distance of M when the latter adopts a tetrahedral geometry, while the Fe-M distance is elongated when M is planar. The Fe-M elongation is subtle in the FeNi case (0.13 Å), but more pronounced for FePd and FePt (~0.30 Å). The relative stability of the two isomers was computed, and the reaction free energies (ΔG°) and free energy barriers (ΔG†) associated with tetrahedral to square planar isomerization are presented in Table 4.4.

### Table 4.4. Calculated Reaction Free Energies and Free Energy Barriers (kcal/mol) for Rotation from Tetrahedral to Square Planar Geometry at the M(dppe) Site in Compounds 1, 2, and 3.

<table>
<thead>
<tr>
<th>Complex</th>
<th>ΔG°</th>
<th>ΔG°⁺/₀</th>
<th>ΔG†</th>
</tr>
</thead>
<tbody>
<tr>
<td>(OC)₃Fe(pdt)Ni(dppe) (1)</td>
<td>−0.67</td>
<td>6.66</td>
<td>9.5</td>
</tr>
<tr>
<td>(OC)₃Fe(pdt)Pd(dppe) (2)</td>
<td>+0.99</td>
<td>3.29</td>
<td>&lt;7.2</td>
</tr>
<tr>
<td>(OC)₃Fe(pdt)Pt(dppe) (3)</td>
<td>−30.42</td>
<td>ND</td>
<td>ND</td>
</tr>
</tbody>
</table>

*The potential energy surface was found to be extremely flat along the isomerization pathway. As a result, the single imaginary frequency for the transition state (TS) was less than 20 cm⁻¹ for the two barriers given in this table. Each TS was verified to lead to the relevant tetrahedral and square planar geometries by following the intrinsic reaction coordinate (IRC) for 4 or 5 steps in both directions and subsequently optimizing the geometries. As the imaginary frequencies may be below the numerical accuracy of the methodology, and the complete IRC was not obtained, the free energy barriers should be viewed with caution. Moreover, multiple TSs were found for 1, suggesting a ruffled potential energy surface connecting the two isomers, and only the highest free energy barrier is reported. The TS was not determined for 3 as the isomerization was found to be overly exergonic. Estimated according to the Gutowsky-Holm relation using the coalescence temperature of 243 K (detailed in SI).*
For the FeNi and FePd congeners, the isomerization is close to thermoneutral ($\Delta G^\circ = -0.67$ and $+0.99$ kcal/mol, respectively). Given the error associated with the calculations, it is not possible to determine a thermodynamic preference for the coordination geometry at Ni/Pd. The reaction free energies calculated using various other density functionals and basis sets with geometry optimization in CH$_2$Cl$_2$ solvent are qualitatively consistent with the results presented here, although certain levels of theory do not predict a stable square planar intermediate for geometry optimizations in the gas phase (see Tables S6-S8). The calculated free energy barriers are relatively low, consistent with the fluxionality of 1 and 2 observed in their room temperature NMR spectra. Because the potential energy surfaces for 1 and 2 were found to be extremely flat along the isomerization pathway, the calculated free energy barriers may not be quantitatively accurate and can be analyzed only qualitatively. The calculated free energy barrier of 6.66 kcal/mol for 1 is qualitatively consistent with, although somewhat lower than, the free energy barrier of 9.5 kcal/mol estimated from the coalescence temperature in the NMR measurement. Moreover, the lower calculated free energy barrier of 3.29 kcal/mol for 2 is consistent with the lack of decoalescence observed in the NMR experiments for 2. The observation of a particularly flat potential energy surface along the isomerization pathway has implications for the enzymatic system, where the geometry at the Ni site is between tetrahedral and square planar.

The similarity in free energies and the relatively low free energy barriers suggest that the two isomers of 1 and 2 may interconvert. In addition, the free energy change associated with oxidation of 1 and 2 is similar for the two different isomers, with the calculated reduction potentials of the square planar and tetrahedral isomers differing by only $\sim$0.03 V, which is within the estimated error of the methodology. It is thus a distinct possibility that the Ni/Pd centers in 1 and 2 ‘preorganize’ into square planar geometries prior to 1e$^-\text{loss.}$ In the case of FePt complex
the square planar isomer is overwhelmingly favored, consistent with the high stability characteristic of square planar Pt(II) centers. Calculations indicate that $\nu_{\text{CO}}$ energies for the tetrahedral isomer of 3, which was found to be a local minimum, are within 10 cm$^{-1}$ of those for 1 and 2. Square-planar isomers of 1 and 2, while not observed experimentally, have calculated $\nu_{\text{CO}}$ bands 55-65 cm$^{-1}$ lower in energy than the respective tetrahedral conformers. This is consistent with the change in conformation inducing a change in oxidation state.

Given the dynamics in the neutral species 1 (as well as 2 and 3), we set about investigating hydride [1H]$^+$ to yield further insight into the mechanism of protonation. At room temperature, [(CO)$_3$FeH(pdt)Ni(dppe)]$^+$ displays a single $^{31}$P NMR peak owing to its $C_v$-symmetry. Given the difficulty with determining whether or not Ni(dppe) twisting is occurring here, we interrogated an analogous hydride of lower symmetry, which entailed preparation of the new bimetallic (CO)$_3$Fe(pdt)Ni(S,S-chiraphos) ($S,S$-chiraphos = 2$S$,3$S$-bis(diphenylphosphino)butane, Scheme 4.4). Analogous to 1, this species gave rise to a single $^{31}$P NMR resonance at room temperature, with decoalescence observed on cooling to $–90^\circ$C, at which temperature Ni($S,S$-chiraphos) twisting is slowed (Figure S49).

![Scheme 4.4.](image-url)

While (CO)$_3$Fe(pdt)Ni($S,S$-chiraphos) twists rapidly at room temperature, its conjugate acid does not; the non-equivalent $^{31}$P nuclei in the chiral hydride [(CO)$_3$FeH(pdt)Ni($S,S$-chiraphos)]$^+$ (Scheme 4.4, center) give rise to two broad signals (resolved into two $^{31}$P-coupled doublets at $–70^\circ$C). If twisting and equivalencing of the $^{31}$P sites were to be facile, it would be
necessary for a tetrahedral hydride (Scheme 4.4, right) to be energetically accessible, which appears not to be the case. In general, it is proposed that twisting of neutral species is associated with ‘on/off’ switching of Fe-M bonding in the FeNi and FePd (but not FePt) complexes. It is now evident that the twisting of hydride species is slower, perhaps reflecting the rigidifying influence of the H···M interactions (despite their weakness).

These observations are supported by computational work: the twisting of hydride [1H]$^+$ was investigated, and two optimized DFT structures characterized as minima were calculated for the isomers of [1H]$^-$. The relative free energies of the structures, which feature either square planar or tetrahedral (pdt)Ni(dppe) fragments (Figure 4.7), were also calculated.

![Image](image.png)

**Figure 4.7.** Energy-minimized structures obtained with DFT for [1H]$^+$ with the (pdt)Ni(dppe) moiety either square planar (red) or distorted tetrahedral (blue). The optimized geometries are superimposed and presented in two views. Other (higher energy) isomers/tautomers proposed can be found in Figure S57.

Twisting of the square planar (pdt)Ni(dppe) site in [1H]$^+$ to adopt a tetrahedral geometry was calculated to be significantly endergonic ($\Delta G^\circ = +11.47 \text{kcal/mol}$). The tetrahedral conformer is predicted to be vastly ($\sim10^8 \times$) more acidic than the square planar state (the experimental $pK_a^{\text{MeCN}}$ for [1H]$^+$ is 10.7). Considering the conjugate bases for the two isomers of [1H]$^+$, these results suggest that 1 has greatly enhanced basicity when in the square planar form, consistent with its description as Fe(0)Ni(II) (Scheme 4.1). Thus, it is reasonable that this
isomer is responsible for the rich acid-base and redox chemistry exhibited by 1 and related complexes.

4.9. Conclusions

Numerous studies have examined the protonation and 1 e⁻ oxidation of complexes of the type (CO)₃Fe(pdt)Ni(diphosphine) and substituted derivatives thereof. This paper presents evidence that the protonations, and possibly the oxidations, proceed via a latent 2 e⁻ mixed-valence intermediate (Scheme 4.5, top).

![Scheme 4.5](image)

Despite previous computational analyses,¹⁹,²²,²⁴ this key square planar species had not been identified as an energetically accessible isomer of 1. The first clue for this unsuspected structure was provided by the ³¹P NMR properties of 1, which implicate a transient or intermediate square planar species that interconverts the ³¹P sites. The new and compelling evidence for the intermediate comes from (i) DFT calculations that suggested thermodynamically accessible rotation at the non-Fe center and (ii) spectroscopic and
electrochemical properties of the FePt complex, the protonation and redox behavior of which mirrors that of Fe(0) diphosphines.\textsuperscript{37,40}

The rotation-induced 2 e\(^{-}\) mixed-valence behavior also has implications for understanding active site mimics of the [FeFe]-H\textsubscript{2}ases. Prior to protonation, model complexes almost invariably feature Fe\(_2\)(SR)\(_2\)L\(_6\) cores wherein terminal coordination sites are \textit{un}available for H\(^+\) binding (Scheme 4.5). Protonation of these Fe(I)Fe(I) complexes has been proposed to proceed \textit{via} a ‘rotated intermediate’ (Scheme 4.6).\textsuperscript{41} Models for this rotated state have been crystallized, revealing that Fe-Fe bonding is weakened and a terminal site on one Fe center is vacant.\textsuperscript{42,43} By analogy to the results in this paper, such rotated diiron complexes should perhaps be described as Fe(II)Fe(0), thus underlining the role of latent, but highly reactive intermediates.

\textbf{Scheme 4.6.}

Overall, the work reported here shows that the geometry at Ni decisively influences the reactivity of a neighboring Fe site. For small molecule [NiFe]-H\textsubscript{2}ase mimics, it has been demonstrated that perturbation of the Ni site greatly alters the acid-base properties of these complexes. With respect to the hydrogenases themselves, the results suggest that the distorted nature of the Ni(Cys)\(_4\) site is likely an important factor in tuning basicity and redox potentials, thereby enabling their remarkable activity.

\textbf{4.10. Experimental Section}

Unless otherwise stated, chemicals were purchased from commercial sources and used as received. Chromatography was performed using SiO\(_2\) (40 - 63 \(\mu\)m, 230 - 400 mesh) as the
stationary phase. CD$_2$Cl$_2$ was distilled from CaH$_2$. The complexes $^1$H, $^{30}$FeI$_2$(CO)$_4$, Cl$_2$Ni(S,S-chiraphos),$^{44}$ (pdt)Pd(dppe) and (pdt)Pt(dppe)$^{45}$ were prepared according to the literature methods. All reactions and purifications were conducted in an MBraun glovebox equipped with a solvent purification system; the concentrations of O$_2$ and H$_2$O in the N$_2$ atmosphere were typically no higher than 2 and 0.2 ppm, respectively. Glassware used in the preparation of deuteride [2D]BF$_4$ was washed with D$_2$O and dried prior to use. Solution IR spectra were recorded on a Perkin-Elmer Spectrum 100 FTIR spectrometer. A Waters Micromass Quattro II spectrometer was used to acquire ESI-MS data for analytes in dilute CH$_2$Cl$_2$ or CD$_2$Cl$_2$ solution. Analytical data were acquired using an Exeter Analytical CE-440 elemental analyzer. Unless otherwise stated, NMR data were acquired at room temperature, with samples under an atmosphere of N$_2$. $^1$H and $^{31}$P{$^1$H} NMR spectra were recorded on a Varian VXR500 or U500 spectrometer at 500 and 202 MHz, respectively. $^2$H NMR spectra were recorded on a Varian UI 500NB spectrometer at 77 MHz. Chemical shifts (δ/ppm) are referenced to CHDCI$_2$/CH$_2$Cl$_2$ (5.32 ppm for $^1$H) and external 85% H$_3$PO$_4$ (0 ppm for $^{31}$P). Cyclic voltammetry was carried out in a single compartment glass cell using a CH Instruments CHI630D electrochemical analyzer. The working, counter and pseudo-reference electrodes were glassy C, Pt and Ag, respectively. The analyte (1 mM) and NBu$_4$PF$_6$ (100 mM) were dissolved in CH$_2$Cl$_2$, and potentials (reported here relative to internal Fe/Fe$^+$) were swept at 0.1 V s$^{-1}$. Crystallographic data were collected using a Siemens SMART diffractometer equipped with a Mo K$_\alpha$ source (λ = 0.71073 Å) and an Apex II detector. EPR spectra of complexes (~1 mM in CH$_2$Cl$_2$/PhMe, 1:1) were recorded on a Varian E-line 12” Century Series X-band CW spectrometer.

$(\text{OC})_3\text{Fe(pdt)Pd(dppe)}$ (2). At –28 °C, a mixture of FeI$_2$(CO)$_4$ (29.5 mg, 70 µmol) and (pdt)Pd(dppe) (42.8 mg, 70 µmol) were dissolved in CH$_2$Cl$_2$ (1 mL) with stirring. After 1 min,
the solution was treated with CoCp₂ (26.5 mg, 140 µmol) in CH₂Cl₂ (1 mL) and allowed to warm to room temperature, before it was concentrated to ~0.5 mL and chromatographed (~10 cm SiO₂, CH₂Cl₂ eluent). The fourth band, deep green in color, was concentrated to ~2 mL, layered with pentane (~20 mL), and allowed to stand overnight at ~28 °C. The solids that formed were isolated by filtration, washed with pentane (2 × 1 mL), and dried briefly to afford the title compound as dark green plates (9.0 mg, 12 µmol, 17%). ¹H NMR (CD₂Cl₂): δ 7.74 (m, 8H, o-Ph), 7.43 (m, 8H, m-Ph), 7.42 (m, 4H, p-Ph), 2.59 (ddd, ²J_HH = 13.5 Hz, ³J_HH = 5.8 Hz, ³J_HH = 3.3 Hz, 2H, equatorial CH₂CH₂CH₂), 2.35 (m, 2H, PCH₂CH₂P), 2.22 (m, 2H, PCH₂CH₂P), 1.94 (m, 2H, axial CH₂CH₂CH₂), 1.29 (m, 2H, CH₂CH₂CH₂). ³¹P{¹H} NMR (CD₂Cl₂): δ 51.3. ESI-MS: m/z 749.8 [M]⁺, 721.8 [M – CO]⁺, 693.8 [M – 2CO]⁺ (ionization induced by addition of Fc[B(C₆H₃-3,5-(CF₃)₂)₄]). Anal. Calcd for C₃₂H₃₀O₃S₂P₂NiPd: C, 50.36; H, 3.99; N, 0.00. Found: C, 50.35; H, 3.57; N, 0.00. Green plates of 2 formed upon slow diffusion of pentane layered onto a concentrated CH₂Cl₂ solution of the title compound at ~28 °C. One crystal (0.377 × 0.232 × 0.058 mm³) was subjected to X-ray diffraction at 168 K. Its space group was determined to be triclinic P-1 with cell parameters: a 8.900 Å, b 12.883 Å, c 14.778 Å, α 104.64°, β 106.73°, γ 90.98°. Integration of 6798 reflections and solution by direct methods using SHELXTL V6.12 afforded a model with R_I = 0.0285 and wR² = 0.0684.

[(OC)₃Fe(pdt)Pd(dppe)]BF₄ ([2]BF₄). A stirred solution of 2 (11.3 mg, 15 µmol) in CH₂Cl₂ (1 mL) was treated with FcBF₄ (4.1 mg, 15 µmol) in CH₂Cl₂ (1 mL). After 1 min, pentane (20 mL) was added and the mixture allowed to stand overnight at ~28 °C. The solids were isolated by filtration, washed with pentane (2 × 1 mL), and dried briefly to afford the title compound as an olive-brown powder (3.5 mg, 4.2 µmol, 28%). ESI-MS: m/z 749.8 [M – BF₄]⁺,
[(OC)$_3$FeH(pdt)Pd(dppe)]BF$_4$ ([$2$H]BF$_4$). A solution of 2 (22.5 mg, 30 µmol) in CH$_2$Cl$_2$ (1 mL) was treated with HBF$_4$ (54% Et$_2$O solution, 14.6 mg, 90 µmol) in CH$_2$Cl$_2$ (1 mL). The solution was evaporated to dryness, and the oily residue triturated with Et$_2$O (3 mL). The resulting solid was isolated by filtration, washed with Et$_2$O (2 × 1 mL), and dried briefly to afford the title compound as an orange powder (21.1 mg, 25 µmol, 84%). $^1$H NMR (CD$_2$Cl$_2$): δ 7.70 (m, 8H, o-Ph), 7.64 (t, $^2$$J_{HH} = 6.7$ Hz, 4H, p-Ph), 7.57 (m, 8H, m-Ph), 2.87 (m, 4H, PCH$_2$CH$_2$P), 2.63 (m, 1H, equatorial CH$_2$CH$_2$CH$_2$), 2.30 (m, 4H, CH$_2$CH$_2$CH$_2$), 1.29 (m, 1H, axial CH$_2$CH$_2$CH$_2$), –4.3 (s, 1H, FeH). $^{31}$P{$^1$H} NMR (CD$_2$Cl$_2$): δ 62.8. ESI-MS: $m$/z 750.8 [M – BF$_4$]$.^+$ Orange blocks of [2$H$]BF$_4$-THF-0.5Et$_2$O formed upon slow diffusion of Et$_2$O vapor into a THF solution of the title compound at –28 °C. One crystal (0.377 × 0.232 × 0.058 mm$^3$) was subjected to X-ray diffraction at 178 K. Its space group was determined to be monoclinic $P2_1/n$ with cell parameters: $a$ 11.466 Å, $b$ 20.692 Å, $c$ 34.720 Å, $\alpha$ 90.00°, $\beta$ 97.41°, $\gamma$ 90.00°. Integration of 9965 reflections and solution by direct methods using SHELXTL V6.12 afforded a model with $R1 = 0.0369$ and $wR2 = 0.0754$.

[(OC)$_3$FeD(pdt)Pd(dppe)]BF$_4$ ([2$D$]BF$_4$). A solution of 2 (22.5 mg, 30 µmol) in CH$_2$Cl$_2$ (1 mL) and CD$_3$OD (0.1 mL) was treated with HBF$_4$ (54% Et$_2$O solution, 14.6 mg, 90 µmol) in CD$_3$OD (0.3 mL). The solution was evaporated to dryness, and the oily residue triturated with Et$_2$O (3 mL). The resulting solid was isolated by filtration, washed with Et$_2$O (2 × 1 mL), and dried briefly to afford the title compound as an orange powder (19.6 mg, 23 µmol, 78%). $^2$H NMR (CH$_2$Cl$_2$): δ –4.3 (m). $^{31}$P{$^1$H} NMR (CH$_2$Cl$_2$): δ 64.3. ESI-MS: $m$/z 751.9
[M – BF$_4$]$^+$. Anal. Calcd for C$_{32}$H$_{30}$DFeO$_2$P$_2$PdS$_2$BF$_4$: 0.25CH$_2$Cl$_2$: C, 44.99; H, 3.69; N, 0.00. Found: C, 45.03; H, 3.82; N, 0.00.

**[Cy$_3$P(OC)$_2$Fe(pdt)Pd(dppe)]BF$_4$ ([2′][BF$_4$]).** Compound 2 (15.0 mg, 20 µmol) and FeBF$_4$ (5.4 mg, 20 µmol) were dissolved in CH$_2$Cl$_2$ (1 mL) with rapid stirring. After 1 min the solution was added dropwise to PCy$_3$ (28.0 mg, 100 µmol) in CH$_2$Cl$_2$ (0.5 mL). The solution was stirred for a further 0.5 min and pentane (–28 °C, 15 mL) was added and the mixture allowed to stand overnight at –28°C. The solids were isolated by filtration, washed with pentane (–28 °C, 2 × 2 mL) and dried briefly to afford the title compound as a yellow-brown powder (5.9 mg, 5.4 µmol, 27%). ESI-MS: m/z 1002.4 [M – BF$_4$]$^+$. 974.3 [M – CO – BF$_4$]$^+$. Anal. Calcd for C$_{49}$H$_{63}$FeO$_2$P$_3$PdS$_2$BF$_4$: C, 53.23; H, 5.76; N, 0.00. Found: C, 53.16; H, 5.82; N, 0.20.

**(OC)$_3$Fe(pdt)Pt(dppe) (3).** A mixture of solid FeI$_2$(CO)$_4$ (25.3 mg, 60 µmol) and (pdt)Pt(dppe) (42.0 mg, 60 µmol) were cooled to –28 °C and dissolved in CH$_2$Cl$_2$ (2 mL) with stirring. After 1 min, the solution was treated with CoCp$_2$ (22.7 mg, 120 µmol) in CH$_2$Cl$_2$ (1 mL) and allowed to warm to room temperature before it was concentrated to ~0.5 mL and chromatographed (~10 cm SiO$_2$, CH$_2$Cl$_2$ eluent). The olive-brown band (following the rapidly eluting orange band containing Fe$_2$(pdt)(CO)$_6$) was concentrated to ~2 mL, layered with pentane (~20 mL), and allowed to stand overnight at –28 °C. The supernatant was removed, the dark residue triturated with pentane (2 × 5 mL), and dried briefly to afford the title compound as a brown powder (17.0 mg, 20.2 µmol, 34%). $^1$H NMR (CD$_2$Cl$_2$): δ 7.86 (m, 8H, o-Ph), 7.48 (m, 8H, m-Ph), 7.46 (m, 4H, p-Ph), 2.57 (m, 2H, CH$_2$CH$_2$CH$_2$), 2.22 (m, 4H, PCH$_2$CH$_2$P), 1.58 (m, 2H, CH$_2$CH$_2$CH$_2$), 1.29 (m, 2H, CH$_2$CH$_2$CH$_2$). $^{31}$P{$^1$H} NMR (CD$_2$Cl$_2$): δ 44.1 (s, d, $^1$J$_{PPt}$ = 3280 Hz). ESI-MS: m/z 838.9 [M]$^+$ (ionization induced by addition of Fc[B(C$_6$H$_3$-3,5-(CF$_3$)$_2$)$_4$]).
Anal. Calcd for C₃₂H₃₆FeO₃P₂PtS₂: C, 45.78; H, 3.60; N, 0.00. Found: C, 46.36; H, 3.22; N, 0.05.

\[(OC)₃Fe(pdt)Pt(dppe)]BF₄ ([3]BF₄).\] This compound was prepared analogously to [2]BF₄, instead using 3 as the precursor. Yield: 81%, yellow-green powder.

ESI-MS: \(m/z\) 838.9 \([M – BF₄]⁺\).

\[(OC)₃FeH(pdt)Pt(dppe)]BF₄ ([3H]BF₄).\] This compound was prepared analogously to [2H]BF₄, instead using 3 as the precursor. Yield: 69%, orange powder. \(^1\)H NMR (CD₂Cl₂): \(\delta\) 7.70 (m, 8H, \(o\)-Ph), 7.64 – 7.54 (m, 12H, \(m, p\)-Ph), 2.88 (m, 2H, PCH₂CH₂P), 2.7 (m, 2H, PCH₂CH₂P), 2.59 (m, 1H, equatorial CH₂CH₂CH₂), 2.43 (m, 4H, CH₂CH₂CH₂), 1.29 (m, 1H, axial CH₂CH₂CH₂), –3.57 (s, d, \(^1\)\(J\)_{HPt} = 173 Hz, 1H, FeH). \(^3\)P\text{\{\(^1\)H\}} NMR (CD₂Cl₂): \(\delta\) 46.4 (s, d, \(^1\)\(J\)_{PPt} = 3220 Hz). ESI-MS: \(m/z\) 839.9 \([M – BF₄]⁺\). Anal. Calcd for C₃₂H₃₆FeO₃P₂PtS₂BF₄ × 0.5CH₂Cl₂: C, 40.25; H, 3.33; N, 0.00. Found: C, 40.10; H, 2.93; N, 0.04.

\[Cy₃P(OC)₂Fe(pdt)Pt(dppe)]BF₄ ([3\ʹ]BF₄).\] This compound was prepared analogously to [2\ʹ]BF₄, instead using 3 as the precursor. Yield: 72%, yellow powder. ESI-MS: \(m/z\) 1091.7 \([M – BF₄⁻]⁺\). Anal. Calcd for C₄₉H₆₃FeO₂P₂PtS₂BF₄·0.75CH₂Cl₂: C, 48.09; H, 5.23; N, 0.00. Found: C, 48.06; H, 5.42; N, 0.52.

\((pdt)Ni(S,S\text{-chiraphos})\). This compound was prepared analogously to (pdt)Ni(dppe), instead using Cl₂Ni(S,S-chiraphos) as the precursor.\(^{44}\) Yield: 95%, orange powder. \(^1\)H NMR (CD₂Cl₂): \(\delta\) 8.15 (m, 4H, \(p\)-Ph), 7.63 – 7.47 (m, 16H, \(o,m\)-Ph), 2.19 (m, 4H, CH₂CH₂CH₂), 2.03 (m, 2H, CH), 1.81 (qu, \(^3\)\(J\)_{HH} = 6.3 Hz, 4H, CH₂CH₂CH₂), 0.96 (dd, \(^3\)\(J\)_{PH} = 10.6 Hz, \(^3\)\(J\)_{HH} = 6.2 Hz, 6H, CH₃). \(^3\)P\text{\{\(^1\)H\}} NMR (CD₂Cl₂): \(\delta\) 56.1. ESI-MS: \(m/z\) 590.4 \([M]⁺\). Anal. Calcd for C₃₁H₃₄S₂P₂Ni: C, 62.96; H, 5.79; N, 0.00. Found: C, 62.22; H, 5.69; N, 0.41.
(CO)$_3$Fe(pdt)Ni(S,S-chiraphos). The complexes (pdt)Ni(S,S-chiraphos) (118.3 mg, 200 µmol) and Fe$_2$(CO)$_9$ (72.8 mg, 200 µmol) were dissolved in CH$_2$Cl$_2$ (3 mL) with stirring. After 18 h, the deep green solution was concentrated to ~0.5 mL and chromatographed (~10 cm SiO$_2$, CH$_2$Cl$_2$ eluent). An olive-green band was collected, treated with MeCN (~5 mL), concentrated to ~1 mL and cooled to −28°C. The solids were isolated by filtration, washed with MeCN (−28°C, 2 mL) and pentane (−28°C, 2 × 2 mL), and dissolved in CH$_2$Cl$_2$ (1 mL) and MeCN (5 mL). The solution was concentrated to ~1 mL and cooled to −28°C. The crystals that formed were isolated by filtration, washed with MeCN (−28°C, 2 mL) and pentane (−28°C, 2 × 2 mL) and dried to afford the title complex as dark green crystals (46.9 mg, 64.1 µmol, 32%). $^1$H NMR (CD$_2$Cl$_2$): δ 8.11 (m, 4H, Ph), 7.62 (m, 4H, Ph), 7.52 – 7.39 (m, 12H, Ph), 2.57 (ddd, $^2$J$_{HH}$ = 13.4 Hz, $^3$J$_{HH}$ = 7.1 Hz, $^3$J$_{HH}$ = 7.1 Hz, 1H, equatorial CH$_2$CH$_2$CH$_2$), 2.17 (ddd, $^2$J$_{HH}$ = 13.2 Hz, $^3$J$_{HH}$ = 7.3 Hz, $^3$J$_{HH}$ = 7.1 Hz, 1H, equatorial CH$_2$CH$_2$CH$_2$), 1.98 (m, 2H, CH), 1.91 (m, 1H, equatorial CH$_2$CH$_2$CH$_2$), 1.82 (m, 2H, axial CH$_2$CH$_2$CH$_2$), 1.31 (m, 1H, axial CH$_2$CH$_2$CH$_2$), 0.87 (m, 6H, CH$_3$). $^{31}$P{$^1$H} NMR (CD$_2$Cl$_2$): δ 63.8. $^{31}$P{$^1$H} NMR (CD$_2$Cl$_2$, −90°C): δ 75.0, 52.6. FTIR (CH$_2$Cl$_2$): ν$_{CO}$ = 2026, 1953 cm$^{-1}$. ESI-MS: m/z 729.3 [M]$^+$ (ionization induced by addition of Fe[B(C$_6$H$_3$-3,5-(CF$_3$)$_2$)$_4$]). Anal. Calcd for C$_{34}$H$_{34}$O$_3$S$_2$P$_2$NiFe: C, 55.84; H, 4.69; N, 0.00. Found: C, 55.15; H, 4.53; N, 0.57.

[(CO)$_3$FeH(pdt)Ni(S,S-chiraphos)]BF$_4$. This compound was prepared analogously to [2H]BF$_4$, instead using (CO)$_3$Fe(pdt)Ni(S,S-chiraphos) as the precursor. Yield: 82%, orange-brown powder. $^1$H NMR (CD$_2$Cl$_2$): δ 7.89 (dd, $^3$J$_{HH}$ = 10.8 Hz, $^3$J$_{HH}$ = 8.0 Hz, 4H, p-Ph), 7.80 – 7.57 (m, 16H, o,m-Ph), 2.60 (m, 2H, CH), 2.37 (ddd, $^2$J$_{HH}$ = 14.4 Hz, $^3$J$_{HH}$ = 2.9 Hz, $^3$J$_{HH}$ = 2.9 Hz, 1H, equatorial CH$_2$CH$_2$CH$_2$), 2.29 (dm, $^2$J$_{HH}$ = 13.8 Hz, 1H, equatorial CH$_2$CH$_2$CH$_2$), 2.15 (dm, $^2$J$_{HH}$ = 13.7 Hz, 1H, equatorial CH$_2$CH$_2$CH$_2$), 1.91 (m, 1H, axial CH$_2$CH$_2$CH$_2$), 1.82 (ddd,
$^2J_{\text{HH}} = 13.7 \text{ Hz, } ^3J_{\text{HH}} = 13.7 \text{ Hz, } ^3J_{\text{HH}} = 2.0 \text{ Hz, 1H, axial } \text{CH}_2\text{CH}_2\text{CH}_2$, 1.30 (m, 1H, axial CH$_2$CH$_2$CH$_2$), 0.87 (m, 6H, CH$_3$)–3.57 (t, $^2J_{\text{HP}} = 4.8 \text{ Hz, 1H, FeH}$). $^{31}\text{P}^1\text{H}$ NMR (CD$_2$Cl$_2$): $\delta$ 71.7, 67.9. $^{31}\text{P}^1\text{H}$ NMR (CD$_2$Cl$_2$, –70°C): $\delta$ 72.6 (d, $^2J_{\text{PP}} = 67 \text{ Hz, } 67.2$ (d, $^2J_{\text{PP}} = 67 \text{ Hz}$).

FTIR (CH$_2$Cl$_2$): $\nu_{\text{CO}} = 2082, 2022 \text{ cm}^{-1}$. ESI-MS: $m/z$ 730.4 [M – BF$_4^-$]$.^+$ Anal. Calcd for C$_{34}$H$_{35}$O$_3$S$_2$P$_2$FeNiBF$_4$$\times$0.5CH$_2$Cl$_2$: C, 48.10; H, 4.21; N, 0.00. Found: C, 48.12; H, 4.11; N, 0.40.

4.11. Calculations

DFT calculations were performed using the B3P86 density functional with the SDD pseudopotential and basis set$^{46}$ for the Fe, Ni, Pd, and Pt atoms, the 6-31G** basis set$^{47}$ for $\mu$-H ligands, and the 6-31G* basis set$^{48,49}$ for all other atoms. The starting geometries for 1,19 [1H]$^+$,30 [2H]$^+$ were obtained from crystal structures. The starting coordinates for atoms in 3 and [3H]$^+$ were obtained from the crystal structure of [2H]$^+$, where the metal center and hydride were manually altered prior to optimization. Solvation effects were included using the C-PCM (conductor-like polarizable continuum model)$^{50,51}$ with Bondi atomic radii$^{52}$ and included the nonelectrostatic contributions of dispersion,$^{53,54}$ repulsion,$^{53,54}$ and cavitation energies.$^{55}$

For the results presented in the main paper, the geometry optimizations were performed in the gas phase. Geometry optimizations were also performed in solution and were found to be consistent with the gas phase optimizations; these results are provided in the Supporting Information. In all cases, the minimum-energy structures were confirmed to have no imaginary frequencies. The $\nu_{\text{CO}}$ values were calculated with DFT and were scaled by the standard factor of 0.9850.$^{56}$ The transition states (TSs) for the tetrahedral to square planar isomerization were identified using the Synchronous Transit-Guided Quasi-Newton method,$^{57,58}$ and the resulting structures were confirmed to have only a single imaginary frequency. Each TS was verified to lead to the relevant tetrahedral and square planar geometries by following the intrinsic reaction
coordinate (IRC) using the local quadratic approximation\textsuperscript{59,60} for 4 or 5 steps in both directions and subsequently optimizing the geometries. Because the potential energy surface was found to be very flat, the single imaginary frequencies were small, and the complete IRC connecting the two isomers through the TS could not be obtained.

Thermochemical data were calculated at $T = 298.15$ K. The reaction free energies ($\Delta G^\circ$) and free energy barriers ($\Delta G^\ddagger$) associated with the tetrahedral to square planar isomerization in solution included zero-point energy, entropic contributions, and solvation effects. The relative reduction potentials and $pK_a$’s were calculated from the corresponding reaction free energies using methodology described elsewhere.\textsuperscript{61-63} Chemical bonding analysis was performed using Natural Bond Orbitals (NBO).\textsuperscript{33} All calculations were performed using the Gaussian 09 electronic structure program.\textsuperscript{64} Structures and energies of the systems studied herein are provided in the Supporting Information.

4.12. References


(33) Glendening, E. D.; Reed, A. E.; Carpenter, J. E.; Weinhold, F.


Chapter 5. H₂ Evolution Mechanism for [NiFe]-Hydrogenase Models*†

5.1. Abstract

The intermediacy of a reduced nickel-iron hydride in hydrogen evolution catalyzed by Ni-Fe complexes was verified experimentally and computationally. In addition to catalyzing hydrogen evolution, the highly basic and bulky (dppv)Ni(μ-pdt)Fe(dppv)(CO) ([I]⁰, dppv = cis-C₂H₅(PPh₂)₂) and its hydride derivatives have yielded to detailed characterization in terms of spectroscopy, bonding, and reactivity. The protonation of [I]⁰ initially produces unsym-[H₁]⁺, which converts by a first order pathway to sym-[H₁]⁺. These species have $C_1$ (unsym) and $C_s$ (sym) symmetry, respectively, depending on the stereochemistry of the octahedral Fe site. Both experimental and computational studies show that [H₁]⁺ protonates at sulfur. The $S = \frac{1}{2}$ hydride [H₁]⁰ was generated by reduction of [H₁]⁺ with Cp*₂Co. DFT calculations indicate that [H₁]⁰ is best described as a Ni(I)-Fe(II) derivative with significant spin density on Ni and some delocalization on S and Fe. EPR spectroscopy reveals both kinetic and thermodynamic isomers of [H₁]⁰. Whereas [H₁]⁺ does not evolve H₂ upon protonation, treatment of [H₁]⁰ with acids gives H₂. The redox state of the "remote" metal (Ni) modulates the hydridic character of the Fe(II)-H center. As supported by DFT calculations, H₂ evolution proceeds either directly from [H₁]⁰ and external acid, or from protonation of the Fe-H bond in [H₁]⁰ to give a labile dihydrogen complex. Stoichiometric tests indicate that protonation-induced hydrogen evolution from [H₁]⁰ initially produces [I]⁺, which is reduced by [H₁]⁰. Our results reconcile the required reductive activation

---

† Author contributions: Dr. Olbelina A. Ulloa, Dr. Casseday P. Richers, Dr. Jeffery A. Bertke, Dr. Mark J. Nilges, and Prof. Thomas Rauchfuss were responsible for all experiments and associated analyses in this chapter.
of a metal hydride and the resistance of metal hydrides toward reduction. This dichotomy is resolved by reduction of the remote (non-hydride) metal of the bimetallic unit.

5.2. Introduction

The crystallographic characterization of the active sites of [NiFe]- and [FeFe]-hydrogenase (H$_2$ases) sparked the synthesis and study of many low molecular weight models. These models replicate some aspects of the biological reactions,\textsuperscript{1} shedding light on how these enzymes operate and perhaps opening avenues for bioinspired if not biomimetic reactivity of interest in the energy sector.

Mechanistic attention is particularly focused on hydride-containing models since hydride is the substrate both in H$_2$ oxidation and proton reduction. Only a small number of hydride-containing containing models for the [NiFe]-enzymes have been designed.\textsuperscript{2-4} A versatile family of models are of the type [(diphos)Ni(μ-pdt)(μ-H)Fe(PR$_3$)$_x$(CO)$_{3-x}$]$^+$ (pdt$^{2-}$ = S$_2$C$_3$H$_6$$^{2-}$; diphos = dppe = 1,2-$C_2$H$_4$(PPh$_2$)$_2$, dppb = 1,2-$C_6$H$_4$(PPh$_2$)$_2$, dppv = cis-1,2-$C_2$H$_2$(PPh$_2$)$_2$). In the case of the Fe(CO)$_3$ and Fe(CO)$_2$(PR$_3$) derivatives, the Ni-Fe-hydrides are notable because they catalyze proton reduction,\textsuperscript{2,3,5,6} as well as, in one case,\textsuperscript{7} hydrogen oxidation.\textsuperscript{8-10}

**Scheme 5.1.** Active site states of [NiFe]-hydrogenases. Only unidirectional conversions are shown for simplicity.
Hydrogen evolution catalyzed both by the enzymes and by the models operates by the intermediacy of Ni(II)-H-Fe(II) species.\textsuperscript{11-13} In the enzyme, this species, the Ni-R state, releases H\textsubscript{2} via protonolysis to give an unsaturated Ni(II)-Fe(II) species (Ni-SI\textsubscript{a}) (Scheme 5.1). The intimate details of H\textsubscript{2} release in the enzyme remain an area of active investigation,\textsuperscript{14,15} one complicating aspect being the protonation state of the terminal thiolate cofactors. The Ni(II)-H-Fe(II) models do not release H\textsubscript{2} upon treatment with proton donors. Instead, hydrogen evolution occurs only upon reduction.\textsuperscript{3,5,16,17} A similar pattern is often observed with Fe(II)-H-Fe(II) (bridging hydride\textsuperscript{17}) and Fe(II)-Fe(II)-H (terminal hydride\textsuperscript{18}) complexes, which also release H\textsubscript{2} only upon reduction.\textsuperscript{19} For the [FeFe]-hydrogenase models, the ET-induced protonation of an Fe(II)-Fe(II)-H species is explained by the involvement of a transient H\textsubscript{red} state featuring an appended [4Fe-4S]\textsuperscript{+} subunit.\textsuperscript{11} Since hydrogen evolution from Ni-H-Fe complexes commences upon reduction, reduced Ni-H-Fe hydrides represent important targets for characterization. Electrochemical reduction of the [(diphos)Ni(\textmu-pdt)(\textmu-H)Fe(PR\textsubscript{3})\textsubscript{x}(CO)\textsubscript{3-x}]\textsuperscript{+} species leads to unstable products, however, probably owing to bimolecular degradation pathways.\textsuperscript{3} Such bimolecular processes should be suppressed in sterically bulky Ni-Fe hydrides. For these reasons, our attention turned to [(dppv)Ni(\textmu-pdt)(\textmu-H)Fe(CO)(dppv)]\textsuperscript{z}, a new family of [NiFe]-hydrogenase models related to the [(dppv)(CO)Fe(\textmu-pdt)(\textmu-H)Fe(CO)(dppv)] platform that has illuminated aspects of the [FeFe]-hydrogenase chemistry.\textsuperscript{17,20,21} The dppv ligand confers stability to these models and minimizes side reactions that can plague more flexible chelating diphosphines.\textsuperscript{22-24} Starting with the preparation and protonation of [(dppv)Ni(\textmu-pdt)(\textmu-H)Fe(CO)(dppv)]\textsuperscript{0}, this report describes previously unseen insights related to hydrogen evolution catalysis. Of particular interest is the observation of the reduced hydride, which is the species that is protonated,
thereby liberating H₂. Our results reconcile the required reductive activation of a metal hydride and the resistance of metal hydrides toward reduction.²⁵,²⁶ This dichotomy is shown to be resolved when ET occurs at the remote (non-hydride) metal of the bimetallic unit.

5.3. Synthesis of [(dppv)Ni(µ-pdt)Fe(CO)(dppv)]⁰ ([1]⁰)

Following precedent for related reactions,² the synthesis of [1]⁰ began with the reaction of Fe(pdt)(dppv)(CO)₂, NiCl₂(dppv), and NaBF₄ to give the Ni(II)-Fe(II) species [(dppv)Ni(µ-pdt)(µ-Cl)Fe(CO)(dppv)]BF₄ ([Cl1]BF₄). Although Fe(pdt)(dppv)(CO)₂ consists of a 1:4 mixture of all-cis and trans isomers, the initially formed [Cl1]⁺ consists only of the unsymmetrical (unsym) isomer wherein the dppv on Fe spans apical and basal coordination sites. Reduction of a THF solution of unsym-[Cl1]BF₄ with KFe(CO)₂Cp gave [1]⁰ (Scheme 5.2).

At room temperature, the ³¹P{¹H} NMR spectrum of [1]⁰ in C₆D₆ solution is simple: signals assigned to the Fe(dppv) center appear as a triplet (δ 89.2), and those assigned to the Ni(dppv) center appear as two doublets (δ 74.2 and 46.2). These assignments are consistent with the ³¹P{¹H} NMR spectrum of the ⁵⁷Fe derivative, which revealed broadening of the signal assigned to the Fe(dppv)(CO) center (Figure S19). The chemical shifts of the signals assigned to the Ni(dppv) group are similar to those for [(dppe)Ni(µ-pdt)Fe(CO)₃]⁰ (δ 81 and 46, at −68 °C).³ The Ni(dppe) center of [(dppe)Ni(µ-pdt)Fe(CO)₃]⁰ is dynamic on the NMR timescale, such that only a single ³¹P signal (δ 63.6) appears at 20 °C.⁶ In contrast, two ³¹P signals assigned to the Ni(dppv) site in [1]⁰ are observed even up to 90 °C, indicating a barrier to rotation at Ni of >14 kcal/mol (Figure S23). Upon cooling a toluene solution of [1]⁰ below −65 °C, four signals appear in the ³¹P{¹H} NMR spectrum. This decoalescence reveals the full complexity of this chiral species.
Crystallographic analysis of [1]$^{0}$ confirms the unsymmetrical structure suggested by the NMR data (Figure 5.1). The Fe-Ni bond of 2.5261(4) Å is 0.06 Å longer than that for [(dppe)Ni($\mu$-pdt)Fe(CO)$_3$]$^0$ (2.466(6) Å), an elongation attributed to steric interactions between the two bulky M(dppv) centers. The iron center of [1]$^{0}$ adopts a distorted octahedral geometry, viewing the Fe-Ni bond as a sixth ligand. The nickel center adopts a distorted trigonal bipyramidal geometry, viewing the Fe-Ni bond as a fifth ligand, in addition to two phosphines and two thiolates. The Fe$_1$-Ni$_1$-P$_2$ angle is 162.2° and the dihedral angle is 80.94° for the S$_1$-Ni-S$_2$ and P$_1$-Ni-P$_2$ planes. For comparison, the corresponding values for [(dppe)Ni($\mu$-pdt)Fe(CO)$_3$]$^0$ are 169.8° and 84.9°. The Ni-P distances are comparable to those in Ni(dppv)$_2$ (2.1414(6) - 2.1586(6) Å) as well as [(dppe)Ni($\mu$-pdt)Fe(CO)$_3$]$^0$ (2.159 - 2.146 Å).
**Figure 5.1.** Molecular structure of the non-hydrogen atoms of [1]$^0$ showing 50% probability ellipsoids.

### 5.4. DFT-Optimized Structure of [1]$^0$

The DFT-optimized structure of [1]$^0$ is in good agreement with the crystal structure. Similar to related dithiolate compounds,$^{6,24}$ [1]$^0$ also contains a Ni-Fe σ-bond that can be localized using either the Natural Bond Orbital (NBO)$^{27}$ or the Pipek-Mezey criterion (Figure S61)$^{28}$.

The $^{31}$P NMR resonance at $\delta$ 89 is assigned to the Fe(dppv)(CO) center. This signal decoalesces at low temperatures in toluene, indicating a rotational barrier of ~9 kcal/mol, in good agreement with the DFT-calculated free energy barrier of $\Delta G^\ddagger = 9.55$ kcal/mol in toluene. Interchange of the PFe sites could conceivably occur via an intermediate wherein the dppv occupies dibasal sites (i.e., sites that are trans to the dithiolate) and the CO ligand occupies the apical site trans to nickel. According to our calculations, the formation of this dibasal intermediate is thermodynamically accessible ($\Delta G^\circ = 5.26$ kcal/mol) but is subject to a high free energy barrier of 27 kcal/mol in toluene. A lower free energy pathway is shown in Scheme 5.3 with a transition state (TS) in which the dppv ligand occupies sites cis to the dithiolate. In the TS structure, the Ni-Fe bond remains ~2.52 Å. A similar pathway has been invoked for Fe$_2$(µ-pdt)(CO)$_4$(diphos) and related complexes.$^{29,30}$
Scheme 5.3. DFT-calculated low-energy pathway for degenerate isomerization of $[1]^0$. Notice that the dibasal isomer is not an intermediate in this process (Ph-P groups omitted for clarity).

The calculations do not reveal a minimum with a square planar Ni center. In contrast, previous calculations on [(dppe)Ni(μ-pdt)Fe(CO)$_3$]$^0$ showed that the valence isomers, with either square planar Ni(II) or tetrahedral Ni(I) centers, are nearly isoergic. Formation of square planar Ni(II) requires an intramolecular electron transfer to convert Fe(I) to Fe(0). The [(dppe)Ni(μ-pdt)Fe(CO)$_3$]$^{+0}$ couple (−0.59 V) is ~300 mV milder than the $[1]^{+0}$ couple, indicating the greater reluctance for the (RS)$_2$Fe(dppe)(CO) vs (RS)$_2$Fe(CO)$_3$ site to undergo reduction from Fe(I) to Fe(0) in $[1]^0$.

The IR spectrum of $[1]^0$ in THF solution shows a dominant ν$_{CO}$ band at 1897 cm$^{-1}$ with a shoulder at ~1912 cm$^{-1}$. These signals probably arise from two conformers ("flippamers"), resulting from the relative orientation of the propanedithiolate backbone. DFT predicts 1905 and 1908 cm$^{-1}$ for the two flippamers of the unsym isomer.

5.5. Redox of $[1]^0$

Cyclic voltammetry (CV) studies on a 1,2-C$_6$H$_4$F$_2$ solution of $[1]^0$ revealed a single redox couple assigned to $[1]^{+0}$ at −1.13 V vs Fe$^{0/+}$. The couple is quasi-reversible on the basis of the current ratio, $i_{pa}/i_{pc} \sim 1.16$, and a fairly linear dependence of $i$ vs (scan rate)$^{1/2}$ (Figures S37). Reflecting the highly reactive nature of $[1]^+$, the CV of a PhCN solution was very different, involving two oxidation and two reductive processes. The latter processes are almost certainly the result of solvent coordination.
5.6. DFT Characterization of [1]$^+$

Using DFT, two valence isomers were characterized for [1]$^+$. These two isomers differ in oxidation states for the two metals: Ni(I)Fe(II) vs Ni(II)Fe(I). The calculations indicate that these two species also differ with respect to the geometry at Ni (Figure 5.2). Moreover, the Fe-Ni bond is disrupted (2.91 Å) in Ni(II)Fe(I) compared to the Ni(I)Fe(II) valence isomer (2.52 Å). The sym isomer of [1]$^+$ was also characterized using DFT, and is qualitatively similar to the unsym isomer presented here.

![Figure 5.2. Spin density plots (isovalue = 0.005) for the Ni(I)Fe(II) (left) and Ni(II)Fe(I) (right) valence isomers of unsym-[1]$^+$.](image)

The spin densities for these two valence isomers are given in Table 5.1. The calculated reaction free energy is small ($\Delta G^\circ = -1.36$ kcal/mol) for the isomerization from Ni(II)Fe(I) to Ni(I)Fe(II), and the calculated free energy barrier for this isomerization is also relatively low ($\Delta G^\ddagger = 2.75$ kcal/mol). Thus, the calculations indicate that both isomers could be thermally accessible at room temperature but do not allow for the identification of a kinetic or thermodynamic preference. The calculate $v_{CO}$ values are 1948 and 1913 cm$^{-1}$ for the Ni(I)Fe(II) and Ni(II)Fe(I) valence isomers, respectively. Solutions of [1]$^+$ were generated oxidation of [1]$^0$ with [Fc]PF$_6$ followed by immediate analysis by FT-IR and EPR spectroscopy.
Table 5.1. Calculated Spin Densities for the Valence Isomers of [1]$^{+a}$

<table>
<thead>
<tr>
<th>Valence Isomer</th>
<th>$\rho$ (Ni)</th>
<th>$\rho$ (Fe)</th>
<th>$\rho$ (2S)</th>
<th>$\rho$ (2P$_{Ni}$)$^b$</th>
<th>$\rho$ (2P$_{Fe}$)$^c$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ni(II)Fe(I)</td>
<td>0.05</td>
<td>0.87</td>
<td>-0.02</td>
<td>0.01</td>
<td>0.05</td>
</tr>
<tr>
<td>Ni(II)Fe(I)</td>
<td>0.05</td>
<td>0.90</td>
<td>-0.02</td>
<td>0.01</td>
<td>0.05</td>
</tr>
<tr>
<td>Ni(I)Fe(II)</td>
<td>0.59</td>
<td>0.04</td>
<td>0.33</td>
<td>0.03</td>
<td>-0.02</td>
</tr>
<tr>
<td>Ni(I)Fe(II)</td>
<td>0.64</td>
<td>-0.05</td>
<td>0.32</td>
<td>0.05</td>
<td>0.00</td>
</tr>
</tbody>
</table>

$^a$For each entry, the first and second entries correspond to the unsym and sym isomers, respectively. $^b$These values refer to the P atoms of the dppv ligand on the Ni center. $^c$These values refer to the P atoms of the dppv ligand on the Fe center.

5.7. Protonation of [1]$^0$

Protonation of [1]$^0$ with HOTf or H(OEt$_2$)$_2$BAR$_4$ afforded a single unsymmetrical isomer of [H1]$^+$ (Scheme 5.2). Three $^{31}$P{$^1$H} NMR signals (intensities: 1:1:2) for unsym-[H1]$^+$ appear at $\delta$ 93.4, 79.4, and 73.5, respectively. The signals at $\delta$ 93.4 and 79.4 are assigned to the Fe(dppv) group, and the signal at $\delta$ 73.5 is assigned to the Ni(dppv) group. Upon cooling the toluene solution of [H1]$^+$ to −40 °C, the Ni(dppv) signal decoalesces into two broad singlets (Figure S33). These observations are consistent with the Fe center being more rigid than the Ni site (Scheme 5.4).

Scheme 5.4. Variable temperature NMR measurements establish that the Ni(II) site is nonrigid in unsym-[H1]$^+$. (Ph-P groups were omitted for clarity.)

Over the course of a few hours in MeCN solution at 25 °C, unsym-[H1]$^+$ partially converts to sym-[H1]$^+$ (Scheme 5.2). The conversion follows first-order kinetics with a rate constant of 5.2 x 10$^{-5}$ s$^{-1}$. Based on transition state theory, this rate for conversion of unsym-[H1]$^+$ to sym-[H1]$^+$ corresponds to $\Delta G^\ddagger$ ~ 23 kcal/mol. The DFT-calculated free energy barrier for this isomerization is 24 kcal/mol. The symmetrical isomer gives rise to two singlets in the $^{31}$P{$^1$H} NMR spectrum. Consistent with its high symmetry, sym-[H1]$^+$ exhibits a $^{31}$P{$^1$H} NMR spectrum that does not
change over the range +35 to −40 °C. At equilibrium, the sym:unsym ratio is 1:5 and 1:8 in MeCN and CH$_2$Cl$_2$ solutions, respectively.

5.8. Crystallographic and DFT Analysis of the Ni(II)-H-Fe(II) Complex

The salt $\text{unsym-}[\text{H}1]\text{BAr}_4^F$ was characterized crystallographically. It features an octahedral iron center and an approximately square pyramidal nickel center (Figure 5.3). The Fe center adopts the unsym geometry, which is the predominant isomer in solution.

![Figure 5.3](image.png)

*Figure 5.3.* Molecular structure of the cation in [H1]BAr$_4^F$ with 50% probability thermal ellipsoids. This structure corresponds to the unsym isomer. Hydrogen atoms that were not refined, the counter ion, and solvent molecules have been omitted for clarity.

The structures of both the sym and unsym isomers of [H1]$^+$ were optimized using DFT, and the unsym isomer was found to be 2.35 kcal/mol lower in free energy than the sym isomer. The optimized unsym structure agrees well with the crystallographic result, as shown by the selected bond lengths and angles provided in Table 5.2. Consistent with the crystal structure of $\text{unsym-}[\text{H}1]^+$, the optimized structures of $\text{sym-}[\text{H}1]^+$ and $\text{unsym-}[\text{H}1]^+$ both feature asymmetrically bound hydrides, with the hydride closer to Fe than to Ni. The DFT-optimized structure of $\text{sym-}[\text{H}1]^+$ is provided in the SI (Table S13). Moreover, consistent with previous DFT calculations on NiFe models, the CO stretching frequencies ($\nu_{\text{CO}}$) for the optimized structures agree well with
the experimental IR data (Table 5.3). The values 1946 and 1957 cm\(^{-1}\) were calculated for the sym and unsym isomers, respectively.

Table 5.2. Selected Structural Metrics (Å, °) for [I]\(^0\), *unsym-[H1]\(^+\) (as BAr\(^F_4\) salt) and *sym-[H1]\(^0\)

<table>
<thead>
<tr>
<th></th>
<th>*unsym-[I](^0)</th>
<th>*unsym-[H1](^+)</th>
<th>*sym-[H1](^0)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Expt</td>
<td>Calc(^a)</td>
<td>Expt</td>
<td>Calc</td>
</tr>
<tr>
<td>Fe-Ni</td>
<td>2.5261(4)</td>
<td>2.523</td>
<td>2.6461(5)</td>
</tr>
<tr>
<td>Fe-H</td>
<td>-</td>
<td>-</td>
<td>1.56(3)</td>
</tr>
<tr>
<td>Ni-H</td>
<td>-</td>
<td>-</td>
<td>1.80(2)</td>
</tr>
<tr>
<td>Ni-P(_1)</td>
<td>2.1557(8)</td>
<td>2.186</td>
<td>2.1379(8)</td>
</tr>
<tr>
<td>Ni-P(_2)</td>
<td>2.1640(6)</td>
<td>2.184</td>
<td>2.1557(8)</td>
</tr>
<tr>
<td>Fe-P(_1)</td>
<td>2.2174(5)</td>
<td>2.203</td>
<td>2.2021(8)</td>
</tr>
<tr>
<td>Fe-P(_2)</td>
<td>2.1923(8)</td>
<td>2.256</td>
<td>2.2250(8)</td>
</tr>
<tr>
<td>P(_1)-Ni-P(_2)/S(_1)-Ni-S(_2) dihedral angle (°)(^b)</td>
<td>80.94</td>
<td>79.19</td>
<td>15.97</td>
</tr>
<tr>
<td>P(_1)-Ni-P(_3) bite angle (°)</td>
<td>89.54</td>
<td>90.09</td>
<td>87.50</td>
</tr>
<tr>
<td>P(_2)-Fe-P(_4) bite angle (°)</td>
<td>86.27</td>
<td>87.08</td>
<td>86.43</td>
</tr>
</tbody>
</table>

\(^a\)The omission of counterions in the calculations may introduce discrepancies between the crystal structure and DFT-optimized structure. \(^b\)Dihedral refers to the angle between the planes formed by the groups of three atoms listed.

Table 5.3. Experimental and Calculated Values for ν\(_{CO}\) (cm\(^{-1}\))

<table>
<thead>
<tr>
<th></th>
<th>ν(_{CO}) (Expt)(^a,b)</th>
<th>ν(_{CO}) (DFT)(^c)</th>
</tr>
</thead>
<tbody>
<tr>
<td>[ClI]BF(_4)</td>
<td>1946(^e)</td>
<td>n/a</td>
</tr>
<tr>
<td>[I](^0)</td>
<td>1912 (sh)(^d), 1897</td>
<td>1905</td>
</tr>
<tr>
<td>[H1](^+)</td>
<td>1953</td>
<td>1946 (sym)</td>
</tr>
<tr>
<td></td>
<td>1960 (sh)(^d)</td>
<td>1957 (unsym)</td>
</tr>
<tr>
<td>[H1H](^2+)</td>
<td>1994 (unsym)</td>
<td>1980 (unsym)</td>
</tr>
<tr>
<td>[H1](^0)</td>
<td>1920</td>
<td>1922</td>
</tr>
</tbody>
</table>

\(^a\)Experiments performed in THF solution unless otherwise noted. \(^b\)Counterions for salts are BAr\(^F_4\)– except in the cases [I]PF\(_6\) and [ClI]BF\(_4\). \(^c\)Central CH\(_2\) of the pdt is oriented toward the Ni center in all cases. The ν\(_{CO}\) values for the structures with the central CH\(_2\) of the pdt oriented toward the Fe center is provided in the SI (Table S15). These values shift by 1-4 cm\(^{-1}\) depending on the pdt conformation. \(^d\)sh = shoulder.

5.9. Acid-Base Properties of [H1]\(^+\)

The pK\(_a\) of [H1]\(^+\) was determined by treating [I]\(^0\) with varying amounts of acetic acid (pK\(_a\)\(^{MeCN}\) = 21.56) to give mixtures of [I]\(^0\) and *unsym-[H1]\(^+\). The \(^{31}\)P\{\(^1\)H\} NMR signals for [I]\(^0\) and *unsym-[H1]\(^+\) are distinct, consistent with slow proton exchange between the metal hydride and
its conjugate base.\textsuperscript{31,32} For these experiments, PhCN was used as the solvent under the assumption that $pK_a^{\text{MeCN}} \sim pK_a^{\text{PhCN}}$.\textsuperscript{33} In MeCN solution, HOAc and OAc$^-$ tend to associate, as indicated by the homoconjugation constant of 3.21.\textsuperscript{32} Our calculations corrected for this effect (Table S7). The $pK_a$ of [H1]$^+$ is estimated to be $16.6 \pm 0.04$.\textsuperscript{31} A $pK_a$ of 17.7 for [H1]$^+$ was calculated by DFT relative to the reference $pK_a$ for [(dppe)Ni(\mu-pdt)(\mu-H)Fe(CO)$_2$(PPh$_3$)]$^+$. As expected for a more electron-rich metal-hydride, the $pK_a$ of [H1]$^+$ is higher than that for related complexes substituted by only two or three phosphines. For example, $pK_a^{\text{PhCN}} = 10.7$ and 14.9 for [(dppe)Ni(\mu-pdt)(\mu-H)Fe(CO)$_3$]$^+$ and [(dppe)Ni(\mu-pdt)(\mu-H)Fe(CO)$_2$(PPh$_3$)]$^+$, respectively.\textsuperscript{3} The DFT calculated relative $pK_a$'s for these species are also in good agreement with the experimental values and are provided in the Supporting Information (Table S18).

In CH$_2$Cl$_2$ solution, [H1]$^+$ is protonated by [H(OEt)$_2$$_2$]BArF$_4$ to give a species denoted [H1H]$^{2+}$. The reaction is indicated by a color change from orange to red-brown. In the IR spectrum, $\nu_{\text{CO}}$ shifts from ~1956 cm$^{-1}$ to 1994 cm$^{-1}$ (Figure 5.4). The same protonation cannot be effected in THF, and addition of THF (100 µL) to a ~9 mM solution of [H1H]$^{2+}$ in CH$_2$Cl$_2$ gave back [H1]$^+$, attesting to the highly acidic nature of the dication. The $^{31}$P NMR spectrum of [H1H]$^{2+}$ shows four multiplets, indicating that the four phosphorus centers are nonequivalent, which indicates that (i) is relatively rigid and (ii) protonation is regiospecific. In contrast, $sym$-[H1]$^+$ is dynamic on the $^{31}$P NMR timescale owing to rotation of the square planar Ni center. Solutions of [H1H]$^{2+}$ in CD$_2$Cl$_2$ are stable for hours. The $^1$H NMR spectrum of [H1H]$^{2+}$ shows a single hydride signal at $\delta$ ~13.1. This hydride signal is shifted upfield of [H1]$^+$ by 7.2 ppm and integrates to 1 relative to the pdt$^{2-}$ ligand.
Figure 5.4. Spectroscopic changes induced by protonation of $[\text{H}1]^+$ at room temperature. Top: FTIR spectra of $[\text{H}1]^+$ (dashed) and $[\text{H1H}]^{2+}$ (solid) in CH$_2$Cl$_2$. Bottom: $^{31}$P NMR spectra of a thermally equilibrated spectrum of $[\text{H}1]^+$ (top) and a freshly prepared solution of $[\text{H1H}]^{2+}$ (bottom) in CH$_2$Cl$_2$ (weak signals at $\delta$ 88 and 68 in the upper spectrum correspond to the sym isomer).

The structure of $[\text{H1H}]^{2+}$ and its interaction with OEt$_2$ were examined by DFT. The inclusion of neutral OEt$_2$ in these calculations is relevant because the acid was $[\text{H(OEt}_2)_2]^+$. The calculations show that $\text{H}^+$ and $\text{H(OEt}_2)^+$ regiospecifically protonate one of the two diastereotopic thiolate sulfur centers. Protonation by $\text{H}^+$ breaks the Ni-S bond trans to the basal Fe-P, as indicated by elongation of the Ni–S distance from 2.25 to 3.66 Å. However, the Ni-S bond is not disrupted when the ether is included in the optimization of $[\text{H1H}]^{2+}$ (Figure 5.5). A possible explanation for the observation that ether stabilizes the bidentate over the monodentate structure is that the hydrogen bonding of ether to the SH group increases the basicity of the sulfur, allowing it to donate...
electrons more easily to the positively charged Ni center. The DFT-calculated values for $\nu_{\text{CO}}$ for the protonated species are 1980 and 1981 cm$^{-1}$ for the structures with and without OEt$_2$, respectively. The S-protonation of bridging dithiolates in diiron complexes by HBF$_4$ and HOTf, as well as the disruption of the metal-sulfur bond, has been observed previously.$^{34-36}$

![Figure 5.5](image_url)

**Figure 5.5.** DFT-calculated structure of [H1H]+$^{2+}$ with (top) and without (bottom) the interaction of the acidic proton with neutral OEt$_2$. Phenyl groups and most hydrogens were omitted for clarity. Another minimum energy structure was obtained by starting with the lower monodentate structure and adding an ether within hydrogen-bonding distance of S$_2$H; this optimized structure retained the monodentate character was ~8 kcal/mol higher in free energy than the upper bidentate structure.

5.10. **Redox Experiments on Hydrides**

Cyclic voltammetry (CV) studies on [H1]$^+$ revealed that the hydride [H1]$^+$ reversibly reduces to [H1]$^0$ (Figure 5.6) at $-1.48$ V vs Fe$^{0/+}$ in MeCN (Table 5.4). In fact, all complexes of the type $[(\text{diphosphine})\text{Ni}^{II}(\mu-\text{pdt})(\mu-H)\text{Fe}^{II}(\text{CO})_{3-x}(\text{PR}_3)_x]^+$ reduce at similar potentials regardless of the substitution at Fe, consistent with a Ni-centered reductions.$^3$ The main difference among
their redox behavior is the reversibility, which improves with steric bulk such that the current ratio, \( i_{pa}/i_{pc} \), for the \([\text{H}1]^{+/0}\) couple is 0.94 at 100 mV/s in MeCN (Table S3). The DFT-calculated potential for the \([\text{H}1]^{+/0}\) couple agrees with the experimental value (Table 5.4).

![Figure 5.6. Cyclic voltammogram of \([\text{H}1]^{+/0}\) at scan rates 50–500 mV s\(^{-1}\). Conditions: 0.1 M \([\text{Bu}_4\text{N}]\text{PF}_6\) in MeCN, 1 mM \([\text{H}1]^{+}\text{OTf}\), glassy carbon working electrode, Ag/AgCl reference electrode, and Pt wire counter electrode.](image)

<table>
<thead>
<tr>
<th>Redox Couple</th>
<th>(E_{1/2}) (expt)</th>
<th>(i_{pa}/i_{pc})</th>
<th>(E_{1/2}) (DFT)</th>
</tr>
</thead>
<tbody>
<tr>
<td>([\text{dppe}]	ext{Ni(µ-pdt)(µ-H)Fe(CO)}_2(\text{P(OPh)}_3)]^{+/0}</td>
<td>–1.44 (^b)</td>
<td>0.93 (^b)</td>
<td>–1.44</td>
</tr>
<tr>
<td>([\text{dppe}]	ext{Ni(µ-pdt)(µ-H)Fe(CO)}_2(\text{PPh}_3)]^{+/0}</td>
<td>–1.49 (^b)</td>
<td>0.06 (^b)</td>
<td>–1.53</td>
</tr>
<tr>
<td>([\text{dppe}]	ext{Ni(µ-pdt)(µ-H)Fe(CO)}_3]^{+/0}</td>
<td>–1.29 (^b)</td>
<td>0.26 (^b)</td>
<td>–1.37</td>
</tr>
<tr>
<td>([\text{dppv}]	ext{Ni(µ-pdt)(µ-H)Fe(CO)(dppv)}]^{+/0} ([\text{H}1]^{+/0})</td>
<td>–1.48 (^b)</td>
<td>0.94 (^b)</td>
<td>–1.45</td>
</tr>
</tbody>
</table>

\(^a\)1 mM complex in MeCN, 0.1 M [NBu\(_4\)]BF\(_4\) electrolyte, glassy carbon electrode, 100 mV/s scan-rate. \(^b\)This couple was used as the reference for all other calculated potentials so the experimental and calculated values agree by construction. Because the experimental reference is \(E_{1/2}\), the calculated potentials are also assumed to be \(E_{1/2}\).

5.11. **Hydrogen Evolution Catalysis**

The hydrides are catalysts for hydrogen evolution. In the presence of added TsOH (\(pK_a^{MeCN} = 8.7\)), CF\(_3\)CO\(_2\)H (\(pK_a^{MeCN} = 12.7\)),\(^{37}\) and ClCH\(_2\)CO\(_2\)H (\(pK_a^{MeCN} = 15.3\)), [\(\text{H}1\)]\(^+\) in MeCN solution shows increased currents at potentials corresponding to the reductive wave near \(-1.4\) V, clear evidence for catalysis. For example, the rate of hydrogen evolution approaches "plateau
catalytic current" at high [H\textsuperscript{+}] and scan rates. The data were instead interpreted by the foot-of-the-wave method, focusing on the slope of the \( i-V \) response at the onset of a catalytic current.\textsuperscript{38} The latter yields a TOF of 480 s\textsuperscript{-1} for TsOH (Figure S42). The production of hydrogen was confirmed by gas chromatographic analysis of the headspace gas after bulk electrolysis. Electrolysis for 1 h at \(-1.8\) V vs Fe\textsuperscript{0/+} (100 mV more negative than the onset potential at 50 mM of CF\textsubscript{3}CO\textsubscript{2}H in MeCN with [Bu\textsubscript{4}N]BF\textsubscript{4} electrolyte) afforded a Faradaic yield of 93\% H\textsubscript{2} (Figure S43). Electrocatalytic production of H\textsubscript{2} was monitored for 2 h. Rapid catalyst decomposition was observed at [CF\textsubscript{3}CO\textsubscript{2}H] > 100 mM.

5.12. Reduced Hydride [(dppv)Ni(\(\mu\)-pdt)(\(\mu\)-H)Fe(CO)(dppv))]\textsuperscript{0}, [H\textsubscript{1}]\textsuperscript{0}

Characterization of the reduced hydride was pursued to gain insight into the mechanism for hydrogen evolution. As indicated above, the [H\textsubscript{1}]\textsuperscript{+/0} couple proceeds with good reversibility. Treatment of THF solutions of [H\textsubscript{1}]\textsuperscript{+} with one equiv of CoCp\textsuperscript{*2} afforded a dark brown species with a single new \( \nu_{\text{CO}} \) band, shifted by 25 cm\textsuperscript{-1} toward lower energy (Figure 5.7).

![Figure 5.7](image)

**Figure 5.7.** FT-IR spectra of THF solutions of [H\textsubscript{1}]BAr\textsubscript{F}\textsubscript{4} (dashed line) and [H\textsubscript{1}]\textsuperscript{0} (solid line).

The stability of THF solutions of [H\textsubscript{1}]\textsuperscript{0} ranges from hours to days at room temperature, the rates apparently depending on the sample quality. Analysis of the headspace gas from a 50 mM solution of [H\textsubscript{1}]\textsuperscript{0} that had been standing at 26 °C for 3.6 h revealed that approximately 0.35 equiv
of H$_2$ (70% of theory, eq 1) was produced for every equiv of [H1]$^0$ consumed. IR spectra in the
v$_{CO}$ region showed isosbestic behavior (Figures S46 and S47). When the decay of [H1]$^0$ was
followed by $^{31}$P{$^1$H} NMR spectroscopy, [1]$^0$ was the exclusive product.

$$2 \text{[H1]}^0 \rightarrow 2 \text{[1]}^0 + \text{H}_2 \quad (1)$$

The EPR spectrum of freshly prepared [H1]$^0$ were simulated as three overlapping sets of
signals with $g_{iso} = 2.059$, 2.056, and 2.030 (Figure 5.8). The ratio changes over the course of
several minutes after the reduction (Figure S51), a process attributed to isomerization. The time-
dependence of the EPR spectrum is mirrored by IR measurements: a freshly reduced sample
showed $v_{CO}$ at 1914 cm$^{-1}$, which shifted to 1920 cm$^{-1}$ after several minutes (Figure S48).

**Figure 5.8.** EPR spectrum of a THF solution of [H1]$^0$ freshly reduced at $-77$ °C (black), together
with a simulated spectra of the three components and their sum.
The rate of isomerization, proceeding with a half life of 2 minutes, is about 1000x faster than the rates for interconversion of sym- and unsym-[H1]^0. The ratio of the three species is initially 1:1.2:1, respectively, reaching 1:0.1:0.1 upon equilibration (Figure 5.9). Each of these three species gives rise to a similar splitting pattern arising from $A \sim 360$ and 160 MHz and two smaller hyperfine splittings around 25 MHz (Table 5.5). The large hyperfine coupling constants are assigned to interactions of the unpaired electron with the two $^{31}$P(Ni) centers, while the two smaller coupling constants are assigned to interactions of the unpaired electron with the two $^{31}$P(Fe) centers. Two of the three species are attributed to two flippamers of unsym-[H1]^0, while the third species is assigned as sym-[H1]^0. The observation of only one flippamer of sym-[H1]^0 is consistent with the unsymmetrical steric environment around the central CH$_2$ of the pdt$^2$- ligand. The EPR spectra of [H1]^0 and [H1-$^{57}$Fe]^0 were indistinguishable, consistent with the spin primarily localized on Ni (Figure S52). Coupling assignable to the hydride was not observed as indicated by the small difference between the experimental and simulated spectra.$^{19}$

![EPR spectrum](image)

**Figure 5.9.** EPR spectrum of a THF solution of [H1]^0 equilibrated at 25 °C in red, together with a simulated spectra in black.
Table 5.5. EPR Simulation Parameters for [H1]$^0$

<table>
<thead>
<tr>
<th>Species</th>
<th>g-factor</th>
<th>$A$ ($^{31}$P) (MHz)</th>
</tr>
</thead>
<tbody>
<tr>
<td>unsym-[H1]$^0$ (flippamer A)</td>
<td>2.059</td>
<td>355, 166, 34, 30</td>
</tr>
<tr>
<td>unsym-[H1]$^0$ (flippamer B)</td>
<td>2.030</td>
<td>348, 164, 33, 28</td>
</tr>
<tr>
<td>sym-[H1]$^0$</td>
<td>2.056</td>
<td>387, 199, 22, 14</td>
</tr>
</tbody>
</table>

The changes in the EPR spectrum suggest that the kinetic product, presumably enriched in the unsym isomer as dominates in [H1]$^+$, isomerizes to a second species, this thermodynamic product presumably being the sym isomer. To independently verify this analysis, a thermally equilibrated MeCN solution of [H1]$^0$ was treated with one equiv of FcBF$_4$ to "freeze" the isomer ratio. The result, measured at 20 min at room temperature, was a 53:47 ratio of unsym-[H1]$^+$:sym-[H1]$^+$. This isomer ratio is significantly enriched in the sym isomer relative to the thermodynamic ratio (85:15 favoring unsym).

5.13. DFT Characterization of [H1]$^0$

DFT calculations were also carried out on the reduced mixed-valence species to obtain further structural, energetic, and mechanistic insights. The calculations indicate that the Ni center changes from square planar to tetrahedral in the conversion of [H1]$^+$ to [H1]$^0$ (Scheme 5.2). This change in geometry is consistent with the description of the reduced complex [H1]$^0$ as Ni(I)Fe(II). Although the geometry at Ni changes dramatically, the bond distances (Ni-Fe, Ni-H, and Fe-H) change by less than 5% (Table 5.2).

Spin densities were analyzed to determine the localization of the unpaired electron in the mixed-valence reduced bridging hydride species. In this species, the spin resides primarily (41-47%) on the Ni center, although spin density is also localized near the Fe, P, and S centers (Table 5.6). As observed for related NiFe dithiolates, the spin densities are relatively insensitive to the orientation of the pdt ring (Table S16). The unpaired spin density plots for both the unsym and sym isomers of [H1]$^0$ are depicted in Figure 5.10. Overall, the reduction of the bridging hydride species is consistently predicted to occur primarily at the Ni center.
Table 5.6. Spin Densities Calculated for the Reduced Hydride Complex \([H1]^0\)

<table>
<thead>
<tr>
<th>Complex</th>
<th>(\rho(\text{Ni}))</th>
<th>(\rho(\text{Fe}))</th>
<th>(\rho(2S))</th>
<th>(\rho(2P_{\text{Ni}}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>sym-[H1](^0)</td>
<td>0.47</td>
<td>0.19</td>
<td>0.14</td>
<td>0.12</td>
</tr>
<tr>
<td>unsym-[H1](^0)</td>
<td>0.41</td>
<td>0.28</td>
<td>0.14</td>
<td>0.12</td>
</tr>
</tbody>
</table>

*These values refer to the P atoms of the dppv ligand on the Ni center.

Similar to the parent Ni(II)Fe(II) derivatives, the hydride ligand for the reduced mixed-valence hydrides is also asymmetrically bound, with a shorter distance to Fe than to Ni (Table 5.2). This asymmetry is consistent with the greater spin density on Ni, which indicates that the species is best described as a Ni(I) paramagnet ligated to an Fe(II)-hydride.

Figure 5.10. Top-down view of the isocontour plot (isovalue = 0.005) of the unpaired spin density of \([H1]^0\) for the unsym (apical-basal-dppv on Fe, left) and sym (dibasal-dppv on Fe, right) isomers showing predominant localization around the Ni coordination sphere, with some localization near Fe as well. This difference is clarified by the numerical atomic spin densities provided in Table 5.6.

5.14. Protonation of \([H1]^0\)

The aim of preparing \([H1]^0\) was to test its role in the hydrogen evolution but in the absence of an electrode. Indeed, when treated with one equiv HBF\(_4\)·Et\(_2\)O, solutions of \([H1]^0\) immediately gave \(H_2\) concomitant with a color change from brown to orange. The \(^{31}\text{P}\{^1\text{H}\}\) NMR spectrum recorded 20 min after the addition of acid showed that a 72:28 mixture of \(\text{unsym-[H1]}^+\) and \(\text{sym-[H1]}^+\) is produced in 96% yield. Based on the rates of the relevant isomerizations, a ratio of ~75:25 is predicted. Analysis of the headspace by gas chromatography revealed that 0.32 equiv of \(H_2\) is
formed. This yield of H\textsubscript{2} corresponds to 64\% of the theoretically predicted value according to the stoichiometry in eq 2.

$$[\text{H}1]^{0} + \text{H}^{+} \rightarrow [\text{H}1]^+ + 0.5 \text{H}_2$$  \hspace{1cm} (2)

In terms of the mechanism for hydrogen evolution, the steps in eq 3-4 are invoked.

**Protonation of [H1]\textsuperscript{0} affords [I]\textsuperscript{+} and H\textsubscript{2} \textit{vide infra}:**

$$[\text{H}1]^{0} + \text{H}^{+} \rightarrow [\text{I}]^{+} + \text{H}_2$$  \hspace{1cm} (3)

**In terms of the mechanism for hydrogen evolution, the steps in eq 3-4 are invoked.**

**Electron transfer from [H1]\textsuperscript{0} to [I]\textsuperscript{+}:**

$$[\text{H}1]^{0} + [\text{I}]^{+} \rightarrow [\text{H}1]^+ + [\text{I}]^{0}$$  \hspace{1cm} (4)

**Protonation of [I]\textsuperscript{0}:**

$$[\text{I}]^{0} + \text{H}^{+} \rightarrow [\text{H}1]^+$$  \hspace{1cm} (5)

The redox process in eq 4 is favorable since our electrochemical measurements show that [H1]\textsuperscript{0} is about 0.5 V more reducing than [I]\textsuperscript{0}.

Three predictions arising from the proposed mechanism were tested. First, protonation of [H1]\textsuperscript{0} with D\textsuperscript{+} sources should give HD. When the reaction of [H1]\textsuperscript{0} and DOTf was monitored by \textsuperscript{1}H NMR spectroscopy, the HD/H\textsubscript{2} ratio was 5.3:1. Second, treatment of [H1]\textsuperscript{0} with a deficiency of H\textsuperscript{+} should give some [I]\textsuperscript{0}. Indeed, addition of 0.5 equiv of H(OEt\textsubscript{2})\textsubscript{2}BArF\textsubscript{4} to a solution of [H1]\textsuperscript{0} gave substantial but variable amounts of [I]\textsuperscript{0}. The variability of these yields is attributed to the sensitivity of [H1]\textsuperscript{0} toward impurities, which accelerate hydrogen evolution according to eq 1. Finally, the ratio \textit{unsym}-[H1]\textsuperscript{+}/\textit{sym}-[H1]\textsuperscript{+} should reflect on the mechanism since oxidation of [H1]\textsuperscript{0} predominantly gives significant sym, and protonation of [I]\textsuperscript{0} generates exclusively \textit{unsym} (see above). Indeed, treatment of a MeCN solution of [H1]\textsuperscript{0} with 1 equiv of HBF\textsubscript{4}·Et\textsubscript{2}O led to a 72:28 ratio of \textit{unsym}-[H1]\textsuperscript{+} to \textit{sym}-[H1]\textsuperscript{+} 20 min after treatment. This ratio is consistent with a mechanism in which half of the [H1]\textsuperscript{0} undergoes a sequence of protonation, reduction, and protonation to form almost exclusively \textit{unsym}-[H1]\textsuperscript{+}, and the other half is oxidized to a mixture of
unsym-[H1]⁺ and sym-[H1]⁺, which was shown independently to reach a 53:47 ratio 20 min after oxidation in MeCN (vide supra).

Several scenarios for the protonolysis of [H1]⁰ were examined with DFT to gain insight into the pathway for H₂ evolution. The free energy diagram for a proposed mechanism is shown in Scheme 5.5. The free energies were calculated with CF₃CO₂H as the acid (pKₐ MeCN = 12.7) and at an applied potential of –1.80 V vs Fc⁰/⁺ in MeCN, consistent with the bulk electrolysis experiments described above. Under these conditions, DFT calculations indicate that the production of H₂ via interaction of [H1]⁰ with CF₃CO₂H acid to give [1]⁺ is thermodynamically favorable (ΔG° = –7.0 kcal/mol, indicated by the red dashed line in Scheme 5.5). This process could occur as a concerted but most likely asynchronous protonation of the hydride and release of H₂, where ‘concerted’ indicates the absence of a thermally stable intermediate. A pathway involving protonation of the Fe-H bond in [H1]⁰ to form an intermediate dihydrogen complex, [H₂1]⁺, is also possible but is thermodynamically unfavorable (gray dashed lines in Scheme 5.5) under these specific conditions. However, this protonation step is more thermodynamically favorable when stronger acids, such as HBArF₄·2Et₂O, are used. Under such strong acid conditions, H₂ is evolved directly after the H₂ adduct is formed upon protonation of [H1]⁰. Thus, both mechanisms may be involved in H₂ evolution. Regardless, evolution of H₂ produces [1]⁺, which is then reduced to form [1]⁰ in the bulk electrolysis experiments. However, note that in the protonation experiments discussed above, an alternative mechanism, in which [1]⁺ is reduced by [H1]⁰ (eq 4), occurs in the absence of an electrode and could in principle also occur in the electrochemical experiments.
Scheme 5.5. Experimentally and computationally-supported pathway for hydrogen evolution catalyzed by 1. The free energies are calculated with CF$_3$CO$_2$H acid (pK$_a$ MeCN = 12.7) and at an applied potential of –1.8 V vs. Fe$^{0/+}$ in MeCN to coincide with the bulk electrolysis experiments. H$_2$ evolution is proposed to occur directly from interaction of [H1]$^0$ and CF$_3$CO$_2$H acid under these conditions (red dashed line), but stronger acids may also evolve H$_2$ from an intermediate dihydrogen complex, [H$_2$1]$^+$ (gray dashed lines). In the last step, an alternative mechanism for reduction of [1]$^+$ by [H1]$^0$ could also occur under certain conditions. The relative free energies for half reactions corresponding to electron transfer are calculated with respect to the applied potential, and the free energy difference between the initial and final states in the pathway corresponds to the potential associated with the double reduction of 2HA to produce 2A$^-$ and H$_2$ relative to this applied potential. An analogous scheme calculated relative to the HA/H$_2$ couple is presented in Figure S5. Ph-P groups omitted for clarity.

5.15. Conclusions

The pathway has been elucidated for hydrogen evolution catalyzed by a synthetic analogue of the [NiFe]-hydrogenase active site. The model system is [(dppv)Ni(µ-pdt)(µ-H)Fe(CO)(dppv)]$^+$, which does not release H$_2$ even upon protonation. Instead, reduction is required to initiate hydrogen evolution. Of particular interest is the spectroscopic and
computational characterization of the reduced hydride complex, which is stabilized by the steric protection afforded by the two dppv ligands. The reduced hydride complex is assigned as a Ni(I)Fe(II) species on the basis of FT-IR and EPR measurements and supported by DFT calculations. These studies revealed that reduction of the Ni center enhances the hydridic character of the Fe-H center. Hydrogen evolution occurs by protonation of the Fe-H bond leading to release of dihydrogen. It is intriguing to contrast the hydrogen evolving properties of [H\textsubscript{1}]\textsuperscript{0} with that the related mixed valence homobimetallic complex [HFe\textsubscript{2}(pdt)(CO)\textsubscript{2}(dppv)\textsubscript{2}], where the hydride is a spectator (it is not involved in H\textsubscript{2} release).\textsuperscript{17}

Comparisons of these NiFe models with the [NiFe]-hydrogenases are instructive. Overall, the biological system operates at one oxidation state higher than these models (Table 5.7). This difference reflects the effects of charge, since the model complexes are neutral or cationic, whereas the active site is highly anionic, which stabilizes higher oxidation states. With seven anionic ligands (4×RS\textsuperscript{−}, 2×CN\textsuperscript{−}, H\textsuperscript{−}), the active site in the Ni-R state is 3\textsuperscript{−} or 2\textsuperscript{−}, depending on the degree of protonation of a terminal thiolate ligand.\textsuperscript{11}

<table>
<thead>
<tr>
<th>formula</th>
<th>Ni-R</th>
<th>reduced hydride</th>
<th>Ni-C</th>
<th>oxidized hydride</th>
</tr>
</thead>
<tbody>
<tr>
<td>(RS)\textsubscript{2}Ni(µ-SR)\textsubscript{2} (µ-H)Fe(CO)(CN)\textsubscript{2} \textsuperscript{3−}</td>
<td>Ni(II), Fe(II)</td>
<td>Ni(III), Fe(II)</td>
<td>Ni(II), Fe(II)</td>
<td>Ni(II), Fe(II)</td>
</tr>
<tr>
<td>(PR\textsubscript{3})\textsubscript{2}Ni(µ-SR)\textsubscript{2} (µ-H)Fe(CO)(PR\textsubscript{3})\textsubscript{2}</td>
<td>Ni(II), Fe(II)</td>
<td>Ni(III), Fe(II)</td>
<td>Ni(II), Fe(II)</td>
<td>Ni(II), Fe(II)</td>
</tr>
<tr>
<td>(RS)\textsubscript{2}Ni(µ-SR)\textsubscript{2} (µ-H)Fe(CO)(CN)\textsubscript{2} \textsuperscript{2−}</td>
<td>Ni(II), Fe(II)</td>
<td>Ni(III), Fe(II)</td>
<td>Ni(II), Fe(II)</td>
<td>Ni(II), Fe(II)</td>
</tr>
<tr>
<td>(PR\textsubscript{3})\textsubscript{2}Ni(µ-SR)\textsubscript{2} (µ-H)Fe(CO)(PR\textsubscript{3})\textsubscript{2}</td>
<td>Ni(II), Fe(II)</td>
<td>Ni(III), Fe(II)</td>
<td>Ni(II), Fe(II)</td>
<td>Ni(II), Fe(II)</td>
</tr>
</tbody>
</table>

Crystallographic and DFT analyses for the Ni-R and Ni-C states in the enzyme reveal that the Ni-H and Fe-H distances are similar, with \(d_{\text{Ni-H}} \approx 1.6\) and \(d_{\text{Fe-H}} = 1.7\) Å.\textsuperscript{13} In synthetic models, the Fe-H distances are always shorter than the Ni-H distances, sometimes by 0.6 Å.\textsuperscript{4,41} X-ray spectroscopic studies indicate that the Fe-H site only weakly interacts with the Ni center in these
models. Studies on both the enzyme and models indicate, however, that these distances are not strongly affected by the metal oxidation state. The most prominent difference between the enzyme and the active site models is the coordination sphere of the nickel center. In the models, the Ni(SR)$_2$(PR$_3$)$_2$ site alternates between square planar and tetrahedral, whereas in the protein, the Ni(SR)$_4$ center maintains a distorted SF$_4$-like geometry. Future studies must address these three gaps: the rigidity of the Ni site, the anionic character of the site, and the fact that redox is confined to Ni. These daunting challenges can only, we believe, be addressed by radical changes in the coordination sphere of the Ni center.

5.16. Experimental Methods

General Considerations. Unless otherwise stated, reactions were conducted using standard Schlenk techniques, reagents were purchased from conventional sources, and solvents were HPLC grade and purified using an alumina filtration system (Glass Contour, Irvine, CA). ESI-MS data were acquired using a Waters Micromass Quattro II or ZMD spectrometer. Analytical data were acquired using an Exeter Analytical CE-440 elemental analyzer. NMR spectra were acquired with Varian U500 and VXR500 spectrometers. Chemical shifts (in ppm) were referenced to residual solvent peaks (for $^1$H, $^{13}$C) or external 85% H$_3$PO$_4$ (for $^{31}$P). Solution IR spectra were recorded on a Perkin-Elmer Spectrum 100 FTIR spectrometer. Crystallographic data were collected using a Siemens SMART diffractometer equipped with a Mo K$_\alpha$ source ($\lambda = 0.71073$ Å) and an Apex II detector. Photolysis employed a Spectroline MB-100 instrument (365 nm). Syntheses of Fe(CO)$_2$(pdt)(dppv)$^{44}$ and K[FeCp(CO)$_2$]$^{45}$ have been reported. CV's were simulated with the CHI630D package. EPR spectra were recorded on a Varian E-line E-12 Century series X-band CW spectrometer. Spectra were collected at room temperature with a microwave frequency of 9.30 GHz, 1 gauss modulation, and 20 mW microwave power. EPR
spectra were simulated using the program SIMPOW6. \(^{46}\) Samples of [1] \(^0\) and [H1] \(^0\) were handled in a glovebox with the exclusion of chlorocarbons.

\[
[(\text{dppv})\text{Ni}(\mu-\text{Cl})(\mu-\text{pdt})\text{Fe(}\text{dppv}\text{)}\text{(CO)})\text{]}\text{BF}_4, \quad [3\text{Cl}]\text{BF}_4. \quad \text{To a solution of}\n\]

Fe(pdt)(CO) \(_2\)(dppv) (1.00 g, 1.63 mmol) in 150 mL of acetone was added solid NiCl \(_2\)(dppv) (0.854 g, 1.63 mmol) and NaBF \(_4\) (0.196 g, 1.78 mmol). The resulting heterogeneous mixture was heated at reflux for 4 h at which point no solid was visible, and the red color had deepened. An FT-IR spectrum confirmed that Fe(CO) \(_2\)(pdt)(dppv) was fully consumed. After cooling in an ice bath, the reaction solution was filtered through a pad of Celite. The solvent was removed from the filtrate, and the solid was extracted into 50 mL of CH \(_2\)Cl \(_2\). The solution was diluted with 150 mL of hexanes, and the mixture was cooled to 0 °C for several days. Crystals of the product were isolated and dried under vacuum. Yield: 1.76 g (93%). \(^1\)H NMR (500 MHz, CD \(_2\)Cl \(_2\), 20 °C): \(\delta\) 8.05 – 7.87 (m, 2H, P(\(\text{CH}_2\)\_2\(\text{P}\)), 7.78 – 7.22 (m, 40H, Ph), 7.14 – 6.95 (m, 2H, P(\(\text{CH}_2\)\_2\(\text{P}\)), 2.05 (s, 1H, (\(\text{SCH}_2\)\_2\(\text{CH}_2\)), 2.02 (s, 2H, (\(\text{SCH}_2\)\_2\(\text{CH}_2\)), 2.00 – 1.93 (m, 2H, (\(\text{SCH}_2\)\_2\(\text{CH}_2\)), 1.94 – 1.85 (m, 1H, (\(\text{SCH}_2\)\_2\(\text{CH}_2\)), 1.04 – 0.91 (m, 2H, (\(\text{SCH}_2\)\_2\(\text{CH}_2\)). \(^{31}\)P\{1H\} NMR (202 MHz, CD \(_2\)Cl \(_2\)): \(\delta\) 73.3 (d, \(J_{\text{PP}}\) = 16.6 Hz, 2 P, Fe\(\text{P}\)), 57.6 (d, \(J_{\text{PP}}\) = 16.5 Hz, 2 P, Ni\(\text{P}\)). IR (acetone): \(v_{\text{CO}}\) = 1946 (s) cm \(^{-1}\).

Although not obtained in analytical purity (%C varied by at least 3%) crude samples were converted to the reduced complex.

\[
(\text{dppv})\text{Ni}(\mu-\text{pdt})\text{Fe(}\text{dppv}\text{)}\text{(CO), [1]}. \quad \text{In a 500-mL Schlenk flask was dissolved [ClI]}\text{BF}_4 \quad (1.25 g, 1.07 mmol) in 200 mL of acetone and stirred at room temperature. This solution was treated dropwise with a solution of K[FeCp(CO)\_2] (0.464, 2.14 mmol) in 60 mL of distilled acetone. The reaction mixture was allowed to stir for ~50 min. and was then concentrated to ~100 mL. The solution was diluted with 100 mL of pentane and was cooled in an ice bath for 4 h to yield a brown precipitate. The precipitate was isolated by cannula filtration while cold and rinsed
with pentane. The precipitate was then extracted into 2 x 30 mL of cold toluene, and these extracts were filtered through a pad of Celite. The toluene was removed under vacuum, and the remaining brown solid was recrystallized from 50 mL of THF and 150 mL of pentane at −40 °C. Yield: 0.972 g (87 %).  

$^1$H NMR (500 MHz, C$_6$D$_6$): $\delta$ 8.08 - 6.93 (44H, Ph and P(CH)$_2$P), 2.32 (m, 2H, (SCH$_2$)CH$_2$), 1.79 (m, 1H, (SCH$_2$)CH$_2$), 1.56 (m, 1H, (SCH$_2$)CH$_2$); 2H, (SCH$_2$)CH$_2$).  

$^{31}$P{$^{1}$H} NMR (202 MHz, C$_6$D$_6$): $\delta$ 89.2 (t, $^2$J$_{PP}$ = 13.6 Hz, 2P, Fe$_2$P), 74.2 (dt, $^2$J$_{PP}$ = 30.1, 15.5 Hz, 1P, NiP), 46.2 (dt, $^2$J$_{PP}$ = 28.3, 12.1 Hz, 1P, Ni-P). IR (THF): $\nu_{CO}$ = 1911 (sh), 1897 (s) cm$^{-1}$. Anal. Calcd for C$_{56}$H$_{50}$FeNiOP$_4$S$_2$: C, 64.58; H, 4.84. Found: C, 64.17; H, 4.70. Single crystals of [1]$^0$ were grown from THF solutions layered with pentane.

[(dppv)Ni($\mu$-H)/($\mu$-pdt)Fe(dppv)(CO)]OTf, [H1]OTf. A solution of [1]$^0$ (0.10 g, 0.10 mmol) in 20 mL of THF was treated with HOTf (10 $\mu$L, 0.11 mmol), causing a color change from brown to orange solution. An FT-IR spectrum confirmed that the starting material was fully consumed. The solution was concentrated to ~10 mL. The concentrated solution was layered with 10 mL of pentane and cooled to −40 °C for four days. The resulting orange crystals were collected, washed with pentane, and dried under vacuum. Yield: 0.114 g (99%). unsym isomer: $^1$H NMR (500 MHz, d$_3$-MeCN, 20 °C): $\delta$ −4.27 (t, $^2$J$_{PH}$ = 35.3 Hz, 1H, Fe($\mu$-H)Ni). $^{31}$P{$^{1}$H} NMR (202 MHz, d$_3$-MeCN, 20 °C): $\delta$ 87.8 (s, 2P, Fe$_2$P), 69.1 (s, 2P, sym.Ni-2P). Unsym isomer: $^1$H NMR (500 MHz, d$_3$-MeCN, 20 °C): $\delta$ −5.80 (t, $^2$J$_{PH}$ = 42.3 Hz, 1H, Fe($\mu$-H)Ni). $^{31}$P{$^{1}$H} NMR (202 MHz, d$_3$-MeCN, 20 °C): $\delta$ 93.4 (s, 1P, FeP), 79.4 (s, 1P, FeP), 73.5 (s, 2P, Ni-2P). IR (THF): $\nu_{CO}$ = 1962 (m), 1952 (s) cm$^{-1}$. Anal. Calcd for C$_{57}$H$_{51}$F$_3$FeNiO$_4$P$_4$S$_3$: C, 57.45; H, 4.31. Found: C, 57.22; H, 4.41.

[(dppv)Ni($\mu$-H)/($\mu$-pdt)Fe(dppv)(CO)]BAr$_4$F, [H1]BAr$_4$F. A solution of [H1]OTf (0.071 g, 0.065 mmol) in 10 mL of THF was treated with NaBAr$_4$F (0.058 g, 0.065 mmol), and the reaction
mixture was heated to 35 °C for 5 min. After cooling to room temperature and filtering, the reaction solution was evaporated, leaving an orange solid residue. **Sym** isomer: $^1$H NMR (500 MHz, C$_6$D$_6$, 20 °C): $\delta$ 4.37 (t, $J_{PH} = 34.0$ Hz, 1H, Fe(µ-H)Ni). $^{31}$P {$^{1}$H} NMR (202 MHz, C$_6$D$_6$, 20 °C): $\delta$ 88.2 (s, 2P, Fe2P), 68.4 (s, 2P, Ni2P). **Unsym** isomer: $\delta$ 8.45 – 6.84 (m, 52 H, Ar), 6.42-6.21 (m, 4 H, P(CH)$_2$P, 1.71 – 1.60 (m, $J_{HH} = 11$ Hz, 1H, (SCH)$_2$CH$_2$), 1.49 (d, $J_{HH} = 14.1$ Hz, 1H, (SCH)$_2$CH$_2$), 1.45 – 1.35 (m, 1H, (SCH)$_2$CH$_2$), 1.03 – 0.95 (m, 2H, (SCH)$_2$CH$_2$), 0.10 (t, $J_{HH} = 13.5$ Hz, 1H, (SCH)$_2$CH$_2$), -5.77 (t, $J_{PH} = 42.5$ Hz, 1H, Fe(µ-H)Ni). $^{31}$P {$^{1}$H} NMR (202 MHz, C$_6$D$_6$, 20 °C): $\delta$ 87.82 (s, 1P, FeP), 80.60 (s, 1P, Fe-P), 72.82 (s, 2P, Ni-2P). IR (THF): $\nu$$_{CO} = 1953$ (s), 1960 (sh).

[(dppv)Ni(µ-H)(µ-pdt)Fe(dppv)(CO)]BF$_4$, [H1]BF$_4$. A solution of [1]$^0$ (0.500 g, 0.48 mmol) in 75 mL of THF was treated with HBF$_4$·Et$_2$O (0.1 M, 0.5 mL, 0.05 mmol) to give an orange solution. After concentrating to half volume, the solution was layered with 100 mL of pentane at 0 °C. Small orange crystals were collected by filtration and dried under vacuum. Yield: 0.485 g (90%). The unsymmetrical isomer crystallized exclusively. $^1$H NMR (500 MHz, CD$_2$Cl$_2$, 20 °C): $\delta$ 8.10 (t, $J_{PH} = 8.3$ Hz, 2H, P(CH)$_2$P)), 7.71-7.12 (m, 40H, Ph), 6.91 – 6.77 (m, 2H, P(CH)$_2$P), 1.54 (t, $J_{HH} = 14.4$ Hz, 1H, (SCH)$_2$CH$_2$), 1.36 – 1.23 (m, 1H, (SCH)$_2$CH$_2$), 0.89 (d, $J_{HH} = 7.3$ Hz, 1H, (SCH)$_2$CH$_2$), 0.09 (d, $J_{HH} = 12.8$ Hz, 1H, (SCH)$_2$CH$_2$), -5.79 (t, $J_{PH} = 42.3$ Hz, 1H, Fe(µ -H)Ni). $^{31}$P {$^{1}$H} NMR (202 MHz, CD$_2$Cl$_2$, 20 °C): $\delta$ 87.82 (s, 1P, Fe-P), 80.60 (s, 1P, Fe-P), 72.82 (s, 2P, Ni-2P). IR (THF): $\nu$$_{CO} = 1963$ (sh), 1951 (s) cm$^{-1}$. Anal. Calcd for C$_{56}$H$_{51}$BF$_4$FeNiOP$_4$S$_2$: C, 59.56; H, 4.55. Found: C, 58.62; H, 4.66.

[(dppv)Ni(µ-H)(µ-pdt)Fe(dppv)(CO)], [H1]$^0$. To a suspension of [H1]BF$_4$ (0.100 g, 0.09 mmol) in Et$_2$O (500 µL) was added a solution of Cp*$^2$Co (0.029 g, 0.09 mmol) in Et$_2$O (500 µL).
After 5 min, the darkened solution was filtered through Celite, which was further rinsed with Et₂O (~2 mL). Solvents were removed under vacuum leaving a dark brown, almost black solid. Yield: 0.075 g (81%). IR (THF): ν\text{CO} = 1920 cm\(^{-1}\). Anal. Calcd for C\text{56}H\text{51}FeNiOP\text{4}S\text{2}: C, 64.51; H, 4.93. Found: C, 64.47; H, 5.28. Samples were handled in a glovebox with the exclusion of chlorocarbons.

5.17. Computational Methods.

The DFT calculations were performed with the BP86 functional\(^{47,48}\) with the SDD pseudopotential and associated basis set\(^{49}\) for Ni and Fe, the 6-31G** basis set\(^{50}\) for the \(\mu\)-H ligands, and the 6-31G* basis set\(^{51,52}\) for all other atoms. The starting coordinates for the geometry optimizations of \([1]^0\) and \([\text{H1}]^+\) were obtained from the available crystal structures, and starting coordinates for species without crystal structures were obtained by manually altering closely related crystal structures prior to optimizations. The conductor-like polarizable continuum model\(^{53,54}\) was used to calculate the solvation free energies using the Bondi atomic radii\(^{55}\) and including non-electrostatic contributions from dispersion,\(^{56,57}\) repulsion,\(^{56,57}\) and cavitation formation.\(^{58}\) The solvent was acetonitrile (MeCN) unless stated otherwise. Structural and energetic information for all studied compounds are provided in the Supporting Information (Tables S19-S54).

Herein, the geometry optimizations were performed in the gas phase; however, a comparison of the structures optimized in the gas and solution phases is provided in the SI (Tables S9-S12). All minimum energy structures were verified to have no imaginary frequencies, and all transition states (TSs) were confirmed to have only a single imaginary frequency. The TSs were located utilizing the synchronous transit-guided quasi-Newton method,\(^{59,60}\) and the intrinsic reaction coordinate (IRC) was followed using the local quadratic approximation\(^{61,62}\) to verify that
the TS leads to the correct reactant and product states. Thermochemical data were calculated at $T = 298.15$ K. Zero-point energy, entropic contributions, and solvation effects were included in the calculated reaction free energies ($\Delta G$) and free energy barriers ($\Delta G^\ddagger$) in solution. The relative reduction potentials and $pK_a$’s were calculated using methodology described elsewhere. Pipep-Mezey localization was performed using keyword IOp(4/9=20212). All calculations were performed with the Gaussian 09 electronic structure program.

5.18. References


Glendening, E. D.; Reed, A. E.; Carpenter, J. E.; Weinhold, F. *NBO Version 3.1*; University of Wisconsin System: Madison, WI, 1996.


Chapter 6. Models of the Ni-L and Ni-SI\textsubscript{a} States of the [NiFe]-Hydrogenase\textsuperscript{*},\textsuperscript{†}

6.1. Abstract

A new class of synthetic models for the active site of the [NiFe] hydrogenases are described. The Ni(I/II)(SCys)\textsubscript{2} and Fe(II)(CN)\textsubscript{2}CO, sites are represented with (RC\textsubscript{5}H\textsubscript{4})Ni(I/II), Fe(II)(diphos)(CO) modules, where diphos = 1,2-C\textsubscript{2}H\textsubscript{4}(PPh\textsubscript{2})\textsubscript{2} (dppe) or cis-1,2-C\textsubscript{2}H\textsubscript{2}(PPh\textsubscript{2})\textsubscript{2} (dppv). The two bridging thiolate ligands are represented by CH\textsubscript{2}(CH\textsubscript{2}S)\textsubscript{2} \textsuperscript{2−} (pdt\textsuperscript{2−}), Me\textsubscript{2}C(CH\textsubscript{2}S)\textsubscript{2} \textsuperscript{2−} (Me\textsubscript{2}pdt\textsuperscript{2−}), and (C\textsubscript{6}H\textsubscript{5}S)\textsubscript{2} \textsuperscript{2−}. Reaction of Fe(pdt)(CO)(dppe) and [(C\textsubscript{5}H\textsubscript{5})\textsubscript{3}Ni\textsubscript{2}]BF\textsubscript{4} affords [(C\textsubscript{5}H\textsubscript{5})Ni(pdt)Fe(dppe)(CO)]BF\textsubscript{4} \textsuperscript{[1a]}BF\textsubscript{4}. Monocarbonyl \textsuperscript{[1a]}BF\textsubscript{4} features an S = 0 Ni(II)Fe(II) center with pentacoordinate iron, as proposed for the Ni-SI\textsubscript{a} state of the enzyme. One-electron reduction of \textsuperscript{[1a]}\textsuperscript{+} affords the S = \textfrac{1}{2} derivative \textsuperscript{[1a]}\textsuperscript{0}, which according to density functional theory (DFT) calculations, EPR spectroscopy, and Mössbauer spectroscopy, is best described as a Ni(I)Fe(II) compound. The Ni(I)Fe(II) assignment matches that for the Ni-L state in [NiFe]-hydrogenase, unlike recently reported Ni(II)Fe(I)-based models. Compound \textsuperscript{[1a]}\textsuperscript{0} reacts with strong acids to liberate 0.5 equiv of H\textsubscript{2} and regenerate \textsuperscript{[1a]}\textsuperscript{+}, indicating that H\textsubscript{2} evolution is catalyzed by \textsuperscript{[1a]}\textsuperscript{0}. DFT calculations were used to investigate the pathway for hydrogen evolution and revealed that the mechanism can proceed through two isomers of \textsuperscript{[1a]}\textsuperscript{0} that differ in the stereochemistry of the Fe(dppe)CO center. Calculations suggest that protonation of \textsuperscript{[1a]}\textsuperscript{0} (both isomers) affords Ni(III)-H-Fe(II) intermediates, which represent mimics of the Ni-C state of the enzyme.


\textsuperscript{†} Author contributions: Dr. Geoffrey M. Chambers, Dr. Yulong Li, Dr. Edward Reijerse, Prof. Thomas Rauchfuss, and Prof. Wolfgang Lubitz were responsible for all experiments and associated analyses in this chapter.
6.2. Introduction

Hydrogenases are enzymes that catalyze the oxidation of hydrogen and reduction of protons at high rates and low overpotentials, reactions that are potentially useful for clean energy applications.\textsuperscript{1-4} Hydrogenases are classified into three main families according to the metals in their active site: [Fe]- [FeFe]- and [NiFe]-hydrogenases.\textsuperscript{5,6} These enzymes are oxygen-sensitive, but the latter can recover from oxygen poisoning.\textsuperscript{7,8} Model compounds have been useful in elucidating catalytic mechanisms or atom identity, notably with [FeFe]-hydrogenases.\textsuperscript{9} [FeFe]-hydrogenase models are fairly advanced and have strong resemblance and relevance to the enzymatic system.\textsuperscript{3,4,10} In contrast, models of the oxygen-tolerant [NiFe]-hydrogenases are less mature, and comparisons to the enzymatic system are sometimes less applicative.\textsuperscript{3,4,11,12}

![Figure 6.1](image_url) States of the active site of the [NiFe]-hydrogenases.

The [NiFe]-hydrogenase active site features a [Ni(SCys)\textsubscript{4}]\textsuperscript{2-} center linked through two thiolate bridges to an Fe(CN)\textsubscript{2}(CO) center (Figure 6.1). In the absence of a third bridging ligand, the geometry at nickel resembles that of SF\textsubscript{4}, while iron is square pyramidal. The active states of the enzyme differ in three ways: (i) the oxidation state of the Ni center, (ii) the presence/absence of a hydride ligand bridging the two metals, and (iii) the state of protonation of at least one terminal cysteinyl thiolate. In contrast to the fluctuating oxidation states at the nickel center, iron appears to remain low-spin Fe(II) throughout all of the known states of the enzyme active site.
Table 6.1. Summary of Models for the Active States of the [NiFe]-Hydrogenase

<table>
<thead>
<tr>
<th>State</th>
<th>Core Description</th>
<th>State of the Art Models</th>
<th>Comment</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ni-R</td>
<td>Ni$^{II}$-H-Fe$^{II}$</td>
<td>[(diphos)Ni(pdt)(µ-H)Fe(CO)$_3$L]$^+$</td>
<td>• models: $d_{Fe-H} &lt; d_{Ni-H}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>[(amine)$_2$Ni(SR)$_2$(µ-H)FeL$<em>3$]$^-</em>{14}$</td>
<td>• enzyme: $d_{Fe-H} &gt; d_{Ni-H}$</td>
</tr>
<tr>
<td>Ni-SI</td>
<td>Ni$^{II}$-Fe$^{II}$</td>
<td>[(diphos)Ni(SR)$_2$Fe(CO)$_2$L$<em>2$]$^-</em>{16}$</td>
<td>• model: six-coord. Fe</td>
</tr>
<tr>
<td>Ni-C</td>
<td>Ni$^{II}$-H-Fe$^{II}$</td>
<td>none</td>
<td>• enzyme: five coord. Fe</td>
</tr>
<tr>
<td>Ni-L</td>
<td>Ni$^{II}$-Fe$^{II}$</td>
<td>[(diphos)Ni(pdt)Fe(CO)$_3$L]$^+$</td>
<td>• no progress</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>• model: Ni(II)Fe(I)</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>• enzyme: Ni(II)Fe(II)$^{15}$</td>
</tr>
</tbody>
</table>

$^a$diphos = R$_2$PCH$_2$CH$_2$PR$_2$; (amine)$_2$(SR)$_2$ = tetradeutate diaminodithiolate; $^b$ L = PR$_3$ or P(OR)$_3$; pdt$^{2-} = CH$_2$(CH$_2$S)$_2$.

Beyond complementing our understanding of the enzyme, synthetic models could underpin the development of new catalysts for processing of hydrogen. The possibility of connecting synthetic complexes to enzyme states was established with [(dppe)Ni(pdt)(µ-H)Fe(CO)$_3$]$^+.^{13,20}$ A related nickel-iron hydride [(amine)$_2$Ni(SR)$_2$(µ-H)FeL$_3$]$^-_{14}$ catalyzes reduction of redox dyes by H$_2$, a characteristic assay for the hydrogenases. With a Ni(II)-H-Fe(II) core, these synthetic complexes represent early models for the Ni-R state of the protein, despite discrepancies (Table 6.1).$^{15}$

Concurrent with the development of these diamagnetic models,$^{20,21}$ efforts have been made to replicate the paramagnetic states of the enzyme. The two most prominent paramagnetic states are Ni-C and Ni-L. Invoked in all catalytic mechanisms is Ni-C,$^{6,22-25}$ which features a hydride bridging the Ni(III) and Fe(II) centers. No synthetic analogues of this Ni(III) hydride has been even detected, but this paper provides a first glimpse of a synthetic model. With a Ni(I)Fe(II) core, Ni-L is typically generated by low-temperature photolysis of Ni-C, and recent evidence points to its role in catalysis.$^{26,27}$ These $S = 1/2$ states deviate most strongly from conventional chemistry of dimetal systems and therefore offer the greatest opportunity for discovery of fundamentally new reactivity. The first reported $S = 1/2$ NiFe complex was [(diphos)Ni(pdt)Fe(CO)$_3$L]$^+$. Spectroscopic, crystallographic, and computational analyses show, however, that these cations are
described as Ni(II)Fe(I) derivatives,\textsuperscript{19,28} not Ni(I)Fe(II) as assigned to the Ni-L state of the enzyme.\textsuperscript{15} The lapse is a consequence of the Fe(CO)\textsubscript{2}L center, which stabilizes Fe(I). In the present report, the Fe(CO)\textsubscript{2}L center is replaced by Fe(diphos)(CO) module, which more strongly stabilizes Fe(II).\textsuperscript{21}

Addressed in this work is a flaw in the first generation of active site models, i.e., the reorganization upon changing the oxidation state of Ni.\textsuperscript{20} Redox at the Ni centers typically induces changes in coordination geometry, such that Ni(I) species (c.f. Ni-L) are tetrahedral and Ni(II/III) species (c.f. Ni-SI\textsubscript{a}, Ni-R, Ni-C) are square planar. In contrast, the NiFe active site changes only subtly as it converts among catalytic states.\textsuperscript{15,27} To minimize reorganizational barriers, the stereodynamic $[\text{Ni(diphos)}]^z$ center is replaced by $[(\text{C}_5\text{H}_5)\text{Ni}]^z$. The $[(\text{C}_5\text{H}_5)\text{Ni}]^z$ center has the advantage that its geometry is relatively insensitive to redox, in contrast to Ni(dppe)-based models (Figure 6.2). Not only does the hapticity of C\textsubscript{5}H\textsubscript{5} not change, crystallographic studies\textsuperscript{29-31} show that the Ni-(C\textsubscript{5}H\textsubscript{5}) distances are relatively invariant (<0.12 Å) for Ni(I)-, Ni(II)-, and Ni(III)-(C\textsubscript{5}H\textsubscript{5}) compounds.

Figure 6.2. Ni-diphosphine-complexes change geometry upon reduction of tetrahedral Ni$^{0/1}$. The Ni(C\textsubscript{5}H\textsubscript{5}) center is structurally invariant over three oxidation states.


The reaction of Fe(pdt)(CO)\textsubscript{2}(dppe) and $[(\text{C}_5\text{H}_5)_3\text{Ni}_2]BF_4$ in nitromethane affords $[(\text{C}_5\text{H}_5)\text{Ni}(\text{pdt})\text{Fe(dppe)(CO)}]BF_4$ ([1a]BF\textsubscript{4}, Scheme 6.1) as a dark-red powder. The derivatives $[(\text{C}_5\text{H}_5)\text{Ni}(\text{pdt})\text{Fe(dpv)(CO)}]BF_4$ ([1b]BF\textsubscript{4}), $[(\text{C}_5\text{H}_5)\text{Ni}(\text{Me}_2\text{pdt})\text{Fe(dppe)(CO)}]BF_4$ ([1c]BF\textsubscript{4}), and $[(\text{CH}_3\text{C}_5\text{H}_4)\text{Ni}(\text{pdt})\text{Fe(dppe)(CO)}]BF_4$ ([1d]BF\textsubscript{4}) where dpv = \textit{cis}-1,2-\text{C}_2\text{H}_2(P\text{Ph}_2)_2 and
Me$_2$pd$^2$ – Me$_2$C(CH$_2$S$^-$)$_2$, were also prepared using the appropriate nickel and iron precursors. These salts are similar to [1a]BF$_4$ in terms of their reactivity and spectroscopic properties.

![Scheme 6.1](image)

The reaction of Fe(pdt)(CO)$_2$(dppe) and [(C$_5$H$_5$)$_3$Ni$_2$]BF$_4$ proceeds via an intermediate with $\nu_{CO} = 2050$ cm$^{-1}$ that converts to [1a]$^+$. We assign this intermediate band to the dicarbonyl complex $\text{trans-}[$(C$_5$H$_5$)Ni(pdt)Fe(dppe)(CO)$_2$]$^+$. The disappearance of the intermediate band follows first order decay (Figures S45 and S46). When the reaction progress is monitored by $^{31}$P NMR spectroscopy, the iron precursor is consumed within ca. 5 min and two new singlet resonances are observed, one of which corresponds to the product [1a]$^+$. Additionally, electrospray ionization mass spectrometry (ESI-MS) of the reaction mixture indicates the presence of the molecular cation [CpNi(pdt)Fe(dppe)(CO)$_2$]$^+$. These observations are consistent with rapid attachment of “(C$_5$H$_5$)Ni$^+$” to the two isomers of the iron dithiolate. The cis-carbonyl derivative appears to release CO rapidly, whereas the isomer with trans-CO ligands is somewhat persistent.

Purified compounds [1a–d]BF$_4$ display a single $\nu_{CO}$ band in the range of 1940–1951 cm$^{-1}$, consistent with the presence of one isomer. These values are similar to those observed in Ni-SI$_a$, which range from 1927 to 1947 cm$^{-1}$ (Table 6.2).
Table 6.2. IR Data for Compounds [1a-d]BF₄ and [2b]BF₄ in CH₂Cl₂ Solution and the Ni-SIₐ State in Two Organisms

| Compound | $\nu_{CO}$ (cm⁻¹) | Expt | Cale
|----------|------------------|------|------
| [(C₅H₅)Ni(pdt)Fe(dppe)(CO)]BF₄, ([1a]BF₄) | 1943 | 1948/1946 |
| [(C₅H₅)Ni(pdt)Fe(dpvp)(CO)]BF₄, ([1b]BF₄) | 1951 | 1951/1949 |
| [(C₅H₅)Ni(Me₂pdt)Fe(dppe)(CO)]BF₄, ([1c]BF₄) | 1940 | 1937/1935 |
| [(CH₃C₅H₄)Ni(pdt)Fe(dppe)(CO)]BF₄, ([1d]BF₄) | 1943 | 1947/1945 |
| [(C₅H₅)Ni(SPh)₂Fe(dppv)(CO)]BF₄ ([2b]BF₄) | 1956 | 1959 |
| Ni-SIₐ (D. vulgaris Miyazaki F)³² | 1943 | n/a |
| Ni-SIₐ (A. vinlandii)³² | 1931 | n/a |
| Ni-SIₐ (D. gigas)³² | 1934 | n/a |

³²The first value corresponds to the conformer in which the central CH₂ of the Fe(pdt)Ni core is oriented toward Fe, and the second value corresponds to the conformer in which the same CH₂ is oriented toward Ni. The calculations were performed in the absence of BF₄⁻.

The structure of [1a]⁺ was confirmed by X-ray crystallography. The Ni–Fe distance is 2.5145(4) Å (Figure 6.3). Although the enzyme in the SIₐ state has not been characterized crystallographically, it is generally assumed that the Ni–Fe distance is ~2.6 Å in all active states.⁶ 

The stereochemistry of the FeL₂(CO) site is biomimetic in the sense that CO occupies an apical site, and the two Lewis basic phosphine ligands are approximately trans to the sulfur centers of the thiolates.

The $^{31}$P NMR spectra of complexes [1a–d]BF₄ are simple, exhibiting only a single resonance, indicating chemically equivalent phosphorus centers or a dynamic process. The diphosphine in complexes [1a–d]BF₄ is only observed in the dibasal arrangement, whereas other Fe(dithiolate)(diphosphine)CO compounds have been shown to adopt both a dibasal and an apical-basal arrangement. The phosphorus centers remain NMR-equivalent at −90 ºC, suggesting either an exceptionally low barrier for isomerization or the presence of a single isomer. In the IR spectrum only a single isomer is observed, which, like the enzyme's active site, has both donor ligands in basal positions.
The reaction of cis,cis-Fe(SPh)\textsubscript{2}(CO)\textsubscript{2}(dppv) and \((C\textsubscript{5}H\textsubscript{5})\textsubscript{3}Ni\)\textsubscript{2}BF\textsubscript{4} gave the dicarbonyl \(((C\textsubscript{5}H\textsubscript{5})\textsubscript{3}Ni(SPh)\textsubscript{2}Fe(dppv)(CO)\textsubscript{2})\)BF\textsubscript{4} (\([2bCO]BF\textsubscript{4}\)), isolated as dark, almost black microcrystals. Its \(^1\)H NMR spectrum exhibits only one Cp signal, and an AB quartet pattern in the \(^{31}\)P NMR spectrum, indicating low symmetry. The structure was confirmed crystallographically (Figure S53). The Ni---Fe distance is 3.308(1) Å, indicating a non-bonding interaction. The stereochemistry at Fe is identical to the precursor, wherein all ligands are cis. The \(\mu\)-SPh groups are diaxial.

Upon treatment with Me\textsubscript{3}NO, the dicarbonyl readily decarbonylated to give \([2b]BF\textsubscript{4}\), isolated as a black solid. The IR spectrum of \([2b]^{+}\) in CH\textsubscript{2}Cl\textsubscript{2} solution displays a single \(\nu_{CO}\) band at 1956 cm\(^{-1}\), vs 2042, 2002 cm\(^{-1}\) observed for the precursor. The monocarbonyl cation, \([2b]^{+}\), decomposes gradually in CH\textsubscript{2}Cl\textsubscript{2} solutions. However, the complex is stable in MeCN and the \(\bar{\nu}_{CO}\) band shifts 22 cm\(^{-1}\) to higher energies, suggesting the formation of an adduct. The \(^{31}\)P NMR spectrum of \([2b]^{+}\) exhibits a singlet at \(\delta 78.8\). Under 1 atm of CO, MeCN solutions of \([2b]^{+}\) partially convert back to \([2bCO]^{+}\). In contrast, \([1a]^{+}\) is not reactive towards CO.
6.4. Electrochemistry

The cyclic voltammetry of \([1\text{a}]\text{BF}_4\) in dichloromethane is very rich (Figure 6.4). Compounds \([1\text{b-d}]^+\) produce similar voltammograms (Table 6.3). Compounds \([1\text{a-d}]^+\) exhibit an irreversible oxidation near +0.65 V, a reversible one-electron reduction near −1.20 V, all vs Fe\(^{0+/+}\). Additionally, a quasi-reversible reduction is observed near −2.15 V, which becomes fully reversible when in THF with \([\text{Bu}_4\text{N}]\text{BAR}_4\) as electrolyte. An irreversible reduction wave is generated at ca. 0.0 V vs Fe\(^{0+/+}\) following the irreversible oxidation.

![Cyclic voltammogram of \([1\text{a}]\text{BF}_4\) at 500 mVs\(^{-1}\). Conditions: ~1 mM in CH\(_2\text{Cl}_2\), 100 mM \([\text{NBu}_4\text{PF}_6\), glassy carbon electrode (d = 3 mm); Ag wire as pseudoreference with internal Fc standard at 0 V; Pt counter electrode.](image)

**Figure 6.4.** Cyclic voltammogram of \([1\text{a}]\text{BF}_4\) at 500 mVs\(^{-1}\). Conditions: ~1 mM in CH\(_2\text{Cl}_2\), 100 mM \([\text{NBu}_4\text{PF}_6\), glassy carbon electrode (d = 3 mm); Ag wire as pseudoreference with internal Fc standard at 0 V; Pt counter electrode.

**Table 6.3.** Electrochemical Properties of Compounds \([1\text{a-d}]^+\) (V vs Fe\(^{+/0}\) in CH\(_2\text{Cl}_2\))

<table>
<thead>
<tr>
<th>Compound</th>
<th>(E_{\text{p,ox}}) (Ni(^{II})Fe(^{II})/Ni(^{III})Fe(^{II}))</th>
<th>Calc(^{b,c})</th>
<th>(E_{\text{p,red}}) (Ni(^{I})Fe(^{II})/Ni(^{I})Fe(^{II}))</th>
<th>Calc(^{b,c})</th>
<th>(E_{\text{p,red}}) (Ni(^{II})Fe(^{III})/Ni(^{II})Fe(^{III}))</th>
<th>Calc(^{b,c})</th>
</tr>
</thead>
<tbody>
<tr>
<td>([1\text{a}]\text{BF}_4)</td>
<td>0.70</td>
<td>0.70</td>
<td>−1.16</td>
<td>−1.16</td>
<td>−2.15</td>
<td>−2.15</td>
</tr>
<tr>
<td>([1\text{b}]\text{BF}_4)</td>
<td>0.72</td>
<td>0.73</td>
<td>−1.15</td>
<td>−1.12</td>
<td>−2.18</td>
<td>−2.05</td>
</tr>
<tr>
<td>([1\text{c}]\text{BF}_4)</td>
<td>0.67</td>
<td>0.76</td>
<td>−1.09</td>
<td>−1.16</td>
<td>−2.15</td>
<td>−2.12</td>
</tr>
<tr>
<td>([1\text{d}]\text{BF}_4)</td>
<td>0.62</td>
<td>0.66</td>
<td>−1.21</td>
<td>−1.23</td>
<td>−2.17</td>
<td>−2.18</td>
</tr>
</tbody>
</table>

\(^a\)For the calculated values, the central CH\(_2\) of the Fe(pdt)Ni core is oriented toward Fe and assumed to remain in this orientation upon reduction or oxidation. Additional calculated values corresponding to the same CH\(_2\) being oriented toward Ni, or flipping from Fe to Ni, are available in the Supporting Information (Tables S1-S2). In all cases, the conformation of the Fe(pdt)Ni core does not affect the potential. \(^b\)[1\text{a}]\text{BF}_4 was used as the reference so the experimental and calculated values agree by construction. \(^c\)The calculations were performed on the positively charged species in the absence of BF\(_4\). \(^d\)This wave is assigned to a Ni\(^{II}\)Fe\(^{III}\)/Ni\(^{I}\)Fe\(^{II}\) couple, where the Ni\(^{I}\)Fe\(^{II}\) state is an open-shell singlet.
The redox couples in the different derivatives are relatively similar. The reversible couple is most affected by Me$_2$pdt substitution. With $\Delta \tilde{\nu}_{CO} = 3$ cm$^{-1}$ (Table 6.1), the donor properties of pdt$^{2-}$ and Me$_2$pdt$^{2-}$ are similar in this system, and the difference in the potential of the Ni$^{II}$Fe$^{II}$/Ni$^{I}$Fe$^{II}$ redox couple may be due to the steric effects of the methyl groups. The reversible reduction in complexes $[1a-d]^+$ shifts to more positive potentials by 50 mV upon replacing Cp for MeC$_5$H$_4$. This effect suggests that the reduction is nickel centered, generating a Ni$^{I}$Fe$^{II}$ species. C$_5$H$_5$C$_5$H$_4$ substitution most affected the oxidation observed at $\sim$0.75 V, shifting this couple by $\sim$80 mV, whereas other substitutions only generate a 20–30 mV effect. This observation suggests that oxidation is Ni-centered. The reduction near $\sim$2 V is negligibly affected by the Ni-based ligands, and is proposed to be an Fe-centered reduction, generating a Ni(I)Fe(I) species.

DFT was used to calculate the potentials associated with the couples $[1a-d]^{2+/+}$, $[1a-d]^{+/-0}$, and $[1a-d]^{0/-}$. The potentials were calculated relative to $[1a]$BF$_4$ and are provided in Table 6.3. The experimental trends are well reproduced by the calculations. The BF$_4^-$ ions were not included in the calculations, which may cause deviations between the calculated and experimental values. The reduction potentials were also calculated for an alternative geometry optimization scheme, as well as using different functionals, and these results are provided in the Supporting Information (Tables S1-S2).

6.5. Models for Ni-L: Synthesis and Structure

The reduced species $[1a-d]^0$ and $[2b]^0$ were produced on a preparative scale by treatment with (C$_5$H$_5$)$_2$Co ($E^0 = -1.23$ V, eq 1).

$$\text{[(C}_5\text{H}_5\text{)}\text{Ni(SR)}_2\text{Fe(diphos)(CO)}]\text{BF}_4 + (\text{C}_5\text{H}_5)_2\text{Co} \rightarrow (\text{C}_5\text{H}_5\text{)}\text{Ni(SR)}_2\text{Fe(diphos)(CO)} + [(\text{C}_5\text{H}_5)_2\text{Co}]\text{BF}_4 \quad (1)$$

Purified samples $[1a-d]^0$ and $[2b]^0$ were isolated as black, air-sensitive solids. Reduction of $[1a-d]^+$ and $[2b]^0$, a Ni$^{II}$Fe$^{II}$/Ni$^{I}$Fe$^{II}$ couple, shifts $\tilde{\nu}_{CO}$ by ca. 40 cm$^{-1}$ to around 1900 cm$^{-1}$. A similar change is observed for reduction of [(dppe)Ni(pdt)Fe(CO)$_3$]$^+$, a Ni$^{II}$Fe$^{I}$/Ni$^{I}$Fe$^{I}$ couple, ($\Delta \tilde{\nu}_{CO} \approx 35$
Oxidation state changes centered at iron typically shift \( \nu_{CO} \) by ca. 100 cm\(^{-1} \) as observed for in (dppe)Pt(pdt)Fe(CO)\(_3\)]\(^{+}/0\), a Pt\(^{II}\)/Fe\(^{I}/\)Pt\(^{II}\)/Fe\(^{0}\) couple.\(^{20} \) For the Ni-L state of the [NiFe]-hydrogenases, values of \( \nu_{CO} \) range from 1911 (\(D.\) vulgaris Miyazaki F)\(^{33} \) to 1862 cm\(^{-1} \) (\(A.\) aeolicus).\(^{26,34} \)

Table 6.4. IR Data for \( \nu_{CO} \) in Compounds [1a-d]\(^0\) and [2b]\(^0\) in CH\(_2\)Cl\(_2\) Solution

<table>
<thead>
<tr>
<th>Compound</th>
<th>Expt</th>
<th>Calc(^a)</th>
</tr>
</thead>
<tbody>
<tr>
<td>[1a](^0)</td>
<td>1901</td>
<td>1901/1901</td>
</tr>
<tr>
<td>[1b](^0)</td>
<td>1903</td>
<td>1909/1906</td>
</tr>
<tr>
<td>[1c](^0)</td>
<td>1897</td>
<td>1898/1899</td>
</tr>
<tr>
<td>[1d](^0)</td>
<td>1898</td>
<td>1900/1900</td>
</tr>
<tr>
<td>[2b](^0)</td>
<td>1915</td>
<td>1911</td>
</tr>
</tbody>
</table>

\(^a\)The first value corresponds to the isomer in which the central CH\(_2\) of the Fe(pdt)Ni core is oriented toward Fe, and the second value corresponds to the isomer in which the same CH\(_2\) is oriented toward Ni.

Figure 6.5. Solid state structure of [1a]\(^0\) shown at the 50% probability level. Hydrogen atoms and solvent have been omitted for clarity. Selected bond distances (Å): Ni1-Fe1, 2.4593(6); Ni1-Cent, 1.795(3); Ni1-S\(_{avg}\), 2.2772(5); Fe1-S\(_{avg}\), 2.2544(5); Fe1-C6, 1.745(3); C6-O1, 1.159(3).

X-ray crystallographic analysis confirmed that the solid state structures of [1a]\(^0\) (Figure 6.5) and [1a]\(^+\) (Figure 6.3) are very similar. The principal changes in these structures are in the coordination sphere of nickel, which suggest that the couple \([1a]^{+}/0\) involves a Ni-centered reduction. The Ni-(C\(_5\)H\(_5\)) and Ni-S bonds elongate by about 5% while the Fe-ligand bond lengths
remain virtually unchanged. The Ni-Fe distance shortens by ~0.1 Å to 2.4593(5) Å, which is well within the sum of the atomic radii (2.56(6) Å), indicating the presence of metal-metal bonding. Although there are no crystallographic data on the metal-metal bonded Ni-L state of the enzyme, DFT calculations predict this distance to be 2.56 Å.\textsuperscript{35} The observed structural changes are consistent with reduction at the nickel center in the transformation of [1a]\textsuperscript{+} to [1a]\textsuperscript{0}. The DFT optimized structures of [1a]\textsuperscript{+} and [1a]\textsuperscript{0} were in excellent agreement with the crystal structures (Table 6.5).

**Table 6.5.** Key Bond Distances (Å) in [1a]BF\textsubscript{4}, [1a]\textsuperscript{0}, and the Difference from Crystal Structures and DFT Calculations

<table>
<thead>
<tr>
<th>Bond</th>
<th>[1a]BF\textsubscript{4}</th>
<th>Calc</th>
<th>[1a]\textsuperscript{0}</th>
<th>Calc</th>
<th>Experimental Difference</th>
<th>DFT Difference</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ni-Fe</td>
<td>2.5145(4)</td>
<td>2.566</td>
<td>2.4593(5)</td>
<td>2.508</td>
<td>−0.055</td>
<td>−0.058</td>
</tr>
<tr>
<td>Ni-centroid</td>
<td>1.722(3)</td>
<td>1.739</td>
<td>1.795(3)</td>
<td>1.814</td>
<td>+0.073</td>
<td>+0.075</td>
</tr>
<tr>
<td>Ni-S\textsubscript{avg}</td>
<td>2.1645(7)</td>
<td>2.168</td>
<td>2.2772(5)</td>
<td>2.259</td>
<td>+0.113</td>
<td>+0.091</td>
</tr>
<tr>
<td>Fe-S\textsubscript{avg}</td>
<td>2.2208(7)</td>
<td>2.236</td>
<td>2.2544(5)</td>
<td>2.252</td>
<td>+0.034</td>
<td>+0.016</td>
</tr>
<tr>
<td>Fe-P\textsubscript{avg}</td>
<td>2.2256(7)</td>
<td>2.249</td>
<td>2.235(8)</td>
<td>2.236</td>
<td>+0.009</td>
<td>−0.013</td>
</tr>
<tr>
<td>Fe-C</td>
<td>1.757(2)</td>
<td>1.721</td>
<td>1.745(3)</td>
<td>1.727</td>
<td>−0.012</td>
<td>+0.006</td>
</tr>
<tr>
<td>C-O</td>
<td>1.147(2)</td>
<td>1.171</td>
<td>1.159(3)</td>
<td>1.174</td>
<td>+0.012</td>
<td>+0.003</td>
</tr>
</tbody>
</table>

**Figure 6.6.** Views of the isocontour plot of the spin density for [1a]\textsuperscript{0}, looking down the Ni-Fe bond (left) and from the side of the Ni-Fe bond (right). The spin density is localized about the Ni coordination sphere. Color scheme: green, Ni; maroon, Fe; yellow, S; purple, P; red, oxygen; white, C. The hydrogens have been removed for clarity.

DFT analysis of the spin density was used to determine the site of reduction of [1a-d]\textsuperscript{+}. The plot of the spin density for [1a]\textsuperscript{0} (Figure 6.6) reveals that the unpaired electron resides
primarily on the Ni center, with some delocalization onto the pdt$^{2-}$ and (C$_5$H$_5$)$^-$ ligands, but no significant delocalization onto the Fe. This result is consistent with the Ni(I)Fe(II) assignment and the experimentally observed elongation of the Ni-ligand bonds upon reduction of $[1a]^+$.

The spin densities for the reduced species $[1a-d]^0$ and $[2b]^0$ are given in Table 6.6 and consistently indicate a Ni-based reduction across all species. As observed for related NiFe dithiolate complexes, the spin densities are not significantly affected by the conformation of the pdt ligand. Furthermore, the calculations also predict that an Fe-based reduction would require potentials 0.2–0.4 V more negative than the experimentally observed Ni-based reduction potentials (Table S1).

<table>
<thead>
<tr>
<th>Compound</th>
<th>Ni</th>
<th>Fe</th>
<th>2S</th>
<th>C$_5$H$_5$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$[1a]^0$</td>
<td>0.71</td>
<td>−0.17</td>
<td>0.24</td>
<td>0.24</td>
</tr>
<tr>
<td></td>
<td>0.71</td>
<td>−0.19</td>
<td>0.23</td>
<td>0.25</td>
</tr>
<tr>
<td>$[1b]^0$</td>
<td>0.69</td>
<td>−0.16</td>
<td>0.22</td>
<td>0.25</td>
</tr>
<tr>
<td></td>
<td>0.69</td>
<td>−0.17</td>
<td>0.22</td>
<td>0.25</td>
</tr>
<tr>
<td>$[1c]^0$</td>
<td>0.69</td>
<td>−0.15</td>
<td>0.22</td>
<td>0.24</td>
</tr>
<tr>
<td></td>
<td>0.71</td>
<td>−0.20</td>
<td>0.24</td>
<td>0.25</td>
</tr>
<tr>
<td>$[1d]^0$</td>
<td>0.69</td>
<td>−0.17</td>
<td>0.23</td>
<td>0.25$^b$</td>
</tr>
<tr>
<td></td>
<td>0.70</td>
<td>−0.19</td>
<td>0.23</td>
<td>0.26$^b$</td>
</tr>
<tr>
<td>$[2b]^0$</td>
<td>0.72</td>
<td>−0.20</td>
<td>0.20</td>
<td>0.26</td>
</tr>
</tbody>
</table>

$^a$For each complex, the first entry corresponds to the isomer in which the central CH$_2$ of the Fe(pdt)Ni core is oriented toward the Fe center, and the second entry corresponds to the isomer in which the same CH$_2$ is oriented toward the Ni center. $^b$For $[1d]^0$, this is the spin density over the CH$_3$C$_5$H$_4$ ligand.

Crystallographic analysis revealed that the structure of $[2b]^0$ differs strongly from that of $[1a]^0$ (Figure 6.7). Specifically the Fe center is inverted such that the CO ligand is poised between the Fe and Ni centers (Scheme 6.2).
The CO ligand is not bridging as indicated by the long Ni---CO distance of 2.878 Å. The DFT optimized geometry of [2b]$^0$ is in good agreement with the crystal structure (Table S8), and the Ni---CO distance is calculated to be 2.916 Å. Furthermore, the $\tilde{\nu}_{\text{CO}}$ for [2b]$^0$ is calculated to be 1911 cm$^{-1}$ for the rotated structure, close to the observed value of 1915 cm$^{-1}$. A small shoulder at approximately 1900 cm$^{-1}$ is also observed in the spectrum that may result from the presence of a second rotamer. For [2b]$^0$, the isomer in which the CO is apical, as it is in the structure of [2b]$^+$, is calculated to be higher in free energy by 3.70 kcal/mol compared to the isomer in which the CO ligand is poised between the Fe and Ni centers. This free energy difference is consistent with the crystal structure of [2b]$^0$ featuring an inverted Fe center.

Figure 6.7. Solid state structure of [2b]$^0$ shown at the 50% probability level. Hydrogen atoms and solvent have been omitted for clarity. Selected bond distances (Å): Ni1-Fe1, 2.4594(3); Ni1-cent, 1.817(2); Ni1-S$_{\text{avg}}$, 2.3067(5); Fe1-S$_{\text{avg}}$, 2.2868(5); Fe1-P$_{\text{avg}}$, 2.1934(5); Fe1-C44, 1.738(2); C44-O1, 1.162(2).

6.6. Models for Ni-L: Spectroscopy

The X-band CW-EPR spectrum of [1a]$^0$ is rhombic with g-values of 1.991, 2.031, and 2.138. No hyperfine coupling was detected (Figure 6.8). Spectra of [1b-d]$^0$ are very similar to [1a]$^0$ (Table 6.7). The absence of significant $^{31}$P hyperfine coupling is consistent with a nickel
centered radical, indicating that these compounds feature \((C_5H_5)Ni(I)\) centers. Large \(^{31}\)P hyperfine coupling was observed in related mixed-valence compounds containing Fe(I) phosphine centers, e.g., \([{(CO)}_2(PR_3)Fe(pdt)Ni(diphosphine)}]^+\).\(^{19,28}\)

Q-band ENDOR experiments were conducted to probe the spin distribution in the mixed-valence state. Consistent with the absence of any hyperfine splitting in the EPR spectrum, the \(^{31}\)P hyperfine couplings were found to be very small (~11 MHz) (Table 6.8 and Figure S6). Several proton hyperfine contributions were detected using ENDOR spectroscopy with hyperfine splittings ranging from 4 to 11 MHz (Figure S5). The variant of complex \([1a]^0\) carrying \(C_5D_5\) ((1a-d5)) in place of \(C_5H_5\) afforded a greatly simplified \(^1\)H ENDOR pattern with the largest proton coupling (~11 MHz) being absent (Figure S4), indicating that the spin density is localized on the Ni and its immediate coordination sphere. The weaker \(^1\)H couplings observed for \([1a-d5]^0\) are assigned to the methylene protons of pdt\(^2\). These observations are consistent with the DFT analysis, which shows that the spin density in \([1a]^0\) and \([2b]^0\) is primarily localized on the \((C_5H_5)Ni(pdt)\) center (Figure 6.6 and Table 6.6). Despite the different stereochemistry at Fe, at least in the solid state, and the presence of SPh\(^-\) vs R\(_2\)pdt\(^-\), the EPR spectrum of all mixed valence compounds are similar.

**Figure 6.8.** X-band CW-EPR spectrum of ~2 mM frozen solution of [1a]\(^0\) in 1:1 CH\(_2\)Cl\(_2\)/THF (110K).
Table 6.7. EPR Parameters for Compounds [1a-1d]$^0$ and [2b]$^0$ (Frozen CH$_2$Cl$_2$/Toluene solution at 110 K) and the Ni-L State in *Desulfovibrio vulgaris* Miyazaki F$^0$

<table>
<thead>
<tr>
<th>Compounds</th>
<th>$g_1$, $g_2$, $g_L$</th>
</tr>
</thead>
<tbody>
<tr>
<td>[1a]$^0$</td>
<td>1.991, 2.031, 2.138</td>
</tr>
<tr>
<td>[1b]$^0$</td>
<td>1.987, 2.031, 2.134</td>
</tr>
<tr>
<td>[1c]$^0$</td>
<td>1.997, 2.031, 2.134</td>
</tr>
<tr>
<td>[1d]$^0$</td>
<td>1.987, 2.039, 2.130</td>
</tr>
<tr>
<td>[2b]$^0$</td>
<td>1.984, 2.039, 2.143</td>
</tr>
<tr>
<td>Ni-L1 (<em>D. vulgaris</em>)</td>
<td>2.046, 2.118, 2.296</td>
</tr>
</tbody>
</table>

The [1a-$d_5$]$^0$ orientation-selective ENDOR patterns (Figure S8) could be simulated with one proton hyperfine interaction. Although, in principle, up to nine protons (C$_5$H$_5$ + CH$_2$(CH$_2$S)$_2$) could contribute to the $^1$H ENDOR spectra of [1a]$^0$, only one additional contribution, assigned to the (C$_5$H$_5$) ligand, could be distinguished and evaluated (see Figure S9). It can be assumed that the proton(s) with their hyperfine tensors are approximately aligned with the $g$-tensor axes leading to the strongest contributions along these axes in the orientation selective ENDOR spectra. The hyperfine tensors of the C$_5$H$_5$ protons are not aligned with the $g$-tensor axes and give rise to additional weaker features in the ENDOR spectra (Figures S9), which have not been simulated. The simulation parameters of the C$_5$H$_5^-$ and pdt$_2^2^-$ (CH$_2$S) contributions are given in Table 6.8.

Table 6.8. Hyperfine Couplings in [1a]$^0$ from Q-band ENDOR and Applied Field Mössbauer Measurements

<table>
<thead>
<tr>
<th>Nucleus</th>
<th>HFI expt. (MHz)</th>
<th>$\Lambda_1$</th>
<th>$\Lambda_2$</th>
<th>$\Lambda_3$</th>
</tr>
</thead>
<tbody>
<tr>
<td>C$_5$H$_5$</td>
<td></td>
<td>4.86</td>
<td>+3.25</td>
<td>−11.25</td>
</tr>
<tr>
<td>CH$_2$(C$_1$H$_2$S)$_2$</td>
<td></td>
<td>7.3</td>
<td>4.4</td>
<td>4.36</td>
</tr>
<tr>
<td>$^{31}$P</td>
<td></td>
<td>11.1</td>
<td>11.3</td>
<td>12.7</td>
</tr>
<tr>
<td>$^{57}$Fe</td>
<td></td>
<td>−3.8</td>
<td>−0.73</td>
<td>−0.9</td>
</tr>
</tbody>
</table>

The measured Mössbauer spectra of [1a–c]$^0$ and the corresponding cations are quite similar. All exhibit low isomer shifts of $\delta = 0.11$–0.14 mm s$^{-1}$ and small quadrupole splittings of $\Delta E_q = 0.72$ to 1.15 mm s$^{-1}$ (Table 6.9). The Mössbauer data are in agreement with a low-spin iron in all complexes. Consistent with other data in this paper, the spectra support the oxidation state assignments of Ni(II)Fe(II) for compounds [1a–c]BF$_4$ and Ni(I)Fe(II) for compounds [1a–c]$^0$. The
isomer shifts observed in the Ni(I)Fe(II) compounds 1a-c are larger than the shift observed for the Ni(II)Fe(I) compound [(dppe)Ni(pdt)Fe(CO)]BF₄ (δ = 0.04 mm s⁻¹), while the quadrupole splittings are comparable (ΔEₚ = 0.57). Mössbauer data for the [NiFe]-hydrogenases are sparse and suffer from difficulties in identifying and correcting for the dominating subspectra of the accessory FeS clusters. Isomer shifts in the range of 0.05–0.15 mm s⁻¹ have been assigned to the iron center of the [NiFe] center.

<table>
<thead>
<tr>
<th>Compound</th>
<th>δ (mm s⁻¹)</th>
<th>ΔEₚ (mm s⁻¹)</th>
<th>η</th>
</tr>
</thead>
<tbody>
<tr>
<td>[1a]BF₄</td>
<td>0.11</td>
<td>1.15</td>
<td></td>
</tr>
<tr>
<td>[1b]BF₄</td>
<td>0.11</td>
<td>1.10</td>
<td></td>
</tr>
<tr>
<td>[1c]BF₄</td>
<td>0.12</td>
<td>0.87</td>
<td></td>
</tr>
<tr>
<td>[1a]⁺</td>
<td>0.14</td>
<td>−0.73</td>
<td>0.13</td>
</tr>
<tr>
<td>[1b]⁺</td>
<td>0.13</td>
<td>−0.72</td>
<td></td>
</tr>
<tr>
<td>[1c]⁺</td>
<td>0.13</td>
<td>−0.76</td>
<td></td>
</tr>
</tbody>
</table>

ΔEₚ Signs for [1a-c]⁺ are obtained from magnetic Mössbauer spectra.

The Mössbauer spectrum of [1a]⁺ was further examined at high magnetic fields. This spectrum shows a very weak internal field contribution with an isotropic ⁵⁷Fe hyperfine coupling of 2 MHz (Table 6.8, Figure S7). For comparison, magnetic Mössbauer spectra of the Fe(I) center in the H₉₀ state of the [FeFe]-hydrogenase indicate ⁵⁷Fe hyperfine couplings in the range 10–20 MHz. The weak coupling is consistent with the spin being localized on the nickel center.

### 6.7. Electronic Structure Analysis

The geometries of the various species were optimized using DFT/BP86 and compared to the available experimental data to further probe the electronic structure of [1a-d]⁺ and [1a-d]⁻. The optimized geometries are in good agreement with the X-ray crystal structures. Relevant bond lengths for [1a]⁺ and [1a]⁻ are compared in Table 6.5. The BF₄⁻ counterions were not included in the geometry optimizations, which may introduce minor discrepancies between the X-ray crystal structure and the DFT optimized structure. Comparisons between the DFT and X-ray crystal
structures for species [1b-d]^+ and [2b]^0 as well as structural and energetic information for all systems studied, are provided in the Supporting Information (Tables S4-S8). Moreover, the CO stretching frequencies (\(\tilde{\nu}_{CO}\)) calculated with DFT are also in good agreement with experimental data (Tables 6.2 and 6.4).

The short Ni–Fe distances in both [1a]^+ and [1a]^0 suggest the presence of metal-metal bonding interactions. According to previous analyses, such Ni-Fe bonds typically arise from overlap of the \(d_{z^2}\) orbitals.\(^{20,41,42}\) Interestingly, however, the nature of the Ni-Fe bonding is quite different between [1a]^+ and [1a]^0, as revealed by analysis of the bonding molecular orbitals localized using the Pipek-Mezey criteria.\(^{43}\) For [1a]^+, the bonding orbital (Figure 6.9, left) reveals a dative Ni→Fe 2-center, 2-electron bond, which resembles a Lewis acid-base type interaction between the \(d^8\) Ni(II) and \(d^6\) Fe(II) centers. In contrast, for [1a]^0, a more covalent 2-center, 2-electron bond is present (Figure 6.9, right), in which the Ni and Fe \(d_{z^2}\) orbitals contribute equally to the bonding orbital. An alternative bonding analysis using Natural Bond Orbitals (NBO)\(^{44}\) is provided in the Supporting Information (pages S37-S40). The formation of a stronger \(\sigma\)-bond in [1a]^0 is consistent with the experimentally observed contraction of the Ni-Fe bond distance from 2.51 Å in [1a]^+ to 2.46 Å in [1a]^0. This Ni-Fe bond length contraction is also found in the DFT-optimized structures (Table 6.5). These chemical bonding patterns are similar for [2b]^+ and [2b]^0 (Figure S58), indicating that the stereochemistry at Fe and the presence of SPh\(^–\) or R\(_2\)pdt\(^2–\) does not disrupt the Ni-Fe bond.
Figure 6.9. Pipek-Mezey localized molecular orbitals showing the Ni-Fe $\sigma$-bonds in $[1a]^+$ and $[1a]^0$. The composition of the MO in $[1a]^+$ is 85% Ni and 15% Fe, and the composition of the MO in $[1a]^0$ is 53% Ni and 47% Fe. Only the $\alpha$ orbital is shown for $[1a]^0$; however, the $\beta$ orbital is similar.


Upon treatment with $\geq 1$ equiv acid, solutions of $[1a]^0$ liberate $H_2$ and afford formation of $[1a]^+$ (eq 2).

\[ [(C_5H_5)Ni(pdt)Fe(CO)(dppe)] + H^+ \rightarrow [(C_5H_5)Ni(pdt)Fe(CO)(dppe)]^+ + \frac{1}{2} H_2 \]  

Hydrogen evolution was observed for HBF$_4$, $p$-toluenesulfonic acid ($pK_a^{MeCN} = 8.01$), and [pyridinium]BF$_4$ ($pK_a^{MeCN} = 12.33$), but not for [NH$_4$]PF$_6$ ($pK_a^{MeCN} = 14.46$). With HBF$_4$ the yield of $H_2$ was 0.44 equiv, as assayed by gas chromatography. With $[1a]^0$ functioning as a 1 $e^-$ reductant, a 0.5 equiv would be predicted.

The protonolysis of $[1a]^0$ to give $[1a]^+$ entails the conversion of an $S = 1/2$ precursor to an $S = 0$ product, which implicates a multistep process involving electron transfer. We anticipated that protonation of $[1a]^0$ would afford $[(C_5H_5)Ni(pdt)HFe(CO)(dppe)]^+$ ($[H1a]^+$), a Ni(III)-Fe(II) species (see DFT section below). However, hydrides were not obtained experimentally; rather, only $[1a]^+$ was obtained. To explain this result, $[H1a]^+$ is proposed to undergo reduction by its conjugate base $[1a]^0$, giving $[H1a]^0$, which in turn undergoes protonolysis to give $[1a]^+$. Similarly,
we found that protonation of $[2b]^0$ afforded $[2b]^+$. The proposed sequence of reactions is summarized in Scheme 6.3.

\[
\text{Step A: } [1a]^0 + H^+ \rightarrow [H1a]^+
\]
\[
\text{Step B: } [H1a]^+ + [1a]^0 \rightarrow [H1a]^0 + [1a]^+
\]
\[
\text{Step C: } [H1a]^0 + H^+ \rightarrow [1a]^+ + H_2
\]

\[\text{Scheme 6.3.}\]

6.9. Proposed H$_2$ Evolution Mechanism

In the presence of excess acid, $[1a]^{0/+}$ is a formal, albeit slow electrocatalyst for hydrogen evolution. Using trifluoroacetic acid, catalytic current is observed at $-1.16V$, corresponding to the $[1a]^{+/0}$ couple, with an acid-independent rate of ca. $4 \text{ s}^{-1}$ (Figure S32). Since the hydride intermediates could not be characterized, they were analyzed with DFT in the context of the mechanism for hydrogen evolution. Additionally, the exclusive observation of the dibasal intermediate prompted theoretical investigation of an apical-basal isomer. These calculations provide insight into the structure and bonding of the Ni(III)-hydride intermediate and also predict the occurrence of two parallel catalytic cycles involving isomers of $[1a]^0$. Similar bimetallic hydrides have been known to adopt different isomers.$^{21,45}$ Furthermore, recent work suggests that the protonation of (dppe)Ni(pdt)Fe(CO)$_3$ proceeds via an unobserved isomer.$^{20}$
Proposed mechanism for H₂ production proceeding through two isomers. Cycle A, on the left, proceeds through structures in which the dppe ligand is in the dibasal position. Cycle B, on the right, proceeds through structures in which the dppe ligand spans apical-basal positions. Isomerization of [1a]₀, which has the lowest free energy barrier, could lead to a transition between Cycle A and Cycle B.

Table 6.10. Spin Densities and Bond Lengths (Å) of [1a]₀, [H1a]+, and [H1a]₀⁺

<table>
<thead>
<tr>
<th>Isomer</th>
<th>ρ(Ni)</th>
<th>ρ(Fe)</th>
<th>ρ(C₅H₅)</th>
<th>d(Ni-Fe)</th>
<th>d(Ni-H)</th>
<th>d(Fe-H)</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>0.71</td>
<td>-0.19</td>
<td>0.23</td>
<td>2.53</td>
<td>1.80</td>
<td>1.61</td>
</tr>
<tr>
<td>B</td>
<td>0.70</td>
<td>-0.16</td>
<td>0.23</td>
<td>2.47</td>
<td>1.79</td>
<td>1.58</td>
</tr>
</tbody>
</table>

For these calculations, the central CH₂ of the Fe(pdt)Ni core is oriented toward the Ni center. An analogous table for the cases where the same CH₂ is oriented toward the Fe center is provided in the Supporting Information (Table S10).

According to the calculations, the dibasal and apical-basal isomers of [1a]₀ are nearly isoergic (ΔG° = +0.48 kcal/mol) and interconvert via a relatively low free energy barrier, ΔG‡ = 11.76 kcal/mol (Table 6.11); however, all other protonation and oxidation states have a much higher barrier for isomerization (ΔG > 20 kcal/mol). This result prompted the investigation of two parallel pathways for hydrogen evolution, labeled A and B in Figure 6.10.
Table 6.11. Calculated Reaction Free Energies and Free Energy Barriers (kcal/mol) for Isomerization of Dibasal to Apical-Basal Isomers for [1a]$^0$ and its Derivatives in the Proposed Catalytic Cycle$^{a,b}$

<table>
<thead>
<tr>
<th>Compound</th>
<th>$\Delta G^\circ$</th>
<th>$\Delta G^\ddagger$</th>
</tr>
</thead>
<tbody>
<tr>
<td>[1a]$^0$</td>
<td>0.48</td>
<td>11.76</td>
</tr>
<tr>
<td>[H1a]$^+$</td>
<td>2.21</td>
<td>20.84</td>
</tr>
<tr>
<td>[H1a]$^0$</td>
<td>5.18</td>
<td>25.64</td>
</tr>
<tr>
<td>[H21a]$^+$</td>
<td>9.86</td>
<td>30.03</td>
</tr>
<tr>
<td>[1a]$^+$</td>
<td>2.03</td>
<td>21.54</td>
</tr>
</tbody>
</table>

$^a$See Figure 6.9 for structures of all species.  
$^b$For the isomerization considered here, the central CH$_2$ of the Fe(pdt)Ni core is oriented toward the Ni. The calculated values corresponding to the cases where the same CH$_2$ is oriented toward the Fe are provided in the Supporting Information (Table S9).  
$^c$The potential energy surface was found to be relatively flat along the isomerization pathway, leading to relatively small imaginary frequencies for the transition states (TS's).

These pathways are initiated by protonation of isomers of [1a]$^0$ that differ in terms of the stereochemistry of the (pdt)Fe(dppe)(CO) site. The adoption of the apical-basal isomer does not disrupt the Ni-Fe $\sigma$-bond (Figure S56) or the localization of the unpaired spin density (Table 6.10 and Figure S57). The isomerization from the dibasal to the apical-basal configuration is expected to occur only for the [1a]$^0$ species because the calculated $\Delta G^\ddagger$ values for the analogous isomerization for all other intermediates in the catalytic cycle shown in Figure 6.9 are more than 20 kcal/mol (Table 6.11). DFT analysis revealed that the transition state for this isomerization disrupts the metal-metal bonding. For [1a]$^0$, the transition state is best described as Ni(II)Fe(I), while the transition state is Ni(II)Fe(II) in [1a]$^+$. The lower isomerization barrier for [1a]$^0$ is attributed to the greater stability of 5-coordinate Fe(I) compared to 5-coordinate Fe(II). The structures of [1a]$^+$ and [1a]$^0$ are similar; however, the main difference is that the Ni center is reduced from Ni(II) to Ni(I) in [1a]$^0$, which contracts the Ni–Fe bond and elongates the Ni-ligand bonds. The isomerization of the Fe(CO)(dppe) subunit from dibasal (Cycle A) to apical-basal (Cycle B) proceeds through a lower free energy barrier for [1a]$^0$ (11.76 kcal/mol) than for [1a]$^+$.
While the structures of $[\text{1a}]^+$ and $[\text{1a}]^0$ are similar, the transition states (TSs) associated with their respective isomerizations are quite different. As seen in Figure 6.11, the TS for $[\text{1a}]^0$ features an elongated Ni-Fe distance and intramolecular electron transfer from the Ni to the Fe center to form a Ni(II)Fe(I) TS. The pentacoordinate Fe(I) center of the TS of $[\text{1a}]^0$ is more accommodating of rotation than the hexacoordinate Fe(II) center in the TS of $[\text{1a}]^+$. Similar arguments could be made for the other intermediates in the catalytic cycle, which also feature hexacoordinate Fe(II) centers and further rigidifying bridging hydride ligands.

![Figure 6.11. Free energy diagram for the isomerization of the Fe(CO)(dppe) subunit from dibasal (Cycle A) to apical-basal (Cycle B) for $[\text{1a}]^+$ (black numbers) and $[\text{1a}]^0$ (blue numbers).](image)

The first step in the catalytic cycle produces isomers of [H$\text{1a}$]$^+$. In both isomers, the unpaired electron is localized mainly in the Ni coordination sphere (Table 6.10), indicating that
these species are best described as Ni(III)HFe(II), analogous to the Ni-C state in the enzyme. Moreover, the DFT-optimized structure of [H1a]$^+$ reveals that the hydride is more closely bound to Fe than to Ni.

Subsequent to its formation, [H1a]$^+$ is reduced by a second equiv. of [1a]$^0$ to form [1a]$^+$ and the Ni(II)Fe(II) hydride [H1a]$^0$. This electron transfer is proposed to occur spontaneously on the basis of the calculated reduction potential of [H1a]$^+$, which is less negative than the calculated reduction potential of [1a]$^+$ by 0.16 (Cycle A) and 0.03 V (Cycle B) vs Fe$^{+}/0$ in CH$_2$Cl$_2$. Reduction to form [H1a]$^0$ mainly affects the Ni center and results in significant elongation of the Ni-H distance by 0.64 and 0.26 Å for the A and B isomers, respectively. Thus, [H1a]$^0$, which would be analogous to the Ni-R state of the enzyme, is more accurately described as a semi-terminal Fe(II)-hydride than a bridging hydride.

The catalytic cycle closes with the protonation of [H1a]$^0$ to produce the H$_2$-adduct intermediate, [H$_2$1a]$^+$, followed by the release of H$_2$ from [H$_2$1a]$^+$ to generate [1a]$^+$. Although the two cycles are very similar energetically, the calculations predict that this protonation is thermodynamically less favorable for the apical-basal isomer [H1a$^B$]$^0$ than for the dibasal isomer [H1a$^A$]$^0$ by ~5 kcal/mol.

In light of these computational results, additional experiments were performed to probe for the presence of two isomers in the catalytic cycle. The calculated $\tilde{\nu}_{CO}$ values for [1a$^A$]$^0$ and [1a$^B$]$^0$ are 1901 and 1904 cm$^{-1}$, respectively, and the calculated $\tilde{\nu}_{CO}$ values for [1a$^A$]$^+$ and [1a$^B$]$^+$ are 1946 and 1949 cm$^{-1}$, respectively. The calculated $\tilde{\nu}_{CO}$ values for all intermediates in the catalytic cycles are provided in the Supporting Information (Table S11). These small calculated $\tilde{\nu}_{CO}$ differences between the isomers are within the numerical accuracy of DFT. As previously mentioned, solutions of [1a]$^+$, which were synthesized directly, were confirmed crystallographically and
spectroscopically to exist only in the dibasal geometry. However, preparation of [1a]$^+$ by the addition of excess acid to [1a]$^0$, presumably going through the catalytic cycle(s) to generate H$_2$, produced a solution containing both isomers of [1a]$^+$. The $^{31}$P NMR spectrum of the [1a]$^+$ solution obtained after presumably proceeding through the catalytic cycle(s) displayed two $^{31}$P NMR signals, one assigned to the dibasal isomer (Cycle A), in which the two $^{31}$P centers are equivalent, and the other assigned to the apical-basal isomer (Cycle B), in which the two $^{31}$P centers are equilibrated by the rocking motion shown in Figure 6.12. The free energy barrier to this rocking motion is calculated to be relatively low, $\Delta G^\ddagger = 9.5$ kcal/mol, and the $^{31}$P-NMR spectrum of the isomer mixture was unchanged when recorded at -100 °C, consistent with a relatively low free energy barrier. These analyses support the proposal that H$_2$ evolution can proceed via either Cycle A or Cycle B shown in Figure 6.10, thereby producing both isomers of [1a]$^+$.

![Figure 6.12. Representation of the Fe-centered rocking motion in [1a$^B$]$^+$.](image)

### 6.10. Conclusions

Experiment and theory indicate that the new complexes reported in this paper replicate aspects of the following enzyme states: Ni-SI$_a$, Ni-L, Ni-C, and Ni-R. Two of these species, the Ni(II)Fe(II) and Ni(I)Fe(II) derivatives, were crystallized and fully characterized. The other two species are implicated in a hydrogen evolution experiment and characterized computationally.

The Ni-SI$_a$ state is distinguished by the Ni(II)Fe(II) bimetallic core and the absence of a third bridging ligand, features that are replicated in compounds [1a-d]BF$_4$. The Ni-L state is distinguished by the formal Ni(I)Fe(II) oxidation states, a feature that is replicated in [1a-d]$^0$. The
oxidation state assignments are supported by EPR and Mössbauer spectroscopic data, crystallographic results, voltammetric measurements, and DFT calculations.

The previous attempts to model states with Ni(II)Fe(II) cores suffered from the presence of hexacoordinate Fe centers.\textsuperscript{16,46} The Ni(II)Fe(II) models reported in this paper in fact arise via such FeL₄(µ-SR)₂-containing intermediates, followed by rapid decarbonylation. DFT analysis of the bonding suggests that decarbonylation is driven by formation of the Ni-Fe bond. Although [\textbf{1a-d}]\textsuperscript{+} and [\textbf{2b}]\textsuperscript{+} are reasonable spectroscopic models for the Ni-SI\textsubscript{a} state of the enzyme, challenges remain. In the Ni-SI\textsubscript{a} state the Ni center binds carbon monoxide to give Ni-SI\textsubscript{a}CO.\textsuperscript{6,47} In contrast, [\textbf{1a-d}]\textsuperscript{+} exhibit no detectable affinity for CO. Substitution of the dithiolate bridge with bridging monothiolates, as in [\textbf{2b}]\textsuperscript{+}, facilitates the binding of carbon monoxide at iron. In contrast, the enzyme binds carbon monoxide at nickel. This difference is understandable because the (C₅H₅)Ni(µ-SR)₂ site has a higher electron count than the (RS)₂Ni(µ-SR)₂ site found in the enzyme. Replicating terminal thiolate ligation at nickel remains one of the central challenges in this area.\textsuperscript{48}

The first attempts to model the mixed valence active site of the Ni-L state focused on the [(CO)₂LFe(pdt)Ni(diphosphine)]\textsuperscript{+}.\textsuperscript{19,28} These cations are described as Ni(II)Fe(I), which is reversed from the Ni(I)Fe(II) states assigned in Ni-L. More recently, we have characterized complexes with the configuration Ni(I)Ru(II), wherein the Ru(II) replaces the redox-inactive ferrous site in active site.\textsuperscript{49} In this work, the Fe(CO)₂(PR₃) center in previous Fe-Ni models was replaced by a Fe(CO)(PR₃)₂ center, which stabilizes the Fe(II) state, as illustrated by [\textbf{1a-d}]\textsuperscript{0} and [\textbf{2b}]\textsuperscript{0}.

Because they are substrates, hydrogenic ligands are a major focus in research concerning the hydrogenases and their models.\textsuperscript{50,51} The DFT-optimized analysis of the Ni(III)Fe(II) hydride
[H1a]+ reveals that the unpaired electron is localized mainly in the Ni coordination sphere, as seen for Ni-C. In the model, the hydride ligand is more closely bound to Fe than to Ni [Δ(M–H) ~ 0.2 Å]. This report is the first suggestion of the accessibility of such a state in [NiFe]-hydrogenase model complexes. In the enzyme, proton transfer reactions are coupled to electron transfer reactions, and a similar effect is seen in the reduction of [H1a]+ by [1a]0. The resulting Ni(II)Fe(II) hydride [H1a]0, analogous to the Ni-R state of the enzyme, reacts with proton sources to liberate H2. Although thwarting efforts to isolate this hydride, this facile protonolysis is analogous to the behavior of the Ni-R/Ni-SI pair. According to our calculations, the Ni(II)Fe(II) hydride is more accurately described as a terminal Fe(II)-hydride than a bridging hydride. In contrast, high resolution X-ray crystallographic analysis of Ni-R from D. vulgaris reveals Ni-H and Fe-H distances of 1.58 and 1.78 Å (error 0.08 Å), respectively, i.e., the hydride is more closely associated with nickel. The structural difference between the model complex and the protein may reflect the influence of terminal thiolate ligands on nickel.

6.11. Experimental

Reactions and manipulations were performed using standard Schlenk techniques at room temperature or in a nitrogen atmosphere glove box. Solvents were HPLC-grade and dried by filtration through activated alumina or distilled under nitrogen over an appropriate drying agent. Bu4NPF6 (GFS Chemicals) was recrystallized multiple times from CH2Cl2 solution by the addition of hexane. Chromatography was performed using Siliflash P60 from Silicycle (230–400 mesh). ESI-MS data for compounds were acquired using a Waters Micromass Quattro II spectrometer. 1H NMR spectra (500 MHz) were referenced to residual solvent relative to TMS. 31P{1H} NMR spectra (202 MHz) were referenced to an external 85% H3PO4. FT-IR spectra were recorded on a Perkin-Elmer 100 FT-IR spectrometer. Crystallographic data for compounds [1a]0, [1c]0, [2b]0,
and [2b]BF_4 were collected using a Siemens SMART diffractometer equipped with a Mo-K_\alpha source (\(\lambda = 0.71073\ \text{Å}\)), and crystallographic data for compounds [1a]BF_4 and [1d]BF_4 were collected using a Bruker D8 Venture diffractometer equipped with a Mo-K_\alpha microfocus-source and a Photon 100 detector. X-band EPR Spectra were recorded on a Varian E-line 12” Century Series X-band continuous-wave spectrometer. Q-Band ENDOR spectra were recorded on a Bruker Elexsys E580 Q-band pulse EPR spectrometer using a homebuilt pulse Q-band ENDOR resonator and an ENI 300L RF amplifier. Cryogenic temperatures were achieved using a Cryogenic Ltd closed cycle cryostat. ENDOR simulations and spectral fits were performed using EasySpin (“salt”) and homwritten scripts in Matlab. Mössbauer spectra were recorded on an alternating constant-acceleration spectrometer. The minimum experimental line width was 0.24 mm/s (full width at half-maximum). The sample temperature was maintained constant in either an Oxford Variox or an Oxford Mössbauer-Spectromag cryostat. The latter is a split-pair conducting magnet system for applying fields of up to 8 T to the samples that can be kept at temperatures in the range 1.5–250 K. The field at the sample is perpendicular to the \(\gamma\)-ray beam. Isomer shifts are quoted relative to metallic iron at 300 K. Mössbauer spectra were simulated with a home-written spin-Hamiltonian program based on the usual nuclear Hamiltonian formalism.

\[ [(\text{C}_5\text{H}_5)_3\text{Ni}_2]\text{BF}_4, \quad [(\text{MeC}_5\text{H}_4)_3\text{Ni}_2]\text{BF}_4, \quad \text{and} \quad [(\text{C}_5\text{D}_5)_3\text{Ni}_2]\text{BF}_4. \]

In a modification of Werner's procedure, a solution of nickelocene (2.08 g, 11.0 mmol) in Et_2O (150 mL) was treated dropwise with an excess of HBF_4·Et_2O (2 mL, 13.8 mol). Over the course of several minutes, a purple solid precipitated from a pale yellow. The solid was collected by filtration and washed thoroughly with ether. Yield: 2.07 g (95%). \(^1\text{H}\) NMR (nitromethane-\(d_3\)): \(\delta 5.48\) (10H, s), \(4.80\) (5H, s). \(^{13}\text{C}\{^1\text{H}\} \text{NMR (nitromethane-}d_3\text{): } \delta 87.8, 56.6.\)
[(C₅H₅)Ni(pdt)Fe(dppe)(CO)]BF₄ ([1a]BF₄). To a solution of [(C₅H₅)₃Ni₂]BF₄ (498 mg, 1.24 mmol) in CH₃NO₂ (50 mL) was added a solution of Fe(pdt)(dppe)(CO)₂⁻¹ (771 mg, 1.25 mmol) in CH₃NO₂ (25 mL). After 30 min., the slurry converted to a dark-red solution. Solvent was removed under reduced pressure, and the residue was washed several times with pentane until the filtrate was colorless. The residue was extracted into minimal volume of CH₂Cl₂ and subjected to column chromatography. Byproducts elute with 100% CH₂Cl₂. The product eluted with 90:10 CH₂Cl₂/THF as a dark-red band. The solvent was removed under reduced pressure to yield a dark-red solid. Yield: 841 mg (85%). ¹H NMR (acetone-δ₆): 8.01 (m, 4H), 7.5 (m, 16H), 4.75 (s, 5H, C₅H₅), 3.83 (m, 2H, PC₆H₄C₆H₄P), 3.62 (m, 2H, PCH₃CH₂P), 2.83 (m, 3H), 2.43 (m, 2H), 1.66 (m, 1H, S₂CH₂CH₂SH₂S₂). ³¹P{¹H} NMR (acetone-δ₆): δ 77.1. ESI-MS: m/z 711 ([M]+), 683 ([M–CO]⁺). IR (CH₂Cl₂): νCO = 1943 cm⁻¹ (THF). Anal. Calcd for C₃₅H₃₅FeNiOP₂S₂: C, 52.61 (52.67); H 4.42 (4.70). Single crystals of [1]BF₄ were obtained by vapor diffusion of pentane into THF solutions.

[(C₅H₅)Ni(pdt)Fe(dpvp)(CO)]BF₄ ([1b]BF₄). Compound [1b]BF₄ was prepared in a similar fashion to [1]BF₄ using Fe(pdt)(dpvp)(CO)₂⁻¹ as the iron reagent. Yield: 66%, dark-red powder. ¹H NMR (CH₂Cl₂): 8.44 (m, 2H, PCH₃CH₂P), 7.91 (br, 4H), 7.54 (s, br, 4H), 7.45 (s, br 12H), 4.31 (s, 5H, C₅H₅), 2.77 (m, 3H) 2.36 (m, 2H), 1.69 (m, 1H, S₂CH₂CH₂SH₂S₂). ³¹P{¹H} NMR (CH₂Cl₂): δ 83.8. ESI-MS: m/z 709 ([M]+), 681 ([M–CO]⁺). IR (CH₂Cl₂): νCO = 1951 cm⁻¹.

[(C₅H₅)Ni(Me₂pdt)Fe(dppe)(CO)]BF₄ ([1c]BF₄). Compound [1c]BF₄ was prepared in a similar fashion to [1]BF₄ using Fe(Me₂pdt)(dppe)(CO)₂⁻¹ as the iron source. Yield: 77%, dark-red powder. ¹H NMR (CH₂Cl₂): 7.77 (m, 4H), 7.46 (m, 16), 4.46 (s, 5H, C₅H₅), 3.65 (m, 2H, PCH₃CH₂P), 3.11 (m, 2H PCH₃CH₂P), 2.71 (m, 2H, S₂CH₃CMeo₂CH₂SH₂S₂), 2.17 (m, 2H, S₂CH₂CMeo₂CH₂SH₂S₂), 1.50 (s, 3H, S₂CH₂CMeo₂CH₂SH₂S₂), 1.19 (s, S₂CH₂CMeo₂CH₂SH₂S₂). ³¹P{¹H} NMR
Single crystals of $[1c]$BF$_4$ were obtained by evaporation of concentrated THF solutions.

$[(\text{CH}_3\text{C}_5\text{H}_4)\text{Ni(pdt)Fe(dppe)(CO)}]BF_4$ ($[1\text{d}]$BF$_4$). Compound $[1\text{d}]$BF$_4$ was prepared in a similar fashion to $[1\text{a}]$BF$_4$ using the $[(\text{CH}_3\text{C}_5\text{H}_4)_3\text{Ni}_2]$BF$_4$ as the nickel source and Fe(pdt)(dppe)(CO)$_2$ as the iron source. Yield: 81%, dark-red powder. $^1$H NMR (CH$_2$Cl$_2$): 7.80 (br, 4H) 7.52 (br, 4H) 7.43(br, 12H), 4.67 (br, 1H) 3.91 (br, 1H), 3.82 (s, 3H), 3.43 (m, 2H, PCH$_2$CH$_2$P), 3.05 (m, 2H, PCH$_2$CH$_2$P), 3.05 (m, 3H), 2.31 (m, 2H), 1.89 (br, 2H), 1.75 (m, 1H, S$_2$CH$_2$CH$_2$S). $^{31}$P{$_1$H} NMR (CD$_2$Cl$_2$): $\delta$ 79.2. ESI-MS: $m$/Z 725 ([M$^+$]), 697 ([M$-$CO$^+$]). IR (CH$_2$Cl$_2$): $\nu_{CO} = 1943$ cm$^{-1}$.

$[(\text{C}_5\text{D}_5)\text{Ni(pdt)Fe(dppe)(CO)}]BF_4$ ($[1\text{a}]$BF$_4$-$d_5$). Compound $[1\text{a}]$BF$_4$-$d_5$ was prepared in a similar fashion to $[1\text{a}]$BF$_4$ using $[(\text{C}_5\text{D}_5)_3\text{Ni}_2]$BF$_4$ as the nickel source. The product was characterized by ESI-MS, which indicated >95% deuteration (see SI).

**Fe(SPh)$_2$(CO)$_2$(dppv).** This compound was synthesized similarly to the published procedure.$^{55}$ To a mixture of FeSO$_4$·7H$_2$O (1.112g, 4.0 mmol) and dppv (1.584 g, 4.0 mmol) in MeOH (80 mL), a solution of PhSNa (1.056 g, 8.0 mmol) in MeOH (20 mL) was added dropwise with stirring under the presence of CO. After stirring for 6 h at room temperature, solvent was removed under reduced pressure. The residue was extracted with CH$_2$Cl$_2$ (40 mL). After filtration to remove Na$_2$SO$_4$, the extract was concentrated (15 mL), diluted with hexane (50 mL). Cooled this mixture at -20 °C for 12h gave a red solid. Yield: 2.23 g (77%). $^1$H NMR (CD$_2$Cl$_2$): $\delta$ 6.85–8.03 (32H, m). $^{31}$P{$_1$H} NMR (CD$_2$Cl$_2$): $\delta$ 81.9 (d, $J = 55$ Hz), 61.1 (d, $J = 55$ Hz). IR (CH$_2$Cl$_2$): $\nu_{CO} = 2023, 1978$ cm$^{-1}$.

$[(\text{C}_5\text{H}_5)\text{Ni(SPh)}_2\text{Fe(dppe)(CO)}]_2$BF$_4$ ($[2\text{bCO}]$BF$_4$). To a solution of $[(\text{C}_5\text{H}_5)_3\text{Ni}_2]$BF$_4$ (200 mg, 0.5 mmol) in CH$_3$NO$_2$ (15 mL) was added a solution of Fe(SPh)$_2$(dppv)(CO)$_2$ (364 mg,
0.5 mmol) in CH$_3$NO$_2$ (10 mL). After 30 min, the slurry converted to a dark-red solution. Solvent was removed under reduced pressure, and the residue was washed with pentane until the filtrate was colorless. The residue was extracted into minimal volume of CH$_2$Cl$_2$ and subjected to column chromatography. After eluting impurities with 100% CH$_2$Cl$_2$, the product eluted with 90:10 CH$_2$Cl$_2$/THF as a dark-red band. The solvent was removed under reduced pressure to yield a dark-red solid. Yield: 380 mg (81%). $^1$H NMR (CD$_2$Cl$_2$): δ 6.85–8.57 (32H, m), δ 4.46 (5H, s). $^{31}$P{$^1$H} NMR (CD$_2$Cl$_2$): δ 75.9 (d, $J$ = 45 Hz), 64.2 (d, $J$ = 45 Hz). ESI-MS: $m$/$Z$ 849 ([M]$^+$), 821 ([M–CO]$^+$), 793 ([M–2CO]$^+$). IR (CH$_2$Cl$_2$): $\nu$$_{\text{CO}}$ = 2042, 2002 cm$^{-1}$. Anal. Calcd for C$_{45}$H$_{37}$BF$_4$FeNiO$_2$P$_2$S$_2$·CH$_2$Cl$_2$: C, 54.05 (53.69); H, 3.85 (3.72). Single crystals of [2bCO]BF$_4$ were obtained by solvent diffusion of pentane into CH$_2$Cl$_2$ solutions.

[(C$_5$H$_5$)Ni(SPh)$_2$Fe(dppe)(CO)]BF$_4$ ([2b]BF$_4$). To a solution of [2bCO]BF$_4$ (280 mg, 0.3 mmol) in CH$_2$Cl$_2$ (30 mL) was added Me$_3$NO·2H$_2$O (34 mg, 0.3 mmol). After 1 h, the solvent was removed under reduced pressure, and the residue was washed several times with pentane. The residue was extracted into minimal volume of CH$_2$Cl$_2$ (5 mL), pentane (30 mL) was added and the mixture was cooled to -20$^\circ$C. The product was collected as a black powder. Yield: 200 mg (74%). $^{31}$P{$^1$H} NMR (CD$_2$Cl$_2$): δ 78.8. ESI-MS: $m$/$Z$ 821 ([M]$^+$), 793 ([M–CO]$^+$). IR (CH$_2$Cl$_2$): $\nu$$_{\text{CO}}$ = 1956 cm$^{-1}$. Anal. Calcd for C$_{44}$H$_{37}$BF$_4$FeNiOP$_2$S$_2$·0.5CH$_2$Cl$_2$: C, 56.16 (56.08); H, 4.02 (4.19).

(C$_5$H$_3$)Ni(pdt)Fe(dppe)(CO) ([1a]$^0$). To a stirred solution of [1a]BF$_4$ (21.2 mg, 26.5 µmol) in THF (3 mL) was added cobaltocene (5.8 mg, 30.7 µmol) in THF (3 mL). The solution immediately darkened and solid [(C$_5$H$_3$)$_2$Co]BF$_4$ precipitated. The solvent is removed under reduced pressure and redissolved in toluene (2 mL). The slurry was filtered through diatomaceous earth to yield a homogeneous, brown-black solution. The product is crystallized upon addition of
pentane (15 mL). Yield: 17.0 mg (90%), black powder. Single crystals were obtained by vapor diffusion of pentane into concentrated toluene solutions. IR (CH$_2$Cl$_2$): $\tilde{\nu}_{CO} = 1901$ cm$^{-1}$. Anal. Calcd. for C$_{35}$H$_{35}$FeNiOP$_2$S$_2$ (found): C, 59.02 (58.72); H 4.95 (4.65).

(C$_5$H$_5$)Ni(pdt)Fe(dppv)(CO) ([1b]$^0$). Compound [1b]$^0$ was prepared in a similar fashion to [1a]$^0$. Yield: 86%, black powder. IR (CH$_2$Cl$_2$): $\tilde{\nu}_{CO} = 1903$ cm$^{-1}$.

(C$_5$H$_5$)Ni(Me$_2$pdt)Fe(dppe)(CO) ([1c]$^0$). Compound [1c]$^0$ was prepared in a similar fashion to [1a]$^0$. Yield: 87%, black powder. IR (CH$_2$Cl$_2$): $\tilde{\nu}_{CO} = 1897$ cm$^{-1}$.

(CH$_3$C$_5$H$_4$)Ni(pdt)Fe(dppe)(CO) ([1d]$^0$). Compound [1d]$^0$ was prepared in a similar fashion to [1a]$^0$. Yield: 80%, black powder. IR (CH$_2$Cl$_2$): $\tilde{\nu}_{CO} = 1898$ cm$^{-1}$.

(C$_5$D$_5$)Ni(pdt)Fe(dppe)(CO) ([1a-d$_5$]$^0$). Compound [1a-d$_5$]$^0$ was prepared in a similar fashion to [1a]$^0$.

(C$_5$H$_5$)Ni(SPh)$_2$Fe(dppv)(CO) ([2b]$^0$). Compound [2b]$^0$ was prepared in a similar fashion to [1a]$^0$. Yield: 61%, black powder. IR (CH$_2$Cl$_2$): $\tilde{\nu}_{CO} = 1915$ cm$^{-1}$. Single crystals of [2b]$^0$ were obtained by vapor diffusion of pentane into concentrated toluene solutions.

6.12. Computational Methods

DFT results reported herein were performed using the BP86 density functional.$^{56,57}$ Additional benchmarking was performed using the B3P86,$^{56,58}$ BP86-D2,$^{56,57,59}$ M06-L,$^{60}$ and $\omega$B97XD$^{59,61-63}$ functionals and is presented in the Supporting Information (Tables S1-S7). The DFT calculations were performed with the Stuttgart pseudopotential and associated basis set of Preuss and co-workers (SDD)$^{64}$ for the Ni and Fe centers, the 6-31G** basis set$^{65}$ for the $\mu$-H ligand, and the 6-31G* basis set$^{66,67}$ for all other atoms. The starting geometries for [1a-d$^+$], [1a]$^0$, and [2b]$^0$ were obtained from their respective crystal structures. The starting coordinates for species without crystal structures were obtained by manually altering the most closely related
crystal structure. Solvation free energies were calculated in dichloromethane (CH$_2$Cl$_2$) using the conductor-like polarizable continuum model (C-PCM)\textsuperscript{68,69} with the Bondi atomic radii\textsuperscript{70} and including the nonelectrostatic contributions of dispersion,\textsuperscript{71,72} repulsion,\textsuperscript{71,72} and cavitation energies.\textsuperscript{73}

For the results presented in the main paper, the geometry optimizations were performed in the gas phase. However, geometry optimizations were also performed in solution and were found to be consistent with the gas phase optimizations. A comparison of the structures optimized in the gas phase and solution phase is provided in the Supporting Information (Tables S4-S7). In all cases, the minimum-energy structures were confirmed to have no imaginary frequencies. The $\nu_{CO}$ frequencies were calculated within the harmonic model. Typically, DFT is more reliable for calculating changes in frequencies than absolute frequencies, and often scaling factors dependent on the functional are used.\textsuperscript{74} Herein, however, the absolute $\nu_{CO}$ values are reported because the BP86 functional gives excellent agreement with the experimental values without the application of scaling factors. Nonetheless, $\Delta\nu_{CO}$ values calculated using different functionals are provided in the Supporting Information (Table S3). The transition states (TSs) for isomerization of the Fe(dppe)(CO) subunit were identified using the synchronous transit-guided quasi-Newton method,\textsuperscript{75,76} and the TSs were confirmed to have only one imaginary frequency. Each TS was verified to lead to the relevant dibasal or apical-basal isomers by following the intrinsic reaction coordinate (IRC) using the local quadratic approximation\textsuperscript{77,78} for 5–10 steps in the forward and reverse directions and subsequently optimizing the geometries.

Thermochemical data were calculated at $T = 298.15$ K. All calculated free energies included zero-point energy, entropic contributions, and solvation effects. The reaction free energies ($\Delta G^\circ$) and free energy barriers ($\Delta G^\ddagger$) associated with the isomerization of the
Fe(dppe)(CO) subunit in solution were calculated from the optimized geometries and TSs. The relative reduction potentials were calculated from the corresponding reaction free energies using methodology described elsewhere.\textsuperscript{79-82} Chemical bonding analysis was performed using the Pipek-Mezey localization criteria\textsuperscript{43}, using keyword IOp(4/9=20212), and Natural Bond Orbitals (NBO).\textsuperscript{44} All calculations were performed using the Gaussian 09 electronic structure program.\textsuperscript{83} Structures and energies of the systems studied herein are provided in the Supporting Information (Tables S12-S73).

6.13. References


Glendening, E. D.; Reed, A. E.; Carpenter, J. E.; Weinhold, F. *NBO Version 3.1*; University of Wisconsin System: Madison, WI, 1996.


Chapter 7. Quinone 1 e\textsuperscript{−} and 2 e\textsuperscript{−}/2 H\textsuperscript{+} Reduction Potentials\textsuperscript{*,†}

7.1. Abstract

Quinones participate in diverse electron transfer and proton-coupled electron transfer processes in chemistry and biology. To understand the relationship between these redox processes, an experimental study was carried out to probe the 1 e\textsuperscript{−} and 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potentials of a number of common quinones. The results reveal a non-linear correlation between the 1 e\textsuperscript{−} and 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potentials. This unexpected observation prompted a computational study of 134 different quinones, probing their 1 e\textsuperscript{−} reduction potentials, pK\textsubscript{a} values, and 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potentials. The density functional theory calculations reveal an approximately linear correlation between these three properties and an effective Hammett constant associated with the quinone substituent(s). However, deviations from this linear scaling relationship are evident for quinones that feature intramolecular hydrogen bonding in the hydroquinone, halogen substituents, charged substituents, and/or sterically bulky substituents. These results, particularly the different substituent effects on the 1 e\textsuperscript{−} versus 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potentials, have important implications for designing quinones with tailored redox properties.

7.2. Introduction

Quinones are ubiquitous redox-active organic molecules that play important roles in processes ranging from photosynthesis and aerobic respiration in biology\textsuperscript{1-3} to chemical oxidation methods and industrial production of hydrogen peroxide.\textsuperscript{4-8} Quinones are also the focus of increased attention in energy storage and conversion applications, including redox flow batteries,

\textsuperscript{†} Author contributions: Colin W. Anson, Andrew C. Cavell, and Prof. Shannon S. Stahl were responsible for all experiments and associated analyses in this chapter.
dye-sensitized solar cells, and water-splitting devices.\textsuperscript{9-12} Studies in our labs have focused on the use of quinones as electron-proton transfer mediators for $O_2$ reduction\textsuperscript{13,14} and as aerobic oxidation catalysts that resemble oxidase enzymes with quinone active sites.\textsuperscript{15-17,18-23}

Quinones participate in a variety of redox processes (Scheme 7.1). For example, the applications noted above feature single electron transfer, proton-coupled electron transfer (PCET), and hydride transfer reactivity. An understanding of the factors that influence the different redox properties of quinones could contribute significantly to the design of new quinones for various applications. The present study was motivated by an interest in the relationship between electron transfer and PCET reactions of quinones.

**Scheme 7.1.** Quinone Electron, Proton, and Proton-Coupled Electron Transfer Processes

Experimental $1\,e^-\text{ and } 2\,e^-/2\,H^+$ reduction potentials of many quinones have been reported previously.\textsuperscript{24-29} However, gaps are present in the data, and many of the values were not determined under a uniform set of conditions. This deficiency is addressed in the present study through the acquisition of a broad set of self-consistent experimental data, thereby facilitating systematic analysis of quinone redox properties. The results described below reveal a heretofore unrecognized non-linear correlation between $1\,e^-$ and $2\,e^-/2\,H^+$ quinone reduction potentials. These unexpected observations prompted computational studies of a much larger set of quinones to explore the basis for this behavior. Overall, quinone redox behavior is found to be well described by a Hammett-
based linear scaling relationship. However, a more thorough assessment of the experimental and calculated reduction potentials and $pK_a$ values of the different quinones draws attention to four different types of substituents that contribute to significant deviations from the global trend: hydrogen-bond acceptors, halogens, charged fragments, and sterically bulky groups. These results provide key insights into quinone structure-activity relationships that may be used in the design and tuning of new quinone-based catalysts, reagents, and/or mediators.

7.3. Experimental Details

All chemicals were used as received unless otherwise stated. 1,4-benzoquinone (1) was purified by sublimation. 2-tert-butyl-1,4-benzoquinone (4), 2,3-dimethyl-1,4-benzoquinone (25), 2,3,5-trimethyl-1,4-benzoquinone (91), 3,5-di-tert-butyl-1,2-benzoquinone, and 4-tert-butyl-1,2-benzoquinone were oxidized from the corresponding hydroquinones following literature procedure.$^{22,30}$

Electrochemical experiments were performed with a standard 3-electrode set-up, using a glassy carbon electrode that was polished with alumina before each experiment, a Ag/Ag$^+$ reference electrode (for non-aqueous measurements) or a Ag/AgCl reference electrode (for aqueous measurements), and a Pt wire counter electrode. To reference the Ag/Ag$^+$ potentials to Fc/Fc$^+$, ferrocene was added to the solution and a cyclic voltammetry (CV) scan was taken to measure this potential. To reference to NHE, +197 mV was added to the Ag/AgCl potentials measured experimentally or ferroin was added and the midpoint potential referenced to NHE. All experimental measurements were obtained using a BASi EC Epsilon potentiostat (note: reported potentials have been corrected for iR drop). A scan rate of 50 mV/s was used in all CV experiments. Experimental CVs are included in the Supporting Information (Figures S1–30).
7.4. Computational Details

The computations were performed with density functional theory (DFT) using the B3LYP functional and the 6-31++G** basis set. The structures were optimized in the solution phase (water or acetonitrile) using the conductor-like polarizable continuum model (C-PCM) with Bondi atomic radii, and including the nonelectrostatic contributions of dispersion, repulsion, and cavitation energies. Thermochemical data were calculated at $T = 298.15$ K, and all calculated free energies included zero-point energy, entropic contributions, and solvation effects. The relative reduction potentials and pK$_a$ values were calculated from the corresponding reaction free energies using methodology described elsewhere. Additional benchmarking was performed using the B3P86, BP86, M06-2X, and ωB97XD functionals (Figure S31). All calculations were performed using the Gaussian 09 electronic structure program.

The reduction potentials and pK$_a$ values were calculated relative to the corresponding values for the reference species (1) using isodesmic reactions. The isodesmic reactions used for this procedure are provided in the SI on pages S14. This procedure is equivalent to determining the difference between the experimental and calculated values for the reference species and adding this correction to all calculated values. The resulting reduction potentials and pK$_a$ values have been shown to be quantitatively accurate for species that are structurally similar to the reference species.

7.5. Calculation of the 2 e$^-$/2 H$^+$ Reduction Potentials

The electron and proton transfer reactions involved in quinone electrochemistry are summarized in the nine-membered thermodynamic cycle shown in Scheme 7.1. The hydroquinone (H$_2$Q) may be generated from the quinone (Q) by proceeding sequentially through two reduction
and two protonation steps. The cationic species are high-energy species not involved in the PCET reactions of interest and therefore were not investigated. Using this thermodynamic cycle, it was possible to determine the reduction potential for the overall 2 e⁻/2 H⁺ process for each quinone in terms of individual reduction and protonation steps:

\[
E^\circ(Q / Q^{2-}) = \frac{1}{2} \left[ E^\circ(Q / Q^\bullet) + E^\circ(Q^\bullet / Q^{2-}) \right]
\]

\[
E^\circ(Q, 2H^+ / H_2Q) = E^\circ(Q / Q^{2-}) + \frac{RT}{2F} \left[ pK_a(H_2Q) + pK_a(HQ^-) \right]
\]

The 2 e⁻/2 H⁺ reduction potential is calculated by considering the following series of four steps (cf. Scheme 7.1): (1) reduction of the neutral quinone (Q) to produce the radical anion (Q⁻⁻), (2) reduction of the Q⁻⁻ to produce the quinone dianion (Q^{2⁻}), (3) protonation of the Q^{2⁻} to produce the protonated anion (HQ⁻⁻), and (4) protonation of the HQ⁻⁻ to produce the neutral hydroquinone (H₂Q). The reduction potentials and pKₐ values associated with these four steps are calculated independently and are combined to generate the 2 e⁻/2 H⁺ reduction potential using Eq. (2).

Overall, the examination of six different states and two to four isomers of 134 quinones required ~2000 DFT geometry optimizations and free energy calculations. Only thermodynamic properties were calculated in this work; consideration of kinetic properties would require calculation of free energy barriers for the proton transfer reactions.

**7.6. Experimental Measurements on Quinones**

Fifteen common 1,4-benzoquinone derivatives were selected for preliminary study (Chart 1), and cyclic voltammetry was used to measure the 1 e⁻ and 2 e⁻/2 H⁺ reduction potentials for each derivative (see Table S1 in the Supporting Information). The 1 e⁻ reduction potentials were measured under aprotic conditions [0.1 M (Bu₄N)PF₆ in MeCN], while the 2 e⁻/2 H⁺ potentials were determined in acidic aqueous solution (1 M p-TsOH). Reversible (or quasi-reversible)
voltammograms were observed under both conditions for most derivatives, and in all cases it was possible to identify an appropriate mid-point potential.\textsuperscript{57,58} \textbf{Error! Hyperlink reference not valid.} Our measured reduction potentials correlate well with available values in the literature (cf. Tables S5 and S9) but, importantly, provide a self-consistent set of data for analysis.

Our measured reduction potentials correlate well with available values in the literature (cf. Tables S5 and S9) but, importantly, provide a self-consistent set of data for analysis.

\emph{Figure 7.1.} Plot of experimental 1 e\textsuperscript{−} (MeCN) versus 2 e\textsuperscript{−}/2 H\textsuperscript{+} (1 M p-TsOH) reduction potentials for the 1,4-benzoquinone derivatives depicted in Chart 8.1.

The plot in Figure 7.1 depicts the relationship between the 1 e\textsuperscript{−} and the 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potentials for the quinones shown in Chart 1. This plot reveals the absence of a linear correlation between these values for the full set of quinones. On the other hand, systematic trends are evident within subsets of related quinones, including (1) the alkyl- and alkoxy-substituted quinones (blue) and (2) halogenated quinones (red). Chloranil (2,3,5,6-tetrachloro-1,4-benzoquinone) is a classical "high-potential" quinone that finds widespread use as an oxidant in organic chemistry.\textsuperscript{6} It exhibits a significantly higher (i.e., less negative) 1 e\textsuperscript{−} reduction potential than 1,4-benzoquinone (BQ) ($\Delta E_{\text{mp}} = 525$ mV); however, the 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potential of chloranil is only slightly higher (i.e., more positive) than that of BQ ($\Delta E_{\text{mp}} = 56$ mV). In comparison, duroquinone (2,3,5,6-
tetramethyl-1,4-benzoquinone) has both a significantly lower 1 e\textsuperscript{−} and 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potential than BQ (\(\Delta E_{\text{mp}}\) = 331 mV and 227 mV, respectively). To our knowledge, this unusual result and the factors contributing thereto have not been highlighted or addressed previously in the literature. In order to gain insights into this behavior, computational studies were initiated to probe the 1 e\textsuperscript{−} and 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potentials for an even wider range of quinones.

7.7. Benchmarking Computational Methods for Quinones

The calculated reduction potentials for numerous quinones were compared to experimental measurements (cf. Chart 8.1 and Chart S1 and Tables S2 and S3 in the Supporting Information). As shown in Figure 7.2, excellent agreement was observed between the calculated and experimental 1 e\textsuperscript{−} and 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potentials. The most significant deviations were observed for anthraquinone, naphthoquinone, and \textit{ortho}-quinone derivatives (gray circles) that are structurally quite different from the reference species (see Supporting Information for details). These deviations decrease if a different reference species (i.e., a different correction constant) is used for each type of structurally distinct quinone. Similarly good agreement between the calculated and experimental reduction potentials was observed when different functionals were used (Figure S31).
Figure 7.2. Plots correlating calculated and experimental reduction potentials for diverse quinones (see Chart S1 and Tables S2 and S3 in the Supporting Information for details). Gray circles indicate anthraquinone, naphthoquinone, and ortho-quinone derivatives that are structurally quite different from the reference species (I).

7.8. Analysis of Linear Correlations

Following these benchmarking studies, we compiled a list of systematically varied mono-, di-, tri-, and tetra-substituted quinones (Chart 8.2). For each quinone derivative, the aqueous reduction potentials and $pK_a$ values involved in the thermodynamic cycle in Scheme 7.1 were then calculated. From these data, the reduction potentials associated with the corresponding PCET reactions (diagonal steps in Scheme 7.1) were determined. The full set of data is compiled in the Supporting Information (Table S11).
Linear correlations between thermodynamic properties have provided insights into many catalytic systems, and the $pK_a$ values and reduction potentials of quinone derivatives have been shown previously to correlate with Hammett parameters. Therefore, the $pK_a$ and $E^o$ data acquired for the 134 quinones were plotted as a function of an effective Hammett constant, defined as the sum of the substituents’ Hammett constants ($\Sigma \sigma$). As in previous studies, the $\sigma_m$ values were used for halogen substituents to better account for inductive effects, the $\sigma_p$– values were used for substituents capable of conjugating with the quinone $\pi$ system, and the $\sigma_p$ values were used for all other substituents.

Figure 7.3 illustrates that the $2 e^-/2 H^+$ potentials are approximately linearly correlated with the effective Hammett constant. The reduction potentials and $pK_a$ values underlying the $2 e^-/2 H^+$ potentials also exhibit approximately linear correlations with the $\Sigma \sigma$ parameter (see Table S11 in the Supporting Information for a full compilation of the individual reduction potentials and $pK_a$ values).
values). As expected, more electron-donating substituents (i.e., lower $\Sigma \sigma$) decrease the 1 e$^-$ reduction potentials because they make it less favorable to add an electron but increase the $pK_a$ values because they enhance the basicity of the quinone oxygen atoms. These offsetting trends are evident in a "leveling" of the PCET reduction potentials,\textsuperscript{68} manifested by a less steep dependence of the PCET reduction potentials relative to the 1 e$^-$ reduction potentials on the effective Hammett constant.

\begin{figure}[h]
\centering
\includegraphics[width=0.5\textwidth]{figure7_3.png}
\caption{Plot of the 2 e$^-$/2 H$^+$ reduction potentials (filled black circles), the sum of the $pK_a$ values (filled gray circles), and the average of the two 1 e$^-$ reduction potentials (open gray circles) as functions of the sum of the Hammett constants ($\Sigma \sigma$) for all substituted quinones in Chart 2. As discussed in the text, intramolecular proton transfer (PT) may occur in the Q$^-$ and Q$^{2-}$ states for the SH, OH, or CO$_2$H substituents, but is not included in the calculated values shown in this figure. This intramolecular PT does not impact the 2 e$^-$/2 H$^+$ reduction potentials but does impact the breakdown into the individual contributions from reductions and external protonations.}
\end{figure}

The full set of data for these 134 quinones provided an opportunity to conduct a thorough assessment of 1 e$^-$ versus 2 e$^-$/2 H$^+$ reduction potentials (Figure 7.4), analogous to the preliminary experimental data shown in Figure 7.1. The data in Figure 7.4 exhibit a reasonably linear overall correlation between the 1 e$^-$ and 2 e$^-$/2 H$^+$ reduction potentials; however, deviations from the linear trendline can be quite significant. For example, as noted above, chloranil (122) and benzoquinone (1) exhibit >500 mV difference in their 1 e$^-$ potentials, but they have very similar and even slightly inverted 2 e$^-$/2 H$^+$ potentials. Conversely, other quinone structures exhibit similar 1 e$^-$ potentials.

193
but significantly different \(2\, e^{-}/2\, H^+\) potentials. For example, the 2,5-bis(methoxycarbonyl)-substituted quinone 61 has a 1\,e\textsuperscript{−} reduction potential 74 mV lower than that of chloranil, but its 2\,e\textsuperscript{−}/2\,H\textsuperscript{+} potential is nearly 400 mV higher.

**Figure 7.4.** Plot of the 1\,e\textsuperscript{−} reduction potentials versus the 2\,e\textsuperscript{−}/2\,H\textsuperscript{+} reduction potentials for the quinones in Chart 2, with colors identifying the physical basis for deviations from the global linear trend. The line was generated by a linear fit to only the gray data points.

### 7.9. Analysis of Deviations from Linear Correlations

Insights into the origin of deviations from the linear global trendline in Figure 7.4 may be gained by reassessment of the Hammett correlations shown in Figure 7.3. Four different classes of substituents were found to contribute significantly to the deviations: (1) hydrogen-bonding and conjugating substituents, (2) halogen substituents, (3) charged substituents, and (4) sterically bulky substituents (see Table S10 of the Supporting Information for a listing of the specific identities of quinones within each class). A variation of Figure 7.3, in which the quinones in each of these classes are color-coded as in Figure 7.4 and the 2\,e\textsuperscript{−}/2\,H\textsuperscript{+} reduction potentials are separated from the contributing 1\,e\textsuperscript{−} reduction potentials and \(pK_a\) values, is provided in Figure 7.5. The influence of the different classes of substituents on the 1\,e\textsuperscript{−} reduction potential, the \(pK_a\) value, and the 2\,e\textsuperscript{−}/2\,H\textsuperscript{+} reduction potential is summarized in Table 7.1 and elaborated below.
Figure 7.5. Correlations between the average $1\,e^{-}$ reduction potentials (a), the sum of the two $pK_a$ values (b), and the $2\,e^{-}/2\,H^+$ reduction potentials (c) of quinones 1–134 and their effective Hammett constants ($\Sigma\sigma$). The gray data points are used to generate the linear fits, and the colored data points were found to exhibit deviations from these linear fits and are defined in the legend, with the specific substituents in each group given in the SI (Table S10). For each substituent, its Hammett constant can be used to determine the expected value of a given property on the line, and the shift from the linear Hammett correlation is the difference between the calculated value for this substituent and the expected value. These shifts are summarized in Table 7.1. As discussed in the text, intramolecular PT may occur in the $Q^-$ and $Q^{2-}$ states for the SH, OH, or CO$_2$H substituents, but this contribution is not included in the calculated values shown in this figure. This intramolecular PT does not impact the $2\,e^{-}/2\,H^+$ reduction potentials but does impact the individual contributions from reductions and external protonations. Because this intramolecular PT occurs spontaneously for the dianion in the hydrogen-bonded conformation for the CO$_2$H substituents (Figure S36), these species are not included in this figure to avoid inconsistencies in the analysis. The analogous figure for the $1\,e^{-}$ reduction potentials, including the CO$_2$H substituents, is given in Figure S37.
Table 7.1. Analysis of Substituent Effects on 1 e− Reduction Potentials, pKₐ values, and 2 e−/2 H⁺ Reduction Potentials Relative to the Linear Hammett Trends in Figure 7.5 and the 2 e−/2 H⁺ Potentials in Figure 7.4

<table>
<thead>
<tr>
<th>Substituent Type</th>
<th>1 e− Reduction Potential</th>
<th>pKₐ Values</th>
<th>2 e−/2 H⁺ Reduction Potential</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen bond accepting (H₂Q) and conjugating</td>
<td>Less positive</td>
<td>Increase</td>
<td>More positive</td>
</tr>
<tr>
<td>Hydrogen bond donating (Q) and conjugating</td>
<td>More positive</td>
<td>Increase</td>
<td>More positive</td>
</tr>
<tr>
<td>Halogen</td>
<td>No shift</td>
<td>Decrease</td>
<td>Less positive</td>
</tr>
<tr>
<td>Charged and hydrogen bonding</td>
<td>Less positive</td>
<td>Increase</td>
<td>More positive</td>
</tr>
<tr>
<td>Sterically bulky</td>
<td>More positive</td>
<td>Decrease</td>
<td>Less positive</td>
</tr>
</tbody>
</table>

The effect of each type of substituent is compared to the expected reduction potential or pKₐ, determined from the linear Hammett correlations (Figure 7.5 and Figure S37). For each substituent, its Hammett constant is used to determine the expected value of a given property on the line given in Figure 7.5 or Figure S37 for the 1 e− reduction potential, and the shift is the difference between the calculated value for this substituent and the expected value.

Figure 7.6. DFT-optimized structures of substituted benzoquinones: (a) 64, showing the hydrogen bonds between the substituent –OH and the C=O in the neutral quinone state (Q); (b) 64, showing the hydrogen bonds between the substituent C=O and the –OH in the hydroquinone state (H₂Q); (c) 61, showing the hydrogen bonds between the substituent C=O and the –OH in the hydroquinone state; (d) 21, showing the hydrogen bond between the SO₃− substituent and the –OH in the hydroquinone state; and (e) 114, showing the nonplanar six-membered ring and the out-of-plane proton positions to illustrate steric effects.
(1) Hydrogen-bonding and conjugating substituents (blue and orange points in Figure 7.4 and blue points in Figure 7.5). These substituents can act as either hydrogen bond donors (orange points) or acceptors (blue points). Due to complications with intramolecular proton transfer (PT), which is discussed further below, the substituents that can act as hydrogen bond donors are not analyzed in the context of Figure 7.5. Quinones with substituents that can act as hydrogen bond acceptors (e.g., R = CHO, COCH₃, CO₂CH₃, CO₂H) stabilize the –OH groups of hydroquinones (Figure 7.6b and 8.6c), leading to the higher-than-expected pKₐ values associated with this class of quinones in Figure 7.5b (blue points). In addition, many of the hydrogen-bonding substituents are carbonyl groups that conjugate with the quinone π system, thereby shifting the 1 e⁻ reduction potentials to less positive values than expected from the Hammett correlation in Figure 7.5a (blue points). The effect on pKₐ values is more pronounced than the effect on 1 e⁻ reduction potentials, resulting in a net shift of the 2 e⁻/2 H⁺ reduction potentials to more positive values, relative to the Hammett correlation in Figure 7.5c. A manifestation of this effect is the unusually high 2 e⁻/2 H⁺ potential for quinone 61 highlighted above and depicted in Figure 7.4. Quinones with substituents that can act as hydrogen bond donors (e.g., R = CO₂H) stabilize the quinone state (Figure 7.6a), leading to more positive than expected 1 e⁻ reduction potentials (orange points in Figure S37). In general, the substituents that can act only as hydrogen bond acceptors will lead to quinones that appear to the right of the trendline in Figure 7.4 (blue points) because the stronger substituent effect on the pKₐ values increases their relative 2 e⁻/2 H⁺ reduction potentials. The substituents that can act as both hydrogen bond donors and acceptors will lead to quinones that appear to the left of the trendline in Figure 7.4 (orange points) because of the dominant effect from the more positive 1 e⁻ reduction potentials.
(2) **Halogen substituents (red points in Figures 7.4 and 7.5).** Halogen substituents exhibit relatively little effect on the 2 e\textsuperscript{-}/2 H\textsuperscript{+} reduction potentials of quinones, as is evident from the nearly vertical trendline for halogenated quinones in Figures 7.1 and 7.4. The strong electron-withdrawing inductive effect of the halogens\textsuperscript{64,65,67} has a significant influence on both 1 e\textsuperscript{-} reduction potentials and pK\textsubscript{a} values; however, the influence is nearly equal and opposite, resulting in a net "redox leveling" effect for the 2 e\textsuperscript{-}/2 H\textsuperscript{+} potentials.\textsuperscript{68} The 1 e\textsuperscript{-} reduction potentials of halogenated quinones fall directly on the Hammett correlation in Figure 7.5a, while the pK\textsubscript{a} values show a more negative slope than the other quinones (inset of Figure 7.5b). The net impact of these effects is evident in a comparison of the properties of halogenated quinones and the parent benzoquinone (1): whereas halogenated quinones have 1 e\textsuperscript{-} reduction potentials that extend nearly 500 mV beyond that of 1, their 2 e\textsuperscript{-}/2 H\textsuperscript{+} reduction potentials differ from 1 by only ~100 mV. These observations suggest that the term "high-potential quinone" for chloranil and related halogenated quinones is only appropriate in the context of their 1 e\textsuperscript{-} reduction potentials.

(3) **Charged substituents (purple points in Figures 7.4 and 7.5).** Anionic substituents, such as sulfonate, can act as hydrogen bond acceptors and thereby increase the pK\textsubscript{a} values, similar to the other hydrogen-bonding substituents discussed above (Figure 7.6d). Electrostatic effects will also increase the pK\textsubscript{a} values because protonation is more favorable for negatively charged molecules (Figure 7.5b). On the other hand, electrostatic effects will destabilize reduced quinones and thereby shift the 1 e\textsuperscript{-} reduction potentials to more negative values (Figure 7.5a). The substituent effect on the pK\textsubscript{a} values is more pronounced than the effect on the 1 e\textsuperscript{-} reduction potentials, accounting for the net shift of the 2 e\textsuperscript{-}/2 H\textsuperscript{+} reduction potentials to more positive values, relative to the Hammett correlation in Figure 7.5c. As a corollary, charged substituents shift the 2 e\textsuperscript{-}/2 H\textsuperscript{+} reduction potentials to the right of the global trendline in Figure 7.4, similar to other
hydrogen-bonding substituents (cf. blue and purple circles in Figure 7.4). Electrostatic effects account for the more negative 1 e\textsuperscript{−} reduction potentials associated with quinones bearing negatively charged substituents, relative to those with other hydrogen-bonding substituents (cf. Figures 7.4 and 7.5a).

(4) Sterically bulky substituents (green points in Figures 7.4 and 7.5). Quinones with sterically bulky substituents, particularly those with large groups in adjacent 2,3- and/or 5,6-positions, can deviate significantly from the Hammett correlations. In most cases, they cause an increase in the 1 e\textsuperscript{−} reduction potentials to more positive values and a decrease in the pK\textsubscript{a} values relative to the linear Hammett correlations (Figures 7.5a and 7.5b). The 1 e\textsuperscript{−} reduction potentials could be affected by sterically induced distortion of the (hydro)quinone six-membered ring (Figure 7.6e). The decrease in pK\textsubscript{a} values could arise from steric clashing between the bulky substituents and the hydroquinone –OH. Because the pK\textsubscript{a} effects tend to be more significant, the 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potentials mostly fall below the linear Hammett correlation in Figure 7.5c. This behavior is also demonstrated by a shift of the 2 e\textsuperscript{−}/2 H\textsuperscript{+} reduction potentials to lower values, namely to the left of the trendline in Figure 7.4.

In addition to these four effects, some substituents (e.g., R = SH, OH, CO\textsubscript{2}H) can undergo intramolecular PT to the quinone oxygen in either the quinone radical anion or dianion. An analysis of this intramolecular PT is provided in Figure S36. These calculations indicate that intramolecular PT is not thermodynamically favorable in the quinone radical anions for most of these substituents, and, if it is favorable, the free energy difference is only ~1 kcal/mol. However, intramolecular PT is more thermodynamically favorable in the quinone dianions for quinones substituted with SH, OH, and CO\textsubscript{2}H. Thus, the 2 e\textsuperscript{−} reduction potentials for these systems may not be experimentally relevant and should be viewed only as analysis tools to examine the
individual contributions to the $2 \, e^-/2 \, \text{H}^+$ reduction potentials, which are not affected by intramolecular PT (Scheme 7.1).

Collectively, these results have important implications for the active and growing field of quinone redox chemistry. Whereas quinone-mediated reactions in organic chemistry often proceed via single electron transfer or hydride transfer (i.e., $2 \, e^-/1 \, \text{H}^+$), applications of quinones in energy storage and conversion processes often involve electrochemical $2 \, e^-/2 \, \text{H}^+$ PCET processes, which will be more strongly influenced by the proton affinity of the (reduced) quinone species. The term "high potential quinone" is commonly used in organic chemistry to describe DDQ, chloranil, and related halogenated quinones, but the results described herein reveal that this term could be misleading in the context of redox transformations that involve PCET processes.

The results summarized in Table 7.1 will be especially valuable in the design of quinones with tailored redox properties. Substituent effects on p$K_a$ values and $1 \, e^-$ reduction potentials typically have offsetting influences on the $2 \, e^-/2 \, \text{H}^+$ reduction potentials, as shown in Figure 7.3. Nevertheless, the deviations of the $2 \, e^-/2 \, \text{H}^+$ reduction potentials from systematic scaling relationships (i.e., Figures 7.4 and 7.5c) tend to follow the direction of the substituent effects on p$K_a$ values. For example, the use of hydrogen-bonding and/or negatively charged substituents appears to be particularly well-suited to shift the $2 \, e^-/2 \, \text{H}^+$ potentials of quinones to more positive values, owing to their increased hydroquinone p$K_a$ values. On the other hand, halogen substituents are especially effective in modulating $1 \, e^-$ reduction potentials, while having minimal impact on the $2 \, e^-/2 \, \text{H}^+$ potentials. The influence of sterically bulky substituents is more complicated, but the results suggest that these groups may be used to shift the $2 \, e^-/2 \, \text{H}^+$ potential to less positive values.

The differential substituent effects on $1 \, e^-$ and $2 \, e^-/2 \, \text{H}^+$ reduction potentials could find utility in quinone-catalyzed aerobic oxidation reactions (Scheme 7.2). For example, chloranil is a
versatile oxidant in organic chemistry (see the right half-reaction in Scheme 7.2, i.e., SubH₂ → Sub^{ox}), and this synthetic utility probably reflects its high 1 e⁻ reduction potential (cf. Figure 7.1). However, chloranil is typically used as a stoichiometric oxidant. Meanwhile, benzoquinone has been implemented successfully as a co-catalyst in a number of aerobic oxidation reactions. The studies described herein suggest that chloranil or other "high potential quinones" could be regenerated by O₂ (cf. Scheme 7.2) at least as readily as benzoquinone if the reaction proceeds via a PCET pathway, on the basis of their similar 2 e⁻/2 H⁺ potentials (cf. Figure 7.1). Recent empirical studies support this prospect, and mechanistic studies of catalytic aerobic oxidation of hydroquinones implicated a PCET pathway.

Scheme 7.2. Simplified Catalytic Cycle for Quinone-Catalyzed Aerobic Oxidation of an Organic Substrate (SubH₂)

7.10. Conclusions

In this work, experimental measurements identified non-linear correlations between the 1 e⁻ and 2 e⁻/2 H⁺ reduction potentials for a set of 15 1,4-benzoquinone derivatives. Subsequent DFT calculations of the 1 e⁻ reduction potentials, pKₐ values, and 2 e⁻/2 H⁺ reduction potentials for a set of 134 substituted quinones revealed approximately linear correlations between these properties and an effective Hammett constant associated with the substituents. More importantly, the calculations identified deviations from these linear scaling relationships for specific classes of substituents, thereby providing an explanation for the experimentally observed non-linear correlations between the 1 e⁻ and 2 e⁻/2 H⁺ reduction potentials. The substituent effects on the pKₐ values tend to be the primary factor governing deviations from the linear Hammett correlations.
Such effects may be used to shift the $2 \, e^{-}/2 \, H^{+}$ reduction potential to more positive values (i.e., using hydrogen-bonding and negatively charged substituents) or to less positive values (i.e., using sterically bulky substituents). Halogenation of quinones may be used to alter the $1 \, e^{-}$ reduction potential with minimal impact on the $2 \, e^{-}/2 \, H^{+}$ reduction potential.

The results of this study, together with recent reports by others,\textsuperscript{11,12} clearly demonstrate the utility of DFT calculations in reproducing and predicting reduction potentials and $pK_a$ values of quinones. One can anticipate that future computational screening studies will play a valuable role in the discovery of new quinones for targeted applications. Meanwhile, the four intuitive parameters identified in this study, namely hydrogen bonding, halogenation, charges, and sterics, may be used to tailor the redox properties of quinones in ways that deviate from standard electronic linear scaling relationships. These parameters especially highlight the importance of considering the effects of quinone substituents on the $pK_a$ values of the reduced quinones. Ongoing studies are now being directed toward implementing these concepts in the development of new catalytic and energy storage and conversion applications.

7.11. References


Chapter 8. Concerted Electron-Proton Transfer Tyr-His Models*†

8.1. Abstract

Nature employs a Tyr$_Z$-His pair as a redox relay that couples proton transfer to the redox process between P680 and the water oxidizing catalyst in photosystem II. Artificial redox relays composed of different benzimidazole-phenol dyads (benzimidazole models His and phenol models Tyr) with substituents designed to simulate the hydrogen bond network surrounding the Tyr$_Z$-His pair have been prepared. When the benzimidazole substituents are strong proton acceptors such as primary or tertiary amines, theory predicts that a concerted two proton transfer process associated with the electrochemical oxidation of the phenol will take place. Also, theory predicts a decrease in the redox potential of the phenol by $\sim$300 mV and a small kinetic isotope effect (KIE). Indeed, electrochemical, spectroelectrochemical, and KIE experimental data are consistent with these predictions. Notably, these results were obtained by using theory to guide the rational design of artificial systems and have implications for managing proton activity to optimize efficiency at energy conversion sites involving water oxidation and reduction.

8.2. Introduction

In photosystem II (PSII) Tyr$_Z$ ($Y_z$) acts as a redox mediator between the oxidized primary electron donor, P680$^+$, and the oxygen evolving complex (OEC), a Mn$_4$CaO$_5$ cluster where water oxidation takes place. By providing an interface between the fast photoinitiated steps of photosynthesis and the slow catalytic process of water oxidation, $Y_z$ is thought to be an

† Author contributions: S. Jimena Mora, Matías Villalba, Marely E. Tejeda-Ferrari, Paul A. Liddell, Brian R. Cherry, Anne-Lucie Teillout, Charles W. Machan, Prof. Clifford P. Kubiak, Prof. Devens Gust, Prof. Thomas A. Moore, and Prof. Ana L. Moore were responsible for all experiments and associated analyses in this chapter.
important factor contributing to the high photochemical quantum efficiency of PSII.\textsuperscript{1,2} It is known that upon oxidation $Y_z$ donates a proton to its hydrogen-bonded partner, His-190, and that the thermodynamic activity of this proton is crucial for preserving the high redox potential necessary for water oxidation. This is one of the most widely cited proton-coupled electron transfer (PCET) processes in natural systems.\textsuperscript{3-9}

This reaction has also been proposed to play an essential thermodynamic role in the generation of proton motive force (PMF). In 2000, Tommos and Babcock wrote, “it is at the level of $Y_z$ that proton currents are switched on” and hypothesized that instead of a single proton transfer upon oxidation of $Y_z$, a series of proton transfers can be associated with the oxidation of the phenol, culminating with proton release into the lumen contributing to the PMF.\textsuperscript{1} In agreement with this hypothesis, a possible proton channel from the OEC through $Y_z$ and extending to the thylakoid lumen was identified in the 1.9 Å-resolution crystal structure of PSII obtained by Umena et al.\textsuperscript{10} However, the issue of proton channels within PSII remains controversial, and the identification of functional proton pathways in the water oxidation cycle is an active area of theoretical and experimental research. Several proton pathways have been proposed, consisting of networks of hydrogen-bonded polar residues and water molecules, where protons can diffuse by a Grotthuss-type mechanism spanning the \~20 Å distance between the OEC and the lumen.\textsuperscript{11-14} It is unambiguous, however, that the protons liberated upon water oxidation find their way to the lumen, where they contribute to the PMF across the thylakoid membrane.

The design, synthesis, and study of simple systems where multiple proton transfers can be associated with the oxidation of a phenol in a concerted mechanism is a challenging but worthwhile endeavor. In addition to providing a deeper understanding of natural systems such as
PSII, these studies help to establish design principles for controlling proton activity at catalytic water oxidation sites in artificial photosynthesis, where managing proton activity is thought to be key to efficient catalysis. Costentin et al. demonstrated two protons transferring concertedly with one electron transfer using an intramolecular contiguous alcohol proton relay induced by oxidation. Herein, we provide theoretical and experimental evidence in biomimetic constructs consisting of benzimidazole phenols (BIPs) substituted with amino groups (see Figure 8.1 for the example of BIP-CH$_2$NEt$_2$) for a concerted one-electron two-proton transfer (E2PT) process that takes place when the phenol is oxidized electrochemically. We call this process E2PT to differentiate it from the EPT process that describes a concerted one-electron one-proton transfer process. The E2PT process shown in Figure 8.1 results in the translocation of protons over a distance of ca. 7 Å and is a starting point for the design of bioinspired proton wires. Although oxidation of the phenol in the amino-BIPs reported here was electrochemically-driven, the E2PT process could be light-driven by substituting the BIP of photochemically activated triad systems previously reported by amino-BIPs.

![Figure 8.1](image)

Figure 8.1. A bioinspired system: substituted benzimidazole phenol (BIP-CH$_2$NEt$_2$), consisting of a phenol internally hydrogen bonded to a benzimidazole proximal nitrogen and a secondary proton acceptor, a tertiary amine, that is hydrogen bonded to the distal NH of the benzimidazole. Upon electrochemical oxidation of the phenol, the system undergoes two concerted proton transfer reactions resulting in a phenoxy radical and an ammonium ion.
8.3. Synthesis and Characterization

Benzimidazole phenol derivatives (BIPs, Figure 8.2) have been chosen to examine the mechanism of proton transfer that accompanies the oxidation of the phenol moiety, and to validate the theoretical prediction of a lowered redox potential for a concerted one-electron two-proton reaction (*vide infra*). Different substitutions on the BIP allow for the interrogation of both one- and two-proton transfer reactions paired with a redox event on a single molecular platform. Substituted BIPs (Figure 8.2) were synthesized by a modification of the standard procedure (Supporting Information, synthesis) with *ortho*-phenylendiamine or its carbomethoxy derivative and the complementary *ortho*-hydroxybenzaldehyde derivative as starting substrates for the Philipps-Ladenburg reaction.\textsuperscript{22,25,26} Amides 4 was prepared by aminolysis of 3, and amide 5 by a coupling reaction starting with 2 and diethylamine with 1-ethyl-3-(3-dimethylaminopropyl)carbodiimide (EDCI) as a coupling reagent. The reduction of 4 and 5 with LiAlH\textsubscript{4} generated the corresponding amines, 6 and 7.

![Chemical structures of BIPs and their derivatives](image)

**Figure 8.2.** Benzimidazole phenol derivatives (BIPs) of the present study.
The presence of a strong intramolecular hydrogen bond involving the phenol and the nitrogen lone pair of benzimidazole (~2.5 Å O–N distance) is well established,2,27 and NMR data for 1–7 support such a bond in these cases. Evidence for the formation of a hydrogen bond between the distal NH of benzimidazole and the benzyamine nitrogen (Figure 8.1) comes from the NMR chemical shift of the distal NH group of benzimidazole. The chemical shift of benzimidazole NH varies considerably with solvent, shifting downfield by about 2.5 ppm from chloroform to acetone. This change is attributed to the formation of hydrogen bonds with the solvent in the case of acetone.2 In 7 the 1H resonance of the benzimidazole NH is found at 11.17 ppm in chloroform, considerably downfield compared to the equivalent NH in 1 (9.34 ppm),22 and does not shift over ten-fold changes in concentration. Additional support for the formation of the hydrogen bond between the distal NH of benzimidazole and the exocyclic amine nitrogen comes from IR data, where a shift to lower frequencies is observed for the stretching and bending vibrations involving the NH compared with 1 (Figure 8.7B and Figures S28 and S31).

Because of the unsymmetrical benzimidazole group, any BIP substituted at the 4 position (i.e., compounds 2–7 in Figure 8.2) consists of a mixture of two isomers due to tautomerization of the imidazole and rotation around the bond between the benzimidazole and the phenol moieties. Because their interconversion is slow on the NMR timescale (lifetime of the isomers is >200 ms as measured for BIP-5-CH3), their presence and ratio can be clearly detected by 1H NMR. Only one of the isomers has the two internal hydrogen bonds necessary for E2PT; it is shown above for 7 (Figure 8.1) where the OH is hydrogen bonded to the lone pair of the benzimidazole nitrogen and a second internal hydrogen bond exists between the lone pair of the exocyclic NEt2 group and the distal NH of the benzimidazole. The other isomer (not shown) lacks the hydrogen bond between the distal NH of the benzimidazole and the NEt2. The ratio of
isomers is approximately 1:1.5 in the case of 7 in acetonitrile as determined by $^1$H NMR (Figure S22).

8.4. Thermodynamics of PCET

The diabatic states and corresponding electron transfer (ET) and proton transfer (PT) reactions are depicted in Figure S37 for BIP (1) and for BIP-CH$_2$NEt$_2$ (7). The analogous schemes for the other BIP system are similar, and the results are summarized in Figure 8.4. The ET steps correspond to an ET between the phenol and the electrode (e.g., $1a \rightarrow 2a$, $1aa \rightarrow 2aa$), while the PT steps correspond to an intramolecular PT between the phenol and the proximal N of the imidazole (e.g., $1a \rightarrow 1b$, $1aa \rightarrow 1ba$) or between the distal N of the imidazole and the exocylic N of an amine group (e.g., $1ab \rightarrow 1bb$). For the concerted electron-proton transfer (EPT), the ET and PT occur concurrently without a stable intermediate. For BIP (1), the concerted EPT process corresponds to the process $1a \rightarrow 2b$. A concerted one-electron, two-proton transfer (E2PT) may also occur for certain systems and would correspond to the process $1aa \rightarrow 2bb$, as for BIP-CH$_2$NEt$_2$ (7).
Figure 8.3. Schematic depiction of the ET, PT, and EPT (or E2PT) reactions involved in the PCET processes for BIP (1) (A) and BIP-CH$_2$NEt$_2$ (7) (B). Similar schemes can be drawn for the other BIP systems.
Figure 8.4. The thermodynamic square/cube schemes with the relative free energies of the different oxidized states drawn for BIP compounds 1–11 (A–K). The free energies ($\Delta G^\circ$) are reported in units of kcal/mol and are calculated relative to either the 2$a$ or 2$aa$ states for the oxidized states. The states are labelled according to the representative structures shown in Figure 8.3.
In all systems 1–11, no minima were located with the proton(s) localized on the acceptor(s) in the reduced state, indicating that PT is not predicted to occur prior to ET (oxidation). For BIP (1) and BIP-PF$_{10}$ (8), the 2$b$ state (the oxidized state with a PT to the proximal imidazole) is more thermodynamically stable than the 2$a$ state, in which no PT occurs (Figure 8.3A and 8.4A for 1). For 2–5, the 2$ba$ state (the oxidized state with only PT to the proximal imidazole) is more thermodynamically stable than the 2$bb$ state (the oxidized state with PT to the proximal imidazole and PT to the exocyclic amine) and the 2$aa$ state, in which neither PT step occurs. This 2$ba$ state is favored over the 2$bb$ state by ~14 kcal/mol in the case of BIP-COOMe (3) (Figure 8.4C). These results predict that only the first PT step will occur for 2–5. In contrast, for BIP-CH$_2$NH$_2$ (6), BIP-CH$_2$NEt$_2$ (7), BIP-imine$^\text{H}$ (9), BIP-imine$^\text{Me}$ (10), and BIP-imine$^\text{Cl}$ (11), the 2$bb$ state (the oxidized state with both PT reactions) is more thermodynamically stable than the 2$ba$ state (the oxidized state with only PT to the proximal imidazole), predicting that both PT steps will occur for these systems. In the case of BIP-CH$_2$NEt$_2$ (7), this 2$bb$ state is favored over the 2$ba$ state by ~6 kcal/mol (Figure 8.4G).

8.5. Electrochemical Evidence for a Concerted E2PT Process in 6 and 7

The redox potentials for compounds 1–8 were calculated with DFT using previously benchmarked methods described elsewhere, and the predicted values were subsequently found to be consistent with the experimental measurements (Table 8.1). For most of the BIP compounds (1–5 and 8), the predicted and experimental redox potentials are ~0.9–1.0 V vs. SCE in acetonitrile and correspond to the concerted oxidation of the phenol and transfer of the phenolic proton to the benzimidazole (an EPT process). For 6 and 7, the oxidation of the phenol was predicted by DFT to occur concertedly with the transfer of both the phenolic and distal benzimidazole protons (an E2PT process) on the basis of a thermodynamic analysis, as discussed...
Moreover, prior to experimental measurements, the DFT calculations predicted that this concerted two-proton transfer would shift the potentials by approximately 300 mV to less positive values (~0.6 V vs. SCE) in 6 and 7.

**Table 8.1.** Calculated and Experimental First Redox Potentials for the Phenol of BIPs

<table>
<thead>
<tr>
<th>Compound</th>
<th>Calculated</th>
<th>Experimental</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 – BIP</td>
<td>1.04(^a)</td>
<td>1.04(^b)</td>
</tr>
<tr>
<td>2 – BIP-COOH</td>
<td>0.98</td>
<td>0.90</td>
</tr>
<tr>
<td>3 – BIP-COOMe</td>
<td>0.96</td>
<td>0.93</td>
</tr>
<tr>
<td>4 – BIP-CONH(_2)</td>
<td>0.94</td>
<td>0.92</td>
</tr>
<tr>
<td>5 – BIP-CONEt(_2)</td>
<td>0.95</td>
<td>0.92</td>
</tr>
<tr>
<td>6 – BIP-CH(_2)NH(_2)</td>
<td>0.66</td>
<td>0.56</td>
</tr>
<tr>
<td>7 – BIP-CH(_2)NEt(_2)</td>
<td>0.57</td>
<td>0.54</td>
</tr>
<tr>
<td>8 – BIP-PF(_10)</td>
<td>0.94</td>
<td>1.00(^b)</td>
</tr>
<tr>
<td>9 – BIP-Ph(_2)imine(^c)</td>
<td>0.84</td>
<td>n/a</td>
</tr>
<tr>
<td>10 – BIP-Ph(_2)Meimine(^c)</td>
<td>0.82</td>
<td>n/a</td>
</tr>
<tr>
<td>11 – BIP-Ph(_2)Climine(^c)</td>
<td>0.88</td>
<td>n/a</td>
</tr>
</tbody>
</table>

\(^a\)This couple was used as the reference for all other calculated potentials, so it agrees with the experimental value by construction. In this procedure, the redox potentials for 2–11 were calculated relative to the redox potential of 1, which was shifted to agree with the experimental value (i.e., the difference between the experimental and calculated values for 1 was added to all calculated potentials). \(^b\)Experimental data from Ref. 2. \(^c\)For the structures of 9–11 see Supporting Information, Figure S1.

This theoretical prediction is corroborated experimentally (see Table 8.1 and Figure 8.5A). The cyclic voltammograms (CVs) of the amino substituted BIPs exhibit a midpoint potential approximately 300 mV less positive than that of the other BIPs. The waves are quasi-reversible in acetonitrile solution with a peak-to-peak separation of ~100 mV. Good evidence to the oxidation of the phenol comes from the comparison with the redox characteristics of the reference compounds lacking the phenol group, benzylamine and diethylbenzylamine, which exhibit irreversible oxidation waves at 1.45 and 0.85 V vs. SCE, respectively, and thus are substantially different from 6 and 7 (Supporting Information, Figures S23A and B). Furthermore, titration with trifluoracetic acid (TFA) of a solution of 7 in dichloromethane (Figure 8.5B and
Supporting Information, Figures S25A and B) clearly shows the disappearance of the wave at ~0.6 V vs. SCE associated with the E2PT process, and at 2 mM TFA the appearance of the wave typical of the EPT process observed for BIPs 1–5 listed in Table 8.1. These results are explained by protonation of the exocyclic amino group by TFA (pK_\text{a} ~13 in acetonitrile) in either 6 (primary amino group pK_\text{a} ~17 in acetonitrile) or 7 (tertiary amino group pK_\text{a} ~19 in acetonitrile), which eliminates the exocyclic amino group as a thermodynamically competent secondary proton acceptor.\textsuperscript{29-31} At even higher concentrations of TFA, the electrochemical oxidation of the protonated benzimidazole group can be detected at ~1.4 and 1.5 V vs. SCE for 6 and 7, respectively.\textsuperscript{32}

The absence of two proton transfers associated with the single electron oxidation of the phenol for 2–5 can be explained on thermodynamic grounds. The pK_\text{a}'s in acetonitrile of the protonated carbonyl groups of 2 and 3 are \leq 0, and those of 4 and 5 are \sim 4,\textsuperscript{33} which is much lower than the pK_\text{a} of the benzimidazolium ion in acetonitrile (~14).\textsuperscript{30} Therefore a second proton transfer is thermodynamically unfavorable. The calculated free energy differences\textsuperscript{28} between the

---

**Figure 8.5.** (A) CVs of BIPs. Concentration: 1 mM of the indicated BIPs, 0.5 M TBAPF_6 in dry acetonitrile. WE: glassy carbon, RE: Ag/AgCl (Fc as internal reference), CE: Pt. Note that 3, 4 and 5 do not oxidize at 0.7 V vs. SCE. (B) CVs of 7 before and after the addition of two equivalents of TFA in dry dichloromethane. Concentration of 7, 1 mM, 0.5 M TBAPF_6 supporting electrolyte, WE: glassy carbon, RE: Ag/AgCl (Fc as internal reference), CE: Pt. Sweep rate, 100 mV s\textsuperscript{−1}.
oxidized product after double proton transfer and the oxidized product after single proton transfer are given in Table 8.2. These values indicate that when the BIPs are substituted with better proton acceptors such as primary or tertiary amines, the second proton transfer becomes thermodynamically favorable, and an E2PT process could occur.

Table 8.2. Calculated $\Delta G^\circ$ Between Double and Single PT Oxidize States$^a$

<table>
<thead>
<tr>
<th>Oxidized species</th>
<th>$\Delta G^\circ$ (kcal/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 – BIP-COOH$^{**}$</td>
<td>15.2</td>
</tr>
<tr>
<td>3 – BIP-COOMe$^{**}$</td>
<td>14.9</td>
</tr>
<tr>
<td>4 – BIP-CONH$_2$$^{**}$</td>
<td>7.1</td>
</tr>
<tr>
<td>5 – BIP-CONEt$_2$$^{**}$</td>
<td>5.6</td>
</tr>
<tr>
<td>6 – BIP-CH$_2$NH$_2$$^{**}$</td>
<td>–4.2</td>
</tr>
<tr>
<td>7 – BIP-CH$_2$NEt$_2$$^{**}$</td>
<td>–6.5</td>
</tr>
<tr>
<td>9 – BIP-PhH$^{imine}$$^{**}$</td>
<td>–2.9</td>
</tr>
<tr>
<td>10 – BIP-PhMe$^{imine}$$^{**}$</td>
<td>–3.1</td>
</tr>
<tr>
<td>11 – BIP-PhCl$^{imine}$$^{**}$</td>
<td>–2.1</td>
</tr>
</tbody>
</table>

$^a$The reported values are the free energy differences between the oxidized state with and without the second proton transfer. The double proton transfer corresponds to proton transfer from the phenol to the proximal imidazole N and proton transfer from the distal imidazole NH to the R group, while the single proton transfer corresponds to only the first proton transfer. Negative values for $\Delta G^\circ$ indicate that the double proton transfer is more thermodynamically favorable than the single proton transfer. No values are reported for 1 and 8 because only a single proton transfer is possible. Using the experimental p$K_a$s (vide supra)$^{29-31}$ for the benzimidazole and exocyclic amines in acetonitrile we estimated $\Delta G^\circ = –3.9$ kcal/mole for 6 and –6.5 kcal/mole for 7, in good agreement with the calculated values.

The thermodynamic price paid for the E2PT process in 6 and 7 is the loss of ~300 mV in the redox potential of the phenol, indicating that neither could be used to oxidize a catalyst for water oxidation at near-neutral to acidic pH. This price must be a function of the $\Delta pK_a$ between the proton acceptor and donor for the second proton transfer (i.e., the p$K_a$ difference between the protonated exocyclic amine and the distal NH of the benzimidazolium ion). Encouragingly, however, theoretical calculations predict that substituents with reduced p$K_a$’s, such as substituted...
imines attached to BIP (see 9–11 in Tables 8.1 and 8.2, and Supporting Information Figure S1), would still undergo concerted E2PT but maintain a considerably higher potential for the phenol oxidation.

It is interesting to consider the implications of these findings for the phenolic proton of the Tyrz-His pair in photosynthesis. If the phenolic proton is pumped into the lumen and the phenol is “reloaded” with a proton from ongoing water oxidation, the ΔpKₐ between the His and lumen must be less than that of 6 and 7 in order for the redox relay to be able to oxidize the OEC. This ΔpKₐ is a function of the proton electrochemical potential (in the Michelian sense) inside the lumen, and could maintain the redox poise sufficient to oxidize the OEC.³⁴ Alternatively, if the phenolic proton is not pumped into the lumen, but remains shared between the phenoxyl radical and its His hydrogen-bonded partner, the surroundings could modulate the ΔpKₐ to maintain the necessary redox potential. This mechanism has been proposed and is known as proton rocking.³⁵,³⁶ As alluded to in the introduction, regardless of mechanistic details, this proton is a key player in water oxidation and the concomitant generation of the PMF in photosynthetic membranes.

8.6. Infrared Spectroelectrochemistry

Infrared spectroelectrochemistry (IRSEC) allows for the identification of bonding changes due to redox-linked protonation/deprotonation of specific sites in the molecules during the PCET process. For example, in the IRSEC spectra obtained for 1 at a potential of 1.0 V vs. SCE, protonation of the benzimidazole moiety upon oxidation of the phenol can be easily detected by the decrease in the intensity of the bands at 3453 and 3415 cm⁻¹, which correspond to the NH stretching of the benzimidazole moiety, and the concomitant appearance of the band at 3323 cm⁻¹, which corresponds to the NH stretching of the benzimidazolium ion (Supporting
Information, Figure S26A).\textsuperscript{27,37} In agreement with this observation, similar changes were detected in this region of the IR spectrum when a solution of 1 was treated with gaseous HCl (Figure S26B).

Figure 8.6. (A) Time course of the IRSEC spectra of 1 (19 mM) obtained at a potential of 1.0 V vs. SCE. The spectrum of 1 obtained at resting potential is shown in black. Solvent dichloromethane, 0.1 M TBAPF\textsubscript{6}. (B) IR spectra of 1, neutral (black) and protonated with gaseous HCl (blue). Solvent dichloromethane.

Figure 8.6A shows the IRSEC spectra of 1 at lower frequencies (1700–1300 cm\textsuperscript{-1}), where large changes in absorbance are indicated with upward (increasing intensity) and downward (decreasing intensity) arrows when a potential of 1.0 V vs. SCE was applied. Many of the changes observed are similar to those observed in the IR spectrum of 1 when it is treated with gaseous HCl (see the bands at 1626, 1556, and 1526 cm\textsuperscript{-1}, red arrows in Figure 8.6B). For example, the band at 1626 cm\textsuperscript{-1} increases considerably under both the applied potential and following protonation by treatment of 1 with HCl. This band has a large component of the C=N stretching vibration of the benzimidazole, for which the dipole moment increases upon protonation.\textsuperscript{37} However, some mismatches are expected between the IR and the IRSEC spectra because the IRSEC detects the bonding changes associated with the formation of the phenoxyl radical and the associated proton transfer (the EPT process), whereas the treatment with HCl only affords a protonation state mimicking the product of the EPT process without the ET. For
example, a new band at 1556 cm\(^{-1}\) grows in both spectra when a potential of 1.0 V \textit{vs.} SCE is applied (Figure 8.6A) and when 1 is protonated by treatment with gaseous HCl (Figure 8.6B); however, the band at \(~1526\text{ cm}\(^{-1}\)) decreases in intensity under the applied potential (Figure 8.6A) but disappears upon protonation (Figure 8.6B). The band at 1556 cm\(^{-1}\) includes an NH in-plane bending vibration in the benzimidazolium ion, while the band at 1526 cm\(^{-1}\) includes an NH in-plane bending mode of the benzimidazole that includes a considerable contribution from the phenoxy radical, as supported by DFT normal mode analysis (Supporting Information Figures S35 and S36).\(^{37-39}\) Both of these bands exhibit clear shifts to lower wavenumbers upon deuteration (Supporting Information, Figures S32A and B). In summary, considering the close correlation between the IRSEC and the IR data, it can be concluded that protonation of BIP by HCl acid and by the EPT process produces similar changes in the bonding around the benzimidazole moiety, namely the formation of the benzimidazolium ion.

Because the redox potentials of 6 and 7 are \(~300\text{ mV}\) lower than those of BIPs 1–5, the changes in the IRSEC spectra of 6 and 7 are expected to occur at significantly lower potentials than those of the other BIPs. As predicted, the IRSEC spectrum of 1 does not show any changes when a potential of 0.7 V \textit{vs.} SCE was applied (Supporting Information, Figure S29A and B), while the IRSEC spectrum of 7 shows significant changes over several minutes at this potential, the results of which are described below.

In the IRSEC spectrum of 7 obtained at zero applied potential, a band is observed at 3375 cm\(^{-1}\), which corresponds to the stretching mode of the benzimidazole NH (Supporting Information, Figure S28A). The shift of the NH stretching mode from \(~3430\text{ cm}\(^{-1}\)) in 1 to 3375 cm\(^{-1}\) in 7 can be interpreted as a weakening of the benzimidazole NH bond due to the formation of the internal hydrogen bond with the exocylic tertiary amine. When a potential of 0.7 V \textit{vs.} SCE was applied (Supporting Information, Figure S29A and B), the IRSEC spectrum of 7 shows significant changes over several minutes at this potential, the results of which are described below.
SCE is applied in the case of 7, two new bands grow in at ~3460 and 3190 cm\(^{-1}\), which are assigned to the stretching modes of the proximal benzimidazole NH and the NH of the exocyclic ammonium ion, respectively. These assignments are supported by the calculated normal modes for 7 (Supporting Information, Figures S35 and S36). It is important to note that no indication of a benzimidazolium NH stretching band, observed at 3323 cm\(^{-1}\) for 1 after EPT, can be detected in the case of 7.

**Figure 8.7.** (A) Time course of the IRSEC spectra of 7 (17 mM) obtained at a potential of 0.7 V vs. SCE. The spectrum of 7 obtained at resting potential is shown in black. Solvent dichloromethane, 0.1 M TBAPF6. (B) Spectra of 7 neutral (black), tertiary exocyclic amine protonated with 2 equivalents of TFA (pink), and protonated with excess gaseous HCl (blue). Solvent dichloromethane.

In the region between 1700 and 1300 cm\(^{-1}\) no difference can be detected in the spectra of 1 taken at 0 and 0.7 V vs. SCE (Supporting Information, Figure S29B), as the EPT process is expected to occur at ~1.0 vs. SCE. Under comparable conditions for 7, four distinct bands at 1622, 1604, 1529, and 1340 cm\(^{-1}\) increase in intensity, while the band at 1515 cm\(^{-1}\) decreases in intensity (Figure 8.7A). The band at 1515 cm\(^{-1}\) has a large component of the NH in-plane bending, comparable to the band at 1526 cm\(^{-1}\) of 1 but at lower frequency, in agreement with the presence of an internal hydrogen bond with the exocyclic amine. Importantly, there is no band growing in at 1556 cm\(^{-1}\) for 7, which was clearly observed for 1 after EPT. This band
corresponds to the formation of the benzimidazolium ion and can be observed in the IR spectrum of 7 only after treatment with excess gaseous HCl, where both the exocyclic amino group and the benzimidazole moiety are expected to be protonated (Figure 8.7B and Supporting Information, Figure S31C).

The data in Figure 8.5B indicate that adding two equivalents of TFA to a solution of 7 protonates the exocyclic amine, which disrupts the hydrogen bond with the distal NH of the benzimidazole and prevents this second proton transfer (Supporting Information, Figure S31B). This process was monitored by the disappearance of the band at 1515 cm\(^{-1}\) and the concomitant appearance of the band at 1529 cm\(^{-1}\) (Figure 8.7B and Supporting Information, Figure S33A). The band at 1529 cm\(^{-1}\) is characteristic of the benzimidazole NH bending vibration and upon deuteration shifts to 1506 cm\(^{-1}\) (Supporting Information, Figure S33B). The IRSEC experiments (Figure 8.7A) show that the bands at 1515 cm\(^{-1}\) and 1529 cm\(^{-1}\) change in the same manner upon oxidation of the phenol as they did in the TFA titration above. This behavior is consistent with protonation of both the exocyclic amine and the proximal N of the benzimidazole (Figure S31D). These observations provide further evidence for the proposed E2PT process (Figure 8.1) in which the exocyclic amine and the proximal N of the benzimidazole are protonated but the distal N of the benzimidazole is not protonated after oxidation of the phenol in 7.

8.7. Analysis of Calculated and Experimental KIE

A kinetic isotope effect (KIE) is often regarded as the hallmark of reactions involving concerted PCET processes, although exceptions are possible.\(^{40,41}\) Experimental KIE values were determined electrochemically and are reported in Table 8.3 along with theoretical values (\textit{vide infra}). For the BIP derivatives 1–5 reported in this work, a KIE of \(~2\) is expected if the PCET takes place by a concerted mechanism, according to studies of related systems.\(^{27,42,43}\) In Table
the theoretical KIE values for compounds 1–5 correspond to a single proton transferred concertedly upon oxidation of the phenol (EPT). In the case of 6 and 7, the calculated KIE values correspond to two protons transferred concertedly upon oxidation of the phenol (E2PT).

We calculated the electrochemical rate constant using a general PCET theory.\textsuperscript{44} In this treatment, the anodic nonadiabatic EPT rate constant for fixed proton donor-acceptor distance ($R$) is given by:\textsuperscript{45–47}

$$k^{\text{EPT}}(\eta; R) = \sum_{\mu, \nu} P_\mu \left( \frac{V^{\text{el}} S_{\mu \nu}}{\beta' \hbar} \right)^2 \sqrt{\frac{\pi}{\lambda k_B T}} \rho_M \int d\varepsilon [1 - f(\varepsilon)] \exp \left[ \frac{-\Delta G^{\dagger}_{\mu \nu}}{k_B T} \right]$$

(8.1)

where the double summation is over all pairs of reactant/product electron-proton vibronic states, $P_\mu$ is the Boltzmann probability for vibronic state $\mu$, $f(\varepsilon)$ is the Fermi distribution function for the electronic states with energy $\varepsilon$ in the electrode, $\rho_M$ is the density of states at the Fermi level, $V^{\text{el}}$ is the electronic coupling, $\beta'$ is a parameter of magnitude $\sim 1$ Å$^{-1}$ representing the exponential decay of $V^{\text{el}}$ with distance, $\lambda$ is the reorganization energy, $S_{\mu \nu}$ is the overlap integral between the proton vibrational wavefunctions $\mu$ and $\nu$, and $\Delta G^{\dagger}$ is the free energy barrier associated with states $\mu$ and $\nu$, which depends on the overpotential $\eta$. In Eq. 8.1, $P_\mu$, $S_{\mu \nu}$, and $\Delta G^{\dagger}$ depend on the proton donor-acceptor distance $R$. A detailed explanation of how the parameters in Eq. 8.1 are defined and calculated is given in Sections 8.12.

To account for the proton donor-acceptor motion, the rate constant is thermally averaged over the proton donor-acceptor distance $R$:\textsuperscript{40,48–50}

$$k^{\text{EPT}}(\eta) = \int P(R) k^{\text{EPT}}(\eta; R) dR$$

(8.2)

Thus, the overall EPT rate constant is determined by calculating the rate constant at a series of $R$ values and numerically integrating over $R$, weighting the rate constant by the probability $P(R)$ of
sampling each $R$ value. For a given pair of reactant/product ($\mu/\nu$) vibronic states, the rate constant is proportional to the square of the proton vibrational wavefunction overlap ($S_{\mu\nu}$), and the KIE is proportional to the square of the ratio of the overlaps for hydrogen and deuterium.\textsuperscript{40,51} Typically, the KIE increases as the overlap decreases. However, the combined summation over all pairs of reactant/product vibronic states and the integration over $R$ leads to more complex behavior that requires the complete calculation of the rate constant for each molecular system. This theoretical treatment has been extended to the case of concerted electron transfer with two proton transfers (E2PT), resulting in similar physical principles but additional complexity.\textsuperscript{49}

Table 8.3. Calculated and Experimental Kinetic Isotope Effects

<table>
<thead>
<tr>
<th>Compound</th>
<th>Calculated$^a$</th>
<th>Experimental</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 – BIP</td>
<td>1.9</td>
<td>n/a (1.4$^b$)</td>
</tr>
<tr>
<td>2 – BIP-COOH</td>
<td>2.0</td>
<td>1.5 ± 0.3</td>
</tr>
<tr>
<td>3 – BIP-COOMe</td>
<td>2.1</td>
<td>1.8 ± 0.3</td>
</tr>
<tr>
<td>4 – BIP-CO\text{H}_2</td>
<td>1.8</td>
<td>1.7 ± 0.5</td>
</tr>
<tr>
<td>5 – BIP-CO\text{N}e\text{t}_2</td>
<td>2.0</td>
<td>n/a</td>
</tr>
<tr>
<td>6 – BIP-\text{CH}_2\text{NH}_2</td>
<td>1.6</td>
<td>n/a</td>
</tr>
<tr>
<td>7 – BIP-\text{CH}_2\text{NE}t\text{t}_2</td>
<td>1.3</td>
<td>0.9 ± 0.5</td>
</tr>
<tr>
<td>8 – BIP-\text{PF}_{10}$^c$</td>
<td>1.0</td>
<td>1.0 ± 0.3</td>
</tr>
</tbody>
</table>

$^a$The uncertainties of the calculated KIEs are approximately ± 0.5 within this theoretical framework (Supporting Information, KIE).

$^b$Data for related compound from reference 43. $^c$Supporting Information, BIP-\text{PF}_{10}.

KIEs were experimentally measured in the electrochemical oxidation of several of the phenols shown in Figure 8.2 (see Table 8.3 and Supporting Information, Figure S34B). Overall, the calculated and experimentally measured KIE values are in good agreement, with both sets of KIE values showing similar trends across the BIP compounds studied. The KIE values obtained for BIPs 1–4 are relatively small but are in line with previously reported values for related systems.\textsuperscript{27,43,52,53} A more detailed analysis of the KIEs is provided in the Supporting Information (pages S54–S61). The theoretical calculations lead to a KIE of unity for 8 because of dominant contributions from excited vibronic states with vibrational wavefunction overlaps that are nearly
unity, mainly arising from the delocalization of the unpaired electron onto the porphyrin ring in the oxidized state (discussed in next section). In contrast, the calculated KIEs of ~2 for BIPs 1–5 arise because of dominant contributions from excited vibronic states with smaller vibrational wavefunction overlaps (~0.5). In the case of 6 and 7, where two proton transfers are associated with the oxidation of the phenol, the calculated KIE values, as well as the experimentally measured KIE for 7, are slightly smaller than those of BIPs 1–5. An in-depth analysis of the calculated rate constants and KIEs is given in the next section.

In principle, the shift in the redox potentials for compounds 6 and 7 could arise from a stepwise mechanism, in which an EPT step is followed by very fast, thermodynamically favorable transfer of the second proton, leading to the E2PT product. Because the CVs cannot clearly distinguish between this stepwise mechanism and the concerted E2PT mechanism, we are unable to rule out this possibility definitively. However, the DFT calculations indicate that the concerted mechanism is thermodynamically favored over the stepwise mechanism by 4–7 kcal/mol for compounds 6 and 7 (Table 8.2). Moreover, for compound 7, the calculated KIE assuming the concerted E2PT mechanism is in agreement with the experimental KIE, which is slightly lower than the KIEs for the BIP compounds undergoing an EPT process (1–5, Table 8.3), also supporting an E2PT mechanism for 7. However, due to the error bars in both the experimental and calculated KIEs, this agreement does not provide conclusive evidence of a specific mechanism. All of these data point to either a concerted E2PT mechanism or a stepwise mechanism with an extremely short-lived intermediate, which could be viewed as virtually concerted on relevant timescales. Regardless of the detailed mechanism, the net process is the transfer of two protons upon oxidation of the BIP, producing the E2PT product, for compounds 6 and 7.
8.8. Analysis of Calculated Rate Constants and KIE

The proton potentials are asymmetric double well potentials, where the donor (left) well is lower in energy for the reduced state, and the acceptor (right) well is lower in energy for the oxidized state. The presence of higher energy minima on the opposite side allows for excited proton vibrational wavefunctions to become localized on the opposite side. Moreover, the separation between the two minima and the potential energy barrier decrease as the donor-acceptor distance decreases. Thus, the overlap integrals between the vibrational wave functions and the relative energies among the vibronic states will also depend on the donor-acceptor distance(s). It is also apparent that the double well proton potentials for dyad BIP-PF$_{10}$ (8) are more symmetric than for the other systems, which will be discussed below.

![Figure 8.8. Harmonic probability distribution function $P(R)$, electrochemical rate constant $k_a^{EPT}(\eta = 0; R)$, and the product of the two quantities for all systems. The quantities are depicted as functions of $\delta R$ or $2\delta R$, the deviations of the proton donor-acceptor distances from their equilibrium values. The vertical line depicts the dominant proton donor-acceptor distances: for 1–5 and 8 the values correspond to the dominant $R_{ON}$ distance, and for BIP-CH$_2$NH$_2$ (6) and BIP-CH$_2$NET$_2$ (7) the first and second values correspond to the dominant $R_{ON}$ and $R_{NN}$ distances, respectively.](image)
The dominant proton donor-acceptor distance in the EPT process is determined by a balance between the probability distribution function $P(R)$ and the electrochemical rate constant $k_{a}^{EPT}(\eta;R)$. $P(R)$ has a maximum at the equilibrium donor-acceptor distance $\bar{R}$, and the rate constant typically increases as the donor-acceptor distance decreases relative to its equilibrium value. Figure 8.8 depicts the probability distribution function $P(R)$, the rate constant $k_{a}^{EPT}(\eta;R)$ calculated from Eq. 8.3, and the product of these two quantities (i.e., the integrand of Eq. 8.2). The dominant donor-acceptor distance, which has the greatest contribution to the total anodic rate constant, is smaller than the equilibrium donor-acceptor distance due to the significant increase in the rate constant as $R$ decreases.

Tables 8.4 and 8.5 provide the main contributions to the rate constants at the dominant proton donor-acceptor distance(s). The contribution of each pair of reduced/oxidized ($\mu/\nu$) vibronic states is predominantly determined by a balance among the Boltzmann probability of the reactant state $P_{\mu}$, the free energy barrier $\Delta G^{4}$, and the overlap integral $S_{\mu\nu}$ between the reduced and oxidized proton vibrational wavefunctions.
### Table 8.4. Main Contributions to the Rate Constant with Hydrogen

<table>
<thead>
<tr>
<th>Compound</th>
<th>$R_{ON}^{a}$</th>
<th>$R_{NN}^{b}$</th>
<th>$\mu/\nu$</th>
<th>$P_{\mu}$</th>
<th>$\Delta \tilde{U}_{\mu\nu}$</th>
<th>$\Delta \tilde{G}^{\ddagger}_{\mu\nu}$</th>
<th>$S_{\mu\nu}^{2}$</th>
<th>% contrib$^{d}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>(1) BIP</td>
<td>2.47</td>
<td>—</td>
<td>1/0</td>
<td>1.07 x 10^{-3}</td>
<td>-4.05</td>
<td>3.52</td>
<td>0.537</td>
<td>63</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>0/0</td>
<td>1.00</td>
<td>0.00</td>
<td>5.35</td>
<td>0.003</td>
<td>17</td>
</tr>
<tr>
<td>(2) BIP-COOH</td>
<td>2.47</td>
<td>—</td>
<td>1/0</td>
<td>4.67 x 10^{-4}</td>
<td>-4.54</td>
<td>3.47</td>
<td>0.447</td>
<td>51</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>0/0</td>
<td>1.00</td>
<td>0.00</td>
<td>5.51</td>
<td>0.002</td>
<td>17</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>0/1</td>
<td>1.00</td>
<td>4.55</td>
<td>8.02</td>
<td>0.095</td>
<td>11</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>2/0</td>
<td>3.32 x 10^{-5}</td>
<td>-6.11</td>
<td>2.88</td>
<td>0.439</td>
<td>10</td>
</tr>
<tr>
<td>(3) BIP-COOMe</td>
<td>2.48</td>
<td>—</td>
<td>1/0</td>
<td>5.25 x 10^{-4}</td>
<td>-4.47</td>
<td>3.47</td>
<td>0.459</td>
<td>53</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>0/0</td>
<td>1.00</td>
<td>0.00</td>
<td>5.46</td>
<td>0.002</td>
<td>18</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>0/1</td>
<td>1.00</td>
<td>4.55</td>
<td>7.99</td>
<td>0.094</td>
<td>10</td>
</tr>
<tr>
<td>(4) BIP-CNH$_2$</td>
<td>2.52</td>
<td>—</td>
<td>1/0</td>
<td>1.87 x 10^{-4}</td>
<td>-5.09</td>
<td>2.63</td>
<td>0.524</td>
<td>70</td>
</tr>
<tr>
<td>(5) BIP-CNEt$_2$</td>
<td>2.48</td>
<td>—</td>
<td>1/0</td>
<td>7.12 x 10^{-4}</td>
<td>-4.29</td>
<td>2.79</td>
<td>0.496</td>
<td>61</td>
</tr>
<tr>
<td>(6) BIP-CH$_2$NH$_2$</td>
<td>2.42</td>
<td>2.71</td>
<td>1/0</td>
<td>7.71 x 10^{-2}</td>
<td>-1.47</td>
<td>4.80</td>
<td>0.511</td>
<td>96</td>
</tr>
<tr>
<td>(7) BIP-CH$_2$NEt$_2$</td>
<td>2.47</td>
<td>2.84</td>
<td>1/0</td>
<td>1.61 x 10^{-1}</td>
<td>-0.98</td>
<td>5.24</td>
<td>0.410</td>
<td>92</td>
</tr>
<tr>
<td>(8) BIP-PF$_{10}$</td>
<td>2.52</td>
<td>—</td>
<td>1/0</td>
<td>1.00</td>
<td>1.51</td>
<td>5.04</td>
<td>0.967</td>
<td>75</td>
</tr>
</tbody>
</table>

$^{a}$Distances given in Å and energies in kcal/mol.
$^{b}$These are the dominant donor-acceptor distances, which have the greatest contribution to the overall standard rate constant.
$^{c}$ $\Delta \tilde{G}^{\ddagger}_{\mu\nu} = (\Delta \tilde{U}_{\mu\nu} + \lambda)^2 / 4\lambda$ and is the effective free energy barrier at $\eta = \varepsilon = 0$.
$^{d}$This is the percentage contribution to the overall standard rate constant at the dominant donor-acceptor distance, $k_{k^{\text{PT}}}^{\Lambda}(\eta = 0; R)$, excluding contributions less than 10%.

### Table 8.5. Main Contributions to the Rate Constant with Deuterium

<table>
<thead>
<tr>
<th>Compound</th>
<th>$R_{ON}^{b}$</th>
<th>$R_{NN}^{b}$</th>
<th>$\mu/\nu$</th>
<th>$P_{\mu}$</th>
<th>$\Delta \tilde{U}_{\mu\nu}$</th>
<th>$\Delta \tilde{G}^{\ddagger}_{\mu\nu}$</th>
<th>$S_{\mu\nu}^{2}$</th>
<th>% contrib$^{d}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>(1) BIP</td>
<td>2.52</td>
<td>—</td>
<td>2/0</td>
<td>1.02 x 10^{-4}</td>
<td>-5.44</td>
<td>3.09</td>
<td>0.682</td>
<td>63</td>
</tr>
<tr>
<td>(2) BIP-COOH</td>
<td>2.52</td>
<td>—</td>
<td>2/0</td>
<td>3.17 x 10^{-5}</td>
<td>-6.14</td>
<td>2.87</td>
<td>0.818</td>
<td>82</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>0/3</td>
<td>1.00</td>
<td>7.79</td>
<td>10.09</td>
<td>0.795</td>
<td>13</td>
</tr>
<tr>
<td>(3) BIP-COOMe</td>
<td>2.48</td>
<td>—</td>
<td>2/0</td>
<td>2.46 x 10^{-5}</td>
<td>-6.29</td>
<td>2.21</td>
<td>0.891</td>
<td>94</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>0/2</td>
<td>1.00</td>
<td>6.15</td>
<td>8.98</td>
<td>0.253</td>
<td>11</td>
</tr>
<tr>
<td>(4) BIP-CNH$_2$</td>
<td>2.57</td>
<td>2.71</td>
<td>1/0</td>
<td>8.64 x 10^{-3}</td>
<td>-1.40</td>
<td>4.83</td>
<td>0.215</td>
<td>81</td>
</tr>
<tr>
<td>(5) BIP-CNEt$_2$</td>
<td>2.53</td>
<td>2.71</td>
<td>2/0</td>
<td>6.14 x 10^{-5}</td>
<td>-5.75</td>
<td>2.26</td>
<td>0.819</td>
<td>89</td>
</tr>
<tr>
<td>(6) BIP-CH$_2$NH$_2$</td>
<td>2.42</td>
<td>2.89</td>
<td>2/0</td>
<td>3.17 x 10^{-3}</td>
<td>-3.40</td>
<td>4.19</td>
<td>0.678</td>
<td>59</td>
</tr>
<tr>
<td>(7) BIP-CH$_2$NEt$_2$</td>
<td>2.52</td>
<td>2.89</td>
<td>1/0</td>
<td>1.05 x 10^{-3}</td>
<td>-2.69</td>
<td>4.49</td>
<td>0.237</td>
<td>41</td>
</tr>
<tr>
<td>(8) BIP-PF$_{10}$</td>
<td>2.52</td>
<td>—</td>
<td>1/0</td>
<td>1.00</td>
<td>1.44</td>
<td>5.00</td>
<td>0.988</td>
<td>84</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>1/0</td>
<td>4.42 x 10^{-3}</td>
<td>-3.21</td>
<td>2.80</td>
<td>0.908</td>
<td>15</td>
</tr>
</tbody>
</table>

$^{a}$Distances given in Å and energies in kcal/mol.
$^{b}$These are the dominant donor-acceptor distances, which have the greatest contribution to the overall standard rate constant.
$^{c}$ $\Delta \tilde{G}^{\ddagger}_{\mu\nu} = (\Delta \tilde{U}_{\mu\nu} + \lambda)^2 / 4\lambda$ and is the effective free energy barrier at $\eta = \varepsilon = 0$.
$^{d}$This is the percentage contribution to the overall standard rate constant at the dominant donor-acceptor distance, $k_{k^{\text{PT}}}^{\Lambda}(\eta = 0; R)$, excluding contributions less than 10%.
A KIE of ~2 is predicted for 1–5. An analysis of these systems illustrates how the different factors contribute to the overall EPT process. Herein, we will focus on BIP-COOMe (3), but the results for the other systems are presented in Tables 8.4 and 8.5. At the dominant donor-acceptor distance ($R_{\text{ON}} = 2.48$ Å), the main contributions to the standard rate constant arise from the 1/0, 0/0, and the 0/1 pairs of states. As depicted in Figure 8.9, the first excited reactant state has substantial delocalization onto the acceptor side (near the nitrogen), resulting in significant overlap with the ground product state, which is localized on the acceptor side (near the nitrogen). This relatively large value of ~0.5 for the square of the overlap for the dominant contribution leads to a moderate KIE of ~2.

![Figure 8.9](image.png)

**Figure 8.9.** Proton potential energy curves and corresponding hydrogen vibrational wavefunctions for the reduced (blue) and oxidized (red) states of BIP-COOMe (3). The data are depicted for the main contributing pairs of reduced/oxidized vibronic states at the dominant proton donor-acceptor distance ($R_{\text{ON}} = 2.48$ Å), with the percentage contribution shown in the pie charts for the 1/0, 0/0, and 0/1 pairs.

In contrast, a KIE of unity is experimentally measured and calculated for the dyad BIP-PF$_{10}$ (8) compound. At the dominant donor-acceptor distance ($R_{\text{ON}} = 2.52$ Å), the main contribution to the standard rate constant arises from the 0/1 and 1/0 pairs of state. The hydrogen vibrational wavefunctions for these states are depicted in Figure 8.10. Perhaps most significant
is that for the dyad BiP-PF$_{10}$ (8), the square of the overlap integral for the dominant contribution is much larger ($S_{\mu\nu}^2 = 0.967$) than it is for the systems mentioned previously. For comparison, the square of the overlap integral for the 1/0 pair in BIP-COOMe (3) is 0.459. The large overlap is presumably a result of the more symmetric double well proton potential. An overlap of nearly unity for the dominant contribution leads to the KIE of unity.

![Proton potential energy curves and corresponding hydrogen vibrational wavefunctions](image)

**Figure 8.10.** Proton potential energy curves and corresponding hydrogen vibrational wavefunctions for the reduced (blue) and oxidized (red) states of BiP-PF$_{10}$ (8). The data are depicted for the main contributing pairs of reduced/oxidized vibronic states at the dominant proton donor-acceptor distance ($R_{ON} = 2.52 \text{ Å}$), with the percentage contribution shown in the pie charts for the 0/1 and 1/0 pairs.

As discussed elsewhere,$^{54}$ the porphyrin ring is expected to play a large role in the structural and electronic properties of the compound. Figure 8.11 depicts the spin density, which represents the localization of the unpaired electron upon oxidation. This figure illustrates that the porphyrin ring can be considered a non-innocent ligand due to the significant delocalization of the spin density onto the porphyrin ring in the oxidized state.
Figure 8.11. Isocontour plots (isovalue 0.005) of the spin density of BIP-COOME (3) and BIP-PF$_{10}$ (8) at the equilibrium geometries of the oxidized state, showing delocalization of the unpaired spin density onto the porphyrin ring of BIP-PF$_{10}$ (8). The spin density is plotted in the cyan (positive values) and magenta (negative values) wireframe. Color key: carbon, gray; hydrogen, white; oxygen, red; nitrogen, blue; chlorine, green.

Figure 8.12. Proton potential energy curve for oxidized state of BIP-PF$_{10}$ (8) at the dominant proton donor-acceptor distance ($R_{ON} = 2.52 \, \text{Å}$) and the relative spin density on the phenol/phenoxyl and porphyrin moieties along the proton transfer coordinate (left). The isocontour plots (isovalue 0.002) of the spin density for [BIP-PF$_{10}$]$^\bullet$ at the two minima along the potential energy curve are shown on the right. The spin density is plotted in the cyan (positive values) and magenta (negative values) wireframe. The left minimum corresponds to the oxidized state in which the proton has not transferred (i.e., on the oxygen), and the right minimum corresponds to the oxidized state in which the proton has transferred (i.e., on the nitrogen). In both cases the spin is delocalized onto the porphyrin ring; however, the predominant localization of the spin onto the porphyrin ring for the oxidized state with the proton localized on the donor stabilizes this state, which features an effectively neutral phenol moiety. Color key: carbon, gray; hydrogen, white; oxygen, red; nitrogen, blue; chlorine, green.

Further analysis of the spin density shows that the non-innocence of the porphyrin ring contributes to the more symmetric double well proton potential. Figure 8.12 depicts the proton potential at the dominant donor-acceptor distance ($R_{ON} = 2.52 \, \text{Å}$) for the oxidized state, along
with the ratio of spin density on the phenol/phenoxyl and porphyrin ring. The right minimum corresponds to the oxidized state in which the proton is localized on the nitrogen (i.e., the proton has transferred). In this state, the spin density is delocalized over the phenoxyl and porphyrin in a 55:45 ratio. The left minimum corresponds to the oxidized state in which the proton is localized on the oxygen (i.e., the proton has not transferred). In this state, the spin density is mainly localized on the porphyrin ring, with a 17:83 ratio. The predominant localization of the spin density on the porphyrin ring indicates that the phenol moiety is effectively neutral in this state (left minimum). In contrast, for the other BIP compounds, the left well in the oxidized state would feature a less stable phenol radical. Thus, the extensive delocalization of the spin density onto the porphyrin stabilizes the left well for the dyad BIP-PF$_{10}$ (8), leading to more symmetric proton potentials, greater wavefunction overlaps, and, ultimately, a smaller KIE of unity.

An intermediate KIE of ~1.5 is predicted for BIP-CH$_2$NH$_2$ (6) and BIP-CH$_2$NEt$_2$ (7). We will focus the analysis on BIP-CH$_2$NEt$_2$ (7), but the results for BIP-CH$_2$NH$_2$ (6) are shown in Tables 8.4 and 8.5. At the dominant donor-acceptor distances ($R_{ON} = 2.47$ Å and $R_{NN} = 2.84$ Å), the dominant contribution (92%) to the standard rate constant arises from the 1/0 pair of states, i.e., the first excited reactant state with the ground product state. As depicted in Figure 8.13, the first excited reactant state has substantial delocalization onto the acceptor sides (near the proximal imidazole nitrogen and the exocyclic nitrogen), resulting in significant overlap with the ground product state, which is localized on the acceptor side (near both nitrogen atoms). The square of the overlap integral ($S_{\mu\nu}^2 = 0.410$) is similar to that of the 1/0 pair for BIP-COOME (3) ($S_{\mu\nu}^2 = 0.459$). However, BIP-COOME (3) features greater participation from other vibronic states (i.e., the 1/0, 0/0, and 0/1 pairs contribute to the total rate constant), leading to smaller rate constants and a larger KIE for BIP-COOME (3) than BIP-CH$_2$NEt$_2$ (7).
Figure 8.13. Two-dimensional proton potential energy surfaces and contour plots (top) and corresponding hydrogen vibrational wavefunctions (bottom) for the reduced and oxidized states of BIP-CH$_2$NEt$_2$ (7). The data are depicted for the main contributing pair of reduced/oxidized vibronic states at the dominant proton donor-acceptor distances ($R_{ON} = 2.47$ Å and $R_{NN} = 2.84$ Å), the 1/0 pair, which contributes 92% to the total rate constant.
8.9. Conclusion

In this work, we explored fundamental principles involved in coupling a redox process to proton transfer(s) in a redox relay inspired by photosynthesis. Theory predicted that a concerted two-proton transfer process associated with the oxidation of a phenol that is hydrogen bonded to a benzimidazole–amine construct would reduce the redox potential of the phenol by ~300 mV and have a low KIE. Electrochemical, spectroelectrochemical, and KIE experimental data provide strong evidence that this E2PT process has indeed been observed. This is a powerful example of the predictive potential of theory, which clearly can be used to guide the rational design of artificial systems for energy conversion.

Protons are important in myriad ways for solar fuel production, beginning with their role in the catalysis of water oxidation to yield oxygen, protons, and reducing equivalents and ending with their role in the production of reduced carbon or hydrogen as fuels.\textsuperscript{15,36} Managing proton activity is key to achieving low overpotentials for proton-linked redox reactions and to providing the “redox leveling” necessary for carrying out redox reactions involving multiple electrons coupled to a fixed redox source.\textsuperscript{55} This study also helps address the thermodynamic consequences associated with the loss of the proton on Y\textsubscript{z} of PSII to the lumen in natural photosynthesis, as well as the thermodynamic cost in redox energy of moving protons in PCET processes via proton diffusion through a hydrogen-bonded network.

8.10. Experimental Methods

\textbf{Cyclic voltammetry.} Cyclic voltammetry was performed with a CHI 650C potentiostat (CH Instrument) with scan rate of 0.1 V s\textsuperscript{−1}. A custom glass cell with Teflon top accommodating a three-electrode setup and 4 mL of solution was used. The working electrode was glassy carbon (3 mm diameter) polished with alumina (1 \textmu m), thoroughly rinsed with ethanol and dried. The
counter electrode was a platinum grid and Ag/AgCl was used as a pseudo-reference electrode. It was placed as close as possible to the working electrode in order to minimize the ohmic drop. All the potentials were corrected using the ferrocenium-ferrocene redox couple as an internal standard (with $E_m$ taken as 0.31 V vs. SCE in acetonitrile and 0.45 V vs. SCE in dichloromethane). The compound of interest was dissolved at a concentration of 1 mM in acetonitrile or dichloromethane previously distilled and kept over molecular sieves and K$_2$CO$_3$. Tetrabutylammonium hexafluorophosphate was kept in an oven overnight and then used as a supporting electrolyte for all experiments at a concentration of 0.5 M. The solution was purged with argon for 2 min prior to any experiment and measurements were then taken under an argon atmosphere.

**Infrared spectroelectrochemistry (IRSEC) and FTIR.** The experimental setup and design of the IRSEC cell has been published by the Kubiak Lab.$^{55,56}$ A Pine Instrument Company model AFCBP1 potentiostat was employed. As the potential was scanned, thin layer bulk electrolysis was monitored by Fourier-Transform Reflectance IR off the working electrode’s surface. All experiments were conducted in 0.1 M TBAPF$_6$/dichloromethane solutions with analyte concentrations of ~3 mM (unless otherwise noted) prepared under a nitrogen atmosphere. The IRSEC cells used (working electrode/reference electrode/counter electrode) were glassy carbon/Ag/Pt, meaning that all potentials were in reference to a pseudo-reference electrode, bare metal Ag/Ag$^+ (~+200$ mV from the Fc/Fc$^+$ couple).$^{56-58}$ FTIR spectra were obtained in a Bruker Vertex 70 using a CaF$_2$ liquid transmission cell. The solvent (DCM) was distilled over CaH$_2$ and kept over molecular sieves and K$_2$CO$_3$. The compounds were dried under high vacuum overnight.

**KIE determination.** Variable scan rate cyclic voltammetry of the compounds at scan
rates varying from 0.1 to 500 V s\(^{-1}\) in dry acetonitrile was used to determine the KIEs with a potentiostat/galvanostat Model 273A (EG&G Princeton Applied Research) using a glassy carbon rod (Tokai 1mm diameter) as a working electrode, a platinum grid as a counter electrode, and a Ag/AgCl pseudo-reference electrode in a conventional three-electrode cell. To measure the apparent rate constant of the redox process, the electron transfer (ET) rate must be separated from the rates of diffusion processes. This is accomplished by increasing the scan rate until ET becomes rate determining. The peak potential separation is measured as a function of the scan rate. This peak separation is used to construct the “trumpet plot” from which the apparent standard rate constant was determined using the DigiSim software. The apparent standard rate constants were determined for the compounds dissolved in acetonitrile containing 2% (v/v) of either CH\(_3\)OH or CH\(_3\)OD methanol. The experimental KIE is estimated by the ratio of these apparent rate constants (\(k_{\text{appH}} / k_{\text{appD}}\)).

**Figure 8.14.** (A) CV of BIP-COOMe (3), with ferrocene as an internal standard. (B) “Trumpet plot”: anodic (upper symbols) and cathodic (lower symbols) peak potentials as a function of the log of scan rate. Simulations of the “trumpet plot” (solid lines) leads to the determination of the apparent standard rate constant. Solvent: acetonitrile + 2% MeOH, experimental peak potential (black squares) and corresponding simulation (black solid lines). Solvent acetonitrile + 2% MeOD, experimental peak potential (red circles) and corresponding simulation (red solid lines). Electrode: glassy carbon.

An example of one of the measurements is shown below. Cyclic voltammetry of BIP-COOMe (3) in acetonitrile exhibits a first reversible wave assigned to the phenoxy/phenol
couple at an apparent standard potential of 0.93 V vs. SCE (Figure 8.14A; black trace includes the phenol oxidation and the reference ferrocene; red and blue traces illustrate the irreversibility of the process(es) after reaching higher potentials).

To measure the apparent rate constant of the redox process, the electron transfer (ET) rate must be separated from the rate of diffusion processes, accomplished by increasing the scan rate until ET becomes rate determining. The peak potential separation increased from 80 mV at 2 V/s to more than 150 mV at 500 V/s. This peak separation is used to construct the “trumpet plot” (Figure 8.14B), and from the simulations the apparent standard rate constants for the electrochemical oxidation of the phenol were determined to be 0.128 ± 0.005 cm/s in the 2% MeOH solution and 0.071 ± 0.007 cm/s in the 2% MeOD solution, leading to a KIE of 1.8 ± 0.3.

8.11. Electronic Structure Calculations

The geometries were optimized using density functional theory (DFT) with the B3LYP density functional\textsuperscript{59,60} and the 6-31G** basis set\textsuperscript{61-64} for all atoms. The structures were optimized in acetonitrile using the conductor-like polarizable continuum model (C-PCM)\textsuperscript{65,66} with the Bondi atomic radii\textsuperscript{67} and including the nonelectrostatic contribution of dispersion,\textsuperscript{68,69} repulsion,\textsuperscript{68,69} and cavitation energies.\textsuperscript{70} For the simple BIP (1) and dyad BIP-PF\textsubscript{10} (8), the starting geometries for the optimizations were the crystal structures for the respective compounds.\textsuperscript{54,71-73} The starting geometries for BIP compounds 2–7 and 9–11 were obtained by modifying the crystal structure of 1 prior to full optimizations. For computational tractability, the C\textsubscript{6}F\textsubscript{5} groups were replaced with Cl groups for the dyad BIP-PF\textsubscript{10} (8) based on the similarity between their Hammett constants (\(\sigma_p = 27\) and 23,\textsuperscript{74} respectively), which represent their electron-withdrawing characters.
Thermochemical data were calculated at $T = 298.15$ K. All calculated free energies included zero-point energy, entropic contributions, and solvation effects. The solution phase free energies are calculated using the standard Gibbs relation $\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$. The zero-point energy and entropic contributions for the molecules are calculated from the vibrational frequencies of the normal modes using the harmonic model, and the solvation free energies are calculated with C-PCM. The relative redox potentials were calculated from the corresponding reaction free energies ($\Delta G^\circ$) using methodology described elsewhere.$^{75-78}$ The redox potentials were also calculated with five different functionals (BLYP,$^{59,79}$ M06-2X,$^{80}$ B3P86,$^{60,81}$ BP86,$^{79,81}$ and ωB97XD$^{82-85}$) and gave qualitatively similar results (Table 8.6). All calculations were performed using the Gaussian 09 electronic structure program.$^{86}$ The structures and energies of the systems studied are provided in Tables S7–S28.

### Table 8.6. Calculated and Experimental Redox Potentials

<table>
<thead>
<tr>
<th>Compound</th>
<th>Expt. $E_{1/2}$ (V vs. SCE)</th>
<th>B3LYP</th>
<th>BLYP</th>
<th>M06-2X</th>
<th>B3P86</th>
<th>BP86</th>
<th>ωB97XD</th>
</tr>
</thead>
<tbody>
<tr>
<td>(1) BIP</td>
<td>1.04$^a$</td>
<td>1.04$^b$</td>
<td>1.04$^b$</td>
<td>1.04$^b$</td>
<td>1.04$^b$</td>
<td>1.04$^b$</td>
<td>1.04$^b$</td>
</tr>
<tr>
<td>(2) BIP-COOH</td>
<td>0.90</td>
<td>0.98</td>
<td>0.99</td>
<td>0.99</td>
<td>0.98</td>
<td>1.00</td>
<td>0.98</td>
</tr>
<tr>
<td>(3) BIP-COOMe</td>
<td>0.93</td>
<td>0.96</td>
<td>0.98</td>
<td>0.97</td>
<td>0.96</td>
<td>0.99</td>
<td>0.96</td>
</tr>
<tr>
<td>(4) BIP-COHNH$_2$</td>
<td>0.92</td>
<td>0.94</td>
<td>0.97</td>
<td>0.93</td>
<td>0.94</td>
<td>0.95</td>
<td>0.93</td>
</tr>
<tr>
<td>(5) BIP-CONEt$_2$</td>
<td>0.92</td>
<td>0.95</td>
<td>0.94</td>
<td>0.94</td>
<td>0.96</td>
<td>0.95</td>
<td>0.94</td>
</tr>
<tr>
<td>(6) BIP-CH$_2$NH$_2$</td>
<td>0.56</td>
<td>0.66</td>
<td>0.71</td>
<td>0.58</td>
<td>0.64</td>
<td>0.64</td>
<td>0.66</td>
</tr>
<tr>
<td>(7) BIP-CH$_2$NEt$_2$</td>
<td>0.54</td>
<td>0.57</td>
<td>0.59</td>
<td>0.47</td>
<td>0.49</td>
<td>0.53</td>
<td>0.52</td>
</tr>
<tr>
<td>(8) BIP-PF$_{10}$</td>
<td>1.00$^a$</td>
<td>0.94</td>
<td>1.03</td>
<td>1.00</td>
<td>1.05</td>
<td>0.98</td>
<td>0.98</td>
</tr>
</tbody>
</table>

$^a$Data from Ref. $^{54}$. $^b$This couple was used as the reference reaction for all other calculated potentials, so the experimental and calculated values agree by construction. Because the experimental reference is $E_{1/2}$, the calculated potentials are also assumed to be $E_{1/2}$.

### 8.12. Calculating Electrochemical Rate Constants $k(R)$

The rate constant expression in Eq. 8.1 in the main text can be expanded to include the definition of the free energy barriers ($\Delta G^\ddagger$). For a fixed proton donor-acceptor distance $R$, the anodic and cathodic nonadiabatic rate constants are.$^{87-90}$
\[ k_{a}^{\text{EPT}}(\eta; R) = \sum_{\mu,\nu} P_{\mu} \left( \frac{V_{e}^{\text{el}} S_{\mu \nu}}{\beta' \hbar} \right)^{2} \sqrt{\frac{\pi}{\hbar k_{B} T \lambda}} \rho_{M} \int d\varepsilon [1 - f(\varepsilon)] \exp \left[ -\frac{(\Delta \tilde{U}_{\mu \nu} + \varepsilon - \eta + \lambda)^{2}}{4 \lambda k_{B} T} \right] \] 

\[ k_{c}^{\text{EPT}}(\eta; R) = \sum_{\mu,\nu} P_{\nu} \left( \frac{V_{e}^{\text{el}} S_{\mu \nu}}{\beta' \hbar} \right)^{2} \sqrt{\frac{\pi}{\hbar k_{B} T \lambda}} \rho_{M} \int d\varepsilon f(\varepsilon) \exp \left[ -\frac{(\Delta \tilde{U}_{\mu \nu} - \varepsilon + \eta + \lambda)^{2}}{4 \lambda k_{B} T} \right] \]  

where the double summations are over all pairs of reactant (\( \mu \)) and product (\( \nu \)) electron-proton vibronic states (i.e., proton vibrational states within the reactant and product ET states). The density of states \( \rho_{M} \) at the Fermi level, electronic coupling \( V_{e}^{\text{el}} \), and \( \beta' \) parameter representing the distance dependence of the electronic coupling are not required for the calculation of KIEs but would be required for the calculation of absolute rate constants.

The total reorganization energy \( \lambda \) is the sum of the outer-sphere (solvent) reorganization energy \( \lambda_{s} \) and the inner-sphere (solute) reorganization energy \( \lambda_{i} \). We calculated \( \lambda_{s} \) using a model in which the solute complex is placed in a spherical cavity of radius \( a \) on the surface of the electrode, where the distance \( d \) from the electrode is the cavity radius (\( d = a \)).

\[ \lambda_{s} = \frac{(\Delta q)^{2}}{2} \left( \frac{1}{\varepsilon_{\infty}} - \frac{1}{\varepsilon_{0}} \right) \left( \frac{1}{a} - \frac{1}{2d} \right). \]  

In this model, \( \varepsilon_{0} \) and \( \varepsilon_{\infty} \) are the static and optical dielectric constants of the solvent, respectively, \( \Delta q \) is the change in charge upon oxidation/reduction (\( \pm 1 \)), and the radius of the cavity \( a \) can be estimated from the volume of the molecular cavity from the C-PCM calculation. In addition, we calculated \( \lambda_{i} \) using the standard expression

\[ \lambda_{i} = \frac{1}{2} \left[ E_{\text{ox}}^{\text{red}} \left( R_{\text{eq}}^{\text{red}} \right) - E_{\text{ox}}^{\text{red}} \left( R_{\text{eq}}^{\text{ox}} \right) + E_{\text{red}}^{\text{red}} \left( R_{\text{eq}}^{\text{ox}} \right) - E_{\text{red}}^{\text{red}} \left( R_{\text{eq}}^{\text{red}} \right) \right]. \]  

The second and fourth terms are the energies at the optimized equilibrium geometry of the oxidized and reduced species, respectively. The first (third) term was calculated at the
optimized reduced (oxidized) geometry with the transferring hydrogen(s) optimized for the oxidized (reduced) state. As shown previously, the KIEs are relatively insensitive to the reorganization energy. The calculated reorganization energies are given in Table 8.7 for all the BIP systems.

Table 8.7. Inner-Sphere ($\lambda_i$), Outer-Sphere ($\lambda_o$), and Total ($\lambda$) Reorganization Energies$^a$

<table>
<thead>
<tr>
<th>Compound</th>
<th>$\lambda_i$</th>
<th>$\lambda_o$</th>
<th>$\lambda$</th>
</tr>
</thead>
<tbody>
<tr>
<td>(1) BIP</td>
<td>11.6</td>
<td>9.8</td>
<td>21.4</td>
</tr>
<tr>
<td>(2) BIP-COOH</td>
<td>12.5</td>
<td>9.5</td>
<td>22.0</td>
</tr>
<tr>
<td>(3) BIP-COOMe</td>
<td>12.5</td>
<td>9.4</td>
<td>21.9</td>
</tr>
<tr>
<td>(4) BIP-CONH$_2$</td>
<td>9.9</td>
<td>9.5</td>
<td>19.4</td>
</tr>
<tr>
<td>(5) BIP-CONEt$_2$</td>
<td>9.8</td>
<td>9.0</td>
<td>18.8</td>
</tr>
<tr>
<td>(6) BIP-CH$_2$NH$_2$</td>
<td>12.6</td>
<td>9.5</td>
<td>22.1</td>
</tr>
<tr>
<td>(7) BIP-CH$_2$N$\text{Et}_2$</td>
<td>14.2</td>
<td>9.0</td>
<td>23.1</td>
</tr>
<tr>
<td>(8) BIP-PF$_{10}$</td>
<td>9.8</td>
<td>7.1</td>
<td>16.9</td>
</tr>
</tbody>
</table>

$^a$Units of kcal/mol.

In Eqs. 8.3 and 8.4, the quantity $\Delta \tilde{U}_{\mu\nu}$ is defined to be $\Delta \tilde{U}_{\mu\nu} = \Delta U_{\mu\nu} + k_B T \ln(Q^{\text{II}} / Q^{\text{I}})$, where $\Delta U_{\mu\nu}$ is the energy difference between the vibronic states $\mu$ and $\nu$, and $Q^{\text{I}}$ and $Q^{\text{II}}$ are the total partition functions of the reduced and oxidized solute compounds, respectively, in bulk solution. In this definition, the vibronic energies are calculated relative to their respective ground states. Moreover, $P_\mu$, $S_{\mu\nu}$, and $\Delta \tilde{U}_{\mu\nu}$ depend on $R$, and we assume all other quantities are independent of $R$. The total rate constant can then be calculated using the results from Eqs. 8.3 and 8.4, along with the probability distribution function $P(R)$, by numerically integrating Eq. 8.2 in the main text. The numerical integration was done over the coordinate $\delta R$ for the single proton transfer systems and over the coordinate $2\delta R$ for the double proton transfer systems.

8.13. Generating Average Structures

We used average structures generated by averaging the optimized reduced ($1a$ or $1aa$), and oxidized ($2b$ or $2bb$) structures to calculate both the anodic and cathodic rate constants. The physical motivation for this procedure is that the electron transfer between the compound and the electrode is expected to occur when thermal fluctuations of the complex lead to a structure that is
in between the equilibrium reduced and oxidized structures. This approach is consistent with Marcus theory and the golden rule formulation, in which the nonadiabatic transition occurs at the crossing point between the initial and final states. To obtain the structures, a series of geometry optimizations was performed for the reduced and oxidized species with the donor-acceptor distance(s) constrained to values displaced incrementally (by 0.05 Å) from the equilibrium distance(s). For each of the donor-acceptor distances, the resulting reduced and oxidized structures were then aligned along the donor-acceptor axis in the same plane and averaged together to generate an average structure at that donor-acceptor distance.


For each of the average structures associated with a specific donor-acceptor distance, the position of the hydrogen was optimized for the reduced state (with the hydrogen on the donor) and the oxidized state (with hydrogen on the acceptor), while all other atoms were held fixed. A line connecting the positions of the hydrogens obtained from these constrained optimizations was then used to define the one-dimensional proton coordinate axis for each donor-acceptor distance. The proton potentials were generated on a one-dimensional grid along this axis for each donor-acceptor distance. Specifically, the hydrogen was moved along a grid of 24 points spanning this axis, a single point DFT calculation was performed at each position, and the resulting energies were then interpolated to generate a smooth potential energy curve. For the double proton transfer system, the two-dimensional proton potential energy surfaces were generated analogously, where each hydrogen moved along a one-dimensional grid.

The shape of the proton potential energy surfaces depends strongly on the proton donor-acceptor distance(s). Figure 8.15 depicts the one-dimensional proton potentials for 1–5 and 8. Figures 8.16 and 8.17 depict the two-dimensional proton potential energy surfaces and contour
plots for BIP-CH$_2$NH$_2$ (6) and BIP-CH$_2$NEt$_2$ (7), respectively, at the equilibrium geometries, as well as the one-dimensional slice along the diagonal of the two-dimensional proton potentials, which corresponds to the two protons moving simultaneously by the same amount. The proton vibrational wavefunctions and all input quantities for the rate constant calculations were determined from the proton potentials shown in Figures 8.15 and the full two-dimensional surfaces shown in Figures 8.16 and 8.17.

![Proton Potential Energy Curves](image)

**Figure 8.15.** Proton potential energy curves for the reduced and oxidized states of BIP (1), BIP-CONH$_2$ (5), BIP-PF$_{10}$ (8), BIP-COOH (2), BIP-COOMe (3), and BIP-CONEt$_2$ (4) for the average structures at fixed donor-acceptor distances ($R_{ON}$) given in the legend.
Figure 8.16. Two-dimensional proton potential energy surfaces and contour plots for the reduced and oxidized states of BIP-CH\textsubscript{2}NH\textsubscript{2} (6) at the average geometry at \(R_{\text{ON}} = 2.57\, \text{Å}\) and \(R_{\text{NN}} = 2.65\) Å, which are the equilibrium donor-acceptor distances for the reduced state (top). Diagonal slices of the proton potential energy surfaces for the reduced and oxidized states of BIP-CH\textsubscript{2}NH\textsubscript{2} (6) for the average structures at fixed donor-acceptor distances (\(R_{\text{ON}}\) and \(R_{\text{NN}}\)) given in the legend (bottom).
Figure 8.17. Two-dimensional proton potential energy surfaces and contour plots for the reduced and oxidized states of BIP-CH$_2$NEt$_2$ (7) at the average geometry at $R_{ON} = 2.57$ Å and $R_{NN} = 2.65$ Å, which are the equilibrium donor-acceptor distances for the reduced state (top). Diagonal slices of the proton potential energy surfaces for the reduced and oxidized states of BIP-CH$_2$NEt$_2$ (7) for the average structures at fixed donor-acceptor distances ($R_{ON}$ and $R_{NN}$) given in the legend (bottom).
8.15. Determining $P(R)$

The Boltzmann probability for sampling each donor-acceptor distance is given by $P(R)$, which is the classical harmonic probability distribution function:

$$P(R) = \frac{\exp\left[-k_{\text{eff}}(R - \bar{R})^2 / (2k_B T)\right]}{\int_{-\infty}^{\infty} \exp\left[-k_{\text{eff}}(R - \bar{R})^2 / (2k_B T)\right] dR}. \quad (8.7)$$

In our approach, the anodic and cathodic probability distribution functions were assumed to be the same, i.e., $P_a(R) \equiv P_c(R)$. We chose $\bar{R}$ to be the average of the donor-acceptor distance in the reduced and oxidized equilibrium structures. The effective force constant ($k_{\text{eff}}$) is obtained by projecting all of the normal modes onto the donor-acceptor axis and summing up the appropriately weighted force constants using methodology described elsewhere.$^{94,95}$ For the double proton transfer system, we include only the symmetric mode, in which both donor-acceptor distances, $R_1$ and $R_2$, increase or decrease concurrently by the same amount. The symmetric motion is expected to exert the greatest effect on the rate constant since both donor-acceptor distances should decrease for the protons to transfer simultaneously. In our approach, we used an average $\bar{k}_{\text{eff}}$ value obtained from the $k_{\text{eff}}$ values for the reduced and oxidized equilibrium structures. The equilibrium donor-acceptor distances and $k_{\text{eff}}$ values are given in Table 8.8.
<table>
<thead>
<tr>
<th>Compound</th>
<th>State</th>
<th>( R_{ON} ) (( \text{Å} ))</th>
<th>( R_{NN} ) (( \text{Å} ))</th>
<th>( \bar{R}_{ON} ) (( \text{Å} ))^a</th>
<th>( \bar{R}_{NN} ) (( \text{Å} ))^a</th>
<th>( k_{eff} ) (au)</th>
<th>( \bar{k}_{eff} ) (au)^a</th>
</tr>
</thead>
<tbody>
<tr>
<td>(1) BIP</td>
<td>Red</td>
<td>2.57</td>
<td>—</td>
<td>2.58</td>
<td>—</td>
<td>0.0451</td>
<td>0.0443</td>
</tr>
<tr>
<td></td>
<td>Ox</td>
<td>2.59</td>
<td>—</td>
<td>2.59</td>
<td>—</td>
<td>0.0435</td>
<td>0.0457</td>
</tr>
<tr>
<td>(2) BIP-COOH</td>
<td>Red</td>
<td>2.58</td>
<td>—</td>
<td>2.59</td>
<td>—</td>
<td>0.0488</td>
<td>0.0440</td>
</tr>
<tr>
<td></td>
<td>Ox</td>
<td>2.59</td>
<td>—</td>
<td>2.59</td>
<td>—</td>
<td>0.0426</td>
<td>0.0452</td>
</tr>
<tr>
<td>(3) BIP-COOMe</td>
<td>Red</td>
<td>2.58</td>
<td>—</td>
<td>2.59</td>
<td>—</td>
<td>0.0463</td>
<td>0.0459</td>
</tr>
<tr>
<td></td>
<td>Ox</td>
<td>2.60</td>
<td>—</td>
<td>2.60</td>
<td>—</td>
<td>0.0440</td>
<td>0.0452</td>
</tr>
<tr>
<td>(4) BIP-CONH₂</td>
<td>Red</td>
<td>2.57</td>
<td>—</td>
<td>2.60</td>
<td>—</td>
<td>0.0488</td>
<td>0.0459</td>
</tr>
<tr>
<td></td>
<td>Ox</td>
<td>2.63</td>
<td>—</td>
<td>2.60</td>
<td>—</td>
<td>0.0430</td>
<td>0.0452</td>
</tr>
<tr>
<td>(5) BIP-CONEt₂</td>
<td>Red</td>
<td>2.58</td>
<td>—</td>
<td>2.59</td>
<td>—</td>
<td>0.0458</td>
<td>0.0447</td>
</tr>
<tr>
<td></td>
<td>Ox</td>
<td>2.60</td>
<td>—</td>
<td>2.60</td>
<td>—</td>
<td>0.0435</td>
<td>0.0447</td>
</tr>
<tr>
<td>(6) BIP-CH₂NH₂</td>
<td>Red</td>
<td>2.57</td>
<td>2.86</td>
<td>2.61</td>
<td>2.85</td>
<td>0.0115</td>
<td>0.0115</td>
</tr>
<tr>
<td></td>
<td>Ox</td>
<td>2.65</td>
<td>2.83</td>
<td>2.61</td>
<td>2.83</td>
<td>0.0115</td>
<td>0.0115</td>
</tr>
<tr>
<td>(7) BIP-CH₂NEt₂</td>
<td>Red</td>
<td>2.57</td>
<td>2.94</td>
<td>2.61</td>
<td>2.93</td>
<td>0.0139</td>
<td>0.0129</td>
</tr>
<tr>
<td></td>
<td>Ox</td>
<td>2.65</td>
<td>2.91</td>
<td>2.61</td>
<td>2.91</td>
<td>0.0119</td>
<td>0.0129</td>
</tr>
<tr>
<td>(8) BIP-PF₁₀</td>
<td>Red</td>
<td>2.57</td>
<td>—</td>
<td>2.56</td>
<td>—</td>
<td>0.0460</td>
<td>0.0446</td>
</tr>
<tr>
<td></td>
<td>Ox</td>
<td>2.55</td>
<td>—</td>
<td>2.56</td>
<td>—</td>
<td>0.0433</td>
<td>0.0446</td>
</tr>
</tbody>
</table>

These values are averages of the values for the reduced and oxidized equilibrium geometries and are used to generate the probability distribution functions \( P(R) \).

### 8.16. Summary of Calculations

- **Step 1**: Optimize the reduced and oxidized structures and determine \( P(R) \).
- **Step 2**: Generate a set of average reduced/oxidized structures for a series of \( R \) values.
- **Step 3**: Determine proton coordinate axes by optimizing the hydrogen position for the reduced and oxidized states on the donor and acceptor, respectively, of the average structures.
- **Step 4**: Generate the proton potentials for the reduced and oxidized states.
- **Step 5**: Calculate the proton vibrational wavefunctions corresponding to the proton potentials in Step 4, and determine \( P_\mu, S_\mu, \) and \( \Delta \hat{U}_\mu \).
- **Step 6**: Calculate the anodic and cathodic rate constants using Eqs. 8.3 and 8.4 for each \( R \) value.
- **Step 7**: Calculate the standard rate constant using the results from Step 6, along with the probability distribution function, by numerically integrating Eq. 8.2. The standard rate
constant is approximated as $k_{\text{a EPT}}^{\text{L}}(\eta = 0)$ because $k_{\text{a EPT}}^{\text{L}}(\eta = 0) = k_{\text{c EPT}}^{\text{L}}(\eta = 0)$ within this formulation for this system.

8.17. References


