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Oxidation of Ferrocene Derivatives with Dibenzoyl Peroxide and Meta-Chloroperoxybenzoic Acid

Jos M. Halstead
Cleveland State University

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OXIDATION OF FERROCENE DERIVATIVES WITH DIBENZOYL PEROXIDE
AND META-CHLOROPEROXYBENZOIC ACID

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CLEVELAND STATE UNIVERSITY
DECEMBER 2018
DEDICATION

This work is dedicated to Dr. John Masnovi, who sadly passed away shortly before its completion. I will remember him for his insightful instruction, kind demeanor, and willingness to help anyone learn. He will be missed.
We hereby approve this dissertation for
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ABSTRACT

The chemical oxidation of ferrocene and related derivatives (RFc) via organic peroxides solvated in acetonitrile was studied spectrophotometrically by varying concentration and temperature to determine kinetics and activation parameters. The reaction rate of ferrocene with dibenzoyl peroxide depends strongly on whether electron withdrawing or donating substituents are present. Products were analyzed and the effect of different solvents on reactivity were studied. The rate law was first order in both oxidant and reductant, and steric and solvent effects are consistent with outer-sphere electron transfer (ET) as the rate-controlling step. B3LYP calculations were conducted to determine reorganization energies using Marcus theory and to examine molecular geometry and steric considerations.
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CHAPTER I

BACKGROUND

1.1 Introduction

Peroxide oxidations of iron compounds are important and ubiquitous processes that can be difficult to study. The multiple stable oxidation states of iron allow it to participate in numerous types of chemical reactions, often resulting in side reactions and byproducts when their compounds are examined directly [1]. In many cases, ferrocene can provide a simpler model to investigate the oxidations of iron(II). Reversible single electron transfers have been observed in ferrocene by electrochemical methods where an electrode serves as the electron acceptor, but the chemical kinetics for these types of oxidations have not been well explored [2]. Novel applications for ferrocene chemistry are emerging at an accelerating rate; it is therefore quite surprising that the reaction mechanisms and kinetics of ferrocenes in solution has received little attention. Ferrocenes are useful in biochemistry and have
been used in applications ranging from the investigation of metalloproteins in biological systems to the creation of new biosensors and probes [3]. Ferrocene also has industrial applications, and has recently been used to synthesize novel stimuli responsive and self-healing polymers [4]. Ferrocene has additional unusual properties as well. For example, straightforward aromatic substitution reactions are irregularly successful even though ferrocene possesses aromatic stability [5]. For example, ferrocene cannot be nitrated directly using nitronium-producing systems [6]. The reasons for this unusual behavior is not completely understood but may be related to the ease by which the products can be oxidized through electron transfer [7]. Given the surprisingly limited information on the chemical oxidation of ferrocenes, the goal for this work is to advance the understanding of these compounds in solution using organic peroxides as the oxidants.

1.2 Ferrocene

Ferrocene is classified as an iron (II) sandwich complex with two cyclopentadienyl (Cp) ligands which undergoes an easily identifiable color change upon oxidation in solution. Ferrocene was first prepared unintentionally in 1951 by Pauson and Kealy by the reaction of cyclopentadienyl magnesium bromide and ferric chloride, in an attempt to prepare fulvalene. Instead, they obtained an unusual orange powder with interesting properties including extreme stability [8]. This new material came to be known as ferrocene, although at the time the true nature of this molecule was not understood.
The structure of ferrocene, seen in Figure 1, was determined in 1952 by Woodward and Wilkinson with their analysis being based on reactivity [9].

![Figure 1: Structure of a ferrocene molecule.](image)

Their work was later confirmed by NMR and X-ray crystallography where the C–C bond distances within the five-membered rings are 1.40 Å, with Fe–C bond distances of 2.04 Å. The Cp rings in ferrocene are in a staggered ($D_{5d}$) conformation in the condensed phase, but it has been shown that in the gas phase they can assume an eclipsed ($D_{5h}$) configuration [10]. However, the energy barrier for interconversion between the eclipsed and staggered conformations is quite low due to the easy by which the Cp rings can rotate in relation to the rest of the molecule [10*]. The unusual and distinctive structure of ferrocene led to an explosion of interest in organometallic chemistry, and ferrocene has since been the subject of many studies looking into its reactivity, properties, and structural features.

The iron atom in ferrocene is usually assigned a +2 oxidation state, with each cyclopentadienyl ring possessing a single negative charge: $Fe^{2+}(Cp^-)_2$. This brings the
total number of π-electrons in each ring to six, making the rings aromatic according to Hückel’s rule. The twelve π electrons are also shared with the metal which, when combined with the six $d$-electrons provided by Fe(II), allow it to attain an 18-electron configuration, making ferrocene extremely stable. The 18-electron rule is used mostly for predicting stable metal complexes and is based on the valence shells of transition metals consisting of nine orbitals, which can collectively accommodate 18 electrons. The combination of these nine atomic orbitals with orbitals from the ligands gives rise to nine molecular orbitals. When all 9 orbitals are filled, the complex is effectively stabilized, with no empty bonding orbitals or unpaired electrons. Because of this, ferrocene is so stable that it can be sublimed without decomposition, withstand temperatures of up to 500°C, and resist reactions with acids and bases.

It has been reported that ferrocene usually reacts with electrophiles by substitution of the cyclopentadienyl rings, but it is possible the mechanism may differ from classical electrophilic aromatic substitution. It has been proposed that the electrophile first attacks at the metal center, but then is quickly transferred to the ligand, which subsequently is deprotonated [11]. This can happen due to the ease by which ferrocene is oxidized to form Fe(III) and Fe(IV) species. Electron transfer processes also occur easily and must be considered. In one example, the viability of electron transfer as a pathway for aromatic nitrations was demonstrated [12]. Ferrocene is particularly apt to undergo electron transfer because, with the removal of a single electron, a stable ferrocenium cation can form. Ferrocenium is easily identifiable due to a color change from orange to blue upon oxidation. Because of this, ferrocene is often used as an internal standard in non-aqueous electrochemistry. It has
been reported, through electrochemical measurements, that substituents on the
cyclopentadienyl rings can change the redox potential of ferrocene: electron-
withdrawing groups, such as acyl groups, shift the potential in the anodic direction and
electron-releasing substituents, such as methyl groups, shift the potential in the
cathodic direction [13].

In recent years bioorganometallic chemistry has developed as a rapidly
growing and maturing area which links classical organometallic chemistry to biology,
medicine, and biotechnology [14-16]. The use of ferrocene has seen considerable
interest due to the stability of the ferrocenyl group in aqueous aerobic media, the
accessibility of a large variety of derivatives, and its favorable electrochemical
properties. Recently ferrocene-based molecules have become very popular for
conjugation with biomaterials and for modeling enzymatic behavior [17]. One
interesting example is horseradish peroxidase, a monomeric, heme-containing
glycoprotein consisting of 308 amino acid residues. Typically, only a narrow range of
molecules can oxidize this peroxidase, whereas a large variety can act as reductants.
Several ferrocene derivatives have been shown to be suitable electron-donor substrates
for studying this glycoprotein [18-21]. Other studies have also suggested that electron-
transfer reactions between inorganic complexes and metalloproteins can occur at the
partly exposed heme edge of cytochrome c, and that small metal complexes with
hydrophobic ligands, into which the metal’s electrons can delocalize rapidly, facilitate
these electron transfers [22]. Ferrocene could make an excellent compound to model
this process as well.
Ferrocene can also be incorporated into functional materials, such as zeolites, glassy materials, oxide surfaces, and others [23]. This has potential applications in fields such as catalysis, sensors, and optical devices. Also of interest are reports that the pyrolysis of ferrocene in an argon atmosphere can yield large amounts of carbon nanotubes [24-26]. Overall interest in ferrocene is exploding. Figure 2 shows that publications involving ferrocene and related applications have experienced rapid and increasing growth since its discovery in 1951.

![Figure 2. Publications involving ferrocene by year according to CAS.](image)

1.3 Redox-responsive polymer gels

Redox-responsive gels (RRGs) are materials that can change their properties in response to external stimuli, and they have attracted considerable attention in recent years. Ferrocene is one of the most widely used redox centers and can play a key role
in the fabrication of new RRGs [27]. Generally, polymer gels are solvent-containing, dispersed systems in which polymer chains are linked to each other to form a network. Some types of polymer gels can change their properties in response to light, temperature, electric or magnetic fields, pH, and redox conditions. Ferrocene is a logical choice for redox responsive polymer systems. The oxidation of ferrocene polymers can alter material properties such as hydrophilicity, color, morphology and more.

There are three different categories of ferrocene-based responsive polymers: main chain-, sidechain-, and end chain-based systems. One of the best methods to create main chain ferrocene polymers is through a ring opening polymerization using strained silicon bridged ferrocenophanes [28], see Figure 3. This process can then be used to produce high molecular weight poly-(ferrocenylsilane) (PFS) polymers, [29]. These PFS chains can then be further cross-linked to improve mechanical properties and enhance thermal stability. Once cross-linked, PFS materials can readily absorb solvents to become gels.

Figure 3. Ring opening polymerization used to create main chain ferrocene polymers.
Side chain ferrocene polymers can be created through the incorporation of ferrocene in the pendant groups of various polymer backbones, such as polyacrylamide, polyacrylic acid, polyethyleneimine, cellulose, and others. This category of materials is of interest due to the broad range of compatible polymers, which allows for specifically targeted material properties. Ferrocene can be incorporated by direct chemical modification of polymer chains, or by the use of ferrocene-containing monomers during initial synthesis. This approach has produced polymers that have shown volume change and changes in phase transition temperature in response to redox stimuli. In one example, Tatsuma et al. reported the synthesis of a redox and thermally responsive hydrogel by the copolymerization of N-isopropylacrylamide and vinylferrocene, with N,N'-methylenebisacrylamide [30].

Below the phase transition temperature of the material, the amine and carbonyl groups form hydrogen bonds with water molecules and the gel swells to an increased volume. However, at higher temperatures the material contracts due to increased hydrophobic interactions. When the ferrocenyl moieties are oxidized, water becomes even more strongly incorporated due to the exceedingly hydrophilic nature of the ferrocenium cation. This significantly increases the phase transition temperature of the polymer. The result of the incorporation of ferrocenyl groups is a material that has a reversible expansion-contraction capability controlled electrochemically. This concept was advanced further by Kaniewska et al., who created an electrically and thermal
responsive gel based on isopropylacrylamide, cysteine, and ferrocene. This material was capable of redox-induced phase transitions in the range of human body temperature [31]. The efficacy of this phase transition is shown in Figure 4, where the volume of the ferrocene gel is changed to a large degree when oxidized.

![Figure 4](image)

**Figure 4.** Stimuli responsive polymer swelling as a function of temperature.

The final category of responsive ferrocene polymers contains terminal ferrocene groups which can offer additional varieties of cross-linking and associative behavior. These types of polymers have even demonstrated self-healing properties. In work done by Nakahata et al., a terminal ferrocene-polymer was fabricated by mixing cyclodextrin-modified poly(acrylic acid) and ferrocene-modified poly(acrylic acid) to form a network [32]. The result is when this polymer is oxidized or reduced it can be
regenerated. This was demonstrated by cutting the polymer into two parts, rejoining
them, and exposing them to a redox stimulus. After 24 hours the cut was gone, and the
original tensile strength was restored. This redox-induced healing is prevented by
stronger chemical interactions such as covalent bonds; it is the weak association
between ferrocene and cyclodextrin that allows for regenerative behavior. These types
of responsive systems show promise for applications such as tissue engineering and
controlled drug release.

Ferrocene-cyclodextrin polymers have also been used to perform mechanical
work through induced volume changes. In one experiment, a weight was raised and
lowered simulating the behavior of a muscle. Shape memory properties are also
possible by incorporating both reversible ferrocene-cyclodextrin-crosslinks with more
rigid covalent crosslinks in the same system. This was demonstrated by Dong et al., by
cross-linking chitosan, which contains cyclodextrin groups, with a ferrocene
containing ethylene imine polymer [33]. Upon reduction and oxidation of the
ferrocene the crosslink density was modified resulting in different volumes and
mechanical properties for the resultant film. To demonstrate shape memory the
material was first oxidized, destroying the ferrocene-cyclodextrin crosslinks. This
showed a visible distortion (bending) of the substrate due to the ferrocenyl moieties
changing their positions relative to the cyclodextrins. The distorted shape was
maintained until a reductive stimulus was applied, upon which the polymer recovered
to its original configuration.
1.4 Fenton’s Reagent

Another area where the oxidative behavior of iron is of considerable interest is in the chemistry of Fenton's reagent. Fenton's reagent was first developed in the 1890s by Henry Fenton for analytical uses, but it has since been found to be exceptionally effective at destroying persistent organic pollutants such as trichloroethylene and tetrachloroethylene in both water and soil [34]. As seen in Equation 1, the Fenton reaction is defined as the oxidation of ferrous iron ($Fe^{2+}$) with hydrogen peroxide producing ferric iron ($Fe^{3+}$) and hydroxyl radicals ($OH^-$), which then further react with $H_2O_2$ to produce superoxide ($O_2^-$). A cascade of additional reactions can then occur which ultimately degrades contaminants to carbon dioxide and water [35].

$$Fe^{2+} + H_2O_2 \rightarrow Fe^{3+} + OH^- + \cdot OH$$
$$Fe^{3+} + H_2O_2 \rightarrow Fe^{2+} + \cdot OO\text{H} + H^+$$
$$H_2O_2 + OH^- \rightarrow \cdot OO\text{H} + H_2O$$
$$\cdot OO\text{H} \rightarrow O_2^- + H^+$$

(1)

Overall the number of reactions associated with Fenton’s reagent are many, and the exact mechanisms by which they occur are not well understood. The reason Fenton’s reagent is so effective for environmental remediation is due to the formation of the hydroxyl radical, which is a powerful oxidant. The destruction of organic compounds by Fenton’s reagent is typically rapid and exothermic, with rate constants for the reaction of hydroxyl radicals with common pollutants typically being in the range of $10^7$ to $10^{10} M^{-1} \cdot s^{-1}$ [36].

Classical Fenton chemistry does have its disadvantages, such as a narrow pH window. The optimal pH range for Fenton's chemistry to take place is between 3 and
6. If the pH becomes more basic then iron can transition from a hydrated ferrous ion to a colloidal ferric state. One this occurs it can cause peroxide to decompose without the formation of reactive radicals, the source of efficacy. Another disadvantage is traditional Fenton’s catalysts cannot be reused. Furthermore, the reactivity of iron with numerous other molecules, including oxygen, can cause the formation of iron sludge after a Fenton’s process completes. When used on a large scale, this byproduct is a kind of pollution itself and can affect water quality. The reactivity of iron also makes studying Fenton-type mechanisms very challenging, due to the many possible side reactions. To avoid some of these problems, recent studies have focused on using ferrocene as the iron source for these reactions. In one example, an immobilized heterogeneous ferrocene catalyst was created that has been shown to be effective for catalyzing the degradation of persistent organic pollutants [37]. In work done by Wang et al., the degradation of methylene blue by a ferrocene catalyst (1.5 mol% loading) with hydrogen peroxide (ten equivalents) was investigated. The results, shown in Figure 5, show that this combination degrades methylene blue quickly whereas there is little to no effect when the two materials are used alone.
Figure 5. Degradation of methylene blue as a function of time by hydrogen peroxide with a ferrocene catalyst.

To address issues of stability and reusability, the immobilized ferrocene catalyst used for an initial round of catalysis were filtered, rinsed with water, dried at room temperature, and reused for second and third rounds of testing. It was observed that the recycled ferrocene particles still displayed high catalytic activity, even after multiple cycles.
CHAPTER II

ELECTRON-TRANSFER PROCESSES

2.1 Oxidation-Reduction processes

An oxidation-reduction (redox) process is a chemical reaction in which the oxidation number or state of a molecule, atom, or ion changes. This change in oxidation state may or may not involve a direct electron transfer step. Discreet electron transfers can be observed in outer sphere type mechanisms, but in other cases redox reactions may only involve closed shell intermediates. An example of the latter is the concerted oxidative addition of metal complexes. Redox reactions are one of the broadest and most important reaction types in chemistry, encompassing all processes that involve molecular oxygen, combustion, corrosion, respiration, photosynthesis, the reactions of electrochemical batteries, and more. An oxidation number is an indicator of the total number of electrons which have been removed or added for an atom to get to its present state. This number may be positive, negative, or zero, and represents the hypothetical charge of an atom if all bonds were 100% ionic, even though this limiting case is never exactly true for real bonds. By convention redox reactions are usually divided into half reactions to aid understanding where the reduced species gains electrons with a decrease in oxidation number, while the oxidized species loses electrons with an increase in oxidation number. Redox reactions can occur at almost any rate, ranging from very slow, in the case of most metallic corrosion, to very fast, as in combustion processes.
Electron transfer between metal complexes is an interesting type of redox reaction that can proceed through either an outer sphere or inner sphere mechanism. Although it is not generally true that thermodynamics influence kinetics, this is often the case for electron transfer reactions in solution. The reasons for this behavior can be explained by Marcus theory, which builds upon the Arrhenius equation and Eyring-Evans-Polanyi transition state theory.

2.2 Transition State Theory

The temperature dependence of reaction rates was an area of interest for Arrhenius, Eyring, and others. The equation Arrhenius proposed was the first of its kind to describe the kinetics of gas-phase reactions. His equation is based on empirical observations, as well as by analogy to the thermodynamic equation

\[ K = e^{-\Delta G^*/RT} \]

and takes the following form [38].

\[ k = Ae^{-Ea/RT} \quad (2) \]

The Arrhenius equation was a significant advancement and successfully described the behavior of gas phase reactions, but a new way of thinking was needed to describe more complex reactions in solution. The concept that emerged was the potential energy surface put forth by René Marcelin in 1913. He theorized that the progress of a chemical reaction could be described as a point on a curve with coordinates in atomic motion and distance. Henry Eyring and Michael Polanyi advanced this idea in 1931 by constructing a three-dimensional potential energy diagram for diatomic hydrogen,
H + H₂ → H₂ + H. Unlike Arrhenius, they based their model on experimental data involving vibrational frequencies and energies of dissociation, as well as on quantum theory. This was further validated a year later when Pelzer and Wigner were able to track the progress of a reaction along a potential energy surface where they concluded that the reaction rate is ultimately determined by the motion of a system though a saddle point [39].

This set the stage for Eyring to derive his famous equation based on the idea that reactants must surpass a high energy intermediate state on the way to form products. This energy barrier determines the minimum energy necessary for a reaction to occur and is thus referred to as Gibbs energy of activation. Figure 6 shows the energy diagram of a molecule as a function of reaction progress.

![Figure 6. Example of a potential energy diagram.](image-url)
As a reaction proceeds from left to right across the diagram, the reactants approach an energy barrier that they must overcome to proceed to products. This state represents an activated complex which can only form if the reactants possess sufficient energy. The activated complex is inherently unstable, and at this point the reaction must either proceed to generate product or fall back down to the reactant state. A key principle is the assumption of a thermodynamic equilibrium between the activated complex and the reactant(s). Thus, the overall rate of reaction is proportional to the concentration of the species in this transition state [39].

The Eyring equation is based on this principle by considering a bimolecular reaction with an equilibrium constant $K$.

$$A + B \rightarrow C$$

$$K = \frac{[C]}{[A][B]} \quad (3)$$

The commonly used form of the Eyring equation is seen below.

$$k = \frac{k_B T}{h} e^{-\frac{\Delta H^\ddagger}{RT} - \frac{\Delta S^\ddagger}{R}} \quad (4)$$

where $k_B$= Boltzmann’s constant, $h$ = Planck’s constant, $T$= absolute temperature, and $R$= ideal gas constant
The Eyring equation can also be applied in its linear form, shown in Equation 5, which is useful because values of $\Delta H^\ddagger$ and $\Delta S^\ddagger$ can be determined experimentally from a linear plot of $\ln(k/T)$ vs $1/T$.

$$\ln \frac{k}{T} = -\frac{\Delta H^\ddagger}{R} \frac{1}{T} + \ln \frac{k_B}{h} + \frac{\Delta S^\ddagger}{R}$$

2.3 Marcus Theory

It could be said that the underlying principles of the kinetics of electron transfer processes were not truly understood until the work of Rudolph A. Marcus [40-46]. He began his pivotal work in 1956, first to explain the rates of outer sphere electron transfer reactions, but later expanding it to include inner sphere contributions as well. Marcus built on the work of Eyring and Arrhenius, with the main difference between their theoretical approaches being that transition state theory is meant to deal with reactions that involve structural changes when the reactants are strongly coupled. In outer-sphere electron transfer, no bonds are formed or broken, only a single electron is transferred, and therefore a different way of thinking was needed. Marcus saw electron transfer as a series of three steps. Step one is the formation of an electron donor-acceptor complex. In step two, the electron is passed from donor to acceptor. Step three is then the separation of products.

A simple example is the Fe$^{2+}$/Fe$^{3+}$ self-exchange reaction, for which the Gibbs energy of activation can be calculated based on the temperature dependence of
the reaction rate. In transition state theory, this would be plotted on a diagram as energy vs. reaction coordinate. The reaction coordinate can be thought of as the change in bond angles and distances during the reaction. Prior to Marcus theory, these were the terms in which electron transfers were usually considered. Alternatively, electron transfers were sometimes described in vibrational terms, as in the work by Franck and Condon [47]. Marcus disagreed with both interpretations. He believed that outer sphere reactions cannot follow a well-defined coordinate path because nuclear motion is too slow when compared to the speed at which the much less massive electrons move. Furthermore, Marcus knew that the vibrational interpretation used in the Franck-Condon principle did not fully explain matters, as electron transfer reactions could occur in the dark, with no source of vibrational excitation at room temperature, even though an activation energy was still observed. The way he resolved these inconsistencies was by highlighting the importance of the solvent environment.

Starting with the theory of dielectric polarization of solvents, Marcus demonstrated that the energy of activation could be calculated using the polarization properties of the solvent, the size of the reactants, and the electron transfer distance alone. He derived a general equation for the transfer of charge between any two bodies of a specified shape and distance by imagining two conducting spheres at a set distance where a variable amount of charge can be reversibly exchanged, as shown in Figure 7.
In the first step, the energy $W_1$ is calculated for the system when both spheres carry half of the amount of charge which is to be transferred. This state is reached by transferring the charge from the donor sphere to the vacuum, and then to the acceptor sphere. This movement of charge induces an electric field which then polarizes the solvent. In a second step, the energy of the reverse transfer, $W_2$, is calculated through the vacuum again, but the atomic orientation and solvent polarization is kept fixed.
The system is now in the required state for an electron transfer, and the sum of energy terms $W_1 + W_2$ is defined as the Gibbs free energy, $\Delta G$. This results in an elegant simplification where the coordinates of all solvent molecules converge into a single polarization state based on the amount of transferred charge. Applying this model at the molecular level creates the problem that the charge can no longer be transferred in any amount, but only as quantized single electron packets. However, since polarization is still determined by the total solvent environment, it can be still treated classically.

When Marcus plotted his reactants and products as two intersecting parabolas, he found that the standard Gibbs free energy $\Delta G^0$, and the reorganization energy $\lambda$, could be considered in terms of the Gibbs energy of activation, $\Delta G^\ddagger$, see Figure 8.

![Figure 8. Potential energy diagram put forth by Marcus.](image-url)
Marcus defined reorganization energy, $\lambda$, as the amount of work it would take to force the reactants and the solvent environment into the same configuration as the products without an electron transfer taking place. He then proposed that $\lambda$ was comprised of the sum of nuclear ($\lambda_i$) and outer sphere ($\lambda_o$) energy contributions, see Equation 6. The nuclear reorganization energy refers to the energy of the inner shell of atoms, close to the redox center, and can be treated as a distance-independent parameter. The outer sphere reorganization energy applies to the solvent environment and is strongly dependent on the distance between donor and acceptor, solvent polarity, and the radius of the reactants, see Equation 7. Marcus also put forth general equations for activation energy and rate constants for electron transfers as a function of $\lambda$ and $\Delta G^\circ$, as shown in Equations 8 and 9.

$$\lambda = \lambda_i + \lambda_o$$  \hspace{1cm} (6)

$$\lambda_o = (\Delta e)^2 \cdot \left( \frac{1}{2a_1} + \frac{1}{2a_2} - \frac{1}{r_{DA}} \right) \cdot \left( \frac{1}{D_{op}} - \frac{1}{D_s} \right)$$  \hspace{1cm} (7)

$a_1 =$ radius of donor \hspace{1cm} $D_{op} =$ optical dielectric constant $= \eta^2$

$a_2 =$ radius of acceptor \hspace{1cm} $D_s =$ static dielectric constant

$r_{DA} =$ distance between donor and acceptor \hspace{1cm} $e =$ transferred charge

$$\Delta G^\dagger = \frac{(\Delta G^\circ + \lambda)^2}{4\lambda}$$  \hspace{1cm} (8)
\[
    k = Ae^{-\frac{(\Delta G^0 + \lambda)^2}{4\lambda RT}}
\]

In most chemical reactions the transition state has to meet both structural and energetic requirements where atomic nuclei shift in such a way that the nuclear configurations are in equilibrium. For these types of reactions, the solvent may only have a minor effect. In outer sphere electron transfers the situation is reversed: nuclear displacement is typically small, and the solvent effect is large. In the same way that atomic nuclei must obtain an equilibrium state between that of product and reactant, so must the solvent molecules for electron transfer. The ideal solvent arrangement should correspond to a situation in with half the charge is on the donor and half on the acceptor. However, an electron cannot be divided, so the real solvent configuration must be in a quasi-equilibrium state with an electron residing asymmetrically on one reactant. This quasi-equilibrium requires energy to create and is typically provided thermally. Furthermore, the creation of the correct solvent arrangement and the electron jump are decoupled and do not happen in a synchronous process. Instead they are stepwise events, with solvent polarization being the rate-limiting step.

In the case of an adiabatic system where reactant coupling is significant, and the energy gap is much larger, Marcus theory is not generally applied. The non-adiabatic systems for which Marcus theory is most useful are those in which state change occurs non-radiatively. These processes are often observed in UV/VIS photochemistry, collisions of electronically excited species, chemiluminescent reactions, and electron transfer processes [48].
For these groundbreaking ideas Marcus was awarded the Nobel Prize in Chemistry in 1992. Since its inception, Marcus theory has been expanded considerably, and it is now used to describe numerous important processes in all branches of chemistry.

2.4 Oxidative Addition Processes

In the most general terms, oxidative addition is the addition of a ligand molecule to a transition metal complex. In this process, the metal center is oxidized by two electrons, increasing its oxidation state by two. In most cases, the coordination number of the metal center increases by two as well. Oxidative addition is the reverse of reductive elimination, and it typically requires a good redox couple with the reactants and products both in stable oxidation states. For example, oxidative addition of Fe(II) to Fe(IV) can occur readily, but oxidative addition of Fe(III) to Fe(V) is quite unfavorable due to the instability of the +5 oxidation state for Fe. Furthermore, oxidative addition cannot occur if the metal is already in its highest oxidation state, as there would be no more valence electrons for it to lose, and the loss of core electrons is energetically implausible. Conversely, a metal center becomes more reactive towards oxidative addition the more reduced it is. In general oxidative addition is usually favored by electron donating ligands because they help stabilize the increased oxidation state of the central metal. Oxidative addition is sometimes thought of an
elementary organometallic reaction, but it is not, because it can occur through a variety of different pathways [48].

One such possibility is the concerted mechanism, as seen in Figure 9, which involves complex formation followed by ligand insertion.

![Figure 9. The concerted mechanism of oxidative addition.](image)

In the first step of the concerted mechanism, the metal approaches the A–B bond and forms a complex. Then if the metal donates enough electron density, new σ bonds are formed, and the metal is formally oxidized. Depending on the nature of the ligands, they may or may not approach the metal symmetrically, and if the metal is a poor nucleophile the reaction may not proceed at all. Although there are many examples of concerted oxidative additions in the literature, one of the most-studied cases is the addition of H2 to the 16-electron, square planar d8 species IrCl(CO)(PPh3)2. This reaction, shown in Figure 10, was discovered by Vaska and DiLuzio in 1961, and gives the 18-electron d6 octahedral dihydride IrCl(H2)(CO)(PPh3)2 as the product [49].
A non-concerted, $S_N2$-type mechanism is also possible for an oxidative addition reaction, as shown in Figure 11. In the $S_N2$ pathway, an electron pair on the metal directly attacks the A–B $\sigma$ bond of the least electronegative atom. This type of mechanism is often seen with the addition of methyl, allyl, acyl, and benzyl halides. Similar to the concerted mechanism, this is a second-order reaction which is favored in polar solvent and often displays a negative $\Delta S^\ddagger$ value, consistent with an ordered transition state. As would be expected in a $S_N2$-type process, inversion of configuration of the atom adding to the metal can be observed. In the first step of this mechanism, the oxidation number of the metal increases by two units but the electron count remains the same.
Figure 11. Mechanism for a $S_N2$-type oxidative addition.

In the second step, the electron count is increased by 2 with no further change in the oxidation state. In the case when an 18e complex is involved, it is possible for the first step to occur without the loss of a ligand, because only the second step requires a vacant 2 electron site. The more nucleophilic the metal, the greater its reactivity towards $S_N2$-type oxidative additions, but steric hindrance can slow the rate of reactions considerably [50].

Radical mechanisms of oxidative additions have been observed, but they are more ambiguous and harder to analyze than the concerted or $S_N2$ pathways [51]. One possible reason for this is that minor changes in the structure of the reactants, the formation of complexes, or even minor impurities in the reagents or solvents, can sometimes be enough to change the mechanism to an entirely different type. Two types of radical additions are possible: chain and non-chain processes. The non-chain variant, as seen in Figure 12, is believed to occur with the additions of certain alkyl halides such as bromoalkanes.

![Figure 12](image)

Figure 12. The mechanism of a radical non-chain oxidative addition.

The key feature is an electron transfer from the metal to the acceptor to form $M^{+\ast}$ and a radical anion $R-X^{-\ast}$. This is the rate limiting step. The $X^{-}$ is then transferred to the
metal and an alkyl radical is released. The two radicals then react with each other to yield the final product. The non-chain radical mechanism is faster for more basic metals, and for systems that can offer additional stability to radical species.

The chain radical process, seen in Figure 13, requires an initiator to start which usually leads to an induction time.

![Figure 13. The mechanism of a radical chain oxidative addition.](image)

Here the radical first associates with the metal and abstracts X• from the halide, leaving behind the chain carrier R•. This results in stereochemical scrambling at the α carbon of the alkyl group in the radical intermediate.

The final possible mechanism of oxidative addition is the ionic pathway, seen in Figure 14. This processes commonly occurs with the addition of hydrogen halides (HX) after they disassociate into their ionic forms. It is possible for the hydrogen and the halide to add in either order, since they add in separate steps. The actual order of addition is based on solvent polarity, how basic the ligands are, and the oxidation state of the metal.
It is more common for the hydrogen to add first due to the electron-rich nature of metal centers able to undergo oxidative addition. In this case, protonation is the rate-determining step. The rarer case in, which the halide ion adds first, is favored by more electrophilic ancillary ligands and follows a rate law dependent upon the halide concentration.

Figure 14. Ionic mechanism of oxidative addition.
3.1 Methodology

There are several plausible mechanisms for the chemical oxidation of ferrocene: oxidative addition, inner sphere electron transfer, and outer-sphere electron transfer. To investigate these processes, a series of redox reactions was conducted with different ferrocene derivatives, varying temperature and concentration for each. In all potential mechanisms, electron-donating substituents on the cyclopentadienyl rings of ferrocene should promote the reactions, and electron withdrawing substituents should inhibit them. However, if a coordination type process is occurring significant steric effects should also be observed if the substituents are sufficiently bulky. In this study, decamethylferrocene and ethylferrocene serve as electron donating moieties, and bromoferrocene and acetylferrocene serve as electron-withdrawing substituents. Unmodified ferrocene is used as a control. Upon oxidation, ferrocenium cations strongly absorb visible light with a wavelength of 617 nm. However, substituent effects can shift absorption peaks to some degree, therefore, reaction rates were monitored using UV-visible spectrophotometry by tracking the change in absorbance at five different wavelengths.

Reaction rates were observed to be first order with respect to ferrocene and peroxide, and second order overall, $d[RFe^+] / dt = k[RFe][peroxide]$. For this reason, reactions were conducted under pseudo-first-order conditions, to determine the rate
constant $k'$. The second-order rate constants, $k = \frac{1}{2} k'[\text{RFc}]$ are obtained from the slopes of $k'$ vs. [RFc] plots. The factor of one-half accounts for the stoichiometry of two moles of [RFc] to one mole peroxide,

\[ 2 \text{RFc} + (\text{C}_6\text{H}_5\text{CO}_2)_2 \rightarrow 2 \text{RFc}^+ + 2 \text{C}_6\text{H}_5\text{CO}_2^- \]  

(10)

Eyring plots of $\ln (k/T)$ vs. $1/T$ were constructed, from which the activation parameters $\Delta H^\dagger$, $\Delta S^\dagger$ and $\Delta G^\dagger$ were derived. Product isolation, conformation, and quantification was conducted, and the effects of different solvents on the reaction rate were determined.

### 3.2 Synthesis of Acetylferrocene

Acetylferrocene was synthesized following the procedure reported by Graham and Lindsey, see Figure 15 and 16 [52]. First, 6.0 g ferrocene and 20 mL acetic anhydride were added to a 200 mL flask. Then 5.0 mL 85% phosphoric acid was added while stirring. The solution was mixed for 20 minutes in a 90 °C water bath, and then poured into a 500 mL beaker containing 80 g ice. After the ice melted, 30 g NaHCO$_3$ was added to raise the pH to 5-6. The solution was then filtered, and the product washed with cold water and dried. The product then underwent separation using a chromatography column with activity grade III alumina and a 1:1 ethyl acetate-hexane solution as the liquid phase. The solvent was removed by evaporation,
and the melting point was measured to be 83-86 °C (literature value = 83-85°C) [52].

The purity was also confirmed by NMR, see Figure 17. The yield was 4.6 g (62%).

Figure 15. Reaction scheme for acetylferrocene synthesis.

Figure 16. Synthetic mechanism for acetylferrocene.
3.3 Synthesis of Bromoferrocene

Bromoferrocene was synthesized following the procedure reported by Fish and Rosenblum, see Figure 18 [53]. A solution was made by dissolving 40 g mercury(II) acetate in 400 mL methanol. This solution was added slowly under agitation using a magnetic stir bar to a second solution made up of 46.5 g ferrocene in 250 mL benzene. All flasks were flushed with argon and a Schlenk line was used to maintain a argon atmosphere for 10 hours. Then, 11.0 g lithium chloride in 100 mL of a 1:1 ethanol-water solution was added slowly. The resulting orange suspension was stirred at room
temperature for 2 hours and then extracted with methylene chloride. The solvent was evaporated, and the extract was washed thoroughly with water and dried over magnesium sulfate. The solid residue was then sublimed to remove unreacted ferrocene. The unsublimed portion resulted in 11.1 g of chloromercuriferrocene. The melting point was measured and found to be 195-197 °C (literature value = 194-198 °C) [53].

In the second stage of the synthesis, a solution of 2.4 g N-bromosuccinimide in 200 mL dimethylformamide was added to a solution of 4.2 g chloromercuriferrocene in 100 mL dimethylformamide. After mixing the reaction was cooled to 0°C and continued under argon for 3 hours, as previously described. Then 200 mL of a 10% aqueous sodium thiosulfate solution was added. The resulting dark solution was poured into 1 L water and the product was extracted 3 times using 200 mL of petroleum ether for each extraction. The solvent was evaporated, and the extract was washed with water and dried over magnesium sulfate. The melting point of the yellow bromoferrocene product was measured and found to be 31-33 °C (literature value = 31-32 °C) [53]. The product was also confirmed by 1H NMR. In the NMR spectrum, seen in Figure 1, the 3 expected peaks for bromoferrocene are present. These peaks have an intensity ratio of 2:2:5, corresponding to the 5 equivalent protons on one ferrocene ring, and 2 sets of 2 equivalent protons on the substituted ring. The synthesis yielded 2.0 g of product (75% based on FcHgCl).
Figure 18. Reaction Scheme for bromoferrocene synthesis.

Figure 19. $^1$H NMR spectrum of bromoferrocene product in CDCl$_3$. 
3.4 Stabilization of Ferrocenium

It was observed during earlier experimental work [54] that the cations of ferrocene, acetylferrocene, and bromoferrocene can degrade during the timescale of kinetic measurements resulting in a reduction in signal intensity in the region of ferocenium absorbance. One possible degradation pathway was proposed by Zotti et al. which involves molecular oxygen, as seen in Figure 20 [55].

![Possible degradation pathway of ferrocene cations.](image)

To counter this, reactions were first attempted under an inert atmosphere of argon and also sparged with argon, but this did not prevent cation degradation. It may be possible that a small amount of oxygen was still present in the solvents, or perhaps a different degradation pathway is taking place entirely. It was found that the addition of 1 μL trifluoromethanesulfonic acid effectively stabilized the ferrocenium ions long enough to complete the reactions under standard atmosphere [54]. Trifluoromethanesulfonic acid was chosen because it does not undergo redox processes itself yet is still a strong acid in organic solvents. Decamethylferrocene and ethylferrocene cations did not degrade as quickly, therefore trifluoromethanesulfonic acid was not used for those reactions.
3.5 Kinetics Procedure

Stock solutions of ferrocene, decamethylferrocene, ethylferrocene (Aldrich), acetylferrocene (see section 3.2), and bromoferrocene (see section 3.3), were made by dissolving the requisite compounds in reagent grade acetonitrile (Aldrich). The oxidants, dibenzoyl peroxide (Aldrich), and meta-chloroperoxybenzoic acid, 77% (Aldrich), were also dissolved in acetonitrile. At the time of reaction, the aforementioned reactants were was stabilized with trifluoromethanesulfonic acid (Aldrich).

The following reactions were performed with peroxide as the limiting reagent. 3 mL of ferrocene solution ranging in concentration from 10-18 mM + 1 µL trifluoromethanesulfonic acid + 6 µL of 60 mM dibenzoylperoxide solution, 3 mL decamethylferrocene solution ranging in concentration from 0.04-0.12 mM + 2 µL of 6 mM dibenzoylperoxide solution, 3 mL ethylferrocene solution ranging in concentration from 2-6 mM + 10 µL of 60 mM dibenzoylperoxide solution, 3 mL bromoferrocene solution ranging in concentration from 30-50 mM + 1 µL trifluoromethanesulfonic acid + 30 µL of 60 mM dibenzoylperoxide solution. Each reaction was conducted in a quartz cuvette, and monitored using a temperature-controlled Hewlett-Packard HP8452 spectrophotometer. The reactions were monitored at a frequency of 1 spectrum per second, at five selected wavelengths, between 300-800 nm in which ferrocenium cations exhibit an absorption. Reactions were tracked by a change in absorbance attributed to formation of RFc⁺. For ferrocene and bromoferrocene, the wavelength of 617 nm was chosen, but for decamethylferrocene
and ethylferrocene 350 nm was used because absorbance values were higher and more repeatable. One explanation for this is that the DMFc⁺ peak is reported to be red-shifted compared to Fc⁺ due to the increased electron density donated from the methyl groups [56]. Furthermore, the rate at which decamethylferrocene and ethylferrocene reacted required the use of low reactant concentrations, and thus an area of greater absorbance change with a higher ε value was chosen. In all cases, measured reaction rates did not depend on the wavelength used. All reactions proceeded until an asymptote was reached at maximum absorbance. Each reaction was repeated at an array of temperatures ranging from 0 °C to 50 °C.

3.6 Product Isolation:

The expected products for the oxidation of ferrocene derivatives with dibenzoyl peroxide are ferrocenium cations are benzoic acid. This may occur according to Equation 11.

\[
\text{Fc} + (\text{C}_6\text{H}_5\text{COO})_2 \rightarrow \text{Fc}^+ + \text{C}_6\text{H}_5\text{COO}^- + \text{C}_6\text{H}_5\text{COO}^- \quad \text{Slow}
\]

\[
\text{Fc} + \text{C}_6\text{H}_5\text{COO}^- \rightarrow \text{Fc}^+ + \text{C}_6\text{H}_5\text{COO}^- \quad \text{Fast}
\]

\[
2\text{Fc} + (\text{C}_6\text{H}_5\text{COO})_2 + 2\text{H}^+ \rightarrow 2\text{Fc}^+ + 2\text{C}_6\text{H}_5\text{COOH} \quad (11)
\]

To confirm that these products were forming, a series of product studies was conducted. Each ferrocene derivative was treated with dibenzoyl peroxide at a 2:1
molar ratio for 30 minutes at the following concentrations using acetonitrile as the solvent: 300 mL of 1.02 mM decamethylferrocene solution + 10 mL of 16.5 mM dibenzoyl peroxide solution. 300 mL of 1.8 mM ferrocene solution + 10 mL of 26.8 mM dibenzoyl peroxide solution + 1 μL trifluoromethanesulfonic acid. 300 mL of 1.5 mM ethylferrocene solution + 10 mL of 23.1 mM dibenzoyl peroxide solution. 300 mL of 1.26 mM bromoferrocene solution + 10 mL of 18.9 mM dibenzoyl peroxide solution + 1 μL trifluoromethanesulfonic acid. The benzoic acid was extracted by precipitation with aqueous sodium bicarbonate, filtered, and the acid was reformed by treatment with HCl. The product was then extracted and dried. The percent yield was determined gravimetrically.

3.7 Solvent Effects

The effect of solvent on the kinetics of the DMFc + DBP reaction were examined at 20.0ºC using 3 mL of a 0.12 mM DMFc solution created using ethanol, acetone, acetonitrile, or 1,2-dichloroethane combined with 3 μL of 6 mM DBP solution. The reaction of DMFc + mCPBA was also investigated at the same temperature with the same solvents, using 3 mL of a 0.24 mM DMFc solution combined with 6 μL of 6 mM mCPBA also in the corresponding solvent. The reactions were then monitored spectrophotometrically as described in section 3.5.
CHAPTER IV

COMPUTATIONAL

4.1 Methodology

Reorganization energy is a key concept in Marcus theory that can be used to describe outer sphere electron transfer mechanisms but is it difficult to determine experimentally. Therefore, to better understand the energetics of ferrocene oxidation, a series of computational studies was conducted. The reorganization energy can be calculated using the experimentally determined $\Delta G^\ddagger$ and computationally determined $\Delta G^\circ$. The analysis of $\lambda$, along with the other activation parameters, can provide insight into what type of reaction mechanism is taking place. Furthermore, it is possible to model $\lambda_i$ computationally, which, when combined with $\lambda$, allows for $\lambda_o$ to be deduced. The ratio of $\lambda_i$ to $\lambda_o$ is a strong indication of the role the solvent environment plays in an electron transfer process and will help complete the picture for the reactions of ferrocene.

A second study to investigate steric hindrance and model a possible inner sphere mechanism for the oxidation of ferrocenes was also conducted. Again, a computational approach is especially well suited for this because it would be difficult to determine experimentally. An examination of steric interactions for ferrocene molecules with bulky substituents coordinating with peroxides can provide additional evidence of whether this reaction pathway is plausible.
An unrestricted B3LYP method was chosen because, when dealing with radicals, it avoids the excessive spin contamination encountered when using a UMP2 method. B3LYP is a hybrid functional developed in the late 1980s that combines elements from density functional theory and Hartree-Fock theory [57]. The goal of both density functional theory and Hartree-Fock theory is to recover electron correlation, but these methods have different shortcomings. Hartree-Fock methods exactly treat the exchange correlation, but have difficulties recovering dynamic electron correlation. In contrast, density functional theory can exactly determine dynamic electron correlation but must approximate exchange correlation. B3LYP combines the strengths of both methods. It is generally regarded as faster than most Hartree-Fock techniques, and more accurate than density functional theory alone.

4.3 Calculation of $\Delta G^\circ$

Computations to determine $\Delta G^\circ$ were performed under solvated conditions in acetonitrile using all-electron (U)B3LYP/6-31+G* with Gaussian 09 [58] through the Ohio Supercomputer Center [59]. Vibrational analysis also verified that optimized geometries were located at true potential-energy minima and provided the free energies. The solvent was treated using the polarizable continuum model, solute in a cavity in the solvent reaction field (Gaussian keyword SCRF) [60]. First initial-guess conformations and geometries were obtained by molecular-mechanics energy minimization.
4.4 Calculation of $\lambda_i$

Using a method developed by Klimkans and Larsson [61], it is possible to calculate the nuclear reorganization energy based on computational geometry optimizations and energy calculations. Using this method, the nuclear reorganization energy $\lambda_i$, is defined as the sum of the energy differences for the reactants before and after an electron transfer, in their initial and final molecular geometries. This idea is further illustrated in Figure 21. Here the energy difference between the two points in the oxidized state and the reduced state is due to molecular geometry. The energy required for this geometry change to take place is defined as $\lambda_i$. The overall relationship between $\lambda$ and $\Delta G$ is shown in Figure 22.
Figure 21. Schematic potential-energy diagram used to determine $\lambda_i$ ($E_1$ and $E_3$ correspond to optimized minima, and $E_2$ and $E_4$ to vertical electron transfer at the geometries $E_1$ and $E_3$ respectively)

![Schematic energy diagram for endergonic electron transfer between donor D and acceptor A.](image)

Figure 22. Schematic energy diagram for endergonic electron transfer between donor D and acceptor A.

Using the same computational parameters as described in section 4.3, the self-consistent field energy was calculated for each reactant and recorded. A single electron was then removed from each ferrocene species to simulate the oxidized state and added to each oxidant to simulate the reduced state. The molecular geometry was then optimized again. A second round of energy calculations was conducted, and the
results recorded. The sum of the energy difference for the different geometries before and after electron transfer was then calculated. This yielded the nuclear reorganization energy for each ferrocene derivative.

4.2 Molecular Geometry

A general geometry optimization was conducted using ORCA 4.0 [62], with an unrestricted hybrid functional B3LYP, and an all-electron 6-31G* basis set. First molecules were drawn in standard molecular modeling software, such as ChemDraw. The atomic coordinates were then imported into Orca, and the initial geometry optimization conducted.

An oxidative addition mechanism was also investigated to evaluate possible steric effects. To accomplish this, both the iron center of ferrocene and decamethylferrocene were coordinated with a benzoyl radical. A geometry optimization was then performed to determine if a stable configuration was possible.
CHAPTER V

RESULTS

5.1 Kinetic Results

The UV-visible spectra of each ferrocene species before and after oxidation can be seen in Figures 23-27. It was observed that the reactions of acetylferrocene were too slow to obtain accurate rate measurements even at the limit of solubility and were therefore disregarded.

![UV-visible spectrum of ferrocene before and after reaction with dibenzoyl peroxide.](image)

Figure 23. UV-visible spectrum of ferrocene before and after reaction with dibenzoyl peroxide (black = before reaction, red = after reaction).
Figure 24. UV-visible spectrum of decamethylferrocene before and after reaction with dibenzoyl peroxide (black = before reaction, red = after reaction).

Figure 25. UV-visible spectrum of ethylferrocene before and after reaction with dibenzoyl peroxide (black = before reaction, red = after reaction).
Figure 26. UV-visible spectrum of bromoferrocene before and after reaction with dibenzoyl peroxide (black = before reaction, red = after reaction).

Figure 27. UV-visible spectrum of decamethylferrocene before and after reaction with mCPBA (black = before reaction, red = after reaction).
An example kinetics trace for each ferrocene reaction can be seen in Figure 28-32.

Figure 28. Absorbance trace showing the growth of ferrocenium peak with dibenzoyl peroxide at 617 nm.
Figure 29. Absorbance trace showing signal growth at 350 nm of decamethylferrocene with dibenzoyl peroxide.

Figure 30. Absorbance trace showing peak growth at 350 nm of ethylferrocene with dibenzoyl peroxide.
Figure 31. Absorbance trace showing peak growth at 617 nm of bromoferrocene with dibenzoyl peroxide.
Figure 32. Absorbance trace showing growth at 350 nm of decamethylferrocene with mCPBA.

The graphs of $k'$ vs. ferrocene concentration which were used to determine $k$ can be seen in Figures 33-37.

Figure 33. Pseudo-first-order rate constants for ferrocene with dibenzoyl peroxide.
Figure 34. Pseudo-first-order rate constants for decamethylferrocene with dibenzoyl peroxide.
Figure 35. Pseudo-first-order rate constants for ethylferrocene with dibenzoyl peroxide.

Figure 36. Pseudo-first-order rate constants for BFc + DBP reactions.
Figure 37. Pseudo-first-order rate constants for decamethylferrocene with mCPBA.

Eyring plots were then creating by graphing $1/T$ vs. $\ln(k/T)$, see Figure 38.

![Eyring graphs for each ferrocene species.](image)

The standard deviation for the rate constants, for each set of reactions, at each temperature, is given in Table I through Table V.

Table I. Average standard deviations of rate constants for reactions of decamethylferrocene and dibenzoyl peroxide.

<table>
<thead>
<tr>
<th>Temperature, °C</th>
<th>Standard Deviation %</th>
</tr>
</thead>
<tbody>
<tr>
<td>8.0</td>
<td>0.40</td>
</tr>
<tr>
<td>15.5</td>
<td>0.76</td>
</tr>
</tbody>
</table>
Table II. Standard deviations of rate constants for reactions of ferrocene and dibenzoyl peroxide.

<table>
<thead>
<tr>
<th>Temperature, °C</th>
<th>Standard Deviation %</th>
</tr>
</thead>
<tbody>
<tr>
<td>14.8</td>
<td>0.05</td>
</tr>
<tr>
<td>22.0</td>
<td>0.26</td>
</tr>
<tr>
<td>33.0</td>
<td>0.15</td>
</tr>
<tr>
<td>39.5</td>
<td>0.42</td>
</tr>
<tr>
<td>44.5</td>
<td>0.51</td>
</tr>
</tbody>
</table>

Table III. Standard deviations of rate constants for reactions of ethylferrocene and dibenzoyl peroxide.

<table>
<thead>
<tr>
<th>Temperature, °C</th>
<th>Standard Deviation %</th>
</tr>
</thead>
<tbody>
<tr>
<td>16.0</td>
<td>0.72</td>
</tr>
<tr>
<td>22.0</td>
<td>0.65</td>
</tr>
<tr>
<td>28.5</td>
<td>0.47</td>
</tr>
<tr>
<td>36.0</td>
<td>0.23</td>
</tr>
<tr>
<td>45.0</td>
<td>0.51</td>
</tr>
</tbody>
</table>

Table IV. Standard deviations of rate constants for reactions of bromoferrocene and dibenzoyl peroxide.

<table>
<thead>
<tr>
<th>Temperature, °C</th>
<th>Standard Deviation %</th>
</tr>
</thead>
<tbody>
<tr>
<td>21.0</td>
<td>0.88</td>
</tr>
<tr>
<td>25.0</td>
<td>0.71</td>
</tr>
<tr>
<td>34.0</td>
<td>0.76</td>
</tr>
<tr>
<td>40.0</td>
<td>0.92</td>
</tr>
</tbody>
</table>
Table V. Standard deviations of rate constants for reactions of decamethylferrocene and mCPBA.

<table>
<thead>
<tr>
<th>Temperature, °C</th>
<th>Standard Deviation %</th>
</tr>
</thead>
<tbody>
<tr>
<td>22.5</td>
<td>1.11</td>
</tr>
<tr>
<td>27.0</td>
<td>0.32</td>
</tr>
<tr>
<td>30.5</td>
<td>1.26</td>
</tr>
<tr>
<td>35.0</td>
<td>1.31</td>
</tr>
</tbody>
</table>

5.2 Activation Parameters

The activation parameters for the oxidation of each ferrocene derivative was calculated (see Table XI) using the slope and intercept from the Eyring graphs by applying Equation 12.

\[
\ln \left( \frac{k_{obs}}{T} \right) = \frac{-\Delta H^\ddagger}{RT} + \ln \left( \frac{k_B}{h} \right) + \frac{\Delta S^\ddagger}{R}
\]  

(12)

This is the linear form of the Eyring equation where \( k_{obs} = 2k \) to account for the stoichiometry of the reaction. The transmission coefficient \( \kappa \) is assumed to be 1, and therefore only the forward reactions are considered. The expressions for the calculation of each activation parameter are seen in Equations 13-15.

\[
\Delta H^\ddagger = -\text{slope} \times R
\]

(13)

\[
\Delta S^\ddagger = R \times \left[ \text{intercept} - \ln(k_B/h) \right]
\]

(14)

\[
\Delta G^\ddagger = \Delta H^\ddagger - T \Delta S^\ddagger
\]

(15)
Table VI. Calculated activation and reorganization energies

<table>
<thead>
<tr>
<th>Reaction</th>
<th>$k$, M$^{-1}$·s$^{-1}$</th>
<th>$\Delta H^\ddagger$, kJ/mol</th>
<th>$\Delta S^\ddagger$, J/(mol·K)</th>
<th>$\Delta G^\ddagger$, kJ/mol$^a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fc + DBP</td>
<td>0.162</td>
<td>59 ± 1</td>
<td>-60 ± 7</td>
<td>76.7</td>
</tr>
<tr>
<td>EtFc + DPB</td>
<td>0.345</td>
<td>49.9 ± 0.5</td>
<td>-85 ± 4</td>
<td>75.0</td>
</tr>
<tr>
<td>DMFc + DBP</td>
<td>269</td>
<td>24.5 ± 0.2</td>
<td>-115 ± 1</td>
<td>58.4</td>
</tr>
<tr>
<td>DMFc + mCPBA</td>
<td>4.85</td>
<td>61.5 ± 0.5</td>
<td>-24 ± 3</td>
<td>68.0</td>
</tr>
<tr>
<td>BrFc + DBP</td>
<td>0.0100</td>
<td>54.6 ± 0.2</td>
<td>-98 ± 1</td>
<td>83.8</td>
</tr>
</tbody>
</table>

$^a22.0^\circ$C.

5.3 Product isolation

The yields of the benzoic acid product, along with recovered unreacted ferrocene, are shown in Table VII. The melting point of the benzoic acid product was measured to be 121-123°C (literature value 122-124°C) [63]. Benzoic acid was also confirmed by GC-MS and NMR. The GC column retention was quite high for an acidic material, and resulted in a trailing peak, seen in Figure 39, but the mass spectrum, seen in Figure 40, does confirm benzoic acid and matches the literature spectrum seen in Figure 41 [64]. The $^1$H NMR spectrum, seen in Figure 42, also matches the literature spectrum [65].
Table VII. Gravimetrically determined % yield of benzoic acid from reactions of RFc with DBP after 30 min reaction time.

<table>
<thead>
<tr>
<th>Ferrocene</th>
<th>Benzoic acid formed, %</th>
<th>Unreacted ferrocene, %</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fc</td>
<td>72.5</td>
<td>25</td>
</tr>
<tr>
<td>EtFc</td>
<td>81.7</td>
<td>17</td>
</tr>
<tr>
<td>DMFc</td>
<td>92.2</td>
<td>5</td>
</tr>
<tr>
<td>BrFc</td>
<td>69.4</td>
<td>29</td>
</tr>
</tbody>
</table>

Figure 39. Gas chromatograph of benzoic acid product from decamethylferrocene and dibenzoyl peroxide reaction.
Figure 40. Mass spectrum of benzoic acid product from decamethylferrocene and dibenzoyl peroxide reaction.

Figure 41. Benzoic acid reference mass spectrum.
Figure 42. $^1$H NMR spectrum of benzoic acid product from decamethylferrocene and dibenzoyl peroxide reaction (left), and library spectrum (right).

Ferrocenium was also extracted qualitatively through aqueous solvent extraction and filtered, yielding a yellow solution of unreacted ferrocene in toluene and a blue-green solution of aqueous ferrocenium ions, see Figures 43 and 44.
Figure 43. Extraction of ferrocenium from reaction mixture (top = toluene layer, bottom = aqueous layer).

Figure 44. Extracted and filtered ferrocenium (left) and unreacted ferrocene (right).
5.4 Solvent Results

The results of the solvent study are seen in Table VIII. For the case of decamethylferrocene + mCPBA in acetone the reaction was found to be too slow to accurately measure the rate constant and was therefore disregarded.

Table VIII. Effect of solvent on $k$ for decamethylferrocene at 20ºC.

<table>
<thead>
<tr>
<th>Solvent</th>
<th>DMFc + DBP, $k$ (M$^{-1}$s$^{-1}$)</th>
<th>DMFc + mCPBA, $k$ (M$^{-1}$s$^{-1}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetone</td>
<td>114</td>
<td>--</td>
</tr>
<tr>
<td>1,2-Dichloroethane</td>
<td>167</td>
<td>10</td>
</tr>
<tr>
<td>Acetonitrile</td>
<td>208</td>
<td>3.3</td>
</tr>
<tr>
<td>Ethanol</td>
<td>306</td>
<td>2.0</td>
</tr>
</tbody>
</table>

5.5 Computational Results

$\Delta G^\circ$ was determined by subtracting the free energy for reduction for the peroxide species from the free energy for oxidation from the ferrocene species, see Table IX. The reorganization energy was calculated using the experimentally derived $\Delta G^\dagger$ and $\Delta G^\circ$ via equation 16, see Table X. $\lambda_i$ was determined by taking the sum of the energy differences for reactants before and after an electron transfer, in initial and final molecular geometries, see Table XI. $\lambda_o$ was calculated by subtracting $\lambda_i$ from $\lambda$. It should be noted that the $\lambda_i$ results are for ferrocene species only, and do not include
the reorganization energy of the peroxide. Due to the dissociative behavior of DBP, \( \lambda \) cannot be accurately modeled using this method.

\[
\Delta G^\ddagger = \frac{(\Delta G^\circ + \lambda)^2}{4\lambda}
\]  

(16)

Table IX. Determination of \( \Delta G^\circ \)

<table>
<thead>
<tr>
<th>Species</th>
<th>( \text{ROOR, au} )</th>
<th>( \text{ROOR}^-, \text{au} )</th>
<th>( \Delta G / \text{au} )</th>
<th>( \Delta G / \text{eV} )</th>
<th>( \Delta G / \text{kJ/mol} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>DBP</td>
<td>-840.223668</td>
<td>-840.419525</td>
<td>-0.19586</td>
<td>-5.33</td>
<td>-514.2</td>
</tr>
<tr>
<td>mCPBA</td>
<td>-955.497222</td>
<td>-955.599445</td>
<td>-0.10222</td>
<td>-2.78</td>
<td>-268.4</td>
</tr>
</tbody>
</table>

\( \text{Free Energies for reduction}^a \)

\( \text{Free Energies for oxidation}^a \)

<table>
<thead>
<tr>
<th>Species</th>
<th>( \text{XFc, au} )</th>
<th>( \text{XFc}^+, \text{au} )</th>
<th>( \Delta G / \text{au} )</th>
<th>( \Delta G / \text{eV} )</th>
<th>( \Delta G / \text{kJ/mol} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fc</td>
<td>-1650.600000</td>
<td>-1650.40719</td>
<td>0.192811</td>
<td>5.25</td>
<td>506.2</td>
</tr>
<tr>
<td>EtFc</td>
<td>-1729.178560</td>
<td>-1728.988554</td>
<td>0.190006</td>
<td>5.17</td>
<td>498.9</td>
</tr>
<tr>
<td>DMFc</td>
<td>-2043.516170</td>
<td>-2043.344326</td>
<td>0.171844</td>
<td>4.68</td>
<td>451.2</td>
</tr>
<tr>
<td>Bfc</td>
<td>-4221.741767</td>
<td>-4221.539720</td>
<td>0.202047</td>
<td>5.50</td>
<td>530.5</td>
</tr>
</tbody>
</table>

\( a (U)B3LYP/6-31+G^* \) in acetonitrile solvent

Table X. Computational results for \( \Delta G^\circ \) and \( \lambda \).

<table>
<thead>
<tr>
<th>Reaction</th>
<th>( \Delta G^\circ ), kJ/mol(^a )</th>
<th>( \lambda ), kJ/mol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fc + DBP</td>
<td>-8.0</td>
<td>323</td>
</tr>
<tr>
<td>EtFc + DPB</td>
<td>-15.4</td>
<td>330</td>
</tr>
<tr>
<td>DMFc + DBP</td>
<td>-63.0</td>
<td>348</td>
</tr>
<tr>
<td>DMFc + mCPBA</td>
<td>183.0</td>
<td>--</td>
</tr>
<tr>
<td>BrFc + DBP</td>
<td>16.3</td>
<td>302</td>
</tr>
</tbody>
</table>

\( a (U)B3LYP/6-31+G^* \) at 25.0°C.
Table XI. Determination of $\lambda_i$ and $\lambda_o$ for ferrocenes solvated in acetonitrile.

<table>
<thead>
<tr>
<th></th>
<th>$\lambda_1$, $E_h$</th>
<th>$\lambda_2$, $E_h$</th>
<th>$\lambda_i$, $E_h^a$</th>
<th>$\lambda_i$, kJ/mol</th>
<th>$\lambda$, kJ/mol</th>
<th>$\lambda_o$, kJ/mol$^b$</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>$E_2-E_3$</td>
<td>$E_4-E_1$</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Fc</td>
<td>0.00040</td>
<td>0.00034</td>
<td>0.00075</td>
<td>2.0</td>
<td>323</td>
<td>321</td>
</tr>
<tr>
<td>EtFc</td>
<td>0.00059</td>
<td>0.00063</td>
<td>0.00122</td>
<td>3.2</td>
<td>330</td>
<td>327</td>
</tr>
<tr>
<td>DMFc</td>
<td>0.00148</td>
<td>0.00128</td>
<td>0.00275</td>
<td>7.2</td>
<td>348</td>
<td>341</td>
</tr>
<tr>
<td>BrFc</td>
<td>0.00099</td>
<td>0.00094</td>
<td>0.00193</td>
<td>5.1</td>
<td>302</td>
<td>297</td>
</tr>
</tbody>
</table>

$^a \lambda_i = \lambda_1 + \lambda_2$, $^b \lambda_o = \lambda - \lambda_i$

5.6 Geometry Results

The optimized structures of the reactants used in this study can be seen in Figures 45-52.
Figure 45. Optimized geometry of a ferrocene molecule.

Figure 46. Optimized geometry of a bromoferrocene molecule.
Figure 47. Optimized geometry of an ethylferrocene molecule.

Figure 48. Optimized geometry of a decamethylferrocene molecule.
Figure 49. Optimized geometry of an acetylferrocene molecule.

Figure 50. Optimized geometry of mCPBA.
Figure 51. Optimized geometry of dibenzoyl peroxide.

Figure 52. Optimized geometry of benzoyloxy radical.
The results of a modeled theoretical coordination processes for ferrocene and decamethylferrocene are seen in Figures 53 and 54. For both cases, the wave functions failed to converge, and the energy continued to increase as the radical approached the iron center. It should then be noted that these figures do not reflect an energy minimum or maximum, just a snapshot of energy at an arbitrary point in the vicinity of coordination distances.

Figure 53. Distortion of ferrocene induced by the approach of a benzoyloxy radical as in a coordination process.
Figure 54. Distortion of decamethylferrocene induced by the approach of a benzoyloxy radical as in a coordination process.
6.1 Kinetics analysis

When considering the kinetics for the oxidation of ferrocene, two observations are especially noteworthy. First, the ten methyl groups on the pentamethylcyclopentadienyl rings of DMFc do not impede the reaction. To the contrary, DMFc reacts faster with DBP than any of the other ferrocenes. This indicates that intramolecular electron donation by the ligands promotes the reaction more than steric effects might hinder it. This represents a strong argument for an electron-transfer mechanism that can occur over a longer distance, eliminating the need for the oxidant to approach the iron center. Second, the reactions are first-order in ferrocene, but consume two molecules of ferrocene per one mole of peroxide. This requires a second ferrocene molecule to participate after the rate-limiting step. In order to further explore the nature of the process, activation parameters derived using the Eyring are considered. The activation entropies $\Delta S^\ddagger$ are consistent with ionization occurring during (or before) the rate-limiting step. Large negative activation entropies are common for ionizations and can be attributed to solvent effects [66]. The activation entropy also shows a general trend in which ferrocene derivatives with larger substituents have a more negative $\Delta S^\ddagger$. This is expected because increasing the molecular radius should result in a larger, more ordered solvent shell [67]. The radius of DMFc (6.00 Å [68]) is the largest, and its entropy of activation is most negative
compared to Fc (4.07 Å [68]) with the smallest radius and the least negative entropy of activation. Similarly, mCPBA is smaller than DBP, and therefore the activation entropy for reaction of DMFc with mCPBA is more favorable. Such a trend supports an outer-sphere electron transfer mechanism for which solvation is a key step. The activation enthalpies generally trend in the same direction as entropy, with the lowest (most favorable) $\Delta H^\ddagger$ found for the reaction of DMFc with DBP, which has the lowest (least favorable) entropy. The fact that DMFc requires the least energy is indicative of electron donation from the ten methyl groups, resulting in a greater reaction rate with a low $\Delta H^\ddagger$. The $\Delta H^\ddagger$ for the reaction of EtFc with DBP is next lowest. BrFc and Fc show a similar $\Delta H^\ddagger$ with BrFc being slightly lower. The rates of the latter compounds correlate with the Hammett $\sigma_p$ constants [69] with $\rho = -3.5$ determined as the slope divided by 2.303, seen in Figure 55. Although Hammett $\sigma_p$ constants are determined for only three compounds, this behavior is consistent with electron donation facilitating the reactions.
A poor correlation was found with Hammett $\sigma_m$ parameters, which is consistent with substituent interactions due to the development of charge. It was also observed that ferrocenes with more electron-rich centers react further towards completion and faster than their less electron-rich counterparts, as evidenced by the trend in the yields of benzoic acid, DMFc > EtFc > Fc > BrFc.

The less favorable activation enthalpy for reaction of DMFc with mCPBA compared to DBP is likely due to mCPBA being a weaker oxidizing agent than DBP. The respective electron affinities for DBP and mCPBA were calculated for their optimized geometries in acetonitrile as part of the computational study, and were found to be 5.33 eV and 2.78 eV respectively. These results reveal that reduction of DBP typically leads to the cleavage of the O-O bond. In fact, the geometry
calculations revealed that the O-O bond distance increases from 1.433 Å for the neutral molecule to 2.201 Å for the radical anion DBP\(^-\) in the optimized structure. This indicates that the latter is essentially a loose complex of benzoate anion and benzoyl radical with little binding energy. This is not unexpected as dialkyl peroxides have been reported to undergo dissociative electron capture \([70]\). On the other hand, the O-O distance in mCPBA increases from 1.440 Å to just 1.451 Å in the radical anion mCPBA\(^-\). The shorter bond distance implies a stronger O-O bond in then in DBP\(^-\). Therefore, electron capture by DBP should be irreversible (or nearly so), but this may not be the case for mCPBA\(^-\). The potential for a reverse electron transfer is one possible explanation for the fact that the experimental \(\Delta G^\ddagger\) for mCPBA is much lower than the calculated \(\Delta G^o\), whereas \(\Delta G^\ddagger\) and \(\Delta G^o\) are in closer alignment for DBP. This may also suggest a different, more competitive mechanism may be taking place in the reaction of mCPBA with ferrocene.

6.2 Solvent Effects

Solvent has surprisingly little effect on the rate constant for reaction of DMFc with DBP. The rates do not correlate directly with dielectric constant of the solvent. Reaction rates rank ethanol > acetonitrile > 1,2 dichloroethane > acetone although the largest difference is less than a factor of 3. The effect of substituents is much larger than this. Correlation is better with spectroscopic estimates of solvent polarity as used for Reichardt’s \(E_T\) and Kosower’s \(Z\) values, seen in Table XII.
Table XII. Effect of solvent on rate constant for DMFc at 20ºC.

<table>
<thead>
<tr>
<th>Solvent</th>
<th>DMFc + DBP, $k$ (M$^{-1}$·s$^{-1}$)</th>
<th>DMFc + mCPBA, $k$ (M$^{-1}$·s$^{-1}$)</th>
<th>Dielectric Constant$^a$</th>
<th>Relative Polarity$^b$</th>
<th>$E_T(30)^{b,c}$</th>
<th>Z value$^d$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetone</td>
<td>114</td>
<td>--</td>
<td>20.7</td>
<td>0.355</td>
<td>42.2</td>
<td>65.7</td>
</tr>
<tr>
<td>1,2-Dichloroethane</td>
<td>167</td>
<td>10</td>
<td>16.7</td>
<td>0.327</td>
<td>41.9</td>
<td>63.8$^e$</td>
</tr>
<tr>
<td>Acetonitrile</td>
<td>208</td>
<td>3.3</td>
<td>37.5</td>
<td>0.46</td>
<td>46</td>
<td>71.3</td>
</tr>
<tr>
<td>Ethanol</td>
<td>306</td>
<td>2.0</td>
<td>24.3</td>
<td>0.654</td>
<td>51.9</td>
<td>79.6</td>
</tr>
</tbody>
</table>

$^a$[71]  $^b$[72]  $^c$[73]  $^d$[74]  $^e$[75]

These scales were developed using the charge-transfer absorption energies for organic salts, which would include specific solvation effects and not just bulk solvent properties. Here the absorption of a photon promotes a vertical transfer of an electron, which serves as a reasonable model for a chemical electron transfer. Also, much like an outer-sphere process, the transfer is non-adiabatic. However, there are some important differences as well, one being that the electron is transferred from polar, uncharged starting materials to form ions, rather than the reverse. It is therefore possible that the role of solvent may differ between chemical and photochemical processes, and the correlation should not be expected to be exact. Furthermore, photoelectron spectroscopy reveals that the highest three occupied molecular orbitals of Fc are principally metal 3$d$ in nature with little ligand character [76]. This means that the charge in ferrocenium should be localized mostly on iron. Since the metal atom is isolated from strong interactions with the solvent by the Cp rings and their substituents, solvation should be relatively ineffective in stabilizing the charge,
especially in DMFc\(^+\). Hence the DMFc/DMFc\(^+\) redox couple is well-defined and only mildly dependent on solvent [77], consistent with this analysis. Ferrocenes are also unusual in that their carbon atoms bear partial negative charge and the central iron some positive charge, even in the neutral compounds [78]. Therefore, ionization, depicted in Figure 56, should decrease the outermost (negative) charge interacting with solvent. In this case, solvent effects may differ from reactions which produce more exposed positive charges, such as the formation of carbocations.

![Figure 56. Diagram depicting the ionization of ferrocene.](image)

Consequently, solvent effects may have more to do with solvation of the anions than of the ferrocenium cations. The negative charge in DBP\(^-\) resides primarily on oxygen, especially the keto O atoms, whereas in the products, the negative charge is more delocalized over the carboxylate anion fragment. Protic solvents like ethanol strongly solvate anions by hydrogen bonding, whereas aprotic solvents like acetone, acetonitrile and ethylene chloride with large dipole moments solvate cations better [68]. Therefore, solvent polarity may be less important for ferrocenium cations but more important for anions that would be stabilized by hydrogen bonding. This would result in a larger rate in ethanol than in the other solvents, as observed.
When compared to the DBP reactions the reaction rates of \( m \)CPBA vary more strongly and with different solvent ordering, with ethanol giving the slowest rate measured. Since \( m \)CPBA has a hydroxyl group but \( m \)-ClC\(_6\)H\(_4\)CO\(_2\)\(^-\) does not, hydrogen bonding with a protic solvent such as ethanol would stabilize the reactant more than the product anion, slowing the overall rate of reaction. Also, electron-donor-acceptor complexes have long been known to form between good metal atoms and ligands. [79-81]. Some of these have been shown to be kinetically competent intermediates in several bimolecular reactions, such as the Diels-Alder reaction [82]. The formation of such complexes between the ferrocenes and peroxides seems likely, considering the polar nature and accepting ability of the ferrocenes. Clearly, the effect of solvent in these processes is complex and warrants further study.

6.3 Reorganization energy

The \( \lambda \) values for the reaction of ferrocenes with peroxide are quite large, in the range of 300-400 kJ/mol. High reorganization energies in this range should be expected for outer-sphere electron transfers [83]. Conversely, reported \( \lambda \) values for the inner-sphere electron transfer of hydroquinones with DBP are significantly lower, in the range of 60-80 kJ/mol [84]. The energy difference between these pathways is mostly attributed to the thermal energy required to rearrange the solvent environment into the correct polarization state to facilitate an electron transfer. In an inner sphere process, the solvent plays a much smaller role. Furthermore, \( \lambda_{ii} \), which is considered to be the energy required to modify bond angles and bond distances from reactant to product, was found to be only a minor contributor to the overall reorganization energy.
This is consistent with the small geometry changes ferrocenes undergo upon electron loss. The $\lambda_o$ component was much more significant, as should be the case for an outer-sphere mechanism where significant solvent ordering must occur. The $\lambda_o$ value is also found to be slightly larger in ferrocene derivatives with electron donating substituents, and smaller in the case of electron withdrawing ones. One explanation for this could be that, when electron withdrawing substituents are present on ferrocene, the Fe atom possesses more partial positive charge and thus requires slightly less polarization of the solvent environment to transition to the product state. Overall, the differences in $\lambda_o$ were small and do not appear to significantly affect reaction rates.

### 6.4 Coordination Model

The coordination model shows that both ferrocene and decamethylferrocene must undergo significant distortion to allow the oxidant to approach closely enough to facilitate a $\sigma$ bond. The coordinated conformation of ferrocene is $1.98\ E_h$ higher in energy then the non-interacting molecules alone, and for decamethylferrocene the coordinated state is $5.19\ E_h$ higher in energy. With energy requirements this considerable, the ten methyl substitutes should be expected to slow the rate of reaction significantly if an oxidative addition process were taking place. However, it should be noted that this was a primarily a qualitative evaluation in an unsolvated system, so the accuracy was not high. Furthermore, the wave functions failed to converge so these models are likely best used only as a means to visualize what types of geometry changes might be required in a coordination mechanism.
6.5 Literature comparison

It is possible to correlate the solvated molecular radius of the reactants with the experimentally determined $\Delta S^{\ddagger}$ values. This is done by considering the theoretical hydrodynamic radius of ferrocene derivatives using the Stokes–Einstein equation, seen in Equation 17. This method is helpful because it uses an experimentally measured quantity, such as a diffusion coefficient in solution, to account for solvated molecular size that would be otherwise difficult to measure directly. When the hydrodynamic radius is compared with the experimentally determined $\Delta S^{\ddagger}$ values, Table XIII, we see that the molecules with the larger radius have more negative entropies of activation. This should be expected, as a larger solvation shell would induce a more ordered environment, thus more negative entropy. In fact, the $\Delta S^{\ddagger}$ value for decamethylferrocene is almost twice that of ferrocene.

\[
D = \frac{k_B T}{6 \pi \eta r}
\]  

(17)

\[ \eta = \text{dynamic viscosity}, \quad r = \text{radius of the spherical particle}, \quad D = \text{diffusion constant}, \]

\[ k_B = \text{Boltzmann's constant}, \quad T = \text{absolute temperature} \]

Table XIII. Theoretical hydrodynamic radius, and $\Delta S^{\ddagger}$ of ferrocene derivatives.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Radius, Å</th>
<th>Compound</th>
<th>$\Delta S^{\ddagger}$, kJ/mol*K</th>
</tr>
</thead>
<tbody>
<tr>
<td>Decamethylferrocene</td>
<td>6.00</td>
<td>Decamethylferrocene</td>
<td>-115</td>
</tr>
<tr>
<td>Ferrocene</td>
<td>4.07</td>
<td>Ferrocene</td>
<td>-60</td>
</tr>
</tbody>
</table>

79
Experimental results can also be correlated with electrochemical studies by comparing the $\Delta G^{\ddagger}$ values with electrochemical ionization and oxidation potentials found in the literature [85]. As shown in Table XIV, ferrocene derivatives with electron donating groups are more easily oxidized and ionized electrochemically when compared to derivatives with electron withdrawing groups. This comparison implies that the mechanism for the chemical oxidation of ferrocenes may be similar to that of electrochemical oxidations.

Table XIV. Electrochemical oxidation and ionization potentials

<table>
<thead>
<tr>
<th>Compound</th>
<th>Oxidation Potential, V, (acetonitrile)</th>
<th>ionization potential, eV</th>
</tr>
</thead>
<tbody>
<tr>
<td>decamethylferrocene</td>
<td>-0.125</td>
<td></td>
</tr>
<tr>
<td>t-butylferrocene</td>
<td>0.305</td>
<td>6.36</td>
</tr>
<tr>
<td>n-amylferrocene</td>
<td>0.31</td>
<td>6.38</td>
</tr>
<tr>
<td>n-butylferrocene</td>
<td>0.31</td>
<td>6.38</td>
</tr>
<tr>
<td>1,1-dimethylferrocene</td>
<td>0.288</td>
<td>6.35</td>
</tr>
<tr>
<td>N,N-dimethylaminomethylferrocene</td>
<td>0.378</td>
<td>6.5</td>
</tr>
<tr>
<td>ferrocene</td>
<td>0.38</td>
<td>6.51</td>
</tr>
<tr>
<td>hydroxymethylferrocene</td>
<td>0.378</td>
<td>6.51</td>
</tr>
<tr>
<td>vinylferrocene</td>
<td>0.385</td>
<td>6.52</td>
</tr>
<tr>
<td>bromoferrocene</td>
<td>0.41</td>
<td></td>
</tr>
<tr>
<td>ferrocenecarboxylic acid</td>
<td>0.625</td>
<td>6.75</td>
</tr>
<tr>
<td>acetylferrocene</td>
<td>0.625</td>
<td>6.81</td>
</tr>
<tr>
<td>ferrocenecarbaldehyde</td>
<td>0.673</td>
<td>6.89</td>
</tr>
</tbody>
</table>

*V vs. SCE.
6.6 Proposed oxidative addition mechanism

The reaction kinetics, steric considerations, activation parameters, reorganization energy, and computational models were all used to deduce the two most plausible reaction mechanisms for the oxidation of ferrocenes with peroxides. The first possibility is a oxidative addition process. Considering all evidence, this is the less likely mechanism, but still warrants discussion, see Equation 18, and Figure 57.

\[
\begin{align*}
\text{Cp}_2\text{Fe} + (\text{C}_6\text{H}_5\text{COO})_2 & \rightarrow \text{Cp}_2\text{Fe}(\text{C}_6\text{H}_5\text{COO})_2 \quad \text{(slow)} \\
\text{Cp}_2\text{Fe}(\text{C}_6\text{H}_5\text{COO})_2 + \text{Cp}_2\text{Fe} & \rightarrow 2 \text{Cp}_2\text{Fe}(\text{C}_6\text{H}_5\text{COO}) \quad \text{(fast)} \\
\text{Cp}_2\text{Fe(OH)} + \text{H}^+ & \rightarrow \text{Cp}_2\text{Fe}^+ + \text{H}_2\text{O} \quad \text{(fast)} \\
2 \text{Fc} + \text{H}_2\text{O}_2 + 2 \text{H}^+ & \rightarrow 2 \text{Fc}^+ + 2 \text{H}_2\text{O} \quad \text{(net)} \\
\end{align*}
\]
Here, the incoming dibenzoyl peroxide molecule binds to the ferrocene as a σ complex, and the oxygen-oxygen bond breaks due to strong back donation from the metal into the σ* orbital. This results in two new iron-oxygen bonds being formed. At this point, the coordination number and oxidation number of the iron is increased by two. A second ferrocene molecule then coordinates with the disubstituted ferrocene intermediate, abstracting one benzoyloxy group. The benzoyloxy groups on the now singly substituted ferrocenes interact with a free proton and are released, yielding benzoic acid and ferrocenium cations. This process follows second order kinetics and the initial binding step is rate limiting. Furthermore, greater electron delocalization
from donating substitutes on the cyclopentadienyl rings should accelerate this mechanism. However steric hindrance should slow it considerably as well, because the oxidant must approach the ferrocene closely to form chemical bonds.

6.7 Proposed Outer Sphere Electron Transfer Mechanism

The second possible mechanism is an outer sphere electron transfer and based on the available evidence, this is the more likely case, see Equation 19, and Figure 58.

\[
\begin{align*}
\text{Fe} + (C_6H_5CO_2)_2 & \rightarrow \text{Fe}^\cdot(C_6H_5CO_2)_2 \quad \text{(fast)} \\
\text{Fe} + (C_6H_5CO_2)_2 & \rightarrow \text{Fe}^{\cdot+} + (C_6H_5CO_2)_2^- \quad \text{(slow)} \\
(C_6H_5CO_2)_2^- & \rightarrow C_6H_5CO_2^- + C_6H_5CO_2\cdot \quad \text{(fast)} \\
\text{Fe} + C_6H_5CO_2\cdot & \rightarrow \text{Fe}^{\cdot+} + C_6H_5CO_2^- \quad \text{(fast)}
\end{align*}
\]

\[
2 \text{Fe} + (C_6H_5CO_2)_2 \rightarrow 2 \text{Fe}^{\cdot+} + 2 C_6H_5CO_2^- \quad \text{(net)} \quad (19)
\]

Figure 58. Proposed outer-sphere electron transfer mechanism.
In this model, the coordination number of the iron center is unchanged, and the oxidation number is increased by one. In the first step, the peroxide approaches the ferrocene molecule forming a weakly coupled complex. The solvent shell surrounding the reactants then assumes the correct geometry and polarization to facilitate an electron jump. This is the rate limiting step. The peroxide then receives an electron from the iron center and becomes a radical anion. This radical anion then quickly disassociates into a benzoyloxy radical and a benzoyloxy anion in a fast step. The radical then interacts with a second ferrocene moiety which also transfers an electron, thus explaining the 1:2 stoichiometry, yielding benzoyloxy anions which are protonated under acidic conditions. In an outer sphere mechanism, electron donating substituents should promote the reaction, and electron withdrawing substituents should inhibit it, as in an oxidative addition, but in this case steric effects are minimal because outer sphere process can occur over much greater distances without close reactant coupling.

6.8 Conclusions

Overall, the rate of oxidation depends strongly on the modification of the ferrocene derivative, with electron donating substituents accelerating the reaction and electron withdrawing substitutes slowing it down. The Gibbs energy of activation for bromoferrocene is lower than that of unsubstituted ferrocene, ethylferrocene and decamethylferrocene. To rationalize these results, an outer-sphere electron transfer
processes was proposed, in which electron-donating substituents promote the reaction, and electron withdrawing substituents inhibit it. This would be the case for other mechanisms, such as oxidative addition, but significant steric effects would also be expected as well, unless geometric changes mitigate them. No evidence of significant geometric change or steric hindrance were observed. This observation was confirmed by computationally modeling the reorganization energy, where \( \lambda_i \) was the minor component and \( \lambda_o \) was the much more significant part of \( \lambda \). This is consistent with solvent polarization as the rate-limiting step. This supports an outer-sphere electron transfer model, as described by Marcus, where ferrocene retains its full coordination shell, and a direct electron transfer from the reductant to the oxidant takes place with no new bonding of ligands that would require kinetically inhibiting geometry changes in the ferrocene. \( \Delta S^\ddagger \) is correlated with the solvated molecular size of reactants in the expected manner, further highlighting the importance of the solvent for this process, and no steric hindrance is observed during the oxidation of even the bulkiest ferrocene derivatives. Based on this evidence a predominantly outer-sphere electron-transfer mechanism for the oxidation of ferrocenes with DBP was proposed.

### 6.8 Final thoughts and future work

The overall goal of this project was to fill a knowledge gap in this subject matter. The literature is ripe with electrochemical studies of ferrocene with little to no data for chemical oxidations. This situation is understandable. Studying these
processes electrochemically can be done much more quickly, easily, and accurately, than the slow, painstaking reactions done for this project. Electrochemical data do not, however, tell the whole story. A coordination processes cannot occur between ferrocene and an electrode, so the reaction is forced down a single path. That path may be the same as the chemical oxidation, but that was not known for certain. While perhaps not providing a definitive answer, it is my hope that this project has shed some additional light on this question.

Perhaps the biggest question left to answer in this work is whether ferrocenes react through a different pathway with mCPBA than DBP. An expanded study on this question would be an interesting area to investigate and could lead to a greater understanding of ferrocene oxidation, and perhaps to new synthetic routes to create ferrocene derivatives. For example, epoxy functionalized ferrocenes have been reported, offering new potential applications in polymer chemistry [86]. mCPBA is known to form epoxy structures with alkenes, [87] and could possibly do so with ferrocene as well under the right conditions. A study that examines rates of reaction and products for mCPBA + RFc by varying solvent, temperature, or even the effects of catalysts could provide insight into this possibility. Furthermore, the study of ferrocenium cation degradation may also be warranted, since the exact pathway is not clear. This could lead to advancements in high-performance lithium ion batteries, where it has been reported that ferrocene can serve as an anode material [88]. Examining ferrocenium decay rates using different solvent environments, ionic conditions, and pH may lead to new ways to stabilizes these ions in electrolyte solutions. Finally, it could be valuable to create a Fenton-type catalyst, as described in
section 1.4, using decamethylferrocene. Based on the results of this study it may be possible to increase the rate at which environmental containments are degraded significantly. Overall, the chemistry of ferrocene is extremely rich, and though it is the subject of an ever-increasing quantity of research, there remains much still to learn.
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