

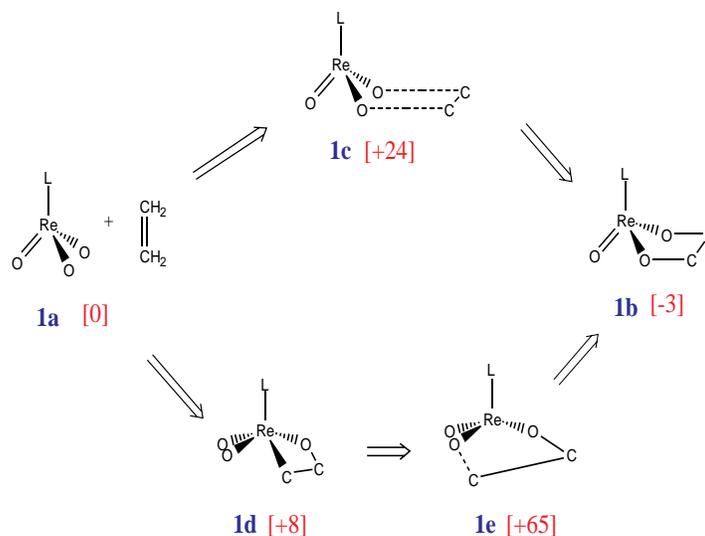
T-12

## Theoretical Chemistry &amp; Molecular Physics

## Rhenium Oxidation Chemistry

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“Concern over the enantiomeric purity of pharmacologically active compounds has made asymmetric synthesis one of the grand challenges of chemical research. Two outstanding examples are the asymmetric *cis* dihydroxylation of olefins and the asymmetric epoxidation of olefins. Both processes utilize high valent transition metal complexes as the catalytic agent, chiral ligands to induce asymmetry in the product, and can use alkylperoxides as oxidizing agents. Complexes of early transition metals such as titanium, vanadium, and chromium preferentially epoxidize olefins. Ruthenium and osmium tetroxide preferentially *cis* dihydroxylate olefins. Why? What is the source of pathway differentiation?”<sup>1</sup>



**Figure 1:** Two paths proposed to be important in the oxidation of ethylene by  $\text{Cp}^*\text{ReO}_3$ . Theoretical energies (kcal/mole) are given in brackets, and strongly support the upper [3 + 2] cycloaddition path.

The recently discovered series  $\text{LReO}_3$  likely holds the key to unraveling this mystery. When  $\text{L} = \text{CH}_3$ , epoxidation occurs. When  $\text{L} = \text{permethyl cyclopentadienyl (Cp}^*)$  a dihydroxylate is formed. When  $\text{L} = \text{O}^{2-}$  the alkene C-C bond is cleaved and an olefin metathesis catalyst is generated.”<sup>1</sup>

An example of the complexity of the mechanistic dilemma is shown in Figure 1. Ethylene reacts with  $\text{Cp}^*\text{ReO}_3$  (**1a**) forming a five-membered ring (diolate, **1b**). In 1933, Criegee suggested that this class of reactions proceeds through a concerted [3 + 2] mechanism (the upper path). In 1977, Sharpless pointed out that

all the known data could as easily be explained by a stepwise mechanism in which an intermediate four-membered ring containing a Re-C bond (metallaioxetane, **1d**) rearranges to form the diolate **1b**. Despite enormous experimental activity, the metallaioxetane has never been directly observed experimentally. A recent review stated it enjoys the distinction of being the most widely invoked and least experimentally validated intermediate in all of oxidation chemistry.

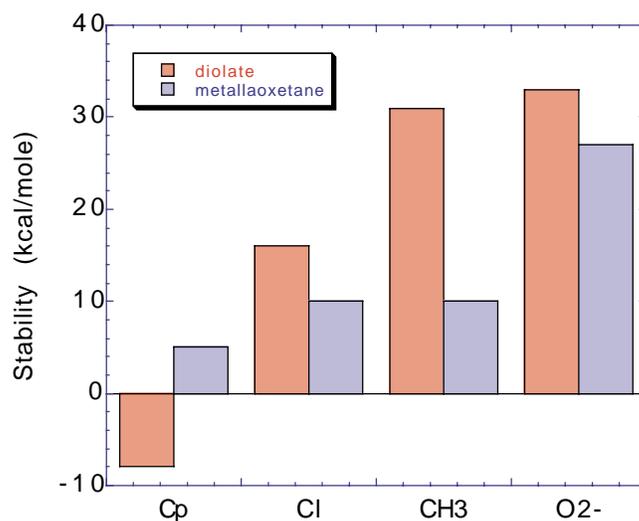
In principle, electronic structure theory possesses an advantage over experiment, since transition states and local minima can be studied regardless of how fleeting their existence in a reaction flask might be. The impediment to such studies is that it is extremely

difficult to obtain reliable heats of formation and activation barriers. This is especially true of reactions involving metal centers. We have recently demonstrated, however, that hybrid density functional theory is capable of providing the necessary thermochemical accuracy in the related (formally  $d^0$ ) metal fluoride series.<sup>2</sup> In addition, we have recently extended our capabilities in relativistic effective core potentials to the analytic evaluation of second derivatives of the energy, thereby dramatically improving our ability to locate transition states.<sup>3</sup> Encouraged by these developments, A. K. Rappe’s group at Colorado State University and I have begun to study this chemistry.

The reaction in Figure 1 was chosen for initial study because of the availability of reliable experimental data with which to test the theory, and the fact that the chemistry takes place in nonpolar solvents so that the complexities of aqueous environments need not be considered. Gable<sup>4</sup> has shown that with  $L = \text{Cp}^*$  the reaction  $1a \rightarrow 1b$  in benzene is slightly exothermic, the diolate being bound by  $5 \pm 3$  kcal/mole. The barrier for extrusion of the ethylene from the diolate is 28 kcal/mole, which, by the principle of microscopic reversibility, can be associated with a reaction barrier in the forward direction of 23 kcal/mole.

The relative energies determined for  $L = \text{Cp}$  by the hybrid density functional method are presented in brackets in Figure 1. When corrected for zero-point energies, the reaction is predicted to be exothermic by 3 kcal/mole, in excellent agreement with experiment.<sup>5</sup> The metallaoxetane lies 8 kcal/mole above the reactants. The transition state for the concerted  $[3 + 2]$  pathway lies 24 kcal/mole above the reactants, consistent with the experimental barrier of 23 kcal/mole. On the other hand, the barrier between the metallaoxetane and the diolate (65 kcal/mole) lies much too high to play a role

in the reaction. Additional support for the  $[3 + 2]$  pathway is found in the measured kinetic isotope effect when  $\text{C}_2\text{D}_4$  is substituted for  $\text{C}_2\text{H}_4$ . Experiment finds a modification of the rate constant of 1.07 per deuterium, whereas theory yields 1.09 for the upper pathway and 1.24 for the lower. This is compelling theoretical evidence for the concerted  $[3 + 2]$  mechanism.



**Figure 2: The relative stabilities of the diolate and metallaoxetane as a function of the ligand. The results for  $L = \text{Cp}$  differ slightly from that in Figure 1 because they have not been corrected for zero-point motion.**

Some preliminary insight into the questions raised in the opening paragraphs has been obtained by comparing the relative energies of the diolate and metallaoxetane for  $L = \text{Cp}$ ,  $\text{Cl}$ ,  $\text{CH}_3$ , and  $\text{O}^{2-}$  (Figure 2). Interestingly, for  $L = \text{CH}_3$ , the metallaoxetane is favored and the diolate product is quite endothermic. Consistent with this, for  $L = \text{CH}_3$  no hydroxylation product is observed experimentally. Instead, with the addition of  $\text{H}_2\text{O}_2$ , epoxidation of the ethylene takes place. We have not yet studied the epoxidation channel, but

additional work mapping out the local minima and transition states of the reactions as a function of the ligand  $L$  may suggest further experimental investigations and help unravel the mechanisms of partial oxidation in these extremely interesting systems.

<sup>1</sup> M. A. Pietsch, T. V. Russo, R. Murphy, R. L. Martin, and A. K. Rappe, submitted to *Angewandte Chemie*.

<sup>2</sup> T. V. Russo, R. L. Martin, and P. J. Hay, *J. Chem. Phys.* **102**, 8023(1995).

<sup>3</sup> T. V. Russo, R. L. Martin, P. J. Hay, and A. K. Rappe, *J. Chem. Phys.* **102**, 9315(1995).

<sup>4</sup> K. P. Gable and T. N. Phan, *J. Am. Chem. Soc.* **116**, 833(1994).

<sup>5</sup> We have shown that the theoretical exothermicity differs by only 2 kcal/mole between  $L = \text{Cp}$  and  $L = \text{Cp}^*$ , and have chosen to examine the transition states based on  $L = \text{Cp}$ .