Mechanistic Study of the Oxidation of a Methyl Platinum(II) Complex with O$_2$ in Water: Pt^{II}Me-to-Pt^{IV}Me and Pt^{II}Me-to-Pt^{IV}Me$_2$ Reactivity

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S Supporting Information

ABSTRACT: The mechanism of oxidation by O$_2$ of (dpms)-Pt^{II}Me(OH$_2$) (**1**) and (dpms)Pt^{II}Me(OH)$^-$ (**2**) [dpms = di(2-pyridyl)methanesulfonate] in water in the pH range of 4−14 at 21 °C was explored using kinetic and isotopic labeling experiments. At pH ≤ 8, the reaction leads to a C_1-symmetric monomethyl Pt^{IV} complex (dpms)Pt^{IV}Me(OH)$_2$ (**5**) with high selectivity ≥97%; the reaction rate is first-order in [Pt^{II}Me] and fastest at pH 8.0. This behavior was accounted for by assuming that (i) the O$_2$ activation at the Pt^{II} center to form a Pt^{IV} hydroperoxo species **4** is the reaction rate-limiting step and (ii) the anionic complex **2** is more reactive toward O$_2$ than neutral complex **1** (pK_a = 8.15 ± 0.02). At pH ≥ 10, the oxidation is inhibited by OH$^-$ ions; the reaction order in [Pt^{II}Me] changes to 2, consistent with a change of the rate-limiting step, which now involves oxidation of complex **2** by Pt^{IV} hydroperoxide **4**. At pH ≥ 12, formation of a C_1-symmetric dimethyl complex **6**, (dpms)Pt^{IV}Me$_2$(OH), along with [(dpms)Pt^{II}(OH)$_2$]$^-$ (**7**) becomes the dominant reaction pathway (50−70% selectivity). This change in the product distribution is explained by the formation of a C_s-symmetric intermediate (dpms)Pt^{IV}Me(OH)$_2$ (**8**), a good methylating agent. The secondary deuterium kinetic isotope effect in the reaction leading to complex **6** is negligible; k_H/k_D = 0.98 ± 0.02. This observation and experiments with a radical scavenger TEMPO do not support a homolytic mechanism. A S$_N$2 mechanism was proposed for the formation of complex **6** that involves complex **2** as a nucleophile and intermediate **8** as an electrophile.

■ INTRODUCTION

The use of dioxygen for selective transition-metal-catalyzed partial C−H oxidation has grown considerably since pioneering works by Hay and Shilov.[1−8] One of the key steps of such transformations may include reaction of O$_2$ with organoplatinum(II) intermediates,[9−14] resulting, in turn, from activation of organic substrates with platinum(II) complexes. Previously, we reported a number of reactions, allowing for a highly selective and facile functionalization of Pt^{II}−C(sp^3) bonds with O$_2$ in water that can lead to methanol, dimethyl ether, ethylene oxide, ethylene glycol, some cycloolefin oxides, and ethanolamines, all enabled by the di(2-pyridyl)-methanesulfonate ligand (dpms).[8] The general reaction sequence involves direct oxidation of (dpms)Pt^{II} monohy-drocarbyl complexes with O$_2$ to form corresponding Pt^{IV} complexes (step a in Scheme 1), their isomerization to produce more electrophilic species featuring the sulfonate group *trans* to the hydrocarbyl (step b in Scheme 1), and subsequent C−O reductive elimination of the isomerized Pt^{IV} complex to form oxygenated organic products (step c in Scheme 1). While the C−O elimination of Pt^{IV} monohydrocarbyls[15] and, in particular, (dpms)Pt^{IV} alkyl complexes (steps b and c in Scheme 1)[11−14] has been studied in some detail, little is known about the

mechanism of reaction of O$_2$ and (dpms)Pt^{II} hydrocarbyl species (step a in Scheme 1). The mechanism of oxidation by O$_2$ of diimine and diamine dimethyl Pt^{II} complexes was introduced by Bercaw et al.[10] The Bercaw mechanism was adapted by us previously for the oxidation of Pt^{II} mono-hydrocarbyls (dpms)Pt^{II}R(OH$_2$), but no detailed experimental testing of this mechanism was performed (Scheme 2).[14] In this work, we studied in greater detail the reaction of O$_2$ with the Pt^{II} methyl aqua complex **1**, (dpms)Pt^{II}Me(OH$_2$), and its conjugate base, hydroxo complex **2**, (dpms)Pt^{II}Me(OH)$^-$, in water. The choice of monomethyl Pt^{II} complexes in this work may be of potential value for the development of catalytic aerobic methane-to-methanol conversion.[8] The mechanism in Scheme 2 was probed by means of kinetic experiments and performing a number of mechanistic tests. Two proton-transfer steps are involved in the oxidation of complex **1**. Hence, a broad range of pH 4.1−14.0 was used in our kinetic study. The results of this study may be valuable for the researchers pursuing aerobic catalytic C−H functionalization using not only platinum but also other metals and, in particular, palladium
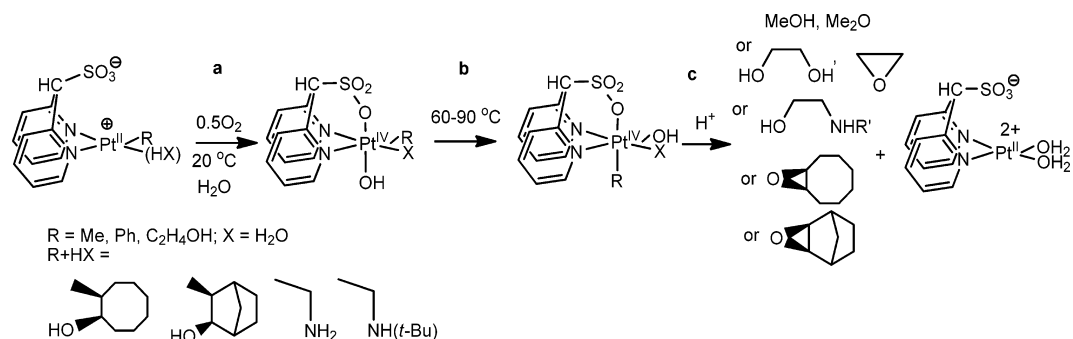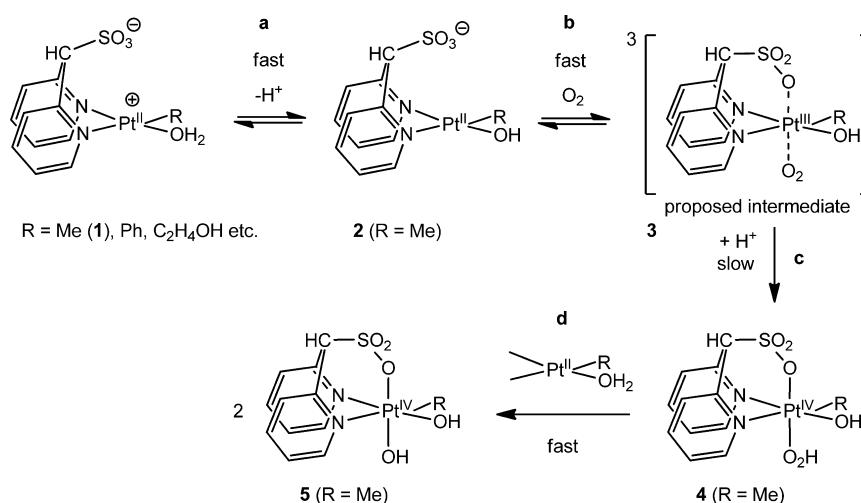
Scheme 1

Scheme 2

Table 1. Distribution of the Reactant Species 1 and 2 and Products[a] of Their Oxidation with O_2 in Water (Equations 1−3) as a Function of the Solution pH at 21 °C

pH	4.1	5.9	8.0	10.0	11.9	14.0
predominant reactant species	1	1	1 and 2	2	2	2
yield of complex 5 (%)	96 ± 1	97 ± 1	99 ± 1	98 ± 1	51 ± 2	30 ± 1
yield of complex 6 (%)	2 ± 1	2 ± 1		2 ± 1	50 ± 2	70 ± 1

[a]Average of two runs. Yields are calculated on the basis of the methyl group balance.

complexes. For instance, oxidation with O_2 of some palladium-(II) hydrocarbyl complexes, where reactivity of the metal center is modulated with help of facially chelating polydentate ligands, was proposed to operate via a mechanism analogous to one in Scheme 2.[3,16a−c] Intriguingly, in this work, we found an alternate reaction direction of oxidation of complexes 1 and 2 leading to Pt-to-Pt methyl group transfer and formation of a dimethyl Pt[IV] product. The latter reaction may be valuable as a model in research targeting oxidative C−H coupling of methane to ethane mediated by platinum or palladium complexes.[16d] The mechanistic results presented here have been supported by extensive first-principles density functional theory studies reported in detail in a companion paper.[17]

■ RESULTS AND DISCUSSION

Equilibrium between Complexes 1 and 2. A quantitative analysis of reactivity of complexes 1 and 2 in the wide pH range 4.1−14.0 used in this work requires knowledge of the equilibrium distribution of complex 1 and its conjugate base 2 as a function of a solution pH, which was maintained in our oxidation experiments using selected buffers (Table 1; see also the Supporting Information). The $pK_a = 8.15 ± 0.02$ for complex 1 was determined using potentiometric titration of the hydroxo complex 2 in water with sulfuric acid. On the basis of the pK_a value found, at pH ≤ 5.9 in ∼10 mM solutions of the aqua complex 1, $(dpms)Pt^{II}Me(OH_2)$, complex 1 is the predominant species (>99%). The fraction of complex 1 drops to less than 1% at pH ≥ 10.0, decreasing linearly with $[H^+]$ at higher pH. In turn, complex 2, $(dpms)Pt^{II}Me(OH)^-$, is the predominant species (>99%) at pH ≥ 10. At pH 8.0, both complexes 1 and 2 are present in solution in comparable concentrations.

Products of Oxidation of Complexes 1 and 2 in the pH Range of 4.1−14.0. The oxidation of complex 1 to form C_1-symmetric $(dpms)Pt^{IV}Me(OH)_2$ (5) in unbuffered solutions, at pH ∼ 5, was reported earlier.[11] We now find that oxidation of complexes 1 (eq 1) and/or 2 (eq 2) to form complex 5 is possible in the broader pH range of 4.1−14.0.

The oxidation of equilibrium mixtures of complexes 1 and 2 is 100% selective at pH 8.0 (Table 1) and ≥97% selective in the

(1)

(2)
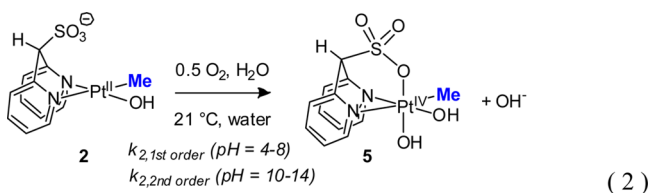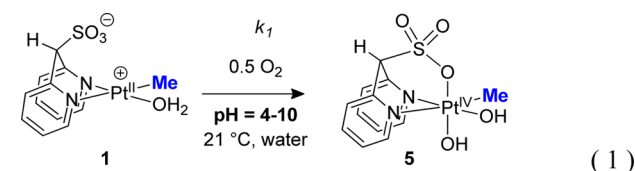
pH range of 4.1−10.0. At pH less than 4.1, the protonolysis of complex **1** with the loss of methane becomes noticeable.

Surprisingly, at pH ≥ 12, the selectivity in complex **5** of the reaction between complex **2** and O_2 drops dramatically to 30−51% (yields are calculated on the basis of the methyl group balance). The decreasing selectivity in complex **5** is solely due to the increasing contribution of the oxidation−methyl transfer reaction 3.

(3)
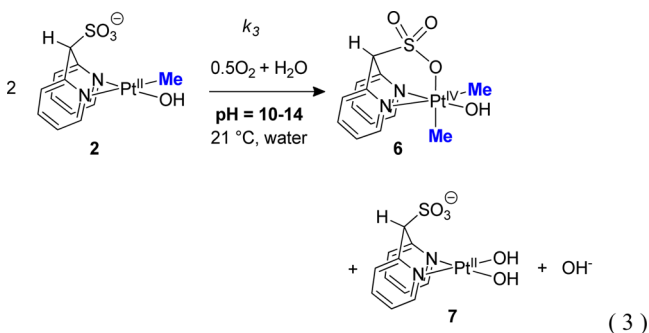
A C_1-symmetric dimethyl Pt^{IV} complex $(dpms)Pt^{IV}Me_2(OH)$ (**6**) is the only other Pt^{IV}-containing product, besides complex **5**, that we observed in our experiments. This product is not detected at pH 8.0; in all other cases at pH ≤ 10.0, complex **6** forms in a trace amount of 2% or less along with either cationic diaqua complex $(dpms)Pt^{II}(OH_2)_2^{+}$[11] (pH ≤ 6.1) or anionic dihydroxo complex $(dpms)Pt^{II}(OH)_2^{-}$ (**7**)[11] (pH 10.0). In more alkaline solutions at pH ≥ 11.9, the fraction of product **6** increases to 50−70%. The existence of the second oxidation pathway leading to complex **6** and the dramatic change in the contribution of reaction 3 to the product distribution at high pH values are intriguing and require a mechanistic explanation.

Possible Reaction Pathways Leading to the Formation of Complex 6 (Reaction 3). The formation of the dimethyl Pt^{IV} complex **6** can be viewed as a result of a Pt^{II}-to-Pt^{IV} or Pt^{II}-to-Pt^{III} aerobic oxidation and a subsequent Pt-to-Pt methyl group transfer, electrophilic or homolytic. The Pt^{IV}-to-Pt^{II} electrophilic methyl group transfer involving methylplatinum(IV) and nucleophilic dimethylplatinum(II) complexes has been reported previously by Puddephatt and co-workers.[18] Such reactions may proceed via a S_N2 mechanism, where the metal atom in the dimethyl Pt^{II} complex acts as a nucleophile with respect to an electrophilic methyl group carbon of the Pt^{IV} methyl complex.

In contrast, a Pt-to-Pt methyl transfer reported by Tilset and co-workers[20] was proposed to occur via a homolytic mechanism involving a Pt^{III} species; the Pt^{III} transients could be generated

via an one-electron chemical or electrochemical oxidation of Pt^{II} dimethyl diimine complexes. A homolytic mechanism of Pt-to-Pt alkyl transfer was also discussed in some reactions of primary (ethyl) or secondary (isopropyl) alkyl halides reacting with dimethyl Pt^{II} complexes.[19] In our systems, no other methyl-containing products besides complexes **5** and **6** were detected in the resulting reaction mixtures and the methyl group balance was virtually 100% (Table 1). Because oxygen can serve as a good radical trap with respect to free alkyl radicals, the absence of such products suggests that free methyl radicals are not involved in the oxidation−methyl transfer reaction 3. Additional tests on the free radical reaction mechanism of oxidation of complexes **1** and/or **2** included additive of a radical scavenger TEMPO[16a] (3−17 equiv). These tests performed at pH 5.9, 10.0, and 14.0 showed neither the reaction rate inhibition nor any change of its product distribution. Hence, a possible reaction pathway leading to dimethyl Pt^{IV} complex **6** may operate via an electrophilic Pt^{IV}-to-Pt^{II} methyl transfer or a non-free radical methyl transfer mechanism involving Pt^{III} species.[20]

To test the possibility of an electrophilic methyl group transfer in reaction 3, ∼10 mM solutions of complexes **5** and **2** taken in a 1:1 molar ratio were reacted at pH 14.0. At this pH, the yield of the methyl transfer product **6** is the highest in our aerobic oxidation experiments (Table 1). A slow first-order reaction leading to the accumulation of complex **6** was observed with a reaction half-life of about 10.6 days, $k_4 = (7.6 \pm 0.1) \times 10^{-7}$ s^{-1}, at 21 °C (eq 4).

(4)
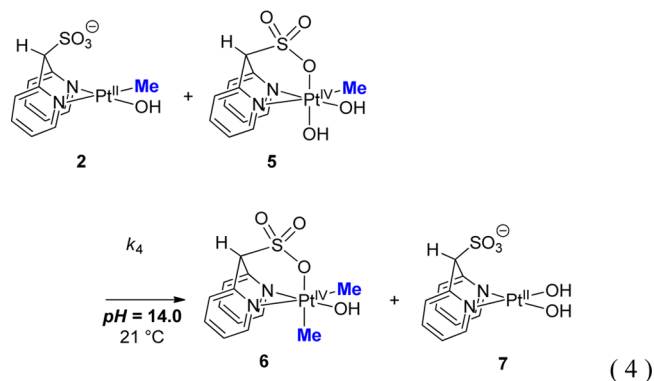
The observed reaction rate is too slow (*vide infra*) to account for the selective formation of complex **6** found in our experiments (Table 1). In addition, a comparison of the order of reactions 2 and 3 in complex **2** also argues against formation of complex **6** via intermediacy of complex **5** as follows. The order in complex **2** is 2 for reaction 2 and 1 for reaction 3 (*vide infra*), which rules out the reaction sequence 5

(5)
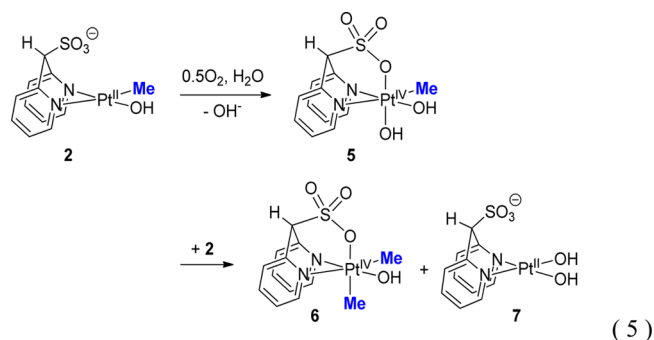
as a mechanism of reaction 3. Another possibility for realization of an electrophilic methyl group transfer in reaction 3 implies involvement of a C_s-symmetric complex **8**, which is known to be much more electrophilic than complex **5**. In particular, complex **8** was shown to eliminate methanol readily in acidic and basic solutions at 20 °C, whereas complex **5** in acidic solutions does not eliminate methanol directly (compare to Scheme 1).[11,12]

To check if complex **8** may be a kinetically viable intermediate responsible for the formation of the dimethyl complex **6** in our oxidation experiments, complex **8** was reacted with complex **2** at pH 14.0. When ∼10 mM complex **8** was combined with complex **2** in a 1:1 molar ratio, a rapid reaction was observed, producing in less than 5 min exclusively the product of the Pt^{IV}-to-Pt^{II} methyl group transfer **6** and an equimolar amount of complex **7** (eq 6).

(6)
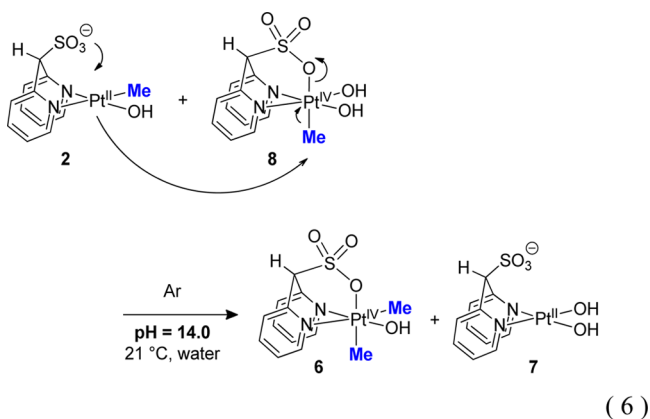
No methanol was formed in this reaction; this fact suggests that the nucleophilicity of the competing Pt^{II} complex **2** is many orders of magnitude greater than that of OH$^-$. Taking the observations above into account, we propose that one of the possible mechanisms of the oxidation–methyl transfer reaction 3 involves the formation of the C_s-symmetric monomethyl Pt^{IV} complex **8** concurrently with its C_1-symmetric isomer **5**. In contrast to complex **5**, complex **8** is not detectable in reaction mixtures by means of nuclear magnetic resonance (NMR) spectroscopy because it is trapped rapidly by the second mole of complex **1** or **2**. This proposed mechanism of formation of complex **6** is summarized in eq 7.

(7)
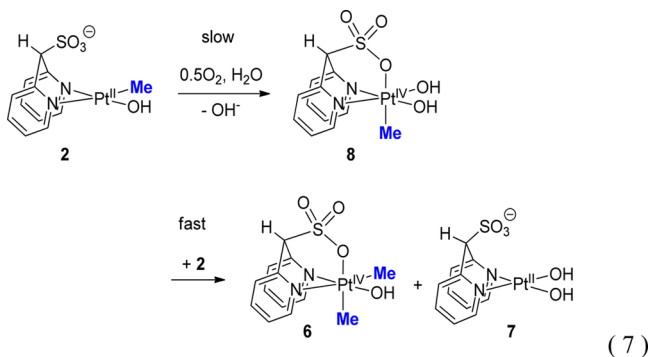
Hence, for the dimethyl Pt^{IV} complex **6** to form as a minor product, the rate of formation of complex **8** should be slow compared to that of complex **5** in the pH range of 4.1–10.0 but, at pH ≥ 11.9, complex **8** should form at a comparable rate.

These hypotheses are tested and validated in the following kinetic study of reactions 1–3.

The second possible mechanistic scenario that could be used to account for the formation of complex **6** includes one-electron oxidation of complexes **1** and/or **2** by O_2 to produce a Pt^{III} intermediate that can transfer its methyl group as a methyl radical to another methyl Pt^{II} species **1** or **2**. The subsequent oxidation of the resulting dimethyl Pt^{III} transient would lead to the observed product **6**. A greater than 1.2 secondary deuterium kinetic isotope effect, k_H/k_D, in reaction 3 may be a sign of a radical mechanism.[19] Such secondary deuterium kinetic isotope experiments were performed in this work.

Reaction Kinetics. Previously, a dependence of the oxidation rate of complex **1** on the partial pressure of O_2 was reported.[21] Hence, solutions containing complexes **1** and **2** (a mixture of complexes **1** and **2** is designated further as **1/2**) were stirred vigorously under 1 atm of O_2 at 21 °C to maintain the concentration of O_2 in reaction solution at a stationary level.

To probe the reaction mechanism in Scheme 2, the order of the net oxidation reaction (eq 1 plus eqs 2 and 3) in [**1/2**] was determined using plots of $\ln([\mathbf{1/2}]_o/[\mathbf{1/2}])$ versus time and $1/[\mathbf{1/2}]$ versus time. The quality of two linear fits was compared, and the corresponding rate constant k_{net} was calculated from the best-fit plot. These plots are given in the Supporting Information, and the net reaction orders in [**1/2**] versus pH are given in Table 2. The observed rate constants for combined reactions 1 plus 2 and individual reactions 2 and 3, $k_{(1+2)}$ (first order), $k_{2,\text{second order}}$ (second order), and k_3, respectively, were calculated as described below and are given in Table 3, along with the net oxidation reaction half-life. The dependence of the reaction half-life (eq 1 plus eqs 2 and 3) versus pH is shown in Figure 1.

pH Range of 4.1−8.0. At pH 4.1, 5.9, and 8.0, the rate of disappearance of the starting monomethyl Pt^{II} complex follows first-order kinetics (Table 2), rate = $k_{net}[\mathbf{1/2}]$. This result is consistent with the mechanism in Scheme 2 if either step a, b, or c but not d is rate-limiting. We postulate here that the rate-limiting step is the spin-forbidden step c. The nature of this reaction step is analyzed in our theoretical paper.[17]

It was also of interest to find the contributions of reactions 1−3 in k_{net} (net oxidation process). In particular, such analysis could allow for a comparison of the relative reactivity of neutral methyl Pt^{II} complex **1** (first-order rate constant k_1; eq 1) and its conjugate base **2** (first-order rate constant $k_{2,\text{first order}}$; eq 2) in oxidation reactions 1 and 2, respectively. The net rate constant k_{net} was separated into the contributions of reactions 1 plus 2 with the observed rate constant $k_{(1+2)}$ and reaction 3 with the observed first-order rate constant k_3 using the product **5/6** ratio given in Table 1

$$k_{net} = k_{(1+2)} + k_3 \quad \text{and} \quad k_{(1+2)}/k_3 = [\mathbf{5}]/[\mathbf{6}]$$
for pH 4.1 and 5.9

$$k_{net} = k_{(1+2)} \quad \text{for pH 8.0}$$

where $k_{(1+2)}$ = (fraction of complex **1**)k_1 + (fraction of complex **2**)$k_{2,\text{first order}}$.

Further separation of $k_{(1+2)}$ into individual rate constants for reactions 1 and 2, k_1 and $k_{2,\text{first order}}$, respectively, could be performed assuming an equilibrium distribution of complexes **1** and **2** in reaction mixtures, which corresponds to the fraction of complex **2** less than 0.01% at pH 4.1, ∼0.7% at pH 5.9, and

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Table 2. Order in [1/2] of the Net Oxidation Reaction, Equation 1 plus Equations 2 and 3, at pH 4.1−14.0, Combined Reaction 1 plus 2 at pH ≤ 8.0, and Reaction 3 in Water, at 21 °C[a]
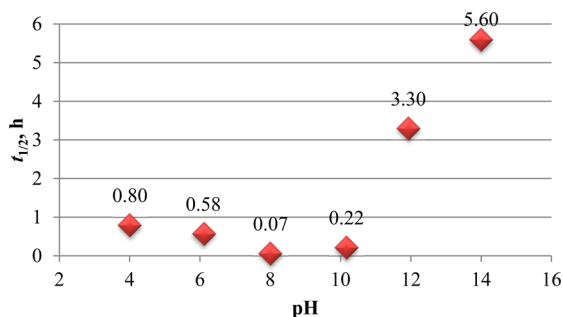
pH	net oxidation reaction, eq 1 plus eqs 2 and 3	combined reactions 1 plus 2 at pH ≤ 8.0	reaction 2 at pH ≥ 10.0	reaction 3
4.1	1	1		(1)
5.9	1	1		(1)
8.0[b]	1	1		
10.0	2		2	1
11.9	1−2		2	1
14.0	1		2	1

[a]The order for reaction 3 at pH ≤ 10 is presumed and given in parentheses. [b]At pH 8.0, reaction 3 is not observed.

Table 3. Half-Life of the Net Oxidation Reaction and the Observed Rate Constants for Combined Reaction 1 plus 2, Reaction 2, and Reaction 3, $k_{(1+2)}$, $k_{2,\text{second order}}$, and k_3, respectively, at [1/2] ∼ 10 mM in Water at 21 °C

entry	pH	net reaction half-life, $t_{1/2}$ (h)	observed reaction 1 plus 2 first-order rate constant, $k_{(1+2)}$ (s^{-1})	second-order reaction 2 rate constant, $k_{2,\text{second order}}$ (M^{-1} s^{-1})	reaction 3 rate constant, k_3 (s^{-1})
1	4.1	0.80 ± 0.03	$(2.4 \pm 0.1) \times 10^{-4}$		$(4.9 \pm 0.2) \times 10^{-6}$
2	5.9	0.58 ± 0.03	$(3.3 \pm 0.2) \times 10^{-4}$		$(6.7 \pm 0.2) \times 10^{-6}$
3	8.0[a]	0.065 ± 0.003	$(2.6 \pm 0.1) \times 10^{-3}$		
4	10.0	0.22 ± 0.01		0.32 ± 0.03	$(9.1 \pm 0.2) \times 10^{-6}$
5	11.9	3.3 ± 0.1		$(7.9 \pm 0.1) \times 10^{-3}$	$(3.2 \pm 0.1) \times 10^{-5}$
6	14.0	5.6 ± 0.3		$(1.13 \pm 0.04) \times 10^{-3}$	$(2.6 \pm 0.1) \times 10^{-5}$

[a]At pH 8.0, reaction 3 is not observed and the net rate constant $k_{\text{net}} = k_{(1+2)}$.

Figure 1. Half-life for the net oxidation reaction (eq 1 plus eqs 2 and 3) versus pH at 21 °C in water. [1/2] ∼ 10 mM.

∼40% at pH 8.0. This analysis resulted in the following values: $k_1 = (2.4 \pm 0.1) \times 10^{-4}$ s^{-1} (eq 1) and $k_{2,\text{first order}} = (6.0 \pm 0.2) \times 10^{-3}$ s^{-1} (eq 2; see details in the Supporting Information). Hence, the rate constant for oxidation of the anionic Pt^{II}Me complex **2**, $k_{2,\text{first order}}$, is about 25 times greater than k_1 for the neutral complex **1**; $k_{2,\text{first order}}/k_1 \sim 25{:}1$. Because complex **2** is more reactive than complex **1**, the net oxidation reaction half-life becomes progressively shorter when going from pH 4.1 to 8.0 (Figure 1).

A similar quantitative analysis of the relative reactivity of complexes **2** and **1** in reaction 3 could not be performed because the latter reaction is not observed at pH 8.0 and the fraction of complex **2** is too low at both pH 4.1 and 5.9 to estimate the reactivity of complex **2** reliably. Qualitatively, complex **2** is more reactive than complex **1** in an oxidation−methyl transfer reaction, such as reaction 3, because the k_3 value calculated at pH 5.9 is greater than that at pH 4.1. Hence, the k_3 values in Table 3 given for this pH range characterize the reactivity of complex **1**.

**Reactions in Strongly Alkaline Solutions, pH ≥ 10.0.** For reactions in strongly alkaline solutions at pH 10.0, 11.9, and 14.0, the contribution of reaction 1 into the net oxidation of **1/2** is negligible because of the very low fraction of complex **1** and because $k_1 \ll k_{2,\text{first order}}$ (see eqs 1 and 2). Two concurrent

reactions, reactions 2 and 3, are involved in the oxidation in this range of pH. Reaction 3 order in complex **2** is 1 at pH 14.0, whereas reaction 2 order in complex **2** is 2 at pH 10.0; the same reaction orders were presumed in the whole pH range.

The individual rate constants for reaction 2, $k_{2,\text{second order}}$, and reaction 3, k_3, were found using a numerical integration of the corresponding rate law (eq 8) and least-squares curve fitting of the experimental data.

$$-\mathrm{d}[\mathbf{2}]/\mathrm{d}t = k_{2,\text{second order}}[\mathbf{2}]^2 + k_3[\mathbf{2}] \qquad (8)$$

pH 10.0. At pH 10.0, the net reaction order in complex **2** is 2 (Table 2). The predominant reactant species present in solution is the anionic complex **2**, and the predominant oxidation reaction is reaction 2 (Table 1). The change in the reaction 2 order at pH 10.0, as compared to the range of pH ≤ 8.0, occurs without a decrease in the reaction selectivity, with complex **5** remaining as the major product. These facts suggest that the rate-determining step of reaction 2 with the observed second-order rate constant $k_{2,\text{second order}}$ is now step d, which involves the second mole of complex **2** (eq 9).

$$(9)$$
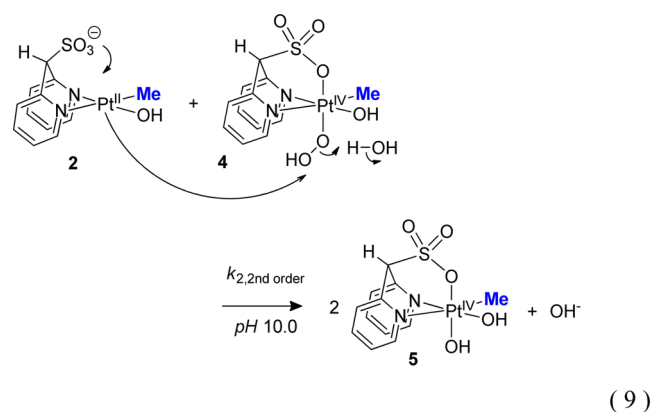
On the basis of the reaction 9 stoichiometry, this reaction is pH-dependent. Some rate deceleration may be expected at

higher pH values. Indeed, at pH > 10.0, the net oxidation reaction half-life increases with increasing pH almost linearly (Figure 1 and Table 3).

pH 11.9 and 14.0. At pH ≥ 11.9, the net oxidation reaction order in complex **2** is intermediate between 1 and 2, as seen, in particular, from plots of $\ln([1/2]_o/[1/2])$ versus time and $1/[1/2]$ versus time for pH 11.9, which are provided in the Supporting Information. The mixed reaction order is a result of the realization of two concurrent processes occurring at comparable rates, reaction 2, which is second-order in complex **2**, and reaction 3, which is first-order in complex **2**. Contribution of the first-order reaction 3 leading to complex **6** increases as pH grows from 11.9 to 14.0, as seen from the increasing **6/5** product ratio (Table 1).

The observed first-order dependence of the reaction 3 rate in complex **2** is consistent with the hypothesis that the rate-limiting step of reaction 3 is the formation of the C_s-symmetric monomethyl Pt^{IV} complex **8**. The subsequent methyl transfer reaction 6 involving complex **2** and intermediate **8** was shown to be very fast under these conditions.

The observed pseudo-first-order rate constant k_3 for the oxidation−methyl transfer reaction 3 increases about 5-fold from pH 4.1 to 14.0, which may be related to the higher reactivity of complex **2** compared to complex **1** in the oxidation step in the reaction sequence 7 and the change of the predominant reactant species from less reactive complex **1** to more reactive complex **2** in this range of pH. Finally, the oxidation−methyl transfer reaction 3 does not involve complex **5** as an intermediate because it follows from the comparison of the rate constant for the oxidation−methyl transfer reaction 3, $k_3 = (2.6 \pm 0.1) \times 10^{-5}$ s^{-1}, found at pH 14.0 and the rate constant for the methyl transfer between complexes **5** and **2** at pH 14.0 (eq 4), $k_4 = (7.6 \pm 0.1) \times 10^{-7}$ s^{-1}. The former is about 34 times greater than the latter. Therefore, the reaction sequence 5 is too slow to account for the observed rate of accumulation of dimethyl Pt^{IV} complex **6**. Hence, two different reaction pathways are responsible for the oxidation of complex **2** to complex **5** and for oxidation−methyl transfer (reaction 3) leading to complex **6**. This hypothesis is validated, and both reaction pathways are analyzed in our theoretical paper.[17]

Secondary Kinetic Deuterium Isotope Effect for Oxidation−Methyl Transfer Reaction 3. To test the possibility of a homolytic methyl transfer in reaction 3 that


(10)
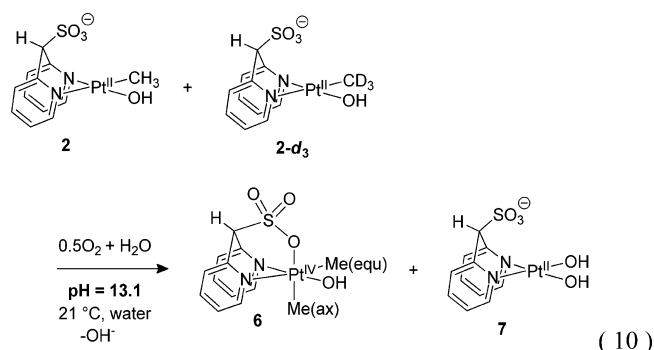
might involve methyl Pt^{III} intermediates, we carried out oxidation of a 1:1 mixture of complexes **2** and **2-d_3**, 13.5 mM each, in water at pH 13.1. The reaction mixture was stirred under O_2 at 20 °C for 1 day, after which time its composition was determined by means of ^1H NMR spectroscopy. The conversion of the isotopologous starting complexes **2** and **2-d_3** was 100%. On the basis of ^1H NMR data, the formation of four isotopologous dimethyl Pt^{IV} complexes **6** was observed with all possible combinations of CH_3 and CD_3 ligands in the axial and equatorial positions (eq 10).

Three of the four dimethyl Pt^{IV} complexes, **6**, **6-d_3** (axial CD_3), and **6-d_3** (equatorial CD_3) produced four partially resolved distinct ^1H NMR signals of their CH_3 ligands in the range of 1.60−2.10 ppm. The isomeric compounds **6-d_3**, one with an axial CD_3 ligand and one with an equatorial CD_3 ligand, were prepared independently by reaction of an appropriate complex **2** or **2-d_3** with CD_3I or CH_3I in water and characterized by ^1H NMR spectroscopy. In both cases, the S_N2 methylation occurs at the axial position of the platinum atom. The assignment of the signals is given in Figure 2.

Two signals of the equatorial CH_3 groups at ∼1.66 ppm correspond to complexes **6** and **6-d_3** (axial CD_3), the products of the CH_3 and CD_3 transfer to complex **2**, respectively. The integration of the CH_3 group signals at 1.663 and 1.659 ppm corresponds to the **6/6-d_3** (axial CD_3) ratio of 0.98 ± 0.02. This very small deuterium kinetic isotope effect does not allow us to conclude whether or not a free methyl radical transfer is involved[19] in the oxidation−methyl transfer reaction 3.

Proposed Reaction Mechanism of Aerobic Oxidation of Complexes 1 and 2. A plausible mechanism for the net
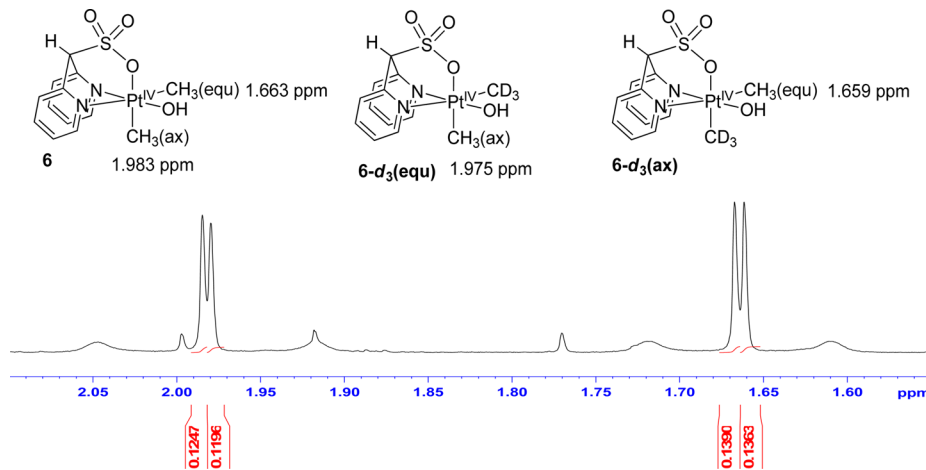
Figure 2. High-field region of a high-resolution 600 MHz ^1H NMR spectrum of the reaction mixture containing **6**, **6-d_3** (axial CD_3), and **6-d_3** (equatorial CD_3) and the assignment of the CH_3 group signals.

Scheme 3
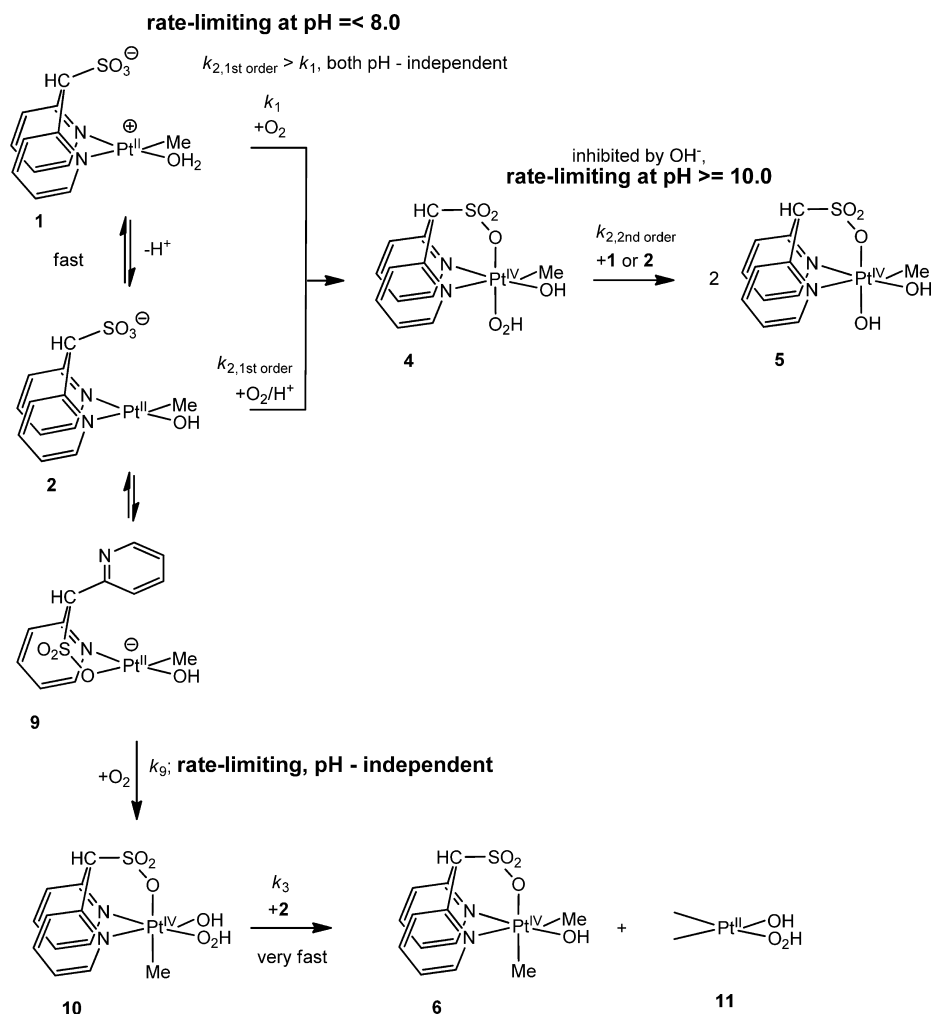
oxidation reaction of complexes 1 and 2 is given in Scheme 3. In our accompanying paper,[17] this mechanism is supported by a theoretical [density functional theory (DFT)] analysis and is discussed in greater detail. In the discussion below, we also provide a brief comment on the consistency of our experimental observations and DFT calculations.

The dioxygen intermediates, such as complex **3** in Scheme 2, have never been observed in platinum(II) chemistry. Accordingly, this species may not exist in our system, or else it may be a relatively high-energy transient. Hence, we do not include complex **3** explicitly in Scheme 3; more work is needed to find experimental evidence for its intermediacy. Our theoretical (DFT) analysis[17] suggests that no stable intermediates, such as complex **3**, are formed along the reaction path from **1/2** to either complex **5** or **6**.

The experimentally observed first-reaction order in [**1/2**] at pH 4.1−8.0 suggests that, in this pH range, the rate-limiting step of the oxidation reactions 1 and 2 is dioxygen activation to form a Pt^{IV} peroxo species **4**. Assuming that the dioxygen activation step by either complex **1** or **2** is pH-independent,[17] the combined reaction 1 and 2 rate increase up to pH 8.0 is due to the increased fraction of the 25-fold more reactive anionic complex **2**. The Pt^{IV}-to-Pt^{II} methyl transfer is slow to occur in this pH range, consistent with our DFT analysis.[17] This is because the major oxidation product, C$_1$-symmetric (dpms)-Pt^{IV}Me(OH)$_2$ (**5**), is a very poor alkylating agent. The

formation of the dimethyl Pt^{IV} complex **6** is presumed to be due to a relatively slow generation of the peroxo species **10** and/or its hydroxo analogue **8** produced in an isomerization/dioxygen activation sequence involving complex **9**, which is an isomer of complex **2**. This hypothesis is supported by theoretical calculations, according to which isomerization of complex **2** to form complex **9** is a relatively low barrier reaction followed by the rate-determining reaction of complex 9 with O$_2$ and water to form complex **10**.[17] On the basis of the net oxidation reaction 1−3 selectivity in monomethyl (**5**) and dimethyl Pt^{IV} (**6**) reaction products, the reaction sequence 7 corresponding to oxidation−methyl transfer (reaction 3) is about 50 times slower than reaction 1 plus 2 at pH 4.1−10.0.

The redox potentials of peroxo compounds, such as complex **4**, depend upon pH: higher pH values correspond to lower potentials. Hence, it is reasonable to assume that the electrophilic hydroxo group transfer between complexes **4** and **2** becomes rate-limiting at pH ≥ 10, consistent with the observed change of the reaction 1 plus 2 order in [**1/2**] from 1 to 2; this assumption was supported by our DFT modeling.[17] Because of the diminished reactivity of complex **4**, formation of complex **5** is inhibited so strongly that reaction 2 is no longer the major direction of the net oxidation reaction at pH ≥ 12.

The reaction that becomes predominant at pH ≥ 12 is the oxidation/Pt-to-Pt methyl transfer (reaction 3). The reaction involves, most likely, the formation of an electrophilic

monomethyl Pt^{IV} hydroperoxo species **10**, followed by a fast S_N2-type Pt^{IV}-to-Pt^{II} methyl transfer from complexes **10** and/or **8** to a Pt^{II} nucleophile **2**. Unlike reactions 1 and 2, where the Pt-to-Pt electrophilic hydroxo group transfer step is severely inhibited at pH $\geq$ 10, the methyl group transfer from complexes **10** and/or **8** to complex **2** to form complex **6** is pH-independent and very fast even at pH 14, in agreement with our DFT calculations. The experimentally measured rate of formation of the dimethyl Pt^{IV} species **6** depends only weakly upon pH. Altogether, because reaction 2 is inhibited by OH^- in strongly alkaline solutions, the rates of reactions 2 and 3 become almost equal and the net oxidation reaction selectivity changes in favor of complex **6** at pH $\geq$ 12.

SUMMARY

The oxidation of the $(dpms)Pt^{II}Me(OH_2)$ complex with O_2 in water includes several pathways, one involving $Pt^{II}Me$-to-$Pt^{IV}Me$ oxidation and another including the oxidation–methyl transfer transformation, $Pt^{II}Me$-to-$Pt^{IV}Me_2$. The second reaction direction contributes only 2% or less in the pH range of 4.1–10.0 but becomes predominant at pH $\geq$ 11.9. At pH < 10.0, the rate-limiting step of the $Pt^{II}Me$-to-$Pt^{IV}Me$ oxidation is the O_2 activation at the Pt^{II} center and the reaction is first-order in $[1/2]$, whereas in more alkaline solutions, the reaction is second-order in complex **2** and the rate-limiting step is presumed to be the electrophilic hydroxo group transfer from a Pt^{IV} hydroperoxo intermediate to the Pt^{II} center. The anionic $Pt^{II}Me$ complex **2** is more reactive than its conjugate acid **1**, which explains the increased reactivity of a mixture of complexes **1** and **2** as the solution pH grows from 4.1 to 8.0. At higher pH, the $Pt^{II}Me$-to-$Pt^{IV}Me$ oxidation is severely inhibited, whereas the rate of the oxidation–methyl transfer sequence $Pt^{II}Me$-to-$Pt^{IV}Me_2$ is accelerated, which explains the dramatic change in the net oxidation selectivity at pH $\geq$ 11.9, where the yield of the $Pt^{IV}Me_2$ product **6** reaches 50–70%. The mechanism of the latter transformation, most likely, involves formation of an electrophilic C_s-symmetric $Pt^{IV}Me$ species **8** and an electrophilic transfer of its methyl group to a nucleophilic $Pt^{II}Me$ complex. The contribution of the radical pathway to the methyl transfer that would include $Pt^{III}Me$ species cannot be excluded. On the basis of the reaction secondary deuterium kinetic isotope effect, we cannot conclude whether or not this pathway contributes significantly to the net oxidation–methyl transfer transformation. The mechanistic results presented here have been supported by extensive first-principles density functional theory studies reported in detail in a companion paper.[17]

ASSOCIATED CONTENT

ⓈSupporting Information

Synthetic procedures and kinetic details. This material is available free of charge via the Internet at http://pubs.acs.org.

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Notes

The authors declare no competing financial interest.

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