Reduction of Permanganate by Thioanisole: Lewis Acid Catalysis

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The oxidation of sulfides and sulfoxides by permanganate in anhydrous acetone solutions is catalyzed by Lewis acids such as iron(III) chloride, zinc chloride, and mercury(II) chloride. The reaction kinetics unequivocally confirm that the function of these catalysts is to activate the oxidant by forming permanganate/Lewis acid complexes analogous to the protonation of MnO_4^- by Brønsted acids. A Hammett analysis of the rate constants for the oxidation of a series of substituted thioanisoles gives a negative ρ value (−1.11) indicative of an electron deficient transition state. No secondary kinetic isotope effect is observed when the hydrogens of the methyl group are replaced by deuterium. Despite previous observations that sulfoxides are preferentially oxidized in competitive experiments, sulfides are oxidized more rapidly when individual rates are measured. All of these observations are most consistent with a mechanism in which the reductant reacts with the oxidant via initial ligation.

Introduction

Lewis acids are effective catalysts for several useful oxidative processes including Oppenauer[1a] and Baeyer−Villiger[1b] oxidations, conversion of thiols to disulfides by reaction with sulfoxides[1c] or manganese dioxide,[1d] oxidative rearrangements catalyzed by selenium dioxide,[1e] the oxidation of unsaturated alcohols by manganate(VI),[1f] and the oxidation of sulfides to sulfones by permanganate.[1g,h] The latter reaction, oxidation of sulfides and sulfoxides to sulfones by permanganate in anhydrous acetone, is possible only if Lewis acids such as ferric chloride or zinc acetate are present. The essential role of Lewis acids in these reactions is exemplified by the observation that thioanisole is converted into the corresponding sulfone in 100% yield when treated with potassium permanganate and ferric chloride in acetone for 30 min at −25 °C, while, in the absence of ferric chloride or another Lewis acid, only a trace of sulfone is produced under identical conditions. The purpose of this work is to investigate the catalytic effect of Lewis acids on these reactions and their relationship to the reaction mechanism.

Three possibilities exist for Lewis acids to exert a catalytic effect on the oxidation of sulfides by permanganate. For example, the effect of a Lewis acid could be to complex the sulfide ($R_2S + FeCl_3 \rightarrow R_2S^+{-}FeCl_3^-$) making it more sulfoxide-like ($R_2S{=}O \leftrightarrow R_2S^+{-}O^-$). Since it is known that sulfoxides are more rapidly oxidized than sulfides under certain conditions,[1g,2] complexation provides a plausible explanation for the observed catalysis. Alternatively, it is also possible that the catalysis could be due to complexing of the oxidant by the Lewis acid. A reaction between permanganate and a Lewis acid would be analogous to the reaction between permanganate and Brønsted acids which results in the formation of permanganic acid, $HMnO_4$,[3] a strong oxidant that is responsible for the acid catalysis which has been observed during several permanganate oxidations.[4] It may also be noted that formation of a Lewis acid/permanganate complex has been proposed to account for the catalysis observed when hydrocarbons are oxidized by permanganate in the presence of boron trifluoride.[5] When the reductant is a hydrocarbon, it is not possible to attribute the catalysis to formation of a Lewis acid/reductant complex. A third reaction mechanism that must also be considered is the possibility that the oxidant and reductant could form an intermediate which undergoes oxidative decomposition with the assistance of a Lewis acid.

It has been possible to distinguish between these three mechanisms from a study of the reaction kinetics with the conclusion that the reaction most likely proceeds by way of a permanganate/Lewis acid complex. Similar results were obtained by using two different Lewis acids, zinc chloride and mercury(II) chloride. The rate constants for a series of substituted thioanisoles and the

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(1) (a) Almeida, M. L. S.; Kocovsky, P.; Backvall, J.-E. *J. Org. Chem.* **1996**, *61*, 6587. (b) Kotsuki, H.; Arimura, K.; Araki, T.; Shinohara, T. *Synlett* **1999**, 462. (c) AbuOmar, M. M.; Khan, S. I. *Inorg. Chem.* **1998**, *37*, 4979. (d) Firouzabadi, H.; Abbassi, M.; Karimi, B. *Synth. Commun.* **1999**, *29*, 2527. (e) Shafer, C. M.; Morse, D. I.; Molinski, T. F. *Tetrahedron* **1996**, *52*, 14475. (f) Firouzabadi, H.; Karimi, B.; Abbassi, M. *J. Chem. Res. S* **1999**, 236. (g) Block, E.; DeOrazio, R.; Thiruvazhi, M. *J. Org. Chem.* **1994**, *59*, 2273. (h) Thiruvazhi, M. Ph.D. Dissertation, SUNY−Albany, Albany, New York, 1993.

(2) (a) Ogura, K.; Suzuku, M.; Tsuchihashi, G. *Bull. Chem. Soc. Jpn.* **1980**, *53*, 1414. (b) Poje, M.; Balenovic, K. *Tetrahedron Lett.* **1978**, 1231. (c) Block, E.; Corey, E. R.; Penn, R. E.; Renken, T. L.; Sherwin, P. E.; Bock, H.; Hirabayshi, T.; Mohmand, S.; Solouki, B. *J. Am. Chem. Soc.* **1982**, *104*, 3119. (d) Rajagopal, S.; Sivasubramanian, G.; Suthakaran, R.; Srinivasan, C. *Proc. Indian Acad. Sci.* **1991**, *103*, 637. (e) Fatiadi, A. J. *Synthesis* **1989**, 85.

(3) Frigerio, N. A. *J. Am. Chem. Soc.* **1969**, *91*, 6200.

(4) (a) Stewart, R. In Oxidation in Organic Chemistry, Part A.; Wiberg, K. B., Ed., Academic Press: New York, 1965; p 35. (b) Verma, R. S.; Reddy, M. J.; Shastry, V. R. *J. Chem. Soc., Perkin Trans. 2* **1976**, 469. (c) Aitken, R. A.; Mesher, S. T. E.; Rosse, B. M. *Synthesis* **1997**, 787.

(5) Lau, T.-C.; Wu, Z.-B.; Bai, Z.-L.; Mak, C.-K. *J. Chem Soc. Dalton Trans.* **1995**, 695.

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corresponding phenyl methyl sulfoxides are also compared, and an attempt has been made to understand the relationship of the relative rates of oxidation of sulfides and sulfoxides with previously reported product studies.

Experimental Section

Materials. Acetone, used as the solvent in all kinetic experiments, was Fisher HPLC grade. Potassium permanganate was BDH AnalaR grade. Anhydrous zinc chloride was obtained from Aldrich and stored in a desiccator.

Thioanisole, *p*-methoxythioanisole, *p*-methylthioanisole, *p*-fluorothioanisole, and *p*-chlorothioanisole, received (from *Aldrich*) as liquids with 97% or greater purity, were carefully distilled under reduced pressure and a middle portion of the distillates collected. GC or HPLC analysis indicated purities of 99.5 to 99.9% for the final products. The observed boiling points were as follows: thioanisole, 114−115 °C/21 Torr; *p*-methoxythioanisole, 139.0−139.5 °C/23 Torr; *p*-methylthioanisole, 114−115 °C/21 Torr; *p*-fluorothioanisole, 89−90 °C/31 Torr; and *p*-chlorothioanisole, 131.0−131.5 °C/23 Torr. *p*-Nitrothioanisole, obtained commercially, was recrystallized from ethanol three times and purified by column chromatography (silica gel; eluent: chloroform/ethyl acetate 1/1). GC analysis indicated a purity of 99.2%, mp 71.8−72.7 °C (lit.[6] 70.5−71.5 °C).

The corresponding sulfoxides were prepared by oxidation of the sulfides with hydrogen peroxide in acetone. In a typical procedure, a flask containing the sulfide (0.12 mol) dissolved in 125 mL of acetone was cooled in an ice−water bath, and 30% H_2O_2 (20 g) was added. The mixture was stirred until TLC showed no sulfide remained. The product was then extracted with chloroform (4 × 50 mL), and the extracts were dried over anhydrous $MgSO_4$. Solvent was removed by flash evaporation, and the remaining mixture of sulfone and sulfoxide was separated by column chromatography (silica gel; eluent: chloroform/ethyl acetate 1/1). Final purification was by recrystallization or distillation under reduced pressure. The products were characterized by use of nuclear magnetic resonance and infrared spectroscopy as indicated below: Phenyl methyl sulfoxide (obtained in 24% yield), bp 161.2−162.0 °C/21 Torr; ^{1}H NMR ($CDCl_3$) δ: 2.70 (s, 3H), 7.50 (m, 3H), 7.65 (d, 2H); IR (cm^{-1}) 1045 (s), 1481 (s), 1582 (s), 2900 (s), 3010 (s). *p*-Methoxyphenyl methyl sulfoxide (obtained in 30% yield) mp 42.1−43.0 °C (lit.[7] 43 °C), ^{1}H NMR ($CDCl_3$) δ: 2.67 (s, 3H), 3.95 (s, 3H), 7.00 (d, 2H), 7.80 (d, 2H); IR (cm^{-1}) 1048 (s), 1255 (s), 1497 (s), 1595 (s), 2837−3200 (s). Methyl *p*-methylphenyl sulfoxide (obtained commercially), mp 41.5−42.5 °C (lit.[7] 42−43 °C), ^{1}H NMR ($CDCl_3$) δ: 2.38 (s, 3H), 2.65 (s, 3H), 7.28 (d, 2H), 7.68 (d, 2H): IR (cm^{-1}) 1056 (s), 1448 (m), 3000 (s). *p*-Fluorophenyl methyl sulfoxide (obtained in 68% yield), bp 135−136 °C /21 Torr, ^{1}H NMR ($CDCl_3$) δ: 2.65 (s, 3H), 7.15 (d, 2H), 7.60 (d, 2H): IR (cm^{-1}) 1049 (s), 1087 (s), 1493 (s), 1590 (m), 2900−3100 (w).

Methyl-d_3 *p*-methylphenyl sulfide was prepared from the reaction of *p*-thiocresol with dimethyl sulfate-d_6. A 100 mL flask fitted with a reflux condenser, a dropping funnel, and a stirrer was immersed in an ice water bath. *p*-Thiocresol (6.47 g, 00.051 mol), tetrabutylammonium bisulfate (0.340 g, 0.001 mol), dichloromethane (20 mL), and 50% NaOH (10.4 g) were added. The solution was stirred vigorously for 30 min, and $(CD_3)_2SO_4$ (7.93 g, 0.06 mol) was added dropwise over 1 h. After stirring at room temperature for an additional 18 h, TLC analysis indicated that no *p*-thiocresol remained. Concentrated NH_4OH (2 mL) was added, and after stirring for another 30 min, water (30 mL) was added to dissolve precipitated white solids. The aqueous layer was separated and extracted with dichloromethane (5 × 20 mL). The organic layers were washed with 33% NaOH solution (2 × 20 mL) and then with water until neutral to pH paper. After being dried over anhydrous magnesium sulfate for 12 h, the solvent was evaporated under reduced pressure leaving a yellow liquid (5.21 g, 74%). This crude product was purified by careful distillation under reduced pressure to give a colorless liquid (2.14 g, 30%), bp 113.5−114.4 °C /21 Torr, ^{1}H NMR ($CDCl_3$) δ: 2.40 (s, 3H), 7.10 (m, 2H), 7.20 (m, 2H); IR (cm^{-1}) 3000 (m), 2050 (m), 1491 (s), 1589 (m).

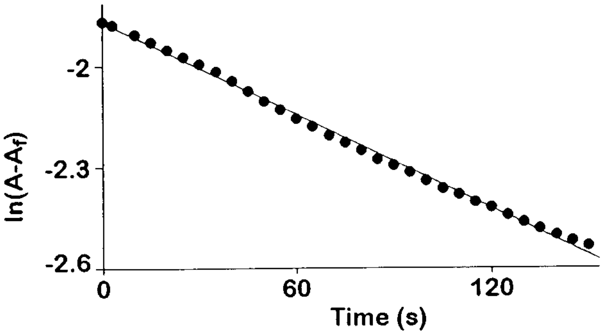

Figure 1. Attempted pseudo-first-order plot.

Kinetic Measurements. It was found experimentally that ferric chloride, the catalyst used for reactions at −25 °C,[1] promoted a rapid reduction of permanganate by the solvent at ambient temperatures, thereby preventing its use in kinetic studies. However, zinc chloride and mercury(II) chloride were found to catalyze the reduction of permanganate by thioanisole at convenient rates without promoting a corresponding rapid reaction between the oxidant and the solvent. The kinetic study of the reduction of permanganate by sulfides and sulfoxides was therefore conducted using anhydrous acetone solutions containing different concentrations of zinc chloride or mercury(II) chloride. In a typical experiment, a solution of zinc chloride in anhydrous acetone was sealed in a 50 mL Erlenmeyer flask and immersed in a constant-temperature bath for 1 h. While the zinc chloride solution was being thermostated, a permanganate solution was prepared by placing a few milligrams of $KMnO_4$ in a 50 mL Erlenmeyer flask and adding 40 mL of anhydrous acetone. After swirling for ca. 30 s, the supernant was transferred, by use of a disposable pipet, to another flask suspended in the constant-temperature bath. The flask was stoppered and sealed to prevent contact with moisture in the air. An aliquot of the zinc chloride solution (2.0 mL) was then transferred to a 10 mm cuvette, and a stock solution of thioanisole in anhydrous acetone (0.10 mL) was added with a microliter syringe. The cuvette was placed in the thermostated cell compartment of an HP8452 diode array spectrophotometer, and after a few minutes permanganate solution (0.50 mL) was added by use of a microliter syringe. The cuvette was quickly inverted several times to ensure good mixing and spectra were collected every five seconds until the reaction was complete. The initial concentration of permanganate was approximately 2 × 10^{-4} M for all experiments, while the concentration of zinc chloride was varied from 5 × 10^{-4} to 7 × 10^{-3} M, and that of thioanisole from 1 × 10^{-2} to 0.12 M. The reaction rates were determined by monitoring the decrease in absorbance at 528 nm and application of standard procedures for kinetic studies.

The Rate Law

Under pseudo-first-order conditions, the approximate linearity of plots of ln($A - A_f$) vs time suggests that the reaction is first order with respect to permanganate. The slight upward curvature of many of these plots (see, for example, Figure 1) indicates that the reaction is subject to autocatalysis, a well-known phenomenon when MnO_2 is one of the products.[8] Possible errors caused by this curvature were avoided by the use of initial reaction rates to determine the rate constants. To prevent the introduction of bias, the initial rates were determined by fitting

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[6] Lindberg, B. J.; Schroder, B. *Acta Chem. Scand.* **1970**, *24*, 3089.

[7] Cerniani, A.; Modena, G. *Gazz. Chim. Ital.* **1953**, *89*, 843.

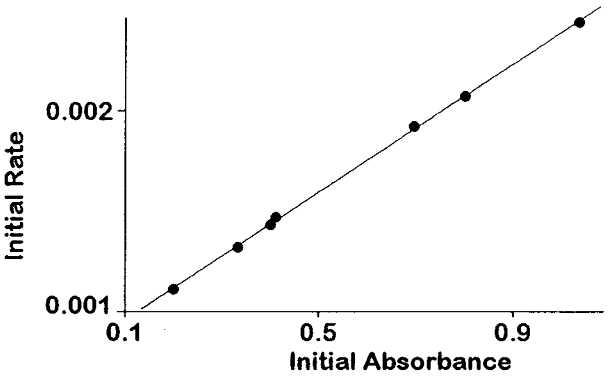

Figure 2. Plot indicating the order with respect to permanganate. Slope = $(1.59 \pm 0.02) \times 10^{-3}\ s^{-1}$; $r = 0.999$.

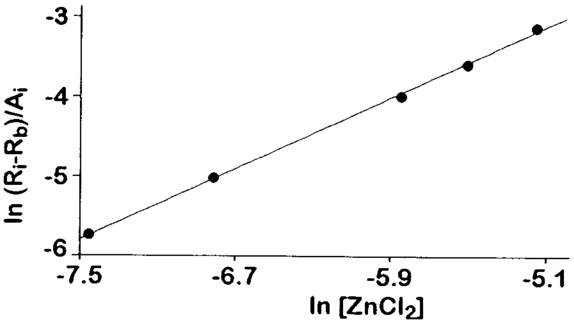

Figure 3. Plot indicating the order with respect to zinc chloride. Slope = 1.12 ± 0.02; $r = 0.999$.

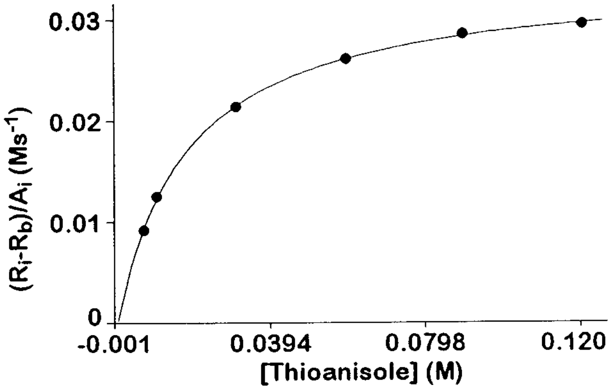

Figure 4. Variation of initial rate with reductant concentration.

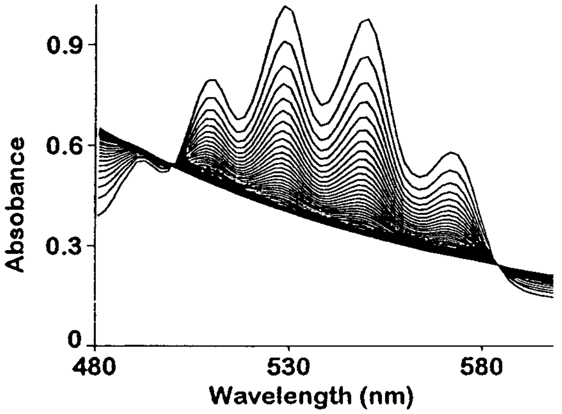

Figure 5. Sequential scans (top to bottom at 528 nm) for the reduction of permanganate (1.62×10^{-4} M) by thioanisole (0.12 M) in anhydrous acetone containing zinc chloride (5×10^{-4} M).

the absorbance/time data to a standard curve, defined by eq 1, as described in previous publications.[9]

$$A = a_0 + a_1t + a_2t^2 + a_3t^3 + a_4t^4 + a_5t^5 \qquad (1)$$

Differentiation with respect to time of the equation obtained from computer modeling of the data, and setting $t = 0$, gives the initial rate, $(-dA/dt)_i$, which has been designated as R_i. Because permanganate is slowly reduced by acetone in the absence of other reductants, it is necessary to correct the observed initial rate for the "blank rate" which has been designated as R_b. This correction did not reduce the initial rate by more than 10% under any conditions. A plot of $R_i - R_b$ (the initial rate) vs A_i (the initial absorbance) is linear (Figure 2), confirming that the reaction is first order with respect to permanganate.

The use of pseudo-first-order conditions and initial reaction rates minimizes the possibility of contributions to the observed rate by consecutive oxidations in which the initial product, a sulfoxide, is converted into the ultimate product, a sulfone.[1,10]

When the concentration of thioanisole was held in a constant large excess while the concentration of zinc chloride was varied, a plot of $\ln(R_i - R_b)/A_i$ vs $\ln[ZnCl_2]$ was linear with a slope of 1.1 ± 0.1 (Figure 3). The rate law is therefore also first order with respect to the Lewis acid.

When the initial rate was plotted against the concentration of thioanisole, the order with respect to thioanisole was observed to change from approximately unity at low concentrations to approximately zero at high concentrations as in Figure 4. A similar observation was made when mercury(II) chloride was used as the catalyst.

Reaction Products. Sequential scanning (Figure 5) indicated that the disappearance of permanganate coincides with the appearance of a product with a spectrum that conforms to Rayleigh's law for light scattering by colloidal particles, $\ln A = -4\lambda + C$.[11] The plot in Figure 6 indicates a good correspondence between the predictions of this law and the final spectrum obtained. Furthermore, if the solutions are allowed to stand undisturbed for 24 h, a brown solid, characteristic of manganese dioxide, slowly precipitates.[12] It appears, therefore, that the initial observable product is colloidal MnO_2.

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(8) (a) Mata-Perez, F.; Perez-Benito, J. *Z. Phys. Chem. Neue Folge* **1984**, *141*, 213. (b) Lee, D. G.; Perez-Benito, J. F. *J. Org. Chem.* **1988**, *53*, 5725. (c) Perez-Benito, J. F.; Arias, C. *Int. J. Chem. Kinet.* **1991**, *23*, 717. (d) Insausti, M. J.; Alvarez-Macho, M. P.; Mata-Perez, F. *Collect. Czech. Chem. Commun.* **1992**, *57*, 2331. (e) Insausti, M. J.; Mata-Perez, F.; Alverez-Macho, M. P. *Int. J. Chem. Kinet.* **1995**, *27*, 507.

(9) (a) Lee, D. G.; Lee, E. J.; Chandler, W. D. *J. Org. Chem.* **1985**, *50*, 4306. (b) Chandler, W. D.; Lee, E. J.; Lee, D. G. *J. Chem. Ed.* **1987**, *64*, 878.

(10) (a) May, B. L.; Yee, H.; Lee, D. G. *Can. J. Chem.* **1994**, *72*, 2249. (b) Ogura, K.; Suzuki, M.; Tsuchihashi, G. *Bull. Chem. Soc. Jpn.* **1980**, *53*, 1414.

(11) (a) Laidler, K. J.; Meiser, J. H. *Physical Chemistry*, 2nd ed.; Houghton Mifflin: Toronto, 1995; p 858. (b) Perez-Benito, J. F.; Arias, C. *J. Colloid Interface Sci.* **1992**, *152*, 70. (c) Perez-Benito, J. F.; Lee, D. G. *Can J. Chem.* **1985**, *63*, 3545.

(12) (a) Perez-Benito, J. F.; Brillas, E.; Pouplana, R. *Inorg. Chem.* **1989**, *28*, 390. (b) Mata-Perez, F.; Perez-Benito, J. F. *Can. J. Chem.* **1985**, *63*, 988.

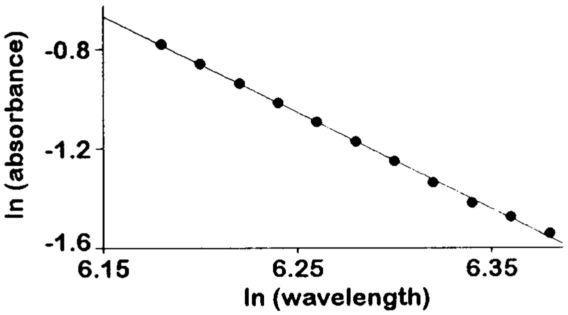

Figure 6. Rayleigh plot for the reaction product (MnO_2). Slope = 3.86 ± 0.04; $r = 0.999$.

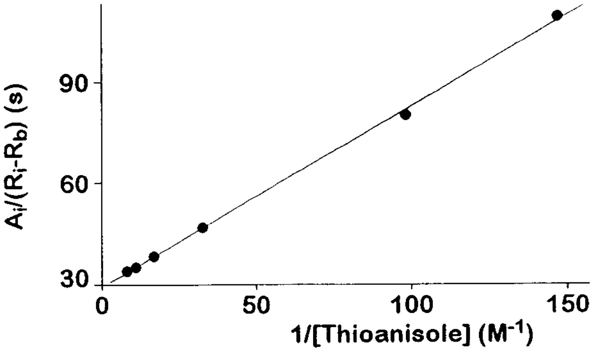

Figure 7. Reciprocal plot in verification of eq 3. Slope = 0.536 ± 0.008 M s; Intercept = 29.1 ± 0.6 s; $r = 0.999$.

Manganate(V) ion, MnO_3^-, formed when an oxygen atom is transferred from permanganate to a sulfide would be a very reactive entity under these conditions; it is known to be stable only in highly alkaline solutions (6–10 M NaOH).[13] If formed in a neutral solution in the presence of an oxidizable solvent such as acetone, reduction to manganese dioxide, the observed product, would be instantaneous.[14]

Results

The experimentally determined rate law (first order in oxidant, first order in Lewis acid, and of variable order between unity and zero in reductant) is of the form expected if zinc chloride acts as a catalyst by combining with permanganate prior to the redox step, as in Scheme 2.

The derived rate law may be restated as in eqs 2 and 3 where $k_{obs} = k_1k_2/k_{-1}$. From eq 3 it is apparent that values for k_1 and k_{obs} can be obtained from the intercepts and slopes, respectively, of plots of $A_i/(R_i - R_b)$ vs 1/[RSR′] as in Figure 7.

$$(R_i - R_b)/A_i = k_{obs}[RSR'][ZnCl_2]/(1 + k_2/k_{-1}[RSR']) \tag{2}$$

$$A_i/(R_i - R_b) = 1/(k_{obs}[ZnCl_2]\,[RSR']) + 1/(k_1[ZnCl_2]) \tag{3}$$

The rate constant, k_1, was found to be independent of the structure of the reductant (Table 1), as would be expected for a reaction between permanganate and zinc chloride as in Scheme 2. The effect of substitutents on the values of k_{obs} in Table 1 must be associated with k_2 because k_1 and k_{-1} would not be dependent on the structure of the reductant.

Discussion

The primary purpose of this investigation is to understand the nature of the observed catalysis that occurs when sulfides are oxidized by permanganate in the presence of Lewis acids. Possible reaction mechanisms involving the complexation of zinc chloride with the reductant, the oxidant, or an intermediate have been summarized in Schemes 1, 2, and 3, respectively. The experimental results, which give a rate law that is first order in oxidant, first order in Lewis acid, and of an order between unity and zero in reductant, clearly indicate that the function of the catalyst is to complex permanganate as in Scheme 2.

Scheme 1. Reaction Sequence and Rate Law When the Intermediate is a Sulfide/Lewis Acid Complex

$$R_2S + ZnCl_2 \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} R_2SZnCl_2$$

$$R_2SZnCl_2 + MnO_4^- \xrightarrow[\text{slow}]{k_2} R_2SO + MnO_3^- + ZnCl_2$$

$$R_2SO + MnO_4^- \xrightarrow{\text{fast}} R_2SO_2 + MnO_3^-$$

$$MnO_3^- \xrightarrow[\text{fast}]{\text{solvent}} MnO_2$$

$$\text{rate} = \frac{k_1k_2[R_2S][MnO_4^-][ZnCl_2]}{k_{-1} + k_2[MnO_4^-]}$$

Scheme 2. Reaction Sequence and Rate Law When the Intermediate is a Permanganate/Lewis Acid Complex

$$MnO_4^- + ZnCl_2 \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} MnO_4ZnCl_2^-$$

$$MnO_4ZnCl_2^- + R_2S \xrightarrow[\text{slow}]{k_2} R_2SO + MnO_3ZnCl_2^-$$

$$R_2SO + MnO_4ZnCl_2^- \xrightarrow{\text{fast}} R_2SO_2 + MnO_3ZnCl_2^-$$

$$MnO_3ZnCl_2^- \xrightarrow[\text{fast}]{\text{solvent}} MnO_2ZnCl_2$$

$$\text{rate} = \frac{k_1k_2[R_2S][MnO_4^-][ZnCl_2]}{k_{-1} + k_2[R_2S]}$$

To understand the catalytic nature of zinc chloride in these reactions, it is necessary to begin with a consideration of the chemical and electronic structure of the oxidant. A tetrahedral ion with four equivalent, highly polarized oxos,[15] permanganate ion may be represented by the use of resonance structures as in Scheme 4.

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[13] Arndt, D. *Manganese Compounds as Oxidizing Agents in Organic Chemistry*; Open Court: LaSalle, Illinois, 1981; p 172.

[14] (a) Zahonyi-Budo, E.; Simandi, L. I. *Inorg. Chim. Acta* **1996**, *248*, 81. (b) Zahonyi-Budo, E.; Simandi, L. I. *Inorg. Chim. Acta* **1992**, *191*, 1. (c) Jaky, M. *Acta Chim. Hung.* **1992**, *129*, 487. (d) Simandi, L. I.; Jaky, M.; Schelly, Z. A. *J. Am. Chem. Soc.* **1984** *106*, 6866.

[15] (a) Sharpless, K. B.; Teranishi, A. Y.; Blackvall, J. E. *J. Am. Chem. Soc.* **1977**, *99*, 3120. (b) Howard, W. A.; Parkin, G. *J. Am. Chem. Soc.* **1994**, *116*, 606.

Table 1. Rate Constants for the Oxidation of Thioanisoles and Methyl Phenyl Sulfoxides by Permanganate in Anhydrous Acetone Solutions

	catalyst			
	$[ZnCl_2] = 1.16 \times 10^{-3}$ M		$[HgCl_2] = 0.120$ M	
reductant	k_1 ($M^{-1}\,s^{-1}$)	k_{obs} ($M^{-2}\,s^{-1}$)	k_1 ($M^{-1}\,s^{-1}$)	k_{obs} ($M^{-2}\,s^{-1}$)
p-methoxythioanisole	25.9 ± 0.1	3630 ± 50	4.68 ± 0.07	317 ± 12
p-methylthioanisole	26.1 ± 0.1	2480 ± 111	4.85 ± 0.05	252 ± 7
thioanisole	26.9 ± 0.4	1580 ± 20	4.83 ± 0.06	203 ± 5
p-fluorothioanisole	27.8 ± 0.4	1430 ± 20	4.72 ± 0.14	183 ± 12
p-chlorothioanisole	26.6 ± 0.1	951 ± 6		
p-nitrothioanisole	23.5 ± 0.2	228 ± 5		
methyl *p*-methoxyphenyl sulfoxide	27.9 ± 0.4	334 ± 5		
methyl *p*-methylphenyl sulfoxide	25.9 ± 1,2	147 ± 4		
methyl phenyl sulfoxide	26.9 ± 1.0	97 ± 1		
methyl *p*-fluorophenyl sulfoxide	26.2 ± 3.1	51 ± 1		

Scheme 3. Reaction Sequence and Rate Law When the Intermediate is a Permanganate/Sulfide Complex

$$R_2S + MnO_4^- \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} R_2SMnO_4^-$$

$$R_2SMnO_4^- + ZnCl_2 \xrightarrow[\text{slow}]{k_2} R_2SO + MnO_3ZnCl_2^-$$

$$R_2SO + MnO_4^- \xrightarrow{\text{fast}} R_2SO_2 + MnO_3^-$$

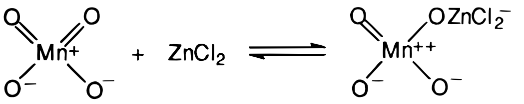
$$MnO_3ZnCl_2^- \xrightarrow[\text{fast}]{\text{solvent}} MnO_2ZnCl_2$$

$$MnO_3^- \xrightarrow[\text{fast}]{\text{solvent}} MnO_2$$

$$\text{rate} = \frac{k_1k_2[R_2S][MnO_4^-][ZnCl_2]}{k_{-1} + k_2[ZnCl_2]}$$

Molecular orbital calculations using the Hartree–Fock–Slater Discrete Variational Method (HFS-DVM), reported by Ziegler, Rauk, and Baerends, indicate that the manganese atom in permanganate has a d-orbital population close to Mn^{2+} in which two electrons have been removed from the 4s orbital.[16] Symmetry adapted cluster (SAC) and SAC-configuration interaction (SAC-CI) calculations have indicated a somewhat smaller charge (+1.04) on manganese with, of course, a charge of −2.04 shared by the oxygens.[17] These calculations confirm the assumption that the metal oxos are polarized as indicated in Scheme 4.[15]

Reaction of a Lewis acid, such as zinc chloride, with permanganate must involve formation of a coordinate covalent bond using electrons from the permanganate HOMO, which has primarily oxygen p-orbital character. Assuming, according to the most recent calculations, that the structure of permanganate is best represented by resonance structure **2**, the reaction with zinc chloride could be depicted as in eq 4.[18]



(4)

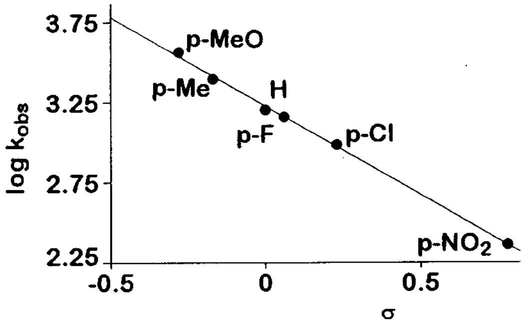


Figure 8. Hammett plot for the oxidation of substituted thioanisoles by permanganate in anhydrous acetone solutions containing zinc chloride as a Lewis acid catalyst. $[ZnCl_2] = 1.16 \times 10^{-3}$ M. $T = 23.0 \pm 0.1$ °C. Slope = −1.11. $r = 0.999$.

The overall effect of this process would be to increase the positive charge on manganese, making it more susceptible to nucleophilic attack by the sulfide. A similar effect may account for the high reactivities of zinc and copper permanganates which have previously been described in the literature.[19]

The validity of the reaction sequence in Scheme 2 was confirmed by the observation that similar kinetics were obtained when mercury(II) chloride was used as the Lewis acid catalyst. Because of the proclivity of mercury(II) to complex sulfides, its use would provide the best opportunity to observe a reaction sequence such as the one depicted in Scheme 1 where the intermediate is a Lewis acid/sulfide complex. The magnitude of the rate constants, k_1 and k_{obs}, were decreased (Table 1) when mercury(II) chloride was used as the catalyst, but an identical rate law was observed.

A quantitative examination of the way in which substituents affect reaction rates can provide an improved understanding of the mechanism of a reaction. For the oxidation of substituted thioanisoles, the Hammett plot is observed to have a negative slope indicating a transition state with decreased electron density. The observation that a better correlation is observed when σ substituent values are used (Figure 8) as compared to σ^+ substituent values (Figure 9) indicates that the transition state is not stabilized by direct resonance with substituents in the *para* position.

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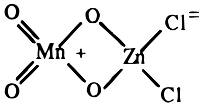
[16] Ziegler, T.; Rauk, A.; Baerends, E. *J. Chem. Phys.* **1976**, *16*, 209.

[17] Jitsuhiro, S.; Nakai, H.; Hada, M.; Nakatsuji, H. *J. Chem. Phys.* **1994**, *101*, 1029.

[18] An alternative representation of the permanganate/zinc chloride complex would be



[19] (a) Wolfe, S.; Ingold, C. F. *J. Am. Chem. Soc.* **1983**, *105*, 7755. (b) Lee, D. G.; Noureldin, N. A. *J. Am. Chem. Soc.* **1983**, *105*, 3188.

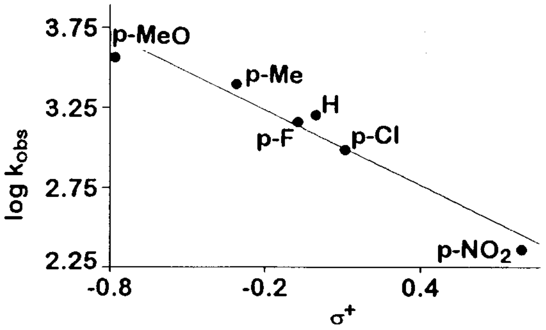
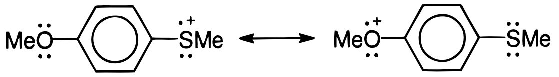
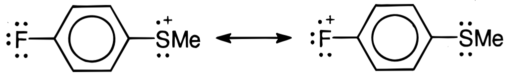


Figure 9. Hammett plot for the oxidation of substituted thoanisoles by permanganate in anhydrous acetone solutions containing zinc chloride as a Lewis acid catalyst. $[\mathrm{ZnCl_2}] = 1.16 \times 10^{-3}$ M. $T = 23.0 \pm 0.1$ °C. Slope = −0.79. $r = 0.973$.

The lack of correlation with σ^+-values indicates that a single electron transfer (SET) is not likely the rate-limiting step of the reaction. The product of such a reaction (eq 5) would be a radical cation which could be stabilized by resonance with substituents such as *p*-methoxy and *p*-fluoro (eqs 6 and 7).[20] Other chemical,[21] electrochemical,[22] and enzyme-catalyzed reactions[23] that involve rate-limiting single electron transfers from sulfur to yield radical cation intermediates, are known to give better Hammett correlations when σ^+ substituent constants are used. If the transition state resembles a radical cation, as predicted by the Hammond postulate[24] for a SET mechanism, a better correlation should have been observed with σ^+-values.

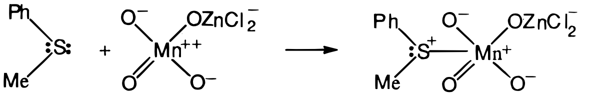
$$\mathrm{Ar\overset{\cdot\cdot}{\underset{\cdot\cdot}{S}}Me} + \mathrm{MnO_4^-} \xrightarrow{-e^-} \mathrm{Ar\overset{\cdot+}{\underset{\cdot\cdot}{S}}Me} + \mathrm{MnO_4^{2-}} \tag{5}$$


(6)


(7)

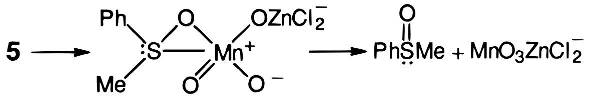
The lack of an observable secondary kinetic isotope effect when the hydrogens on the carbon adjacent to the sulfur are replaced by deuterium is also an indication that the rate-limiting step is not a single electron transfer. The reactions of thioanisole with other oxidants that have been proposed to proceed via SET mechanisms show substantial secondary isotope effects when the methyl hydrogens are replaced by deuterium.[25] In those cases it is believed that the secondary isotope effects are related to the ability of the methyl group to stabilize the radical cation by hyperconjugation.[26,27] This effect, which arises from the possibility for stronger orbital overlap between a sulfur p-orbital containing a single electron and an adjacent C−H σ-bond as compared to an adjacent C−D σ-bond, can be illustrated by use of resonance structures as in Scheme 5. If the reaction is initiated by a single electron transfer, it must therefore occur as a fast step prior to the rate-determining step for the reaction.[28]

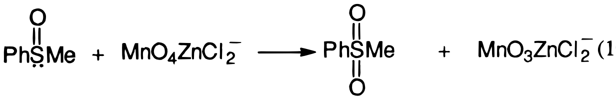
An alternative approach to an understanding of the mechanism for sulfide oxidation by permanganate is to consider the rate-limiting step to be ligand formation; i.e., the attachment of the sulfide to manganese by a coordinate covalent bond as in eq 8.[29]


(8)

5

The product of this reaction, **5**, would be the intermediate that most closely resembles the transition state. The development of a positive charge on sulfur as the reaction progresses is consistent with the observed negative slope of the Hammett plot (Figure 8), and the inability of this intermediate to delocalize the charge via resonance with substituents on the aromatic ring agrees with the observation that a better correlation is observed with σ-values as compared to σ^+-values (Figures 8 and 9). This intermediate, **5**, would also be analogous to the agostic complex that has recently been proposed as an intermediate in the reaction of hydrocarbons with permanganate.[30] The absence of a secondary isotope effect when the methyl hydrogens are replaced by deuterium is also consistent with this intermediate because hyperconjugation of the type depicted in Scheme 5 would not be possible. This intermediate could also be easily converted into products as indicated in eqs 9−11.


(9)



$$\mathrm{Ph\overset{\overset{O}{\|}}{\underset{\cdot\cdot}{S}}Me} + \mathrm{MnO_4ZnCl_2^-} \longrightarrow \mathrm{Ph\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}Me} + \mathrm{MnO_3ZnCl_2^-} \tag{10}$$

$$\mathrm{MnO_3ZnCl_2^-} \xrightarrow{\text{Acetone}} \mathrm{MnO_2ZnCl_2} \tag{11}$$

It has been assumed, as proposed in the literature,[29a] that the first stable product of this reaction sequence is a sulfoxide which would subsequently be oxidized to a sulfone, the observed product of the reaction. Perman-

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[20] Aplin, J. T.; Bauld, N. L. *J. Chem. Soc., Perkin Trans. 2* **1997**, 853.

[21] Ganesan, T. K.; Rajagopal, S.; Bharathy, J. B.; Sheriff, A. I. M. *J. Org. Chem.* **1998**, *63*. 21.

[22] (a) Bauld, N. L.; Aplin, J. T.; Yueh, W.; Loinaz, A. *J. Am. Chem. Soc.* **1997**, *119*, 11381. (b) Bauld, N. L.; Aplin, J. T.; Yueh, W.; Endo, S.; Loving, A. *J. Phys. Org. Chem.* **1998**, *11*, 157.

[23] Oae, S.; Watanbe, Y.; Fujimori, K. *Tetrahedron Lett.* **1982**, *23*, 1189.

[24] March, J. *Advanced Organic Chemistry*, 4th ed.; Wiley: New York, 1992; p 215.

[25] Acquayne, J. H.; Muller, J. G.; Takeuchi, K. J. *Inorg. Chem.* **1993**, *32*, 160.

[26] Carroll, F. A. *Perspectives on Structure and Mechanism in Organic Chemistry*; Brooks/Cole: Toronto, 1998; p 262.

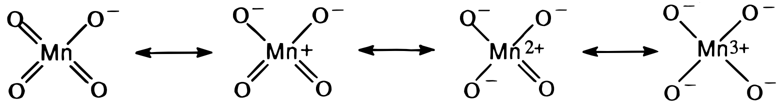
[27] Hess, R. A.; Hengge, A. C.; Cleland, W. W. *J. Am. Chem. Soc.* **1998**, *120*, 2703.

[28] (a) Kochi, J. K. *Angew. Chem., Int. Ed. Engl.* **1988**, *27*, 1227. (b) Bockmann, T. M.; Hubig, S. M.; Kochi, J. K. *J. Am Chem. Soc.* **1998**, *120*, 2826.

[29] (a) Bohra, A.; Sharma, P. K.; Banerji, K. K. *J. Org. Chem.* **1997**, *62*, 3562. (b) Lee, D. G.; Chen, T. *J. Org. Chem.* **1991**, *56*, 5346.

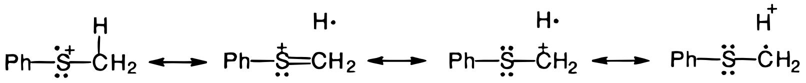
[30] Collman, J. P.; Chien, A. S.; Eberspacher, T. A.; Brauman, J. I. *J. Am. Chem. Soc.* **1998**, *120*, 425.

Scheme 4. Resonance Structures Depicting Permanganate Ion

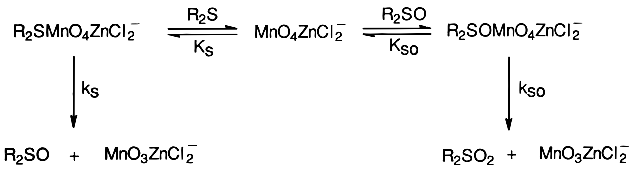


1 **2** **3** **4**

Scheme 5. Resonance Structures That Illustrate the Stabilization by Hyperconjugation of the Radical Cation Formed by a Single Electron Transfer from Thioanisole

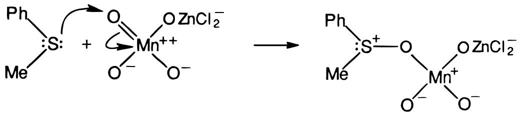


Scheme 6. Competitive Oxidation of Sulfides and Sulfoxides by Permanganate in Anhydrous Acetone Solutions Containing Zinc Chloride as a Catalyst



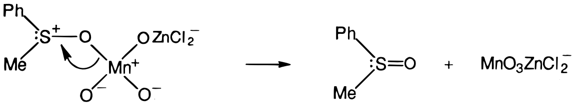
ganate would concurrently be reduced to trioxomanganate(V), MnO_3-, as in eqs 9 and 10. Such manganate(V) compounds are known to be extremely vigorous oxidants which would be rapidly reduced by the solvent, acetone, to give the observed product, manganese dioxide as in eq 11.[14] Since the product of the initial reaction, a sulfoxide, would be readily oxidized to the corresponding sulfone,[10] it is likely that some of the permanganate would be reduced by secondary oxidation, as in eq 10. However, the use of initial reaction rates ensures that the measured rate constants are for the reaction depicted in eq 8; no sulfoxide is present initially.

It has also been suggested that the initial interaction of the oxidant and reductant could actually be between sulfur and one of the permanganate oxygens.[31] Nucleophilic attack at an electron rich oxygen may seem improbable; however, electrostatic repulsions between two electron-rich atoms could be reduced if electrons were simultaneously transferred from the oxo into vacant manganese d-orbitals as depicted in eq 12. Although this process has been referred to in the literature as an "S_N2 type mechanism", it is more properly viewed as a remote nucleophilic attack on the metal. The resulting intermediate, **6**, would be easily converted into products as indicated in eq 13. A similar mechanism has been proposed for the oxidation of sulfides by ruthenium(IV) oxo complexes.[32]



6

(12)



6

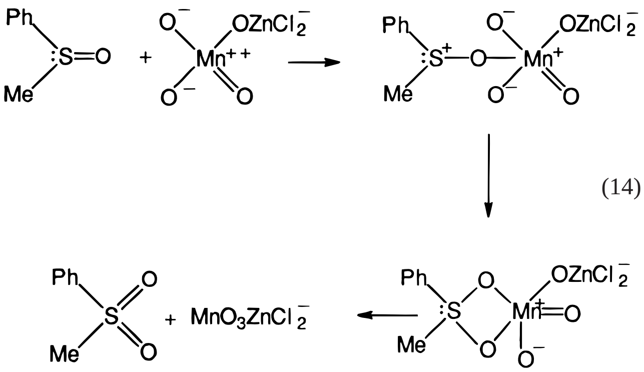
(13)

Both of the mechanisms depicted in eqs 8 and 12 are consistent with the Hammett analysis which indicates that the electron density at sulfur has decreased in the transition state and with the experimentally determined rate law. Both mechanisms also accommodate product formation directly. On the basis of the available kinetic data it is, therefore, not possible to distinguish between them.

Holm[33] has suggested a mechanism similar to the one depicted in eq 8 for the transfer of oxygen from a molybdenum(VI) oxo complex to triethyl- and triphenylphosphine. The reaction is described as a nucleophilic attack on the Mo=O bond by interaction with its vacant π^* orbital.[34] Since π^* orbitals have the same shape and signs as the permanagnate 2e molecular orbital, the reactions appear to be mechanistically similar.

The asymmetric oxidation of thioanisole,[35] and other sulfides, by (oxo)(salen)manganese(V) chloride is also consistent with the idea that attack by a nucleophile on manganese oxos involves some interaction with the metal center. It is not likely that direct interaction of a sulfur p-orbital and an orbital localized on oxygen would lead to an enantioselective reaction.

Since a similar rate law and substitutent effects are observed it seems reasonable to expect that a similar mechanism must pertain for the oxidation of sulfoxides. Because the sulfur−oxygen double bond is quite polar ($R_2S{=}O \leftrightarrow R_2S^+{-}O{-}$), a mechanism analogous to eq 8 would likely occur with ligand formation through oxygen rather than sulfur as in eq 14.



(14)

A mechanism analogous to the one proposed in eq 12 would, however, likely involve an initial interaction

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[31] Banerji, K. K. *Tetrahedron* **1988**, *44*, 2969.

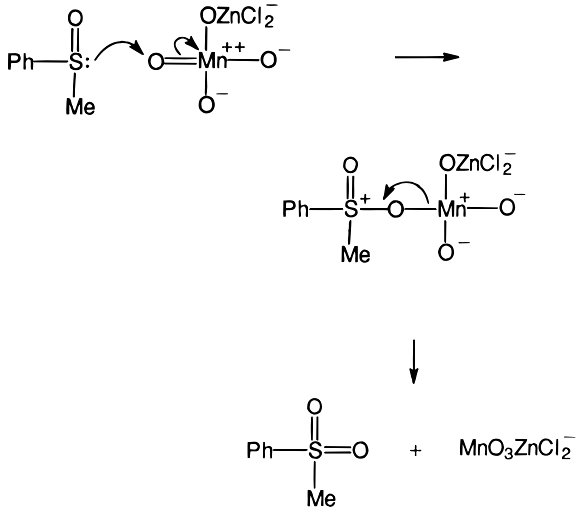
[32] Roecher, L.; Dobson, J. C.; Vining, W. J.; Meyer, T. J. *Inorg. Chem.* **1987**, *26*, 779.

[33] (a) Schultz, B. E.; Gheller, S. F.; Muetterties, M. C.; Soctt, M. J.; Holm, R. H. *J. Am. Chem. Soc.* **1993**, *115*, 2714. (b) Holm, R. H. *Chem. Rev.* **1987**, *87*, 1401.

[34] Holm, R. H.; Berg, J. M. *Acc. Chem. Res.* **1986**, *19*, 363.

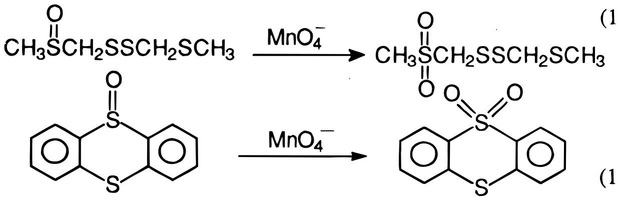
[35] (a) Palucki, M.; Hanson, P.; Jacobsen, E. N. *Tetrahedron Lett.* **1992**, *33*, 7111. (b) Chellamani, A.; Alhaji, N. I.; Rajagopal, S.; Sevvel, R.; Srinivasan, C. *Tetrahedron* **1995**, *51*, 12677.

between sulfur and a manganese oxo as in eq 15.



(15)

The negative Hammett ρ value observed for the oxidation of sulfoxides is consistent with both of the mechanisms presented in eqs 14 and 15, as is the observation that a slightly better correlation is obtained with σ rather than σ^+ values; i.e., $\rho = -2.18$, $r = 0.97$ and $\rho^+ = -0.85$, $r = 0.96$. A distinction between these two mechanisms on the basis of kinetics alone is therefore not possible. However, evidence from product studies can be of assistance since it is known from competitive studies that permanganate oxidizes sulfoxides to sulfones in the presence of sulfides[1a,10] as indicated by the examples in eqs 16 and 17. Such preferential oxidations, despite the fact that sulfoxides react slower than sulfides when individual rates are measured, suggests a mechanism in which the sulfoxide undergoes ligation more completely, but converts to products more slowly. With reference to Scheme 6, K_{so} must be much larger than K_s, while k_s is larger than k_{so}. This description of the mechanism is consistent with all of the experimental evidence including both kinetic and product studies. It is also consistent with the knowledge that sulfoxides are, in general, much better ligands for simple metal cations than sulfides. It therefore appears that the mechanisms presented in eqs 8 and 14 are most consistent with the experimental evidence.



$$CH_3\overset{O}{\overset{\|}{S}}CH_2SCH_2SCH_3 \xrightarrow{MnO_4^-} CH_3\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}CH_2SCH_2SCH_3$$

(16)

(17)

Conclusions

1. A kinetic study has shown conclusively that the catalytic effect of Lewis acids on the oxidation of sulfides and sulfoxides by permanganate is due to complexation of the oxidant with the catalyst prior to reaction with the reductant.
2. Substituent and isotope effect studies indicate that a single electron transfer mechanism is not likely for the oxidation of sulfides by permanganate.
3. Permanganate reacts faster with sulfides than sulfoxides when individual rates are measured, but in competitive reactions, sulfoxides react preferentially.
4. The mechanisms summarized in eqs 8 and 14 and in Scheme 6 are consistent with all of the available experimental evidence.

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