Mechanistic Studies of O_2-Based Sulfoxidations Catalyzed by NO_x/Br Systems

Zhen Luo, Yurii V. Geletii, Daniel A. Hillesheim, Yimin Wang, and Craig L. Hill*

Department of Chemistry, Emory University, 1515 Dickey Dr., Atlanta, Georgia 30322, United States

Ⓢ *Supporting Information*

ABSTRACT: Many studies have noted that mixtures of nitrogen oxides (or oxyanions; $NO_x = NO_3^-$, NO_2^-, NO^-, etc.) and halide (particularly bromine) with or without a transition metal salt (most frequently, copper and iron complexes) are fast and highly selective catalysts for aerobic sulfoxidation and decontamination of many oxidizable toxic compounds. However, the lack of extensive and rigorous kinetic information has kept unequivocal mechanistic inferences on these important catalyst systems to a minimum. Through a series of spectroscopic and kinetics studies on NO_x/bromine systems in acetonitrile, the following clarifying technical points have been established: (1) bromine is the reactive species in these oxidations; (2) nitrosonium and hypobromite are not kinetically important, (3) tribromide (Br_3^-) functions as a stable reservoir for Br_2; (4) salts of Br_3^- and nitrate are stable, storable decontamination catalyst precursors; and (5) most importantly, the multiple complex roles of water (required for catalysts yet inhibitory at most concentrations) have been clarified.

KEYWORDS: air-based sulfoxidation, bromine, tribromide, nitrate, mechanism
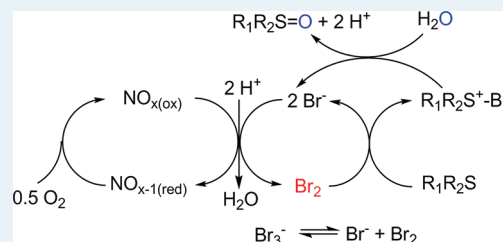
INTRODUCTION

The development of new catalysts for the selective O_2/air-based oxidation of organic compounds is a general goal in green synthetic organic chemistry and decontamination technologies.[1−10] Selective O_2/air-based oxidation of sulfur compounds is highly important in both areas.[2,3,6,8−15] The selective transformation of sulfides to sulfoxides (no production of sulfone even at high conversion of reactant) continues to be a major goal in organic synthesis, and the facile aerobic oxidative removal of H_2S and other oxidizable toxic industrial chemicals (TICs) as well as the chemical warfare agents, mustard (HD), $(ClCH_2CH_2)_2S$, and VX, $(CH_3CH_2O)P(O)(CH_3)$-$SCH_2CH_2N(2-C_3H_7)_2$, continues to be the subject of considerable ongoing research. A host of catalysts have been reported for O_2-based oxidations of sulfur compounds. Sulfoxidation catalysts alone include those based on the 3d metals: manganese,[16] iron,[9,17,18] cobalt,[19] and copper;[16,17] 4d metals: palladium,[20] ruthenium;[21] and 5d metals: rhenium[22] and gold.[8,23] Different families of polyoxometalates (POMs)[12−15,24−35] and combinations of nitrogen oxides (neutral and anionic),[3,36−40] henceforth "NO_x", are also highly effective catalysts.

While product selectivities are obviously crucial in synthetic organic transformations, selectivities, high rates, and conversions are vital for effective decontamination. Indeed, the best catalytic purification systems for indoor air, water, and most human environments ideally would facilitate rapid oxidative removal of the toxic substance(s) using only ambient air. An examination of the literature from the vantage of turnover rates, as well as selectivity, shows that the NO_x/halide catalyst systems are among the fastest. Apart from some seminal work by Bosch and Kochi on the NO^+/NO_2 (halide-free)[3,36] system there is

almost no thorough kinetic and mechanistic information on any NO_x/halide catalyzed O_2-based oxidation reaction. We report here the first systematic kinetic and mechanistic study of the highly selective NO_x/halide catalysts for air-based oxidations, using sulfoxidation of the mustard simulant, 2-chloroethyl ethyl sulfide (CEES), as an exemplary reaction. The reversible inhibition of catalytic oxidation in such systems by water, and other mechanistic features are explained.

EXPERIMENTAL SECTION

Materials. $Na_9[A-\alpha-PW_9O_{34}]$ was prepared by the literature method,[41] and its purity was confirmed by FT-IR. The other reagents including tetra-*n*-butylammonium nitrate (TBANO$_3$), TBANO$_2$, NOBF$_4$, Br$_2$, TBABr, TBABr$_3$, $(CH_2)_6N_4HBr_3$, HClO$_4$, acetonitrile, 1,3 dichlorobenzene, bromocyclohexane, and 2-chloroethyl ethyl sulfide (CEES) were purchased from Aldrich and used without further purification. Dry acetonitrile was purchased from Aldrich; it was further purified by passage through a freshly activated alumina column and maintained under Ar.

Instruments. The IR-spectra were recorded on a Nicolet 510 FT-IR spectrometer. The electronic absorption spectra were taken on an Agilent 8453 UV−visible spectrometer equipped with an Agilent 89090A temperature controller. Oxidation products were identified by gas chromatography—mass spectrometry

(GC/MS; Hewlett-Packard 5890 series II gas chromatograph connected to a Hewlett-Packard 5971 mass selective detector) and quantified by gas chromatography (GC; Hewlett-Packard 5890 series gas chromatograph equipped with a flame ionization detector, a 5% phenyl methyl silicone capillary column, and a Hewlett-Packard 3390A series integrator).

Preparation of the Catalytic Solutions. *1. Preparation of NO_x/Br Catalytic Solutions.* In general, two stock solutions of NO_x and Br were mixed in CH_3CN and stirred until complete dissolution of the reagents. In an exemplary case, 0.16 g (0.5 mmol) of TBABr and 0.15 g (0.5 mmol) of TBANO_3 were mixed in 5 mL of CH_3CN. After the reagents were completely dissolved, the solution was transferred to a 10 mL volumetric flask and diluted by CH_3CN to a mark. This freshly prepared catalytic solution was colorless and used immediately for kinetic studies. Usually the reaction kinetics with such freshly prepared solutions showed distinctive induction periods. If the solution was aged (from hours to days) it became yellow and the induction periods disappeared. Henceforth, this latter solution is referred to as "aged."

Alternatively, the catalytic solution was prepared as follows. A 0.16 g (0.5 mmol) portion of TBABr and 0.145 g (0.45 mmol) of TBANO_3 were dissolved in 5 mL of CH_3CN. To this solution was added 14 mg (0.05 mmol) of TBANO_2. This mixture was transferred to a 10 mL volumetric flask, and 25 μL of aqueous 4 M $HClO_4$ (0.1 mmol) was added. The resulting solution immediately turned yellow and had the same activity as the aged solutions described above. No induction period was observed in this case.

2. Modified Preparation of $TBA_3H_2[Fe(NO_3)PW_{11}O_{39}]$. The preparation of $TBA_3H_2[Fe(NO_3)PW_{11}O_{39}]$ is similar to the method used by Okun et al.,[18] except that TBANO_3 was used instead of TBABr to obtain the precursor, $[\alpha\text{-}PW_{11}O_{39}]^{7-}$. A solid sample of $Na_7[\alpha\text{-}PW_{11}O_{39}]\cdot74H_2O$ (1 g, ca. 0.34 mmol) was dissolved in 10 mL of deionized water and mixed with 8 mL of an aqueous solution containing TBANO_3 (3.8 g, ca. 12.4 mmol). The mixture was stirred for 30 min at room temperature. The resulting precipitate was separated by filtration, redissolved in 15 mL of CH_3CN, then solid $Fe(NO_3)_3\cdot9H_2O$ (0.16 g, 0.4 mmol) was added under vigorous agitation. A dark reddish-brown solution and a small amount of oily precipitate were formed. The solution was separated from the oily product, filtered through the fine filter paper (Fisher brand P2) to remove the suspended precipitate, stirred for 1.5 h at ambient temperature, and filtered again through the P2 paper. The filtrate was left in air and light brown crystals (0.5 g; 39% yield; henceforth denoted as "**1**") were produced after 2 days. These crystals were dissolved in CH_3CN to obtain a stock solution of the catalyst, **1**.

General Procedure of Catalytic Oxidation of CEES. In a typical CEES sulfoxidation reaction, 1 mL of catalytic NO_x/Br solution, (e.g., 0.05 mmol NO_3^- and 0.05 mmol Br^-) was diluted with 1 mL of acetonitrile in a 15 mL vial. To this mixture, 25 μL of 1,3-dichlorobenzene or bromocyclohexane (internal standards for GC) was added, and then the vial was sealed with a PTFE septum. After stirring for 1 min, 25 μL of CEES was injected by syringe into a vial. To avoid depletion of O_2 in the headspace, access to air during the experiment was provided through a needle in the septum on the cap. To accelerate the reaction, the temperature was elevated to 70 °C. The consumption of CEES was analyzed by GC every 15—30 min, and typically the reaction was monitored for 10—20 h.
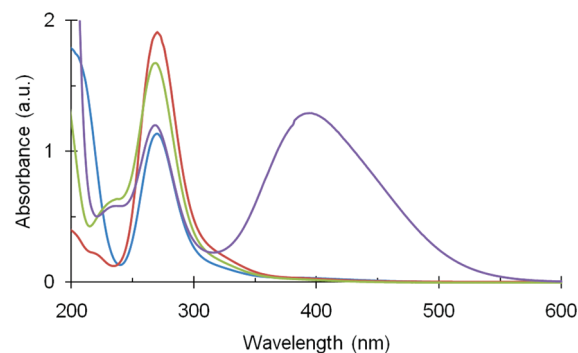

Figure 1. Absorption spectra of different NO_x/Br systems in acetonitrile. From top to bottom at the 270 nm peak: red curve, TBABr_3; green curve, TBABr/NOBF_6; purple curve, Br_2; blue curve, TBABr/TBANO_2.

■ RESULTS AND DISCUSSION

**Tribromide Is a Key Intermediate in NO_x/Br-Catalyzed Oxidations and Several Other Bromide Species Are Not.** When both TBANO_3 and TBABr are dissolved in acetonitrile solution no significant changes in color are observed in the first few hours. However, the addition of very small amounts of strong acid ($HClO_4$ or *p*-toluenesulfonic acid (TsOH)) instantaneously turns the colorless solution to yellow. If TBANO_2 or NOBF_4 are used instead of TBANO_3, the color change is rapid even without acid. UV—vis spectra of the acetonitrile solution of these NO_x/Br systems show that the yellow color is due to a strong absorption peak near 270 nm which tails to λ > 400 nm. The absorption maximum at 270 nm exactly matches the spectrum of Br_3^- in CH_3CN. Figure 1 compares the spectra of authentic TBABr_3 and other NO_x and Br containing systems. The red curve is the spectrum of authentic TBABr_3, which has the strong characteristic peak at 270 nm assigned to Br_3^-. In the gas phase Br_2 has a single peak near 420 nm. When Br_2 dissolves in CH_3CN a broad peak around 400 nm and two additional peaks in the UV-range are observed (Figure 1, purple curve): the additional peak at <200 nm is assigned to Br^- and the other peak at 270 nm is assigned to Br_3^-. These two latter species are formed in small amounts via a disproportionation of Br_2 in the presence of trace amounts of H_2O in CH_3CN (eqs 1—2). (The K_a of hypobromous acid in water is 2.2×10^{-9}).[42] A secondary reaction, where hypobromite (BrO^-) spontaneously disproportionates to Br^- and bromate (BrO_3^-), takes place rapidly in water at 20 °C and slowly at 0 °C. Bromates, including bromic acid, are relatively unstable and undergo further decomposition to Br^- and O_2.

The equilibrium in eq 2 is significantly shifted to the right side, $K_2 = 9 \times 10^6$ M^{-1} in acetonitrile.[43] The Br_3^- at low concentration is seen as a intense peak because of its very high extinction coefficient ($\varepsilon = 3.8 \times 10^4$ M^{-1} cm^{-1})[44,45] which is 2 orders of magnitude larger than that of Br_2. Both the NO_x/Br systems prepared from TBANO_2/TBABr (blue curve) and from NOBF_4/TBABr (green curve) have the same peak at 270 nm, indicating the formation of Br_3^-.

$$Br_2 + H_2O \rightarrow Br^- + BrOH + H^+ \qquad (1)$$

$$Br_2 + Br^- \xrightleftharpoons{K_2} Br_3^- \qquad (2)$$

Previous work performed by Lengyel et al.[46] also supports the generation of Br_2/Br_3^- in the NO_x/Br system. They studied the

oxidation of bromide by nitric acid to bromine in aqueous solution at room temperature in the absence of air/oxygen, eqs 2–3, and found that the experimental value of $K_3 = 1.6 \times 10^{-6} \text{ M}^{-4}$.

$$NO_3^- + 2Br^- + 3H^+ \overset{K_3}{\rightleftharpoons} Br_2 + HNO_3 + H_2O \qquad (3)$$
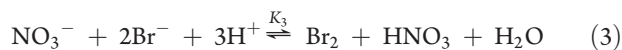
The reaction appeared to be autocatalytic. A clearly observed induction period was strongly dependent on the initial concentration of NO_2^- (HNO_2 in acidic conditions, $pK_a = 3.35$).[42] On the basis of eq 3, an increase in proton concentration should shift the equilibrium in favor of Br_2 formation, while an increase in H_2O concentration should act in the opposite direction. The HNO_2 generated in situ catalyzes the formation of Br_2,[46] which makes the reaction autocatalytic. It should be noted that in aerobic sulfoxidation catalyzed by the freshly prepared TBA-NO_3/TBABr system (no acid added), an induction period is also observed (see Supporting Information, Figure S1). Additionally, in our earlier work we had repeatedly noted an inhibition of intermediate yellow color formation by the presence of water. Both these observations are consistent with the formation of Br_2/Br_3^- in the first step of the catalytic reaction.

The stoichiometry for aerobic sulfoxidation in previously studied homogeneous catalytic systems and in the NO_x/Br catalyst examined here is eq 4. Sulfide, and sulfoxide were quantified by GC methods (see Experimental Section). This dioxygenase stoichiometry is more attractive because it is 100% atom efficient in oxygen, but more importantly it does not generate potentially deleterious byproduct (see Supporting Information, Figure S2) as in monooxygenase and some other catalyzed O_2-based oxidation reactions. From the decontamination perspective, formation of the sulfoxide of mustard is greatly preferred relative to overoxidation to the sulfone because the latter is far more toxic than the former.

$$R_2S + 0.5O_2 \xrightarrow[\text{ambient conditions}]{Br^-/NO_3^-} R_2SO \qquad (4)$$
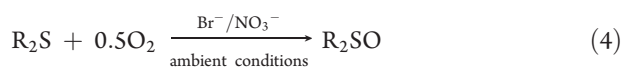
Previous research in our group led to identification of the monoiron-substituted POM, $TBA_3H_2[Fe(NO_3)PW_{11}O_{39}]$ (**1**), as a highly efficient catalyst for eq 4 without addition of any bromide.[18] However, the synthesis of catalytically active solutions in these studies[18] used a large excess of TBABr to prepare the precursor, $TBA_{7-x}Na_xPW_{11}O_{39}$. Without thorough purification, some Br^- could be retained in the final catalyst as an impurity. Therefore, in this work we replaced TBABr with $TBANO_3$ to avoid any contamination with Br^- species (see Experimental Section, preparation 3) in the synthesis of **1**. Importantly, no sulfoxidation of the mustard simulant, CEES) in the presence of **1** prepared in this manner (no bromine source) is observed. This modified (bromine-free) system exhibits no catalytic activity for aerobic sulfoxidation in general (see details in Supporting Information, Figure S3). However, if some TBABr is added to complex **1**, such a system becomes catalytically active again. If $TBANO_3$ is also added the **1**/TBABr solution, the catalytic activity is increased further. This collective evidence strongly suggests that the most active sulfoxidation catalyst contains both some form (oxide) of nitrogen and bromine, that bromine is the likely the substrate (sulfide) oxidizing specie, and that tribromide constitutes a bromine storage reservoir. This evidence led to a series of experiments we now describe.

Although the experiments above afforded fairly compelling evidence that Br_2/Br_3^- is both abundant and reactive under turnover conditions in systems containing NO_x and bromide/bromine, the specific NO_x species involved in these aerobic oxidations was
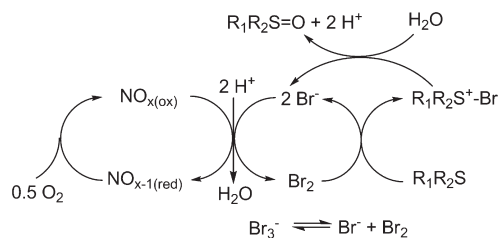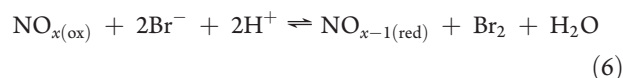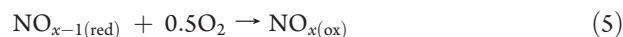

Figure 2. Proposed mechanism for the selective aerobic oxidation of sulfide to sulfoxide catalyzed by the NO_x/Br systems.

far from clear at the outset of this study. Since nitrosonium cation is present in acidic NO_x mixtures, we initially assessed the catalytic aerobic sulfoxidation activity of NOBr, a very volatile yellow compound which readily dissociates into NO and bromine atom under ambient conditions.[47] NOBr is the evident product when $NOBF_4$ and TBABr are mixed in CH_3CN. However, isolation of the nonvolatile catalytically active yellow solid, largely the Br_3^- salt, cast doubt on the intermediacy of volatile and unstable NOBr in these reactions. Addition of nitrosonium, nitrite, or nitrate to the bromide containing solutions significantly increases the rates for aerobic sulfoxidation. This follows from the O_2-based oxidation of Br^- to Br_2 catalyzed by nitrite or nitrate.

Important Elementary Processes in NO_x/Br-Catalyzed Oxidations. One group that has published several recent papers on NO_x/Br/TM systems, where TM = Fe(III) or Cu(II), as aerobic oxidation catalysts is that of Rossi, Martin and co-workers.[9,10,48–50] They proposed that the selective aerobic oxidation of sulfides to sulfoxides catalyzed by Fe and/or Cu nitrate and bromide salts takes place at the metallic center with nitrate/nitrite and bromide/bromine redox cycles as oxidant and mediator, respectively. However, no kinetics or other analyses that proved these mechanistic proposals were provided. Both their group and ours noted that the presence of transition metals (such as Fe or Cu) significantly affects the reaction, and therefore the reaction kinetics could well be too complicated for interpretation. In addition, the spectra of Br_3^- and Cu and/or Fe bromides are overlapping, which prohibits the quantification of Br_3^-. Tribromide salts are stable (the tetrabutylammonium salt is commercially available, CAS Number: 38932-80-8). Given this point, we can combine all the data from the Martin and Rossi studies and our data, and propose a reaction mechanism without involving transition metals (Figure 2) that accounts for much of the reactivity seen in the many NO_x/bromide catalytic systems with and without a transition metal complex present.

To elucidate the mechanism of the sulfoxidation (eq 4) catalyzed by NO_x/Br, the known related or component processes must be considered. In addition to the nitrate-bromide/nitrite-bromine equilibrium, eq 3, sulfides are oxidized by Br_2, which is not one elementary process but several, vide infra. Dioxygen is involved in oxidation of reduced $NO_{x-1(red)}$.

Earlier Lengyel et al.[46] showed that several different reduced forms of NO_x (NO, NO_2, NO_2^- and BrNO) are involved as intermediates in eq 3. Therefore, the formation of Br_2 and regeneration of oxidant can be represented by eqs 5 and 6.
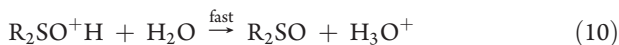
$$NO_{x-1(red)} + 0.5O_2 \rightarrow NO_{x(ox)} \qquad (5)$$

$$NO_{x(ox)} + 2Br^- + 2H^+ \rightleftharpoons NO_{x-1(red)} + Br_2 + H_2O \qquad (6)$$

where $NO_{x(ox)}$ and $NO_{x-1(red)}$ represent the oxidized and reduced forms of NO_x, respectively.

The regeneration of Br_2 is further simplified by the reaction in eq 7. In other words, the regeneration of Br_2 is a catalytic oxidation of Br^- by O_2 catalyzed by NO_x.

$$2Br^- + 0.5O_2 + 2H^+ \xrightarrow{NO_x} Br_2 + H_2O \tag{7}$$
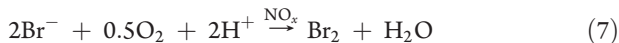
In view of the fact that in acetonitrile, K_2 (eq 2) is much higher than that of in water $(9 \times 10^6 \, M^{-1} \text{ vs } 17 \, M^{-1})$,[43,51] Br_3^- is likely to be the main species present in the Br_2/Br_3^- equilibrium under catalytic conditions.

Thioether Sulfoxidation by Br₂. The oxidation of organic sulfides to sulfoxides with molecular halogens is a well-known process,[52] and is likely to proceed via eqs 8–10.[53–56]

$$R_2S + Br_2 \overset{K_{8A}}{\rightleftharpoons} R_2SBr_2 \overset{k_8}{\underset{k_{-8}}{\rightleftharpoons}} R_2S^+Br + Br^- \tag{8}$$

$$R_2S^+Br + H_2O \xrightarrow[\text{slow}]{k_9} R_2SO^+H + H^+ + Br^- \tag{9}$$

$$R_2SO^+H + H_2O \xrightarrow{\text{fast}} R_2SO + H_3O^+ \tag{10}$$

Mixing sulfide with bromine initially forms a sulfide-bromine complex, eq 8. The sulfide in the complex is oxidized to a sulfonium cation by an intramolecular electron transfer from the sulfur center to a bromine atom. The resulting sulfonium-bromine complex donates another electron to a second bromide atom in the reaction assisted by water, eq 9. This reaction is likely to be the slowest step. Subsequent hydrolysis in eq 10 produces a sulfoxide product. Two bromides are formed in overall reaction.

The reaction rate law can be deduced by applying the approximation of steady-state concentrations with respect to R_2S^+Br (eqs 11).

$$-\frac{d[R_2S]}{dt} = k_9 K_{8A} \frac{k_8[H_2O]}{k_{-8}[Br^-] + k_8[H_2O]}[R_2S][Br_2] \tag{11}$$

Because eq 9 is the slowest step, it is reasonable to assume that $k_8[H_2O] \ll k_{-8}[Br^-]$. Then eq 11 simplifies to eq 12.

$$-\frac{d[R_2S]}{dt} \approx k_9 K_{8A} K_8 \frac{[H_2O][Br_2]}{[Br^-]}[R_2S] \tag{12}$$

Since Br_3^- is the main form of bromine in the Br_2/Br_3^- equilibrium, and because Br_3^- has the highest extinction coefficient, we have chosen the absorption of Br_3^- to follow the reaction kinetics.

A reaction mass balance with respect to all forms of Br is given in eq 13.

$$[Br^-] = [Br^-]_o - 3[Br_3^-] - 2[Br_3^-]/(K_2[Br^-]) \tag{13}$$

Under typical experimental conditions for an aged system, a considerable amount of tribromide forms which results in too high an initial absorbance at $\lambda = 270$ nm, $A_{270} > 2$. For this reason, we withdrew small aliquots from the reaction solution at desired time intervals, diluted it fifty-fold with CH_3CN, and then immediately measured A_{270} of these diluted solutions. Subsequently in these diluted solutions the absorbance at 270 nm slowly increased because of an oxidation of Br^- by dioxygen to Br_2/Br_3^-. Under the experimental conditions in Figure 3, the A_{270}, after an initial quick drop, reaches a plateau with $A_{270} \approx 0.04$ ($l = 1$ cm). After taking into account the dilution factor the A_{270} value at
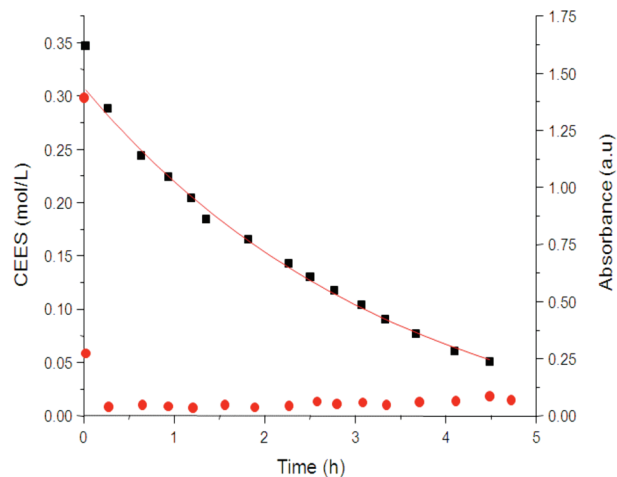

Figure 3. Kinetics (black squares) and fittings (red line) of CEES oxidation by air catalyzed by the $TBANO_3/TBABr$ system, and the time profile of A_{270} (red circles) determined after 50-fold dilution. The reaction system: $[CEES]_o = 0.35$ M, $[TBANO_3] = 0.018$ M, $[TBABr] = 0.11$ M, $[TsOH] = 0.018$ M, $T = 70\,°C$.

initial 110 mM $[Br^-]_o$ corresponds to $[Br_3^-] \approx 0.05$ mM and $[Br_2] \approx 2.2 \times 10^{-7}$ mM under turnover conditions. The total amount of bromine and tribromide is only about 0.14% of $[Br^-]_o$. Consequently, the steady-state concentration of $[Br^-]_t$ remains almost constant and nearly equals to $[Br^-]_o$ during the catalytic reaction. Accordingly, combining eq 2 and eq 12 results eq 14.

$$-\frac{d[R_2S]}{dt} = \frac{k_9 K_{8A} K_8}{K_2} \frac{[H_2O]}{[Br^-]_o^2}[R_2S][Br_3^-] \tag{14}$$

Under steady-state conditions, the $[Br_2]/[Br_3^-]$ ratio is constant and equal to $1/(K_2[Br^-]_o)$. Below we describe the time profile of both $[R_2S]$ and $[Br_3^-]$.

Figure 3 shows the kinetics of CEES consumption and the time profile of $[Br_3^-]$ in sulfoxidation catalyzed by aged NO_x/Br solution. In the first few minutes the concentration of Br_3^- drops dramatically likely because of the reaction of Br_2 with CEES to form R_2S^+Br as an intermediate in eq 8. This reaction proceeds with a typical time $\sim 10^2$ s. After this initial period, the Br_3^- concentration then remains low and nearly unchanged throughout the rest of the reaction. After reaching a constant Br_3^- concentration, the consumption of CEES is exponential in agreement with eq 14.

The above analysis does not take into account the regeneration of Br_2/Br_3^- by O_2. It only takes into account the experimental observation that $[Br_3^-]$ remains constant with time under given experimental conditions. The overall sulfoxidation reaction relies on the regeneration of Br_2/Br_3^- which might depend on $[O_2]$ and $[NO_x]$. As a result, the above analysis is not complete and does not describe the complete rate law.

Regeneration of Br₂/Br₃⁻ by O₂. Unlike the well-known sulfide oxidation by halogens,[52] the kinetics of the regeneration of Br_2/Br_3^- by O_2 catalyzed by NO_x is much more complex and cannot be described by eqs 5–6. Therefore, we attempted to find experimentally the rate law for regeneration of Br_2/Br_3^-. As mentioned above, after a 50-fold dilution of the reaction mixture by CH_3CN it was observed that A_{270} (which is proportional to

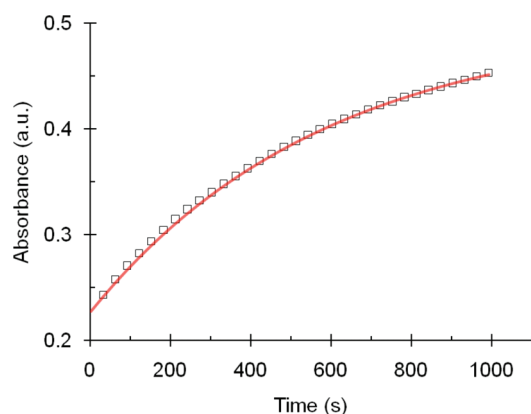**Figure 4.** Time profile of A_{270} (or $[Br_3^-]$) in the presence of air after 50-fold dilution of the reaction mixture. Red curve is calculated using eq 15. Reaction conditions: $[CEES]_o = 108$ mM, $[TBANO_3] = 25$ mM, $[TBABr] = 25$ mM, $[HClO_4] = 5$ mM, $CH_3CN = 2.0$ mL, $T = 70\,°C$, $l = 1$ cm.

$[Br_3^-]$) slowly increases in time. Figure 4 represents a typical Br_3^- regeneration time profile of A_{270} measured by UV−vis.

As discussed above, the data in Figure 3 reveal that the reaction in eq 8 is not extremely fast: it takes $\sim 10^2$ s to establish the equilibrium. Upon dilution the concentrations of $[R_2S]$ and $[Br_2]$ decrease 50-fold, which results in more than a 10^3-fold drop in the rate of eq 8. For this reason, a dilution actually "freezes" the consumption of Br_3^-. The increase of A_{270} with time represents the kinetics of Br_2/Br_3^- regeneration by O_2. From data in Figure 4 we attempted to determine a reaction rate law that describes these kinetics. For this purpose we plotted the instant reaction rate at a given time t as a function of Br_3^- concentration at this time. As seen from Supporting Information, Figure S4, the simplest rate law for Br_3^- formation in Figure 4 might be written as eq 15.

$$\frac{d[Br_3^-]}{dt} = R^0_{15} - R^1_{15}[Br_3^-] \tag{15}$$

where $R^0_{15} = 2.3 \times 10^{-8}$ M/s, $R^1_{15} = 1.7 \times 10^{-3}$ s^{-1}.

The independence of the rate on $[Br_3^-]$ is consistent with the rate limiting step being the reaction of NO_{x-1}(red) with O_2 to regenerate Br_3^- (one step in the complex reaction in eq 7). At the same time the reaction is inhibited by the product because Br_2 binds Br^-, which is involved in eq 7. As a result the reaction does not proceed to 100% conversion but reaches an equilibrium similar to that in eq 3.

Thus, although the mechanism of Br_2/Br_3^- regeneration is complex, the experimental rate law is relatively simple and describes the experimental kinetic curve well. This reaction rate law is obtained at significantly lower concentrations of reactants, but it could well be a plausible rate law under turnover conditions.

Using the data for Br_2/Br_3^- regeneration, eq 15, the equation for CEES consumption can be deduced as follows. At steady-state, the rate of sulfide consumption equals the rate of Br_3^- regeneration, which results in eq 16.

$$[Br_3^-] = \frac{R^0_{15}}{R^1_{15} + \frac{k_9 K_{8A} K_8}{K_2} \frac{[H_2O]}{[Br^-]_o^2}[R_2S]} \tag{16}$$

Set $k_{17} = (k_9 K_{8A} K_8/K_2) \times ([H_2O]/[Br^-]_o^2)$ and combining eqs 16 and 14 affords eq 17.

$$-\frac{d[R_2S]}{dt} = k_{17}[R_2S]\frac{R^0_{15}}{R^1_{15} + k_{17}[R_2S]} \tag{17}$$

When $R^1_{15} \gg k_{17}[R_2S]$, then eq 17 simplifies to eq 18. In this case the rate limiting step is eq 9, and the rate of CEES consumption should be a first-order with respect to $[R_2S]$ as in eq 18.

$$-\frac{d[R_2S]}{dt} = \frac{R^0_{15}k_{17}}{R^1_{15}}[R_2S] \tag{18}$$

On the other hand, if a large amount of water is present in the system, then $k_{17}[R_2S] \gg R^1_{15}$, and eq 17 simplifies to eq 19. The reaction becomes a zero order with respect to $[R_2S]$.

$$-\frac{d[R_2S]}{dt} = R^0_{15} \tag{19}$$

In dry acetonitrile eq 9 is slow, resulting in first order kinetics with respect to $[R_2S]$. However, as the associate constant of Br_3^-, K_2, becomes smaller in water than in acetonitrile[43,51] the addition of water can greatly increase $k_{17} = (k_9 K_{8A} K_8/K_2) \times ([H_2O]/[Br^-]_o^2)$ and makes $k_{17}[R_2S]$ comparable with R^1_{15}. In this case the integration of eq 17 gives the solution in eq 20.

$$\frac{-1}{R^0_{15}}\left\{\frac{R^1_{15}}{k_{17}}\ln\left(\frac{[R_2S]_t}{[R_2S]_0}\right) + [R_2S]_t - [R_2S]_0\right\} = t \tag{20}$$

In the following paragraphs eq 20 will be used to simulate the effect of water on the catalytic sulfoxidation.

Effect of Water on Catalytic Activity. Water is an important factor that greatly affects the activity of the NO_x/Br catalytic systems. Our previous work[57] showed that large amount of H_2O significantly inhibits the activity of the catalyst, and that this inhibition is reversible. However, these highly reproducible observations have not been explained to date.

This paradox is explicable by our mechanism. In accordance with eq 14, water should accelerate the rate-limiting step in eq 9 and thus increase the overall reaction rate. The data acquired by Lengyel and co-workers[46] indicates that water inhibits the autocatalytic reaction in eq 3. In addition, water increases the induction period if added to a freshly prepared system. Water also affects regeneration of the oxidized form of NO_x. Indeed, nearly all rate constants in eq 20 are dependent on water concentration. For that reason, we examined the effect of water on the kinetics of aerobic sulfoxidation catalyzed by the NO_x/Br system.

The slow step in eq 9,[8,58] includes water as a reagent, so the reaction was carried out in a water-free system. The preparation of a water-free NO_x/Br system is similar to that of the aged NO_x/Br catalyst except that freshly dried acetonitrile and concentrated H_2SO_4 were used instead of aqueous $HClO_4$. The effect of water is shown in Figure 5. Clearly, the reaction in the absence of water is much slower. Interestingly, a drop in $[CEES]$ concentration is clearly seen in first ~ 30 min of the reaction both in the absence and in the presence of water. This drop is caused by the stoichiometric reaction of Br_2 (already present in the system) with CEES. After that CEES oxidation in a water-free system proceeds very slowly because the R_2S^+Br intermediate is quite stable in this case. The maximum conversion of CEES in the beginning of the reaction does not exceed the total amount of Br^-, which is a source of Br_2. Indeed, after 24 h the amount of
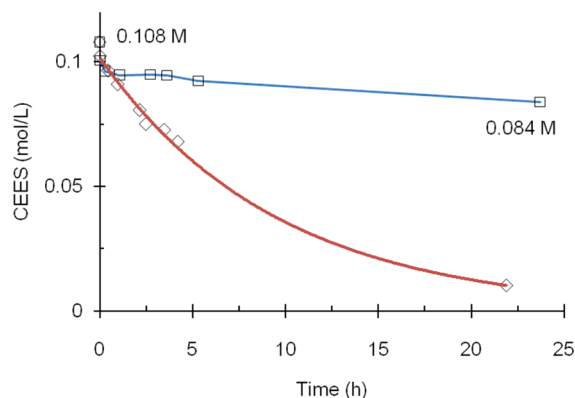
Figure 5. Kinetics of CEES oxidation by air in NO_x/Br system in the absence ($\square$, 2.5 mM of dry H_2SO_4) and in the presence of water ($\diamondsuit$, 5 mM $HClO_4$, and 277 mM H_2O). Conditions: $[CEES]_o = 108$ mM (25 μL), $[TBANO_3] = 25$ mM, $[TBABr] = 25$ mM, $CH_3CN = 2.0$ mL, $T = 70\ °C$.
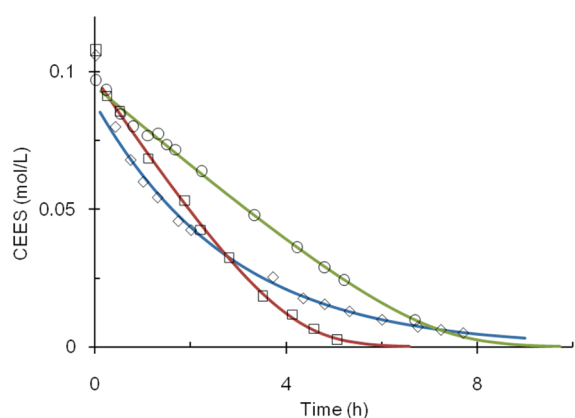

Figure 6. Kinetics of CEES oxidation by air catalyzed by aged NO_x/Br system containing 0.28 ($\diamondsuit$),1.38 M ($\square$) and 2.77 M ($\bigcirc$) of water at 70 °C in acetonitrile. Conditions: $[CEES]_o = 108$ mM, $[TBANO_3] = 25$ mM, $[TBABr] = 25$ mM, $[HClO_4] = 5$ mM. Solid lines are obtained from the least-squares fitting to eq 20.

Table 1. Reaction Rate and Associated Values Obtained by Least-Squares Fitting of Aerobic CEES Oxidation Catalyzed by the NO_x/Br Systems Shown in Figure 6

$[H_2O]$, (M)	0.28	1.38	2.77
R^0_{15}, (s^{-1})	1.0×10^{-4}	8.3×10^{-6}	4.5×10^{-6}
k_{17}/R^1_{15}, $(M^{-1})^a$	1.2	78	1.2×10^2

^{a} The values of k_{17} and R^1_{15} cannot be obtained from the fitting of eq 20.

CEES consumed is equal to the TBABr used in the reaction (0.048 and 0.05 mmol, respectively, Figure 5). In the presence of water the kinetics of CEES consumption is exponential as described above.

Figure 6 shows the strong effects of water on the reaction kinetics: an increase of water concentration results in change in the reaction rate order with respect to [CEES]. At low water content the consumption of CEES is exponential and becomes almost linear with increase of water concentration (Figure 6). Thus the reaction rate order with respect to [CEES] changes

from the first to zero. The least-squares fitting of kinetic curves to eq 20 gave values for k_{17}, R^0_{15}, and R^1_{15} presented in Table 1.

A decrease in R^0_{15} with water addition suggests that the regeneration of Br_2 is inhibited by water. At the same time, k_{17}/R^1_{15} increases with water resulting in an increase the overall reaction rate in agreement with eq 18. The significant changes in k_{17}/R^1_{15} with addition of water is likely to indicate that R^1_{15} is also water dependent.

CONCLUSIONS

(1) Many spectroscopic and kinetics studies indicate that the much studied, fast, and highly selective catalytic systems for O_2-based oxidation of organic substrates (most studied for sulfoxidation) comprising a nitrogen oxide or oxyanion ($NO_x = NO_3^-$, NO_2^-, NO^-, etc.) plus bromide with or without a transition metal involve the generation of tribromide, Br_3^-, a result not noted previously. Hypobromite is not important.

(2) Tribromide salts can be isolated from the reaction mixtures and stored for apparently indefinite periods. Upon dissolution in low-H_2O-content organic media, in the presence of nitrate or other NO_x species, they catalyze rapid and highly selective aerobic sulfoxidation and other oxidation reactions.

(3) Water has been noted to inhibit aerobic oxidations catalyzed by NO_x/halide systems but to do so reversibly. Kinetic and product isolation studies herein provide an explanation for the reversible inhibition of catalytic aerobic sulfoxidation by water, a result long observed but never understood.

(4) The rate limiting step depends on water concentration and changes with reaction conditions.

ASSOCIATED CONTENT

S **Supporting Information.** Additional kinetics information on O_2-based oxidation of a sulfide (CEES) by a range of catalysts under varying conditions. This material is available free of charge via the Internet at http://pubs.acs.org.

AUTHOR INFORMATION

Corresponding Author
*E-mail: chill@emory.edu.

REFERENCES

(1) Mashkina, A. V. *Catal. Rev.* **1990**, *32*, 105–161.
(2) Srivastava, R. S.; Milani, B.; Alessio, E.; Mestroni, G. *Inorg. Chim. Acta* **1992**, *191*, 15–17.
(3) Bosch, E.; Kochi, J. K. *J. Org. Chem.* **1995**, *60*, 3172–3183.
(4) Komatsu, N.; Uda, M.; Suzuki, H. *Chem. Lett.* **1997**, 1229–1230.
(5) Iwahama, T.; Sakaguchi, S.; Ishii, Y. *Tetrahedron Lett.* **1998**, *39*, 9059–9062.
(6) Xu, L.; Boring, E.; Hill, C. *J. Catal.* **2000**, *195*, 394–405.
(7) Shii, Y.; Sakaguchi, S.; Iwahama, T. *Adv. Synth. Catal.* **2001**, *343*, 393–427.

(8) Boring, E.; Geletii, Y. V.; Hill, C. L. *J. Am. Chem. Soc.* **2001**, *123*, 1625–1635.

(9) Martín, S. E.; Rossi, L. I. *Tetrahedron Lett.* **2001**, *42*, 7147–7151.

(10) Rossi, L. I.; Martín, S. E. *Appl. Catal., A* **2003**, *250*, 271–278.

(11) Boring, E.; Geletii, Y.; Hill, C. L. *J. Mol. Catal. A: Chem.* **2001**, *176*, 49–63.

(12) Okun, N. M.; Anderson, T. M.; Hardcastle, K. I.; Hill, C. L. *Inorg. Chem.* **2003**, *42*, 6610–6612.

(13) Okun, N. M.; Anderson, T. M.; Hill, C. L. *J. Am. Chem. Soc.* **2003**, *125*, 3194–3195.

(14) Okun, N. M.; Anderson, T. M.; Hill, C. L. *J. Mol. Catal. A: Chem.* **2003**, *197*, 283–290.

(15) Okun, N. M.; Ritorto, M. D.; Anderson, T. M.; Apkarian, R. P.; Hill, C. L. *Chem. Mater.* **2004**, *16*, 2551–2558.

(16) Song, G. Q.; Wang, F.; Zhang, H.; Lu, X. L.; Wang, C. *Synth. Commun.* **1998**, *28*, 2783–2787.

(17) Firouzabadi, H.; Iranpour, N.; Zolfigol, M. A. *Synth. Commun.* **1998**, *28*, 1179–1187.

(18) Okun, N. M.; Tarr, J. C.; Hilleshiem, D. A.; Zhang, L.; Hardcastle, K. I.; Hill, C. L. *J. Mol. Catal. A: Chem.* **2006**, *246*, 11–17.

(19) Dell'Anna, M. M.; Mastrorilli, P.; Nobile, C. F. *J. Mol. Catal. A: Chem.* **1996**, *108*, 57–62.

(20) Aldea, R.; Alper, H. *J. Org. Chem.* **1995**, *60*, 8365–8366.

(21) Zhou, X. T.; Ji, H. B.; Cheng, Z.; Xu, J. C.; Pei, L. X.; Wang, L. F. *Bioorg. Med. Chem. Lett.* **2007**, *17*, 4650–4653.

(22) Huang, R. L.; Espenson, J. H. *J. Mol. Catal. A: Chem.* **2001**, *168*, 39–46.

(23) Boring, E.; Geletii, Y. V.; Hill, C. L. In *Advances in Catalytic Activation of Dioxygen by Metal Complexes*; Simandi, L. I, Ed.; Kluwer Academic Publishers: Dordrecht, The Netherlands, 2003; Vol. 26, pp 227–264.

(24) Pope, M. T.; Müller, A. *Angew. Chem.* **1991**, *103*, 56–70.

(25) Hill, C. L.; Prosser-McCartha, C. M. *Coord. Chem. Rev.* **1995**, *143*, 407–455.

(26) Okuhara, T.; Mizuno, N.; Misono, M. *Adv. Catal.* **1996**, *41*, 113–252.

(27) Hill, C. L. *Chem. Rev.* **1998**, *98*, 1–390.

(28) Neumann, R. *Prog. Inorg. Chem.* **1998**, *47*, 317–370.

(29) *Polyoxometalate Chemistry From Topology via Self-Assembly to Applications*; Pope, M. T., Müller, A., Eds.; Kluwer Academic Publishers: Dordrecht, The Netherlands, 2001.

(30) *Polyoxometalate Chemistry for Nano-Composite Design*; Yamase, T., Pope, M. T., Eds.; Kluwer Academic/Plenum Publishers: New York, 2002; Vol. 2.

(31) Kozhevnikov, I. V. *Catalysis by Polyoxometalates*; Wiley: Chichester, England, 2002; Vol. 2.

(32) Pope, M. T. In *Comprehensive Coordination Chemistry II: From Biology to Nanotechnology*; Wedd, A. G., Ed.; Elsevier Ltd.: Oxford, U.K., 2004; Vol. 4, pp 635–678.

(33) Hill, C. L. In *Comprehensive Coordination Chemistry-II: From Biology to Nanotechnology*; Wedd, A. G., Ed.; Elsevier Ltd.: Oxford, U.K., 2004; Vol. 4, pp 679–759.

(34) Neumann, R. *Mod. Oxid. Methods* **2004**, 223–251.

(35) Long, D.-L.; Burkholder, E.; Cronin, L. *Chem. Soc. Rev.* **2007**, *36*, 105–121.

(36) Bosch, E.; Kochi, J. K. *J. Chem. Soc., Perkin Trans.* **1995**, *1*, 1057–1064.

(37) Chen, C. Y.; Zhang, H.; Zhang, L. X.; Li, L. D.; Yan, Y. L. *Chin. J. Org. Chem.* **2008**, *28*, 1978–1981.

(38) Zhang, H.; Chen, C.-Y.; Liu, R.-H.; Xu, Q.; Liu, J.-H. *Synth. Commun.* **2008**, *38*, 4445–4451.

(39) Hill, C. L.; Okun, N. M.; Hillesheim, D. A.; Geletii, Y. V. In *Anti-Terrorism and Homeland Defense: Polymers and Materials*; Reynolds, J. G., Lawson, G. E., Koester, C. J., Eds.; ACS Symposium Series 980; American Chemical Society: Washington, DC, 2007; Chapter 12, pp 198–209.

(40) Zhang, H.; Chen, C.; Liu, R.; Xu, Q.; Zhao, W. *Molecules* **2010**, *15*, 83–92.

(41) Domaille, P. J. In *Inorganic Syntheses*; Ginsberg, A. P., Ed.; John Wiley and Sons: New York, 1990; Vol. 27, pp 96–104.

(42) *CRC Handbook of Chemistry and Physics*, 81st. ed.; Lide, D. R., Ed.; CRC Press: Boca Raton, FL, 2000.

(43) Allen, G. D.; Buzzeo, M. C.; Villagrán, C.; Hardacre, C.; Compton, R. G. *J. Electroanal. Chem.* **2005**, *575*, 311.

(44) *Bromine Compounds Chemistry and Applications*; Raphael, L.; Price, D., Iddon, B., Wakefield, B. J., Eds.; Elsevier: New York, 1988; pp 369–384.

(45) Domínguez, A.; Iglesias, E. *Langmuir* **1988**, *14*, 2677.

(46) Lengyel, I.; Nagy, I.; Bazsa, G. *J. Phys. Chem.* **1989**, *93*, 2801–2807.

(47) Blair, C. M.; Brass, P. D.; Yost, D. M. *J. Am. Chem. Soc.* **1934**, *56*, 1916–1918.

(48) Rossi, L. I.; Suárez, Á. R. *Sulfur Lett.* **2002**, *25*, 123–127.

(49) Suárez, Á. R.; Baruzzi, A. M.; Rossi, L. I. *J. Org. Chem.* **1998**, *63*, 5689.

(50) Rossi, L. I.; Rossi, R. H. d. *Appl. Catal., A* **2004**, *267*, 267–272.

(51) Scaife, D. B.; Tyrrell, H. J. V. *J. Chem. Soc.* **1958**, 386–392.

(52) Kowalski, P.; Mitka, K.; Ossowska, K.; Kolarska, Z. *Tetrahedron* **2005**, *61*, 1933–1953.

(53) Higuchi, T.; Gensch, K.-H. *J. Am. Chem. Soc.* **1966**, *88*, 5486.

(54) Gensch, K.-H.; Pitman, I. H.; Higuchi, T. *J. Am. Chem. Soc.* **1968**, *90*, 2096–2104.

(55) Miotti, U.; Modena, G.; Sedea, L. *J. Chem. Soc. B* **1970**, 802.

(56) Choudhary, K.; Suri, D.; Kothari, S.; Banerji, K. K. *J. Phys. Org. Chem.* **2000**, *13*, 283.

(57) Luo, Z. Ph.D. Dissertation, Emory University, Atkanta, GA, 2009.

(58) Smith, S. G.; Winstein, S. *Tetrahedron* **1958**, *3*, 317–319.