Ferrate(VI) Oxidation of Recalcitrant Compounds: Removal of Biological Resistant Organic Molecules by Ferrate(VI)

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The oxidation of recalcitrant compounds, ethylenediaminetetraacetate (EDTA) and sulfamethoxazole (SMX) by ferrate(VI) (Fe(VI), Fe\(^{VI}\)O\(_4^{2-}\)) is presented. Kinetics of the reactions were determined as a function of pH at 25 °C by a stopped-flow technique. The rate law for the oxidation of these compounds by Fe(VI) is first-order with respect to each reactant. The observed second-order rate constants for the reaction of Fe(VI) with SMX decreased non-linearly with increase in pH and are possibly related to protonations of Fe(VI) and compounds. The individual rate constants, \(k\) (M\(^{-1}\)s\(^{-1}\)) of Fe(VI) species, HFeO\(_4^{-}\) and FeO\(_4^{2-}\) with protonated and deprotonated forms of compounds were estimated. The HFeO\(_4^{-}\) species reacts much faster with these compounds than does the FeO\(_4^{2-}\) species. The results showed that Fe(VI) has the potential to serve as an environmentally-friendly oxidant for removing biologically resistant organic molecules within minutes and converting them to relatively less toxic by-products in water.

Introduction

Chemical structures that are resistant to microbiological degradation are referred as recalcitrant compounds, which cause potential problems in the environment (1). Such compounds have been found in groundwater, treated wastewater effluent, landfill leachate, and soils irrigated with reclaimed water (2). Ethylenediaminetetraacetic acid (EDTA) and sulfamethoxazole (SMX) are nitrogen-containing molecules (Figure 1) and are not readily biodegradable (3, 4).
EDTA is a hexadentate chelating ligand and is widely employed in many commercial and industrial applications such as leather manufacturing, textile finishing, electroplating, and paper processes (2,5). The use of EDTA can be found in household items such as detergents and food items. EDTA itself does not pose significant problem, but its strong ability to complex metal ions causes an increase in the mobility of toxic metals in the aquatic environment and an adverse effect on metal removal in wastewater treatment plants (6). EDTA is not removed during conventional treatment methods and is thus found in the treated effluents and receiving waters. For example, a wide use of EDTA and its slow removal under many environmental conditions has led its status as an anthropogenic compound in many European surface waters (7). In the United States, EDTA concentrations in groundwater receiving effluent discharge have been reported as 1-72 µg/L (7).

Sulfamethoxazole (SMX) consists of two moieties, an aniline ring and a five membered heterocyclic group, connected to both sides of the sulfonamide linkage (-NH-SO₂-) (Figure 1). Sulfonamides are among the most frequently detected antibiotics in surface waters and are used in both human therapy and animal husbandry. Disposal of domestic and hospital waste, fields treated with animal manure, and runoff and infiltration from confined animal feeding operation result in the entry of sulfonamides into the aquatic environment (8-10). Among the sulfonamides, SMX has been repeatedly detected at concentrations of 200-2000 ng/L in secondary wastewater effluents and 70-150 ng/L in surface waters (10-12). Detection of sulfonamides is of concern due to the possibility of increased bacterial resistance (13). Aerobic biodegradation of sulfonamide is limited, which may be a reason explaining the finding of SMX in wastewater effluent and surface water bodies (11,14-17).

Various chemical oxidation processes and advanced oxidation methods have been applied to remove EDTA and SMX (18-34). Conventional method using chlorine may create and leave disinfection by-products (DBPs) in treated water. Ozonation has shown great potential to remove EDTA and SMX (30,33). However, ozone can form the potent carcinogenic bromate ion by reacting with bromide present in water. Another advanced oxidation process involves photocatalysis, which can transform both EDTA and SMX on titanium dioxide
Yet another promising method is the possible application of potassium ferrate(VI) (K$_2$FeO$_4$), which can address some of the concerns of currently used methods, in treating sulfonamide antimicrobials.

Ferrate(VI) (Fe$^{VI}$O$_4^{2-}$, Fe(VI)) is a powerful oxidizing agent in aqueous media with reduction potential of 2.20 V and 0.70 V in acidic and alkaline solutions, respectively (35). The spontaneous decomposition of Fe(VI) in water gives molecular oxygen and Fe(III) thus making Fe(VI) a green chemistry chemical for coagulation, disinfection, and oxidation for multipurpose treatment of water and wastewater (36-38). Furthermore, the application of Fe(VI) can improve the removal of natural organic matter or DPBs precursors. Unlike ozone, Fe(VI) does not react with bromide ion; thus the carcinogenic bromate ion would not be formed in the treatment of bromide containing water by Fe(VI).

In the last few years, it has been demonstrated that the destruction of sulfur- and nitrogen-containing contaminants by Fe(VI) can be accomplished in seconds to minutes with the formation of relatively non-toxic by-products (39-41). In this paper, the kinetics of the oxidation of EDTA and SMX is presented. Products of the oxidation reactions are briefly discussed. The results suggest that ferrate(VI) has a potential to treat recalcitrant organic compounds in water and wastewater.

Experimental Section

Potassium ferrate (K$_2$FeO$_4$) of high purity (98% plus) was prepared by oxidizing Fe(III) with hypochlorite in an alkaline medium (42). A stopped-flow spectrophotometer (SX.18 MV, Applied Photophysics, U.K.) equipped with a photomultiplier (PM) detector was used to make Fe(VI) kinetic studies. The kinetic measurements were performed under pseudo-first-order conditions in which substrates were in excess. The details are given elsewhere (43,44). Various analytical techniques were applied to evaluate the effect of the oxidation process on the fate of EDTA and SMX in water. Experimental conditions and analytical procedures used in different techniques are reported in earlier papers from this group (45,46).

Results and Discussion

Ethylenediaminetetraacetate (EDTA)

Initially, the reaction rates of Fe(VI) with EDTA were measured at 25 °C at different pH in alkaline media. The rate expression for the oxidation reaction between Fe(VI) and EDTA is given by
\[-d[\text{Fe(VI)}]/dt = k[\text{Fe(VI)}]^m[\text{EDTA}]^n\] 

(1)

where \([\text{Fe(VI)}]\) and \([\text{EDTA}]\) are concentrations of Fe(VI) and EDTA, respectively. The \(k\) represents the rate constant of the reaction, while the exponential variables \(m\) and \(n\) are the order of the reaction with respect to the assigned reactant. Excess EDTA was used in the kinetic studies to ensure that reactions were measured under pseudo-first-order conditions. Under these conditions equation (1) can be re-written as:

\[-d[\text{Fe(VI)}]/dt = k_1[\text{Fe(VI)}]^m\] 

(2)

where \(k_1 = k[\text{EDTA}]^n\) 

(3)

The absorbance versus time profile for Fe(VI) gave a single-exponential decay curve, indicating the reaction was first-order with respect to Fe(VI). The \(k_1\) values for the reaction were determined at various concentrations of EDTA. The plot of \(k_1\) versus \([\text{EDTA}]\) was linear with a correlation coefficient, \(r^2 = 0.99\) (Figure 1). The resulting slopes of the plot of log \(k_1\) vs. log[EDTA] were nearly one, indicating that the reaction with respect to EDTA is essentially first order. The rate law for this reaction can be expressed as

\[-d[\text{Fe(VI)}]/dt = k[\text{Fe(VI)}][\text{EDTA}]\] 

(4)

**Figure 2.** A plot of pseudo-first-order rate constant \((k_1, \text{s}^{-1})\) versus \([\text{EDTA}]\) at 25 °C.
Next, kinetic measurements of oxidation of amines by ferrate(VI) were carried out at pH = 9.0 and at 25 °C. Rates were found to be first order with respect to each reactant. The rate constants for oxidation of different amines with ferrate(VI) are given in Table 1.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Structure</th>
<th>$k$, M$^{-1}$s$^{-1}$</th>
<th>$t_{1/2}$, min</th>
</tr>
</thead>
<tbody>
<tr>
<td>Glycine</td>
<td>NH$_2$(CH$_2$COO')</td>
<td>93.7 ± 4.7</td>
<td>0.36</td>
</tr>
<tr>
<td>Iminodiaminetetraacetate</td>
<td>NH(CH$_2$COO')$_2$</td>
<td>18.9 ± 1.1</td>
<td>1.76</td>
</tr>
<tr>
<td>Nitrilotriaminetetraacetate</td>
<td>N(CH$_2$COO')$_3$</td>
<td>0.71 ± 0.05</td>
<td>46.9</td>
</tr>
<tr>
<td>EDTA</td>
<td>CH$_2$N(CH$_2$COO')$_2$</td>
<td>1.72 ± 0.08</td>
<td>19.4</td>
</tr>
</tbody>
</table>

Glycine is a primary amine ($1^\circ$), IDA is a secondary amine ($2^\circ$), and NTA is a tertiary amine ($3^\circ$). The order of reactivity was $1^\circ > 2^\circ > 3^\circ$. EDTA reacted faster with ferrate(VI) than NTA. The order of reactivity of amines with ferrate(VI) suggests that that ferrate(VI) attacks at the nitrogen atom sites of the amines. For example, successive cleavages of NTA at the nitrogen – carbon bond lead to the formation of ammonia (eq 1) (47). Excess ammonia can also be removed by adding additional amount of ferrate(VI) (48).

$$\begin{align*}
\text{FeO}_4^{2-} & \rightarrow \text{NH}_2\text{CH}_2\text{COO}^- \\
\text{N(CH}_2\text{COO')}_3 & \rightarrow \text{NH}_2\text{CH}_2\text{COO}^- \\
\text{IDA} & \rightarrow \text{NH}_3
\end{align*}$$

The preliminary work in the present study gave a reduction of total organic carbon (TOC) in the reaction between Fe(VI) and EDTA. The TOC reduction was approximately 20 % at a molar ratio of 6:1 ([Fe(VI)]:[EDTA]). It appears that similar to NTA, successive cleavages of EDTA at the nitrogen – carbon bond lead to the formation of acetate, which further reacts with ferrate(VI) to give reduction in total organic carbon (TOC).
The half-life of oxidation reactions were calculated considering concentration of ferrate(VI) (500 μM) was in excess than amine concentrations (Table 1). Amines can thus be removed by ferrate(VI) in second to minute time scales. It should be pointed out that the reaction rates are pH dependent; so are the half-lives of the reactions. Overall, the destruction of APC in water to relatively non-hazardous products by ferrate(VI) can be accomplished in minutes.

**Sulfamethoxazole (SMX)**

The rate law for the oxidation of SMX by Fe(VI) was found to be first-order for each reactant (46). The rate law may be written as

\[-d[\text{Fe(VI)}]/dt = k[\text{Fe(VI)}][\text{SMX}]\]  \hspace{1cm} (6)

The observed second-order rate constant, \(k\), decreased non-linearly with pH (Figure 3).

![Figure 3. The rate constant, \(k\) (M\(^{-1}\)s\(^{-1}\)) versus pH at 25 °C.](http://pubs.acs.org)

A change in \(k\) with pH can be described by considering the equilibrium of the mono protonated Fe(VI) (HFeO\(_4\)) and SMX (SH) (SH ⇌ H\(^+\) + S\(^-\); pK\(_{a,SH}\) = 5.7) (49). Two forms of Fe(VI) species (HFeO\(_4\)\(^-\) and FeO\(_4\)^{2-}\) can possibly react with two forms of SMX (SH and S\(^-\)) in the studied pH range (Model 1). In the non-linear regression of the data, it was determined that reaction between FeO\(_4\)^{2-}\)
and $S^-$ did not contribute to the rate constant, $k$ \((46)\). The rate constants for other reactions were obtained as \((46)\)

\[
\begin{align*}
HFeO_4^- + SH & \rightarrow Fe(OH)_3 + \text{Product(s)} \quad k_{15} = 3.0 \times 10^4 \text{ M}^{-1}\text{s}^{-1} \quad (7) \\
HFeO_4^- + S^- & \rightarrow Fe(OH)_3 + \text{Product(s)} \quad k_{16} = 1.7 \times 10^2 \text{ M}^{-1}\text{s}^{-1} \quad (8) \\
FeO_4^{2-} + SH & \rightarrow Fe(OH)_3 + \text{Product(s)} \quad k_{17} = 1.2 \times 10^0 \text{ M}^{-1}\text{s}^{-1} \quad (9)
\end{align*}
\]

A faster reaction rate constant of $HFeO_4^-$ with the neutral sulfonamide species ($SH$) \((7)\) than with the negatively charged ionized species ($S^-$) \((8)\) was obtained. This is in contrast with the reactivity of ozone and hydroxyl radical with amines and amino acids in which higher reactivity was observed as a result of deprotonation \((22,30,50)\). The Hoigne and Bader \((50)\) observed experimentally that ozone is almost unreactive with amines and amino acids in the highly acidic region. Contrary to this phenomena, ferrate(VI) reactivity with substrates increases in the acidic region. Reactivity of ozone and hydroxyl radicals increases with increase in pH, which is opposite to the reactivity of Fe(VI) with SMX.

In analogy with the reactivity of ozone, the pH dependence data in Figure 4 can be empirically fitted by considering interactions of ferrate species ($H_2FeO_4$, $HFeO_4^-$, and $FeO_4^{2-}$) with only deprotonated sulfonamide ($S^-$) \((eq\, 10-12)\).

\[
\begin{align*}
H_2FeO_4 + S^- & \rightarrow Fe(OH)_3 + \text{Product(s)} \quad k_{18} = 4 \times 10^6 \text{ M}^{-1}\text{s}^{-1} \quad (10) \\
HFeO_4^- + S^- & \rightarrow Fe(OH)_3 + \text{Product(s)} \quad k_{19} = 2 \times 10^2 \text{ M}^{-1}\text{s}^{-1} \quad (11) \\
FeO_4^{2-} + S^- & \rightarrow Fe(OH)_3 + \text{Product(s)} \quad k_{20} = 1 \times 10^{-1} \text{ M}^{-1}\text{s}^{-1} \quad (12)
\end{align*}
\]

According to this fitting approach (Model 2), it was assumed that ferrate species are only reactive enough with anionic SMX species ($S^-$) and other six reactions of three ferrate species with positively charged ($SH_2^+$) and neutral species ($SH$) had no role to play. However, the mechanism by which Fe(VI) oxidizes SMX may have proton involvement from the SMX (i.e. sulfonamide). This was the case in the oxidation of aniline by ferrate(VI) \((51)\) in which involvement of substrate proton was demonstrated by performing isotopic studies. Additionally, DFT calculations modeled the abstraction of hydrogen from alcohol in reaction with ferrate species. Thus fitting of pH dependence of $k$ values using Model 2 may not be appropriate. However, large difference in reactivity of $HFeO_4^-$ with $SH$ and $S^-$ cannot solely be from the electrostatic attraction phenomenon. A lesser degree of repulsion in the reaction of $HFeO_4^-$ with $SH$ facilitates the involvement of the substrate proton in the reaction that ultimately increases the rate.
Next, the stoichiometry of the oxidation of SMX by Fe(VI) was determined at neutral and basic pH (Figure 4). Slopes of linear relationships between [SMX] and [Fe(VI)] were found to be 0.27 ± 0.02 and 0.24 ± 0.01, respectively, for pH 7.0 and 9.0. The formation of oxygen and Fe(III) were determined as products of the reaction. A slope between Fe(VI) consumed and oxygen formation was found to be 0.23 ± 0.01.

![Figure 4. A plot of SMX decrease and formation of oxygen in the reaction of sulfamethoxazole with Fe(VI) under anoxic conditions. (sulfamethoxazole (△-pH 7.0, □-pH 9.0), ●-oxygen (pH 9.0)) (Data was taken from ref. 46).](image)

The oxidation of SMX by Fe(VI) thus follows a stoichiometry of 4:1 ([Fe(VI)]:[SMX]) which leads to the evolution of one mole of oxygen per mole of SMX (eq 13).

\[
4 \text{HFeO}_4^- + \text{SMX} \rightarrow 4 \text{Fe(III)} + \text{O}_2 + \text{Product(s)} \quad (13)
\]

Finally, the products analysis of the oxidation of SMX by Fe(VI) at pH 9.0 gave three products A, B, and C (Scheme 1) (46).

Scheme 1 suggests the oxidation of both the isoxazole moiety and the aniline unit of SMX by Fe(VI). At higher molar ratios of Fe(VI) to SMX, oxidation of both the isoxazole and the aniline units may occur. We may have indeed formed some products that were oxidized at both units, however, due to the increased polarity of such products, they may have been retained on the column during chromatography. Therefore the three products described in Scheme 1 may not necessarily represent all products of ferrate oxidation, but rather the ones that were isolated and characterized. Product A indicates...
opening of the isoxazole while the presence of predominantly a nitroso and a nitro group in product C and B, respectively. Importantly, the oxidation process by Fe(VI) is the destruction of the aromatic ring, isoxazole ring, which will undoubtedly render the oxidized product a differing biological binding property. It is expected that an oxidation of the amino group and/or an oxidation of the isoxazole ring (which leads to its potential opening/destruction) will change its binding properties sufficiently rendering it less of a mimic for the important p-aminobenzoic acid. The latter is necessary in the synthesis of the essential vitamin: folic acid. Thus Fe(VI) not only removes SMX in water, but also produces by-products that are expected to be less toxic.

Acknowledgment

We wish to thank Dr. Y. Lee for useful comments.

References


