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Nitrite Reduction by Siderite

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Nitrite Reduction by Siderite

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Nitrate-dependent Fe(II) oxidation is an important process in the inhibition of soil Fe(III) reduction, yet the mechanisms are poorly understood. One proposed pathway includes chemical reoxidation of mineral forms of Fe(II) such as siderite [FeCO$_3$(s)] by NO$_2$~-. Accordingly, the objective of this study was to investigate the reactivity of FeCO$_3$(s) with NO$_2$~-. Stirred-batch reactions were performed in an anoxic chamber across a range of pH values (5.5, 6.5, and 7.9), initial FeCO$_3$(s) concentrations (5, 10, and 15 g L$^{-1}$) and initial NO$_2$~- concentrations (0.83–9.3 mmol L$^{-1}$) for kinetic and stoichiometric determinations. Solid-phase products were characterized using x-ray diffraction (XRD). Siderite abiotically reduced NO$_2$~- to N$_2$O. During the process, FeCO$_3$(s) was oxidized to lepidocrocite [γ-FeOOH(s)] based on the appearance of XRD peaks located at 0.624, 0.329, and 0.247 nm. The rate of NO$_2$~- reduction was first order in total NO$_2$~- concentration and FeCO$_3$(s), with a second-order rate coefficient (k) of 0.55 ± 0.05 M$^{-1}$ h$^{-1}$ at pH 5.5 and 25°C. The reaction was proton assisted and k values increased threefold as pH decreased from 7.9 to 5.5. The influence of pH on NO$_2$~- reduction was rationalized in terms of the availability of FeCO$_3$(s) surface sites (>FeHCO$_3$~-, >FeOH$_2$~-, and >CO$_3$Fe~+) and HNO$_2$ concentration. These findings indicate that if FeCO$_3$(s) is present in an Fe(III)-reducing soil where fertilizer NO$_3$~- is applied, it can participate in secondary chemical reactions with NO$_2$~- and lead to an inhibition in Fe(III) reduction. This process is relevant in soil environments where NO$_3$~-, and Fe(III)-reducing zones overlap or across aerobic–anaerobic interfaces.

Abbreviations: XRD, x-ray diffraction.

Iron is the fourth most abundant element in mineral soils and is subject to changes in oxidation state (Essington, 2004). Microbial Fe(III) reduction to Fe(II) is an important process in anoxic soil environments because of its influence on organic C oxidation, soil physicochemical properties, and contaminant mobility (Lovley, 2004; Favre et al., 2002). During reduction of Fe(III) (oxy)hydroxides or phyllosilicate Fe(III), Fe(II) is released to solution and can undergo secondary processes such as adsorption and precipitation.

Siderite [FeCO$_3$(s)] is a common Fe(II) mineral produced as a result of secondary precipitation during microbial Fe(III) reduction under anoxic conditions (Coleman et al., 1993; Fredrickson et al., 1998; Zachara et al., 1998; Williams et al., 2005). Siderite has been shown to control Fe(II) solubility in anoxic sediments (Suess, 1979; Postma, 1982), rice paddy soil (Ratering and Schnell, 2000), subsoil peat horizons in close association with plant material (McMillan and Schwertmann, 1998), and coal overburden (Frisbee and Hossner, 1995; Haney et al., 2005). Oxidants for FeCO$_3$(s) include O$_2$ (Frisbee and Hossner, 1995; Duckworth and Martin, 2004), Cr(VI) (Wilkin et al., 2005), H$_2$O$_2$ (Jambor et al., 2003), and KMnO$_4$ (Haney et al., 2006).

Nitrate inhibits Fe(III) reduction to Fe(II) in soils and sediments (Sorensen, 1982; Lovley, 2000). One explanation for this inhibition is NO$_3$~--dependent Fe(II) oxidation, resulting in the anoxic production of Fe(III) (Lovley, 2000). Environments where NO$_3$~--dependent Fe(II) oxidation is important include zones where NO$_3$~--reduction and Fe(III) reduction overlap (Weber et al., 2006). Concurrent NO$_3$~-- and Fe(III)-reducing zones have been reported in laboratory incubations of field soils and pure cultures (Komatsu et al., 1978; Obuekwe et al., 1981; DiChristina, 1992). Where concomitant NO$_3$~-- and Fe(III) reduction occur, the biologically produced Fe(II) and NO$_3$~-- can react chemically, producing Fe(III) and N$_2$O (Moraghan and Buresh, 1977). The Fe(III) product resulting from solution Fe(II) oxidation by NO$_3$~-- was shown to be a poorly crystalline Fe(III) oxide mineral that was capable of affecting U cycling (Senko et al., 2005). The Fe(II)–NO$_3$~-- chemical process has been invoked to explain the apparent inhibition of Fe(III) reduction in the presence of NO$_3$~-- in pure cultures (Obuekwe et al., 1981) and anoxic paddy soil slurries (Komatsu et al., 1978). Cleemput and Baert (1983) showed that this reaction was more rapid as pH decreased. This may be attributed to the greater proportion of HNO$_3$ species. Protonation promotes N–O bond breaking, thus HNO$_2$ is a stronger oxidant than NO$_2$~-- (Shriver et al., 1977). The presence of mineral surfaces such as Fe(III) (hydr) oxide minerals can readorb Fe(II), leading to an acceleration in the electron transfer reaction rate to NO$_3$~-- because surface Fe(II) is more reactive than dissolved Fe(III) (Sorensen and Thorling, 1991; Cooper et al., 2003).

Previously, NO$_3$~-- was added to an agricultural surface soil under Fe(III)-reducing conditions and NO$_3$~-- dependent
from the stock solution to the stirred siderite suspensions. In one set of experiments designed to evaluate the influence of initial FeCO$_3$($s$) concentration on the rate of NO$_2^-$ reduction, FeCO$_3$($s$) was varied between 5 and 15 g FeCO$_3$($s$) L$^{-1}$ [corresponding to Fe(II) concentrations of 0.04–0.12 mol L$^{-1}$], while NO$_2^-$ concentration was held constant (4.6 mmol L$^{-1}$) at pH 5.5. In another set of experiments, the initial NO$_2^-$ was varied between 0.83 and 9.3 mmol L$^{-1}$ at a constant FeCO$_3$($s$) concentration of 10 g L$^{-1}$ at pH 5.5. In addition, the rate of NO$_2^-$ reduction was followed at pH 6.0, 6.5, and 7.9. The biological buffers MES [2-(N-morpholinoo)ethane sulfonic acid], PIPES (1,4-piperazine diethane sulfonic acid), and HEPES [1-piperazineethane sulfonic acid, 4-(2-hydroxyethyl)-monosodium salt] with concentrations of 0.3 mol L$^{-1}$ were added to control the pH (Alowitz and Scherer, 2002). The pH was monitored periodically and found constant throughout the experiment. All experiments were conducted at 25°C.

Blank experiments were performed [FeCO$_3$($s$) free] by shaking 4.6 mmol L$^{-1}$ NO$_2^-$ in MES at pH 5.5 to account for possible self-decomposition of NO$_2^-$ (Cleemput and Baert, 1983). We performed another experiment in which 4.6 mmol L$^{-1}$ NO$_2^-$ was added to dissolved Fe(II) extracted from the dissolving FeCO$_3$($s$) mineral to determine if dissolved Fe(II) could be responsible for the reaction. This experiment is referred to as “dissolved Fe(II).” Suspensions were removed periodically and filtered through a 0.2-μm membrane filter (Fisher Scientific, Hampton, NH). Ferrozine was added to the filtrates to complex and quantify dissolved Fe(II). The residue on the filter paper was washed with anoxic water to remove any NO$_2^-$ present and treated with 0.5 mol L$^{-1}$ HCl to dissolve Fe(II) present in the solid phase. The moles of solid-phase Fe(II) reacted were determined by comparing the Fe(II) concentrations in the solution and solid phases of reacted samples with that of a control FeCO$_3$($s$) experiment (NO$_2^-$ free).

Nitrite concentration was measured using a Metrohm 792 Basic ion chromatograph (Herisau, Switzerland) with a Metrosep A column and Metrosep RP guard disk holder. The eluent was a mixture of 3.2 mmol L$^{-1}$ Na$_2$CO$_3$ and 1 mmol L$^{-1}$ HCO$_3^-$ with conductivity detection. The retention time for NO$_2^-$ was 6 min. The indophenol-blue method (Ngo et al., 1982) was used to measure NH$_4^+$. In separate experiments, N$_2$O was measured in the head space of capped

**MATERIALS AND METHODS**

**Siderite Synthesis and Characterization**

All the solutions were prepared using deionized water (18 Ω) that was made anoxic by purging with Ar for 3 h. To ensure anoxic environments, siderite synthesis and reactivity studies were conducted in an Ar–H$_2$ purged anaerobic chamber (Coy Laboratory Products, Grass Lake, MI). Siderite was synthesized by adding Na$_2$CO$_3$ to a stirred 0.5-L solution of FeCl$_2$ in equimolar amounts (0.5 mol L$^{-1}$) to form a pale gray precipitate, as described by Thamdrup et al. (1993). The siderite precipitate was washed with anoxic water to remove salts until the electrical conductivity of the wash water was lowered to background levels. The washed siderite was stored in suspension. The solid concentration of the siderite suspension was determined by weighing replicate subsamples. Subsamples of siderite were removed for characterization of alkalinity and solution Fe(II) at the native pH of the washed siderite (pH 8.0) and at pH 6.0. The pH and alkalinity values were used to calculate total carbonate concentration using MINEQL+ (Schecher and McAvoy, 1998), assuming a system closed to atmospheric CO$_2$. Solution Fe(II) was measured using the ferrozine-[3-(2-pyridyl)-5,6 bis(4-phenylsulfonic acid)-1,2,4-triazine, monosodium salt] method by following absorbance at a wavelength of 562 nm (Stookey, 1970) with a ultraviolet–visible–near-infrared scanning spectrophotometer (Shimadzu, UV-3101 PC, Columbia, MD).

X-ray diffraction was used to characterize the gray precipitate. A slurry of the siderite mineral was mixed with Ar-degassed glycerol to prevent Fe(II) oxidation and dried under Ar. Scans were taken from 2 to 80° 2θ with CuKα radiation using a Siemens D-500 powder diffractometer fitted with a graphite monochromator and NaI (TI) scintillation detector. The XRD peaks located at 0.359, 0.279, 0.234, 0.213, 0.196, 0.179, 0.173, 0.152, and 0.144 nm confirmed the presence of siderite (Sharp, 1960).

**Stirred-Batch Experiments**

Stock solutions of NO$_2^-$ were prepared by dissolving certified ACS-grade NaNO$_2$ in deoxygenated, deionized water in a glove box. Reactions were performed in stirred-batch mode in duplicate 30-mL glass bottles. Experiments were initiated by adding aliquots of NO$_2^-$ to the stock solution to the stirred siderite suspensions. In one set of experiments designed to evaluate the influence of initial FeCO$_3$($s$) concentration on the rate of NO$_2^-$ reduction, FeCO$_3$($s$) was varied between 5 and 15 g FeCO$_3$($s$) L$^{-1}$ [corresponding to Fe(II) concentrations of 0.04–0.12 mol L$^{-1}$], while NO$_2^-$ concentration was held constant (4.6 mmol L$^{-1}$) at pH 5.5. In another set of experiments, the initial NO$_2^-$ was varied between 0.83 and 9.3 mmol L$^{-1}$ at a constant FeCO$_3$($s$) concentration of 10 g L$^{-1}$ at pH 5.5. In addition, the rate of NO$_2^-$ reduction was followed at pH 6.0, 6.5, and 7.9. The biological buffers MES [2-(N-morpholinoo)ethane sulfonic acid], PIPES (1,4-piperazine diethane sulfonic acid), and HEPES [1-piperazineethane sulfonic acid, 4-(2-hydroxyethyl)-monosodium salt] with concentrations of 0.3 mol L$^{-1}$ were added to control the pH (Alowitz and Scherer, 2002). The pH was monitored periodically and found constant throughout the experiment. All experiments were conducted at 25°C.

Blank experiments were performed [FeCO$_3$($s$) free] by shaking 4.6 mmol L$^{-1}$ NO$_2^-$ in MES at pH 5.5 to account for possible self-decomposition of NO$_2^-$ (Cleemput and Baert, 1983). We performed another experiment in which 4.6 mmol L$^{-1}$ NO$_2^-$ was added to dissolved Fe(II) extracted from the dissolving FeCO$_3$($s$) mineral to determine if dissolved Fe(II) could be responsible for the reaction. This experiment is referred to as “dissolved Fe(II).” Suspensions were removed periodically and filtered through a 0.2-μm membrane filter (Fisher Scientific, Hampton, NH). Ferrozine was added to the filtrates to complex and quantify dissolved Fe(II). The residue on the filter paper was washed with anoxic water to remove any NO$_2^-$ present and treated with 0.5 mol L$^{-1}$ HCl to dissolve Fe(II) present in the solid phase. The moles of solid-phase Fe(II) reacted were determined by comparing the Fe(II) concentrations in the solution and solid phases of reacted samples with that of a control FeCO$_3$($s$) experiment (NO$_2^-$ free).

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XRD was performed as described above. 30-mL glass vials using a Varian 3700 gas chromatograph with a 2 mol L$^{-1}$ packed column, Porapak Q, with a thermal conductivity detector and 20 mL min$^{-1}$ He carrier gas.

Solid-phase reaction products were collected for samples reacted with 4.6 mmol L$^{-1}$ NO$_2^-$ for 24 and 96 h and compared with a control siderite sample. Slurries were mixed with glycerol, dried, and XRD was performed as described above.

RESULTS AND DISCUSSION

Stoichiometry

Siderite readily reduced NO$_2^-$ (Fig. 2A–2C). For example, approximately 60% of the initial NO$_2^-$ was lost from solution after 23 h at pH 5.5 and initial NO$_2^-$ and FeCO$_3$(s) levels of 4.6 mmol L$^{-1}$ and 10 g L$^{-1}$ (Fig. 2A). No significant NO$_2^-$ loss occurred in the blank experiments [FeCO$_3$(s) free] at pH 5.5 during the same time frame. This rules out self-decomposition of NO$_2^-$ to be important at pH < 5.0 (Bartlett, 1981). The major product identified for NO$_2^-$ reduction was N$_2$O based on gas chromatography. No NH$_4^+$ was detected. Wet chemical extractions of Fe(II) showed that 0.50 ± 0.25 mmol L$^{-1}$ was oxidized and 0.24 ± 0.1 mmol L$^{-1}$ of NO$_2^-$ was reduced after 1 h of reaction time in experiments containing 10 g L$^{-1}$ FeCO$_3$(s) and 4.6 mmol L$^{-1}$ NO$_2^-$ (Fig. 2A). This indicates a 2:1 consumption of Fe(II) per mole of NO$_2^-$ reduced and is consistent with the formation of N$_2$O as the major product of NO$_2^-$ reduction. Similarly, Sorensen and Thorling (1991) found N$_2$O as the major product of NO$_2^-$ reduction in the presence of Fe(II) bound to lepidocrocite. Hansen et al. (1994) reported production of N$_2$O during reduction of NO$_2^-$ by green rust.

Solid-phase products were characterized using XRD in reacted and control (NO$_2^-$ free) samples. The diagnostic d-spacings at 0.359, 0.279, 0.234, 0.196, 0.179, 0.152, and 0.144 nm revealed that FeCO$_3$(s) was the sole mineral present after 96 h in control experiments (Fig. 3a). Siderite reacted with NO$_2^-$ under identical conditions as in Fig. 2A [10 g L$^{-1}$ FeCO$_3$(s), 4.6 mmol L$^{-1}$ NO$_2^-$, pH 5.5] showed a decrease in the intensity of the 104 reflection at 0.279 nm and the 116 reflection at 0.173 nm after 24 h of reaction (Fig. 3b), where approximately 2.8 mmol L$^{-1}$ NO$_2^-$ was reduced (Fig. 2A). At longer times (96 h) in the NO$_2^-$-reacted samples, the appearance of peaks at 0.624, 0.329, and 0.247 nm indicated the production of lepidocrocite [$\gamma$-FeOOH(s)] (Fig. 3c). The weak feature at 0.173 nm could represent unreacted siderite or the 151 reflection of lepidocrocite. In addition to the lepidocrocite peaks, one diffraction peak was observed for goethite [$\alpha$-FeOOH(s)] at 0.420 nm.

Past studies have documented the important role of O$_2$ in siderite oxidation. Resulting oxidation products include ferrihydrite (Duckworth and Martin, 2004), lepidocrocite, and goethite (Senkayi et al., 1986; Frisbee and Hossner, 1995; Haney et al., 2006). Our results show that NO$_2^-$ can function as an oxidant of siderite under anoxic conditions and produce lepidocrocite as the primary mineral, which coexisted with goethite (Schwertmann and Taylor, 1972; Ishikawa et al., 2005). It appears that poorly crystalline ferrihydrite formed in our experiments as a precursor to lepidocrocite based on the broad feature from 15 to 40° 2θ after 24 h (Fig. 3b). These results suggest the following overall reaction to be operative under our experimental conditions:

$$4\text{FeCO}_3(s) + 2\text{NO}_2^- + 2\text{H}^+ + \text{H}_2\text{O} \rightleftharpoons 4\gamma\text{-FeOOH}(s) + \text{N}_2\text{O}(g) + 4\text{CO}_2(g)$$  \[1\]

It is possible that a homogeneous reaction involving dissolved Fe(II) in equilibrium with the solid FeCO$_3$(s) and NO$_2^-$ could occur in addition to the heterogeneous reaction with FeCO$_3$(s).

We tested this possibility by reacting NO$_2^-$ with dissolved Fe(II) extracted from the dissolving FeCO$_3$(s) mineral, referred to as “dissolved Fe(II)” in Fig. 2A. There was negligible reactivity in the reaction of NO$_2^-$ with dissolved Fe(II) from the FeCO$_3$(s) mineral within a period of 23 h (Fig. 2A). This suggests that structural Fe(II) in FeCO$_3$(s) or adsorbed Fe(II)–FeCO$_3$(s) species are involved in reducing NO$_2^-$.

Fig. 2. Time course of NO$_2^-$ reduction at (A) varying siderite level (5, 10, and 15 g L$^{-1}$) compared with a blank (siderite free) and dissolved Fe(II) sample with an initial NO$_2^-$ concentration of 4.6 mmol L$^{-1}$ at pH 5.5; (B) varying initial NO$_2^-$ concentration at pH 5.5 and a siderite concentration of 10 g L$^{-1}$; and (C) varying pH values, with a siderite concentration of 10 g L$^{-1}$ and NO$_2^-$ of 4.6 mmol L$^{-1}$. Error bars represent the standard deviation from the mean for duplicate runs.
Kinetic Analysis

Kinetic data were analyzed using the method of initial rates and isolation (Lasaga, 1981). One can write the overall rate equation for NO$_2^-$ reduction by FeCO$_3$(s) as

\[ -\frac{d[NO_2^-]}{dt} = k[FeCO_3(s)] \left[ NO_2^- \right]_T^x \]  \[ \text{[2]} \]

where \(-d[NO_2^-]/dt\) is the rate of disappearance of [NO$_2^-$]$_T$, the sum of dissolved species of NO$_2^-$; \(k\) is the overall rate coefficient; and \(x\) and \(y\) are reaction orders for FeCO$_3$(s) and NO$_2^-$, respectively. The initial reaction rate was determined by regression analysis of the initial linear phase of [NO$_2^-$]$_T$ removal. For experiments where an excess of NO$_2^-$ is present at pH 5.5 and the initial concentration of NO$_2^-$ is varied, Eq. [2] reduces to

\[ -\frac{d[NO_2^-]}{dt} = k_I[NO_2^-]_T^y \]  \[ \text{[3]} \]

where \(k_I = k[FeCO_3(s)]^x\). Taking the logarithm of both sides of Eq. [3] allows one to calculate \(y\), the reaction order for NO$_2^-$, which is the slope of the least squares linear fit. In the same way, [FeCO$_3$(s)] is varied at constant [NO$_2^-$]$_T$ and pH to calculate \(x\).

There was a first-order dependence on [NO$_2^-$]$_T$ based on the slope of the regression, which is close to unity (Fig. 4A). These results differ from those of Hansen et al. (1994), where a zero-order dependence of NO$_2^-$ was observed in the reduction by sulfate green rust. These differences may be explained by the differences in mineral structure. Sulfate green rust has a layered structure containing external and internal sites for NO$_2^-$, with SO$_4^{2-}$ functioning as a charge-balancing interlayer anion (Hansen et al., 1994). Siderite, however, has a rhombohedral unit cell (Sharp, 1960) and possesses external reactive sites. The reaction order for [FeCO$_3$(s)] was 1.02 ± 0.02, which indicates a first-order dependence on FeCO$_3$(s) (Fig. 4B). This implies a surface-controlled process and may be explained by the fact that at higher FeCO$_3$(s) concentrations, more surface sites are available for reaction.

Thus, the reduction of NO$_2^-$ by siderite can be described by an overall second-order rate expression:

\[ -\frac{d[NO_2^-]}{dt} = k[FeCO_3(s)] \left[ NO_2^- \right]_T \]  \[ \text{[4]} \]

The average rate coefficient (\(k\)) calculated using Eq. [4] at pH 5.5 for the NO$_2^-$ and siderite concentration ranges used (0.83–9.3 and 42–120 mmol L$^{-1}$, respectively) was 0.55 ± 0.05 M$^{-1}$ h$^{-1}$.

The reduction of NO$_2^-$ by siderite was pH dependent. The second-order rate coefficients increase threefold as pH...
5.5 to 7.9, the >FeHCO₃ and >FeOH₁ and >FeO₂⁻ where
 kombined with their corresponding mass action
 reactions and are characterized by stability constants (Table 1). The
 equilibrium expressions (Van Cappellen et al., 1993; Pokrovsky and
 Schott, 2002). The >FeOH₀ and >CO₃H₀ groups on the
 surface sites and NO₂⁻ is related to surface complexes through
 speciation.

Fig. 5. Second-order rate coefficients for the reduction of NO₂⁻ by
 siderite at different pH values, for 10 g L⁻¹ siderite and an initial
 NO₂⁻ concentration of 4.6 mmol L⁻¹.

decreased from 7.9 to 5.5 (Fig. 5). The influence of pH can be
 rationalized on the basis of the relative distribution of siderite
 surface sites and NO₂⁻ speciation.

Siderite bears several types of surface sites in the presence
 of water, based on comparisons from its isostrophic counterpart, calcite (Van Cappellen et al., 1993). In the presence of
 water, siderite forms two primary sites on the surface, >FeO₀ and >CO₃H₀ groups. The speciation of siderite surface sites has
 been described using the surface complexation model, where
 solution chemistry is related to surface complexes through
 equilibrium expressions (Van Cappellen et al., 1993; Pokrovsky and
 Schott, 2002). The >FeOH₀ and >CO₃H₀ groups on the
 siderite surface undergo protonation and deprotonation reactions and are characterized by stability constants (Table 1). The
 total number of reactive sites for >FeO₀ groups, denoted as
 [>Fe]ₜ, is related to the sum of protonated and deprotonated surface species:

\[ [>\text{Fe}]_\text{t} = [>\text{FeOH}^+] + [>\text{FeOH}_2^-] + [>\text{FeO}^2-] + [>\text{FeHCO}_3^-] + [>\text{FeCO}_3^-] \]  

One can rearrange and express Eq. [5] in terms of [>FeOH₀]:

\[ \frac{[>\text{FeOH}^0]}{[\text{H}^+]^2} = K_1 [>\text{FeOH}^+] \text{[H}^+\text{]} + K_2 [>\text{H}_2\text{O}] \text{[CO}_3\text{]} + K_3 [>\text{FeHCO}_3^-] \text{[H}^+\text{]} + K_4 [>\text{FeCO}_3^-] \]  

where \( K_1 \) to \( K_4 \) correspond to the stability constants for
 Reactions 1 to 4 (Table 1). Expressions can be derived for
 [>FeO⁻⁻], [>FeOH₂⁺], [>FeHCO₃⁻], and [>FeCO₃⁻] using Eq.
 [5] and [6] combined with their corresponding mass action
 equations. The concentration of CO₂⁻ was calculated using
 MINEQL+ with the average of all alkalinity determinations,
 assuming a system closed with respect to atmospheric CO₂. We
 assumed a value for [>Fe]ₜ of 4 x 10⁻⁴ mol sites L⁻¹ based on
 crystallographic data as described by Wersin et al. (1989) under
 comparable experimental conditions [10 g FeCO₃(s) L⁻¹].

Figure 6A shows the speciation of >FeOH₀ surface sites as a
 function of pH. Under our experimental conditions of pH
 5.5 to 7.9, the >FeHCO₃⁻ and >FeOH₂⁺ sites are dominant
 fractions, followed by >FeCO₃⁻ surface species (Fig. 6A). The
 >FeO₀ and >FeO⁻⁻ species do not become important until
 much greater pH values, as observed elsewhere (Van Cappellen
 et al., 1993; Pokrovsky and Schott, 2002). A similar approach
 was taken for the >CO₃H₀ surface sites using Reactions 5 and
 6 (Table 1). Calculations for >CO₃Fe⁺ were based on mean
 solution Fe(II) concentrations measured in control bottles at
 pH 6 and 7.9. Figure 6B shows that negatively charged species,
 >CO₃⁻ sites, were predicted to be abundant, followed by
 >CO₃Fe⁺ and >CO₃H₀.

The speciation of dissolved NO₂⁻ must be accounted for to help explain the pH dependence in its reduction rate. Total NO₂⁻
 concentration ([NO₂⁻]ₜ), which was measured in our experiments using ion chromatography, is the sum of protonated (HNO₃) and
deprotonated (NO₂⁻) forms and is expressed by the following mass balance expression:

\[ [\text{NO}_2^-]_\text{t} = [\text{HNO}_3^-] + [\text{NO}_2^-] \]  

Nitrite can undergo protonation and deprotonation reactions
 depending on pH:

\[ \text{HNO}_2^- \leftrightarrow \text{NO}_2^- + \text{H}^+ K_a = 5 \times 10^{-4} \]  

where \( K_a \) is the acid dissociation constant. Equations [7] and
 [8] can be rearranged to solve for [HNO₃⁻] and [NO₂⁻] as a
 function of pH:

\[ [\text{NO}_2^-]_\text{t} = [\text{HNO}_3^-] \left( 1 + \frac{K_a}{[\text{H}^+]} \right) \]  

There are several possible combinations of precursor surface
 complexes that may form in the transition state between
 siderite surface sites and different chemical species of NO₂⁻
 to explain the pH dependence in Fig. 5. Of the >FeOH₀ surface
 sites, we assumed >FeHCO₃⁻ and >FeOH₂⁺ to be most
 important based on their abundance (see Fig. 6A). The >CO₃H₀ Fe⁺ site, where Fe(II) is bound on FeCO₃(s), was considered because
 past studies have shown that Fe(II) bound on lepidocrocite was
 reactive toward NO₂⁻ (Sorensen and Thorling, 1991). In addition,
 the possibility of >CO₃Fe⁺ playing a role was suggested based
 on the lack of reaction between dissolved Fe(II) [in equilibrium
 with FeCO₃(s)] and NO₂⁻ (see Fig. 2A).

Possible precursor surface complexes were calculated as a
 product of their concentrations and as a function of pH (Fig.
 7). This approach has been used elsewhere to describe sul-
 fide oxidation by Fe(III) and Mn(IV) minerals and assumes that
 precursor surface complexation is rate limiting (Yao and
 Millero, 1993; Poulton, 2003). The products [>FeHCO₅⁻]
 [HNO₃⁻] and [>CO₃Fe⁺][HNO₃⁻] sharply increased below pH

Table 1. Surface complexation reactions and corresponding
 stability constants (K) considered in the siderite–water inter-
 face (temperature T = 298 K, ionic strength I = 0)†.

<table>
<thead>
<tr>
<th>Reaction no.</th>
<th>Reaction</th>
<th>Log K</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>&gt;FeOH₀⁻⁺ ↔ &gt;FeO⁻⁻ + H⁺</td>
<td>-10.4</td>
</tr>
<tr>
<td>2</td>
<td>&gt;FeOH₀⁻⁺ + H⁺ ↔ &gt;FeOH⁺⁺</td>
<td>10.2</td>
</tr>
<tr>
<td>3</td>
<td>&gt;FeOH₀⁻⁺ + CO₂⁻⁻ + 2H⁺ ↔ &gt;FeHCO₃⁻ + H₂O</td>
<td>22.75</td>
</tr>
<tr>
<td>4</td>
<td>&gt;FeOH₀⁻⁺ + CO₂⁻⁻ + H⁺ ↔ &gt;FeCO₃⁺ + H₂O</td>
<td>14.65</td>
</tr>
<tr>
<td>5</td>
<td>&gt;CO₃H₀⁻ ↔ &gt;CO₃⁻ + H⁺</td>
<td>-4.4</td>
</tr>
<tr>
<td>6</td>
<td>&gt;CO₃H₀⁻ + Fe⁺⁺ ↔ &gt;CO₃Fe⁺ + H⁺</td>
<td>-1.6</td>
</tr>
</tbody>
</table>

† Reactions from each site are located on the (104) plane (Pokrovsky and
 Schott, 2002).
6.5 (Fig. 7A and 7C), which is similar to the trend in the reaction rate (Fig. 5). The [>FeHCO₃⁺][NO₂⁻] product increased with a decrease in pH as well and was in greater concentration than the other complexes (Fig. 7A, right-hand y axis).

The [>FeOH₂⁺][HNO₂] form increased linearly with a decrease in pH (Fig. 7B). The Fe(II)–OH₂ bond is labile and would result in dissociation of H₂O. In solution chemistry, the water exchange rate for hexaaquo Fe(II) [Fe(H₂O)₆(aq)²⁺] is rapid, estimated to be 3.2 × 10⁶ s⁻¹ (Shriver et al., 1994). The HNO₂ could bond directly to an exposed Fe(II) surface site and undergo electron transfer reactions. It is probable that the bonding mode of HNO₂ on Fe(II) is by the N atom, because this forms a stronger complex (Hitchman and Rowbottom, 1982; Shriver et al., 1994; Figgis and Hitchman, 2000). This configuration would allow the HNO₂ to serve as a π acceptor; thus, it would be able to accept electron density from the π system of Fe(II) (Luther et al., 1992).

The variations in [>FeOH₂⁺][NO₂⁻] and [>CO₃Fe⁺][NO₂⁻] are sensitive to pH, but in the opposite direction to the reaction rate (compare Fig. 5 with Fig. 7B and 7C). Therefore, they were ruled out as possible precursor surface complexes.

It is noteworthy that three out of the four possible precursor surface complexes ([>FeHCO₃⁺][HNO₂], [>FeOH₂⁺][HNO₂], and [>CO₃Fe⁺][HNO₂]) involve HNO₂. Nitrous acid is a reactive oxidant toward Fe(II) species in solution. For example, oxidation of solution Fe(II) complexed with ethylenediaminetetraacetate exhibited a sharp increase in reaction rate with a decrease in pH. Kinetic modeling of the rate data revealed a second-order dependence on HNO₂ concentration (Zang et al., 1988). Protonation on the O weakens the N–O bond, allowing it to break (Shriver et al., 1994). Although [HNO₂] comprises only a small percentage of [NO₂⁻] at even at the lowest experimental pH of 5.5 (~0.7%), our data suggest that it is an important species kinetically.

Nitrite reduction by FeCO₃(s) is a secondary reaction in the overall process of NO₃⁻–dependent Fe(II) oxidation, a process relevant in environments where NO₃⁻ and Fe(III)-reducing zones overlap or across aerobic–anaerobic interfaces. These environments could be present in poorly drained soils containing a shallow fragipan or in freshwater wetlands. Further experiments in our lab showed that NO₃⁻ reactivity was negligible with FeCO₃(s) for time periods up to 30 d (data not shown), despite the favorable thermodynamics (Fig. 1). Thus, where fertilizer NO₃⁻ is added to soil under Fe(III)-reducing conditions [where FeCO₃(s) controls Fe(II) solubility], it is probable that bacteria containing the nitrate
The main products of the reaction were N$_2$O and a solid. This project was supported by National Research Initiative Competitive.


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