Inorganic Chemistry
Article

Inorg. Chem. **2009**, *48*, 6979–6985 **6979**
DOI: 10.1021/ic900688g

Oxidation of a Water-Soluble Phosphine and Some Spectroscopic Probes with Nitric Oxide and Nitrous Acid in Aqueous Solutions†

Andreja Bakac,* Margaret Schouten, Alicia Johnson, Wenjing Song, Oleg Pestovsky, and Ewa Szajna-Fuller

Iowa State University, Ames, Iowa 50011

Received April 8, 2009

In acidic aqueous solutions, nitrogen monoxide oxidizes monosulfonated triphenylphosphine, $TPPMS^-$, to the corresponding phosphine oxide. The NO-derived product is N_2O. This chemistry parallels that reported for the reaction of NO with the unsubstituted triphenylphosphine in nonpolar organic solvents, but the rate constant measured in this work, $5.14 \times 10^6\ M^{-2}\ s^{-1}$, is greater by several orders of magnitude. This makes $TPPMS^-$ a useful analytical reagent for NO in aqueous solution. The increased rate constant in the present work appears to be a medium effect, and unrelated to the introduction of a single sulfonate group in the phosphine. The reaction between nitrous acid and $TPPMS^-$ has a 2:1 $[TPPMS^-]/[HNO_2]$ stoichiometry and generates NH_2OH quantitatively. The rate law, rate = $4k_d[HNO_2]^2[TPPMS^-]$, identifies the second-order self-reaction of HNO_2 as the rate-limiting step that generates the active oxidant(s) for the fast subsequent reaction with $TPPMS^-$. It appears that the active oxidant is N_2O_3, although the oxides NO and NO_2 derived from it may be also involved. Bimolecular self-reaction of HNO_2 also precedes the oxidations of $ABTS^{2-}$ and TMPD. Competing with this path are the acid-catalyzed oxidations of both reagents via NO^+.

Introduction

Nitrite and nitrogen oxides play important roles in various biological processes in mammals[1] and plants,[2,3] in nitrogen dynamics in soil,[4] atmospheric (photo)chemistry,[5] and in the laboratory. The first example of nitrite acting as electron donor for anoxygenic photosynthesis has also been discovered recently.[6]

In view of their widespread involvement in nature, nitrogen oxides and nitrite/nitrous acid have been a subject of numerous mechanistic studies.[7−11] One of the mechanisms by which nitrous acid reacts with substrates involves the disproportionation of HNO_2 followed by the reaction of one or both nitrogen oxides with the substrate, as shown in eq 1 and 2.[12,13]

For a number of reactive substrates the step(s) in eq 2 is fast relative to both forward and reverse steps in eq 1. Under those circumstances, the kinetics become second order in $[HNO_2]$ and zero-order in [substrate], eq 3, where $k_d = (10.9 + 175\ [H^+])\ M^{-1}\ s^{-1}$ in acidic aqueous solutions at 25 °C,[13] and n is the stoichiometric factor, i.e., the number of equivalents of substrate consumed per rate-determining step.

$$2HNO_2 \overset{k_d,\ -H_2O}{\rightleftharpoons} N_2O_3 \rightleftarrows NO + NO_2 \quad (1)$$

$$N_2O_3 \text{ and/or } NO \text{ and/or } NO_2 + \text{substrate} \rightarrow \text{products} \quad (2)$$

$$-d[\text{substrate}]/dt = nk_d[HNO_2]^2 \quad (3)$$

Another common mechanism for oxidations with nitrous acid involves acid-catalyzed formation of NO^+ (or $H_2NO_2^+$) followed by the oxidation of substrates with NO^+, eqs 4 and 5. The rate law, eq 6, is first-order in each $[H^+]$, $[HNO_2]$, and [substrate], which makes this mechanism kinetically distinguishable from the disproportionation mechanism above.

$$HNO_2 + H^+ \rightleftarrows NO^+ + H_2O \quad K_4 \quad (4)$$

$$NO^+ + \text{substrate} \rightarrow \text{products} \quad k_5 \quad (5)$$

$$-d[\text{substrate}]/dt = k_5K_4[HNO_2][H^+][\text{substrate}] \quad (6)$$

Nitrogen dioxide, generated in eq 1, might be expected to act as an oxygen atom donor toward convenient acceptors,

† Reported in part at the 235th ACS National Meeting in New Orleans

*To whom correspondence should be addressed. E-mail: bakac@ameslab.gov.

(1) *Nitric Oxide. Biology and Pathobiology*; Ignarro, L. J., Ed.; Academic: San Diego, 2000.
(2) Neill, S. J.; Desikan, R.; Hancock, J. T. *New Phytol.* **2003**, *159*, 11–35.
(3) Crawford, N. M. *J. Exp. Bot.* **2006**, *57*, 471–478.
(4) Van Cleemput, O.; Samater, A. H. *Fert. Res.* **1996**, *45*, 81–89.
(5) Vione, D.; Maurino, V.; Minero, C.; Pelizzetti, E.; Harrison, M. A. J.; Olariu, R.-I.; Arsene, C. *Chem. Soc. Rev.* **2006**, *35*, 441–453.
(6) Griffin, B. M.; Schott, J.; Schink, B. *Science* **2007**, *316*, 1870.
(7) Ford, P. C.; Laverman, L. E. *Coord. Chem. Rev.* **2005**, *249*, 391–403.
(8) Goldstein, S.; Czapski, G. *Inorg. Chem.* **1995**, *34*, 4041–4048.
(9) Roncaroli, F.; Videla, M.; Slep, L. D.; Olabe, J. A. *Coord. Chem. Rev.* **2007**, *251*, 1903–1930.
(10) Wolak, M.; van Eldik, R. *Coord. Chem. Rev.* **2002**, *230*, 263–282.
(11) Bakac, A. *Adv. Inorg. Chem.* **2004**, *55*, 1–59.
(12) Stedman, G. *Adv. Inorg. Chem. Radiochem.* **1979**, *22*, 113–170.
(13) Pestovsky, O.; Bakac, A. *Inorg. Chem.* **2002**, *41*, 901–905.

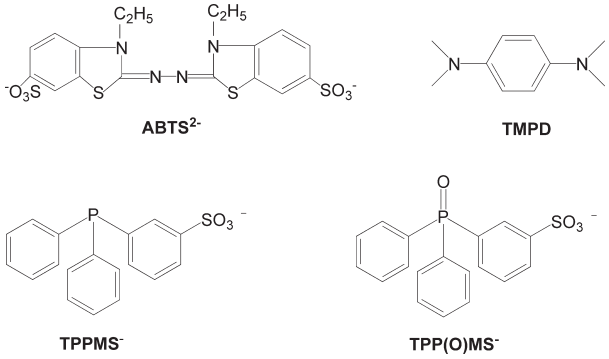
Published on Web 06/22/2009

but the available data do not support that notion. Even though a number of cases of photochemically induced oxygen atom transfer (OAT) from NO_2 have been published,[14−16] the data on thermal OAT to typical oxygen acceptors are quite limited. Even when net OAT does occur, as in the oxidation of thioethers[17] or olefins,[18] the reaction requires initiation by electron transfer. In the former reaction, the nitrosonium donor−acceptor complex [R_2S, NO^+] NO_3^- is a critical intermediate generated by thioether-induced disproportionation of NO_2.[17] The oxidation of olefins[18] requires electron transfer to generate olefin cation radicals, followed by OAT from NO_2. We are not aware of any kinetic data even for reactions with "standard" oxygen atom acceptors, such as PPh_3.

The NO_2/PPh_3 reaction to generate NO and $OPPh_3$ is thermodynamically favorable. For a gas-phase reaction, a standard enthalpy change ΔH_r^0 is about −50 kcal/mol, as calculated from the standard thermochemical data and assuming that the enthalpies of sublimation of PPh_3 and PPh_3O are not greatly different. Oxidation of PPh_3 to the corresponding oxide is among the most common and convenient reactions that allow one to gauge the ability of an oxidant to act as an oxygen atom donor. One-electron oxidation of PPh_3, which will ultimately also yield the oxide $OPPh_3$, is limited to strong 1-e oxidants because of the rather high reduction potential for the $PPh_3^{\bullet+}/PPh_3$ couple.[19−21]

The reaction of nitrogen monoxide with PPh_3, on the other hand, has been studied, although not in aqueous solutions.[22,23] The products, N_2O and $OPPh_3$, were generated in a kinetically third-order process, rate = k[PPh_3][NO]^2, that exhibited strong dependence on solvent polarity and phosphine basicity in a series of para-substituted triarylphosphines.

The complete lack of data on the reactivity of NO_2 toward oxygen atom acceptors, and the availability of NO/PPh_3 data only in nonaqueous solutions prompted us to examine the reaction of nitrous acid as a convenient source of both NO and NO_2 with a phosphine. Because the biological chemistry of nitrite and nitrogen oxides typically takes place in an aqueous milieu, the phosphine that we selected for this study is the water-soluble $P(C_6H_5)_2(3\text{-}SO_3\text{−}C_6H_4)^-$ (hereafter $TPPMS^-$). The reactions of nitrous acid with two one-electron donors, tetramethylphenylene diamine (TMPD) and 2,2′-azinobis(3-ethylbenzothiazoline-6-sulfonate ($ABTS^{2-}$), were also examined. The structures of various reagents used in this work are shown below.



Experimental Section

2,2'-Azino-bis(3-ethylbenzothiazoline-6-sulfonic acid) diammonium salt ($(NH_4)_2$ABTS, 98% pure), *N*,*N*,*N*′,*N*′-tetramethyl-*para*-phenylenediamine (TMPD, 99% pure), and ninhydrin were purchased from Sigma Aldrich. Diphenylphosphinobenzene-3-sulfonic acid sodium salt, $Na[P(C_6H_5)_2(meta\text{-}SO_3\text{−}C_6H_4]$ (NaTPPMS, >90% pure), was obtained from TCI America and used as received. In-house distilled water was further purified by passage through a Barnstead EASYpure III setup. The ionic strength was maintained at 0.10 or 1.0 M with $HClO_4$ and $LiClO_4$.

Solutions of nitrous acid were generated in situ by addition of the appropriate amount of 0.10 M $HClO_4$ into a spectrophotometric cell containing sodium nitrite (99.999%, Fisher). Gaseous NO (Matheson) was purified by passage through Ascarite, sodium hydroxide and water.[24] Stock solutions of NO were prepared by bubbling the purified gas through argon-saturated solutions of desired acidity ($HClO_4$) for 30 min. Such a solution typically contained 1.7 mM NO and 0.2−0.3 mM total nitrite.[25] Hydroxylamine was detected and quantified with ninhydrin, which reacts with primary amines to yield the intensely blue product, λ_{max} 580 nm.[26−28] The molar absorptivity of the product, determined in calibration experiments with known amounts of hydroxylamine, was found to be 4910 M^{-1} cm^{-1}. The results were confirmed with another test that involved the reduction of $Fe(phen)_3^{3+}$. The reaction solution was mixed with aqueous Fe(III) and phenanthroline (1:3 ratio) in aqueous acetate buffer (pH 4.2), and the absorbance increase was monitored at the 510 nm maximum of $Fe(phen)_3^{2+}$ ($\varepsilon = 1.14 \times 10^4$ M^{-1} cm^{-1}). The rate constant matched that measured independently with genuine NH_2OH under the same conditions. The concentration of NH_2OH was calculated from the absorbance increase using the independently established 1:1 [Fe(III)]:[NH_2OH] stoichiometry.

Kinetic measurements were initiated by injecting the desired substrate ($ABTS^{2-}$, TMPD, or $TPPMS^-$) into a spectrophotometric cell containing all the other reagents. The absorbance change was monitored by conventional spectrophotometry (Shimadzu 3101 PC) at a wavelength of maximum absorbance of $ABTS^{\bullet-}$, $TMPD^{\bullet+}$, or $TPPMS^-$. The kinetics of the NO/$TPPMS^-$ reaction were too fast for this method and required a stopped flow (Applied Photophysics DX 17 MV). Ionic strength was adjusted with $HClO_4$ and $LiClO_4$. All of the kinetic data were obtained at 25.0 ± 0.2 °C under air-free conditions (argon atmosphere). Kinetic analyses were performed with KaleidaGraph 3.6 PC software.

The solid NaTPPMS used in this work was found to contain ∼9% (by molar ratio) unreactive impurities, most

(14) Tanaka, N.; Oike, J.; Shibuya, K.; Kudoh, S.; Nakata, M. *Res. Chem. Int*. **1998**, *24*, 893–903.
(15) Harrison, J. A.; Frei, H. *J. Phys. Chem*. **1994**, *98*, 12152–12157.
(16) Nakata, M. *Spectrochim. Acta, Part A* **1994**, *50A*, 1455–1465.
(17) Bosch, E.; Kochi, J. K. *J. Org. Chem*. **1995**, *60*, 3172–3183.
(18) Bosch, E.; Kochi, J. K. *J. Am. Chem. Soc*. **1996**, *118*, 1319–1329.
(19) Yasui, S.; Kenji, I.; Ohno, A. *Heteroatom Chem*. **2001**, *12*, 217–222.
(20) Nakamura, M.; Miki, M.; Majima, T. *J. Chem. Soc., Perkin Trans. 2* **2000**, 1447–1452.
(21) Shinro, Y.; Tsujimoto, M.; Kosei, S.; A., O. *Chem. Ber./Recueil* **1997**, *130*, 1699–1707.
(22) Lim, M. D.; Lorkovic, I. M.; Ford, P. C. *Inorg. Chem*. **2002**, *41*, 1026–1028.
(23) Longhi, R.; Ragsdale, R. O.; Drago, R. S. *Inorg. Chem*. **1962**, *1*, 768–770.
(24) Pestovsky, O.; Bakac, A. *J. Am. Chem. Soc*. **2002**, *124*, 1698–1703.
(25) Song, W.; Bakac, A. *Chem.—Eur. J*. **2008**, *14*, 4906–4912.
(26) Moore, S.; Stein, W. *J. Biol. Chem*. **1948**, *176*, 367.
(27) Moore, S.; Stein, W. *J. Biol. Chem*. **1954**, *211*, 907.
(28) Moore, S. *J. Biol. Chem*. **1968**, *243*, 6281.

Table 1. Spectral Data for Some Species in this Work[a]

	ε ($\times 10^3$ M^{-1} cm^{-1}) ([H^+]/M)		
species	$\lambda = 259$ nm	$\lambda = 266$ nm	$\lambda = 369$ nm
$TPPMS^-$ [a]	9.69 (0.10) 11.2 (0.005)		
$TPP(O)MS^-$	1.59 (0.10) 1.70 (0.005)	2.08 (0.10)	
HNO_2			0.050 (0.10)

[a] Mixture of $TPPMS^-$ and TPP(H)MS.

likely moisture and water of crystallization. Such solutions were only moderately stable toward autoxidation (~5% degradation in 24 h by UV−vis). Most experiments were carried out with freshly prepared solutions and/or under anaerobic conditions.

Molar absorptivities of $TPPMS^-$ (λ_{max} 259 nm) and its oxidized form, $TPP(O)MS^-$, were obtained by oxidizing $TPPMS^-$ (~160 μM) with a limiting amount of hydrogen peroxide (33.5 μM, standardized with $TiOSO_4$)[29] under anaerobic conditions in 0.10 M $HClO_4$. This experiment provided the values of $\Delta\varepsilon_{259} = 8.10 \times 10^3$ M^{-1} cm^{-1} (0.10 M $HClO_4$) and 9.5×10^3 M^{-1} cm^{-1} (5.0 mM $HClO_4$). Another sample of $TPPMS^-$ was oxidized with a slight excess of H_2O_2. The absorbance change at 259 nm was used to calculate the absolute concentration of $TPPMS^-$ in the sample, which led to the molar absorptivities in Table 1. All of the $TPPMS^-$ solutions in this work were standardized by UV−vis spectrophotometry at 259 nm.

The acidity constant K_{aP} of TPP(H)MS was determined from the effect of [H^+] (0.01−0.10 M $HClO_4$) on the absorbance of the phosphine (100 μM) at 259 nm under anaerobic conditions and 0.10 M ionic strength ($HClO_4 + NaClO_4$). The data were fitted to eq 7, where Abs_{obs} represents the measured absorbance in the presence of added H^+, Abs_P is the absorbance in the absence of added H^+, and ε_{HP} and ε_P are the molar absorptivities of TPP(H)MS and $TPPMS^-$, respectively.

$$\frac{Abs_{obs}}{Abs_P} = \frac{1}{[H^+]+K_{aP}} \times (K_{aP} + \varepsilon_{HP}/\varepsilon_P \times [H^+]) \qquad (7)$$

Figure 1 shows the fit of the experimental data to eq 7 which yielded $K_{aP} = 0.49 \pm 0.05$ M and $\varepsilon_{HP}/\varepsilon_P = 0.30 \pm 0.06$. Similar, but much less precise data were obtained in an earlier study at variable ionic strength which resulted in H^+-dependent K_{aP}.[30]

Results

TMPD/HNO_2 Reaction. Upon mixing of the two reagents under air-free conditions, the solution turned dark blue, indicative of the formation of the radical cation $TMPD^{\bullet+}$. The kinetics were monitored at 610 nm where the molar absorptivity of the radical cation is 8600 M^{-1} cm^{-1}.[31] The reaction stoichiometry, expressed as $\Delta[TMPD^{\bullet+}]/\Delta[HNO_2]$, was cleanly 1.0 when TMPD was used in large excess. When TMPD was limiting, the initial build-up of the blue radical cation was followed by its disappearance. We interpret these data to mean that the radical cation also reacted with HNO_2 to generate a much less absorbing species, $TMPD^{2+}$,[32] in agreement with the known reaction between NO_2 and $TMPD^{\bullet+}$, $k = 1.3 \times 10^4$ M^{-1} s^{-1} at pH 5.5, 40% CH_3CN/H_2O.[33]

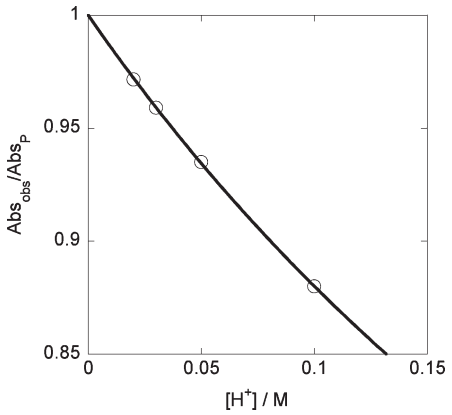


Figure 1. Plot of Abs_{obs}/Abs_P against [H^+] for the protonation of $TPPMS^-$. Conditions: $\lambda = 259$ nm, 100 μM $TPPMS^-$, 0.02 − 0.10 M $HClO_4$, 0.10 M ionic strength, argon atmosphere. The solid line is a fit to eq 7.

To avoid interference from the secondary $TMPD^{\bullet+}/HNO_2$ reaction, we studied the kinetics by the initial rate method. Concentrations were varied over a wide range and included runs with either one or the other reagent in excess. The ionic strength was kept at 1.0 M. In the acidity range used, 5 mM ≥ [H^+] ≥ 2 mM, TMPD exists as a mixture of singly and doubly protonated forms ($K_{a1} = 6.3 \times 10^{-3}$ M),[34] and nitrous acid as a mixture of HNO_2 and NO_2^- ($K_a = 6.3 \times 10^{-4}$ M).[35]

The initial rates were independent of the concentration of TMPD, and linearly dependent on $[HNO_2]^2$. By expressing the concentration of HNO_2 in terms of total concentration ($[HNO_2]_{tot} = [HNO_2] + [NO_2^-]$), eq 8, one obtains the relation in eq 9, which suggests a plot described by eq 10, where R_i stands for initial rate.

$$[HNO_2] = \frac{[H^+]}{K_a+[H^+]}[HNO_2]_{tot} \qquad (8)$$

$$R_i = k_{obs}[HNO_2]^2 = k_{obs}\left(\frac{[H^+]}{(K_a+[H^+])}[HNO_2]_{tot}\right)^2 \qquad (9)$$

$$\frac{R_i(K_a+[H^+])^2}{[H^+]^2} = k_{obs}[HNO_2]_{tot}^2 \qquad (10)$$

The fit, shown in Figure 2 yielded $k_{obs} = 12.3 \pm 0.3$ M^{-1} s^{-1}. At the acid concentrations used, the values of k_d calculated for the mechanism in eqs 1−3 from the parameters established in our earlier work[13] range from 11.2 M^{-1} s^{-1} (2 mM H^+) to 11.8 M^{-1} s^{-1} (5 mM H^+). The experimental value derived above agrees with this calculation to within less than 10% and strongly supports the mechanism in eqs 1−3. Because NO does not react with TMPD, see below, the observed 1:1 [TMPD]/[HNO_2] stoichiometry requires that each molecule of NO_2 oxidizes two molecules of TMPD.

$ABTS^{2-}$/HNO_2 Reaction. Kinetic traces were exponential when [$ABTS^{2-}$] (0.2−0.8 mM) was present in large

(29) Pestovsky, O.; Bakac, A. *J. Am. Chem. Soc.* **2004**, *126*, 13757–13764.
(30) Wright, G.; Bjerum, J. *Acta Chem. Scand.* **1962**, *16*, 1262–1270.
(31) Rao, P. S.; Hayon, E. *J. Phys. Chem.* **1975**, *79*, 1063–1066.
(32) Chaka, G.; Bakac, A. *Dalton Trans.* **2009**, 318–321.
(33) Pestovsky, O.; Bakac, A. *Inorg. Chem.* **2003**, *42*, 1744–1750.
(34) Neta, P.; Huie, R. E. *J. Phys. Chem.* **1985**, *89*, 1783–1787.
(35) Greenwood, N. N.; Earnshaw, A. *Chemistry of the Elements*, 2nd ed.; Oxford University Press: New York, 1997.

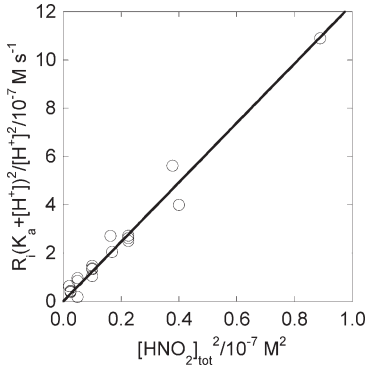


Figure 2. Plot of the initial rate of formation of $TMPD^{\bullet+}$, adjusted for the proportion of $[HNO_2]$ present as nitrite, against the square of the total concentration of HNO_2. Conditions: $[TMPD] = (1-30)\times 10^{-5}$ M, $[HNO_2] = (5-30)\times 10^{-5}$ M, $[H^+] = (2.0-5.0)\times 10^{-3}$ M, ionic strength 1.0 M.

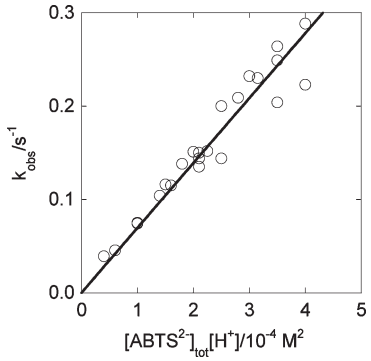


Figure 3. Plot of k_{obs} vs $[ABTS^{2-}]_{tot}[H^+]$ for the reaction of HNO_2 (0.01−0.1 mM) with excess $ABTS^{2-}$ (0.2−0.8 mM) at 0.20−0.80 M H^+ and 1.0 M ionic strength.

excess over $[HNO_2]$ (0.01−0.10 mM), and yielded pseudo-first-order rate constants that exhibited first-order dependence on each $[ABTS^{2-}]$ and $[H^+]$. A plot of k_{obs} against the product $[ABTS^{2-}]_{tot} \times [H^+]$ at 1.0 M ionic strength, shown in Figure 3, yielded the value of the third-order rate constant, $k_A = 695\pm 16$ M^{-2} s^{-1}. The concentration of H^+ was varied in these experiments over the range 0.20−0.80 M. Under these conditions, most of $ABTS^{2-}$ exists as a singly protonated species, $HABTS^-$, $pK_a = 2.08$.[36] Surprisingly, when a correction was made for the protonated fraction, and the corrected value used in a plot analogous to that in Figure 3, i.e., assuming that only the dianion $ABTS^{2-}$ was reactive, the data strongly deviated from linearity. Clearly, $ABTS^{2-}$ is not much more reactive than $HABTS^-$.

The data are consistent with the mechanism shown in eqs 4−6, with both $ABTS^{2-}$ and $HABTS^-$ reacting with NO^+, and $k_A = K_4k_5$. For $K_4 = 3\times 10^{-7}$ M^{-1},[37] one obtains $k_5 = (2.32 \pm 0.05) \times 10^9$ M^{-1} s^{-1}. In view of the rate constant being close to diffusion-controlled value, it is not surprising that the reactivity is not greatly affected by converting $ABTS^{2-}$ ($E^0 = 0.68$ V) to a somewhat less reducing $HABTS^-$ ($E^0 = 0.81$ V).[36]

(36) Scott, S. L.; Chen, W.-J.; Bakac, A.; Espenson, J. H. *J. Phys. Chem.* **1993**, *97*, 6710–6714.
(37) Bayliss, N. S.; Dingle, R.; Watts, D. W.; Wilkie, R. J. *Aust. J. Chem.* **1963**, *16*, 933–942.

The kinetic behavior changed at higher concentrations of HNO_2 when $ABTS^{2-}$ became the limiting reagent. The appearance of the kinetic plots was close to second-order, especially at low $[H^+]$ (< 0.05 M). Initial rates were used to calculate the second-order rate constants k_{HNO_2} under the assumption that the mechanism in eqs 1−3 dominates. Data analysis, analogous to that outlined above for TMPD, yielded the second-order rate constants of 16.5 M^{-1} s^{-1} at 0.010 M H^+ and 19.8 M^{-1} s^{-1} at 0.05 M H^+, close to the expected values, 12.7 and 19.6 M^{-1} s^{-1}, respectively. These calculations assumed an overall 1:1 $[HNO_2]/[ABTS^{2-}]$ stoichiometry based on the lack of reactivity of $ABTS^{2-}$ toward NO and the 2:1 $[ABTS^{2-}]/[NO_2]$ stoichiometry.

$TPPMS^-/HNO_2$ Reaction. The stoichiometry was determined with either reagent in excess. In experiments with excess HNO_2, the absorbance change corresponding to the oxidation of $TPPMS^-$ to $TPP(O)MS^-$ was measured at 259 nm ($\Delta\varepsilon_{259} = 8.10\times 10^3$ M^{-1} cm^{-1}), 290 nm ($\Delta\varepsilon_{290} = 3.48\times 10^3$) or 300 nm ($\Delta\varepsilon_{300} = 1.43\times 10^3$). When excess $TPPMS^-$ was employed, the stoichiometry was determined from the absorbance change at either 369 nm ($\varepsilon_{HNO_2} = 50$ M^{-1} cm^{-1}) or a convenient wavelength below 300 nm where $TPPMS^-$ absorbs. Throughout the range of concentrations used, i.e., 1.0 mM $\geq [HNO_2] \geq$ 0.02 mM and 1.0 mM $\geq [TPPMS^-] \geq$ 0.1 mM, and regardless of which reagent was in excess, the overall stoichiometry, $[TPP(O)MS^-]_{inf}/\Delta[HNO_2]$, was in the range 1.9 ± 0.1.

Consistent with this stoichiometry, the HNO_2-derived product was hydroxylamine. In an experiment starting with 1.0 mM $TPPMS^-$ and 1.0 mM HNO_2 in 5.0 mM $HClO_4$, all of the $TPPMS^-$ was oxidized, and the final concentration of NH_2OH was 0.51 ± 0.03 mM (average of two determinations) by ninhydrin method, and 0.41 ± 0.01 by iron phenanthroline method. The reaction is thus described by eq 11.

$$HNO_2 + 2TPPMS^- + H_2O \rightarrow NH_2OH + 2TPP(O)MS^- \quad (11)$$

The kinetic data were obtained by the method of initial rates at 2.0−5.0 mM $HClO_4$. The rate of the formation of $TPP(O)MS^-$ was second order in $[HNO_2]$ and independent of $[TPPMS^-]$ as shown in Figure 4 which yielded $k = 39.0 \pm 1.6$ M^{-1} s^{-1}, close to the calculated value of $4k_d$ (defined in eq 1) in this acidity range (i.e., 45 M^{-1} s^{-1} at 2 mM H^+, and 47 M^{-1} s^{-1} at 5 mM H^+). The rate law is given in eq 12.

$$-d[TPPMS^-]/dt = -2d[HNO_2]/dt = 4k_d[HNO_2]^2 \quad (12)$$

At higher acid concentrations, the appearance of the traces was still second-order, but the rate constants increased more than expected for the mechanism in eqs 1−3. As in the $ABTS^{2-}$ reaction, it appears that the mechanism in eqs 4−6 begins to contribute as $[H^+]$ increases. These conditions were not explored further.

$TPPMS^-$/NO Reaction. The 2:1 stoichiometry was determined from the absorbance changes at 259 nm caused by the addition of limiting amounts of NO to an

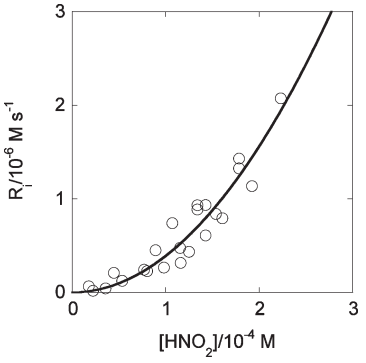


Figure 4. Dependence of initial rate for the $TPPMS^-/HNO_2$ reaction on $[HNO_2]$. Conditions: $[TPPMS^-] = (2-30) \times 10^{-5}$ M, $[HNO_2] = (2-25) \times 10^{-5}$ M, $[H^+] = 2-5$ mM, ionic strength 0.10 M. The line is a fit to second-order kinetics with $k = 39.0 \pm 1.6\ M^{-1}\ s^{-1}$.

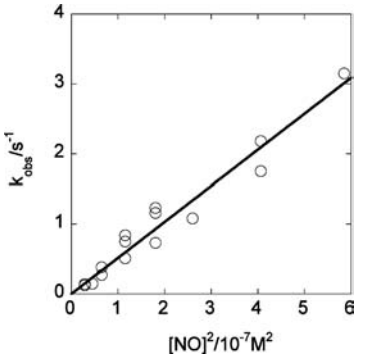


Figure 5. Plot of k_{obs} versus $[NO]^2$ for the reaction with $TPPMS^-$ in 2.0 mM $HClO_4$. $[TPPMS^-] = 5 \times 10^{-6}$ M.

excess of $TPPMS^-$. The formation of N_2O as reaction product was confirmed by GC-MS. The overall reaction is shown in eq 13.

$$2NO + TPPMS^- \rightarrow N_2O + TPP(O)MS^- \qquad (13)$$

The kinetics were studied under pseudo-first-order conditions using a large excess of NO over $TPPMS^-$. A plot of the observed pseudo-first-order rate constants against $[NO]^2$ is linear, Figure 5, confirming the second-order dependence on [NO], eq 14, and yielding $k_{13} = (5.14 \pm 0.20) \times 10^6\ M^{-2}\ s^{-1}$.

$$-d[TPPMS^-]/dt = k_{13}[TPPMS^-][NO]^2 \qquad (14)$$

$TPPMS^-$ as Analytical Reagent for NO. The large rate constant and substantial associated absorbance changes in the UV, see Table 1, are attractive features for an analytical method for quantitative determination of small concentrations of NO. This idea was tested by using $TPPMS^-$ to quantify NO generated in the TMPD/HNO_2 reaction.

On the basis of the observed 1:1 TMPD/HNO_2 stoichiometry, and the value of the rate constant for the disappearance of TMPD approaching k_d, we concluded that only one of the nitrogen oxides generated in eq 1 reacted with TMPD. If both NO_2 and NO were reactive, then both the stoichiometry and the kinetics would be twice larger than observed. Moreover, TMPD does not have the potential to reduce NO, so that one should expect NO to accumulate in TMPD-HNO_2 solutions.

The reaction between TMPD (1×10^{-4} M) and HNO_2 (5×10^{-5}M) in 5.0 mM $HClO_4$ at 0.10 M ionic strength was allowed to go to completion (9600 s), as established by monitoring the absorbance increase at the 610-nm maximum of $TMPD^+$. At that point, $TPPMS^-$ was injected (final concentration 7.86×10^{-5} M) and the reaction with NO was monitored at 259 nm. From the absorbance decrease, we obtained $[NO] = 3.81 \times 10^{-5}$ M, and from the initial rate we calculated $k = 5.33 \times 10^6\ M^{-2}\ s^{-1}$. The excellent agreement between the measured rate constant and that determined independently for the $TPPMS^-$/NO reaction confirms that the species reacting with $TPPMS^-$ was indeed NO. The concentration of NO was 76% of that calculated (5×10^{-5} M) from the 1:1 TMPD/HNO_2 stoichiometry. We consider the agreement satisfactory in view of the slowness of the step that generated the NO. Despite all the precautions, some loss of NO to traces of air over such long times is unavoidable.

Discussion

The NO/$TPPMS^-$ reaction in water closely parallels the NO/PPh_3 reaction in organic solvents.[22] In both cases the overall stoichiometry is [NO]/[Phosphine] = 2, kinetics are second order in NO, and N_2O is produced, as in eq 13,14. The rates are, however, quite different; the oxidation of $TPPMS^-$ in H_2O is orders of magnitude faster than the NO/PPh_3 reaction in toluene and chloromethanes, Table 2. Even though electron-donating substituents in the *para*-position have an accelerating effect on the PPh_3 reaction ($\rho = -1.5$, recalculated from data in ref 22 which reported $\rho = -4.5$), even the most reactive phosphine studied, $P(4\text{-}OCH_3\text{-}C_6H_4)_3$, has a rate constant of only $1.6 \times 10^3\ M^{-2}\ s^{-1}$ in toluene.

The much greater reactivity of $TPPMS^-$ appears to be mostly solvent effect and not the electronic effect of a single sulfonate group on one of the phenyls. We base this conclusion on the magnitude of the effect, which is disproportionately large when compared to that observed in $P(4\text{-X-}C_6H_4)_3$, see above. In those cases, the replacement of the most electron donating substituent (X = OCH_3) with the most electron-withdrawing one (Cl) in all three phenyls caused the rate constant in toluene to decrease 260-fold. Upon introducing a sulfonate in a single phenyl in PPh_3 and changing the solvent to water, the rate constant increased 5.6×10^4-fold. Even though the sulfonate renders the phosphine anionic, the electron density at phosphorus does not appear to be dramatically greater than in PPh_3, as shown by the similarity of acidity constants for the protonated forms of PPh_3 (ca. 1 M, estimated from the data in H_2O/CH_3CN)[38] and $TPPMS^-$ (0.49 M, see above). The difference between the acidity constants for $(4\text{-X-}C_6H_4)_3P$ (X = Cl, H) are much greater, at least in the mixed H_2O/CH_3CN solvent.[38]

The strong solvent effect as well as the electronic effects in phenyl-substituted phosphines are both consistent with the two mechanisms already considered in the literature,[22] i.e., the reaction of the phosphine with NO dimer, eq 15, or the formation of a phosphine/NO adduct followed by the reaction with another molecule of NO, eq 16. This step generates

(38) Pestovsky, O.; Shuff, A.; Bakac, A. *Organometallics* **2006**, *25*, 2894–2898.

Table 2. Summary of Kinetic Data Obtained in This Work[a]

reaction	rate law[b]	stoich.	conditions	products
$TPPMS^-$ [c] + NO	$k[TPPMS^-][NO]^2$ $k = (5.14 \pm 0.20) \times 10^6\ M^{-2}\ s^{-1}$	1:2	2−5 mM H^+	$TPP(O)MS^- + N_2O$
PPh_3 + NO [d]	$k[PPh_3][NO]^2$ $k = (1.57 \pm 0.02) \times 10^3\ M^{-2}\ s^{-1}$	1:2	CH_2Cl_2	$OPPh_3 + N_2O$
PPh_3 + NO [d]	$k[PPh_3][NO]^2$ $k = (1.05 \pm 0.01) \times 10^3\ M^{-2}\ s^{-1}$	1:2	$CHCl_3$	$OPPh_3 + N_2O$
PPh_3 + NO [d]	$k[PPh_3][NO]^2$ $k = 91 \pm 1\ M^{-2}\ s^{-1}$	1:2	$PhCH_3$	$OPPh_3 + N_2O$
TMPD + HNO_2	$k_d[HNO_2]^2$ [e]	1:1	2 - 5 mM H^+	$TMPD^{\bullet+}$ + NO
$ABTS^{2-}$ + HNO_2	$k_d[HNO_2]^2$ [e]	1:1	2 - 10 mM H^+	$ABTS^{\bullet-}$ + NO
$ABTS^{2-}$ + HNO_2	$k_A[ABTS]_{tot}[HNO_2][H^+]$ $k_A = 695 \pm 16\ M^{-2}\ s^{-1}$ $k_5 = (2.32 \pm 0.05) \times 10^9\ M^{-1}\ s^{-1}$ [f]	1:1	0.2−0.8 M H^+	$ABTS^{\bullet-}$ + NO
$TPPMS^-$ + HNO_2	$4k_d[HNO_2]^2$ [e]	2:1	2−5 mM	$TPP(O)MS^- + NH_2OH$

[a] In acidic ($HClO_4$) aqueous solutions at 298 K, unless stated otherwise. [b] For disappearance/formation of absorbing species ($TPPMS^-$, PPh_3, $ABTS^{\bullet-}$, $TMPD^{\bullet+}$). The rate of disappearance of coreactants is related to the given rate law by the stoichiometric factor given in third column. [c] $TPPMS^- = P(C_6H_5)_2(3\text{-}SO_3{-}C_6H_4^-)$. [d] 294 K, see ref 22. [e] $k_d = (10.9 + 175\ [H^+])\ M^{-1}\ s^{-1}$, ref 13, see eq 1. [f] For reaction of NO^+ with $ABTS^{2-}$/$HABTS^-$.

$TPP(N_2O_2)MS^-$ intermediate, which eliminates N_2O and forms $TPP(O)MS^-$.

$$2NO \overset{K_{dim}}{\rightleftarrows} (NO)_2 \xrightarrow{+TPPMS^-,\ k_{15}} TPP(O)MS^- + N_2O \qquad (15)$$

$$NO + TPPMS^- \overset{K_{NO}}{\rightleftarrows} TPP(NO)MS^- \xrightarrow{+NO,\ k_{16}}$$

$$[TPP(N_2O_2)MS]^- \rightarrow TPP(O)MS^- + N_2O \qquad (16)$$

The preference for the second mechanism was expressed[22] on the basis of the small estimated equilibrium constant for the formation of $(NO)_2$ from NO in the gas phase ($K_{dim} < 3 \times 10^{-4}\ M^{-1}$),[39] and the assumption that K_{dim} has a similar value in toluene. In the meantime, the binding energy for $(NO)_2$ in the gas phase has been reported as only 696 cm^{-1},[40] again suggesting that the equilibrium concentration of N_2O_2 in NO is extremely small and would require an exceptionally large rate constant, $\sim 1 \times 10^5\ M^{-1}\ s^{-1}$, for the $(NO)_2/PPh_3$ reaction to explain the experimental data. On this basis, the $(NO)_2$ route would seem to be disfavored.

The much greater rate constant for the $NO/TPPMS^-$ reaction observed in this work would seem to strengthen the antidimer case even more, provided that K_{dim} does not increase greatly in water. If this is the case, then the rate constant for the $(NO)_2/TPPMS^-$ reaction would have to be unreasonably high, about $1 \times 10^{10}\ M^{-1}\ s^{-1}$, which would rule out the $(NO)_2$ route. Unfortunately, the value of K_{dim} is probably not constant across various solvents, and may well be much greater in water, leaving the $(NO)_2$ mechanism a (weak) possibility.

We prefer the mechanism in eq 16, not only by disfavoring eq 15, but also because the highly polar intermediate, $[TPP^{\delta+}(NO)_2{}^{\delta-}MS]^-$, would be stabilized by polar solvents, thus explaining the strong positive effect of solvent polarity on the rate.

The rate constant for the $NO/TPPMS^-$ reaction was determined in this work in 2−5 mM $HClO_4$, but the value obtained should remain unchanged at higher pH, including pH 7, because neither reactant exhibits any acid/base chemistry in this pH range.

Regardless of the details, the fast $NO/TPPMS^-$ reaction in aqueous solutions does provide a convenient method for the determination of NO, as demonstrated in the HNO_2/TMPD reaction which generates 1 equiv. of NO.

In its reactions with TMPD and $ABTS^{2-}$, HNO_2 utilizes both mechanisms outlined in the Introduction. As expected, the mechanism in eqs 1−3 is prevalent at high $[HNO_2]$ and low $[H^+]$, i.e., conditions that favor second-order kinetics, whereas high $[H^+]$ activates the mechanism in eqs 4−6.

Both $ABTS^{2-}$ and TMPD are known to be oxidized by NO_2 with rate constants $2.2 \times 10^7\ M^{-1}\ s^{-1}$ (pH 7)[41] and $1.8 \times 10^8\ M^{-1}\ s^{-1}$ (pH 5.5, 40% CH_3CN/H_2O),[33] respectively. This fast one-electron oxidation draws the disproportionation equilibrium of eq 1 to the right and generates the observed products. As one would expect, the NO that is also generated in eq 1 does not react with either TMPD or $ABTS^{2-}$, as shown by the overall 1:1 stoichiometry for the reaction of each with HNO_2, and by finding close to one equivalent of NO in solution after completion of the TMPD/HNO_2 reaction.

The oxidation by NO^+ of both $ABTS^{2-}$ and TMPD is extremely rapid. The rate constant for $ABTS^{2-}$/$HABTS^-$ obtained here is $2.32 \times 10^9\ M^{-1}\ s^{-1}$, and that for TMPD is expected to also be close to diffusion controlled, given that TMPD is a much better reductant than $ABTS^{2-}$. These rapid reactions with NO^+ make the mechanism in eqs 4−6 functional under the conditions that promote the formation of NO^+, i.e., at high $[H^+]$.

Oxidation of $TPPMS^-$ with HNO_2. In view of the second-order dependence on HNO_2, some variant of the mechanism in eqs 1−3 must be operative. In a temptingly simple scenario, the disproportionation of HNO_2 would be followed by the reactions of $TPPMS^-$ with the nitrogen oxides formed. One might expect NO_2 to react by oxygen atom transfer, and NO as in eq 13. This mechanism, summarized in Scheme 1, predicts a 1:1 stoichiometry and N_2O as product, clearly contradicting experimental facts, i.e., 2:1 stoichiometry and formation of NH_2OH with a rate constant approaching $4k_d$.

Undoubtedly, the second-order self-reaction of HNO_2 is still involved, but subsequent chemistry deviates from

(39) Forte, E.; Van den Bergh, H. *Chem. Phys.* **1978**, *30*, 325–331.
(40) Wade, E. A.; Cline, J. I.; Lorenz, K. T.; Hayden, C.; Chandler, D. W. *J. Chem. Phys.* **2002**, *116*, 4755–4757.
(41) Forni, L. G.; Mora-Arellano, V. O.; Packer, J. E.; Willson, R. L. *J. Chem. Soc., Perkin Trans 2* **1986**, 1–6.

Scheme 1

$$2\,HNO_2 \underset{}{\overset{-H_2O,\,k_d}{\rightleftharpoons}} N_2O_3 \rightleftharpoons NO + NO_2$$

$$NO_2 + TPPMS^- \rightarrow NO + TPP(O)MS^-$$

$$2\,NO + TPPMS^- \rightarrow N_2O + TPP(O)MS^-$$

$$2HNO_2 + 2TPPMS^- \rightarrow N_2O + 2\,TPP(O)MS^- + H_2O$$

the simple picture in Scheme 1. The reaction of the phosphine with NO is known, and that with NO_2 might be reasonably assumed to involve intermediates. In the NO reaction, the evidence comes from the kinetics. Regardless of the exact nature of the intermediate, either $(NO)_2$ or phosphine-NO adduct, the reaction would be impeded at low, steady-state concentrations of NO, i.e., when the sole source of NO is eq 1. Under such conditions, the NO-phosphine reaction may take a different course than that shown in eq 13. In particular, the intermediate may react with additional phosphine when the supply of NO is extremely low. A similar situation would apply to the NO_2-phosphine adduct.

It is also feasible, and perhaps more likely, that N_2O_3, formed in the first step in Scheme 1, reacts with the phosphine directly, without prior dissociation to NO and NO_2. Several fast subsequent steps would follow to generate the observed products, Scheme 2. As pointed out by a reviewer, the first two-electron step in N_2O_3/$TPPMS^-$ reaction may generate HNO followed by the reaction of HNO with additional $TPPMS^-$ to yield NH_2OH, similar to the reaction of HNO with thiolates.[42]

To some extent, our results also bear similarity to metal-mediated enzymatic[43] and electrocatalytic[44] reduction of nitrite, where successive addition of electrons and protons to coordinated nitrogen species generates ammonia, hydroxylamine, and N_2O. In contrast to those reactions, which are believed to be initiated by coordination of nitrite to the metal followed by coupled proton−electron transfers, the initial interaction with the phosphine does not involve HNO_2/NO_2^- but a product(s) of its bimolecular self-reaction, as established by the second order rate law of eq 12.

Scheme 2

$$2\,HNO_2 \underset{}{\overset{k_d}{\rightleftharpoons}} N_2O_3 + H_2O$$

$$N_2O_3 \xrightarrow[\text{fast}]{4\,TPPMS,\,3\,H_2O} 2\,NH_2OH + 4\,TPP(O)MS$$

Conclusions

In the reactions with TMPD and $ABTS^{2-}$, HNO_2 utilizes two mechanisms, disproportionation and NO^+-mediated oxidation, either one of which can dominate depending on the conditions. The disproportionation route generates NO and NO_2 followed by the fast NO_2/substrate reaction. Nitrogen monoxide does not react with either of these substrates, but instead accumulates in solution, as confirmed with $TPPMS^-$ which we found to be a good analytical reagent for quantitative determination of NO in aqueous solutions. The reaction of nitrous acid with $TPPMS^-$ also involves bimolecular self-reaction of HNO_2, but produces NH_2OH, suggesting rapid follow-up steps between the phosphine and N_2O_3. No evidence was found for direct OAT from NO_2.

Acknowledgment. We are grateful to the reviewers for useful and constructive comments. This work was supported by a grant from National Science Foundation, CHE 0602183. Some of the work was conducted with the use of facilities at the Ames Laboratory.

(42) Wong, P. S.-Y.; Hyun, J.; Fukuto, J. M.; Shirota, F. N.; DeMaster, E. G.; Shoeman, D. W.; Nagasawa, H. T. *Biochemistry* **1998**, *37*, 5362–5371.
(43) Einsle, O.; Messerschmidt, A.; Huber, R.; Kroneck, P. M. H.; Neese, F. *J. Am. Chem. Soc.* **2002**, *124*, 11737–11745.
(44) Barley, M. H.; Takeuchi, K. J.; Meyer, T. J. *J. Am. Chem. Soc.* **1986**, *108*, 5876–5885.