JOC The Journal of Organic Chemistry

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J. Org. Chem., **Just Accepted Manuscript** • DOI: 10.1021/acs.joc.7b01965 • Publication Date (Web): 02 Oct 2017

ACS Publications

The Journal of Organic Chemistry is published by the American Chemical Society.
1155 Sixteenth Street N.W., Washington, DC 20036

Enhancement of the Oxidizing Power of an Oxoammonium Salt by Electronic Modification of a Distal Group

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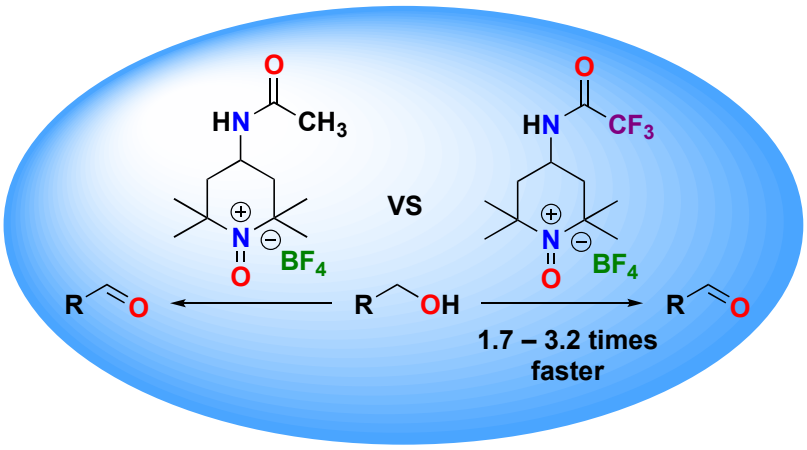


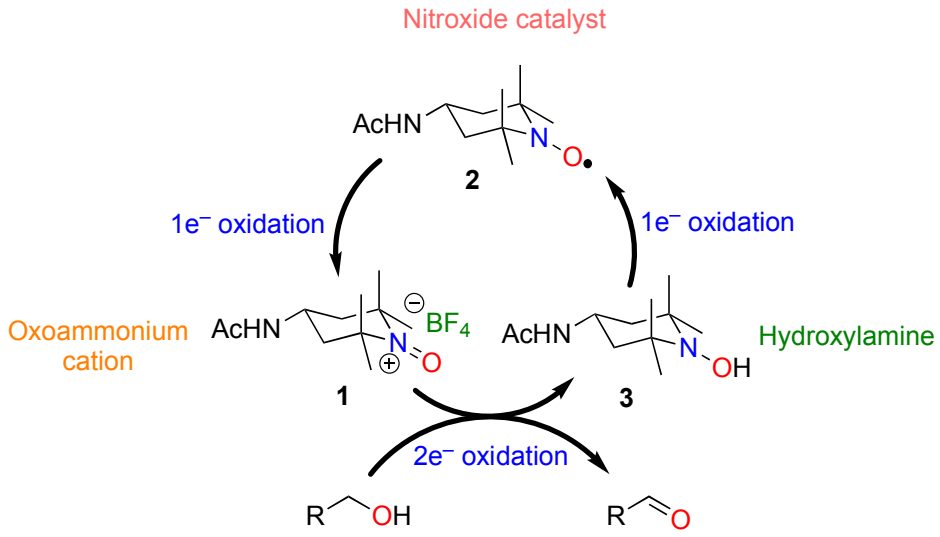
ABSTRACT: The multi-gram preparation and characterization of a novel TEMPO-based oxoammonium salt, 2,2,6,6-tetramethyl-4-(2,2,2-trifluoroacetamido)-1-oxopiperidinium tetrafluoroborate (**5**), and its corresponding nitroxide (**4**), are reported. The solubility profile of **5** in solvents commonly used for alcohol oxidations differs substantially from that of Bobbitt's salt,

4-acetamido-2,2,6,6-tetramethyl-1-oxopiperidinium tetrafluoroborate (**1**). The rates of oxidation of a representative series of primary, secondary and benzylic alcohols by **1** and **5** in acetonitrile solvent at room temperature have been determined and oxoammonium salt **5** has been found to oxidize alcohols more rapidly than does **1**. The rate of oxidation of *meta* and *para*-substituted benzylic alcohols by either **1** or **5** displays a strong linear correlation to Hammett parameters ($r > 0.99$) with slopes (ρ) of –2.7 and –2.8, respectively, indicating that the rate-limiting step in the oxidations involves hydride abstraction from the carbinol carbon of the alcohol substrate.

INTRODUCTION

Stable nitroxide radicals have drawn considerable interest among chemists for use as spin labels, antioxidants, and mediators in polymerization reactions.[1] They have also found a growing role in oxidation chemistry as precursors to oxoammonium cations.[2] Such cations may be isolated as stable salts, such as 4-acetamido-2,2,6,6-tetramethyl-1-oxopiperidinium tetrafluoroborate (**1**, Scheme 1), more commonly known as "Bobbitt's salt",[3] or they may be generated from a stable nitroxide, such as 4-acetamido-TEMPO (**2**), in a catalytic cycle (Scheme 1) involving the corresponding hydroxylamine (**3**). A number of stable nitroxide species have been identified,[4] the most well-known of which is 2,2,6,6-tetramethylpiperidine 1-oxyl (TEMPO, Figure 1). The stability of such radicals arises from the absence of an α–hydrogen on the carbons adjacent to the nitrogen possessing the oxygen radical and, in the case of TEMPO and its analogues, additional kinetic stabilization of the radical is achieved by the steric constraints imposed by the four neighboring methyl groups.

Scheme 1. Redox Cycle of Oxoammonium Salt Oxidants



In an effort to design better catalysts for alcohol oxidation, less hindered nitroxides, such as those shown in Figure 1, have been prepared.[5] While modification of the structural motif of nitroxide catalysts has been a primary focus of catalyst design, a less commonly addressed factor is the potential influence of remote substituents on the oxidizing capability of the oxoammonium cations derived from nitroxides.

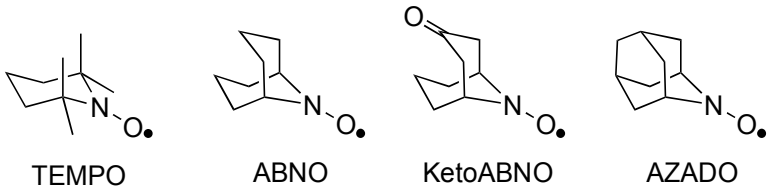


Figure 1. Selected nitroxides.

Several reports have noted that when the steric difference between two nitroxide catalysts is negligible, catalytic activity correlates with the nitroxide / oxoammonium cation redox

potential and, in general, nitroxides bearing electron-withdrawing substituents were found to have higher redox potentials. Kanai and co-workers, for example, reported that ketoABNO was more effective than ABNO in the copper-catalyzed oxidation of secondary amines to imines and they attributed this observation to the electron-withdrawing ability of the carbonyl group.[6] More recently, computational studies by Coote and Iwabuchi,[7] as well as structure-activity relationships reported by both Stahl[8] and Sigman,[9] demonstrated that electron-donating substituents appended to a nitroxide decrease its redox potential while electron-withdrawing substituents increase the potential. These results correlate well with experimental measurements of nitroxide / oxoammonium cation redox potentials.

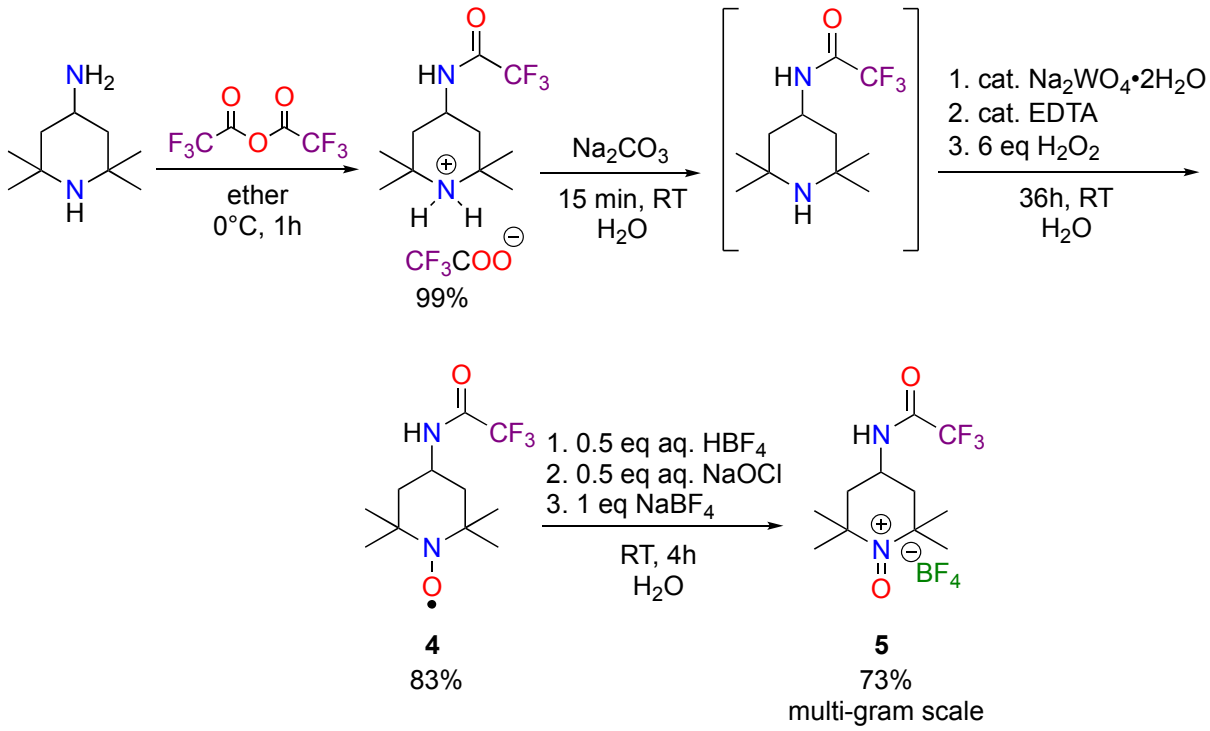
As a result of our interest in the use of Bobbitt's salt (**1**) for a variety of oxidations,[10] we made note of a 2014 paper by Szolcsányi and coworkers reporting that a TEMPO-based nitroxide containing a 4-trifluoromethylacetamido group (**4**) exhibits a large difference in anodic and cathodic peak potentials, almost double that of the 4-acetamido derivative (**2**).[11] We were interested in exploring whether the oxoammonium salt derived from this CF_3 - containing nitroxide (**4**) would oxidize alcohols more rapidly than does Bobbitt's salt (**1**). As demonstrated by the results presented below, this electronic modification of the amide moiety at the 4-position of **1** leads to a significant increase in the rate of oxidation of alcohols.

RESULTS AND DISCUSSION

The 4-trifluoroacetamido oxoammonium salt **5** was readily prepared on a multi-gram scale in good overall yield from 4-amino-2,2,6,6-tetramethylpiperidine and trifluoroacetic anhydride as illustrated in Scheme 2 by modification of the procedure used for the preparation of **1**.[3] Oxoammonium salt **5** is isolated as a yellow, crystalline solid that melts with decomposition

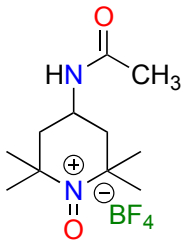
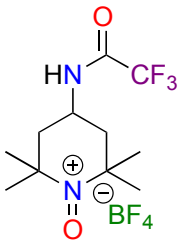
at 160°C. When left open to ambient air on the lab bench, samples of **5** turned white in color over the course of a few days. This white solid was identified as 1-hydroxy-2,2,6,6-tetramethyl-4-(2,2,2-trifluoroacetamido)piperidinium tetrafluoroborate (**6**), the protonated, reduced form of **5**, by comparison to an authentic sample. Initially it was thought that **6** resulted from oxidation of adventitious moisture by **5**. Endo for example, has reported that that the oxoammonium bromide salt derived 4-methoxy-TEMPO is reduced over the course of a few hours when dissolved in basic water.[12] However, **5** is stable for weeks when dissolved in water and sealed under argon; thus, it is not likely that residual moisture is the source of decomposition of **5**. Indeed, **5** can be stored for periods of several months without degradation in a sealed plastic container under an atmosphere of argon or nitrogen.

Scheme 2. Preparation of 2,2,6,6-Tetramethyl-4-(2,2,2-trifluoroacetamido)-1-oxopiperidinium tetrafluoroborate (5) and its Corresponding Nitroxide (4)



The solubilities of oxoammonium salts **1** and **5** in various solvents were determined and the results are summarized in Table 1. On the whole, the 4-trifluoromethylacetamido derivative **5** is more soluble in polar organic solvents such as CH_2Cl_2, EtOAc, and CH_3CN than is the parent 4-acetamido salt **1**. In water, the situation is reversed; **1** is some 2.7 times more soluble than is **5**. With the relative solubility of the two oxoammonium salts in hand, it was of interest to evaluate the relative rate of oxidation of a variety of alcohol substrates by each of the two salts.

Table 1. Solubility of Oxoammonium Salts 1 and 5 in Various Solvents.[a]

solvent[b]	 **1**	 **5**
cyclohexane	insoluble	insoluble
toluene	0.10 (3.3 x 10^{-4} M)	0.10 (2.8 x 10^{-4} M)
diethyl ether	insoluble	insoluble
dichloromethane	0.50 (1.7 x 10^{-3} M)	0.75 (2.1 x 10^{-3} M)
ethyl acetate	insoluble	0.85 (2.4 x10^{-3} M)
acetonitrile	36.0 (0.12 M)	61.7 (0.17 M)
water	70.4 (0.23 M)	30.4 (0.086 M)

[a] Solubility is expressed as grams of solute per liter of solvent and, within parentheses, as molarity. [b] All organic solvents were dried prior to use by either distillation from a dark purple solution of Na/benzophenone or calcium hydride or $MgSO_4$.

The colorimetric properties of oxoammonium salts allows for determination of the rate of disappearance of each oxoammonium salt during an oxidation by UV-Vis spectroscopy. Although both salts **1** and **5** are yellow in color, each has a unique UV-Vis spectrum in acetonitrile solvent as illustrated in Figures 2 and 3 for salts **1** and **5**, respectively. There is, however, substantial overlap in the spectra of the two salts. This overlap was dealt with, as detailed in the Experimental Section, by employing the method described by Blanco and coworkers[13] with further simplification by application of a least-squares analysis of the spectra using Excel Solver as described by Harris.[14]

The oxidation of an alcohol by an oxoammonium salt is clearly a second-order process. However, the kinetic analysis was simplified by employing a 100-fold excess of alcohol relative

to the oxoammonium salt. Under these conditions, the oxidations were found to be cleanly pseudo-first order in each salt, and plots of ln [salt] vs. time (s) gave a linear fit with the slope = – k. Rate studies were conducted in competition using both salts in approximately equal concentration; accurately determined individual concentrations of each salt in CH_3CN (~0.005 M) afforded a total oxoammonium salt concentration of ~0.01 M. The individual concentrations of each salt were monitored during alcohol oxidations at wavelengths between 370 nm and 530 nm: both **1** and **5** exhibit local absorption maxima in this wavelength range (**1** at 456 nm and 475 nm; **5** at 460 nm, and 479 nm). Neither the alcohol substrates used in the study nor the protonated hydroxylamines (the HBF_4 salt of **3**) and (**6**), the co-products of the oxidation, absorb within this range, ensuring minimal interference. Control experiments, detailed in the Supporting Information, demonstrated that aldehydes or ketones were the only products generated in the oxidations by either salt.

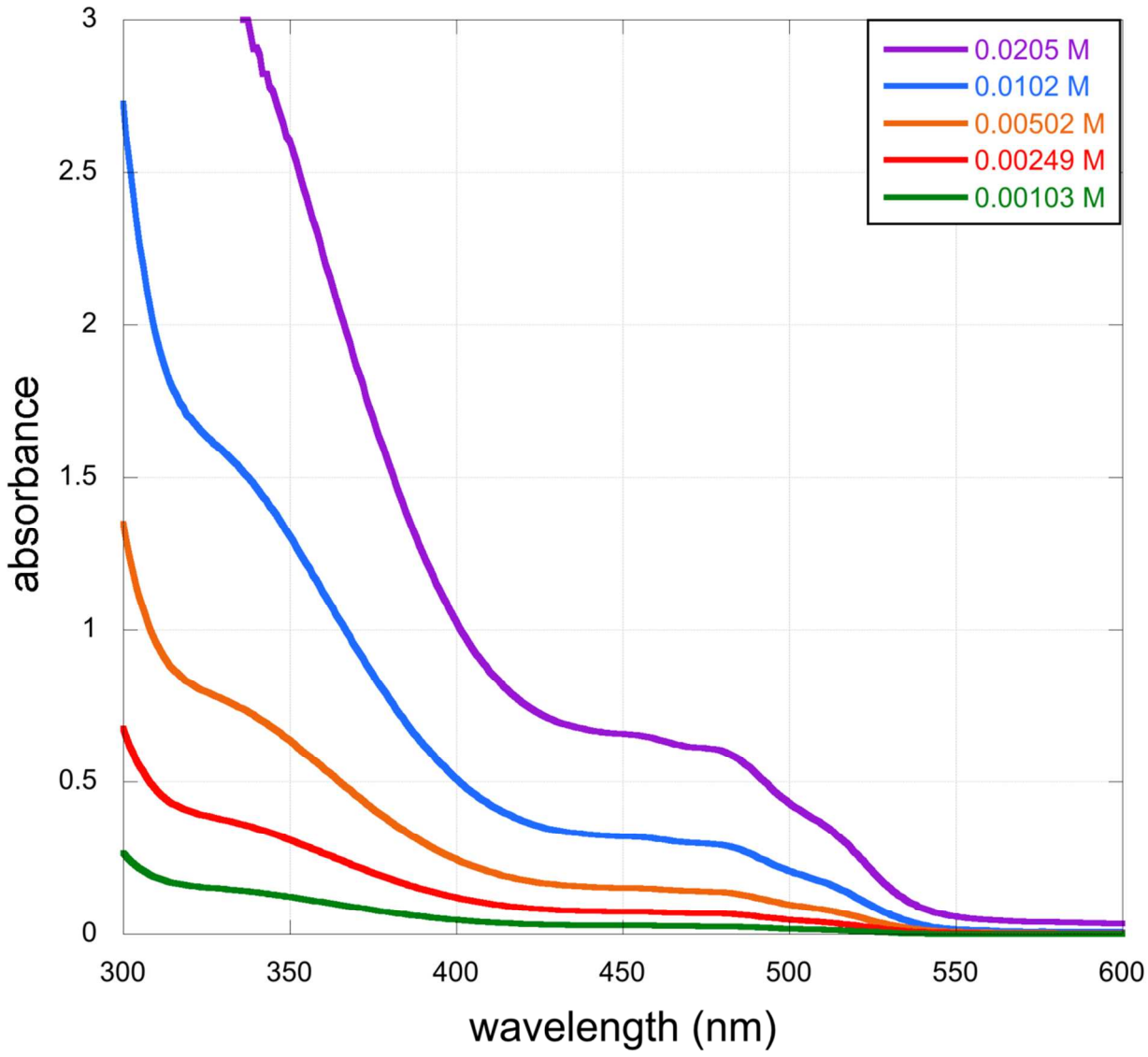


Figure 2. UV-Vis spectrum of Bobbitt's salt (**1**) in acetonitrile as function of concentration.

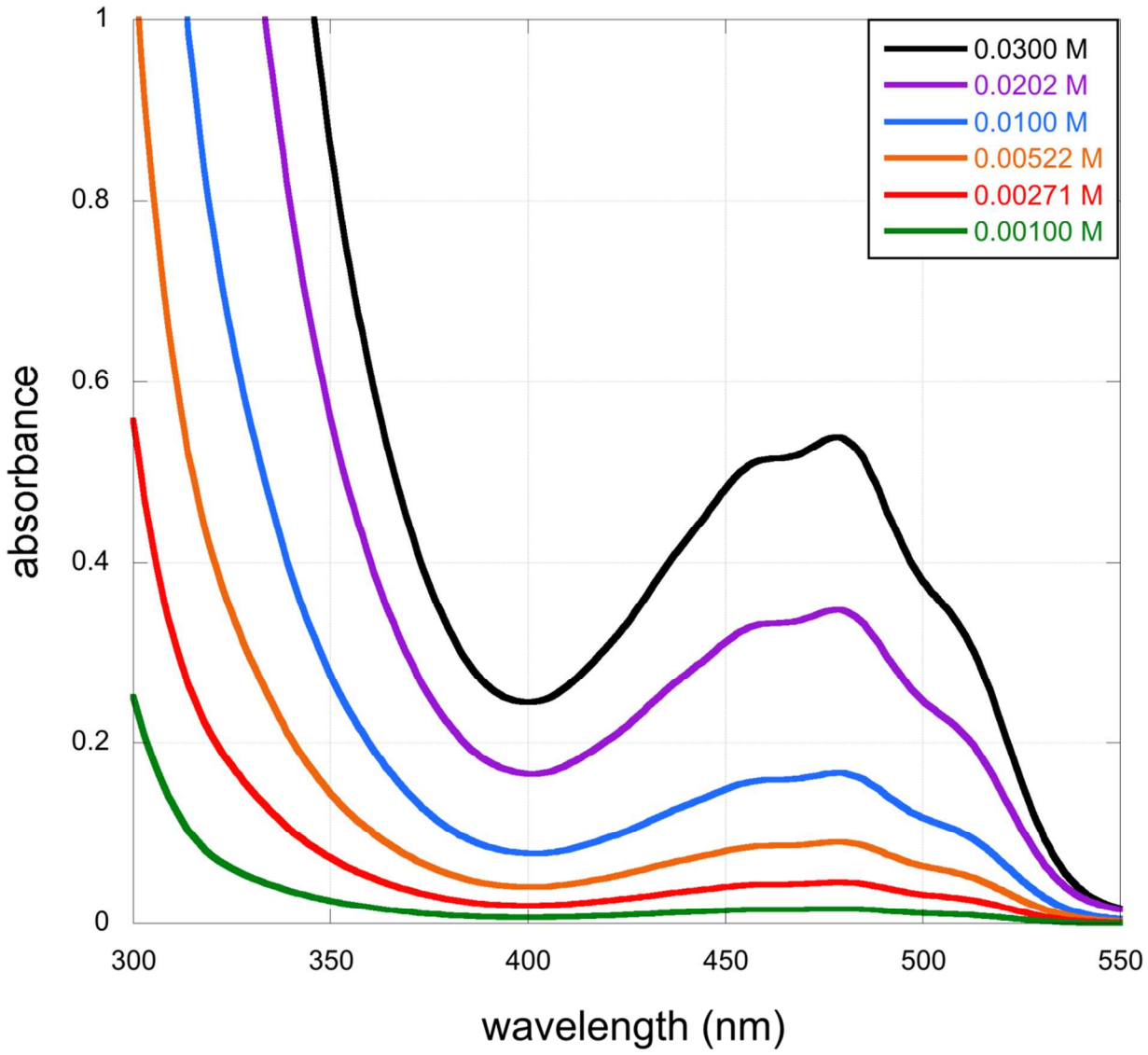


Figure 3. UV-Vis spectrum of oxoammonium salt **5** in acetonitrile as a function of concentration.

The Supporting Information includes UV-Vis spectra used in the kinetic analyses, determination of molar absorptivities, as well as plots of ln ([**1**]) and ln ([**5**]) vs. time (s) for all substrates investigated. A representative example of the kinetic data is shown below for the oxidation of *p*-methoxybenzyl alcohol (Figure 4); the pseudo-first order rate constants for

oxidation by **1** and **5** were determined from the slope of the linear least squares fit of these data (Figure 5).

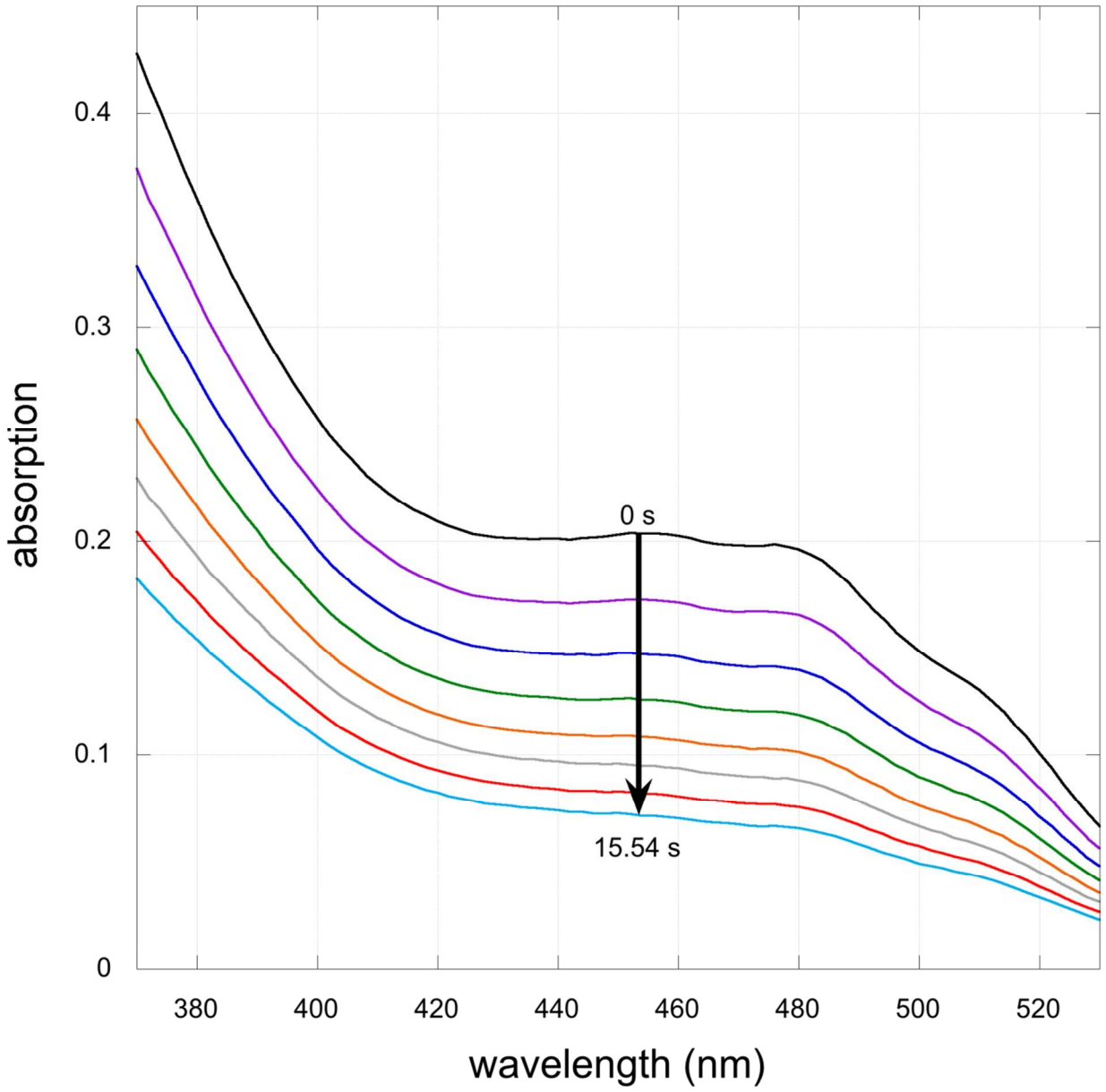


Figure 4. Change in the UV-Vis spectra of equimolar quantities of both salts (**1** and **5**) during the oxidation of *p*-methoxybenzyl alcohol.

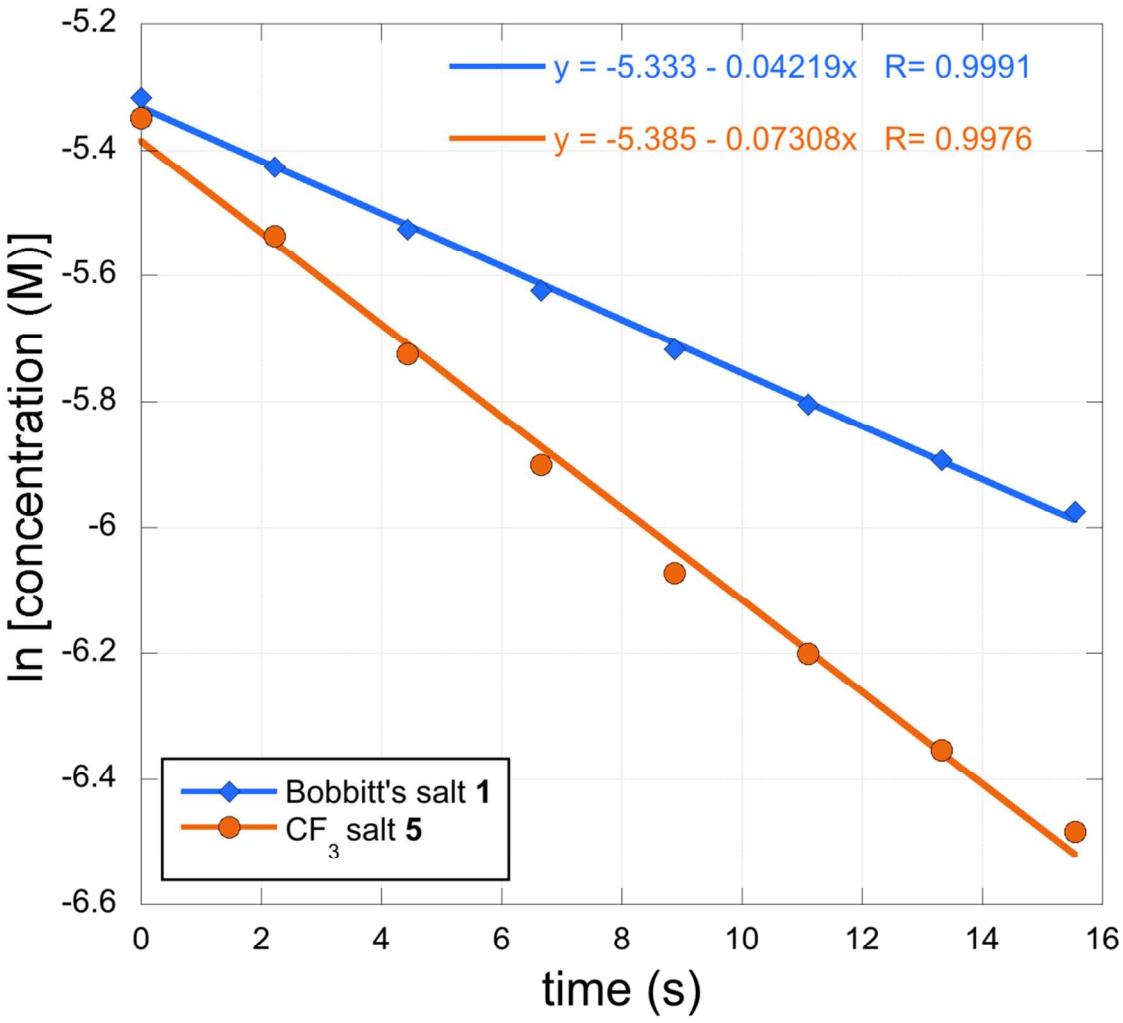
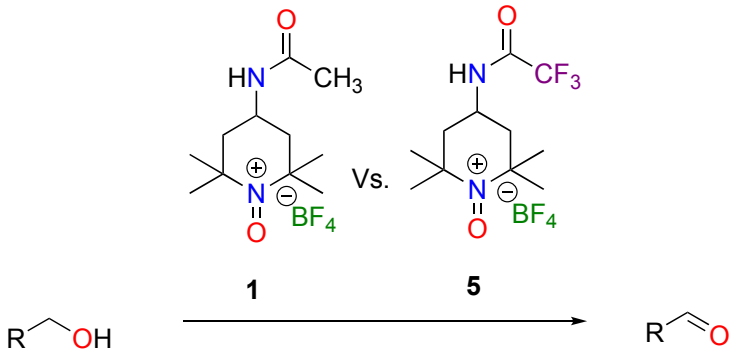


Figure 5. Rates of oxidation of *p*-methoxybenzyl alcohol by salts **1** and **5**.

The individual pseudo-first order rate constants for oxidation of a representative series of alcohols in CH_3CN solvent by **1** and **5**, as well as the half-times for the oxidations and the relative rates of oxidation of each alcohol by **1** and **5**, are summarized in Tables 2 and 3. The estimated error in the rate constants is < 5%. In all cases, **5** was found to oxidize a given alcohol at a rate generally 1.7 – 3.2 times faster than **1**. Table 2 summarizes the data for various aliphatic and unsaturated alcohols and, as expected,[15] α,β-unsaturated or benzylic substrates are oxidized

more quickly than aliphatic alcohols and secondary alcohols are oxidized faster than are unactivated primary alcohols by either oxoammonium salt. Not surprisingly, there is a considerable steric component that influences the rate of oxidation of aliphatic alcohols: neopentyl alcohol is oxidized 5 times slower by **1** and 4 times slower by **5** than is 1-heptanol. The oxidation of cinnamyl alcohol is something of an anomaly as it is apparently oxidized ~18 times faster by **5** than **1**. The origin of this significant rate difference is currently unknown, but it may simply be an artifact of the very rapid oxidation of the substrate by **5** having a half-time of only ~3 s.

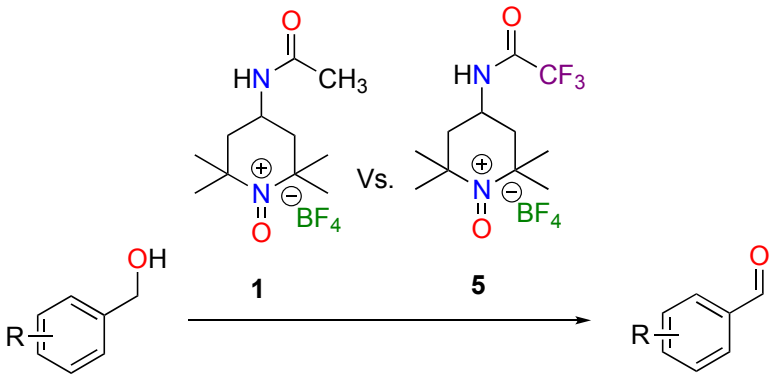
Table 2. Rates of Oxidation of Various Aliphatic and Unsaturated Alcohols by 1 and 5 in CH_3CN Solvent.



alcohol	rate of oxidation by Bobbitt's salt **1** (k x 10^4 s^{-1})	half-time of oxidation by Bobbitt's salt **1**	rate of oxidation by CF_3 salt **5** (k x 10^4 s^{-1})	half-time of oxidation by CF_3 salt **5**	relative rate of oxidation (**5**/**1**)
1-heptanol	1.73	66.8 min	3.71	31.1 min	2.2
2-hexanol	9.14	12.6 min	1.67	6.9 min	1.8
neopentyl alcohol	0.385	5.0 h	0.965	2.0 h	2.5
cyclohexanol	18.9	6.1 min	33.3	3.5 min	1.8
cyclopentanol	19.1	6.1 min	33.3	3.5 min	1.7
allyl alcohol	28.2	4.1 min	58.0	2.0 min	2.1
cinnamyl	145	48 s	267	2.6 s	18

alcohol					
piperonyl alcohol	365	19.0 s	980	7.1 s	2.7

Table 3. Rates of Oxidation of Various Substituted Benzyl Alcohols by 1 and 5 in CH_3CN Solvent.



R	rate of oxidation by Bobbitt's salt **1** (k x10^4 s^{-1})	half-time of oxidation by Bobbitt's salt **1**	rate of oxidation by CF_3 salt **5** (k x 10^4 s^{-1})	half-time of oxidation by CF_3 salt **5**	relative rate of oxidation (**5**/**1**)
p-OCH_3	422	16.4 s	731	9.5 s	1.7
p-(*t*-Bu)	165	42.0 s	332	20.9 s	2.0
p-CH_3	208	33.3 s	389	17.8 s	1.9
H	75.1	1.5 min	93.4	1.2 min	1.2
p-F	32.5	3.6 min	46.6	2.5 min	1.4
p-Br	11.7	9.9 min	37.1	3.1 min	3.2
p-CF_3	2.42	47.7 min	5.31	21.8 min	2.2
m-OCH_3	28.3	4.1 min	80.0	1.4 min	2.8
m-CH_3	10.1	1.1 min	22.4	30.9 s	2.2
m-F	9.32	12.4 min	15.4	7.5 min	1.7
m-Cl	6.53	17.7 min	10.8	10.7 min	1.7
m-CF_3	3.65	31.7 min	7.78	14.8 min	2.1

The influence of substituents on the rate of oxidation of a representative sample of *para*- and *meta*-substituted benzyl alcohols is summarized in Table 3. Cursory examination of the data presented in Table 3 demonstrates that the rate of oxidation of a benzylic alcohol by either **1** or **5** is noticeably affected by the nature of the substituent. Notably, plots of the rate of oxidation of substituted benzyl alcohols by **1** and by **5** vs. the appropriate Hammett parameters,[16] derived from the pK_as of substituted benzoic acids, give strongly linear correlations ($r = 0.99$) as illustrated in Figure 6. The slopes of these plots are essentially identical; $\rho = -2.8$ for oxidations by **1** and –2.7 for oxidations by **5**. A negative ρ-value of this magnitude indicates the development of a substantial positive charge during the rate-limiting step of the oxidation. This result provides convincing experimental support for the mechanism proposed to account for the oxidation of alcohols by an oxoammonium cation in neutral or slightly acidic media:[15] namely, the oxidation proceeds via removal of a hydride from the carbinol carbon of the alcohol resulting in the development of a positive charge on that carbon.

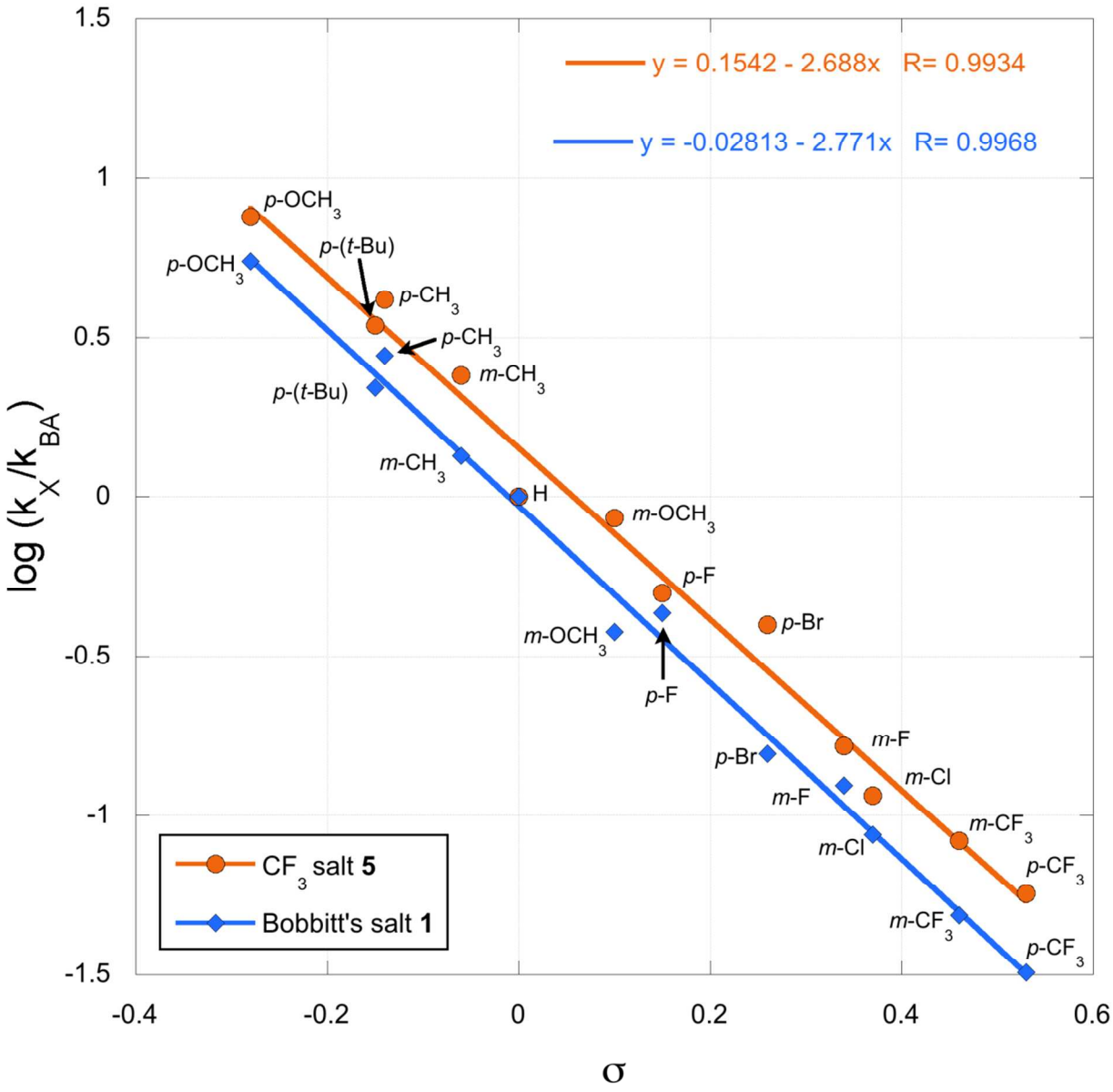
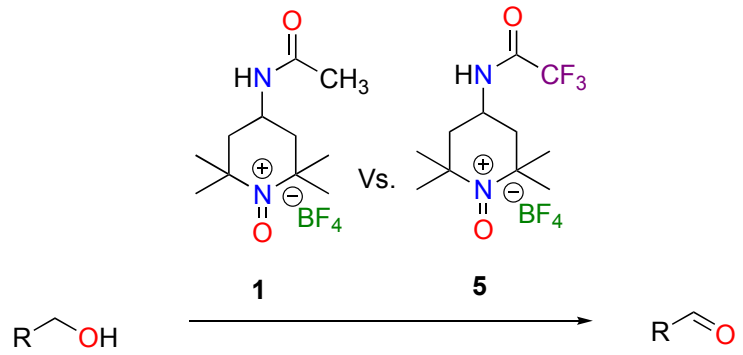


Figure 6. Hammett plot of the rates of oxidation (Table 3) of substituted benzyl alcohols by **1** and **5**; k_{BA} = rate constant for oxidation of benzyl alcohol and k_X = rate constant for oxidation of substituted benzyl alcohol.

To ensure that the relative rate differences between oxidations conducted using **1** and **5** in CH_3CN solution were not simply a result of solubility differences between the oxoammonium

salts (Table 1), a number of oxidations were conducted in water in which **1** is more soluble than **5**. The study was perforce limited to three fairly water-soluble secondary alcohols as over oxidation to the carboxylic acid occurs when primary alcohols are oxidized under aqueous conditions.[10e] In the event, the same results were obtained (Table 4): oxoammonium salt **5** oxidizes alcohols at a rate about twice that of Bobbitt's salt (**1**). The rates of oxidation of the secondary alcohols are somewhat slower in aqueous solution than when conducted in acetonitrile.

Table 4. Rates of Oxidation of Various Aliphatic Alcohols by **1** and **5** in Water.



alcohol	rate of oxidation by Bobbitt's salt **1** (k x 10^4 s^{-1})	half-time of oxidation by Bobbitt's salt **1**	rate of oxidation by CF_3 salt **5** (k x 10^4 s^{-1})	half-time of oxidation by CF_3 salt **5**	relative rate of oxidation (**5**/**1**)
2-propanol	2.44	47.3 min	5.09	22.7 min	2.1
cyclohexanol	8.80	13.1 min	1.40	8.3 min	1.6
2-butanol	3.21	36.0 min	5.37	21.5 min	1.7

CONCLUSIONS

In summary, a seemingly minor electronic modification of Bobbitt's salt (**1**), involving replacement of the 4-acetamido group with a 4-trifluoromethylacetamido group, affords an

oxoammonium salt (**5**) that oxidizes alcohols significantly more rapidly than does **1**: alcohols are oxidized 1.7 to 3.2 times more quickly by **5** than by **1**. The enhanced oxidizing power of **5** vis-à-vis **1** is attributable to the presence of the electron-withdrawing trifluoromethylacetamido group at the 4-position of the salt. The rate of oxidation of substituted benzyl alcohols by either **1** or **5** exhibits a very strong linear correlation with the Hammett parameters giving a ρ of –2.8. This observation provides experimental corroboration of the hydride-transfer mechanism proposed to account for the oxidation of alcohols by an oxoammonium cation under neutral or slightly acidic conditions.[15] In broader perspective, the results presented above suggest that development of new nitroxides, and their corresponding oxoammonium salts, might profitably benefit from consideration of the activating effect of remote electron-withdrawing substituents.

EXPERIMENTAL SECTION

General Procedures. All reactions were carried out in clean, dried glassware. NMR spectra (^{1}H, ^{13}C, ^{19}F) were recorded in either deuterated acetonitrile containing 1 v/v% tetramethylsilane, a solution of 70% trifluoroacetic acid / 30% deuterated chloroform, or a solution of 70% acetic acid / 30% deuterated chloroform, as noted below, on a 400 MHz NMR spectrometer. Chemical shifts are reported in parts per million (ppm) relative to TMS at δ = 0.00. ^{1}H and ^{13}C NMR spectra were referenced at δ = 7.26 and 77.23, respectively, for the residual ^{1}H resonance of the solvent and the center line of the ^{13}C adsorption for deuterated chloroform or at δ = 0.00 for the ^{1}H or ^{13}C resonance of TMS. ^{19}F NMR spectra were recorded in the appropriate non-fluorinated solvent, as noted below, with fluorobenzene used as an internal standard (δ = –

113.15), and chemical shifts are reported relative to CCl_3F at $\delta = 0.00$. HRMS molecular mass determinations were performed on a TOF mass spectrometer; ionization methods are noted below for individual compounds. Acetonitrile solvent was dried over calcium hydride and distilled directly prior to use in the rate studies. Bobbitt's salt (**1**) was prepared as previously described[3] and was dried overnight in a drying pistol at 56 °C (acetone, 30 mm) prior to use. Melting points are not corrected.

2,2,6,6-Tetramethyl-4-(2,2,2-trifluoroacetamido)piperidinium trifluoroacetate. To a 1.5-L Erlenmeyer flask equipped with a large stirbar was added 78.14 g (500 mmol) of 4-amino-2,2,6,6-tetramethylpiperidine and 400 mL of dry diethyl ether. The solution was cooled to 0 °C by placing the flask in an ice / water bath, stirring was initiated, and a solution of 126.0 g (600 mmol) of trifluoroacetic anhydride in 100 mL of dry diethyl ether was added dropwise over the course of 1 h. Upon addition of the anhydride a white precipitate forms. The mixture was stirred for an additional 1 h and then filtered through a Büchner funnel. The resulting off-white solid was rinsed with a 200-mL portion of diethyl ether and then dried in a vacuum oven at 120 ºC to afford 180.7 g (99% yield) of the title compound as a white powdered solid: mp 256 °C (dec), [lit.[17] mp 220 °C (dec)]; ^{1}H NMR (400 MHz, CD_3CN) δ 1.43 (s, 6H), 1.50 (s, 6H), 1.75 (t, J = 13.0 Hz, 2H), 1.96 (dd, J = 3.8 Hz, J = 13.0 Hz, 2H), 4.23–4.33 (m, 1H), 7.74 (d, J = 5.3 Hz, 1H); ^{13}C NMR (100 MHz, CD_3CN) δ 25.0, 30.5, 40.7, 42.8, 58.1, 117.0 (q, $J_{C\text{-}F}$ = 287.6 Hz), 118.1 (q, $J_{C\text{-}F}$ = 294.6 Hz), 157.4 (q, $J_{C\text{-}F}$ = 36.7 Hz), 161.2 (q, $J_{C\text{-}F}$ = 33.5 Hz); ^{19}F NMR (376 MHz, CD_3CN) δ –74.2 (3F), –74.7 (3F); HRMS (DART-TOF) *m/z* calcd for $C_{11}H_{20}F_3N_2O$ [M – $CF_3COO]^+$ 253.1528, found 253.1522.

2,2,6,6-Tetramethyl-4-(2,2,2-trifluoroacetamido)piperidine-1-oxyl (4). To a 2-L beaker equipped with a large magnetic stir bar was added 176.0 g (0.48 mol) of 2,2,6,6-

tetramethyl-4-(2,2,2-trifluoroacetamido)piperidinium trifluoroacetate and 600 mL of deionized water. The slurry was stirred for 10 min and then 71.0 g (0.67 mol) of solid sodium carbonate was added in portions over a period of 30 min.[18] Following complete addition of the Na_2CO_3 the pH of the mixture was ~ 9. At this point, 8.59 g (260 mmol) of sodium tungstate dihydrate and 12.07 g (290 mmol) of EDTA were added to the mixture followed by the slow addition over period of 3 h of 360 mL (3.2 mol) of a 30 wt% aqueous solution of hydrogen peroxide. The reaction mixture was stirred vigorously for 12 h at room temperature, an additional 120 mL (1.1 mol) of the hydrogen peroxide solution was then added dropwise over a period of 2 h, and the mixture was then stirred at room temperature for an additional 24 h. The orange reaction mixture was cooled to –10 °C in an ice/salt bath and filtered through a large Büchner funnel to give 107.5 g (83%) of the title compound as an orange solid. An analytical sample, prepared by recrystallization from EtOAc, afforded a glistening, orange-red, crystalline solid that displayed the following properties: mp 149.6–151.8 °C, (lit.[17] mp 149–152 °C); HRMS (DART-TOF) *m/z* calcd for $C_{11}H_{18}F_3N_2O_2$ $[M]^+$ 267.1320, found 267.1335.

2,2,6,6-Tetramethyl-4-(2,2,2-trifluoroacetamido)-1-oxopiperidinium tetrafluoroborate (5). To a 250-mL beaker equipped with a magnetic stir bar was charged with 20.0 g (74.9 mmol) of 2,2,6,6-tetramethyl-4-(2,2,2-trifluoroacetamido)piperidine-1-oxyl (**4**) and 50 mL of deionized water. The slurry was stirred at room temperature and 15.0 mL (7.50 g, 85.4 mmol) of a 50 wt% aqueous solution of tetrafluoroboric acid was added dropwise over the course of 15 min. The mixture was stirred for ~ 0.5 h until the solution was a yellow color, at which point, 33.8 mL (2.79 g, 37.4 mmol, NaOCl) of an 8.25 wt% solution of commercial bleach was added dropwise to the vigorously stirred mixture over a period of 0.5 h. Following the addition of bleach, 8.22 g (74.9 mmol) of sodium tetrafluoroborate was added and the reaction

mixture was stirred for an additional 1 h. The reaction mixture was then cooled to –10 °C in an ice / salt bath and filtered through a large Büchner funnel to afford a yellow solid. The solid was dried in a drying pistol overnight [acetone (56 ºC, 30 mm)], to yield 19.30 g (73%) of the title compound in greater than 90% purity; this material, containing traces sodium carbonate and sodium bicarbonate, is sufficiently pure for most purposes. For the kinetic studies, the salt was recrystallized as follows. The yellow solid was dissolved in 40 mL of water at 90 °C, the hot solution was quickly filtered, cooled to room temperature and then to –10 °C in an ice / salt bath. The cold slurry was filtered through a Büchner funnel and the resulting yellow solid was rinsed with 75 mL of dry diethyl ether and dried overnight in a drying pistol [acetone (56 ºC, 30 mm)] to yield 11.04 g (55%) of the title compound as a yellow crystalline solid that was powdered with a mortar and pestle and stored under argon in a plastic bottle: mp 160 °C (dec); ^{1}H NMR (400 MHz, 70% CF_3COOH / 30% $CDCl_3$) δ 1.64 (s, 6H), 2.03 (s, 6H), 2.84 (apparent dd, J = 4.0 Hz, J = 14.1 Hz, 2H), 2.97 (t, J = 13.1 Hz, 2H), 5.37–5.47 (m, 1H), 8.12 (d, J = 7.2 Hz, 1H); ^{13}C NMR (100 MHz, 70% CF_3COOH / 30% $CDCl_3$) δ 28.4, 32.4, 41.2, 43.3, 104.3, 115.5 (q, $J_{C\text{-}F}$ = 285.4 Hz), 160.3 (q, $J_{C\text{-}F}$ = 39.6 Hz); ^{19}F NMR (376 MHz, 70% CH_3COOH / 30% $CDCl_3$) δ –75.5 (3F), –148.1 (4F); HRMS (DART-TOF) *m/z* calcd for $C_{11}H_{18}F_3N_2O_2$ $[M - BF_4]^+$ 267.1320, found 267.1336.

1-Hydroxy-2,2,6,6-tetramethyl-4-(2,2,2-trifluoroacetamido)piperidinium tetrafluoroborate (**6**). To a 50-mL round bottomed flask was added 618 mg (1.70 mmol) of 2,2,6,6-tetramethyl-4-(2,2,2-trifluoroacetamido)-1-oxopiperidinium tetrafluoroborate (**5**), 364 mg (6.3 mmol) of allyl alcohol and 30 mL of dichloromethane. The reaction mixture was stirred for 2 h at room temperature until the solution became colorless and gave a negative starch-KI test for the presence of **5**. Solvent and excess alcohol were removed under reduced pressure to afford

537 mg (87%) of the title compound as a white solid: mp 170 °C (dec); ^{1}H NMR (400 MHz, CD_3CN) δ 1.46 (s, 6H), 1.47 (s, 6H), 1.98 (t, J = 13.2 Hz, 2H), 2.21 (apparent dd, J = 2.4 Hz, J = 13.4 Hz, 2H), 4.25–4.35 (m, 1H), 7.66 (br s, 1H), 8.98 (br s, 1H); ^{13}C NMR (100 MHz, CD_3CN) δ 20.1, 28.0, 41.3, 41.7, 70.7, 116.9 (q, J_{C-F} = 287.8 Hz), 157.5 (q, J_{C-F} = 37.4 Hz); ^{19}F NMR (376 MHz, CD_3CN) δ –74.8 (3F), –149.0 (4F); HRMS (DART-TOF) *m/z* calcd for $C_{11}H_{20}F_3N_2O_2$ $[M - BF_4]^+$ 269.1477, found 269.1463.

Rate Studies. Standard solutions of each oxoammonium salt were prepared in dry CH_3CN (0.00500 M for **1** and 0.00502 M for **5**) and their UV-Vis spectra were recorded on a Cary 60 spectrophotometer from Agilent Technologies (version 2.00) scanning from 550–350 nm at a rate of 4600 or 9600 nm/min. The baseline was corrected prior to obtaining readings. The molar absorptivity of each salt was calculated from the average of five individual determinations of the absorbance of these standard solutions at 10 nm intervals from 370–530 nm. This allowed for the calculation of an expected absorbance value for the mixture of two salts. During a kinetic run, a least-squares minimization of the difference between the expected absorbance and the experimentally measured absorbance at 17 individual wavelengths between 370 and 530 nm using an iterative procedure described by Harris[14] and implemented using the "Solver" function in Excel provided the concentration of each oxoammonium salt as a function of time.

A typical experiment was conducted as follows. A dry 100-mL volumetric flask was charged with 147.2 mg (0.4905 mmol) of Bobbitt's salt (**1**), 168.3 mg (0.4753 mmol) of the 4-trifluoromethylacetamido derivative (**5**), and dry CH_3CN. The flask was sonicated for 2-3 min to ensure the oxoammonium salts had completely dissolved and additional CH_3CN was added to give 100.0 mL of solution. The UV-Vis spectrum was taken of this stock solution to measure an

initial absorbance reading. An accurately determined quantity of the alcohol substrate (~0.50 mmol) was added to a clean, dry 3.5 mL x 10 mm x 10 mm quartz cuvette, the cuvette was placed in the UV-Vis spectrophotometer and an accurately measured 2.50 mL of the stock solution was quickly added to the cuvette. The UV-Vis spectrum was recorded for a given period of time (5.76 s – 300 s) depending on the substrate and rate of oxidation. The first UV-Vis spectrum recorded was treated as the initial reading (t = 0 s). Any inconsistencies between the initial absorbance readings of the oxidation and those of the stock solution were accounted for by subtracting the two, and then subtracting this difference from all absorbance readings measured to afford a corrected absorbance reading at each time (t). This corrected reading was used in the least squares analysis described by Harris.[14] UV-Vis spectra as well as the plots of ln ([**1**]) vs. time (s) and ln ([**5**]) vs. time for all substrates investigated can be found in the Supporting Information; the rate of oxidation was determined from the slope of the linear least squares fit of these data.

Solubility Studies (Table 1). The oxoammonium salts were recrystallized from water and dried for 12 h in an Abderhalden (56 °C, 1 mm, $CaCl_2$) prior to use. The solvent of interest was measured into a volumetric flask and the oxoammonium salt was then added portion-wise to the flask, while noting the amount added each time. Periodically, the flask was sonicated for a few s to ensure complete dissolution at room temperature (23 °C). This process was repeated until no more oxoammonium salt appeared to dissolve. The flask was left standing for 2–3 h, at which point, if solid remained the maximum solubility was determined to be reached; if not, the process was repeated. Solvent was then decanted into a tared round-bottomed flask, the solvent was removed under reduced pressure, and the mass of oxoammonium salt was recorded. This mass was divided by the volume of solvent to afford the solubility.

ASSOCIATED CONTENT

Supporting Information

UV-Vis spectra of all oxidations; plots of ln ([**1**]) vs. time (s) and ln ([**5**]) vs. time (s) for all oxidations; NMR spectra of all products. This material is available free of charge via the Internet at http://pubs.acs.org

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ACKNOWLEDGEMENTS

The authors wish to dedicate this work to Professor James M. Bobbitt with thanks for helpful discussions and for being an excellent mentor. The work conducted at the University of Connecticut was supported by grants from Procter and Gamble Pharmaceuticals, Mason, OH; the University of Connecticut's Office of Undergraduate Research (ZDS); and the National Science Foundation (REU site grant 1359081; SMK). The authors would like to thank Drs. Jie He and Jing Zhao for use of their UV-Vis spectrophotometers, and Dr. Christopher Kelly for helpful discussions.

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