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Equilibrium Acidities and Homolytic Bond Dissociation Energies of N−H and/or O−H Bonds in *N*-Phenylhydroxylamine and Its Derivatives

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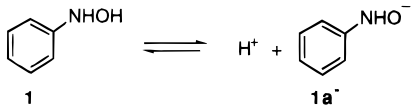
Abstract: The equilibrium acidities (pK_{HA} values) in DMSO of the following hydroxylamines have been measured: (1) *N*-phenylhydroxylamine and its *p*-bromo and *p*-cyano derivatives, (2) *N*-benzyl-*N*-phenylhydroxylamine and its *p*-bromo and *p*-cyano derivatives, (3) *O*-benzyl-*N*-phenylhydroxylamine, (4) *N*-benzoylphenylhydroxylamine and its *p*-bromo and *p*-cyano derivatives, and (5) *N*-hydroxylpiperidine. The BDEs of the O−H and/or N−H bonds in these 11 weak acids have been estimated by combining their pK_{HA} values with the oxidation potentials of their conjugate bases according to the following equation: $BDE_{HA} = 1.37pK_{HA} + 23.1E_{ox}(A^-) + 73.3$ kcal/mol.

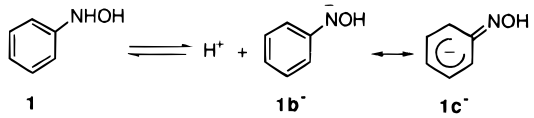
Introduction

A comparison of the acidities of six series of analogous oxygen, nitrogen, and carbon acids in DMSO solution and the gas phase has shown that an intrinsic element effect usually causes nitrogen acids to be more acidic than carbon acids by an average of 17 ± 5 kcal/mol and oxygen acids to be more acidic than nitrogen acids by a like amount.[1] (Henceforth, kcal/mol will be abbreviated as kcal.) It is conceivable, then, that either (a) an element effect could cause *N*-phenylhydroxylamine (**1**) to ionize as an oxygen acid in DMSO, i.e.,



1 **1a**⁻

or (b) that the delocalization of the negative charge in the $PhN(OH)^-$ nitranion into the phenyl ring could cause it to ionize as a nitrogen acid, i.e.,



1 **1b**⁻ **1c**⁻

One purpose of the present paper was, therefore, to compare the heterolytic cleavages of the N−H and O−H bonds (acidities) in PhNHOH and its derivatives to see whether it ionizes in DMSO as an N−H acid, an O−H acid, or both an N−H and O−H acid.

Benzoylation of PhNHOH gives *N*-phenylbenzohydroxamic acid, PhCON(Ph)OH. Both hydroxylamines[2] and their acyl derivatives, hydroxamic acids,[3] are known to give strongly stabilized nitroxyl radicals, $PhN(R)O^\bullet$ and $PhCON(Ph)O^\bullet$, on homolytic cleavage of their O−H bonds. A second purpose of the present paper was to compare the stabilities of these two

[⊗] Abstract published in *Advance ACS Abstracts*, September 1, 1996.

(1) Bordwell, F. G.; Fried, H. E.; Hughes, D. L.; Lynch, T.-Y.; Satish, A. V.; Whang, Y. E. *J. Org. Chem.* **1990**, *55*, 3330−3336.

(2) (a) Sutherland, J. O. In *Comprehensive Organic Chemistry*; Barton, D., Ollis, W. P., Eds.; Pergamon Press: New York, 1979; Vol. 22, pp 185−218. (b) Aurich, H. G. In *The Chemistry of Amino, Nitroso and Nitro Compounds and Their Derivatives*; Patai, S., Ed.; Wiley: New York, 1987; Part I, p 565.

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Table 1. Equilibrium Acidities of *N*-Phenylhydroxylamines, *N*-Phenylbenzohydroxamic Acids, and Benzyl Alcohols at 25 °C Determined by the Overlapping Indicator Method[a]

no.	acid	indicator (pK_{In})[b]	$pK_{HA\pm SD}$[c]
1.	PhNHOH	TNH (24.3)	24.2 ± 0.1
2.	*p*-BrC_6H_4NHOH	9-methylfluorene (22.6)	22.9 ± 0.07
3.	*p*-NCC_6H_4NHOH	CNAH (18.9)	18.6 ± 0.1
4.	PhN(Bz)OH	TNH (24.3)	23.87 ± 0.08
5.	PhNHOBz	TNH (24.3)	23.48 ± 0.05
6.	*p*-$BrC_6H_4N(Bz)OH$	$PhSO_2Bz$ (23.43)	22.7 ± 0.1
7.	*p*-$NCC_6H_4N(Bz)OH$	*p*-$CF_3C_6H_4CH_2SO_2Ph$ (20.2)	19.4 ± 0.1
8.	PhCON(Ph)OH	*p*-$NCC_6H_4CH_2SO_2Ph$ (18.5)	19.0 ± 0.1
9.	PhCON(*p*-BrC_6H_4)OH	Ph_2CHCN (17.5)	17.85 ± 0.1
10.	PhCON(*p*-NCC_6H_4)OH	$(MeSO_2)_2CH_2$ (15.01)	14.65 ± 0.1
11.	*c*-$C_5H_{10}NOH$	TH (30.6)	29.80 ± 0.01
12.	$C_6H_5CH_2OH$	PXH (27.90)	26.93 ± 0.01
13	*p*-$ClC_6H_4CH_2OH$	MCLPXH (26.6)	26.92 ± 0.03
14	*p*-$CF_3C_6H_4CH_2OH$	MCLPXH (26.6)	26.71 ± 0.03

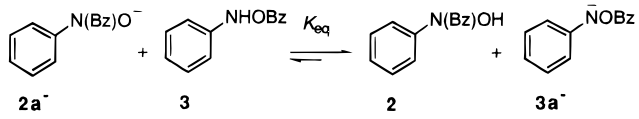
[a] See ref 4 and Experimental Section. [b] Indicators or standard acids used with their pK_a values: TNH = 1,1,3-triphenyl-2-azapropene, CNAH = 4-chloro-2-nitrophenol, TH = triphenylmethane, PXH = 9-phenylxanthene, MCLPXH = 9-(*m*-chlorophenylxanthene. [c] Measured pK_a corrected for homo-H-bonding[5] and corresponding standard deviation.

types of nitroxyl radicals by measuring the BDEs of the O−H bonds in their N−O−H precursors.

Results and Discussion

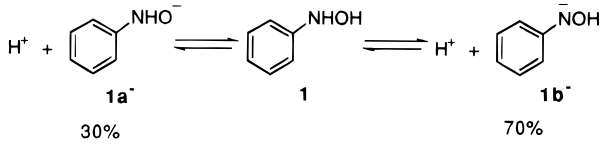
Comparison of the Acidities of the N−H and O−H Bonds in *N*-Phenylhydroxylamine and Its Derivatives. The equilibrium acidities of *N*-phenylhydroxylamine and some of its derivatives and related compounds that are summarized in Table 1 provide the information needed to answer the question concerning the relative N−H and O−H acidities of *N*-phenylhydroxylamine.

The pK_{HA} values for *N*-benzyl-*N*-phenylhydroxylamine, PhN(Bz)OH (**2**), an O−H acid, and *O*-benzyl-*N*-phenylhydroxylamine, PhNHOBz (**3**), an N−H acid (entries 4 and 5, respectively, in Table 1), differ but little, the N−H acid being only about 0.4 pK_{HA} unit stronger than the O−H acid. Both are slightly more acidic than PhNHOH (**1**) (entry 1), an effect attributable to the electron-withdrawing field/inductive effect of the benzyl group.[6] These results indicate that in DMSO solution the following equilibrium will be established.



K_{eq} ~ 2.5 in DMSO solution

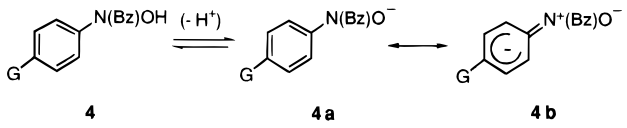
If the acidity difference of N−H bond and O−H bond in PhNHOH can be assumed to be approximately equal to the acidity difference of N−H bond in PhNHOBz and O−H bond in PhN(Bz)OH, i.e., 0.4 pK_{HA} unit, the true pK_{HA} values of N−H and O−H bonds in PhNHOH can be estimated to be 24.3 (= 24.2 + log 3.5/2.5) and 24.7 (= 24.2 + log 3.5/1), respectively, i.e., about 30% of **1a** and 70% of **1b** will be present at equilibrium:



This is consistent with the observation that PhNHOH failed to exibit appreciable homo-H-bonding during the titration, which indicates it acts more like an N−H acid than an O−H acid. (Experiments have shown that anilines fail to exibit homo-H-bonding,[7] in contrast to O−H acids, which exibit strong homo-H-bonding[5]).

The conclusion that PhNHOH behaves primarly as an N−H acid was supported further by the similarity of the effects of *p*-Br and *p*-CN substituents on the acidities of PhNHOH and $PhNH_2$, thus, the slope of a three-point Hammett plot observed for the pK_{HA} values for *p*-GC_6H_4NHOH vs σ_p^-,[6] with G = H, Br, and CN, gave a rho (ρ) value of 5.6, which is the same as the ρ value obtained for a plot of the pK_{HA} values for anilines in DMSO[7] vs σ_p^-.

The slope of a plot of the pK_{HA} values for *p*-$GC_6H_4N(Bz)OH$ acids vs σ_p^- is not much smaller (ρ = 4.5) despite the separation of the acidic site from the benzene ring by an intervening nitrogen atom. In contrast, the damping effect of the CH_2 group in $GC_6H_4CO_2H$ and $GC_6H_4CH_2CO_2H$ reduces the Hammett ρ value from 1.0 to 0.43.[8] The unexpectedly high sensitivity of the pK_{HA}'s of *p*-$GC_6H_4N(Bz)OH$ acid to *p*-substituent effects may be due (a) to better transmission of charge by N than CH_2 or (b) to four-electron repulsions in the *p*-$GC_6H_4N(Bz)O^-$ anion (**4a** ↔ **4b**), that encourages delocalization of the electrons on nitrogen into the substituent-bearing benzene ring.



Comparison of the BDEs of the N−H and O−H Bonds in *N*-phenylhydroxylamine with Those in Anilines and Alcohols. The literature values of 88,[9a] 86,[9b] and 89[9c] (except 90.6[9d]) kcal for the N−H bond in aniline are not consistent with a recent value of 87.3 kcal[10] reported for diphenylamine, because the substitution of a phenyl group for a hydrogen atom on nitrogen in aniline is expected to cause a sizable decrease in BDE. Our BDE for aniline of 92.4 kcal, estimated by eq 1, is the average of the BDEs of aniline itself (92.3 kcal) and its *p*-Me (92.0 kcal), *p*-Cl (92.5 kcal), *m*-Cl (92.5 kcal), and *p*-Br (92.4 kcal) derivatives.[11]

$$BDE_{HA} = 1.37pK_{HA} + 23.1E_{ox}(A^-) + 73.3 \text{ kcal/mol} \quad (1)$$

(3) (a) Bordwell, F. G.; Harrelson, J. A., Jr.; Lynch, T. Y. *J. Org. Chem.* **1990**, *55*, 3337−3341. (b) Perkins, M. J.; Berti, C.; Brooks, D. J.; Grierson, L.; Grimes, J. A.-M.; Jenkins, T. C.; Smith, S. L. *Pure Appl. Chem.* **1990**, *62*, 195−200, and references cited therein.

(4) (a) Matthews, W. S.; Bares, J. E.; Bartmess, J. E.; Bordwell, F. G.; Cornforth, F. J.; Drucker, G. E.; Margolin, Z.; McCallum, R. J.; McCollum, G. J.; Vanier, N. R. *J. Am. Chem. Soc.* **1975**, *97*, 7006−7014. (b) Bordwell, F. G.; McCallum, R. J.; Olmstead, W. N. *J. Org. Chem.* **1984**, *49*, 1424−1427.

(5) Olmstead, W. N.; Margolin, Z.; Bordwell, F. G. *J. Org. Chem.* **1980**, *45*, 3295−3299.

(6) Hansch, C.; Leo, A.; Taft, R. W. *Chem. Rev.* **1991**, *91*, 165−195.

(7) Bordwell, F. G.; Algrim, D. J. *J. Am. Chem. Soc.* **1988**, *110*, 2964−2968.

(8) Hammett, L. P. *Physical Organic Chemistry*, 2nd ed.; McGraw-Hill: 1970; p 367.

(9) (a) McMillen, D. F.; Golden, D. M. *Annu. Rev. Phys. Chem.* **1982**, *33*, 483−532. (b) Colussi, A. J.; Benson, S. W. *Int. J. Chem. Kint.* **1978**, *10*, 1139−1149. (c) Jonsson, M.; Lind, J.; Eriksen, T. E.; Merenyi, G *J. Am. Chem. Soc.* **1994**, *116*, 1423. (d) Denisov, E. T.; Khudyakov, I. V. *Chem. Rev.*, **1987**, *87*, 1313.

(10) Varlamov, V. J.; Denisov, E. T. *Dokl. Akad. Nauk. USSR* **1987**, *293*, 126−128.

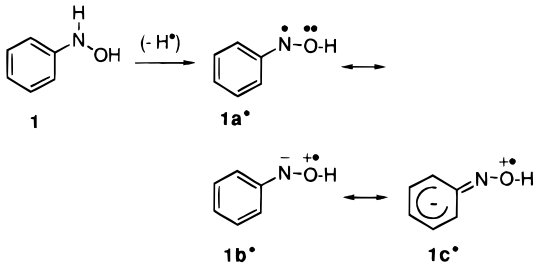
Table 2. Homolytic Bond Dissociation Energies of *N*-Phenylhydroxylamines and *N*-Phenylbenzohydroxamic Acids

no.	acid	pK_{HA}	$E_{ox(A^-)}$[a]	BDE_{HA}
1.	PhNHOH	24.2	−1.255	77.5
2.	*p*-BrC_6H_4NHOH	22.9	−1.146	78.2
3.	*p*-NCC_6H_4NHOH	18.58	−0.898	78.0
4.	PhN(Bz)OH	23.87	−1.354[b]	74.8
5.	PhNHOBz	23.48	−1.284	75.8
6.	*p*-$BrC_6H_4N(Bz)OH$	22.7	−1.259[b]	75.4
7.	*p*-$NCC_6H_4N(Bz)OH$	19.4	−0.976[b]	77.3
8.	PhCON(Ph)OH	19.0	−0.649[b]	84.2
9.	PhCON(*p*-BrC_6H_4)OH	17.85	−0.600[b]	84.0
10.	PhCON(*p*-NCC_6H_4)OH	14.65	−0.395[b]	84.2
11.	*c*-$C_5H_{10}NOH$	29.80	−1.610[c]	77.0

[a] Irreversible oxidation potentials measured in DMSO with 1−3 mM concentrations of HA and 0.1 M $Et_4N^+BF_4^-$ electrolyte at a sweep rate of 100 mV s^{-1}, and referred to the ferrocene/ferrocenium couple unless otherwise indicated. [b] Partially reversible at 10 V/s. [c] *n*-$Bu_4N^+PF_6^-$ as electrolyte.

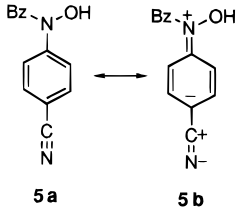
Our BDE for diphenylamine of 87.4 kcal is the average of the BDEs of diphenylamine itself (87.5 kcal) and its *p*-Me (86.9 kcal), *m*-Me (87.6 kcal), and *p*,*p*′-diBr (87.8 kcal) derivatives plus iminodibenzyl (87.0 kcal). The 5 kcal difference in the average values for PhNH−H and Ph_2N−H agrees well with the 6 kcal difference between the BDEs of $PhCH_2$−H (88 kcal) and Ph_2CH-H (82 kcal).[11]

The BDE of 75.8 kcal for the N−H bond in PhNHOBz (**3**) and the BDE of ∼77.5 kcal for the N−H bond in PhNHOH (**1**) (Table 2) are about 16 kcal weaker than the N−H bond in aniline. This is attributed to stablization of the corresponding radical by delocalization of the odd electron (e.g., **1a**•− **1c**•).

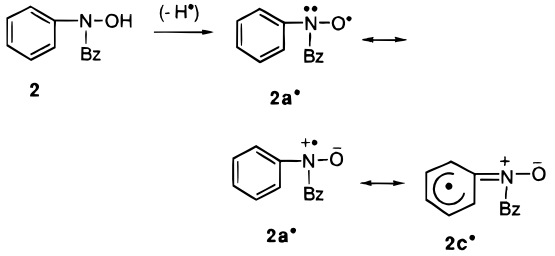


The BDEs of the O−H bonds in hydroxylamines can also be estimated by eq 1, but these estimates may need to be corrected for a solvation effect. Griller *et al.* have suggested that in dipolar solvents the BDEs of O−H bonds, such as that in phenol, will be higher than the gas-phase BDE by an amount equal to that required to break the polar bonds to the solvent.[14] Examination of the literature has shown that the strengths of the solvent hydrogen-bonds to the O−H bonds in benzoic acid, acetic acid, phenol, and *tert*-butyl alcohol are linearly related to their pK_{HA} values in DMSO.[15] The hydrogen bond strength of PhN(Bz)OH (**2**) in DMSO (pK_{HA}= 24) should on this basis be about midway between that of phenol, 7.2 kcal[15a] (pK_{HA}= 18) and *tert*-butyl alcohol, 4.7 kcal[15a] (pK_{HA}= 32.2), i.e., about 5−6 kcal. The BDE of the O−H in PhN(Bz)OH (**2**) estimated to be 74.8 kcal by eq 1 could, on this basis, be as low as 70 kcal.

The BDEs of the O−H bonds in the *p*-Br and *p*-CN derivatives are 0.6 and 2.5 kcal higher than that of **2**. Examination of the two terms in eq 1 shows that structural changes in a series of weak acids can cause an increase in the BDE either (a) by increasing the pK_{HA} or (b) by causing an anodic (positive) shift in $E_{ox}(A^-)$. The *p*-Br and *p*-CN substituents both cause a *decrease* in pK_{HA} accompanied by anodic shift in $E_{ox}(A^-)$. For the stronger electron-withdrawing *p*-CN group, the pK_{HA} decrease is 4.5 pK_{HA} units (6.1 kcal) and anodic $E_{ox}(A^-)$ shift is 378 mV (8.7 kcal). The 2.5 kcal increase in BDE is therefore caused by the anodic shifts in $E_{ox}(A^-)$, which overshadows the decrease in pK_{HA}. Introduction of a *p*-CN group into phenol, thiophenol, or aniline also causes increases in BDE. These increases have been attributed to decrease in ground-state energies resulting from the stablizing effects of dipole−dipole interactions of the *p*-CN group with the acidic function,[16] following the lead of Clark and Wayner.[17,18] This rationalization can also be applied to the present case although the O−H group is not directly involved, (**5a** and **5b**) (the remote substituent effect on the highly stablized nitroxide radical is expected to be relatively small).



The BDE of the O−H bond in *N*-benzyl-*N*-phenylhydroxylamine is weaker than that in an analogous alcohol, $C_6H_5CH_2$-OH,[9a] by about 30 kcal, which means that the delocalizing effect of an adjacent nitrogen in weakening the O−H bond is much greater than the weakening of an adjacent N−H bond by an O−H group (16 kcal). This is consistent with the much greater donor properties of nitrogen than oxygen together with the much greater ability of oxygen than nitrogen to accommodate a negative charge.



Entry 11 in Table 1 shows that the equilibrium acidity (pK_{HA}) of the O−H bond in *N*-hydroxypiperidine (**6**) is 29.8 in DMSO, which is the same as that of the O−H bond in ethanol and only

(11) Bordwell, F. G.; Cheng, J.-P.; Ji, G.-Z.; Satish, A. V.; Zhang, X.-M. *J. Am. Chem. Soc.* **1991**, *113*, 9790−9795. (Originally the $E_{ox}(A^-)$ values were referenced to the Ag/AgI electrode and the constant 56 kcal was used; when referenced to the Fc/Fc^+ couple the constant becomes 73.3 kcal). Irreversible $E_{ox}(A^-)$ values were usually found to agree with reversible $E_{1/2}$ values within 100 mV or better when such comparsons are possible.[12] In this and subsequent papers,[13] BDEs using irreversible $E_{ox}(A^-)$ values have been shown to give agreement with gas phase BDEs to ±3 kcal/mol, or better in most instances (the error between relative values in a series of closely related compounds will be much smaller, *ca.* ±0.5 kcal/mol experimental error).

(12) (a) Venimadhavan, S.; Amarnath, K.; Harvey, N. G.; Arnett, E. M. *J. Am. Chem. Soc.* **1992**, *114*, 221. (b) Bordwell, F. G.; Harrelson, J. A.; Lynch, T.-Y. *J. Org. Chem.* **1990**, *55*, 3337−3341.

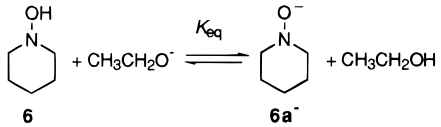
(13) (a) Bordwell, F. G.; Zhang, S. *J. Am. Chem. Soc.* **1995**, *117*, 4858−4861. (b) Bordwell, F. G.; Satish, A. V.; Zhang, S.; Zhang, X.-M. *Pure Appl. Chem.* **1995**, *67*, 735−740 and references cited therein.

(14) (a) Kanabus-Kaminska, J. M.; Gilbert, B. C.; Griller, D. *J. Am. Chem. Soc.* **1989**, *111*, 3311. (b) Wayner, D. D. M.; Lusztyk, E.; Page, D.; Ingold, K. U.; Mulder, P.; Laarhonen, L. J. J.; Aldrich, H. S., *J. Am. Chem. Soc.* **1995**, *117*, 8737−8744.

(15) (a) Arnett, E. M.; Jaris, L.; Mitchell, E. J.; Murty, T. S. S. R.; Gorrie, T. M.; Schleyer, P. v. R. *J. Am. Chem. Soc.* **1970**, *92*, 2365. (b) Rosenberg, M. S.; Iogansen, A. V.; Mashkovsky, A. A.; Odinokov, S. E. *Spectroscopy Lett.* **1972**, *5*, 75−80.

(16) Bordwell, F. G.; Zhang, X.-M.; Satish, A. V. *J. Am. Chem. Soc.* **1994**, *116*, 6605−6610.

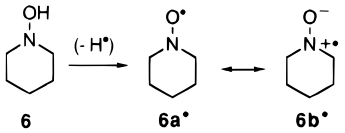
0.4 pK_{HA} unit greater than that of the O−H bond in isopropyl alcohol.[5] It was at first sight surprising to see that the greater electronegativity of the nitrogen atom bonded to oxygen in **6** had not induced a greater acidity on the O−H bond than that caused by the carbon atom joined to the oxygen atom in ethanol.



K_{eq} ~ 1 in DMSO solution

It follows that an acid-weakening effect must be present in **6** that offsets this N vs C (acid strengthening) element effect. The four-electron repulsion in the N−OH moiety is expected to increase the ground-state energy of the undissociated acid (**6**), relative to that in ethanol, which should be acid strengthening, but the increase in ground-state energy in the anion, **6a**$^-$, will be greater than that in the undissociated acid because of the negative charge, and this acid-weakening effect evidently offsets the nitrogen vs carbon element effect.

N-Hydroxypiperidine **6** (entry 11) has a pK_{HA} 5.6 unit (7.7 kcal) higher than that of PhNHOH and 6.3 pK_{HA} unit (8.7 kcal) higher than that of PhN(Bz)OH, but the BDE of its O−H bond is almost identical with that of PhNHOH and only 2.2 kcal higher than that of PhN(Bz)OH. This means that, although the O−H bond in **6** is 7.7−8.7 kcal stronger toward heterolytic cleavage than the O−H bond in **1** or **2**, the strength of the O−H bond in **6** is not much different from that of the O−H bonds in **1** or **2** toward homolytic cleavage.



The Effect of *N*-Benzoylation on the BDE of the O−H Bond in Phenylhydroxylamine. Examination of entry 8 in Table 2 shows that *N*-benzoylation of PhNHOH to give *N*-phenyl-benzohydroxamic acid decreases the pK_{HA} by 5.2 units (7.1 kcal) but *increases* the BDE by about 6.5 kcal caused by a 14 kcal anodic shift in $E_{ox}(A^-)$. The effect of substituting the benzoyl group on nitrogen contrasts sharply with the effect of substituting a benzyl (Bz) group which has but little effect on the pK_{HA} of PhNHOH and slightly weakens the O−H bond. The powerful electron-withdrawing effect of the carbonyl group is of course responsible. It mitigates the donor properties of nitrogen and introduces a strong acid-strengthening field/inductive effect. The strength of the O−H bond is increased homolytically as a consequence of the decreased ground-state energy of the molecule but still remains weaker than the O−H bond in alcohols by about 20 kcal. The presence of a *p*-CN group in the phenyl ring of the PhCON(*p*-NCC$_6$H$_4$)OH increases the acidity of the O−H group by about 6 kcal but has no effect on the homolytic bond energy.

Summary and Conclusions

The N−H bond in PhNHOH (**1**) is estimated to have a pK_{HA} of 24.3 compared to 24.7 for the O−H bond, based on the relative acidities of PhN(Bz)OH (**2**) and PhNH(OBz) (**3**). The acidity of the N−H bond in **1** is about 6.4 pK_{HA} units (8.8 kcal) stronger than the N−H bond in aniline, and the acidity of the O−H bond in **1** is about 3 pK_{HA} units (4 kcal) stronger than that in benzyl alcohol. The BDE of the N−H bond in **1** is about 15 kcal weaker than that in aniline and the O−H bond in **1** is about 30 kcal weaker than that in benzyl alcohol. In other words an adjacent nitrogen atom weakens an O−H bond by about 15 kcal/mol more than an adjacent oxygen atom weakens an N−H bond. The O−H bond in *N*-benzylphenylhydroxylamine (**2**) is about 10 kcal weaker than that in *N*-phenylbenzohydroxamic acid.

Experimental Section

NMR spectra were recorded on a Gemini XL-300 (300 MHz) or XLA 400 (400 MHz) spectrometer. Melting points were measured on a Thomas Hoover capillary melting point apparatus and are uncorrected.

Equilibrium acidities in DMSO were determined by the overlapping indicator method[4] at ambient temperatures and are summarized in Table 1. The concentrations of the acid and the indicator used in pK_{HA} measurements were typically between 0.5−5.0 mM. The pK_{HA} and pK_{AHA^-} values should be independent of the concentrations used.[4,5] All the pK_{HA} measurements were based on 2−4 titrations; each had 3−5 points with different HA/A$^-$ ratios. Measurement of pK_{AHA^-} value requires at least 5 points with HA/A$^-$ ratio within a range from 1:3 to 3:1.[5]

The O−H acids were found to form strong homo-hydrogen-bonds (the observed pK_{HA} values for an acid will depend on the ratio of the acid and its conjugate anion if the acid and its conjugate anion form strong homo-H-bond[5]), e.g., the homo-H-bonding constant (pK_{AHA^-}) for PhN(OH)Bz, *c*-C$_5$H$_{10}$NOH, PhCH$_2$OH, *p*-ClC$_6$H$_4$CH$_2$OH, and *p*-CF$_3$-C$_6$H$_4$CH$_2$OH were found to be 2.97 ± 0.01, 3.27 ± 0.03, 3.65 ± 0.02, 3.73 ± 0.12, and 3.59 ± 0.11, respectively. Since the corresponding program was not available for titrations requiring the use of standard acids, the pK_{AHA^-} values for *p*-G-C$_6$H$_4$N(Bz)OH (G = Cl, CN) and *p*-G-C$_6$H$_4$N(COPh)OH (G = H, Cl, CN) acids were not obtained. The effect of homo-hydrogen-bonding on pK_{HA} values for these acids was avoided by measuring the pK_{HA} values at points where the ratios of the concentration of the acid and the conjugate anion were close to 1:1.[5] The N−H acids did not exibit homo-H-bonding as expected,[7] e.g., the pK_{AHA^-} values for PhNHOBz were found to be less than 0.5.

Sutherland reported that hydroxylamines are quite unstable in alkaline aqueous solutions.[19] However, we found that the hydroxylamines and hydroxamic acids we used were stable enough to perform titrations and oxidation potential measurements in DMSO solution.

Oxidation potentials were measured by a conventional cyclic voltammetric instrument, as described previously.[11] The working electrode (BAS) consists of a 1.5 mm diameter platinum disk embedded in a cobalt glass seal. It was polished with 0.05-μm Fisher polishing alumina and rinsed with ethanol and dried before each run. The counter electrode was platinum wire (BAS). The reference electrode was Ag/AgI, and the redox potentials reported were referenced to ferrocene/ferrocenium couple. Tetraethylammonium tetrafluoroborate or tetra-*n*-butylammonium hexafluorophosphate was used as the supporting electrolyte. The electrochemical experiments were carried out under an argon atmosphere.

Material. Nitrobenzene, *p*-bromonitrobenzene, *p*-cyanonitrobenzene, *N*-hydroxylpiperidine were purchased from Aldrich and used as received. *N*-Phenylbenzohydroxamic acid (Eastman Organic Chemicals) was purified by recrystallization from benzene−hexane solution. Tetraethylammonium tetrafluoroborate was recrystallized from absolute alcohol and dried at 110 °C at 0.1τ for 24 h.

Substituted *N*-phenylhydroxylamines were prepared by a modified method of Crumbliss and Brink.[20] Typically, into an aqueous solution (30 mL) of ammonium chloride (20 mmol) was added a hot alcohol solution (20 mL) of substituted nitrobenzene (20 mmol), and the mixture was stirred vigorously to form a milky suspension. Then 22 mmol of

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(18) Rüchardt, C. *Angew. Chem., Int. Ed. Engl.* **1970**, *9*, 830−843 and Nicholas, A. M. P.; Arnold, D. R. *Can. J. Chem.* **1984**, *62*, 1850 have also called attention to the importance of changes in ground state energies in affecting bond dissociation energies.

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zinc dust was added during 20 min at 70 ± 5 °C, and the mixture was stirred for an additional 20 min. The zinc oxide formed was removed by filtration, and the resulting solution was extracted with 30 mL of ether. The organic layer was dried with $CaCl_2$. After the solvent was stripped off, the crude substituted *N*-phenylhydroxylamine was recrystallized from benzene−hexane solution and dried at 0.1τ / 25 °C for 1 h. The samples thus obtained were found to be stable at −10 °C for several weeks. The yield and the properties of the *N*-phenylhydroxylamines were characterized as follows: PhNHOH, 60%, mp 82 °C (lit.[21] 81−2 °C), ^{1}H NMR ($CDCl_3$): δ 6.5 (2H, br, NHOH), 7.02 (3H, m), 7.30 (2H, m). *p*-BrC_6H_4NHOH, 75%, mp 92 °C (lit.[21]91 °C), ^{1}H NMR ($CDCl_3$): δ 6.1 (2H, br, NHOH), 6.88 (2H, d, J = 8.7 Hz), 7.39 (2H, d). *p*-NCC_6H_4NHOH, 49%, mp 87 °C (lit.[21] 86 °C), ^{1}H NMR ($CDCl_3$): δ 5.6 (1H, s, OH), 7.03 (1H, s, NH), 7.01 (2H, d, J = 8.5 Hz), 7.54 (2H, d).

Substituted *N*-benzyl-*N*-phenylhydroxylamines were prepared by the method of Utzinger.[22] Typically, 5 mmol of substituted *N*-phenylhydroxylamine was dissolved in 5 mL of ether, mixed with 5 mmol of benzyl chloride and 5 mmol pyridine, and allowed to stand at room temperature for 48 h. The resulting solution was mixed with 20 mL of water extracted with 10 mL of ether and dried with $CaCl_2$. After removal of the ether, the crude product was purified by column chromatography (1:5 ether and hexane as eluent). The properties and (yield) of the substituted *N*-benzyl-*N*-phenylhydroxylamine were characterized as follows: PhN(OH)Bz (56%), mp 85 °C (lit.[22] 86 °C), ^{1}H NMR ($CDCl_3$): δ 4.39 (2H, s, CH_2), 5.62 (1H, s, OH), 7.04 (1H, t, J = 7.4 Hz), 7.22 (2H, d, J = 7.5 Hz), 7.28−7.45 (7H, m). *p*-$BrC_6H_4N(OH)Bz$ (40%), mp 92−92.5 °C, ^{1}H NMR ($CDCl_3$): δ 4.36 (2H, s, CH_2), 5.46 (1H, s, OH), 7.03 (2H, d, J = 8.7 Hz), 7.15−7.25 (7H, m). Calculated for $C_{13}H_{12}BrNO$: C, 56.14; H, 4.35; N, 5.04. Found: C, 55.90; H, 4.35; N, 5.21. *p*-$NCC_6H_4N(OH)Bz$ (40%), mp 100−1 °C, ^{1}H NMR (C_6D_6): δ 4.1 (2H, s, CH_2), 6.02 (1H, s, OH), 6.68 (2H, d, J = 8.7), 7.0−7.3 (7H, m). Calculated for $C_{14}H_{12}N_2O$: C, 74.98; H, 5.39; N, 12.49. Found: C, 74.85; H, 5.43; N, 12.51.

***N*-Phenyl-*O*-benzylhydroxylamine** was prepared by the method of Perronnet *et al*.:[23] To a DMSO (10 mL) solution of *N*-phenylbenzohydroxamic acid (10 mmol) were added anhydrous K_2CO_3 (20 mmol), and benzyl chloride (15 mmol), and the solution was stirred at room temperature for 48 h. The resulting mixture was poured into 30 mL of water and extracted with 30 mL of ether. The ether layer was washed with 30 mL of water, dried with $CaCl_2$, and concentrated under vacuum. The residue was purified by column chromatography (2:5 ether:hexane) to give 1.5 g (50%) of *O*-benzyl-*N*-phenylhydroxamic acid, mp 85−6 °C, ^{1}H NMR ($CDCl_3$): δ 4.84(2H, s, CH_2), 7.1−7.7(15H, m). A mixture of DMF (6 mL), *O*-benzyl-*N*-phenylhydroxamic acid (1.2 g, 4 mmol) and hydrazine hydrate (1 mL) was stirred at 85 °C for 12 h. The resulting mixture was poured into 20 mL water of and extracted with 20 mL of ether. The ether layer was washed with 20 mL of water, dried with $CaCl_2$, and concentrated under vacuum. The residue was purified by column chromatography (1:5 ether:hexane) to give 0.4 g (50%) of *O*-benzyl-*N*-phenylhydroxylamine, mp ∼ −10 °C, ^{1}H NMR ($CDCl_3$): δ 4.95 (2H, s, CH_2), 6.98 (3H, m), 7.03 (1H, s, NH), 7.31 (2H, m), 7.4−7.55 (5H, m).

Substituted *N*-phenylbenzohydroxamic acids were prepared by the method of Ayyanger *et al*.[24] The identities of the compounds were confirmed by their ^{1}H NMR spectrum. Their yields and the spectral data were as follows: *p*-$BrC_6H_4N(OH)COPh$, 40%, mp 163 °C (lit.[24] 163 °C), ^{1}H NMR ($CDCl_3$): δ 7.08 (2H, d, J = 8.8 Hz), 7.32 (2H, d), 7.42 (5H, m), 9.15 (1H, s, OH). *p*-$NCC_6H_4N(OH)COPh$, 45%, mp 147−8 °C, ^{1}H NMR ($CDCl_3$): δ 7.33 (2H, d), 7.38 (2H, m), 7.47 (3H, m), 7.57 (2H, d, J = 8.5 Hz), 9.15 (1H, s, OH).

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