Spectrophotometric determination of protonation constant of N-phenylbenzohydroxamic acid in mineral acids

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The UV spectra of N-phenylbenzohydroxamic acid (C_6H_5CON(OH)C_6H_5) in sulphuric, perchloric and hydrochloric acid solutions have been analyzed by the characteristic vector method in order to separate protonation from the medium effect. Using the first vector coefficient values, proton at ion constant ($pK_{BH^+}$) and solvation parameters have been estimated. In the high acidity range, results obtained according to the Cox-Yates excess acidity function, Marziano Mc function, Bunnett-Olsen and Hammett-acidity function methods are in good agreement.

The protonation equilibria of polyfunctional acids and bases in concentrated solutions of mineral acids, have been the subject of a number of studies. On the other hand, only scanty data are available for hydroxamic acids both in solution and in the gas phase. N-Phenylbenzohydroxamic acid (PBHA) has been widely employed in analytical chemistry. One of its characteristic properties is that it forms complex compounds in strongly mineral acid media with several highly charged, easily hydrolysed elements, an effect that sharply increases the specificity and selectivity of analytical reactions. For the last few years we have been investigating the hydrolysis of a large number of C- and N-substituted hydroxamic acids. Before any interpretation of these data can be made, it is necessary to correct the constant of hydrolysis for the degree of protonation and solvation parameters of the substrate. Hence precise measurement of protonation constant ($pK_{BH^+}$) of the substrate (Eq. 1) is essential.

$$B + H^+ \xrightleftharpoons{K_{BH^+}} BH^+$$

There is some confusion over what exactly protonation constant, or acidity constant refers to when a compound is potentially amphoteric. As the name implies, on the one hand hydroxamic acids behave like a weak acids ($pK_a = 8-9$) and on the other, like most carbonyl compounds, they are also weak bases ($pK_{BH^+} = -1$ to $-2$). The suppression of acid character may be attributed to the intramolecular hydrogen bonding. Thus, we may refer to $pK_{BH^+}$ of B or $pK_a$ of BH$^+$ (conjugate species). The experimental studies generally rely on indirect methods, like changes in NMR chemical shifts, UV or IR spectral changes and solvent extraction, etc. Protonation studies of carbonyl compounds are handicapped by uncertainties, which arise from the solvent sensitivity (medium effect) of the UV spectra. Various methods of compensating for the medium effect have been advanced, but the most powerful method is the characteristic vector analysis (CVA) or principal component analysis. The encouraging results we obtained using vector analysis procedure for the correction of medium effects in UV spectral data of some hydroxamic acids prompted us to adopt this method for estimating the $pK_{BH^+}$ of N-phenylbenzohydroxamic acid C_6H_5CON(OH)C_6H_5 (PBHA) in sulphuric, perchloric and hydrochloric acids. The Hammett acidity function method (HAFM), the Cox-Yates excess acidity method (EAM), Marziano Mc method and the Bunnett-Olsen (BOM) method have been compared in order to rationalize the differences observed between $pK_{BH^+}$ values estimated by each classical methods.

**Experimental**

PBHA was prepared by standard procedure. Sulphuric, perchloric and hydrochloric acid solutions were prepared by diluting commercial (Qualigens Excelar) concentrated acids and titrated with standard NaOH.

The stock solution of PBHA was prepared in ethanol ($1.5 \times 10^{-3}$ mol dm$^{-3}$). An aliquot of this solution (20 ml) was diluted to 50 ml using distilled water. The samples were prepared by diluting a small portion (1-2 ml) of the above solution of substrate in water with the appropriate mineral acid solutions. Practically no corrections were necessary for the chemical reactions of PBHA in acids during the measurement. The sam-
ples were cooled in an ice bath before and during the addition to prevent the substrate from any hydrolysis effect. The resulting cold solution was then thermostated to 25±0.1°C, and the volumetric flask was eventually filled up to the mark with water. The solutions were transferred into 1 cm quartz cells and the spectra recorded against an acid solution of the same concentration using Unicam UV-2 300 spectrophotometer. The final concentration of PBHA in the medium was 6.0×10⁻⁵ mol dm⁻³.

**Characteristic vector analysis of UV data**

UV data were collected into a matrix, each row corresponding to a complete spectrum at a given acidity and each column to a particular wavelength. The parameters on which pK⁺ BH⁺ was assumed to be dependent are protonation and medium (solvation/H-bonding). This data matrix was then subjected to CVA, according to the procedure originally described by Simonds. The output was a matrix resulting from the coefficients of the first and second characteristic vectors only, a column of which was processed as usual.

**Results and discussion**

The UV spectra of PBHA at different acidities (H₂SO₄ = 0.0-16.2, HClO₄ = 0.0-9.0, HCl = 0.0-9.3 mol dm⁻³) are very similar for all the three mineral acids. However, in the spectrum of PBHA in hydrochloric acid solutions, no significant shifts in λmax is observed with respect to acid concentration.

In perchloric acid at low acidities (water-5 mol dm⁻³), a peak is present at about 256 nm, which shifts to 268 nm in strongly acidic solutions. In sulphuric acid a clear cut bathochromic shift (250 nm to 276 nm) is observed. Protonation and medium effects contribute equally to this spectral variation. Solutions of PBHA in mineral acids are colourless. They undergo no visible changes on standing. In all the three acids the molar extinction coefficients increased successively with increasing acid concentrations: H₂SO₄ (λmax 276) = 5841 to 15094, HClO₄ (λmax 268) = 5807 to 15859, HCl (λmax 262) = 8487 to 12730] mol⁻¹ cm⁻¹ at their respective λmax.

We have already studied kinetics of hydrolysis of PBHA in the mineral acids. For the hydrolysis of PBHA the catalytic order of strong mineral acids is HCl > H₂SO₄ > HClO₄. It may be concluded that strong acids having anions of high charge density (SO₄²⁻, Cl⁻) would be more effective catalysts, i.e., protonating power is greater. It seems probable that an acid having anion of low charge density (ClO₄⁻) stabilizes carbonium ion like transition state. It can be seen from the results that H₂SO₄ is the suitable medium for studying protonation behaviour. The bathochromic shifts indicate that more than one process is occurring; it is apparent that each of the two spectral processes contributes approximately equally to the total spectral variation. No clear isosbestic point has been observed. At low sulphuric acid concentration (water-5.0 mol dm⁻³), no significant differences in the spectra have been observed. At any given acid concentration, ionization ratios (I = [BH⁺]/[BJ]) were calculated from Eq. 2,

\[ I = (A - A_B)/(A_{BH⁺} - A) \]  

where \( A_B \) is the absorbance of free base (molecular species), \( A \) is the absorbance value at intermediate acidity and \( A_{BH⁺} \) is the absorbance of the protonated base. This study allows us to select the most suitable wavelength (280 nm maximum variation for H₂SO₄ and HClO₄ and 260 nm for HCl) and \( A_B \) and \( A_{BH⁺} \) as the limits.

In strong acid media the convention to write protonation equilibria (Eq. 1) for the determination of pK⁺ BH⁺ is as follows (Eq. 3)

\[ pK⁺ BH⁺ = \log I - \log C_H⁺ - \log f_B f_H/f_BH⁺ \]  

where \( C_H⁺ \) is the hydrogen ion concentration and \( f_B f_H/f_BH⁺ \) represents molar activity coefficient ratio. Generally the \( pK⁺ BH⁺ \) values are calculated by Hammett acidity function (HAFM) (Eq. 4) and Bunnett-Olsen method (BOM) (Eq. 5).

\[ \log I = m (-H₀) + pK⁺ BH⁺ \]  
\[ \log I + H₀ = pK⁺ BH⁺ + \phi_e (H₀ + \log C_{H⁺}) \]

where \( H₀ \) is the Hammett acidity function. Bunnett-Olsen equation may also be represented as Eq. 6.

\[ \log I - \log C_{H⁺} = (φ-1) (H₀ + \log C_{H⁺}) + pK⁺ BH⁺ \]
The $pK_{BH+}$ values were obtained by plotting $\log I$ against $H_0$ (HAFM) and ($\log I - \log C_{H^+}$) against ($H_0 + \log C_{H^+}$) (BOM). Thus, only one acidity function ($H_0$) is needed for the purpose of estimating $pK_{BH+}$.

$$\log f_{B^+}/f_{H^+} = M_c \text{ or } X \quad \cdots \ (8)$$

Cox and Yates\(^1\) excess acidity method is denoted by Eq. 7

$$\log I - \log C_{H^+} = m^*X + pK_{BH^+} \quad \cdots \ (7)$$

According to this method $pK_{BH^+}$ values were determined as intercepts by plotting the $\log I - \log C_{H^+}$ against $X$ or $Mc$ values. Theoretical and practical aspects of the above methods have been treated extensively. Marziano et al.\(^18\) demonstrated the differences between activity coefficient function and acidity function, as well as the parameter related to protonating ability of the solvent.

The values of $pK_{BH}$, calculated by the different methods closely resemble each other (Table 1). It is difficult to comment on the validity of each method.

Each method has its equivalent for kinetic studies. In all these plots the standard deviations in slope and intercept were in the range $0.01 < s < 0.11$. The $pK_{BH^+}$ values of PBHA in HCl and HClO$_4$ solution were low. Amongst these acids, PBHA seems to be stronger base in sulfuric acid solutions.

**Characteristic vector analysis**

CV (principle component analysis) analysis can be used to obtain an abstract solution where all data are expressed as linear sums of product terms\(^12\). Thus any absorbance, $A_{n.p}$ may be represented by Eq. 8; in which $c_1$ represents the principal component for wavelength $p$, $v_1$ represents characteristic vector for the $n$th spectrum, subscripts 1 and 2 describe first and second components (weighing factors) and finally $p$ and $n$ describe certain wavelengths and acidities, respectively. $\bar{A}$ is mean absorbance at a given wavelength.

$$A_{n.p} = \bar{A}_{n.p} + c_{1.p} v_{n.1} + c_{2.p} v_{n.2} \quad \cdots \ (8)$$

In the present case the first vector accounted for 94% of the total data variation and the second vector

<table>
<thead>
<tr>
<th>Mineral acid (mol dm$^{-3}$)</th>
<th>HAFM</th>
<th>BOM</th>
</tr>
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<tbody>
<tr>
<td></td>
<td>$H_0$</td>
<td>$H_A$</td>
</tr>
<tr>
<td></td>
<td>$m$</td>
<td>$pK_{BH^+}$</td>
</tr>
<tr>
<td>HCl (0.0-9.28)</td>
<td>0.60</td>
<td>-1.37</td>
</tr>
<tr>
<td>H$_2$SO$_4$ (0.0-16.2)</td>
<td>0.38</td>
<td>-2.12</td>
</tr>
<tr>
<td>HClO$_4$ (0.0-9.00)</td>
<td>0.71</td>
<td>-2.07</td>
</tr>
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</table>

<table>
<thead>
<tr>
<th>Mineral acid (mol dm$^{-3}$)</th>
<th>EAM</th>
<th>MCP</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>$m^*$</td>
<td>$pK_{BH^+}$</td>
</tr>
<tr>
<td>HCl (0.0-9.28)</td>
<td>0.44</td>
<td>-1.39</td>
</tr>
<tr>
<td>H$_2$SO$_4$ (0.0-16.2)</td>
<td>0.31</td>
<td>-2.41</td>
</tr>
<tr>
<td>HClO$_4$ (0.0-9.00)</td>
<td>0.31</td>
<td>-1.99</td>
</tr>
</tbody>
</table>

$r$ = Regression coefficient

HAFM = Hammett acidity function method; BOM = Bunnett-Olsen Method; EAM = Excess acidity method; MCP = Marziano-Cimino method

\(^1\)Values in parentheses indicate data after applying CV analysis.
accounting for 6%. The first vector (principal component) is associated predominately with protonation and the second reflects only the medium effects, i.e., hydrogen bonding and solvation. The spectra reconstructed using vectors and coefficients obtained are in good agreement with the originals. The coefficients of the two vectors are plotted as a function of $-H_0$, (for $H_2SO_4$ only) for illustrative purposes, (Fig 1). Now simple protonation would give a sigmoid titration curve, changing from a constant $C_B$ to a constant $C_{BH^+}$, when plotted in this way. It is evident that $c_1$ behaves like this initially but does not reach a stable final value and is roughly linear, with $-H_0$. $c_2$ has no obvious physical chemical counterpart. Reconstructed spectra show good isosbestic points. The mean curve for the spectrum of PBHA in $H_2SO_4$ is shown in Fig. 2A. There is a contribution at each wavelength from the first characteristic vector, given in Fig. 2 B weighted by the factor $c_1$ from Table 2, and another contribution at this wavelength from the second characteristic vector, given in Fig. 2C, weighted by the factor $c_2$ from Table 2. It should be emphasized that the curves of Fig. 2 are obtained by a purely mathematical processing of the raw data by the methods of vector analysis, with no assumptions about the nature of the effects (other than their orthogonality). From these reconstituted curves, the ionization ratios $[BH^+]/[B]$ at varying acid concentrations could be obtained. These ionization ratios are independent of the wavelength chosen to calculate them, as they should be in the absence of medium effects. It must be noted that our data analysed by all the methods generally gave nearly coincident results. The Bunnett-Olsen and excess acidity methods have been shown to be essentially equivalent, giving the same $pK_{BH^+}$ and $[m^* = (1 - \phi_e)]$ values. Now the question arises; which method should be used? Most of the authors recommend the excess acidity method.

It is well known that $m^*$ measures the solvation requirement of BH+, i.e., the extent to which the positive charge is externally dispersed by interactions with the solvent especially through hydrogen bonding. More practically, $m^*$ measures the rate of increase of log I with acidity. Amides display characteristic $m$

<table>
<thead>
<tr>
<th>$[H_2SO_4]$</th>
<th>$c_1$</th>
<th>$c_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.00</td>
<td>-0.342</td>
<td>-0.479</td>
</tr>
<tr>
<td>1.80</td>
<td>-0.332</td>
<td>-0.226</td>
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<td>3.60</td>
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<tr>
<td>5.40</td>
<td>-0.252</td>
<td>-0.068</td>
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<tr>
<td>7.20</td>
<td>-0.209</td>
<td>0.172</td>
</tr>
<tr>
<td>9.00</td>
<td>-0.011</td>
<td>0.315</td>
</tr>
<tr>
<td>9.90</td>
<td>0.055</td>
<td>0.400</td>
</tr>
<tr>
<td>10.8</td>
<td>0.137</td>
<td>0.499</td>
</tr>
<tr>
<td>14.4</td>
<td>0.278</td>
<td>-0.291</td>
</tr>
<tr>
<td>15.3</td>
<td>0.416</td>
<td>-0.239</td>
</tr>
<tr>
<td>16.2</td>
<td>0.555</td>
<td>-0.185</td>
</tr>
</tbody>
</table>
values (Table 1) which are similar to those reported in the literature for amides\(^\text{19}\). Values in this range are typical for oxygen bases where the positive charge is partly delocalized to heteroatoms. Hydroxamic acids are more acidic than amides. However, the solvent may play an important role in the proton dissociation process. The N-substituent with a negative effect strengthens the acid properties of amides greatly. The electron donating ability of \(\text{C}_6\text{H}_5\) helps to stabilize resonance form of PBHA and BHA by 0.48 \(pK_{\text{BH}^+}\) units. Some of it must be due to higher electronegativity of the \(sp^3\) carbons in BHA, causing electron withdrawal and making protonation more difficult. Benzamide (\(pK_a = 13.0\) and \(pK_{\text{BH}^+} = -1.54\))\(^\text{20}\) is more basic than BHA. Protonated benzamide can partially delocalize its positive charge into the solvating water molecules. This will be more difficult when \(\text{NH}_2\) becomes NHOH, and there is a phenyl group present (RNHOH, PBHA).

**Site of protonation**

The site of protonation of a base or acid possessing more than one possible site may be difficult to determine experimentally. Hydroxamic acids contain three electrophilic centres. The carbonyl oxygen and nitrogen atom are two suitable sites for protonation. The oxygen has two lone pairs; and, unlike the nitrogen’s lone pair, these electrons are not involved in resonance. Protonation at nitrogen destroys hydroxamic acid resonance while protonation at oxygen does not. Thus, the O-protonated intermediate should be a more stable species. Available experimental evidences, and also existing theoretical calculations support this view. The results of the present study support this theory. The present study shows that \(pK_{\text{BH}^+}\) values are independent of the methods used for their calculation but are dependent on mineral acids. Although there are various experimental (NMR, FT–ICR–CD, Raman, solvent partitioning, cryoscopy, electrochemical) and theoretical methods available in the literature for the determination of \(pK_{\text{BH}^+}\), the present UV method is simple rapid and reliable. Characteristic vector analysis may be used to separate medium effects from protonation equilibria (clear isosbestic point). The strength of a polyfunctional base can be expressed in terms of two factors, its \(pK_{\text{BH}^+}\) and sensitivity of the protonated base to the solvating power of the medium.

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**References**