Ionic Characterization of a Synthetic Amphoteric Surfactant: N,N-((Butyloxy)propyl)amino Diacetic Acid

ARTICLE *in* LANGMUIR · AUGUST 1998

Impact Factor: 4.46 · DOI: 10.1021/la980291d

5 AUTHORS, INCLUDING:

Jean-Joseph Max
Université du Québec à Trois-Rivières
55 PUBLICATIONS **1,329** CITATIONS

SEE PROFILE

Gervais Bérubé
Université du Québec à Trois-Rivières
81 PUBLICATIONS **862** CITATIONS

SEE PROFILE

Camille Chapados
Université du Québec à Trois-Rivières
91 PUBLICATIONS **1,500** CITATIONS

SEE PROFILE

Ionic Characterization of a Synthetic Amphoteric Surfactant: *N,N*-((Butyloxy)propyl)amino Diacetic Acid

Jean-Joseph Max, Gervais Bérubé, Michel Trudel, Serge Groleau, and Camille Chapados*

Département de Chimie-Biologie, Université du Québec à Trois-Rivières, Trois-Rivières, Québec, Canada G9A 5H7

Received March 10, 1998. In Final Form: June 4, 1998

N,N-((Butyloxy)propyl)amino diacetic acid is an amphoteric surfactant of the glycinate family. The synthesis of this glycinate was achieved via the ester route to obtain a pure compound. The infrared (IR) spectra of the aqueous solutions of this compound were obtained in the pH range 0.2−14 in order to determine the ionic distribution of the molecule as a function of pH. After subtracting the water bands in the IR spectra, factor analysis (FA) was used to separate the spectra of the ionic species and determine their abundance. The pK_a's were retrieved from the volumetric titration as a function of pH. The values were used to calculate the volumetric abundance of each ionic species as a function of pH. The distribution curves thus obtained were compared with the normalized distribution curves obtained from the IR data. There is a good agreement between the two curves. The following species were obtained for the glycinate in water: the cation (pH 0−3.4), the zwitterion (pH 0−5.4), the monocarboxylate (pH 1−12), and the dicarboxylate (pH 7−14).

1. Introduction

Amphoteric surfactant is a class of surfactant that can dissolve acidic, basic, or neutral substances. Because of this ability, there is a growing demand for this kind of surfactant. The industrial amphoteric surfactants synthesized in water are tinted with side products leading to commercial products that are ill-defined.[1] For instance HOPA [*N,N*-((hexyloxy)propyl)amino diacetic acid, CAS Registry Number 200136-73-4] is an industrial amphoteric surfactant of the glycinate family manufactured by Laboratoires Choisy, Ltée. Since the linear amine is not available commercially, HOPA is made from a mixture of isomeric amines which also contains a small amount of the alcohol used to prepare the amine. HOPA manufactured commercially showed some solubility problems at low pH where, depending on the ionic strength of the solutions, we observed a phase separation. To surmount the difficulties of characterizing a surfactant obtained in an aqueous medium that contains numerous side products, we decided to synthesize the glycinate next to HOPA in this series of amphoteric surfactant. The new compound which is *N,N*-((butyloxy)propyl)amino diacetic acid [hereafter BOPA = AH_2, Figure 1] was synthesized via the organic path, starting from the amine which is available commercially in pure form and using the ester route to obtain a well-characterized pure product. We define the organic path as one where the reactions are made in a nonaqueous media, whereas the industrial surfactants are usually made in an aqueous media. Presenting similarities to the well-known glycine,[2] the neutral BOPA is expected to form a zwitterion in water.

Because of its (butyloxy)propyl aliphatic chain, BOPA is an ideal molecule to be used as a hydrotrope in nonionic and ionic surfactant formulations. This quality comes from the natural ability of amphoteric surfactants and the relatively short hydrocarbon chain of BOPA. We are mainly concerned with the ionic properties of the polar headgroup since it is the site where the action is taking place when the pH is varied. It governs the fundamental properties of this type of molecules. Very little quantitative information is available on this type of molecule.

IR spectroscopic studies of a mixture of dodecyltrimethylammonium chloride and sodium dodecyl sulfate[3]

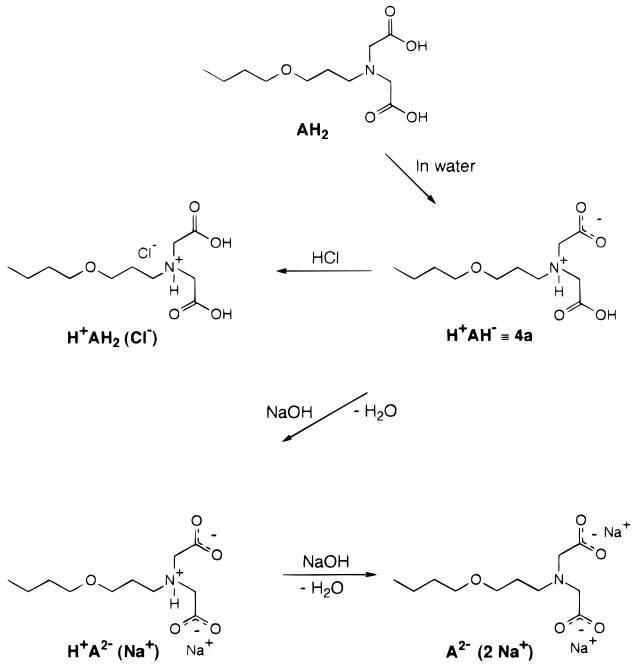

Figure 1. Molecular structures of the ionic species of **4a**: hydrochloride diacid (H^+AH_2); zwitterion (H^+AH^-); sodium monocarboxylate (H^+A^{2-}); and sodium dicarboxylate (A^{2-}).

* Corresponding author. E-mail: Camille_Chapados@uqtr.uquebec.ca.

(1) Ricca, J. M.; Derian, P. J.; Marcenac, F.; Vukov, R.; Tracy, D.; Dahanayake M. *New generation of imidazoline derived amphoteric surfactants*; Special Publications; Royal Society of Chemistry: Letchworth, U.K., 1996; Vol. 187, pp 302−320.

(2) Max, J.-J.; Trudel, M.; Chapados, C. *Appl. Spectrosc.* **1998**, *52*, 226−233.

(3) Scheuing, D. R.; Weers, J. G. *Langmuir* **1990**, *6*, 665−671.

S0743-7463(98)00291-1 CCC: $15.00
Published on Web 08/15/1998

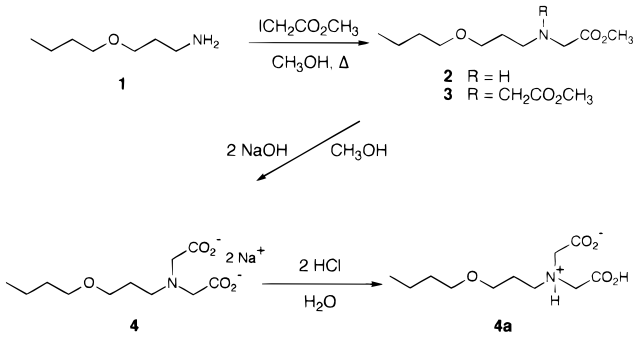

Figure 2. Synthesis of *N*,*N*−(3-(butyloxy)propyl)amino diacetic acid **4a**.

and betaines[4] have previously been conducted. In these studies where the micellar properties of the surfactants have been evaluated, only two ionic situations are encountered. Not so with BOPA which contains four different ionic species and not knowing about the distribution of these species as a function of pH, we may anticipate that some of them may not be encountered in pure form at any pH. To understand properly the solubilizing properties of BOPA, it is necessary to first characterize each ionic species and determine their relative abundance as a function of pH.

The objective of this paper is therefore to determine the ionic structure of BOPA. We will do this by infrared (IR) and volumetric titrations in the pH range 0−14. For the IR titration, we have developed a scheme that combines some theoretical development with precise IR measurements and data treatments. Factor analysis (FA) was used to separate the different spectra of concomitant ionic species encountered throughout the pH range studied. The theory used to obtain "real" spectra with FA is given in ref 5. This method was used successfully with the electronic absorption spectra of amphotericin B to separate the spectra of the aggregated species[6] and with the fluorescence spectra of porphyrins to obtain the spectra of the different ionic species.[7,8] More recently we have used FA on the IR spectra of aqueous glycine to obtain the IR spectra of each ionic species.[2]

2. Theoretical Considerations

When BOPA (AH2) is solubilized in water, it forms the zwitterion H^+AH^-. When 1 equiv of aqueous HCl is added to the latter, it forms the cation $H^+AH_2Cl^-$. When 1 equiv of aqueous NaOH is added to the zwitterion, it forms the monoanion $H^+A^{2-}Na^+$, and with another equivalent of NaOH it forms the dianion $A^{2-}(Na^+)_2$. The different ions are presented in Figure 1. The counterions indicated in parentheses in the figure, although present with the parent ions, have been omitted in the text for clarity.

2.1. Synthesis of *N*,*N*-(3-(Butyloxy)propyl)amino Diacetic Acid (4a, BOPA Diacid). The synthetic route to BOPA is given in Figure 2. The starting materials are 3-(butyloxy)propylamine **1** and methyl iodoacetate, which are available commercially. Addition of the amine **1** to methyl iodoacetate in methanol at 22 °C for 15 h gave a mixture of monoester **2** and diester **3**, which were separated by flash column chromatography. The pure diester **3** was hydrolyzed to give the disodium salt **4**, which was neutralized to give the diacetic acid **4a**. The mixture of esters allowed us to separate easily the mono- and disubstituted products.

2.2. Theoretical Calculations of the Volumetric Titration Curves. We have developed for aqueous glycine[2] the equations for the titration based on the dissociation equilibrium equations, the equations of conservation of the species, and the equation of electroneutrality. For the titration of BOPA the principles of calculation are the same. Since BOPA is a substance containing three pK_a's and was titrated starting from an ionic form, the equations derived for the titration of glycine which contains two pK_a's cannot be used as reported. The problem must be reworked from the starting point. For clarity, we give all the equations pertinent to BOPA.

For BOPA the different ionic species in equilibrium are believed to be as given in eq 1. The abbreviations as defined in Figure 1 are given in the second line. The equilibrium constants relate two adjoining ionic species.

$$\begin{array}{c} RNH^+(CH_2COOH)_2 \overset{K_1}{\rightleftharpoons} \\ H^+AH_2 \end{array} \quad \begin{array}{c} RNH^+(CH_2COOH,CH_2COO^-) \overset{K_2}{\rightleftharpoons} \\ H^+AH^- \end{array} \quad \begin{array}{c} RNH^+(CH_2COO^-)_2 \overset{K_3}{\rightleftharpoons} RN(CH_2COO^-)_2 \\ H^+A^{2-} \qquad\qquad A^{2-} \end{array} \quad (1)$$

2.2.1. Definition of the Problem. We consider here the two cases of the titration: the acid and the base in the aqueous solutions. We write the problem as if both titrants were added simultaneously, which is not the case in the experiments, but the mathematics are easier to deal with when the two cases are taken together. The two cases are separated at the end of the resolution of the problem.

The following substances are present in the aqueous solutions: (1) the product to titrate (H^+AH^-); (2) the aqueous HCl which is the strong acid; (3) the aqueous NaOH which is the strong base; and (4) the counterions Na^+ and Cl^-. The base and acid are assumed to dissociate completely in water so that only the product to titrate is in equilibrium with its different ionic species.

The mean number of counterions, Na^+ and Cl^-, associated with dry BOPA is indicated by the parameter α. We write α_{-1} for the negative counterion associated with H^+AH_2 and α_1 for the positive counterion associated with H^+A^{2-} and A^{2-}. We have $0 \leqslant \alpha_1 \leqslant 2$ and $0 \leqslant \alpha_{-1} \leqslant 1$.

The equilibrium state reached by the mixture is given by eq 2 in which the left member represents the starting point and the right one, the equilibrium state of the mixture.

$$\begin{array}{c} AH_2 + Cl^- + Na^+ + H_2O + NaOH + HCl \rightarrow \\ A \quad \alpha_{-1}A \quad \alpha_1A \quad D \quad C_1 \quad C_{-1} \end{array}$$

$$\begin{array}{c} H^+AH_2 + H^+AH^- + H^+A^{2-} + A^{2-} + H^+ + OH' + \\ y_1 \quad y_2 \quad y_3 \quad y_4 \quad x \quad x' \end{array}$$

$$\begin{array}{c} Na^+ + Cl^- + H_2O \\ c \quad c' \quad d \end{array} \quad (2)$$

The second line in (2) represents the quantities involved during the titration measurements. The quantities A, C_1, C_{-1}, and D are known. The quantity of water does not vary much during the reactions and is considered constant

(4) Weers, J. G.; Rathman, J. F.; Axe, F. Y.; Crichlow, C. A.; Foland, L. D.; Scheuing, D. R.; Wiersema R. J.; Zielske, A. G. *Langmuir* **1991**, *7*, 854−867.
(5) Chapados, C.; Trudel, M. *Biophys. Chem.* **1993**, *47*, 267−276.
(6) Chapados, C.; Barwicz, J.; Gruda, I. *Biophys. Chem.* **1994**, *51*, 71−80.
(7) Chapados, C.; Girard, D.; Trudel, M.; Ringuet, M. *Biophys. Chem.* **1995**, *54*, 165−174.
(8) Chapados, C.; Girard, D.; Trudel, M.; Ringuet, M. *Biophys. Chem.* **1995**, *55*, 289−300.

as a first approximation. The problem contains 11 unknowns. The values of the parameters α_1 and α_{-1} are either known *a priori* or are obtained by fitting the experimental data with the theoretical titration curves. Therefore, nine unknowns remain to be determined.

2.2.2. The Equations. As mentioned above, the usual equations are used for the following calculation of the theoretical titration curves: the dissociation equilibrium equations, the equations of conservation of the species, and the equation of electroneutrality. References are given as examples.[9−11]

2.2.2.1. Equilibrium Relations. As a first approximation, we consider the activity coefficients constant throughout the series of samples. The total volume of the solution is V. Using the quantity of each species as defined in eq 2, the dissociation constants between the four ionic forms of AH_2 in equilibrium with water are written as follows:

$$\frac{[H^+AH^-][H^+]}{[H^+AH_2]} = \frac{(y_2/V)(x/V)}{(y_1/V)} = K_1 \quad (3)$$

$$\frac{[H^+A^{2-}][H^+]}{[H^+AH^-]} = \frac{(y_3/V)(x/V)}{(y_2/V)} = K_2 \quad (4)$$

$$\frac{[A^{2-}][H^+]}{[H^+A^{2-}]} = \frac{(y_4/V)(x/V)}{(y_3/V)} = K_3 \quad (5)$$

$$[H^+][OH^-] = \frac{x}{V}\frac{x'}{V} = K_0 \quad (6)$$

in the latter $[H^+]$ represents $[H_3O^+]$ for simplicity. The dissociation constants used above are the effective constants corresponding to the aqueous solution studied. These effective constants are not the thermodynamic equilibrium constants. Equations 3−6 are rewritten as follows:

$$y_2 = y_1 K_1 \frac{V}{x} \quad (7)$$

$$y_3 = y_1 K_1 K_2 \left(\frac{V}{x}\right)^2 \quad (8)$$

$$y_4 = y_1 K_1 K_2 K_3 \left(\frac{V}{x}\right)^3 \quad (9)$$

$$x' = K_0 \frac{V^2}{x} \quad (10)$$

2.2.2.2. Conservation of Species. The relations for the conservation of the species lead to

$$A = y_1 + y_2 + y_3 + y_4 \quad (11)$$

$$C_{-1} + \alpha_{-1} A = c' \quad (12)$$

$$C_1 + \alpha_1 A = c \quad (13)$$

$$x' + d = C_1 + D \quad (14)$$

The latter comes from the application of the relation for the conservation of species to OH^- in eq 2.

2.2.2.3. Electroneutrality Relation. Since the solutions are considered electrically neutral at all times, we apply this condition to the equilibrium state. This will give us the remaining necessary equation to solve the problem:

$$y_1 + x + c = y_3 + 2y_4 + x' + c' \quad (15)$$

2.2.3. Resolution of the Problem. 2.2.3.1. Abundance of the Species as a Function of pH. Using eqs 7 to 9 and 11, one obtains the following relations:

$$y_1 = A \Big/ \left[1 + K_1\left(\frac{V}{x}\right) + K_1K_2\left(\frac{V}{x}\right)^2 + K_1K_2K_3\left(\frac{V}{x}\right)^3\right] \quad (16)$$

$$y_2 = A \Big/ \left[\frac{1}{K_1}\left(\frac{x}{V}\right) + 1 + K_2\left(\frac{V}{x}\right) + K_2K_3\left(\frac{V}{x}\right)^2\right] \quad (17)$$

$$y_3 = A \Big/ \left[\frac{1}{K_1K_2}\left(\frac{x}{V}\right)^2 + \frac{1}{K_2}\left(\frac{x}{V}\right) + 1 + K_3\frac{V}{x}\right] \quad (18)$$

$$y_4 = A \Big/ \left[\frac{1}{K_1K_2K_3}\left(\frac{x}{V}\right)^3 + \frac{1}{K_2K_3}\left(\frac{x}{V}\right)^2 + \frac{1}{K_3}\left(\frac{x}{V}\right) + 1\right] \quad (19)$$

The relative abundance of the four ionic forms of the molecule AH_2 as a function of pH is obtained by dividing eqs 16−19 by A and replacing x/V by 10^{-pH}.

2.2.3.2. Evaluation of the Quantity of Titrant from the Concentration of H^+, (x/V). For a given amount of BOPA in solution, it is only the difference between the quantities of acid and base that will influence the value of the pH and the repartition of the species. We therefore calculate the value of $C_1 - C_{-1}$. Combining eqs 12 and 13 gives

$$C_1 - C_{-1} = c - c' - A(\alpha_1 - \alpha_{-1}) \quad (20)$$

Then using eqs 14, 7−10, and 15, we obtain

$$C_1 - C_{-1} = A\left[\left[\left(-1 + K_1K_2\left(\frac{V}{x}\right)^2 + 2K_1K_2K_3\left(\frac{V}{x}\right)^3\right) \Big/ \left(1 + K_1\left(\frac{V}{x}\right) + K_1K_2\left(\frac{V}{x}\right)^2 + K_1K_2K_3\left(\frac{V}{x}\right)^3\right)\right] - (\alpha_1 - \alpha_{-1})\right] - x + K_0\left(\frac{V^2}{x}\right) \quad (21)$$

2.2.3.3. Evaluation of the Quantity of Matter (mol) from the Mass of the Products. The molar quantities of the base NaOH (C_1) and the acid HCl (C_{-1}) are written as follows:

$$C_1 = m_1\frac{\epsilon_1}{M_1} \quad \text{and} \quad C_{-1} = m_{-1}\frac{\epsilon_{-1}}{M_{-1}} \quad (22)$$

where the m's are the masses of the titrants, the M's their molar masses, and the ϵ's the relative masses of their solutions. The subscripts 1 and −1 are for the base and acid, respectively. In the same way, the relative mass of the initial aqueous solution of BOPA is denoted ϵ_2. The molar mass of H^+AH^- is M_2. Depending on the initial ionic state of the BOPA used, we have to take into account the presence of the counterion associated with the dry BOPA in the balance of mass. To simplify the following text, we use the index Δ for the charge of the counterion: $\Delta = \pm 1$. Let α_Δ represent the relative initial amount of the counterion associated with the product to titrate and $M_{\Delta 3}$ the mass added to the molar mass of the neutral form H^+AH^- when associated with the counterion. For the positive counterion, Na^+, the mass added is 21.99 g/mol

(9) Barret, A. J.; Kennedy, J. C.; Jones, R. E.; Nadeau, P.; Pottier, R. H. *J. Photochem. Photobiol. B: Biol.* **1990**, *6*, 309−323.
(10) Levie, R. D. *J. Chem. Educ.* **1993**, *70*, 209−217.
(11) Glaister, P. *J. Chem. Educ.* **1997**, *74*, 760−761.

(Na replaces one H). For the negative counterion, Cl^-, the mass added is 36.45 g/mol (since H^+ is linked to the molecule and Cl^- is the counterion, see Figure 1) so that $M_3 = 21.99$ g/mol and $M_{-3} = 36.45$ g/mol. The mass of the solution of BOPA in the sample is denoted m_2. When dry, BOPA is associated to one type of counterion at a time. We therefore have the relation

$$A = m_2 \frac{\epsilon_2}{M_2 + \alpha_\Delta M_{\Delta 3}} \tag{23}$$

Combining eqs 21 to 23 we obtain

$$m_1\frac{\epsilon_1}{M_1} - m_{-1}\frac{\epsilon_{-1}}{M_{-1}} = m_2\frac{\epsilon_2}{M_2 + \alpha_\Delta M_{\Delta 3}}\left[\left[\left(-1 + K_1K_2\left(\frac{V}{x}\right)^2 + 2K_1K_2K_3\left(\frac{V}{x}\right)^3\right)\Big/\left(1 + K_1\left(\frac{V}{x}\right) + K_1K_2\left(\frac{V}{x}\right)^2 + K_1K_2K_3\left(\frac{V}{x}\right)^3\right)\right] - \Delta\alpha_\Delta\right] - x + K_0\left(\frac{V^2}{x}\right) \tag{24}$$

2.2.3.4. Experimental Parameters. To simplify the following text, we use the index for the titrant: $\delta = 1$ and -1 for the solutions of the base and acid, respectively.

The experiments are done by measuring the volumes and masses involved in the titration. Two methods can be used to make the measurements: (i) by taking the same mass of the solution to titrate in each sample so that in eq 24 the mass m_2 is constant, then the total volume will not be constant; or (ii) by taking the same total volume for each sample, then the mass of the solution to titrate, m_2, will not remain constant.

With the first method, m_2 is constant (A is also constant) and the total volume V is variable. Therefore, eq 24 gives the masses m_δ as a function of both the volume (V) and the quantity of H^+ (x), while the pH is a function of both V and x (pH is a function of x/V). Such a problem is not simple to resolve analytically since it is difficult to isolate the variable m_δ as a function of x or x/V only.

With the second method, the variable m_δ can be isolated from eq 24. The total volume V is kept constant, and the mass m_2 is a simple function of both the total mass m and the mass of titrant m_δ. The remaining two variables in the right side of eq 24 are m_2 and x. It is then necessary to express the variable m_2 as a function of x and m_δ. As it will be shown below, m_δ as a function of x is easy to isolate. This will in turn give the solution to the problem. The second method, being easier than the first, will be used to resolve the problem.

2.2.3.5. Mass of Titrant. Each sample was prepared to have the same volume, V. Performing the titration, we measured the mass m_δ of the titrant solution and the total mass m of the resulting mixture when the volume was completed to V. We therefore obtain $m_2 = m - m_\delta$. We noted that the total mass m was not the same for each sample. This indicates that the density of the mixture depends on the quantity of the titrant. The density of the samples can be approximated by a linear function of the mass of the titrant m_δ:

$$m = \rho V = \rho_0 V + \Delta\rho V = \rho_0 V + \rho_\delta m_\delta \tag{25}$$

where ρ_δ is introduced as a nondimensional parameter ($\rho_\delta = \Delta\rho V/m_\delta$) to further simplify the mathematical expressions. (The value of ρ_δ may not be the same for both titrants.) Using eqs 23 and 25 and the fact that $m_2 = m - m_\delta$ the value of A is written as follows:

$$A = \frac{\epsilon_2}{M_2 + \alpha_\Delta M_{\Delta 3}}[\rho_0 V - (1 - \rho_\delta)m_\delta] \tag{26}$$

It should be noted that the value of A depends only on the variable m_δ (ρ_δ is a constant parameter deduced from the measurements of m and m_δ taken together).

We now introduce the restriction that each titrant is used one at a time. Using eq 26 to derive the value of the mass m_2 and introducing it into eq 24, we obtain

$$\delta m_\delta\frac{\epsilon_\delta}{M_\delta} = [\rho_0 V - (1 - \rho_\delta)m_\delta]\frac{\epsilon_2}{M_2 + \alpha_\Delta M_{\Delta 3}}\left[\left[\left(-1 + K_1K_2\left(\frac{V}{x}\right)^2 + 2K_1K_2K_3\left(\frac{V}{x}\right)^3\right)\Big/\left(1 + K_1\left(\frac{V}{x}\right) + K_1K_2\left(\frac{V}{x}\right)^2 + K_1K_2K_3\left(\frac{V}{x}\right)^3\right)\right] - \Delta\alpha_\Delta\right] - x + K_0\left(\frac{V^2}{x}\right) \tag{27}$$

Isolating m_δ from eq 27 and replacing the quotient x/V by $[H^+]$ we obtain

$$m_\delta = V\left[\left(\rho_0\frac{\epsilon_2}{M_2 + \alpha_\Delta M_{\Delta 3}}\left[\left[\left(-1 + \frac{K_1K_2}{[H^+]^2} + 2\frac{K_1K_2K_3}{[H^+]^3}\right)\Big/\left(1 + \frac{K_1}{[H^+]} + \frac{K_1K_2}{[H^+]^2} + \frac{K_1K_2K_3}{[H^+]^3}\right)\right] - \Delta\alpha_\Delta\right] - [H^+] + \frac{K_0}{[H^+]}\right)\Big/\left(\delta\frac{\epsilon_\delta}{M_\delta} + (1 - \rho_\delta)\frac{\epsilon_2}{M_2 + \alpha_\Delta M_{\Delta 3}}\left[\left[\left(-1 + \frac{K_1K_2}{[H^+]^2} + 2\frac{K_1K_2K_3}{[H^+]^3}\right)\Big/\left(1 + \frac{K_1}{[H^+]} + \frac{K_1K_2}{[H^+]^2} + \frac{K_1K_2K_3}{[H^+]^3}\right)\right] - \Delta\alpha_\Delta\right]\right)\right] \tag{28}$$

Equation 28 gives the value for the mass of titrant m_δ as a function of $[H^+]$. The latter can be transformed directly into pH. Then the quantity of titrant, m_δ, can be expressed as a function of pH to obtain the plot in the familiar way. We have reported the titration of glycine which does not form any dianion, starting from the nonionic form.[2] Equation 28 can be reduced to the result obtained for glycine by using $\alpha_\Delta = 0$ and $K_3 = 0$ (so that $A^{2-} = 0$). This indicates that the above calculation is more general than the one reported for glycine.

Considering the high concentration of BOPA studied here (27.0% (w/v)), the concentration of water is below the standard value of 55.6 M at 25 °C, 1 atm. Then the exact value of the dissociation constant of water K_0 differs from the usual value of 10^{-14}.[12] This will affect the result for m_δ. However, eq 28 shows that this correction concerns only the very low concentration of H^+ (high pH). In this case, the slope of m_δ against $1/[H^+]$ depends directly on K_0. This correction is small for a 27% (w/v) aqueous solution of BOPA: the water concentration is about 43.5 M instead of 55.6 M for pure water. The value of K_0 to be used here is near 8.0×10^{-15}.

2.2.4. Concentration of Water. The quantity of water, d, at equilibrium is calculated using eq 14. The original quantity of water, D, is first evaluated from the content of water in the mixture:

$$D = m_2\frac{1 - \epsilon_2}{M_{H_2O}} + m_\delta\frac{1 - \epsilon_\delta}{M_{H_2O}} \tag{29}$$

(12) Vogel, A. I. *Quantitative Inorganic Analysis*; Longmans: London, 1951.

In the latter equation, m_2 is replaced by using eq 24. Then, C_1, x', and D are replaced in eq 14 using eqs 10, 22, and 29. After rearrangement, the concentration of water is obtained by

$$[H_2O] = \frac{d}{V} = \rho_0\left(\frac{1-\epsilon_2}{M_{H_2O}}\right) - \frac{K_0}{[H^+]} + \frac{m_\delta}{V}\left[\frac{\epsilon_\delta(1+\delta)}{2M_\delta} + \frac{\rho_\delta(1-\epsilon_2)+\epsilon_2-\epsilon_\delta}{M_{H_2O}}\right] \quad (30)$$

2.3. Factor Analysis. The mathematical formulation of factor analysis (FA) can be found in the book by Malinowski and Howery,[13] where the development leads to abstract factors. The method has been modified to give the *real* spectra and the *real* multiplication factors, MFs.[5] With the proper normalization procedure, the MFs in infrared spectroscopy give the concentration of the species. The FA procedure that we have used in this study is made up in two phases.

We give here a brief account of the first phase.[5] In the latter, the different steps in FA that we use to retrieve the real spectra and the real MF of the individual species from a series of experimental spectra are the following. The *n* IR spectra are first transformed into a data matrix [**D**], the starting point in the FA procedure. The data matrix is transformed into the covariance matrix, diagonalized, and decomposed into the abstract factors, which are used to obtain the eigenvalues.[14] The number of different species in a set of spectra are determined, and the eigenvectors are obtained for the MF. For the orthogonal spectra an optimization program[15,16] is used to (1) obtain the minimization of the negative values of the MF and of the spectra, (2) scan for the simplest spectra using the spectrum, its first and second derivatives;[17] and (3) constrain the MF to have a maximum and to have the extremities at zero. This last constraint is due to the fact that MF may not take a value above unity (the current value when the considered species is alone in the mixture). On the other hand, if the MF value is not restricted to be as high as possible (below unity), the related spectrum will present artificially high absorption bands. The observed spectrum is the product of the MF and the related spectrum. Physical evidence does not permit MF values to be negative, but experimental errors may conduct to some low negative MF values. The above optimizations, which are parametrizable, are made in succession. This method gave us the real IR spectrum of all species in a solution with their real MF at each pH.

Although the results at the first stage were good we, nevertheless, pursued the investigation into the next stage. Once we obtained the pure spectra of each ionic species of BOPA and their distribution as a function of pH, we knew at what pH the pure ionic forms are situated. Then, after transporting the spectroscopic data to an Excel spreadsheet program, we can select the pH regions where the pure spectra can be retrieved. At the pH where a species is the most abundant but not in pure form we subtract the spectrum of the species that we have identified previously as pure. By repetitive use of this procedure we obtain nonredundant spectra. With the pure spectra at hand one determines the amount of each species at all pH's.

3. Experimental Section

3.1. Synthesis of *N,N*-((Butyloxy)propyl)amino Diacetic Acid (4a). *3.1.1. Materials.* Anhydrous reactions were performed under an inert atmosphere; the setup was assembled and cooled under dry nitrogen. Unless otherwise noted the starting materials, reactants, and solvents were obtained commercially from Aldrich and used as such or purified and dried by standard procedures.[18] Organic solutions were dried over magnesium sulfate ($MgSO_4$) and evaporated under reduced pressure on a rotatory evaporator. All reactions were monitored by thin-layer chromatography (TLC). The plates were visualized by UV fluorescence or by staining with iodine. Commercial TLC plates were polyester silica gel 60 Å, 0.25 mm (Sigma T 6145). Flash chromatography was performed according to the method of Still *et al.* on Merck grade 60 silica gel, 230−400 mesh.[19] All solvents used in the chromatography had been distilled. The IR spectra were taken on a Nicolet model 205 FT-IR. Mass spectral assays (MS, *m/e*) were obtained using a VG Micromass 7070 HS instrument using an ionization energy of 70 eV. Nuclear magnetic resonance (NMR) spectra were obtained in $CDCl_3$ solutions, unless otherwise noted, on a Bruker AMX-2 (500 MHz) instrument. Chemical shifts were measured relative to internal standards: tetramethylsilane (TMS, δ 0.0 ppm) for ^{1}H and $CDCl_3$ (δ 77.0 ppm) for ^{13}C NMR. For spectra obtained in a mixture of D_2O/dioxane-d_8 the chemical shifts were measured relative to internal standards: dioxane (δ 3.77 ppm) for ^{1}H and (δ 66.65 ppm) for ^{13}C NMR. For ^{13}C NMR the number of carbons associated with a signal is presented in brakets if this number is greater than 1. Multiplicities are described by the following abbreviations: s (singlet), d (doublet), t (triplet), and so on.

*3.1.2. Preparation of N,N-(3-(Butyloxy)propyl)amino Dimethyl Acetate (**3**).* A solution of 3-(butyloxy)propylamine (**1**) (1.7 g, 2.0 mL, 13 mmol) and methyl iodoacetate (1.55 g, 7.78 mmol) in methanol was stirred at 22 °C for 15 h under a nitrogen atmosphere. The reaction mixture was cooled and the solvent evaporated. The residue, dissolved in diethyl ether, was washed with an aqueous solution of sodium bicarbonate (5%, 1 × 3 m) and with water (3 × 3 mL). The organic phase was evaporated and the crude material purified by flash chromatography using a mixture of hexane/acetone (9:1 and 7:3) to give the diacetate **3** (0.35 g, 33%) and monoacetate **2** (0.44 g, 28%), respectively. With a mixture of hexane/acetone (4:1) the R_f of compounds **2** and **3** are 0.3 and 0.5, respectively. Both compounds were fully characterized. However, only the spectral data of compound **3** are presented below as it will lead to the final product **4a**. IR, (ν_{max}, thin film on KBr): 1753 and 1741 (2 C=O), 1225−1100 (C−O) cm^{-1}. ^{1}H NMR (ppm): δ 3.70 (6H, s, 2 $-OCH_3$), 3.56 (4H, s, 2 $-CH_2CO_2-$), 3.45 (2H, t, J = 6.7 Hz, $-CH_2O-$), 3.39 (2H, t, J = 6.4 Hz, $-CH_2O-$), 2.80 (2H, t, J = 7.5 Hz, $-CH_2N-$), 1.73 (2H, p, J = 7.0 Hz, $-NHCH_2CH_2CH_2O-$), 1.53 (2H, p, J = 7.0 Hz, $CH_3CH_2CH_2-$), 1.35 (2H, h, J = 7.3 Hz, CH_3CH_2-), 0.91 (3H, t, J = 7.0 Hz, CH_3CH_2-). ^{13}C NMR (ppm): δ 171.5 (2), 70.4, 68.3, 54.7 (2), 51.3, 51.2 (2), 31.6, 28.0, 19.1, 13.7. MS (*m/e*): 275 (M^+), 216 ($M^+ - CO_2CH_3$). High-resolution MS. Calcd for $C_{13}H_{25}NO_5$: 275.1733. Found: 275.1728.

*3.1.3. Preparation of N,N-(3-(Butyloxy)propyl)amino Diacetic Acid, Sodium Salt (**4**).* A solution of diester **3** (1.18 g, 4.3 mmol) in methanol (30 mL) was treated with 2 equiv of an aqueous solution of sodium hydroxide (2.44 M, 3.54 mL, 8.66 mmol). The reaction mixture was stirred at 22 °C for 15 h. Then, the methanol and water were evaporated under reduced pressure to give the dicarboxylate **4** (1.24 g, 99.6%). IR (ν_{max}, KBr pellet): 1599 (C=O), 1115 (C−O) cm^{-1}. ^{1}H NMR (D_2O, dioxane-d_8, ppm): δ 3.42 (4H, q apparent (2 t), J = 7.9 Hz, $-CH_2OCH_2-$), 3.10 (4H, s, 2 $-CH_2CO_2Na$), 2.52 (2H, t, J = 7.4 Hz, $-CH_2N-$), 1.67 (2H, p, J = 7.5 Hz, $-NHCH_2CH_2CH_2O-$), 1.45 (2H, p, J = 7.4 Hz, $CH_3CH_2CH_2-$), 1.24 (2H, h, J = 7.4 Hz, CH_3CH_2-), 0.81 (3H, t,

(13) Malinowski, E. R.; Howery, D. G. *Factor analysis in chemistry*; Robert E. Krieger Publishing Co.: Malabar, Fl, 1989.
(14) Shurvell, H. F.; Bulmer, J. T. *Vibrational spectra and structure*; Elsevier Scientific Publishing Co.: Amsterdam, 1977.
(15) Powel, M. J. D. *Optimization techniques with FORTRAN*; McGraw-Hill: New York, 1973.
(16) Friedrich, J. B.; Yu, J. P. *Appl. Spectrosc.* **1987**, *41*, 227−234.
(17) Savitsky, A.; Golay, M. J. E. *Anal. Chem.* **1964**, *36*, 1627−1639.
(18) Perrin, D. D.; Armarego, W. L. F. *Purification of Laboratory Chemicals*, 3rd ed.; Pergamon Press: Oxford, U.K., 1988.
(19) Still, W. C.; Kahn, M.; Mitra, A. *J. Org. Chem.* **1978**, *43*, 2923−2925.

J = 7.4 Hz, CH_3CH$_2$−). ^{13}C NMR (D_2O, dioxane-d_8, ppm): δ 179.5 (2), 70.6, 69.0, 58.8 (2), 51.7, 30.9, 26.3, 18.8, 13.3.

*3.1.4. Preparation of N,N-(3-(Butyloxy)propyl)amino Diacetic Acid (**4a**).* The acidification of derivative **4** was done using a standardized solution of hydrochloric acid. Briefly, to a solution of carboxylate (5 mmol) dissolved in water (5 mL) was added the necessary amount of HCl(aq) (2 equiv). The solution was stirred for 1 h. Afterward, the solvent was evaporated, and the residue was dissolved in a minimum of ethanol. The mixture was left undisturbed for 24 h, during which time sodium chloride precipitated out of the solution. The salt was filtered, and the filtrate was evaporated to give the crude acid. Azeotropic evaporations (3−7 times) with dry acetone yielded the pure acid **4a** (99%). IR (ν_{max}, KBr pellet): 3200−2100 (OH, acid), 1730 (C=O), 1111 (C−O) cm^{-1}. ^{1}H NMR (D_2O, dioxane-d_8, ppm): δ 4.04 (4H, s, 2 −CH_2CO$_2$−), 3.70 (2H, t, J = 5.4 Hz, −CH_2O−), 3.55 (2H, t, J = 6.7 Hz, −CH_2O−), 3.46 (2H, t, J = 6.7 Hz, −CH_2N−), 2.07 (2H, p, J = 5.8 Hz, −NHCH$_2$CH_2CH$_2$O−), 1.57 (2H, p, J = 7.2 Hz, CH$_3$CH$_2$CH_2−), 1.34 (2H, h, J = 7.5 Hz, CH$_3$CH_2−), 0.91 (3H, t, J = 7.4 Hz, CH_3CH$_2$−). ^{13}C NMR (D_2O, dioxane-d_8, ppm): δ 169.2 (2), 71.3, 68.6, 55.7, 55.6 (2), 30.8, 23.7, 18.8, 13.3. MS (*m/e*): 247 (M^+), 229 ($M^+ - H_2O$). High-resolution MS. Calcd for $C_{11}H_{21}NO_5$: 248.1420. Found: 247.1424.

3.2. Aqueous Measurements. *3.2.1. Chemical and Solutions.* BOPA diacid (**4a**) was synthesized as described above. The product was partially neutralized with 1 equiv NaOH. Deionized water was used to prepare the aqueous solutions. Aqueous NaOH, 50.8% (w/w), and concentrated HCl, 37.0% (w/w, density 1,19 g/mL, Fisher Scientific), were used for the titration. The titrants were used at these concentrations to maintain a high concentration of BOPA (27.0% (w/v)). Both acid and base were calibrated by standard methods. A stock solution of BOPA was prepared by dissolving 67.5 g of solid in 250 mL of water. The pH of this solution was 7.99.

Each sample was prepared first by weighing the titrant in an empty 10 mL volumetric flask. The volume was then completed to 10 mL with a part of the stock solution of BOPA. The resulting total mass was measured. A series of 27 samples in the pH range 0−14 was prepared. The samples gave homogeneous solutions. Each sample was divided into two parts: one part for the IR measurements and the other part for the pH measurements.

3.2.2. pH Measurements. The pH was measured at ambient temperature (24 ± 1 °C) with a pH meter (Omega model PHH-253) equipped with a combination electrode (Analytical Sensors, Inc., model PH10107B-03-B). Prior to any series of measurement a calibration was carried out at pH = 4.00, 7.00, and 10.00.

3.2.3. IR Measurements. The IR measurements were obtained on a Nicolet FT-IR spectrometer model 510P. Two KBr windows isolate the measurement chamber from the external atmosphere. The samples were contained in a "Circle cell" (SpectraTech, Inc.) equipped with a ZnSe crystal rod (making 11 internal reflections). The spectra were taken under a flow of nitrogen to ensure low residue of CO_2 and water vapors in the spectrometer. For each spectrum, an accumulation of 20 scans with a resolution of 4 cm^{-1} was performed. All measurements were taken at 26.5 ± 0.3 °C. The cell was carefully washed and dried before each measurement.

3.2.4. Treatment of the Data. The IR data points [$I(\tilde{\nu})$ vs $\tilde{\nu}$ (cm^{-1})] stored on diskettes were transferred to a central computer (IBM RS6000) for the numerical treatment. Parallel calculations were made using a Microsoft Excel spreadsheet program on a personal computer. The subtraction of the water bands was done using four eigenspectra of water; one of pure water, one of NaCl solvated water (2.60 M); one of NaOH solvated water (2.23 M); and one of HCl solvated water (1.54 M). The rationale for this subtraction is described in ref 20. Briefly, each solution has its own signature that does not vary when the solutions are mixed together except for a dilution effect. This permits careful subtraction of the water bands from every spectrum. The resulting spectra needed no further baseline adjustment. Usually, with cylindrical ATR cells[21] and/or with samples in aqueous media[22] a baseline adjustment is made. The FA procedure is applied to these resulting spectra.

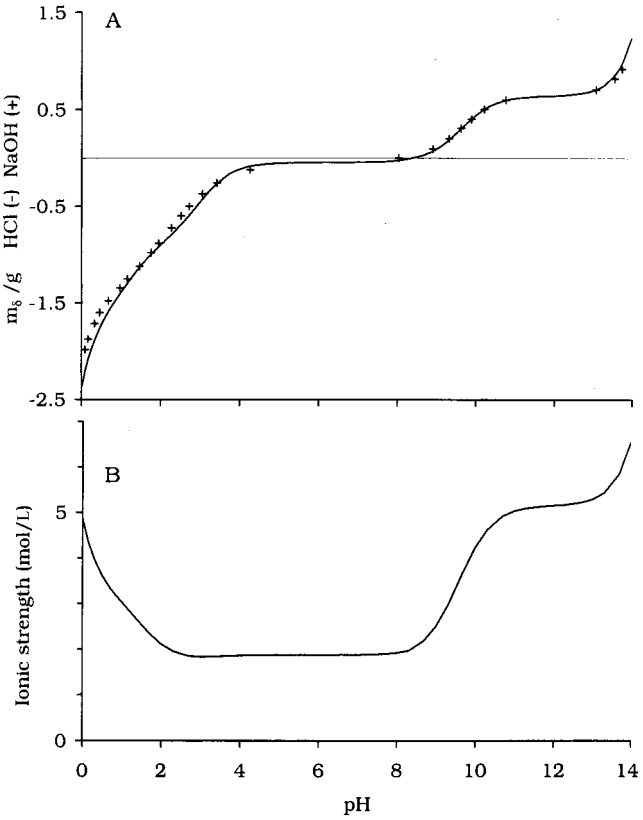

Figure 3. (A) Volumetric titration curve of 10 mL of 1 M aqueous BOPA. (—) calculated and (+) observed. The quantity of HCl (37.0% (w/w)) or NaOH (50.8% (w/w)) added to the solution of BOPA is in grams (see text for details). (B) Ionic strength as a function of pH.

Table 1. Parameter Values for Theoretical Volumetric Titration Curve Calculation of BOPA[a]

M_1 (g)	40.00	ρ_0 (g/L)	1065.9
M_{-1} (g)	36.46	ρ_1	0.30
M_2 (g)	247.30	ρ_{-1}	0.07
ϵ_1 (g/g)	0.508	V (L)	0.01
ϵ_{-1} (g/g)	0.37	K_0	8.0×10^{-15}
ϵ_2 (g/g)	0.248		

[a] M_1, M_{-1}, and M_2 are the molar masses of the base, acid, and neutral form of BOPA, respectively. ϵ_1, ϵ_{-1}, and ϵ_2 are the concentrations (w/w) of the base, acid, and the stock solution of BOPA, respectively. ρ_0 is the density of the stock solution of BOPA, ρ_1 and ρ_{-1} are nondimensional parameters related to the variation of the density (see text), V is the volume of each sample, and K_0 is the dissociation constant of water.

4. Results and Discussion

4.1. Volumetric Titration. Figure 3A shows the plot of the experimental volumetric titration of BOPA: m_δ as a function of pH. The experimental values (+ + +) are presented without further treatment. The theoretical values for the titration curve are calculated using eq 28, and the results are presented in Figure 3A by the full line (—). The values of the parameters used for this calculation are given in Table 1. Using eq 30, we evaluated the variation of the concentration of water in the different samples reported here. The results showed that the water concentration varied from 44.5 to 42.9 M. This small

(20) Max, J.-J.; Trudel, M.; Chapados, C. *Appl. Spectrosc.* **1998**, *52*, 234−239.
(21) Doyle, W. M. *Appl. Spectrosc.* **1990**, *44*, 50−59.
(22) Powell, J. R.; Wasacz, F. M.; Jakobsen, R. J. *Appl. Spectrosc.* **1986**, *40*, 339−344. Dousseau, F.; Therrien, M.; Pézolet, M. *Appl. Spectrosc.* **1989**, *43*, 538−542.

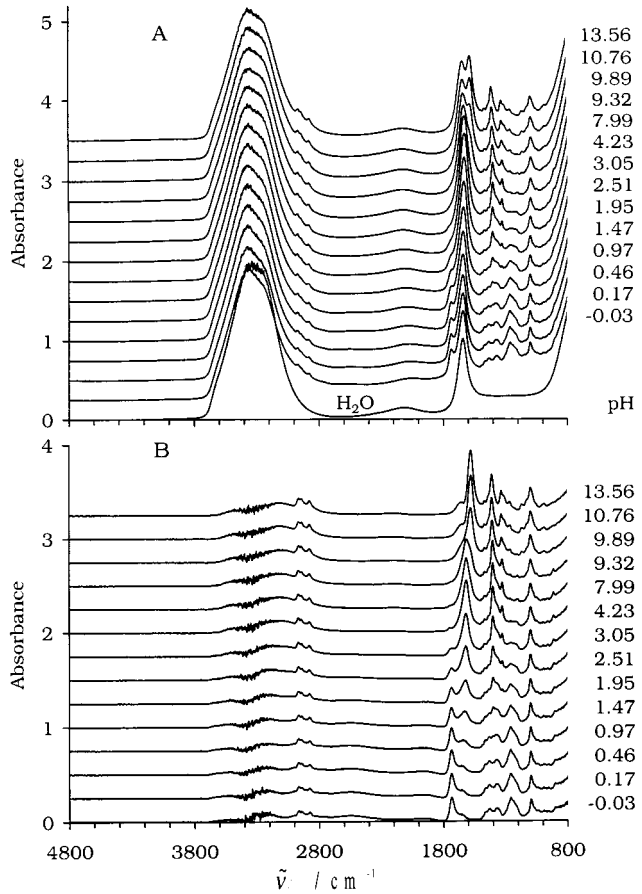

Figure 4. (A) IR spectra of aqueous BOPA at different pH. The bottom spectrum is that of pure water. (B) IR spectra after subtraction of the water spectra. In A each spectrum except that of pure water is offset by 0.25 Abs; in B each spectrum is offset by 0.25 Abs except that at pH − 0.03.

variation of less than 4% justifies not to consider it in the theoretical calculations. For the mean concentration of water (43.5 M) the dissociation constant $K_0 = 8.0 \times 10^{-15}$.

Good agreement between experimental measurements and theoretical values derived from eq 28 (Figure 3A) is obtained by taking $pK_1 = 1.30$, $pK_2 = 2.90$, and $pK_3 = 9.65$. These pK_a values represent the observed pK_a's of BOPA in the present experimental conditions. The mean number of counterion (α_1) per BOPA molecule obtained by the best fit between experimental data and theoretical values is $\alpha_1 = 1.05 \pm 0.05$. The dry BOPA used here contains 1.05 Na^+ per molecule as a mean value. Therefore, the neutralization of BOPA diacid (**4a**) had been made with 1.05 equiv NaOH to give 95% molarity of monoanion H^+A^{2-} and 5% of dianion A^{2-}. Furthermore, the molar weight of the dry BOPA equals 270.39 g/mol, and the stock solution is 1.00 ± 0.01 M. The small deviation between experimental and theoretical values at low pH (Figure 3A) is attributed to the activity coefficient of the protons (see section 4.5).

4.2. IR Spectra. From the 27 spectra of BOPA in water obtained in the pH range 0−14, Figure 4A shows 14 typical spectra from 4800 to 800 cm^{-1}. Great variations are observed in the carbonyl region (1750 to 1550 cm^{-1}) which reflect the perturbation induced to the carbonyl groups during titration. From these spectra, the four eigenspectra of water were carefully subtracted (section 3.2.4). The resulting spectra are given in Figure 4B. The position and relative intensity of many bands vary with pH due to the change in the ionic composition of BOPA.

4.3. Results from Factor Analysis. *4.3.1. IR Spectra of Separated Species.* Because of our experimental method of titration, the amount of BOPA is not constant throughout the series of samples. This situation requires that we use a two-step FA procedure. In the first step, the variation of the amount of BOPA is not taken into consideration. This gives us the pK_a values and the minimal number of species involved. As indicated by eqs 3−5, the value of pH gives the value of the pK_a when the two species in equilibrium are of equal amounts (1:1). The IR results are in agreement with eq 1 (vide supra). The pK_a values obtained from the IR spectra are in reasonable agreement with the values obtained from the volumetric titration (Table 2).

The real spectra of the different species are obtained with the FA procedure. We evaluated the real amount of BOPA in each sample and used the relative values as normalization factors in the FA process. The parameters needed are the following (see Table 1): $M_2 + \alpha_\Delta M_3 = 270.39$ g mol^{-1} (the molar mass of BOPA with its counterion); $\epsilon_2 = 0.248$; $\rho_0 = 1065.9$ g/; $\rho_1 = 0.30$; $\rho_{-1} = 0.07$; and $V = 10.0$ mL. Figure 5 shows the IR spectra of the four separated species of aqueous BOPA obtained from FA: the cation (H^+AH_2), the zwitterion (H^+AH^-), the monoanion (H^+A^{2-}), and the dianion (A^{2-}). The results shown in Figure 5 are normalized to the molar concentration of each species.

4.3.2. Relative Abundance of the Isolated Species: IR Titration Curves. Using the spectra of the isolated species normalized to the same amount of BOPA, the MFs were derived relative to each species, and therefore the variation of the MFs as a function of the quantity of titrant was obtained. The MF values were determined using a least-squares fit between the calculated and experimental spectra. Because of the high noise level in the 3200−3500 cm^{-1} region due to the strong absorption of water (Figure 4A), this region was omitted in the least-squares fit. The MF values obtained have a precision of ±0.015 with a mean error below 0.0038 Abs.

Figure 6 shows the normalized MFs as a function of the equivalent titrant. In the plot, the latter are the abscissas given by

$$X = \delta \frac{(m_\delta \epsilon_\delta / M_\delta)}{[(m - m_\delta)\epsilon_2 / (M_2 + \alpha_\Delta M_{\Delta 3})]} \qquad (31)$$

where m_δ is the mass of the titrant and m the mass of the 10 mL solution. The difference between these two quantities is the mass m_2, the amount of the BOPA solution in the mixture. The other terms in eq 31 are defined in Table 1. The ordinate in the plot is the normalized MFs since the amount of BOPA is not constant in all the samples (section 2.2.3.4). The normalization is made to the maximum amount of BOPA in the series: i.e., the sample without the titrant ($m_\delta = 0$). The normalized MFs of the four ionic species of BOPA which are given by Y_i (where $i = 1-4$ for the four species) are obtained with

$$Y_i = \frac{\mathrm{MF}_i}{[(m - m_\delta)/m_0]} \qquad (32)$$

where m_0 represents the mass of BOPA corresponding to the sample without any titrant. The result given in Figure 6 shows that the Y_i are linear relations with X. This result is in agreement with the classical neutralization process of a weak acid and weak base by a strong acid and strong base, respectively. When the amount of the base titrant is added to aqueous acidic BOPA, the cation (H^+AH_2: ▲) is transformed to the zwitterion (H^+AH^-: ×), then to the

Table 2. Comparison of the pK_a Values of BOPA and Glycine

glycine			BOPA (this work)		
volumetric[12]	volumetric at zero ionic strength[24,25]	volumetric and IR[2]	volumetric	IR	mean
			1.30 ± 0.15	1.50 ± 0.05	
2.34	2.350	2.35 ± 0.15			2.19 ± 0.15
			2.90 ± 0.15	3.05 ± 0.05	
9.60	9.778	10.00 ± 0.15	9.65 ± 0.10	9.60 ± 0.05	9.63 ± 0.10

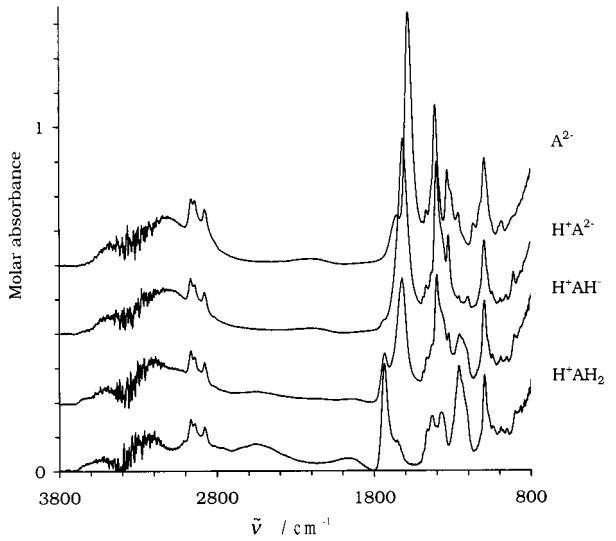

Figure 5. Molar spectrum of the four ionic species of BOPA retrieved by FA. The spectra of H^+AH^-, H^+A^{2-}, and A^{2-} are offset by 0.2, 0.4, and 0.6 absorbance, respectively.

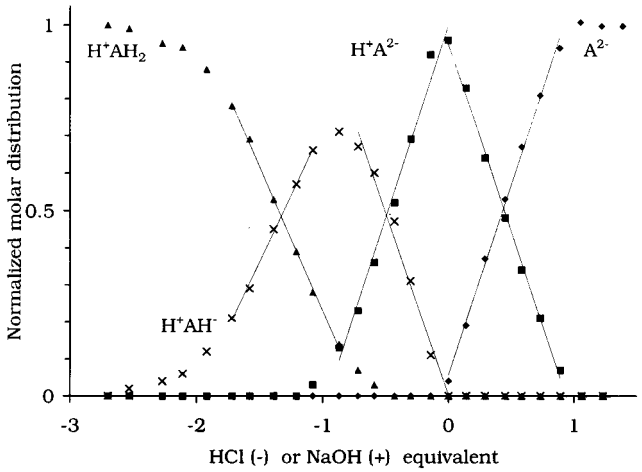

Figure 6. Normalized distribution of the four ionic species of aqueous BOPA as a function of the quantity of titrant (HCl (—) or NaOH (+) equivalent). H^+AH_2 is the cation (▲); H^+AH^- is the zwitterion (×); H^+A^{2-} is the monoanion (■); and A^{2-} is the dianion (♦). The points (▲, ×, ■, ♦) are the experimental values, and the lines are the best linear fits.

monoanion (H^+A^{2-}: ■), and finally to the dianion (A^{2-}: ♦). These transformations are directly proportional to the amount of titrant.

In Figure 6 except for the first two slopes (starting from the left) of species H^+AH_2 and H^+AH^- the observed slopes for the other species are close to unity (Figure 6 and Table 3, in absolute terms), which are the expected values for direct transformation of one species to another. For the first two slopes (H^+AH_2 and H^+AH^-) obtained at pH less than 2, the small difference between observed and theoretical slopes (Table 3) is attributed to the activity coefficient of the proton which, at the lowest pH, could differ substantially from unity. At pH < 1.0 the deviation

Table 3. Results of the Linear Fit of the Normalized Distribution of BOPA as a Function of Equivalent Titrant (See Figure 6)

titrant	ionic species	symbols	slope	ordinate at origin	correlation coeff
HCl	H^+AH_2	▲	−0.807	−0.585	0.999
HCl	H^+AH^-	×	0.734	1.453	0.998
HCl	H^+AH^-	×	−0.998	0.001	0.987
HCl	H^+A^{2-}	■	1.050	0.997	0.985
NaOH	H^+A^{2-}	■	−1.024	0.952	0.997
NaOH	A^{2-}	♦	1.032	0.056	0.997

of linearity of these two species at HCl equivalent to greater than 1.7 (<−1.7 on Figure 6) is also attributed to the activity coefficient of the proton. The activity coefficients do not modify the spectra of the ionic species,[23] and its influence on pH has not been taken into consideration in the theoretical calculations. Notwithstanding the latter effect, the slopes of unity indicate that the neutralization of aqueous 1 M BOPA follows strictly the dissociation equilibrium relation given in eq 1.

4.3.3. Efficacy of the FA Procedure. Figure 7 gives two examples that demonstrate the effectiveness of the FA procedure by calculating the difference between calculated and observed spectra. The former is obtained when the spectra of the species retrieved by FA (Figure 5) are multiplied by the MFs at the pH considered (Figure 6) and adding the results together. At pH = 9.32, two species are present: H^+A^{2-} and A^{2-}. The top part of Figure 7 shows the situation at this pH where the experimental spectrum (top) and the calculated spectra (middle) are illustrated. The bottom traces show the difference spectrum. The bottom part of Figure 7 shows the situation at pH = 1.95, where three species are present: H^+AH_2, H^+AH^-, and H^+A^{2-}. For these two examples which are typical of all 27 solutions studied, the low residues observed in the difference spectra indicate that the FA procedure is fully functional for the IR titration of an aqueous 1 M BOPA solution. Thus the spectra and the MFs retrieved by FA for the ionic species of BOPA are reliable.

4.3.4. Relative Amount of Each Species as a Function of pH. The normalized MFs (Y_i) of each sample can be expressed as a function of pH using eq 28. The result for BOPA is given in Figure 8 as four series of points: the cation (H^+AH_2: ▲); the zwitterion (H^+AH^-: ×); the monoanion (H^+A^{2-}: ■); and the dianion (A^{2-}: ♦). The theoretical MF values are calculated with eqs 16−19 and 26. The values of the parameters used are given in Table 1. The three dissociation constants obtained from the pK_a values deduced above (section 4.3.1) are also used for these calculations. The results are given in Figure 8 as the four full lines. These calculations are normalized relative to the amount of BOPA in each sample. From

(23) Menichelli, C.; Max, J.-J.; Trudel, M.; Chapados, C. Titrage de H_2SO_4 par infrarouge. 43rd International Conference on Analytical Sciences and Spectroscopy, Montréal, 1997. Scientific Program, p 35.

(24) Neckers, D. C.; Doyle, M. P. *Organic chemistry*; John Wiley and Sons: New York, 1977.

(25) Martell, A. E.; Smith, R. M. *Critical Stability Constants*, Vol. 1 Plenum Press: 1974. Cited in: Harris, D. C. *Quantitative Chemical Analysis*, 4th ed.; Freeman: New York, 1996.

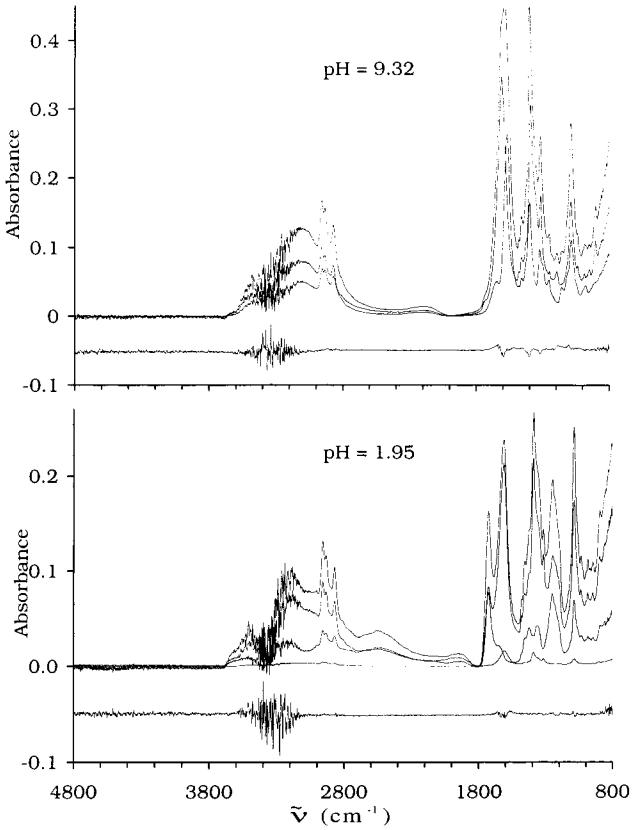

Figure 7. Comparison between experimental spectra and those retrieved by FA. At pH = 9.32, the top spectrum is the experimental spectrum; the middle spectra are the two retrieved spectra (from Figure 5) multiplied by the MFs (from Figure 6). After adding the middle spectra, they are subtracted from the top one to give the bottom spectrum (shifted by −0.0.05 Abs). At pH 1.95, the same pattern is used except that three retrieved spectra are used.

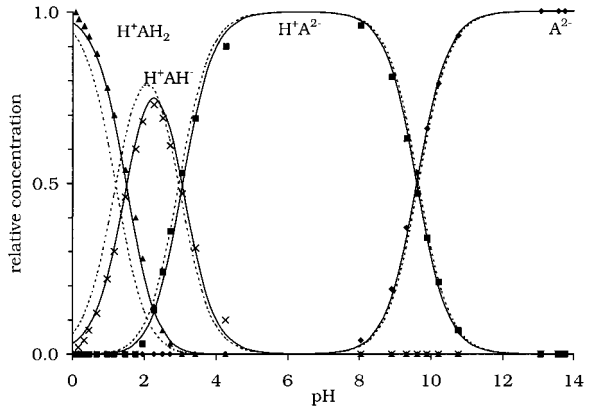

Figure 8. Relative distribution of the ionic species of BOPA as a function of pH. The symbols (▲, ×, ■, ♦) indicate the values obtained by the IR measurements. The solid lines (—) are obtained by the theoretical equilibrium equations. The dashed lines (- -)are obtained with the pK_a's given by the volumetric measurements.

Figure 8, it can be seen that there is a good agreement between the experimental values deduced from the spectra by the FA procedure and theoretical values calculated with the dissociation constants that were obtained from the experimental spectra. This indicates that FA is an effective method for obtaining the dissociation constants of a compound that possesses a characteristic IR spectrum for each ionic state of the compound.

On Figure 8 we observe that the cation H^+AH_2 and the zwitterion H^+AH^- are never alone in the pH range where they are observed. The quantity of the zwitterionic species in the solutions reaches its maximum value of 0.750 ± 0.015 at pH 2.2. During the titration near this pH, the cation and the monoanion are present along with the zwitterion. This complication indicates why the spectrum of the pure zwitterion is not observed at any pH in the original spectra, even after subtracting the water spectra. To separate the spectra of the different ions, it is necessary to have recourse to the FA procedure.

After obtaining Figure 8 in the first phase of the FA procedure, we verify the results in the second phase. Between pH 12 and 14 only the species A^{2-} exists; at pH 6.2 ± 0.5 only the species H^+A^{2-} exists; at pH 0 only the species H^+AH_2 exists. Of the three species that exist at pH 2.2, two are known. By subtracting the latter from the experimental spectrum, we obtain the pure zwitterion, H^+AH^-. With all the pure spectra of the ionic species at hand and since each spectrum has an original signature, the amounts of the pure species at the other pH's are determined. This is done by repetitive subtractions until the residues obtained by the difference between the calculated and experimental spectra are minimized (Figure 7).

4.4. IR Spectra of the Ionic Species of BOPA. The spectra of the four ionic species of BOPA are given in Figure 5 and the assignment of the bands are given in Table 4. Although there are some similarities from one spectrum to the other, a careful look shows that the spectra are genuine and nonredundant. The CH stretch bands are situated between 2980 and 2860 cm^{-1}. The bands that are typical of aliphatic CH groups do not change from one ionic species to the other. This indicates that the aliphatic chain is not affected by the ionic situation of the molecule. Therefore all the differences observed on the four spectra are due to the modifications that the head (or hydrophilic) groups undergo when passing from one ionic state to the other.

Because of the broad absorption and the high noise level, the OH stretch region is the most difficult region to assign although some qualitative interpretation can be done. For the dianion, A^{2-}, there is no NH stretch, and no acidic OH stretch bands so the broad absorption from 3600 to 2600 cm^{-1}, notwithstanding the CH stretch bands, can only be assigned to solvated water. This assignment is corroborated by the presence of the small typical water band situated near 2190 cm^{-1} which cannot be assigned to ordinary water nor to basic water that absorbs near 2120 cm^{-1}. The solvated water should be situated in the hydrophilic portion of the surfactant. For the cation, H^+AH_2, the OH stretch spans from 3680 to 2690 cm^{-1}. Some NH stretch is also situated in this region. There is no clear indication of the presence of solvated water although such a presence cannot be dispelled because there is some low absorption near 2120 cm^{-1} where a combination band of water is usually observed. The bands situated near 2540 and near 1960 cm^{-1} are typical bands of organic acids in water. For the zwitterion and the monoanion the situation in the OH stretch region is intermediate between the dianion and the cation.

The carbonyl stretch bands are also of interest. These absorb in the 1700 cm^{-1} region. For the cation, H^+AH_2, the carbonyl stretch band is situated at 1734 cm^{-1}, and the C−O stretch of the COH group is situated at 1256 cm^{-1}. These two bands being so far apart indicates that there is little resonance between the two groups. For the dianion, A^{2-}, the asymmetric C═O stretch is situated at 1574 cm^{-1}, while the symmetric C═O stretch is situated at 1406 cm^{-1}. For the monoanion and the zwitterion, these two bands are situated near 1617 cm^{-1} and at near 1398

Table 4. Positions (cm^{-1}) of the Ionic Species of BOPA (See Figure 5)[a]

type of ion	H^+AH_2	H^+AH^-	H^+A^{2-}	A^{2-}	assignment[b]
	3680−2689	3680−2689	3680−2689	3680−2689	OH s, NH s, H_2O[c]
	2965	2965	2965	2961	CH_3 as st
	2942	2944	2942	2939	CH_2 as st
	2878	2878	2878	2876	CH_3, CH_2 s st
	∼2540 b	∼2540 w,b			Acidic OH s
			∼2193 w,b	∼2193 w,b	H_2O combination
	∼1956 w,b	∼1956 w,b			H_2O combination
			1738 w		combination ?
	1734	1732			C═O st
	1650 b			1650	combination ?
		1619	1615	1574	CO_2^- as st
	1485 w, sh	1485	1485	1487	CH_3, CH_2 def
	1460 sh	1464	1466	1465	CH_3, CH_2 def
	1429	1428	1430 sh	1430 sh	CH_3, CH_2 def
	1406 vw				CH_2N^+ def
		1397	1399	1406	CO_2^- s st
	1370	1364 bsh	1364 sh	1370 w sh	CH_3 s def
		1324	1323	1329	CH_2 wag
				1308 sh	CH_2 wag ?
	1296 w sh				CH_2 wag ?
	1276 sh	1270 sh			CH_2 rock
	1260	1256			carboxylic C−O st
			1257 w	1262 w	CH_2 rock
	1239 sh	1237 sh	1235 vvw	1230 sh	CH_2 rock
	1210 sh	1210 sh	1201		NH^+ def
				1164	CCO^- wag
	1140 vw, sh	1140 vw sh	1140 w sh	1140 sh	C−N st
	1096	1096	1095	1094	C−O st of C−O−C

[a] Only the intensity of the weak and shoulder bands are indicated. For the other intensities, one should refer to the related spectra: w, weak; b, broad; sh, shoulder; v, very. [b] st, stretch; s st, symmetric stretch; as st, asymmetric stretch; def, deformation; s def, symmetric deformation. [c] Solvated water.

cm^{-1}, respectively. The displacements of the bands indicate a small loss of resonance of the COO^- group due in part to the influence of the N^+H group nearby when the dianion is transformed to the monoanion and to the zwitterion. For the latter, an acidic C═O band is also observed at 1734 cm^{-1} which is at the same position as the carbonyl groups of the acid H^+AH_2. This indicates that there is no interaction between the carboxylate and carboxylic groups of the zwitterion. Overall, the carbonyl bands truly reflect the ionization situation of each ionic species.

4.5. pK_a Values. The calculated ionic strength in the present measurements is given in Figure 3B. The values vary from 1.8 to about 7 mol/L. It is not possible to determine exactly how the observed pK_a's of BOPA would be affected at high concentration in aqueous solution by the ionic strength although those of aqueous glycine are not affected by a high concentration of 20% (w/v).[2] Therefore due to the similarity of the active structures between glycine and BOPA, the measured pK_a's of BOPA should not be affected by the high concentration used in the present work. However, such an ionic strength could affect the activity of proton in the low-pH region. This would explain the deviation observed at low pH between the theoretical and experimental titration curves shown in Figure 3A. Such a deviation has been observed on the IR titration of H_2SO_4.[23]

The values of pK_a's of BOPA obtained by IR and volumetric titrations are given in Table 2. For comparison we give in the same table the values of pK_a of glycine. The pK_a value of BOPA related to the amino group is 9.63. This value, which is almost the same as that of glycine, indicates that the presence of a second acetic group in BOPA does not modify the pK_a value of the amino group. The two acetic groups in BOPA give two pK_a at low pH, the mean value of which is 2.19. This value is the same (within experimental errors) as the pK_a value of glycine. The similarities of the pK_a's of glycine and BOPA indicate that the ionic behavior of the latter is similar to that of glycine, which makes it a true member of the glycinate family.

When we compare the pK_a values of BOPA obtained by IR and volumetric measurements, we notice that the pK_a's related to the amino group are the same within experimental errors: 9.60 and 9.65, respectively. The pK_a's related to the first acidic group although slightly different (3.05 and 2.90) are still the same within experimental errors. The pK_a's related to the second acidic group are at the limit of experimental errors. When these pK_a's are used to calculate the relative concentrations of the different ionic species (eqs 16−19), the deviations between IR and volumetric measurements of pK_a's are more evident. In Figure 8 we observe that from pH 14 to 4 good agreement is obtained between IR and volumetric values. From pH 4 to 0, the deviation grows as the pH decreases. The IR measurements of pK_a's give the true values of the abundance of the ionic species because this method observes them directly. The volumetric values are related to an electrical potential measurement influenced by the amount of protons in solution. The activity coefficient of the protons has not been taken into account in the equations developed here. This activity coefficient can explain the growing deviation of the two sets of curves as the pH decreases because the amount of protons increases as the pH decreases.

5. Conclusion

We have synthesized a pure surfactant of the glycinate family, BOPA. The synthesis was done in an organic solvent to ensure that the product would be pure and free of side products that may be difficult to get rid of when the synthesis is done in water. In an aqueous solution of BOPA and with no *a priori* information, we have identified, characterized, and determined the abundance of the cation, zwitterion, monoanion, and dianion in the IR

spectra. The ionization scheme proposed in Figure 1 has been verified.

Water being a strong absorber in IR could be considered an obstacle difficult to surmount in IR titration studies. However, water is strongly influenced by the presence of NaCl, HCl, and NaOH, which are the other molecules present in the solution. Each of these species interacts with water in a specific manner giving *per se* four eigenspectra when pure water is included. The four eigenspectra of water were subtracted from every aqueous solution, giving a series of spectra free of identifiable water with a flat baseline that necessitate no adjustment.

The FA procedure is then applied on the identifiable water free spectra to separate the spectra of the four ionic species and obtain their concentration in each solutions. The procedure worked perfectly well with BOPA giving the *real* IR spectra and *real* concentration of the four ionic species. The normalized distribution curves obtained by the FA procedure is the same as that obtained with volumetric titration notwithstanding the small deviation at low pH which we attribute to the activity coefficient of the proton that influences the volumetric titration but not the IR titration. The method that we have used for IR titration should work with more complicated systems as well as with other spectroscopic techniques such as Raman.

The results that we have obtained for BOPA indicates that IR titration of aqueous solutions is possible. This titration necessitates no color indicator and is not influenced by the activity coefficient of the proton nor by the ionic strength of the solution. During titration, the zwitterionic species is never alone and, at its maximum relative quantity of 0.75 reached at pH of 2.2, two other species are present: the cation H^+AH_2 and the monoanion H^+A^{2-}. This method gives the true IR spectra of the ionic species in their natural environment, water. The position of the carbonyl bands are good internal indicators of the ionic composition of the solutions.

Good agreement have been obtained between IR and volumetric titrations of aqueous BOPA. The pK_a's obtained by the two methods are the same when the activity coefficient of the protons are taken into account and comparable to the parent compound: glycine. The added advantage of the IR titration over other titration methods is that individual functional groups can be followed as the pH of the solution is modified. This advantage could be used to determine the aggregated properties of amphoteric surfactants.

Acknowledgment. This work was supported by Laboratoires Choisy Ltée and by the National Research Council Canada through its IRAP program.

List of Symbols

BOPA ≡ AH_2	*N*,*N*-((butyloxy)propyl)amino diacetic acid: $CH_3(CH_2)_3O(CH_2)_3N(CH_2COOH)_2$
H^+AH_2	*N*,*N*-((butyloxy)propyl)amino diacetic acid cation: $RNH^+(CH_2COOH)_2$
$AH_2 \equiv H^+AH^-$	*N*,*N*-((butyloxy)propyl)amino diacetic acid zwitterion: $RNH^+(CH_2COOH,CH_2COO^-)$
H^+A^{2-}	*N*,*N*-((butyloxy)propyl)amino diacetic acid monoanion: $RNH^+(CH_2COO^-)_2$
A^{2-}	*N*,*N*-((butyloxy)propyl)amino diacetic acid dianion: $RN(CH_2COO^-)_2$
α_1	mean number of counterions Na^+ associated with dry BOPA: $0 \leqslant \alpha_1 \leqslant 2$
α_{-1}	mean number of counterions Cl^- associated with dry BOPA: $0 \leqslant \alpha_{-1} \leqslant 1$
Δ	Δ selects the counterion associated with dry BOPA: $\Delta = \pm 1$
V	total volume of the sample (L)
A	total amount (mol) of BOPA in the sample
D	total amount (mol) of water initially in the sample
C_δ	amount (mol) of titrant added in the sample
y_1	amount (mol) of H^+AH_2 present in the sample at equilibrium
y_2	amount (mol) of H^+AH^- present in the sample at equilibrium
y_3	amount (mol) of H^+A^{2-} present in the sample at equilibrium
y_4	amount (mol) of A^{2-} present in the sample at equilibrium
x	amount (mol) of H^+ present in the sample at equilibrium
x'	amount (mol) of OH^- present in the sample at equilibrium
c	amount (mol) of Na^+ ions present in the sample at equilibrium
c'	amount (mol) of Cl^- ions present in the sample at equilibrium
d	amount (mol) of water present in the sample at equilibrium
K_1, K_2, K_3	dissociation constants of BOPA in water
K_0	dissociation constant of water
M_1	molar mass of the strong base (M)
M_{-1}	molar mass of the strong acid (M)
M_2	molar mass of BOPA in the neutral form (M)
$M_{\Delta 3}$	mass added to the molar mass of BOPA when associated with the counterion
δ	δ selects the titrant used: base, $\delta = +1$; acid, $\delta = -1$
$\epsilon\delta$	relative concentration of the solution of titrant (w/w)
ϵ_2	relative concentration of the stock solution of BOPA (w/w)
ρ_0	density of the stock solution of BOPA (g/L)
ρ	density of the sample (g/L)
ρ_δ	variation of the total density of the sample divided by the inverse of the partial
m	total mass of the sample (g)
m_δ	mass of the titrant added to the sample (g)
m_2	mass of the stock solution of BOPA added to the sample (g)
m_0	mass of the stock solution of BOPA in the sample without titrant (g)

LA980291D