Acid−Base Equilibrium Constants for Glycine in $NaClO_4$, KCl, and KBr at 298 K. Dependence on Ionic Strength

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The dependence of the acid−base equilibrium constants for glycine on the ionic strength of $NaClO_4$, KCl, and KBr solutions was systematically studied. The Guggenheim, Scatchard, Pitzer, and quasi-lattice models were used to express the activity coefficients for the species involved in the equilibrium.

Introduction

Glycine, the most simple α-amino acid, is specially significant as a structural prototype for other amino acids and proteins. As such, it has been the subject of much theoretical and experimental work. Recently, IUPAC (Kiss et al., 1991) revised its stability constants and, obviously, its acidity constants among them. The study clearly revealed that most such constants had been determined at a specific ionic constant and that the influence of this quantity on the acid−base equilibrium constants had scarcely been examined in systematic manner, particularly at moderate to high ionic strengths. One exception is the paper by Gergely et al. (1974), not cited in the previous revision.

Our group has for several years investigated the influence of the ionic medium on the equilibria of various amino acids. Specifically, we have studied glycine in both pure electrolytes (Herrero et al., 1993a,b; Fiol et al., 1994; Vilariño et al., 1997) and complex mixtures such as artificial seawater (Fiol et al., 1995). Moreover the effect of ionic strength of acid−base equilibria has also been reviewed (Sastre de Vicente, 1997).

In this work, we extended previous investigations to the equilibrium constants for glycine in various saline media at 298 K. We compare our results with those obtained by Gergely et al. and the data collected by Kiss et al. (1991).

Thermodynamic equilibrium constants (K^T) are related to stoichiometric constants (K^*) via activity coefficients based on expressions of the following type

$$K^* = K^T \prod_i (\gamma_i)^{\nu_i} \tag{1}$$

where the product is extended to each species i present in the equilibrium considered and ν_i is the stoichiometric coefficient for species i.

Equation 1 can be used in conjunction with some available models for the activity coefficients of the species present in an equilibrium to account for the experimental dependence of stoichiometric constants on ionic strength.

In this work, activity coefficients were expressed on the basis of the well-known Guggenheim, Scatchard, and Pitzer models, as well as the quasi-lattice model, to analyze the dependence of pK on I (see Appendix).

Experimental Section

All reagents used were Merck p.a. grade. Solutions were made in water purified by passage through a Millipore-Milli Q system.

Glycine was titrated by adding standarized KOH (0.1000 M) over a solution of the amino acid in HCl, both 0.0100 M; the necessary amount of inert electrolyte was added to all solutions in order to maintain a constant ionic strength. The electrode was calibrated prior to each titration by adding HCl to a solution of the supporting electrolyte, both with the same ionic strength, see Fiol et al. (1992). Calibrations were carried out for the p[H^+] range between 2.3 and 2.9, as recommended by May et al. (1982). The emf/p[H^+] data obtained were fitted to the equation $E = E^{\mathrm{ref}} - s\mathrm{p}[H^+]$, and the reference potential, E^{ref}, was obtained. The slope of the fit allowed us to check the behavior of the electrode. It was close to its theoretical value in every case ($s = 59.2$ mV).

Experiments were conducted in a dual-wall cell that was maintained at constant temperature by passing water at 25 ± 0.1 °C from a thermostat. Purified (99.999%) N_2 was bubbled through the solutions it in order to homogenize them and remove CO_2. The titrant was added with the aid of a Crison microBU 2030 autoburet furnished with 2.5 mL syringes; additions were computer-controlled. A Radiometer GK2401C electrode connected to a Crison micropH 2002 pH-meter, also interfaced to a computer, was used to measure pH values. This is a combined glass electrode with reference electrodes of silver/silver chloride; when the inert electrolyte was $NaClO_4$, the solution of the external reference electrode, which is in contact with the problem solution through a ceramic plug, was 0.95 M in $NaClO_4$ and 0.05 M in NaCl, instead of saturated KCl, to avoid precipitation of $KClO_4$.

Results and Discussion

The program Miniquad (Vacca and Sabatini, 1985) was used to calculate pK^* values. The molar data thus obtained were converted into molal values (the type used by the models). The molar−molal relation for KCl and KBr was taken from the literature (Weast, 1986), whereas for

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S0021-9568(98)00040-5 CCC: $15.00

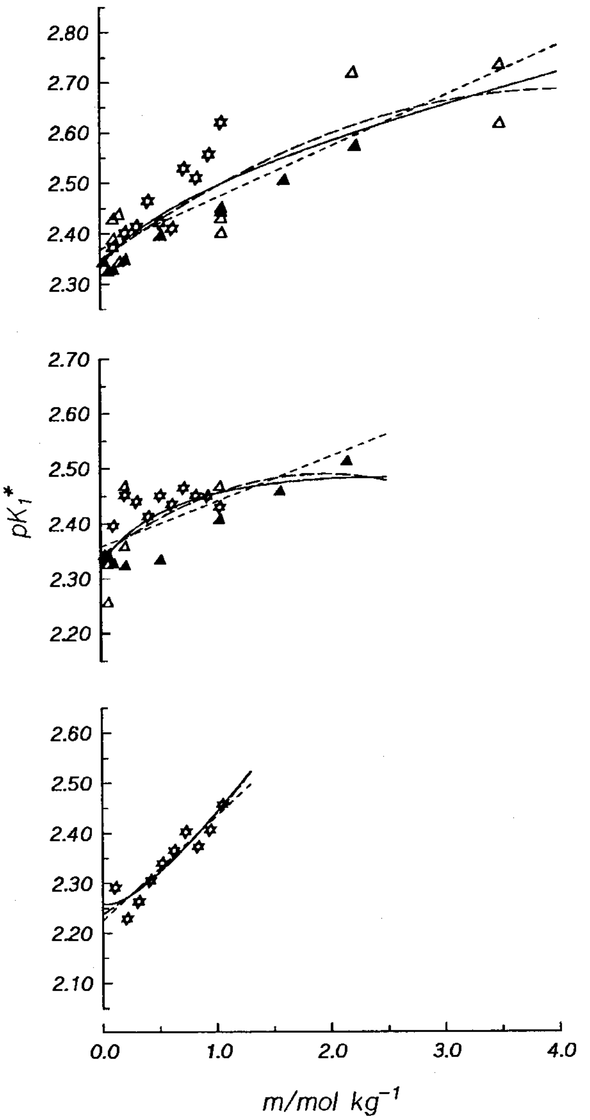

Figure 1. pK_1^* vs I according to several models for glycine in $NaClO_4$ (top graph), KCl (middle graph), and KBr (bottom graph), data from Table 1, are represented by stars, data from Kiss et al. (1991) by triangles, and data from Gergely et al. (1974) by solid triangles, at 25 °C. Models: - - -, Guggenheim; — — —, Scatchard; —, Pitzer.

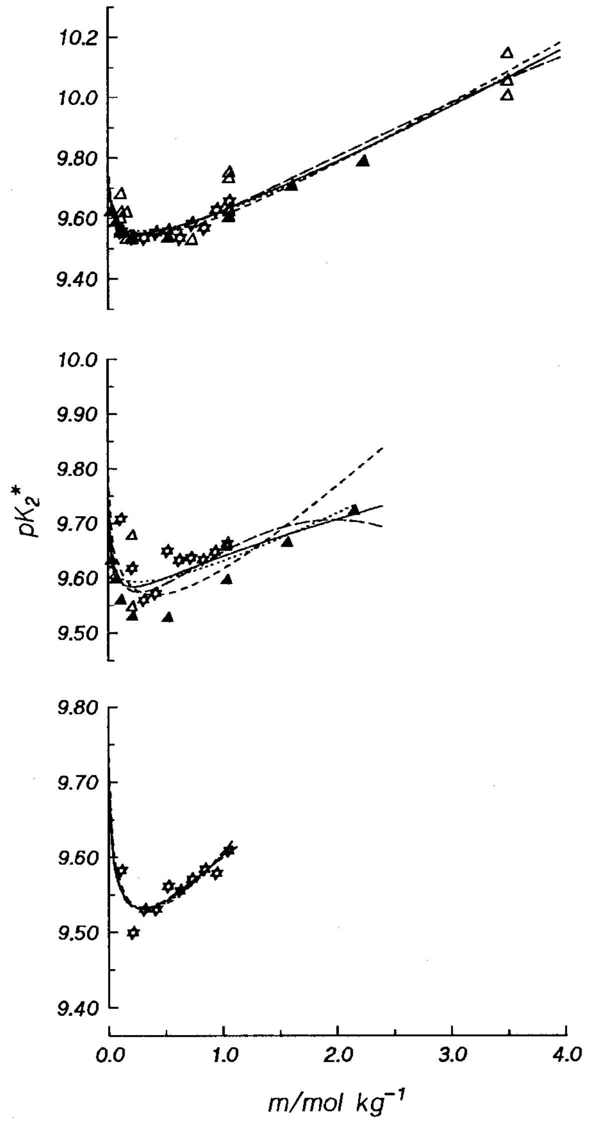

Figure 2. pK_2^* vs I according to several models for glycine in $NaClO_4$ (top graph), KCl (middle graph), and KBr (bottom graph), data from Table 1, are represented by stars, data from Kiss et al. (1991) by triangles, and data from Gergely et al. (1974) by solid triangles, at 25 °C. Models: - - -, Guggenheim; — — —, Scatchard; —, Pitzer; • • •, quasi-lattice.

Table 1. Molal pK_i^* Values for Glycine in (a) $NaClO_4$, (b) KCl, and (c) KBr at a Variable Ionic Strength

$NaClO_4$			KCl			KBr		
I	pK_1^*	pK_2^*	I	pK_1^*	pK_2^*	I	pK_1^*	pK_2^*
0.103	2.363	9.546	0.101	2.396	9.708	0.101	2.292	9.582
0.197	2.413	9.547	0.202	2.452	9.618	0.202	2.230	9.500
0.291	2.434	9.555	0.303	2.440	9.560	0.304	2.263	9.531
0.385	2.490	9.577	0.405	2.413	9.572	0.407	2.305	9.531
0.480	2.450	9.586	0.508	2.451	9.650	0.509	2.340	9.561
0.574	2.442	9.566	0.611	2.435	9.633	0.614	2.365	9.556
0.668	2.564	9.616	0.715	2.465	9.637	0.719	2.403	9.571
0.762	2.548	9.606	0.820	2.451	9.634	0.824	2.373	9.584
0.856	2.597	9.668	0.925	2.450	9.648	0.931	2.407	9.579
0.951	2.663	9.698	1.031	2.430	9.663	1.039	2.457	9.609

$NaClO_4$ the equation of density as a function of molality was taken from Chen and Chopin (1995). The conversion was performed for our data and data taken from literature.

The molal pK_1^* and pK_2^* values thus obtained (Table 1) were subject to average errors of 0.001−0.050 units for the measurements made in the three electrolytes. Our data are represented, together with data taken from Kiss et al. (1991) and Gergely et al. (1974) in Figures 1 and 2.

All these values were fitted to ionic strength functions, using the different models described in the Appendix, so we obtain equations to explain the behavior of pK^* vs I, for pK_1^* and pK_2^* of glycine in three electrolytes (KCl, $NaClO_4$, and KBr). The parameters for the used models are given in Table 2.

From the graphs obtained by applying the different models (Figures 1 and 2) and the analysis of the tabulated data, it follows that all the fitting equations tested accurately reproduced the experimental facts.

As can be seen, some parameters were subject to large errors, which can be ascribed to the dispersion between the experimental data from the different authors as to possible problems of multicollinearity, also commonly termed collinearity or ill-conditioning (Fiol et al., 1994; Herrero et al., 1993c). However, the error of the fit (σ_1 and σ_2 for the first and second equilibrium, respectively) was in most cases within accepted limits (0.05 pK^* units) for this type of measurement (Albert and Serjeant, 1984).

Table 2. Interaction Parameters and pK^T Values for Glycine in $NaClO_4$, KCl, and KBr Obtained by Use of Guggenheim's, Scatchard's, Pitzer's Theories and Quasi-Lattice Model

	$NaClO_4$	KCl	KBr
Guggenheim Model (Eqs A5 and A7)			
pK_1^T	2.369 ± 0.015	2.358 ± 0.014	2.225 ± 0.019
ϵ_1	0.102 ± 0.012	0.082 ± 0.019	0.209 ± 0.029
σ_1	0.061	0.050	0.028
pK_2^T	9.784 ± 0.011	9.818 ± 0.017	9.760 ± 0.012
ϵ_2	0.229 ± 0.009	0.206 ± 0.022	0.254 ± 0.019
σ_2	0.050	0.050	0.018
Scatchard Model (Eqs A9 and A10)			
pK_1^T	2.347 ± 0.018	2.341 ± 0.018	2.238 ± 0.034
A_1^S	0.165 ± 0.036	0.151 ± 0.053	0.149 ± 0.137
B_1^S	−0.020 ± 0.011	−0.039 ±0.027	0.053 ± 0.117
σ_1	0.058	0.048	0.029
pK_2^T	9.769 ± 0.015	9.783 ± 0.019	9.744 ± 0.021
A_2^S	0.277 ± 0.028	0.350 ± 0.055	0.329 ± 0.086
B_2^S	−0.014 ± 0.008	−0.079 ± 0.028	−0.066 ± 0.074
σ_2	0.047	0.048	0.018
Pitzer Model (Eqs A15 and A16)			
pK_1^T	2.332± 0.025	2.320 ± 0.022	2.260 ± 0.049
A_1^P	0.029 ± 0.013	−0.010 ± 0.026	0.165 ± 0.079
B_1^P	0.176 ± 0.095	0.259 ± 0.127	0.249 ± 0.320
σ_1	0.058	0.046	0.029
pK_2^T	9.756 ± 0.019	9.755 ± 0.025	9.744 ± 0.032
A_2^P	0.250 ± 0.018	0.101 ± 0.055	0.255 ± 0.105
B_2^P	0.317 ± 0.067	−0.607 ± 0.140	0.302 ± 0.213
σ_2	0.046	0.047	0.019
Quasi-Lattice Model (Eq A24)			
pK_2^T	9.709 ± 0.038	9.663 ± 0.055	9.738 ± 0.090
a	- 0.338 ± 0.069	- 0.161 ± 0.110	−0.450 ± 0.185
b	0.250 ± 0.023	0.129 ± 0.056	0.317 ± 0.104
σ	0.047	0.047	0.020

Acknowledgment

The authors wish to thank Dr. T. Kiss for his useful comments.

Appendix

The acid base equilibrium for an amino acid containing two dissociating groups (e.g., glycine) can be expressed as (Albert and Serjeant, 1984; Scatchard, 1943)

$$AH_2^{\ +} \leftrightarrows AH + H^+ \tag{A1}$$

$$AH \leftrightarrows A^- + H^+ \tag{A2}$$

where AH denotes the zwitterion, AH_2^+ the protonated amino acid molecule, and A^- its anion. These equilibria have constants such as

$$K_1^T = \frac{(AH)(H^+)}{(AH_2^{\ +})} = \frac{\gamma_{AH}\gamma_{H^+}}{\gamma_{AH_2^+}}\frac{[AH][H^+]}{[AH_2^{\ +}]} = K_1^\gamma\, K_1^* \tag{A3}$$

$$K_2^T = \frac{(A^-)(H^+)}{(AH)} = \frac{\gamma_{A^-}\gamma_{H^+}}{\gamma_{AH}}\frac{[A^-][H^+]}{[AH]} = K_2^\gamma\, K_2^* \tag{A4}$$

where K^T is the thermodynamic constant for the amino acid, which depends on the temperature, pressure, and solvent; K^* its stoichiometric constant; and K^γ a constant that includes the influence of activity coefficients. The latter two constants depend on the temperature, pressure, solvent, and ionic strength of the solution.

By replacing the activity coefficients, γ_i, with the ionic strength function of the Guggenheim, Scatchard, Pitzer, and quasi-lattice models, one can obtain the pK_1^*-I and pK_2^*-I functions for any amino acid.

The Guggenheim expression for the first equilibrium (Brandariz et al., 1993; Salvatore et al., 1986) is a linear function of the ionic strength, the intercept of which gives the thermodynamic pK value

$$pK_1^* = pK_1^T + \epsilon_1 I \tag{A5}$$

where

$$\epsilon_1 = \frac{2}{\ln 10}[\lambda - \beta_{AH_2^+,X^-} + \beta_{H^+,X^-}] \tag{A6}$$

The second equilibrium cannot be represented by a linear function owing to the presence of the Debye−Hückel term

$$pK_2^* = pK_2^T - 2A'\frac{\sqrt{I}}{1 + c\sqrt{I}} + \epsilon_2 I \tag{A7}$$

where

$$\epsilon_2 = \frac{2}{\ln 10}[-\lambda + \beta_{A^-,K^+} + \beta_{H^+,X^-}] \tag{A8}$$

The Scatchard expressions for the first and second equilibrium are similar to the Guggenheim expressions albeit more complex as they include serial expansions of the quadratic term for the logarithms of the mean ion activity coefficients as a function of the ionic strength, $\gamma_\pm$ (Brandariz et al., 1993; Scatchard, 1961).

$$pK_1^* = pK_1^T + A_1^S I + B_1^S I^2 \tag{A9}$$

$$pK_2^* = pK_2^T + \frac{2S\sqrt{I}}{\ln 10(1 + a_J\sqrt{I})} + A_2^S I + B_2^S I^2 \tag{A10}$$

where

$$A_1^S = [-a_{AH_2X}^{(1)} - b_{AH_2X,XK}^{(0,1)} + a_{HX}^{(1)} + b_{HX,XK}^{(0,1)} + h_0]\,\frac{1}{\ln 10} \tag{A11}$$

$$B_1^S = [-a_{AH_2X}^{(2)} + a_{HX}^{(2)} - b_{AH_2X,XK}^{(0,2)} + b_{HX,XK}^{(0,2)} - b_{HX,XK}^{(1,2)} + b_{AH_2X,XK}^{(1,2)}]\,\frac{1}{2\ln 10} \tag{A12}$$

$$A_2^S = [a_{AH}^{(1)} + b_{AH,XK}^{(0,1)} + a_{KX}^{(1)} - h_0]\,\frac{1}{\ln 10} \tag{A13}$$

$$B_2^S = [a_{AH}^{(2)} + 2a_{KX}^{(2)} + b_{AH,XK}^{(0,2)} - b_{AH,XK}^{(1,2)}]\,\frac{1}{2\ln 10} \tag{A14}$$

The Pitzer expressions (Pitzer, 1991; Brandariz et al., 1993) are mathematically more complex owing to the presence of exponential functions and the new term included in the equation for the limiting law, f^γ, for pK_2^*

$$pK_1^* = pK_1^T + 2IA_1^P + gB_1^P \tag{A15}$$

$$pK_2^* = pK_2^T + \frac{2f^\gamma}{\ln 10} + IA_2^P + gB_2^P + \frac{2I}{\ln 10}\beta_{XK}^1\, e^{-2\sqrt{I}} \tag{A16}$$

where

$$g = 1 - (1 + 2\sqrt{I})e^{-2\sqrt{I}} \tag{A17}$$

$$A_1^{\mathrm{P}} = -2\{\beta^0_{\mathrm{XAH_2}} - \lambda_{\mathrm{AH,XK}} - \beta^0_{\mathrm{HX}}\}/\ln 10 \tag{A18}$$

$$B_1^{\mathrm{P}} = -\{\beta^1_{\mathrm{XAH_2}} - \beta^1_{\mathrm{HX}}\}/\ln 10 \tag{A19}$$

$$A_2^{\mathrm{P}} = 2\{\beta^0_{\mathrm{AK}} - \lambda_{\mathrm{AH,XK}} - \beta^0_{\mathrm{HX}}\}/\ln 10 \tag{A20}$$

$$B_2^{\mathrm{P}} = \{\beta^1_{\mathrm{AK}} + \beta^1_{\mathrm{HX}} - \beta^1_{\mathrm{XK}}\}/\ln 10 \tag{A21}$$

$$f^{\gamma} = -A_{\Phi}\left[\frac{\sqrt{I}}{1 + 1.2\sqrt{I}} + \frac{2}{1.2}\ln(1 + 1.2\sqrt{I})\right] \tag{A22}$$

By contrast, the quasi-lattice model (Vilariño et al., 1997; Chen and Choppin, 1995; Pitzer, 1991) is an alternative to empirical and semiempirical equations based on the limiting law. The model assumes a lattice structure for dissolved ions based on diffraction studies on concentrated electrolyte solutions. It uses the following relation between log $\gamma_{\pm}$ and the cubic root of the ionic strength

$$\log \gamma_{\pm} = AI^{1/3} + BI \tag{A23}$$

where A and B are two constants associated with the model.

The quasi-lattice expression for the second equilibrium is as follows

$$\mathrm{p}K_2^* = \mathrm{p}K_2^{\mathrm{T}} + aI^{1/3} + bI \tag{A24}$$

where $a = 2A$, $b = 2B - c$, and $\log \gamma_{\mathrm{AH}} = c{\cdot}I$.

More detailed applications of these models for acid−base equilibria are provided in several papers published by our group (e.g., Brandariz et al., 1993; Vilariño et al., 1997 and references therein).

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Received for review February 3, 1998. Accepted June 20, 1998. The financial support of the Xunta de Galicia (Project XUGA10304B97) is gratefully acknowledged.

JE9800407