J. Chem. Eng. Data **2004,** *49,* 1193−1196

Effect of Ionic Strength, Aquo-Organic Solvents, and Temperature on the Stabilities of *N*-[Tris(hydroxymethyl)methyl]glycine + Metal Complexes

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The stability constants of complexes of *N*-[(tris-hydroxymethyl)methyl]glycine (Tricine) with the transition metal ions Co^{2+}, Ni^{2+}, Cu^{2+}, and Zn^{2+} were studied potentiometrically in the presence of (0.02, 0.05, 0.10, and 0.15) mol dm^{-3} KNO_3; in the presence of (60, 65, and 75) vol % acetone, 75 vol % ethanol, and 75 vol % dioxane at 0.05 mol dm^{-3} KNO_3 at 25 °C; and at (5, 15, 25, 35, 45) °C in the presence of 0.1 mol dm^{-3} KNO_3. The thermodynamic stability constants for tricine and for tricine + M(II) [where M(II) = Co^{2+}, Ni^{2+}, Cu^{2+}, and Zn^{2+}) have been obtained from the measurements.

Introduction

N-[Tris(hydroxymethyl)methyl]glycine, $(HOCH_2)_3CNHCH_2COOH$ (tricine),[1,2] is a useful biological buffer of pH range 7.2−8.5 that is of considerable physiological and chemical interest.[3−7] Metal + ligand complexes of tricine have been reported.[8−11] A mathematical model for calculating apparent acid dissociation constants (pK_a) in aqueous organic mixtures with respect to the concentration of organic solvent in a binary mixture has been proposed.[12] The behavior of tricine in mixed aqueous solvents and the effect of the medium on the ionization process of tricine have been reported.[13,14] Mixed solvents have been used for studying complex reactions and developing a theoretical understanding of the ionization process in these media.[15−18] In a continuation of our work on tricine,[14] the stability constants of metal + tricine binary complexes (where the metals were Th^{4+}, La^{3+}, Ce^{3+}, and UO_2^{2+})[19] and ternary systems of metal + tricine + thiazolidines (where the metals were Co^{2+}, Ni^{2+}, and Cu^{2+}) were studied.[20]

The present paper deals with the measurement of the stability constants of the binary complexes of Co^{2+}, Ni^{2+}, Cu^{2+}, and Zn^{2+} with tricine (i) in the presence of (0.02, 0.05, 0.1, and 0.15) mol dm^{-3} KNO_3 and (ii) in the presence of acetone (60, 65, and 75 vol %), ethanol (75 vol %), and dioxane (75 vol %) at $I = 0.05$ mol dm^{-3} KNO_3 at 25 °C.

The thermodynamic parameters (ΔG, ΔH, and ΔS) were evaluated in the presence of 0.1 mol dm^{-3} KNO_3 in pure aqueous media in the temperature range of 5−45 °C. The Irving and Rossotti[21,22] pH-titration technique was used to determine the acid dissociation constants as well as the formation constants for the various complexes.

Experimental Section

Materials and Solutions. *N*-[Tris(hydroxymethyl)methyl]glycine (tricine) was a Sigma A. R. product and was used without further purification. The organic solvents were ethanol (Merck, 99.5%), acetone (BDH, 99.5%), and dioxane (Aldrich, 99.5%), which were purified by the method described by Vogel.[23] Solutions of the bivalent metal ions Co^{2+}, Ni^{2+}, Cu^{2+}, and Zn^{2+} were prepared from their nitrate salts (chloride in the case of Zn^{2+}). A stock solution of tricine was prepared by dissolving a known mass in bidistilled water. NaOH solution (∼1 mol dm^{-3}) was prepared by dissolving a known mass of the Analar pellets in CO_2-free bidistilled water, and the solution was standardized with a standard potassium hydrogen phthalate. HNO_3 solution (∼0.06 mol dm^{-3}) was prepared and used after standardization. KNO_3 was used as a supporting electrolyte. Solutions of lower concentrations were prepared by appropriate volumetric dilutions of the corresponding stock solutions.

pH-metric Titration. The experimental procedure involved the titration of the following solutions (total volume = 50 cm^3) against a standard CO_2-free (0.1846 mol dm^{-3}) NaOH solution: (a) 3 cm^3 of HNO_3 (0.055 mol dm^{-3}) + 2.5 cm^3 of KNO_3 (2 mol dm^{-3}), (b) solution a + 1 cm^3 of tricine (0.1 mol dm^{-3}), and (c) solution b + 1 cm^3 of metal solution (0.1 mol dm^{-3}).

These titrations were repeated for (i) ionic strengths $I = (0.02, 0.05, 0.1, \text{and } 0.15$ mol $dm^{-3})$ KNO_3 at 25 °C, (ii) various temperatures (5, 15, 25, 35, and 45 °C) at $I = 0.1$ mol dm^{-3} KNO_3, and (iii) various aquo-organic media (60, 65, and 75 vol % acetone; 75 vol % ethanol; and 75 vol % dioxane). The NaOH solution used in experiment iii contained the same volume percentage of the appropriate organic solvent.

The pH measurements were carried out using a Fisher Scientific, Acumen pH Meter model 825 MAP with a glass combined calomel electrode. The water thermostat was a Fisher Scientific Isotemp Refrigerated Circulator model 9000 accurate to ±0.1 °C. The solutions were equilibrated in the thermostat for about 15 min before titration. The equations[21,22] used for the various calculations were programmed into an IBM computer.

Results and Discussion

Titration curves were obtained for the titrations of Co^{2+}, Ni^{2+}, Cu^{2+}, and Zn^{2+} in the presence of different molarities of KNO_3, in the presence of different solvents, and at different temperatures. No precipitates formed, indicating that there was no tendency to form hydroxo complexes as the number of moles of NaOH consumed was equivalent to the number of moles of HNO_3 and tricine. The stability

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10.1021/je030228c CCC: $27.50 © 2004 American Chemical Society
Published on Web 08/04/2004

Table 1. Stability Constants of Bivalent Metal Complexes with Tricine at (25 ± 0.1) °C at Different Ionic Strengths *I* in Pure Aqueous Medium

		I			
		(mol dm^{-3} KNO$_3$)			
system	stability constant	0.15	0.10	0.05	0.02
tricine	$\log K^{H}$	8.02 ± 0.07	8.20 ± 0.02 (8.18)^{b}	8.41 ± 0.14	8.74 ± 0.05
Cu^{II} + tricine	$\log K^{Cu}_{Cu+tricine}$	7.01 ± 0.03	7.70 ± 0.08 (7.70),^{a} (7.30)^{b}	8.00 ± 0.09	8.51 ± 0.12
Ni^{II} + tricine	$\log K^{Ni}_{Ni+tricine}$	5.11 ± 0.02	5.72 ± 0.07 (5.72)^{a} (5.45)^{b}	5.95 ± 0.06	6.26 ± 0.07
Co^{II} + tricine	$\log K^{Co}_{Co+tricine}$	4.06 ± 0.08	4.51 ± 0.12 (4.51),^{a} (4.71)^{b}	4.66 ± 0.14	4.89 ± 0.10
Zn^{II} + tricine	$\log K^{Zn}_{Zn + tricine}$	4.94 ± 0.10	5.11 ± 0.04 (5.29)^{b}	5.32 ± 0.06	5.68 ± 0.05

[a] From ref 24. [b] From ref 9.

constants of the formed complexes were investigated in the pH range of 4−6. Also, the calculated stability constants of copper + tricine complexes (7.01−8.51) are higher than that of Cu(OH)$^{+}$ (log K = 6.1), indicating that the formed complexes are more stable than the soluble hydroxo species. However, a buffer zone was observed at high pH values, indicating the deprotonation of the amide group of tricine, which is not involved in chelation and is ionized at higher pH values of hydrolysis region.[9,24]

The mean values of n_H (the average number of protons associated with the ligand) for tricine at different pH values were calculated from the titration curves of solutions i and ii by employing the relationship derived by Irving and Rossotti.[21,22] On plotting $\log[n_H/(1 - n_H)]$ vs pH, a straight line having an intercept equal to pK_a on the pH axis is obtained. From the titration curves of solutions i−iii, the metal + ligand formation number values $\bar{n}$ (the average number of ligand molecules coordinated to the metal ion) of the metal complexes were obtained at various pH values. The $\bar{n}$ values were found to be less than or equal to 1, indicating the formation of 1:1 complexes. The $\bar{n}$ values were calculated using the Irving and Rossotti formulation.[21,22] The free ligand exponent pL was calculated from the so-obtained values of $\bar{n}$ by the equation

$$pL = \log\left[\frac{\sum_{y=0}^{y=1}\beta_y^{H}\left(\frac{1}{10^{B}}\right)}{C_L - n_H C_M}\,\frac{V_o + V_c}{V_o}\right]$$

Here, C$_M$ is the concentration of ions M^{n+} used, C_L is the concentration of the ligand, y is the number of dissociable protons (y = 1 for tricine), and V_o is the original volume (50 cm^3). V_c is the volume of alkali (NaOH) consumed to reach the same pH values in curve c corresponding to the titration of solution iii. β_y^{H} is the formation constant values of tricine, and B is the pH value. The mean pK_a and stability constant values obtained from the corresponding different experimental formation curves using the average value and straight line methods along with the error as obtained by least-squares fits are reported in Tables 1−4.

Effect of Ionic Strength on Stability Constants. The values of pK_a and $\log K^{M}_{M+tricine}$ were found to decrease with increasing ionic strength of the medium (Table 1) in agreement with the Debye−Hückel equation.[25] The thermodynamic stability constants are obtained by extrapolating the straight line plot of pK_a and/or $\log K^{M}_{M+tricine}$ versus $\sqrt{I}$ to zero ionic strength. The log $K°$ (thermodynamic stability constant) values were (9.18, 6.95, 5.40, 9.23, and 5.98) for tricine and M + tricine binary complexes, where M = Co^{2+}, Ni^{2+}, Cu^{2+}, and Zn^{2+}, respectively (Figure 1). The equation $\Delta G = -2.303\,RT \log K°$ gives the relationship between the thermodynamic stability constant and the free energy change according to complex formation. ΔG values were (−52.38, −39.66, −30.79, −52.68, and −34.10) kJ·mol^{-1} for tricine and M + tricine binary complexes, with M = Co^{2+}, Ni^{2+} Cu^{2+}, and Zn^{2+}, respectively.

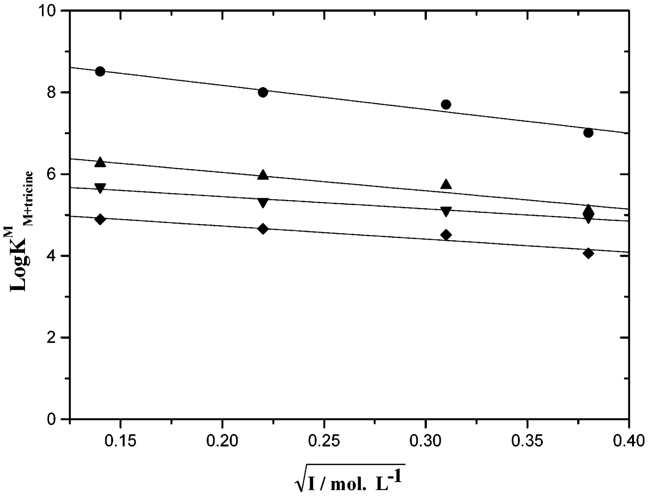

Figure 1. Plots of $\log K^{M}_{M+tricine}$ vs $\sqrt{I/\text{mol}\cdot\text{L}^{-1}}$, where M(II) = ◆, Co; ▲, Ni; ●, Cu; ▼, Zn.

Effect of Solvents on Stability Constants. In mixed aqueous solvents, the proton ligand formation constant is influenced by factors such as (a) the relative permittivity of the mixed solvent, (b) the change in hydrogen-bonding ability of organic solvents in water, and (c) the protonation of organic solvents. Bates[26] and Rorabacher[27] have explained the change of log K_H with solvent composition considering both electrostatic and nonelectrostatic effects. They concluded that the nonelectrostatic phenomenon become increasingly important in solvents containing an organic solvent at more than 50% concentration. Protonation is viewed[27] as a two-step process:

$$\underset{\text{solvent−separated ion pair}}{H^{+} + L^{-b} \rightleftharpoons H^{+}\cdots L^{-b} \rightleftharpoons HL^{+(1-b)}} \qquad (1)$$

where

$$K_H = K_{os}k_{H-L} \qquad (2)$$

The values of K_{os} represent the diffusion-controlled equilibrium constant for ion pair formation and can be calcu-

Table 2. Mean pK_a Values of Tricine in Different Organic Solvent (1) + Water (2) I = 0.05 Mol dm^{-3} KNO$_3$ and at 25 ± 0.1 °C

$100\phi_1{}^a$	$\chi_1{}^a$	$10^{-2}/(\epsilon \times 10^{-4})^b$	pK_a
–	–	1.276	8.41 ± 0.14
60	0.218	1.52	8.69 ± 0.06
65	0.247	1.55	8.81 ± 0.12
75	0.317	1.65	8.98 ± 0.08
75	0.363	1.69	8.65 ± 0.05
75	0.420	2.2	9.22 ± 0.14

[a] χ_1 is the mole fraction and ϕ_1 is the volume fraction on a solvent + water basis. [b] ϵ is the relative permittivity.

lated from the Fuoss equation[28]

$$K_{os} = \frac{4}{3}\pi a^3 N_A 10^{-3} \exp\left(\frac{e^2}{\epsilon a \kappa T}\right) \tag{3}$$

where a is the distance of closest approach between the solvated proton and base in the ion pair, e is the electron charge, κ is Boltzmann's constant, and N_A is Avogadro's number. The term k_{H-L} in eq 2 represents the proton jump equilibrium occurring within the ion pair, presumably a nonelectrostatic term that primarily depends on the relative basicity of the solvent and the base with a contribution from the nature and orientation of the solvent molecules separating the solvated proton and base species in the ion pair. It also appears that the nonelectrostatic part of the interaction is related to the proton-accepting property of the medium and the proton solvation of the organic solvent. It is understood that, with decreasing relative permittivity ϵ = (2.2, 20.7, and 24.3) for dioxane, acetone and ethanol, respectively,[29] log K_H increases.

The mole fractions, relative permittivities of the mixed solvents, and pK_a values of tricine in mixed aqueous solvents at 25 °C and 0.05 mol dm^{-3} KNO$_3$ are reported in Table 2. According to Braude and Stern,[30] the tetrahedral lattice structure of water is gradually broken down upon the addition of an organic solvent, and because of the denser packing and lower extent of hydrogen bonding between water molecules, the stability of the hydroxonium ion increases, and the proton-donating property of the medium falls. This might imply that the proton-accepting property of the solvent increases. It is also said that the hydrogen-bond structure is less prevalent in pure ethanol in comparison to water and largely absent in pure acetone and pure dioxane. Thus, the extent of hydrogen bonding in ethanol + water is greater than that in dioxane + water or acetone + water.

The foregoing solvent effects influence the pK_a of ligand conjugate acid in the following manner: (i) With decreasing solvent relative permittivity, the pK_a of the ligand increases and vice versa. (ii) As the extent of hydrogen bonding in water is decreased by the organic solvent, the proton-accepting property of water increases, so the pK_a of the ligand decreases. (iii) Increasing proton solvation by the organic solvent decreases the pK_a of the ligand and vice versa.

From Table 2, for a giving mixed aqueous solvent (acetone + water; 60, 65, and 75 vol %), it is clear that the pK_a of the ligand increases with increasing organic content of the aquo-organic solvent.

Table 3. Stability Constants (log $K^{M}_{M+tricine}$) of Bivalent Metal Ion Complexes with Tricine in the Presence of Different Amounts (vol %) of Solvents at I = 0.1 mol dm^{-3} KNO$_3$ and T = (25 ± 0.1) °C

	vol % of acetone				
metal ion	60	65	75	75 vol % of ethanol	75 vol % of dioxane
Cu(II)	7.77 ± 0.05	7.95 ± 0.07	8.14 ± 0.04	7.98 ± 0.09	8.35 ± 0.70
Ni(II)	5.93 ± 0.09	6.04 ± 0.15	6.22 ± 0.10	6.12 ± 0.06	6.37 ± 0.14
Co(II)	4.81 ± 0.12	4.95 ± 0.23	5.19 ± 0.19	5.10 ± 0.11	5.39 ± 0.20
Zn(II)	5.45 ± 0.07	5.63 ± 0.10	5.84 ± 0.08	5.70 ± 0.05	5.97 ± 0.15

Table 4. Stability Constants of Bivalent Metal Complexes with Tricine in the Presence of I = 0.1 mol dm^{-3} KNO$_3$ and Some Thermodynmic Parameters at Different Temperatures

system	T (°C)	stability constant	ΔH (kJ·mol^{-1})	$-\Delta S$ (J·mol^{-1}·K^{-1})	ΔG (kJ·mol^{-1})
tricine	5	8.62 ± 0.02	40.96	17.70	45.90
	15	8.46 ± 0.05		19.75	46.65
	25	8.20 ± 0.02		19.54	46.78
	35	7.98 ± 0.03		19.80	47.07
	45	7.75 ± 0.01		19.85	47.20
log $K^{Cu}_{Cu+tricine}$	5	8.08 ± 0.03	−41.17	+6.57	−43.01
	15	7.96 ± 0.07		+9.46	−43.89
	25	7.70 ± 0.08		+9.29	−43.93
	35	7.43 ± 0.04		+8.58	−43.81
	45	7.22 ± 0.08		+8.74	−43.97
log $K^{Ni}_{Ni+tricine}$	5	6.30 ± 0.07	−42.51	32.30	−33.51
	15	5.89 ± 0.05		34.81	−32.47
	25	5.72 ± 0.09		33.14	−32.64
	35	5.53 ± 0.04		32.13	−32.59
	45	5.37 ± 0.12		30.84	−32.68
log $K^{Co}_{Co+tricine}$	5	4.81 ± 0.05	−32.34	24.23	−25.61
	15	4.66 ± 0.07		23.10	−25.69
	25	4.51 ± 0.12		22.18	−25.73
	35	4.29 ± 0.15		22.89	−25.31
	45	4.15 ± 0.08		22.26	−25.27
log $K^{Zn}_{Zn+tricine}$	5	5.46 ± 0.16	−33.30	15.27	−29.08
	15	5.25 ± 0.07		15.10	−28.95
	25	5.11 ± 0.04		13.93	−29.16
	35	4.95 ± 0.06		13.35	−29.20
	45	4.78 ± 0.08		13.22	−29.12

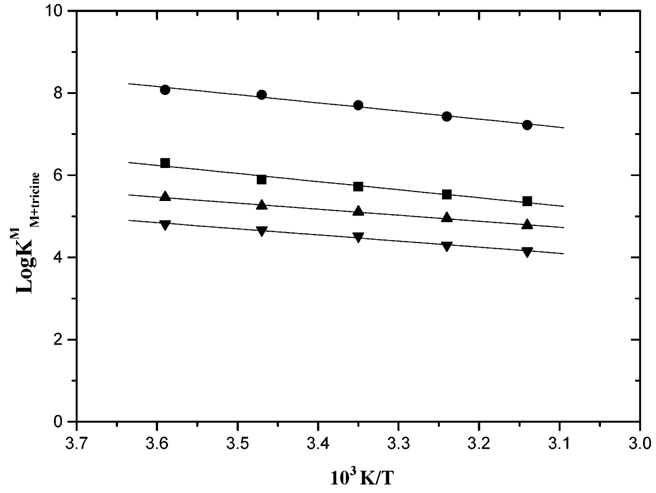

Figure 2. log $K^{\mathrm{M}}_{\mathrm{M+tricine}}$ vs 10^3 K/T plots in the presence of 0.01 mol dm^{-3} KNO_3, where M(II) = ▼, Co; ■, Ni; ●, Cu; ▲, Zn.

The data in Table 3 show that the metal + ligand formation constants increase with increasing percentage of organic solvent in the medium, i.e., acetone + water (75 vol %) > acetone + water (65 vol %) > acetone + water (60 vol %).

For a particular composition (75 vol %) of the solvent and water mixture, the $1/\epsilon$ values of the various aquo-organic solvents vary in the following order: dioxane + water > acetone + water > ethanol + water.

The stability constants of metal + tricine complexes and pK_a values follow the same order; see Tables 2 and 3.

The apparent increase in stability of the complex in an aqueous mixed solvent compared to that in water alone could be related to the lower concentration of water in the mixed solvent compared to that in pure water. Thus, the solvation of the reacting species is controlled by water and cosolvent molecules.

Effect of Temperature on Stability Constants. From the pK_a and log $K^{\mathrm{M}}_{\mathrm{M+tricine}}$ values and their temperature dependence, the values of the thermodynamic functions ΔG, ΔH, and ΔS were derived.[31]

The values of stability constants in Table 4 reveal that the stability constants decrease with increasing temperature, along with the pK_a value (Figure 2).

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Received for review September 30, 2003. Accepted June 17, 2004.

JE030228C