Potentiometric Studies on the Binary and Ternary Complexes of Copper(II) Containing Dipicolinic Acid and Amino Acids

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The binary and ternary complexes of Cu(II) with dipicolinic acid as a primary ligand and some selected mono- and dicarboxylic amino acids as biologically important secondary ligands were studied potentiometrically. The acidity constants of the ligands were measured and used for determining the stability constants of the complexes formed in aqueous solutions at different temperatures at an ionic strength of 0.1 mol dm^{-3} $NaNO_3$. The formation of 1:1:1 ternary complexes is inferred from the potentiometric titration curves. The order of stability of the binary or ternary complexes in terms of the nature of the amino acid is investigated and discussed. The values of $\Delta \log K$ for the binary and ternary complexes involving amino acids have been evaluated along with the thermodynamic parameters $\Delta H°$, $\Delta S°$, and $\Delta G°$.

Introduction

Pyridinecarboxylic acids are of great interest to medicinal chemists because of the wide variety of their physiological properties displayed by the natural as well as synthetic acids. These acids are present in many natural products such as alkaloids, vitamins, and coenzymes. Pyridinecarboxylic acid metal complexes are therefore, especially interesting model systems. Over the years there has been a growing interest in studying the binary and ternary complexes involving pyridinecarboxylic acids (Napoli, 1968; Napoli and Magri, 1987; Seleim et al., 1987; Shelke and Jahagirdar, 1979a,b; Ullah and Bhattacharya, 1991; Choppin et al., 1992; Won et al., 1994; Kidani et al., 1976; Casassas et al., 1998). Amino acids and their metal complexes have frequent utilization in biological and chemical applications (Bagger, 1987; Ho and Yut, 1982; Crans et al., 1989; Wright et al., 1988). Recently, Khalil et al. (1997) studied the complexation equilibria and determination of stability constants of binary and ternary complexes of some alkaline earth and transition metal(II) ions involving dipicolinic acid (2,6-pyridinedicarboxylic acid) and glycine using potentiometric and conductometric techniques. The values of equilibrium constants are evaluated at 25 °C and ionic strength $I = 0.1$ mol dm^{-3} $NaNO_3$.

As a continuation of our research program oriented to the solution equilibria of binary and ternary complexes of biological importance (Khalil et al., 1985, 1997; Khalil et al., 1994; Khalil and Radalla, 1998), this paper reports on the complex formation of Cu(II) binary and ternary complexes involving dipicolinic acid as a primary ligand and some selected mono- and dicarboxylic amino acids as secondary ligands. The stability constants of the investigated complexes are studied at 25, 35, 45, and 55 °C and the corresponding thermodynamic parameters are evaluated, by the temperature coefficient method, and discussed.

Experimental Section

Materials and Solutions. Dipicolinic acid was an A. R. product (Fluka). The reagent was repeatedly recrystallized from water, dried at 115 °C, and checked by its melting point (250−255 °C). Stock solutions were prepared by dissolving precisely weighed amounts of the anhydrous acid in suitably bidistilled water. Glycine, norvaline, aspartic acid, and glutamic acid were also provided by Fluka and used without further purification. The metal salt was provided by BDH as nitrates. All solutions of Cu(II) metal ions were prepared and standardized complexometrically by EDTA (Welcher, 1965). Carbonate-free sodium hydroxide (titrant, prepared in 0.1 mol dm^{-3} $NaNO_3$ solution) was prepared by dissolving the Analar pellets in CO_2-free bidistilled water, and the solution was standardized potentiometrically with KH-phthalate (Merck AG). A HNO_3 solution (≈0.04 mol dm^{-3}) was prepared and used after standardization. A total of 0.5 mol dm^{-3} of $NaNO_3$ was used as a supporting electrolyte. HNO_3, NaOH, and $NaNO_3$ were from Merck p.a.

Apparatus. Potentiometric pH measurements were made on solutions in a double-walled glass vessel using a Griffin pH J-300-010 G digital pH meter. The temperature was controlled by circulating water through the jacket, from a constant-temperature bath. The cell was equipped with a magnetic stirrer and a tightly fitting rubber stopper, through which an Amel 882 delivery dispenser, readable to 1 μL, and electrode system were inserted. The electrode system was calibrated in terms of hydrogen ion concentrations instead of activities. It is to be assumed that the activity coefficient is constant, an assumption usually justified by working in a medium of high ionic strength (Ringbom, 1963). The electrode system was calibrated by periodic titrations of HNO_3 (or NaOH) solution (0.1 mol dm^{-3} in $NaNO_3$) with standard NaOH (or HNO_3) solution, under the same temperature as that used for the test solution titration. The resulting titration data were used to calculate the standard electrode potential $E°$ and the dissociation constant of water. These values were then used in the calculation of hydrogen ion concentration from potential readings. Thus, all constants determined in this work are concentration constants.

Procedure. The following solutions were prepared (total volume 50 cm^3) and titrated potentiometrically against standard carbonate-free NaOH (0.2003 mol dm^{-3}) solution:

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Table 1. Potentiometric Titration Data for the Cu^{2+}−DPA−Aspartic Acid System at 25 °C and I = 0.1 mol dm^{-3} $NaNO_3$

V/cm^3 NaOH	pH					
	a	b	c	d	e	f
0.00	2.14	2.14	2.14	2.20	2.22	2.10
0.05	2.16	2.16	2.15	2.23	2.24	2.12
0.10	2.18	2.18	2.17	2.25	2.26	2.15
0.15	2.20	2.20	2.18	2.29	2.29	2.17
0.20	2.25	2.22	2.20	2.32	2.32	2.20
0.25	2.27	2.25	2.22	2.35	2.36	2.22
0.30	2.31	2.28	2.25	2.38	2.39	2.24
0.35	2.34	2.31	2.27	2.42	2.42	2.26
0.40	2.38	2.34	2.30	2.46	2.46	2.30
0.45	2.42	2.37	2.32	2.51	2.49	2.33
0.50	2.47	2.41	2.36	2.55	2.54	2.35
0.55	2.53	2.44	2.39	2.60	2.58	2.38
0.60	2.59	2.48	2.43	2.65	2.64	2.41
0.65	2.66	2.52	2.46	2.71	2.70	2.45
0.70	2.75	2.58	2.50	2.77	2.78	2.50
0.75	2.84	2.64	2.55	2.86	2.85	2.54
0.80	2.97	2.70	2.60	2.95	2.95	2.59
0.85	3.22	2.79	2.67	3.07	3.00	2.64
0.90	3.62	2.91	2.75	3.24	3.15	2.70
0.95	4.50	3.02	2.83	3.40	3.25	2.77
1.00	6.56	3.17	2.95	3.59	3.45	2.84
1.05	8.98	3.40	3.07	3.85	3.60	2.96
1.10	9.60	3.74	3.27	4.30	3.80	3.08
1.15	9.96	4.16	3.59	5.15	3.90	3.21
1.20	10.18	4.50	4.27	5.95	4.15	3.35
1.25	10.35	4.90	5.02	7.19	4.50	3.51
1.30	10.51	5.30	5.53	8.46	4.85	3.79
1.35	10.63	5.68	5.85	8.93	5.30	4.13
1.40	10.72	6.08	6.09	9.26	5.75	4.63
1.45	10.79	6.40	6.37	9.60	6.20	5.06
1.50	10.86	6.97	6.73	9.71	6.85	5.38
1.55	10.90	8.59	7.37	9.83	7.50	5.67
1.60	10.99	9.05	8.46	9.93	8.65	5.88
1.65	11.00	9.32	8.91	10.04	9.10	6.05
1.70	11.05	9.50	9.18	10.13	9.30	6.23
1.75	11.09	9.68	9. 36	10.22	9.50	6.39
1.80	11.12	9.79	9.50	10.30	9.65	6.55
1.85	11.15	9.93	9.62	10.39	9.74	6.72
1.90	11.17	10.02	9.72	10.46	9.85	6.98
1.95	11.19	10.10	9.82	10.51	9.95	7.28
2.00	11.21	10.17	9.89	10.59	10.07	7.72
2.05	11.23	10.24	9.96	10.64	10.15	8.35
2.10	11.25	10.30	10.02	10.69	10.24	8.73
2.15	11.27	10.37	10.09	10.74	10.30	8.93
2.20	11.29	10.43	10.14	10.78	10.37	9.05
2.25	11.31	10.49	10.18	10.81	10.45	9.20
2.30	11.33	10.54	10.23	10.86	10.52	9.29
2.35	11.35	10.59	10.28	10.90	10.59	9.36
2.40	11.37	10.64	10.32	10.93	10.66	9.43
2.45	11.39	10.68	10.35	10.96	10.71	9.49
2.50	11.41	10.72	10.40	10.99	10.77	9.55

(a) HNO_3 (0.0041 mol dm^{-3}) + $NaNO_3$ (0.10 mol dm^{-3})

(b) solution a + (0.001 mol dm^{-3}) dipicolinic acid

(c) solution b + (0.0004 mol dm^{-3}) copper ion

(d) solution a + (0.001 mol dm^{-3}) amino acid

(e) solution d + (0.0004 mol dm^{-3}) copper ion

(f) Solution a + (0.001 mol dm^{-3}) copper ion + (0.001 mol dm^{-3}) dipicolinic acid + (0.001 mol dm^{-3}) amino acid

Each of the above solutions was thermostated at 25, 35, 45, and 55 °C with an accuracy of ±0.1 °C, where the solutions were left to stand for about 15 min before titration. The equations of Irving and Rossotti (1953, 1954) were used to determine the protonation constants of the ligands and formation constants of the metal complexes. Multiple titrations have been performed for each system.

Table 2. Potentiometric Titration Data for the Cu^{2+}−DPA−Aspartic Acid System at 35 °C and I = 0.1 mol dm^{-3} $NaNO_3$

V/cm^3 NaOH	pH					
	a	b	c	d	e	f
0.00	2.22	2.05	1.98	2.22	2.05	2.10
0.05	2.25	2.08	2.00	2.27	2.08	2.15
0.10	2.29	2.11	2.04	2.34	2.13	2.20
0.15	2.34	2.16	2.07	2.40	2.19	2.23
0.20	2.40	2.20	2.11	2.47	2.26	2.25
0.25	2.46	2.24	2.15	2.55	2.36	2.27
0.30	2.53	2.32	2.18	2.66	2.48	2.30
0.35	2.60	2.39	2.23	2.76	2.60	2.35
0.40	2.72	2.47	2.29	2.91	2.74	2.39
0.45	2.84	2.55	2.36	3.10	2.93	2.43
0.50	3.09	2.65	2.45	3.32	3.24	2.47
0.55	3.66	2.82	2.55	3.65	3.50	2.55
0.60	5.90	3.09	2.67	4.22	3.60	2.60
0.65	8.69	3.69	2.83	5.58	4.00	2.70
0.70	9.40	4.32	3.11	6.61	4.75	2.75
0.75	9.81	5.08	4.03	8.33	5.60	2.85
0.80	10.09	5.92	5.33	9.01	6.05	3.00
0.85	10.28	8.38	6.02	9.36	7.30	3.20
0.90	10.43	9.29	8.38	9.61	8.20	3.50
0.95	10.55	9.62	9.17	9.81	9.05	4.00
1.00	10.64	9.91	9.56	10.05	9.45	4.40
1.05	10.71	10.10	9.83	10.12	9.60	4.85
1.10	10.77	10.23	9.99	10.26	9.73	5.25
1.15	10.83	10.35	10.13	10.36	9.85	5.70
1.20	10.88	10.45	10.26	10.45	9.94	6.30
1.25	10.92	10.52	10.38	10.52	10.04	6.90
1.30	10.96	10.58	10.46	10.58	10.10	8.70
1.35	10.99	10.63	10.52	10.65	10.17	9.25
1.40	11.02	10.67	10.57	10.70	10.35	9.65
1.45	11.05	10.70	10.62	10.75	10.27	9.75
1.50	11.08	10.74	10.67	10.79	10.45	10.00
1.55	11.10	10.77	10.72	10.83	10.50	10.15
1.60	11.12	10.81	10.76	10.86	10.55	10.25
1.65	11.14	10.84	10.80	10.89	10.60	10.38
1.70	11.16	10.87	10.84	10.92	10.65	10.45
1.75	11.18	10.89	10.87	10.94	10.70	10.52
1.80	11.19	10.92	10.91	10.97	10.75	10.60
1.85	11.21	10.95	10.94	10.99	10.80	10.62
1.90	11.22	10.99	10.98	11.02	10.82	10.70
1.95	11.23	11.02	11.00	11.05	10.84	10.72
2.00	11.24	11.05	11.01	11.07	10.86	10.75

Results and Discussion

Potentiometric titration data obtained according to the sequence mentioned in the Experimental Section, for Cu(II)−DPA−aspartic acid systems, at 25, 35, and 45 °C are shown in Tables 1−3. Figure 1 displays a representative set of experimental titration curves, for the Cu(II)−DPA−aspartic acid system, at 55 °C. The second proton association constants of DPA and amino acids have been determined under identical conditions from the titration curves a, b for DPA and a, d for monocarboxylic amino acid. The third association constants of dicarboxylic amino acids were also determined under the same conditions. The constructed titration curves clearly reveal that the different binary Cu(II)−DPA complexes are formed at lower pH (∼2.0). This is attained from the appeared divergence of the 1:2 binary titration curve c from that of the corresponding free DPA solution, curve b. The complex solutions of such binary systems do not show any precipitation due to hydrolysis up to higher pH's, where nearly complete complex formation occurs. This behavior strongly suggests that the ligand DPA is characterized by a high tendency to form stable metal complexes in solution. With respect to the titration curves of the different amino acid complexes

Table 3. Potentiometric Titration Data for the Cu^{2+}−DPA−Aspartic Acid System at 45 °C and I = 0.1 mol dm^{-3} $NaNO_3$

V/cm^3 NaOH	pH					
	a	b	c	d	e	f
0.00	2.19	1.95	1.88	2.15	2.12	1.97
0.05	2.23	1.99	1.91	2.16	2.16	2.00
0.10	2.26	2.01	1.94	2.22	2.21	2.03
0.15	2.31	2.06	1.99	2.27	2.26	2.06
0.20	2.36	2.11	2.02	2.33	2.32	2.10
0.25	2.43	2.16	2.07	2.39	2.39	2.13
0.30	2.52	2.22	2.13	2.46	2.46	2.17
0.35	2.62	2.28	2.19	2.52	2.55	2.22
0.40	2.74	2.36	2.26	2.65	2.67	2.28
0.45	2.90	2.45	2.33	2.82	2.82	2.33
0.50	3.18	2.59	2.43	3.00	3.04	2.41
0.55	4.13	2.77	2.58	3.23	3.28	2.49
0.60	6.10	3.06	2.74	4.00	3.62	2.59
0.65	7.59	3.57	3.14	5.50	4.01	2.72
0.70	9.27	4.24	3.95	6.75	5.00	2.87
0.75	9.65	4.82	5.33	7.61	5.60	3.10
0.80	9.88	5.61	6.08	8.74	6.41	3.37
0.85	10.06	6.38	7.97	9.40	7.15	3.84
0.90	10.21	8.67	9.07	9.60	8.85	4.44
0.95	10.33	9.19	9.39	9.80	9.35	5.14
1.00	10.41	9.52	9.62	10.00	9.65	5.68
1.05	10.50	9.77	9.80	10.10	9.85	6.17
1.10	10.56	9.96	9.98	10.20	10.05	6.82
1.15	10.61	10.09	10.11	10.28	10.19	8.21
1.20	10.67	10.20	10.19	10.36	10.29	9.01
1.25	10.72	10.28	10.27	10.46	10.35	9.37
1.30	10.75	10.36	10.35	10.52	10.40	9.63
1.35	10.78	10.45	10.42	10.57	10.50	9.83
1.40	10.81	10.50	10.49	10.63	10.55	9.98
1.45	10.84	10.55	10.55	10.70	10.64	10.12
1.50	10.88	10.59	10.60	10.74	10.69	10.25
1.55	10.92	10.63	10.65	10.79	10.70	10.33
1.60	10.96	10.67	10.69	10.83	10.75	10.41
1.65	10.99	10.70	10.74	10.87	10.80	10.48
1.70	11.03	10.74	10.79	10.91	10.85	10.54
1.75	11.06	10.77	10.83	10.95	10.90	10.60
1.80	11.09	10.81	10.88	10.99	10.95	10.65
1.85	11.12	10.85	10.93	11.02	11.00	10.69
1.90	11.15	10.88	10.97	11.05	11.05	10.73
1.95	11.17	10.91	11.01	11.08	11.10	10.76
2.00	11.19	10.94	11.05	11.11	11.15	10.79

investigated, it is evident that these complexes begin to form at a pH of 3.0−4.5.

The existence of a ternary complex is proved by comparison of the mixed ligand titration curve (f) with the composite curve, obtained by graphical addition of the amino acid titration data to that of the 1:1 Cu(II)−DPA titration curve. Therefore, it is assumed that, in the presence of both ligands, dipicolinic acid is ligated to the metal ion and then followed by ligation of amino acid; i.e., the ternary complex formation could be considered in stepwise equilibria (eqs 1 and 2).

$$M + A \rightleftharpoons MA \quad (1)$$

$$MA + L \rightleftharpoons MAL \quad (2)$$

$$K^{MA}_{MAL} = \frac{[MAL]}{[MA][L]} \quad (3)$$

From the titration curves of the solutions c and f, $\bar{n}'_{mix}$, the average number of moles of the secondary ligand amino acid coordinated to the 1:1 binary complex, [Cu(dipicolinic acid], is calculated as in the original paper (Irving and Rossotti, 1953, 1954) using the relationship:

$$\bar{n}'_{mix} = \frac{(V_f - V_c)[E^\circ + N^\circ + T^\circ_L(Y - \bar{n}_H)]}{(V_o + V_c)[T^\circ_M(DPA)]\bar{n}_H}$$

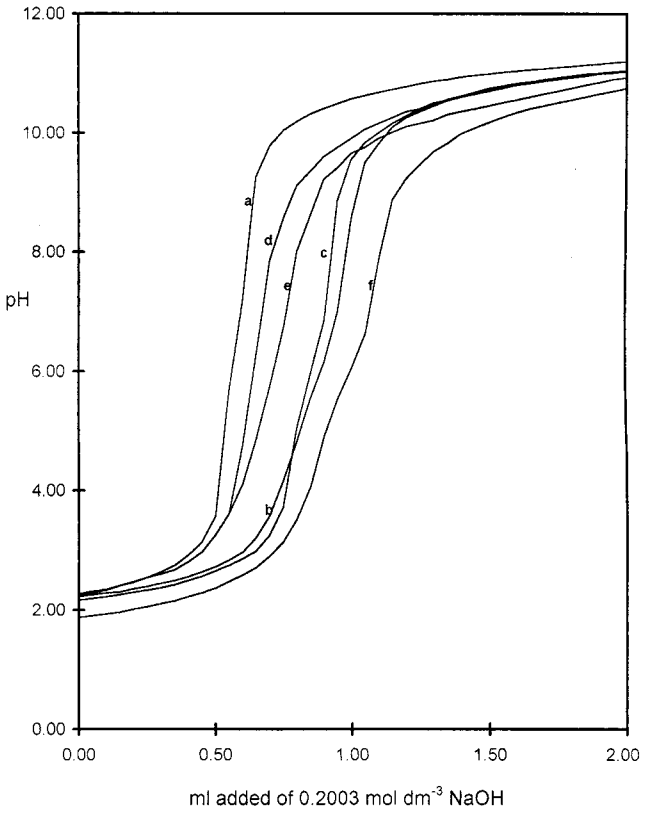

Figure 1. Potentiometric titration curves for the Cu(II)−DPA−aspartic acid system at 55.0 °C and I = 0.10 mol dm^{-3} $NaNO_3$.

Table 4. Acidity Constants of Glycine and Stability Constants of 1:1 and 1:2 Binary Complexes and 1:1:1 Ternary Complexes with Dipicolinic Acid at Different Temperatures, I = 0.1 mol dm^{-3} $NaNO_3$

system	for given t/°C			
	25	35	45	55
pK_{2a}[a]	9.76 ± 0.08	9.62 ± 0.05	9.45 ± 0.05	9.05 ± 0.07
$\log K^{M}_{ML}$	7.85 ± 0.06	7.70 ± 0.06	7.55 ± 0.04	7.03 ± 0.02
$\log K^{M}_{ML_2}$	6.75 ± 0.04	4.89 ± 0.07	5.48 ± 0.06	3.92 ± 0.08
$\Delta \log K'$	−1.10	−2.81	−2.07	−3.11
$\log K^{MA}_{MAL}$	5.80 ± 0.07	5.68 ± 0.04	5.51 ± 0.08	4.96 ± 0.06
$\Delta \log K''$	−2.05	−2.02	−2.04	−2.07

[a] For dipicolinic acid: pK_{2a} = 4.53 ± 0.06, 4.26 ± 0.04, 4.02 ± 0.07, and 3.76 ± 0.02 at 25, 35, 45, and 55 °C, respectively.

Table 5. Acidity Constants of Norvaline and Stability Constants of 1:1 and 1: 2 Binary Complexes and 1:1:1 Ternary Complexes with Dipicolinic Acid at Different Temperatures, I = 0.1 mol dm^{-3} $NaNO_3$

system	for given t/°C			
	25	35	45	55
pK_{2a}	9.70 ± 0.03	9.50 ± 0.05	9.22 ± 0.08	9.00 ± 0.05
$\log K^{M}_{ML}$	7.50 ± 0.06	7.35 ± 0.05	7.07 ± 0.05	6.95 ± 0.05
$\log K^{ML}_{ML_2}$	5.02 ± 0.07	4.92 ± 0.08	4.58 ± 0.09	4.05 ± 0.04
$\Delta \log K'$	−2.48	−2.43	−2.49	−2.90
$\log K^{MA}_{MAL}$	5.02 ± 0.06	4.83 ± 0.04	4.56 ± 0.05	4.35 ± 0.07
$\Delta \log K''$	−2.48	−2.52	−2.51	−2.60

and since $E^\circ + N^\circ \gg T^\circ_L$

$$\bar{n}'_{mix} = \frac{(V_f - V_c)[E^\circ + N^\circ]}{(V_o + V_c)[T^\circ_{M(DPA)}]\bar{n}_H} \quad (4)$$

where V_f and V_c are the volumes of NaOH consumed to reach the same pH value in the curves f and c, respectively.

Table 6. Acidity Constants of Aspartic Acid, Stability Constants of 1:1 and 1:2 Binary Complexes and 1:1:1 Ternary Complexes with Dipicolinic Acid at Different Temperatures, I = 0.1 Mol dm^{-3} NaNO$_3$

	for given t/°C			
system	25	35	45	55
pK_{3a}	9.80 ± 0.02	9.75 ± 0.06	9.33 ± 0.04	8.93 ± 0.04
log K^{M}_{ML}	8.85 ± 0.04	8.60 ± 0.04	8.50 ± 0.04	8.03 ± 0.05
log $K^{ML}_{ML_2}$	7.88 ± 0.08	6.46 ± 0.05	6.40 ± 0.07	4.66 ± 0.08
Δ log K'	−0.97	−2.14	−2.10	−3.37
log K^{MA}_{MAL}	9.60 ± 0.02	9.46 ± 0.04	9.38 ± 0.07	8.74 ± 0.07
Δ log K''	0.75	0.86	0.88	0.71

Table 7. Acidity Constants of Glutamic Acid and Stability Constants of 1:1 and 1:2 Binary Complexes and 1:1:1 Ternary Complexes with Dipicolinic Acid at Different Temperatures, I = 0.1 mol dm^{-3} NaNO$_3$

	for given t/°C			
system	25	35	45	55
pK_{3a}	9.42 ± 0.04	9.25 ± 0.07	9.00 ± 0.02	8.75 ± 0.05
log K^{M}_{ML}	8.37 ± 0.05	8.35 ± 0.05	8.10 ± 0.03	7.85 ± 0.02
log $K^{ML}_{ML_2}$	7.42 ± 0.04	7.10 ± 0.04	6.28 ± 0.05	6.50 ± 0.05
Δ log K'	−0.95	−1.25	−1.82	−1.35
log K^{MA}_{MAL}	8.69 ± 0.06	8.56 ± 0.05	8.33 ± 0.08	8.06 ± 0.04
Δ log K''	0.32	0.21	0.23	0.21

[[$T°_M$(DPA)] is the concentration of the binary copper−dipicolinic acid complex which is equivalent to the initial copper ion concentration $T°_M$, $\bar{n}_H$ is the average number of protons associated with the secondary ligand amino acid, $T°_L$ is the initial concentration of the secondary ligand amino acid, and Y is equal to 1. V_o = original volume, $E°$ = HNO_3 concentration in the titrated solution, and $N°$ = NaOH concentration, in mol dm^{-3}. From the values of n'_{mix} so obtained, free secondary ligand exponent, pL'_{mix}, was calculated using the equation:

$$pL'_{mix} = \log\left(\sum_{y=0}^{y=1} \frac{B_y^H \left(\frac{1}{10^B}\right)^y}{T°_L - \bar{n}_{mix} T°_M} \frac{V_o + V_f}{V_o}\right) \qquad (5)$$

B_y^H = second or third formation constant values for mono- or dicarboxylic amino acid, respectively. B = the pH-meter reading.

The mean pK_a, log K^{M}_{ML}, log $K^{ML}_{ML_2}$, and log K^{MA}_{MAL} values are determined from the corresponding experimental formation curves using the average value and straight line methods. The values obtained along with the estimated error using least-squares refinement are given in Tables 4−7.

The stability constants of copper(II)−dipicolinic acid binary complexes could not be determined using the pH-metric technique, due to the combination of the acid strength of the complexing agent with the high stabilities of the complexes, which made the determination very inaccurate.

It was found that the stability constants of the 1:2 binary systems of amino acids investigated are lower than those of the corresponding 1:1 ones, as expected from statistical considerations. The Δ log K' (log $K^{ML}_{ML_2}$ − log K^{M}_{ML}) values are negative for all systems studied.

The stabilities of 1:1 amino acid binary or 1:1:1 ternary complexes at each temperature decrease in the following order: aspartic acid > glutamic acid > glycine > norvaline (Tables 4−7). The relative stabilities of the binary and ternary complexes of glycine and norvaline follow their relative basicities, since it is well-known that the increase in basicity of a ligand increases the stability of its metal complexes.

Concerning dicarboxylic amino acids, the observed little difference in stability of the binary or ternary complexes of α,β-dicarboxylic aspartate compared to the corresponding ones of α,γ-dicarboxylic glutamate can be attributed to steric effects.

The relative stability of the ternary complexes, as compared to that of the corresponding binary amino acid systems, can be quantitatively expressed (Martin and Prados, 1974) in terms of Δ log K''. Tables 4−7 demonstrate the difference in stabilities of the binary and ternary complexes in terms of Δ log K'' as defined by eq 6.

$$\Delta \log K'' = \log K^{MA}_{MAL} - \log K^{M}_{ML} \qquad (6)$$

It can be observed that the values of Δ log K'' are positive for aspartic acid and glutamic acid systems, which means that ternary complexes of these acids are more stable than the corresponding binary ones. In contrast, negative values of Δ log K'' for glycine and norvaline ternary systems are obtained; i.e., their binary complexes are more stable than their corresponding ternary systems.

It is shown from Tables 4−7 that the values of the stability constants of 1:1 and 1:2 binary copper(II)−amino acid complexes and 1:1:1 ternary ones decrease with an increase in temperature. This can be ascribed to the exothermic nature of the investigated complex formation process.

The values of the thermodynamic quantities associated with (a) dissociation of the ligands studied, (b) formation of binary 1:1 copper(II)−amino acid complexes, and (c) formation of 1:1:1 ternary complexes were also determined (Table 8).

Table 8. Thermodynamic Quantities Associated with the Dissociation of the Ligands Studied, the Interaction of Cu^{2+} with Amino Acids at 1:1 Molar Ratio, and the Interaction of Cu^{2+} with Dipicolinic Acid and Amino Acids at a 1:1:1 Molar Ratio, I = 0.1 mol dm^{-3} NaNO$_3$

system	$\Delta H°$/(kJ·mol^{-1})	$\Delta G°$(25°C)/(kJ·mol^{-1})	$\Delta S°$/(J·K^{-1}·mol^{-1})
dipicolinic acid	46.05 ± 0.90	26.04 ± 0.06	75.32 ± 0.90
glycine	28.47 ± 0.80	56.06 ± 0.08	−92.57 ± 0.80
norvaline	34.77 ± 0.70	55.72 ± 0.03	−53.50 ± 0.70
aspartic acid	44.04 ± 0.50	56.31 ± 0.02	−41.15 ± 0.50
glutamic acid	40.19 ± 0.80	54.09 ± 0.04	−46.64 ± 0.80
Cu^{2+}−glycine	−27.42 ± 0.60	−45.09 ± 0.06	59.28 ± 0.60
Cu^{2+}−norvaline	−39.69 ± 0.90	−43.08 ± 0.06	11.38 ± 0.90
Cu^{2+}−aspartic acid	−52.33 ± 0.90	−50.82 ± 0.04	−5.06 ± 0.90
Cu^{2+}−glutamic acid	−48.73 ± 0.80	−48.06 ± 0.05	−2.26 ± 0.80
Cu^{2+}−dipicolinic acid−glycine	−29.47 ± 0.70	−33.32 ± 0.07	12.89 ± 0.70
Cu^{2+}−dipicolinic acid−norvaline	−42.24 ± 0.50	−28.84 ± 0.06	−44.92 ± 0.50
Cu^{2+}−dipicolinic acid−aspartic acid	−71.80 ± 0.60	−55.14 ± 0.02	−55.89 ± 0.60
Cu^{2+}−dipicolinic acid−glutamic acid	−63.93 ± 0.50	−49.90 ± 0.06	−47.05 ± 0.50

The enthalpy change for the dissociation of the ligands are positive (endothermic). The strong interaction between Cu(II) and amino acids is evidenced by the large magnitude of stability constants and the exothermic nature of $\Delta H°$. The negative values of $\Delta G°$, and hence the spontaniety of binary complex reactions of Cu(II) with monocarboxylic amino acids (glycine and norvaline) are attributed to the high positive $\Delta S°$ term; hence the reaction is entropy favored. The negative values of $\Delta G°$ and hence the spontaniety of both binary and ternary complexes of Cu(II) with the amino acids studied are ascribed to the high negative values of $\Delta H°$ compared with $\Delta S°$ values; i.e., complexation reactions are generally enthalpy favored.

Literature Cited

Bagger, S. Metal−ligand amino acids and peptides as substrates for proteases. *Int. J. Pept. Protein Res.* **1987**, *30*, 99−101.

Casassas, E.; Tauler, R.; Fonrodona, G. Mixed ligand complex formation of copper(II) and zinc(II) ions with 3-hydroxynaphthoate and picolinate ions as ligands in dioxane-water mixtures. *Polyhedron* **1998**, *7*, 1335−1340.

Choppin, G. R.; Rao, L. F.; Rizkalla, E. N.; Sullivan, J. C. Calorimetric studies of neptunium(V) complexation by polycarboxylate ligands. *Radiochim. Acta* **1992**, *57*, 173−175.

Crans, D. C.; Bunch, R. L.; Theisen, L. A. Interaction of trace levels of vanadium (IV) and vanadium (V) in biological systems. *J. Am. Chem. Soc.* **1989**, *111*, 7597−7607.

Ho, B.; Yut, K. Stabilizers for vinyl halide polymers. Eur. Pat. Appl. EP 62, 362, 1982; *Chem. Abstr.* **1983**, *98*, 55125y.

Irving, H. M.; Rossotti, H. S. Methods for computing successive stability constants from experimental formation curves. *J. Chem. Soc.* **1953**, 3397−3405.

Irving, H. M.; Rossotti, H. S. The calculation of formation curves of metal complexes from pH-titration curves in mixed solvents. *J. Chem. Soc.* **1954**, 2904−2910.

Khalil, M. M.; Radalla A. M. Binary and ternary complexes of inosine. *Talanta* **1998**, *46*, 53−61.

Khalil, M. M.; Tanase I.; Luca C. A polarographic study of some complexes of TI(I) with polyoxa macrocyclic ligands. *Talanta* **1985**, *32*, 1151−1152.

Khalil, M. M.; Elghandour, A. H. H.; Mostafa, M.; Shoukry, M. M. Metal chelates of some 1-substituted-3-thiazole-2-ylthiourea. *Polyhedron* **1994,** *13*, 3295−3297.

Khalil, M. M.; Mohamed S. A.; Radalla A. M. Potentiometric and conductometric studies on the binary and mixed ligand complexes in solutions: M^{II}-dipicolinic acid-glycine systems. *Talanta* **1997**, *44*, 1365−1369.

Kidani, Y.; Hirose, J.; Koike, H. Coordination chemical studies on metalloenzymes. I. Kinetics and mechanism of the zinc(II) exchange reaction between chelating agent and apo-bovine carbonic anhydrase. *J. Biochem. (Tokyo)* **1976**, *79*, 43−51.

Martin, R. B.; Prados, R. J. Some factors influencing mixed complex formation. *J. Inorg. Nucl. Chem.* **1974**, *36*, 1665.

Napoli, A. The system aluminium(III)-dipicolinic acid in aqueous 0.5 M sodium perchlorate medium. *Talanta* **1968**, *15*, 189−198.

Napoli, A.; Magri A. L. Oxovanadium(IV) complexes with α-pyridinedicarboxylic acids. *Ann. Chim. (Rome)* **1987**, *77*, 783−788.

Ringbom, A. Complexation in analytical chemistry; Wiley-Interscience: New York, **1963**.

Seleim, M. M.; Idriss, K. A.; Saleh, M. S.; Sedaira, H. Reaction of mercury(II)-thiosalicylate complex with picolinic acid and ultraviolet spectrophotometric determination of mercury(II). *Analyst* **1987**, *112*, 1685−1688.

Shelke, D. N.; Jahagirdar, D. V. Equilibrium study of the mixed complexes of uranyl ion with pyridine dicarboxylic and aliphatic dicarboxylic. acids. *J. Inorg. Nucl. Chem.* **1979a**, *41*, 929−931.

Shelke, D. N.; Jahagirdar D. V. Ternary chelates of Zn(II) with some biological active ligands. *J. Inorg. Nucl. Chem.* **1979b**, *41*, 1635−1638.

Ullah, M. R.; Bhattacharya, P. K. Study of various factors affecting the stability of ternary Cu(II) and Ni(II) complexes. *Bull. Chem. Soc. Jpn.* **1991**, *64*, 3659−3661.

Welcher, F. J. *The analytical uses of ethylenediaminetetraacetic acid*; Von Nostrand: Princeton, **1965**.

Won, T. J.; Sudan, B. M.; Thompson, R. C. Characterization of oxoperoxo(2,6-pyridinedicarboxylato)molybdenum(VI) and oxoperoxo(nitrilotriacetato)molybdate(VI) in aqueous solution and a kinetic study of their reduction by a (thiolato)cobalt(III) complex, dimethyl sulfoxide, and iron(II). *Inorg. Chem.* **1994**, *33*, 3804−3810.

Wright, A. P. G.; Wen, B. P. L.; Schenk, T.G.; Maron, A. Preparation and testing of aluminium hydroxide carboxylic acid complexes anticides. U.S. Patent 4, 772, 724, 1988; *Chem. Abstr.* **1989**, *110*, 154563n.

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