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Prediction of Stability Constants.
II. Metal Chelates of Natural Alkyl Amino Acids and their Synthetic Analogs

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Abstract

The currently available critical stability constants of amino acid complexes of metal ions are employed as the basis for estimating the constants that have not been measured so as to provide a comprehensive data base describing the reactions of metal ions with twenty two types of aliphatic amino acid ligands. The latter include diaminomonocarboxylic and monoaminodicarboxylic acids, but are restricted to aliphatic naturally occurring ligands, and closely related synthetic compounds. In addition stability data are presented for fourteen structural types of monoprotonated amino acids, and six structural types of diprotonated amino acids. For more limited numbers of amino acid types data are presented and trends are discussed for the alkaline earths, the lanthanides, and the first row transition metal series. The need for more measured (experimental) data for the development of a comprehensive data base of measured and estimated constants for such groups of metal ions as the lanthanides, transition metals, the actinides, and a few heavy metal ions is pointed out.

Introduction

Procedures have been described [1, 2] for extending the limited metal–organic ligand equilibrium data currently available for the large number of ligands and metal ions present in environmental systems and biological fluids. The expansion of the data base on stability constants is needed for the determination of the speciation (*i.e.*, concentrations of the complexes actually formed) of metal ions and ligands in these complex multicomponent systems at equilibrium. Once a data base containing equilibrium constants for all the metal complexes that may be formed becomes available, calculation of speciation of the complete system may be carried out by computer with the program SPE [3] for homogeneous media, or program MINEQL [4] when solid phases are formed. This information is necessary for the determination of the physical states (solution *vs.* precipitate) of the complexes, their possible involvement in metal and ligand transport, and estimation of their toxicities. The biological action, including toxic effects, of both ligands and metal ions have been shown to be profoundly influenced by the stabilities and chemical nature of the complexes involved [1, 5, 6].

The previous paper by the authors on this subject consists of the expansion of existing data to cover the equilibrium constants for complexes of the basic, type A metal ions with a large number of aliphatic carboxylates [2]. This work was based on the relatively small number of critical stability constants of metal carboxylates available [7, 8]. In this paper these methods are employed for the development of a data base for another important class of compounds found to be wide spread in the environment as the result of breakdown of proteinaceous materials of plant or animal origin, the α-amino acids. Because amino acids contain a moderately hard donor (the carboxylate group) and a donor that is intermediate between hard and soft (the amino-group) it is essential that estimates be made for both hard and soft metal ions, including the first row transition metal ions. Ligands are restricted to natural amino acids and their closely related analogs. Studies of EDTA and related chelating agents involve a large number of synthetic ligands, and will be described in a subsequent report.

Experimental

The procedures employed have been described previously [1, 2] and are given here only in outline form. On the basis of known variations of reliable measured equilibrium constants [7, 8] as a function of the nature of the metal ion and ligand, estimates are made for metal ions with the ligands for which data exist on stabilities of the complexes of the same ligand with closely-related metal ions. Conversely, affinities of new ligands for a metal ion are estimated

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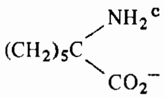
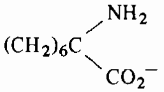
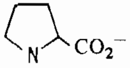
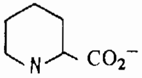
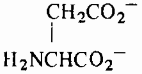
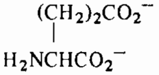
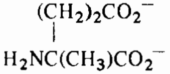
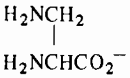
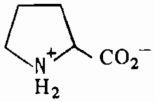
TABLE I. Logarithms of Stability Constants for Metal Ion Complexes of Amino Acids at 25.0 ℃ and 0.10 M Ionic Strength. Values in [ ] are estimated.

	H^+	Mg^{2+}[a]	Ni^{2+}[b]		Cu^{2+}		Ag^+		CH_3Hg^+		Fe^{3+}
	Log K_1	Log K_1	Log K_1	Log β_2	Log K_1	Log β_2	Log K_1	Log β_2	Log K_1	Log β_2	Log K_1
Unprotonated Amino Acids											
$H_2NCH_2CO_2^-$	9.6 ± 0.1	1.5 ± 0.0	5.8–0.1	10.6–0.1	8.1 + 0.1	15.0 ± 0.0	3.2 + 0.1	6.6 + 0.1	7.8 ± 0.1	9.5	[10.1]
$H_2NCHRCO_2^-$	9.6 ± 0.1	1.6	5.4 ± 0.1	9.8 ± 0.2	8.1 ± 0.1	14.9 ± 0.2	3.2 ± 0.2	6.6 ± 0.4	7.6 ± 0.1	9.3 ± 0.2	10.1 ± 0.2
$H_2NC(CH_3)_2CO_2^-$	10.1 ± 0.0	[1.6]	5.1	9.5 ± 0.2	8.2	15.0	[3.4]	[7.1]	[8.1]	[9.9]	10.6
 $(CH_2)_5C(NH_2)^{c}CO_2^-$	10.1 ± 0.1	[1.6]	5.5 + 0.1	10.0 + 0.1	8.6 ± 0.0	15.8 ± 0.0	[3.4]	[7.1]	[8.1]	[9.9]	[10.6]
 $(CH_2)_6C(NH_2)CO_2^-$	10.3	[1.6]	5.3	9.7	8.6	15.9	[3.5]	[7.3]	[8.2]	[10.0]	[10.8]
$H_2NCHRCH_2CO_2^-$	10.1–0.1	[1.1]	4.6 ± 0.1	7.9 ± 0.1	7.1 ± 0.1	12.7 ± 0.2	3.4	7.2	7.3	8.9	
$H_2N(CH_2)_3CO_2^-$	10.3 + 0.1	[1.1]	[3.4]	[6.2]	5.5	[10.2]	3.5	7.2	7.6	[9.3]	
$H_2N(CH_2)_nCO_2^-$	10.5 + 0.1	[1.1]	[3.6]	[6.5]	[5.7]	[10.6]	3.6 ± 0.0	7.5 ± 0.1	7.8 ± 0.0	[9.5]	
$CH_3NHCH_2CO_2^-$	10.0–0.1	[1.6]	5.5–0.1	9.8 ± 0.1	7.7 ± 0.0	14.2	[3.3]	[6.9]	[6.7]	[8.2]	10.0
$RNHCH_2CO_2^-$	10.1–0.1	[1.6]	4.8 ± 0.0	8.5–0.1	7.3 ± 0.0	13.4 ± 0.1	[3.4]	[7.1]	[6.8]	[8.3]	[10.1]
$(CH_3)_2CHNHCH_2CO_2^-$	10.1	[1.6]	3.9	[7.1]	6.7	12.4	[3.4]	[7.1]	[6.8]	[8.3]	[10.1]
	10.5–0.1	[1.7]	5.9 ± 0.0	10.9 ± 0.0	8.8 + 0.1	16.4 ± 0.1	[3.6]	[7.5]	[6.9]	[8.4]	10.3
	10.5	[1.6]	[5.0]	[9.1]	7.4	14.0	[3.6]	[7.5]	[6.9]	[8.4]	[10.3]
$(CH_3)_2NCH_2CO_2^-$	9.8 ± 0.0	[1.5]	4.8	8.6	7.3–0.1	13.7 + 0.1	[3.3]	[6.9]	[5.4]	[6.6]	[10.0]
$(CH_3CH_2)_2NCH_2CO_2^-$	10.5	[1.4]	4.2	[7.6]	6.9	12.9	[3.6]	[7.5]	[5.7]	[7.0]	[10.3]
 $H_2NCH(CH_2CO_2^-)CO_2^-$	9.6 ± 0.0	2.4	7.2–0.1	12.4 ± 0.1	8.9 ± 0.1	15.9 ± 0.1	[3.8]	[7.9]	[7.7]	[9.4]	12.0
 $H_2NCH((CH_2)_2CO_2^-)CO_2^-$	9.6 ± 0.1	1.9	5.6	9.8	8.3 + 0.1	14.8 ± 0.1	3.8	[7.9]	[7.7]	[9.4]	12.7
 $H_2NC(CH_3)((CH_2)_2CO_2^-)CO_2^-$	9.6	[1.9]	[5.0]	[9.1]	8.0	[14.3]	[3.8]	[7.9]	[7.7]	[9.4]	[12.7]
 $H_2NCH(CH_2NH_2)CO_2^-$	9.4 ± 0.0	[2.4]	8.2–0.1	15.2 + 0.1	10.6–0.1	19.8 ± 0.0	[4.3]	[9.0]	[7.9]	[9.6]	

$H_2N(CH_2)_2$ │ $H_2NCHCO_2^-$	10.2 ± 0.0	[1.9]	8.9 + 0.1	16.0	10.6–0.1	19.0	[5.9]	[12.3]	[7.4]	[9.0]
$H_2N(CH_2)_3$ │ $H_2NCHCO_2^-$	10.5 ± 0.0	1.9	7.1	12.1	10.0	15.7–0.1	[5.9]	[13.3]	[7.6]	[9.3]
$H_2N(CH_2)_4$ │ $H_2NCHCO_2^-$	10.7 ± 0.0	[1.9]	5.8	10.4	[8.4]	14.9 ± 0.1	[6.0]	[12.5]	[7.8]	[9.5]

	H^+ Log K_1	Na^+ Log K_1	Mg^{2+} Log K_1	$La^{3+\,d}$ Log K_1	Am^{3+} Log K_1	UO_2^{2+} Log K_1	Ni^{2+} Log K_1	Cu^{2+} Log K_1	CH_3Hg^+ Log K_1	Fe^{3+} Log K_1
Monoprotonated Amino Acids										
$H_3N^+CH_2CO_2^-$	2.4–0.1	[–0.8]	[0.4]	0.8	1.0	1.4	[1.0]	1.6 ± 0.1	[2.0]	[1.8]
$H_3N^+CHRCO_2^-$	2.3–0.1	[–0.8]	[0.4]	0.6	[1.0]	[1.4]	[0.8]	1.2 + 0.1	[2.0]	[1.7]
$H_3N^+C(CH_2)_4CO_2^-$	2.4	[–0.8]	[0.4]	[0.7]	[1.0]	[1.4]	[1.2]	1.9	[2.0]	[1.8]
$H_3N^+C(CH_2)_6CO_2^-$	2.6	[–0.8]	[0.4]	[0.8]	[1.1]	[1.5]	[0.7]	1.1	[2.2]	[2.0]
$H_3N^+CHRCH_2CO_2^-$	3.5 ± 0.1	[–0.6]	[0.5]	[1.2]	[1.6]	[2.0]	[0.9]	[1.4]	2.3	[2.7]
$H_3N^+(CH_2)_3CO_2^-$	4.1–0.1	[–0.5]	[0.6]	[1.6]	[2.0]	[2.4]	[1.0]	[1.6]	2.9	[3.1]
$H_3N^+(CH_2)_nCO_2^-$	4.3 ± 0.1	[–0.4]	[0.6]	[1.7]	[2.1]	[2.5]	[1.1]	1.8	3.3–0.1	[3.3]
	1.9 ± 0.1	[–0.9]	[0.3]	[0.4]	[0.7]	[1.1]	[0.7]	1.1	[1.8]	[1.4]
$CH_2CO_2^-$ │ $H_3N^+CHCO_2^-$	3.7 ± 0.0	[0.0]	[1.0]	[2.1]		2.6	[1.9]	3.0 ± 0.1	[3.1]	[5.4]
$(CH_2)_2CO_2^-$ │ $H_3N^+CHCO_2^-$	4.2 ± 0.1	[0.0]	[0.8]	[1.9]		2.7	[1.8]	2.9 ± 0.1	[3.3]	[6.1]
$H_3N^+CH_2$ │ $H_2NCHCO_2^-$	6.7 ± 0.0	[–0.9]	[1.5]				4.0 ± 0.0	6.2 ± 0.1	[6.0]	[7.0]
$H_3N^+(CH_2)_2$ │ $H_2NCHCO_2^-$	8.2 ± 0.0	[–0.8]	[1.6]				4.6 ± 0.0	7.0 ± 0.1	[6.7]	[8.6]
$H_3N^+(CH_2)_3$ │ $H_2NCHCO_2^-$	8.8–0.1	[–0.8]	1.7				4.6 ± 0.1	7.4–0.1	[7.0]	[9.2]

(Continued overleaf)

TABLE I. (*Continued*)

	H^+ Log K_1	Na^+ Log K_1	Mg^{2+} Log K_1	La^{3+d} Log K_1	Am^{3+} Log K_1	UO_2^{2+} Log K_1	Ni^{2+} Log K_1	Cu^{2+} Log K_1	CH_3Hg^+ Log K_1	Fe^{3+} Log K_1
$H_3N^+(CH_2)_4$ $H_2NCHCO_2^-$	9.2–0.1	[–0.8]	[1.7]				4.9–0.1	7.6 + 0.1	[7.2]	[9.7]
Diprotonated Amino Acids										
CH_2CO_2H $H_3N^+CHCO_2^-$	1.9 + 0.1	[–0.9]	[0.3]	[0.4]	[0.7]	[1.1]	[0.7]	[1.1]	[1.4]	[1.4]
$(CH_2)_2CO_2H$ $H_3N^+CHCO_2^-$	2.2 ± 0.1	[–0.8]	[0.4]	[0.5]	[0.9]	[1.3]	[0.8]	[1.2]	[1.6]	[1.7]
$H_3N^+CH_2$ $H_3N^+CHCO_2^-$	1.3 ± 0.0	[–1.0]	[0.2]	[0.1]	[0.4]	[0.8]	[0.6]	[1.0]	[1.0]	[1.0]
$H_3N^+(CH_2)_2$ $H_3N^+CHCO_2^-$	1.7 ± 0.2	[–0.9]	[0.3]	[0.3]	[0.6]	[1.0]	[0.7]	[1.1]	[1.3]	[1.3]
$H_3N^+(CH_2)_3$ $H_3N^+CHCO_2^-$	1.9 ± 0.2	[–0.9]	[0.3]	[0.4]	[0.7]	[1.1]	[0.7]	[1.1]	[1.4]	[1.4]
$H_3N^+(CH_2)_4$ $H_3N^+CHCO_2^-$	2.2 ± 0.0	[–0.8]	[0.4]	[0.5]	[0.9]	[1.3]	[0.8]	[1.2]	[1.6]	[1.7]

[a] Internal trends for alkaline earth complexes are in Table II. [b] Internal trends for some + 2 transition and post-transition metal ions are in Table IV. [c] Includes $H_2NC(CH_2)_4CO_2^-$. [d] Internal trends for lanthanide complexes are in Table III.

from known stabilities of complexes of a given metal ion with similar ligands. Such estimates are frequently made by linear extrapolation or interpolation. The ratios of log K values that are found to hold for pairs of metal ions with a variety of ligands has also been found to be very useful. No estimates are made for systems for which baseline data on related complexes are not available.

The ionic strength relationships described in the previous paper [2] were employed for converting equilibrium data from one ionic strength to another. The standard conditions used in this research, in order to make comparisons between stability constants of various complexes, are those most commonly employed for stability constant determination, 25.0 °C, and 0.10 M ionic strength. The protonation constants and hydrolysis constants involving hydrogen ion equilibria are expressed in terms of hydrogen ion concentration, $[H^+]$.

Because of the wide range of ligand types and metal ions for which stability constants are needed, and the grouping of closely-related ligands so as to consider them to be a single ligand, it was necessary to round off the logarithms of the equilibrium constants to a single decimal number. This degree of accuracy should be sufficient for speciation determinations for most environmental and biological systems. Greater accuracy may be achieved by more detailed studies of closely related groups of complexes.

Results and Discussion

The critical constants of amino acid complexes obtained from the literature [7, 8], and the new values of estimated constants obtained by the procedures described above, are presented in Table I. All of the mono-α-aminomonocarboxylic acids, containing simple alkyl chains are grouped together. Cyclic α-amino acids, and amino acids with the amino and carboxyl groups separated by more than one carbon atom have considerably different properties and are considered separately, as are those that contain secondary and tertiary amino groups. The mono-aminodicarboxylic and the diaminomonocarboxylic acids are of course also listed individually. The second and third parts of the table contain data for mono- and diprotonated amino acids, respectively. The constants given are for the formation of complexes with the same number of protons on the metal complex as were on the ligand to start with. There are, of course, no implications concerning the nature of the protonated sites on the protonated metal complex finally formed.

Ligand Protonation Constants

The first protonation constant reflects the combination of a proton with the amino group of the most basic form of the ligand to form a dipolar ion. The magnitude of the constant is decreased relative to that of the corresponding amine by the electron-withdrawing effect of the neighboring deprotonated carboxylate. The effect decreases as we consider ligands with the carboxylate farther removed from the amino group and ultimately the protonation constants approach the values observed with alkylamines.

With amino acids containing a secondary or tertiary amino group or with α-amino acids which are disubstituted at the carbon, the protonation constants are higher than those of the less-substituted amino acids.

The second protonation constant involves adding a proton to the carboxylate. The greatly reduced magnitudes of these constants compared to those of alkylcarboxylates are due to the strong electron-withdrawing effect of the protonated amino group. Its values increase in magnitude in ligand molecules where the separation between amino and carboxylate groups is increased.

The presence of a second carboxylate in the molecule does not affect the magnitude of the protonation constant of the amino group. The second protonation constant involves the more isolated carboxylate and its basicity is less influenced by the protonated amino group in molecules where the two groups are more widely separated. The third protonation constant involves the protonation of the carboxylate close to the protonated amino group. The basicity of the second carboxylate is generally very low, primarily because of the electron withdrawal effect (–I) of the protonated amino group, and to a lesser extent because of the –I contribution of the protonated carboxylate group.

With a second amino group in the molecule the first protonation reaction occurs at the more isolated amino group. The magnitude of its protonation constant is decreased somewhat by the other neutral amino group only when the two are on adjacent carbon atoms, and is not appreciably decreased when they are farther separated from each other. The second protonation constant for amino acids containing two amino groups reflects the addition of a proton to the amino group closer to the carboxylate. The magnitude of this constant is greatly reduced because of the large electron-withdrawing effect of the other protonated amino group, but only when the separation between these groups is small.

Alkali Metal Complexes

No experimental values are available for the stability constants of alkali metal complexes of α-amino acids so that estimates must be made from those of the alkylcarboxylic acids. No complexes are expected with deprotonated amino acids because when these are formed at high pH ($pK_a = 10.0 \pm 0.5$) the hydroxide

TABLE II. Alkaline Earth Internal Trends, Log K_1 Relative to Mg^{2+}.

	Mg^{2+}	Ca^{2+}	Sr^{2+}	Ba^{2+}	Ra^{2+}
$H_2NCH_2CO_2^-$	0.0	−0.5 ± 0.2	−0.9–0.1	−1.1	[−1.2]
$H_2NCHRCO_2^-$	0.0	−0.8	[−1.1] − 0.2	−1.2	[−1.2]
$\overset{CH_2CO_2^-}{\overset{\vert}{H_2NCHCO_2^-}}$	0.0	−0.8	−0.9	−1.3	−1.5
$\overset{(CH_2)_2CO_2^-}{\overset{\vert}{H_2NCHCO_2^-}}$	0.0	−0.5–0.1	−0.5	−0.6	[−0.6]
$\overset{H_2N(CH_2)_3}{\overset{\vert}{H_2NCHCO_2^-}}$	0.0	0.0			

ion effectively competes with dilute amino acids for alkali metal ions since the stability constants of the alkali metal hydroxides are about the same as those of the alkylcarboxylic acids (log $K = -0.4$). With large concentations of amino acid anions complex formation is possible, but the complexes formed would be expected to have stability constants that are similar in magnitude but slightly higher than those of the mono-protonated amino acids.

Mono-protonated amino acids are dipolar and consequently the carboxylate group is available for coordination. Protonation because of its higher affinity for amines makes the amino group unavailable for metal ion coordination and the stabilities of alkali metal complexes would be expected to be similar to those of the alkylcarboxylic acids modified by the electron-withdrawing effect of the protonated amine. As the protonated amine becomes more distant from the carboxylate, the stability constants should approach those of the alkylcarboxylates. The constants for Na^+ in Table I were estimated from those of the alkylcarboxylates, with appropriate allowances for the inductive (−I) effect of the protonated amino group.

The complexes of the other alkali metal ions would be expected to show similar trends relative to the stability constants of the sodium complexes, as those reported for the alkylcarboxylates [2]. It would be of value to have available measured values of the stability constants of sodium ion with aspartic acid and other more structurally complex amino acids to test the predictions made in Table I.

Alkaline Earth Complexes Including Beryllium(II)

No stability constants of beryllium(II) complexes with alkyl amino acids have been measured and the best estimates for these values would be from those of the magnesium complexes and the Mg^{2+}/Be^{2+} log stability constant ratio observed with alkylcarboxylic acids of 0.38 [2]. Ligands with a second carboxylic acid would be expected to have a larger ratio.

Only a few magnesium(II) complexes of alkylamino acids have been measured. The magnitudes of the formation constants with α-amino acids are larger than those of the monoalkylcarboxylic acids and are similar to those of substituted malonic acids [2], which implies that a metal chelate is formed. Amino acids with the functional groups more separated may be estimated on the basis of stabilities of analogous complexes containing two isolated carboxylate donors [2]. No stability constants have been measured for amino acids containing secondary or tertiary amino groups and these were assumed to be similar to those having primary amino groups, but modified because of expected steric effects, especially with the tertiary amines. The enhanced values of the stability constants of aspartate compared to glutamate imply the presence of two chelate rings in the complexes of both ligands, with that of aspartate being the more stable.

Internal trends of the stability constants of the alkaline earth complexes relative to those of magnesium(II) are given in Table II. Estimated values are based on the observed trends.

Scandium, Yttrium, and Lanthanide Complexes

Most of the constants reported for scandium(III), yttrium(III) and the lanthanides(III) are apparent constants which assume that the amino proton is displaced by the metal ion and that hydrolysis is insignificant [9]. The displacement of the amino proton is contradicted by a crystal structure of neodymium glycinate [10] and an infrared study of lanthanide glycinates [11]. Species distribution calculations as a function of pH based on reported constants indicate that precipitation occurs before the normal deprotonated complex (ML) is formed.

Circular dichroism spectra [12], NMR spectra, and UV spectra [13] have been interpreted as indicating some degree of chelate formation with a large excess of ligand in basic solutions. The apparent constants reported [9] are probably a result of base uptake by

TABLE III. Lanthanide Internal Trends Relative to La^{3+}.

		La^{3+}	Ce^{3+}	Pr^{3+}	Nd^{3+}	Pm^{3+}	Sm^{3+}	Eu^{3+}	Gd^{3+}
$H_3N^+CHRCO_2^-$	Log K_1 [a]		0.00[a]		+0.11[a]	+0.14[a]		+0.17[a]	+0.20[a]
$CH_2CO_2^-$ $H_2NCHCO_2^-$	Log K_1	0.00	+0.29	+0.39	+0.54 ±0.02		+0.73	+0.80	+0.92
	Log β_2	0.0	+0.5	+0.8	+1.1 ±0.1		+1.4	+1.5	+2.2
		Tb^{3+}	Dy^{3+}	Ho^{3+}	Er^{3+}	Tm^{3+}	Yb^{3+}	Lu^{3+}	
$H_3N^+CHRCO_2^-$	Log K_1								
$CH_2CO_2^-$ $H_2NCHCO_2^-$	Log K_1	+0.98	+1.03	+1.09	+1.26	+1.28	+1.36	+1.43	
	Log β_2	+2.0	+2.5	+2.5	+2.6	+2.8	+3.2	+3.3	

[a] Relative to Ce^{3+}.

metal ion hydrolysis, and the completely deprotonated amino acid lanthanide complexes apparently do not exist under normal reaction conditions.

Complexation with the carboxylate groups of protonated amino acids (MHL) has been measured with a few lanthanides. Estimates in Table I are based on the observed trends of the stability constants of lanthanum(III) carboxylates with basicity of the carboxylate group. Lanthanide internal trends relative to lanthanum(III) are shown in Table III.

The stability constants of scandium(III) complexes of protonated amino acids can be estimated from the lanthanum values and the observed log K ratio La^{3+}/Sc^{3+} of 0.60 observed for carboxylate complexes [2]. Yttrium(III) constants can be estimated in a similar way from the La^{3+}/Y^{3+} ratio of logarithms of stability constants of 1.03 for mono-carboxylates and 0.85 for di-carboxylates [2].

Actinide Complexes

Actinide complexes are apparently similar to the lanthanide complexes in that they hydrolyze at p[H] values lower than that required to deprotonate the amino group in the presence of the metal ion and only complexation to the carboxylate group is observed. The uranyl ion has been the most widely investigated actinide ion, and several studies indicate that the amino proton is not displaced at p[H] values below about 4.3, above which precipitation becomes a problem [14]. Am^{3+} and UO_2^{2+} values in Table I were estimated from trends of americium(III) and uranyl carboxylates with basicity of the carboxylate group. NpO_2^+ values can be estimated from the UO_2^{2+} values and the NpO_2^+/UO_2^{2+} ratio of 0.40 for the logarithms of the stabilities of the carboxylate complexes. The only trivalent actinide value other than that of Am^{3+} indicates that Cm^{3+} values are 0.11 log units higher than those of Am^{3+}, which is similar to the behavior observed with the corresponding alkylcarboxylates.

Transition and Post-Transition Metal Complexes

The stability constants of the Ni^{2+} complexes of amino acids are listed in Table I. Log β_2 values were estimated from the observed log K_1/log β_2 ratio of 0.55, which was found to hold quite closely for the entire transition and post transition metal series, including Cu(II). The decrease in magnitude going from α-amino acids to β-amino acids suggests the presence of chelates with 5- and 6-membered rings. Adding a second carboxylate or amine shows an additional chelate effect. Steric effects are important as shown by the decrease in stability constants as substitution increases at the methylene carbon of α-amino acids, in going to secondary and tertiary nitrogens, and by increasing the branching of the N-alkyl groups.

Internal trends of the stability constants Cr^{2+}, Mn^{2+}, Fe^{2+}, Co^{2+}, Zn^{2+}, and Cd^{2+} relative to Ni^{2+} are shown in Table IV. The magnitudes increase going from Mn^{2+} to Ni^{2+} as would be expected from ligand field theory and then decreases to Zn^{2+} and Cd^{2+}. As another carboxylate or especially another amino group is added to the molecule, the magnitudes of the stability constants of Ni^{2+} complexes are increased more than those of other ions. The variations of logarithms of the stability constants relative to Ni(II) is illustrated in Fig. 1 for simple alkyl monoamino monocarboxylic acids, and for aspartic acid. The sharp maximum at Cu(II) is well known to be due to the stabilizing contribution of the Jahn-Teller effect. The broken lines indicate the general trend expected in the absence of this effect (*i.e.*, for an octahedral ligand field). Especially obvious from the plots in Fig. 1 is the much greater variation of β_2 than for β_1 for both types of amino acids. Apparently the affinity of the second ligand for the metal ion is much more sensitive to the nature of the metal ion than that of the first. This effect may be due to steric considerations as well as the differences in coordinate bond strength and ligand field effects.

TABLE IV. Internal Trends of Transition^{2+}, Zn^{2+}, and Cd^{2+} Complexes Relative to Ni^{2+}.

		Cr^{2+}	Mn^{2+}	Fe^{2+}	Co^{2+}	Ni^{2+}	Zn^{2+}	Cd^{2+}
$H_2NCHRCO_2^-$	log K_1	−1.4	−3.00 ± 0.02	−1.7 ± 0.1	−1.10 ± 0.04	0.0	−0.80 ± 0.05	−1.55 ± 0.05
	log β_2	−2.9	−5.6 ± 0.2	−2.9	−2.1 ± 0.1	0.0	−1.3 ± 0.1	−2.8 ± 0.2
$H_2NCHRCH_2CO_2^-$	log K_1	−0.5			−1.0 ± 0.1	0.0	−0.5 ± 0.1	
	log β_2				−1.9 ± 0.1	0.0	−0.3	
$CH_2CO_2^-$	log K_1		−3.4	−2.4	−1.20 ± 0.03	0.00	−1.33 ± 0.04	−2.80 ± 0.05
$H_2NCHCO_2^-$	log β_2		[5.2]	[3.6]	−2.2 ± 0.1	0.0	−2.3 ± 0.0	−4.9
H_2NCH_2	log K_1				−1.88	0.00		
$H_2NCHCO_2^-$	log β_2				−3.8	0.0		
$H_2N(CH_2)_2$	log K_1				−2.16	0.00	−2.27	
$H_2NCHCO_2^-$	log β_2				−4.0	0.0	−4.1	

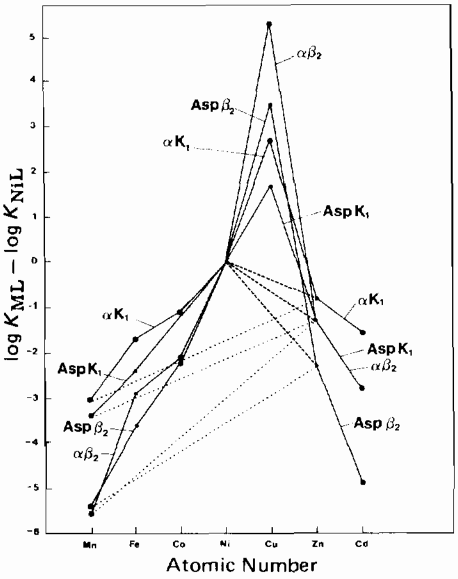


Fig. 1. Trends of stability constants of some transition and post-transition divalent metal complexes; αK_1 = log K_1 with alkyl-α-aminoacids; $\alpha\beta_2$ = log β_2 with alkyl-α-amino acids; Asp K_1 = log K_1 with aspartic acid; Aspβ_2 = log β_2 with aspartic acid; ———, experimental or estimated values; — — —, variation expected in absence of Jahn-Teller effect; ·······, variation expected on the basis of electronegativity changes alone (*i.e.*, in the absence of ligand field stabilization).

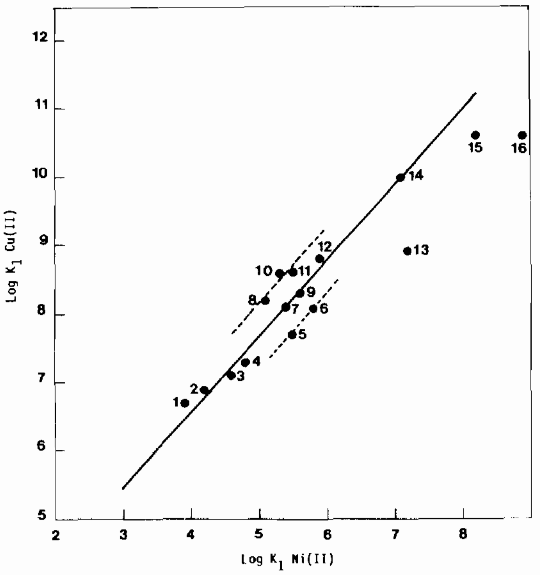


Fig. 2. Linear relationships of stability constants of amino acid complexes of copper(II) and nickel(II). 1. N-2-propylglycine, 2. N,N-diethylglycine, 3. β-alanine, 4. N-alkylglycine, 5. Sarcosine, 6. Glycine, 7. Alkylglycine (alanine, etc.), 8. 2-Methylalanine, 9. Glutamic acid, 10. 1-aminocycloheptanecarboxylic acid, 11. 1-aminocyclohexanecarboxylic acid, 12. Proline, 13. Aspartic acid, 14. Ornithine, 15. 2,3-Diaminopropanoic acid, 16. 2,4-Diaminobutanoic acid.

The initial rationale for the separate listing of Cu(II) and Ni(II) in Table I was the expectation that their stabilities would not correlate well with each other because of differences in coordination number and coordination geometry. This idea is tested in Fig. 2, which plots graphically the logarithms of the stability constants of the Cu(II) complexes *vs.* those of Ni(II). It is seen that nearly all amino acids selected show a nearly uniform ratio for log stabilities of Cu(II) to Ni(II). Accordingly, the slope of the linear correlation in Fig. 2 is close to unity, and the intercept is close to zero on both axes. It is apparent that the differences in coordination geometry of copper(II) and nickel(II) would have little influence on the relative binding of the bidentate amino acids. With aspartate however, octahedral Ni(II) would easily accommodate all three donor groups, whereas square planar Cu(II) would not. It is seen that the ligands that deviate considerably from the linear relationship in Fig. 2 are those that require an octahedral coordination sphere to bind all donor groups simultaneously. Minor deviations involving the degree of substitution at the methylene carbon are indicated by dotted lines in Fig. 2. Nickel(II) stability constants decrease in the order: glycine > alanine > 2-methylalanine while copper(II) values are approximately constant. This probably involves steric effects which are important with octahedral nickel(II) bonding but not with square planar copper(II) bonding.

Iron(III) Complexes

Iron(III) forms strong complexes with α-amino acids because of chelate formation with the amino and carboxylate donor groups. Estimates indicated in Table I were made based on the measured values and the expected trend with the basicity of the amino group. Estimates were not made with ligands containing the amino group at a greater distance from the carboxylate or with two amino groups in the molecule because of lack of data from iron(III) amine complexes. It should be noted that none of these amino acid complexes of Fe(III) are sufficiently stable to prevent precipitation of Fe(III) hydroxide from neutral or alkaline solution. Existence of the Fe(III) complexes in Table I is therefore restricted to acid solutions.

Silver(I) Complexes

The measured stability constants of the silver(I) ion with amino acids are proportional to the amine protonation constants except for glutaric acid, for which the observed enhancement is apparently due to the extra carboxylate group. Estimates for log K_1 are based on the amine protonation constants except for the dicarboxylates, which are based on the glutaric acid value, and the diamines, which are based on those measured with primary aliphatic diamines. Log β_2 values are estimated from the observed log K_1/log β_2 ratio of 0.48.

Comparison of stability constants obtained with alkylamines indicates that those with primary amino groups are 0.50 log units larger than those formed with secondary amines which are 0.35 log units larger than those formed with tertiary amines, indicating steric interference by the alkyl groups in complex formation with Ag(I).

The primary amino acids have silver(I) stability constants 0.20 log units larger than those of the alkylamines, which suggests the possibility of weak coordination of the carboxylates. However, given the linear coordinating tendency of Ag(I), and its coordination number 2, it is obvious that α-amino acids can coordinate only very weakly through the carboxylate donor, taking into consideration the strong binding to the amino group.

The stability constants for α-amino acids show approximately the same variations with basicity of the amino group as those where the two donor groups are more widely separated, indicating that there is little or no chelate effect involved in the formation of these complexes.

Methylmercury Complexes

Methylmercury stability constants with amino acids (Table I) show a linear dependence on the amine protonation constants (*i.e.*, with amino group basicity). Estimated values for log K_1 are based on the amine protonation constants except for those containing secondary or tertiary amino groups which are adjusted for the expected steric interference. Those containing an additional carboxylate or amino group have been adjusted for the expected enhanced intramolecular coordination. Log β_2 values are estimated from the observed log K_1/log β_2 ratio of 0.82.

Other Metal Complexes

Some stability constants have been measured with other metal ions and a few amino acids which are not listed in the tabulated data in this paper, but there are uncertainties about their trends with structure and constitution, and only orders of magnitude are available now. The following values with glycine may be taken as samples of these data: $(CH_3)_3Pb^+$, log K_1 = 1.5; Hg^{2+}, log K_1 = 9 ± 3; log β_2 = 16 ± 1; Pb^{2+}, log K_1 = 4.8 ± 0.2, log β_2 = 8.1 ± 0.1; Ga^{3+}, log K_1 = 9.6 ± 0.3.

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References

1 A. E. Martell, R. J. Motekaitis and R. M. Smith, in K. J. Irgolic and A. E. Martell (eds.), 'U.S.-Italy Workshop on Environmental Chemistry', Verlag-Chemie International, Deerfield Beach, Fla., 1984, in press.

2 R. M. Smith, A. E. Martell and R. J. Motekaitis, *Inorg. Chim. Acta, 99,* 207 (1985).

3 R. J. Motekaitis and A. E. Martell, to be submitted.

4 J. C. Westall, J. L. Zachary and F. M. M. Morel, 'Technical Note No. 18', EPA Grant NO. R-803738, 1976.

5 A. Popovici, C. F. Geschickter, A. Reinovsky and M. Rubin, *Proc. Soc. Exp. Biol. Med., 74,* 415 (1950).

6 M. Rubin, S. Gignac, S. P. Bessman and E. L. Belknap, *Science, 117,* 659 (1953).

7 R. M. Smith and A. E. Martell, 'Critical Stability Constants, Vols. 1–5', Plenum, New York, 1974, 1975, 1976, 1977, 1982.

8 R. M. Smith and A. E. Martell, 'Critical Stability Constants, Vol. 6', in preparation.
9 M. Cefola, A. S. Tompa, A. V. Celiano and P. S. Gentile, *Inorg. Chem.*, *1*, 290 (1962).
10 F. Belyaeva, M. A. Porai-Koshits, T. I. Malinovskii, L. S. Aslanov, L. S. Sukhanova and L. I. Martynenko, *J. Struct. Chem. (USSR)*, *10*, 470 (1969); *Zh. Struk. Khim.*, *10*, 557 (1969).
11 B. S. Mathur and T. S. Srivastava, *J. Inorg. Nucl. Chem.*, *32*, 3277 (1970).
12 L. I. Katzin and E. Gulyas, *Inorg. Chem.*, *7*, 2442 (1968).
13 R. D. Hancock, *J. Chem. Soc., Dalton Trans.*, 1384 (1979).
14 N. C. Li, E. Doddy and J. M. White, *J. Am. Chem. Soc.*, *80*, 5901 (1958); I. Feldman and L. Koval, *Inorg. Chem.*, *2*, 145 (1963); V. V. Ramanujam, K. Rengaraj and B. Sivasankar, *Bull. Chem. Soc. Jpn.*, *52*, 2713 (1979); A. Bismando, L. Rizzo, G. Tomat, D. Curto, P. DiBernardo and A. Cassol, *Inorg. Chim. Acta*, *74*, 21 (1983).