ligand donor groups that must be brought together to form metal chelate rings with the metal ion. When two or more phosphonate groups are present in the ligand, apparently these repulsions become great enough to decrease the formation constants for the combination of the ligand with divalent metal ions.

It is possible that this effect is overcome sufficiently with metal ions of higher charge, however, so that increased stabilities would be obtained with two or more phosphonate groups on the chelating ligand. The data needed to test this interpretation are being gathered and will be the subject of a subsequent publication.

New Multidentate Ligands. VIII. Metal Complexes of N,N′-Diglycylethylenediamine in Aqueous Solutions[1]

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Abstract: The specific coordinating tendencies of the synthetic ligand, N,N′-diglycylethylenediamine, toward Cu(II) and Ni(II) ions have been studied as a model for peptide–metal binding. The coordination compounds formed are unique in that hydrogen ions are displaced by these metal ions from the amide linkages, and in the formation of a low pH polynuclear copper(II) chelate that depolymerizes to a mononuclear chelate in alkaline solution. The copper(II) chelate, CuL^{2+}, releases amide protons in two distinct steps, to give the species $Cu_2(H_{-1}L)_2^{2+}$ and $CuH_{-2}L$. On the other hand, the nickel(II) chelate of the neutral ligand, NiL^{2+}, loses both amide protons simultaneously to give the chelate $NiH_{-2}L$. Cobalt(II) and zinc(II) ions do not displace the peptide protons, forming only the "ordinary" complexes, ML^{2+}, in which the metal ion is coordinated to the neutral ligand. In the types of reactions with metal ions, the ligand resembles triglycine and tetraglycine, but contrasts with these and other peptides in the manner in which it forms polynuclear copper(II) chelates. Electronic spectra of the Cu(II) and Ni(II) chelates formed are measured and discussed.

Metal complexes of N,N′-diglycylethylenediamine (DGEN) were first studied by Pfeiffer and Saure[2] in their investigation of the biuret reaction with a series of ligands containing amide groups. They found that the copper(II)–DGEN complex, as well as biuret complexes, has a deep violet color in alkaline solution. For such complexes, they suggested a structure in which a proton is displaced from the amide group coordinated to the copper(II) ion. They also suggested that nickel(II) chelates of these compounds have similar structures. The stability constants of DGEN complexes of various metal ions were first reported by Chakraburtty and coworkers from their potentiometric studies.[3] These authors[3] first observed that the "ordinary" complexes of DGEN with copper(II) and nickel(II) react further with alkali hydroxide, and interpreted the reaction as hydrolysis of the metal complexes.

Subsequently, similar reactions were observed in systems of copper(II)[4–14] and nickel(II)[6,15,16] complexes of many amino acid peptides, and it is now established that, in acidic or near neutral pH, the peptide protons, not those of coordinated water molecules, are being dissociated in such systems.[4–16] The displacement of the peptide or amide protons is also known in the solid state of both copper(II)[6,17,18] and nickel(II)[6,17] chelates. Direct evidence for the peptide proton dissociation was obtained from aqueous (D_2O) infrared studies[13,14,16] for complexes in solution. The result of X-ray crystallographic studies[18] has shown unequivocally the peptide proton displacement in the solid and the coordination of the resulting negative-peptide groups to metal ion.

All of the copper(II) and nickel(II) complexes of peptides and amides known so far, in which the peptide or amide protons are displaced, have at least one amino group in the ligand. Thus, it appears that the presence of the strongly coordinating amino group(s) may be necessary for strong chelation which stabilizes the complex resulting from displacement of the amide protons. In this respect, DGEN appeared to be an ideal ligand for the study of metal–peptide interactions because it has two amino groups, one on each end of the chain, and the chelate rings resulting from the displacement of the peptide protons would all be five membered. Thus DGEN is expected to form the negative-peptide

(1) This work was supported by the U. S. Atomic Energy Commission under Contract No. AT (40-1) 3621 with Texas A&M Research Foundation.
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Table I. Equilibrium Constants for the Interaction of DGEN with Metal Ions[a]

Equilibrium quotient,[b] K	Log K				
	H^+	Co^{2+}	Ni^{2+}	Cu^{2+}	Zn^{2+}
$K_1^H = \frac{[HL^+]}{[L][H^+]}$	8.22				
$K_2^H = \frac{[H_2L^{2+}]}{[HL^+][H^+]}$	7.48				
$K_1 = \frac{[ML^{2+}]}{[M^{2+}][L]}$		3.30	5.38	7.50	3.95
$\beta_2 = \frac{[ML_2^{2+}]}{[M^{2+}][L]^2}$			8.50		
$K_{1Ad} = \frac{[(MH_{-1}L)_2^{2+}][H^+]^2}{[ML^{2+}]^2}$				−9.2	
$K_{1Bm} = \frac{[MH_{-2}L]^2[H^+]^2}{[(MH_{-1}L)_2^{2+}]}$				−18.40	
$K_{1AB} = \frac{[MH_{-2}L][H^+]^2}{[ML^{2+}]}$			−16.04	−13.8	

[a] Medium contains 0.10 M KNO_3; 25°. [b] L represents the neutral ligand DGEN.

chelates readily. Moreover, the polypeptide complexes of nickel(II) in which the peptide protons are displaced are known to be planar[15,16] as are most copper(II) complexes.

In addition to reactions of this ligand that are analogous to those of peptides, this paper reports for the first time a unique polynuclear–mononuclear chelate equilibrium in which depolymerization occurs with an increase in pH.

Experimental Section

Reagents. DGEN was synthesized by the method of Cottrell and Gill.[19] The compound was obtained in the form of the dihydrochloride salt, mp 245–247° dec. The assay was 99.8% pure as determined by titration with standard base. All of the metal salts used were the usual reagent grade chemicals, and their solutions were standardized by titration with EDTA by the method of Schwarzenbach.[20] Aliquots of the metal solutions were also passed through Dowex 50WX8 cation-exchange resin, and the effluent solutions were titrated with a standard sodium hydroxide solution. In all cases, the two determinations agreed with each other within 0.1%. Tetraglycine was Mann Research Laboratories' MA grade and was used without further purification.

Measurements. A Beckman Research Model 101900 pH Meter fitted with extension glass and calomel electrodes was used for all pH measurements. The pH meter–electrode system was standardized in terms of hydrogen ion concentration using standard solutions of hydrochloric acid, acetic acid–sodium acetate buffers, and sodium hydroxide, adjusted to an ionic strength of 0.10 with potassium nitrate as described by Courtney, *et al.*[21]

Aliquots of $DGENH_2Cl_2$ solutions were titrated with a standard sodium hydroxide solution in the absence and in the presence of metal ions. In all titrations, the ionic strength was maintained relatively constant by using a supporting electrolyte of 0.10 M KNO_3. All experiments were carried out at 25.0 ± 0.1°. During titration, nitrogen gas, passed initially through dilute alkali solution and then through 0.1 M KNO_3 solution, was bubbled through the reaction mixture.

The reaction of copper(II) with DGEN was fast as is generally the case with other ligands. In the case of nickel(II)–DGEN, the dissociation of the peptide protons was very slow; parts of the titration curves were obtained by batchwise titration. Here, each point required 1–3 hr to reach equilibrium.

The visible–near-infrared spectra were obtained with a Cary Model 14 recording spectrophotometer. Here, potassium chloride was substituted for potassium nitrate. In the case of nickel(II)–DGEN, spectra were taken several hours after the preparation of the solutions to ensure complete equilibration. All spectra were obtained with matched 1.00-cm cells.

Results

DGEN. The potentiometric titration curves of DGEN are shown in Figure 1 (curves 1 and 2). It is seen that the two ammonium protons dissociate in closely overlapping steps. The two acid dissociation

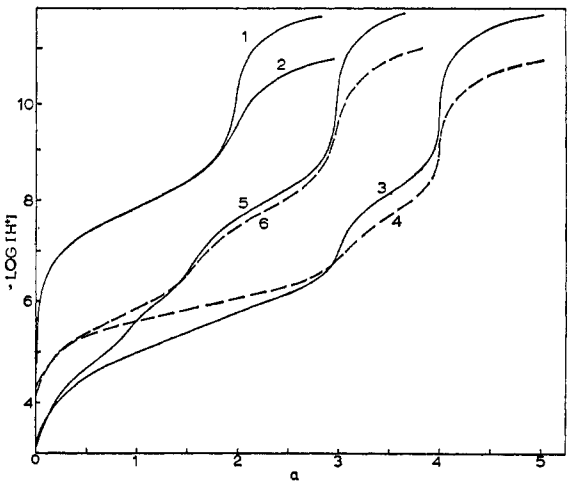

Figure 1. Potentiometric equilibrium curves for DGEN and its copper(II) complex systems. Test solutions: (1) 40.0 ml, 0.0100 M $DGENH_2Cl_2$; (2) 100 ml, 0.00100 M $DGENH_2Cl_2$; (3) 20.0 ml, 0.0100 M in both $Cu(NO_3)_2$ and $DGENH_2Cl_2$; (4) 100 ml, 0.00100 M in both $Cu(NO_3)_2$ and $DGENH_2Cl_2$; (5) 20.0 ml, 0.0100 M in $Cu(NO_3)_2$ and 0.0200 M in $DGENH_2Cl_2$; and (6) 100 ml, 0.00100 M in $Cu(NO_3)_2$ and 0.00200 M in $DGENH_2Cl_2$. All solutions are 0.10 M in KNO_3; titrant: 0.1006 M NaOH. a = number of moles of sodium hydroxide added per mole of ligand.

constants (K_n^H), defined as shown in Table I, where L represents the neutral ligand DGEN, were obtained by a simple least-squares method in −log [H^+]. The values obtained from curve 1 are log K_1^H = 8.23 and

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Table II. Equilibrium Constants for the Reaction in Eq 3 Obtained from Various Titration Data

Figure 1, curve no.	Initial concn, T_M, M	Ratio $T_M:T_L$	Range, a, employed	No. of pts taken	pK_{1Bm}
3	0.01	1:1	3.2–3.8	7	18.39 ± 0.03
4	0.001	1:1	3.2–3.8	7	18.40 ± 0.04
5	0.01	1:2	2.0–2.9	5	18.44 ± 0.17
6	0.001	1:2	2.0–2.9	4	18.47 ± 0.35

log K_2^H = 7.49. The corresponding values obtained from one-tenth as concentrated solution (curve 2) are 8.20 and 7.47, respectively. The average values were taken as the best values and are listed in Table I.

Copper(II)–DGEN. The titration curves for copper(II)–DGEN systems are shown in Figure 1 (curves 3–6). The reactions are fast. During titration, the color of the solutions gradually changed from light blue to deep blue to violet and finally to purple. For the 1:1 copper(II)–DGEN solutions (curves 3 and 4) two definite breaks are seen, one at a = 3 and the other at a = 4, where a represents the number of moles of sodium hydroxide added per mole of ligand present in the reacting solution. In the first sloping buffer region, the two protons attached to the terminal groups of the ligand and one of the two peptide protons are being dissociated by the copper(II) ion in overlapping steps.

The following complex equilibria were found to take place in the reaction of copper(II) with DGEN

$$M^{2+} + L \rightleftharpoons ML^{2+} \qquad K_1 = \frac{[ML^{2+}]}{[M^{2+}][L]} \tag{1}$$

$$2ML^{2+} \rightleftharpoons (MH_{-1}L)_2^{2+} + 2H^+ \qquad K_{1Ad} = \frac{[(MH_{-1}L)_2^{2+}][H^+]^2}{[ML^{2+}]^2} \tag{2}$$

$$(MH_{-1}L)_2^{2+} \rightleftharpoons 2MH_{-2}L + 2H^+ \qquad K_{1Bm} = \frac{[MH_{-2}L]^2[H^+]^2}{[(MH_{-1}L)_2^{2+}]} \tag{3}$$

where the negative subscript to H represents the number of peptide protons removed from the complex. The first two constants, K_1 and K_{1Ad}, were evaluated by the method of Schwarzenbach, *et al.*,[14,22] as outlined below.

Assuming that reactions 1 and 2 take place in the region from a = 0 to a = 3 of the 1:1 titration curves, the following relationship between K_1 and K_{1Ad} can be obtained.

$$K_{1Ad} = \frac{(T_L - [L]\{\alpha + K_1(T_M - T_L + \alpha[L])\})[H^+]^2}{2K_1^2[L]^2(T_M - T_L + \alpha[L])^2} \tag{4}$$

and

$$a_1[L]^2 + b_1[L] + c_1 = 0 \tag{5}$$

where

$$a_1 = K_1\alpha$$

$$b_1 = K_1(T_M - T_L) + \alpha + \alpha'$$

$$c_1 = -3T_L + T_{OH} + [H^+] - K_w/[H^+]$$

$$\alpha = [H^+]^2K_1^HK_2^H + [H^+]K_1^H + 1$$

$$\alpha' = 2[H^+]^2K_1^HK_2^H + [H^+]K_1^H$$

The two quantities T_M and T_L represent total concentrations of metal ion and DGEN, respectively, and T_{OH} represents the total concentration of sodium hydroxide added. The value used for K_w is the concentration constant of water obtained from the data of Harned and Owen.[23] The constants were evaluated in the vicinities of $a \sim 2$.

Reaction 3 takes place between a and 3 and a = 4 in the 1:1 titration (curves 3 and 4 of Figure 1). For this reaction, the following two simultaneous equations are obtained.

$$\alpha'^2[L]^2 + \alpha'\left(2T_M - 2S + \frac{K_{1Bm}}{2[H^+]^2}\right)[L] + T_M^2 + S\left(S - 2T_M - \frac{K_{1Bm}}{2[H^+]^2}\right) = 0 \tag{6}$$

$$\alpha_d^2[L]^2 + \left\{2\alpha_d(T_L - S) + \frac{\alpha'K_{1Bm}}{2[H^+]^2}\right\}[L] + T_L^2 + S\left\{S - 2T_L - \frac{K_{1Bm}}{2[H^+]^2}\right\} = 0 \tag{7}$$

where

$$\alpha_d = \alpha' - \alpha$$

$$S = 2T_M + 2T_L - T_{OH} - [H^+] + K_w/[H^+]$$

The only unknowns in eq 6 and 7 are [L] and the unknown constant K_{1Bm}. These two equations were solved simultaneously by successive approximation at each titration point. The results are summarized in Table II. The value pK_{1Bm} = 18.40 was taken as the best and is given in Table I. The fact that the same value was obtained for K_{1Bm} from both levels of concentration indicates that the model in reaction 3 is correct. It is also the only reaction taking place in the 1:2 system in the region from a = 1.5 to a = 3. This is in accord with the previous finding that, below a = 3 in the 1:1 titration (or below a = 1.5 in the 1:2 titration), the most basic copper(II) species present is $(CuH_{-1}L)_2^{2+}$, not the monomer $CuH_{-1}L^+$. The absence of the monomer is thus established in both ranges of a values. No other models for the reaction sequences in the Cu(II) system gave constant equilibrium constants over a range of reaction conditions.

Above a = 4, curves 3 and 4 of Figure 1 are the same as those calculated assuming no further reaction between $CuH_{-2}L$ and the hydroxide ion. The same is true for curves 5 and 6 above a = 3.

Visible Spectra of Copper(II)–DGEN. The absorption spectra of copper(II)–DGEN solutions are shown in Figure 2. Before addition of base, little complex formation takes place (curves 1 and 2). As the concentration of base is increased, the band maximum shifts to shorter wavelengths and increases in intensity. Curve 5, at a = 3 for a 1:1 ratio, is almost entirely due to the species $(CuH_{-1}L)_2^{2+}$. Curves 5, 6,

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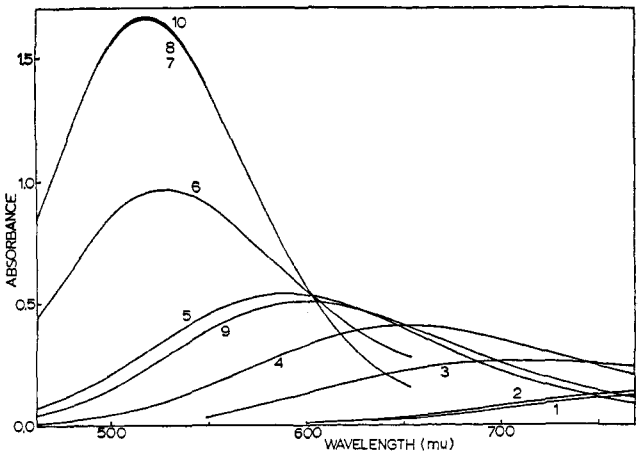


Figure 2. Visible absorption spectra of copper(II)–DGEN solutions. All solutions are 0.0100 M in $Cu(NO_3)_2$ and 0.10 M in KCl. Molarities of $DGENH_2Cl_2$ and a values are, respectively: (1) 0.00, 0.00; (2) 0.0100, 0.00; (3) 0.0100, 1.006; (4) 0.0100, 2.01; (5) 0.0100, 3.02; (6) 0.0100, 3.52; (7) 0.0100, 4.02; (8) 0.0100, 5.03; (9) 0.0200, 1.51; and (10) 0.0200, 3.02. a = number of moles of sodium hydroxide added per mole of ligand.

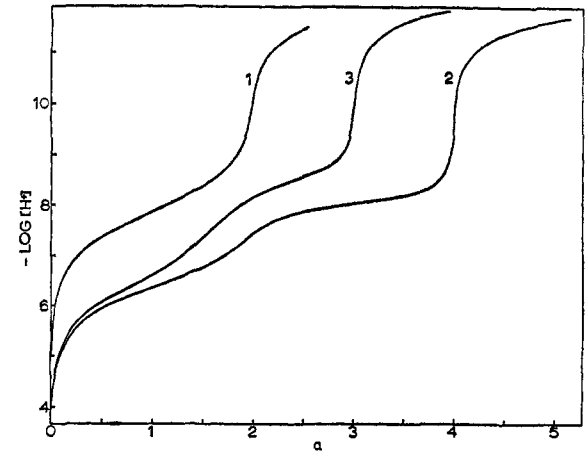


Figure 3. Potentiometric equilibrium curves for nickel(II)–DGEN systems. Test solutions: (1) 40.0 ml, 0.0100 M $DGENH_2Cl_2$; (2) 20.0 ml, 0.0100 M in both $Ni(NO_3)_2$ and $DGENH_2Cl_2$; (3) 20.0 ml, 0.0100 M in $Ni(NO_3)_2$ and 0.0200 M in $DGENH_2Cl_2$. All solutions are 0.10 M in KNO_3; titrant: 0.1006 M NaOH.

and 7 give an isosbestic point (at 603 mμ with A = 0.535) which confirms the potentiometric finding that only two species, $(CuH_{-1}L)_2^{2+}$ and $CuH_{-2}L$, are present in the region from a = 3 to a = 4 in the 1:1 solution. Curve 8 is identical with curve 7. This supports the potentiometric observation that no further reaction takes place between $CuH_{-2}L$ and hydroxide ion. Curve 9, at a = 1.5 in the 1:2 system, shows slightly lower intensity and appears at slightly longer wavelengths than does curve 5. This is due to the presence of a small amount of CuL^{2+} in the solution in addition to the main species $(CuH_{-1}L)_2^{2+}$. In this solution, most of the metal ion and the ligand exist as $(CuH_{-1}L)_2^{2+}$ and H_2L^{2+} at a 1:1 molar ratio, in accord with the inflection at a = 1.5 in the 1:2 potentiometric determination (Figure 1, curves 5 and 6). There also exists a small amount of CuL^{2+} and HL^+ (in 1:1 proportion) which still agrees with the inflection at a = 1.5 in the 1:2 titration. Since the value of $-\log [H^+]$ at a = 1.5 is only about 0.9 log unit lower than log K_2^H, the presence of the latter pair is obvious.

Curve 10 is identical with curves 7 and 8. This supports the potentiometric observation that only $CuH_{-2}L$ exists in both the 1:1 solution at a = 4 and the 1:2 solution at a = 3. The inflections at a = 3 in the 1:2 potentiometric titrations agree with the formation of $CuH_{-2}L$ from half of the ligands and the complete neutralization of the uncomplexed half of the ligands.

Nickel(II)–DGEN. The potentiometric titration curves for nickel(II)–DGEN systems are shown in Figure 3. Below $a \sim 2$ in the 1:1 solution, the system required several minutes to reach equilibrium after each addition of titrant. Between a = 2 and a = 4, 1–3 hr was required for the solution to reach equilibrium after addition of titrant. Above a = 4, the system reached equilibrium very rapidly. As titration proceeds, the initial pale green solution turns bluish. At a = 2, the solution exhibits a greenish tint. Beyond a = 2, the solution has a yellow color. That the reaction is slow and that the solution turns yellow on titration of the peptide protons are the same as observed in some other nickel(II)–peptide systems.[15,16] The reaction in the 1:2 system (curve 3) was more rapid; in the region from a = 1.5 to a = 3, each point required about 30 min to 1 hr. The difference in the rate of the reaction between 1:1 and 1:2 systems is probably a pH effect.

Equations 1, 8, and 9 were found to describe the reactions of nickel(II) with DGEN.

$$M^{2+} + L \rightleftharpoons ML^{2+} \qquad K_1 = \frac{[ML^{2+}]}{[M^{2+}][L]} \tag{1}$$

$$ML^{2+} + L \rightleftharpoons ML_2^{2+} \qquad K_2 = \frac{[ML_2^{2+}]}{[ML^{2+}][L]} \tag{8}$$

$$ML^{2+} \rightleftharpoons MH_{-2}L + 2H^+ \qquad K_{1AB} = \frac{[MH_{-2}L][H^+]^2}{[ML^{2+}]} \tag{9}$$

Below $a \sim 2$ for the 1:1 titration (Figure 3, curve 2), only reaction 1 takes place. From a Bjerrum plot below about a = 1.6, a value of 5.38 was obtained for log K_1. Approximately the same value (5.46) was obtained from the 1:2 titration curve below about a = 0.9. From this 1:2 curve between about a = 1.3 and a = 1.7, a value of 3.04 was obtained for log K_2. Reaction 9 takes place in the region from a = 2 to a = 4 in 1:1 solution. The value of K_{1AB} was calculated at each titration point from a = 2.3 to a = 3.8 from the equation

$$K_{1AB} = [H^+]^2\left[\frac{2T_M}{S + (T_M - T_L)\alpha'/\alpha} - 1\right] \tag{10}$$

The result of calculations for the seven points taken is $pK_{1AB} = 16.04 \pm 0.16$.

Reactions 8 and 9 describe the equilibria in the region from a = 1.5 to a = 3 for the 1:2 titration. The following two simultaneous equations may be derived for this system.

$$f_1[L]^2 + f_2[L] + f_3 = 0 \tag{11}$$

$$g_1[L]^2 + g_2[L] + g_3 = 0 \tag{12}$$

where

$$f_1 = 2(\alpha' - \alpha)K_2$$

$$_2 = 2(T_L - S)K_2 - 2\alpha + \alpha'(1 + K_{1AB}/[H^+]^2)$$

$$f_3 = 2T_L - S(1 + K_{1AB}/[H^+]^2)$$

$$g_1 = \alpha' K_2$$

$$g_2 = \alpha'(1 + K_{1AB}/[H^+]^2) + (2T_M - S)K_2$$

$$g_3 = 2T_M - S(1 + K_{1AB}/[H^+]^2)$$

In these two equations, everything is known except [L] and the unknown parameter K_{1AB}. These two equations were solved simultaneously by successive approximation at each point in the region from $a \sim 2.0$ to $a \sim 2.9$. Below about $a = 2.0$, the concentration of $NiH_{-2}L$ is very small. The calculation gave $pK_{1AB} = 16.16 \pm 0.24$ which is approximately equal to the value obtained from the 1:1 titration. Therefore, no species other than Ni^{2+}, NiL^{2+}, NiL_2^{2+}, and $NiH_{-2}L$ exists in the 1:2 solution.

As in the case of copper(II)–DGEN systems, above $a = 4$ for the 1:1 solution or above $a = 3$ for the 1:2 solution, the titration curves are the same as those calculated assuming no reaction between $NiH_{-2}L$ and the base.

Visible Spectra of Nickel(II)–DGEN. The absorption spectra of nickel(II)–DGEN systems are shown in Figure 4. Before addition of base, practically no

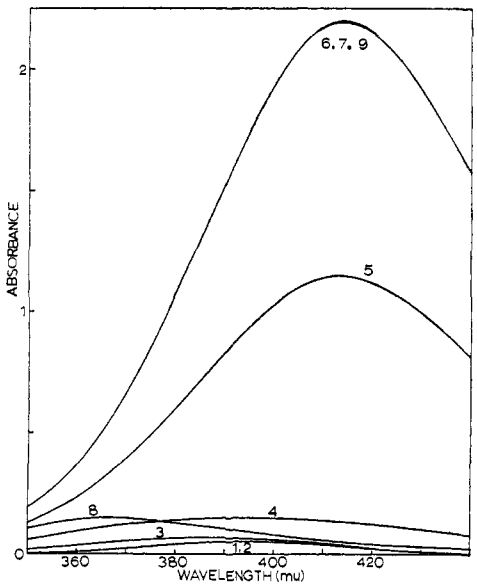


Figure 4. Absorption spectra of nickel(II)–DGEN solutions. All solutions are 0.0100 M in $Ni(NO_3)_2$ and 0.10 M in KCl. Molarities of $DGENH_2Cl_2$ and a values are, respectively: (1) 0.00, 0.00; (2) 0.0100, 0.00; (3) 0.0100, 1.006; (4) 0.0100, 2.01; (5) 0.0100, 3.02; (6) 0.0100, 4.02; (7) 0.0100, 5.03; (8) 0.0200, 1.51; and (9) 0.0200, 3.02.

complex formation takes place (curves 1 and 2). As base concentration is increased to $a = 1$, the band shifts slightly toward shorter wavelength with slight increase in intensity (curve 3). Curve 4, at $a = 2$ for the 1:1 solution, is also seen to have a maximum at about 395 mμ. This is on the longer, not shorter, wavelength side of curve 3. Thus, the presence in this solution of a species with a geometry different from the octahedral NiL^{2+} is apparent. As will be shown below, the small concentration of $NiH_{-2}L$ present in this solution influenced the band maximum because $NiH_{-2}L$ has very high molar absorbance compared to those of NiL^{2+} and aqueous Ni^{2+}.

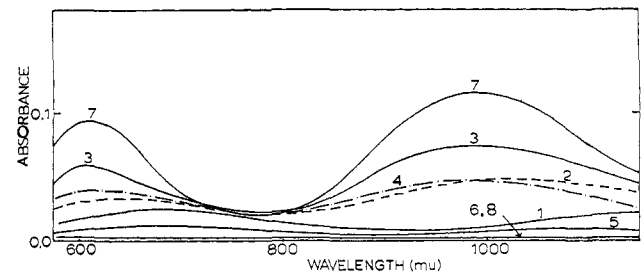


Figure 5. Absorption spectra of nickel(II)–DGEN solutions. All solutions are 0.0100 M in $Ni(NO_3)_2$ and 0.10 M in KCl. Molarities of $DGENH_2CL_2$ and a values are, respectively: (1) 0.0, 0.0; (2) 0.0100, 1.0; (3) 0.0100, 2.0; (4) 0.0100, 3.0; (5) 0.0100, 4.0; (6) 0.0100, 5.0; (7) 0.0200, 1.5; and (8) 0.0200, 3.0.

As base concentration is further increased beyond $a = 2$, the band sharply increases in intensity with a shift to longer wavelength. Reversal in direction of shift is indicative of structural change. Curve 6, at $a = 4$ for the 1:1 solution, is almost entirely due to the species $NiH_{-2}L$. At any wavelength, the intensity of curve 5, at $a = 3$ for the 1:1 solution, is almost exactly one-half of the sum of the intensities of curves 4 and 6 at that wavelength. This suggests that only two species are present between $a = 2$ and $a = 4$ in the 1:1 solutions, as observed in the potentiometric titration. The molar absorbance of the hypothetical species $NiH_{-1}L^+$ is not necessarily expected to be exactly one-half of the sum of those of NiL^{2+} and $NiH_{-2}L$. Curve 7 is identical with curve 6. This shows either that $NiH_{-2}L$ does not react with hydroxide ion or that coordination of hydroxide ion to $NiH_{-2}L$ does not alter the spectra. The potentiometric result is in accord with the former interpretation.

Curve 8, at $a = 1.5$ for the 1:2 solution, gives a band maximum at about 365 mμ which indicates the presence of NiL_2^{2+} in this solution, as observed in the potentiometric titration. As base concentration is increased beyond $a = 1.5$ in the 1:2 solution, reversal in the direction of the shift is again observed. Curve 9, at $a = 3$ in the 1:2 solution, is identical with curves 6 and 7. This also agrees with the potentiometric observation in that part of the titration curve.

The two longer wavelength bands of nickel(II)–DGEN solutions are shown in Figure 5. As base concentration is increased in the 1:1 solutions, both bands increase in intensity with slight shifts toward shorter wavelengths. Both bands attain maximum intensity at $a = 2$, beyond which the bands decrease gradually. At $a = 4$ or greater, both bands disappear completely. For the 1:2 solution both bands have their highest intensities at $a = 1.5$ and disappear completely at $a = 3$. The region scanned is from 550 to 1500 mμ.

Zinc(II)- and Cobalt(II)–DGEN. The potentiometric equilibrium curves for the zinc(II)- and cobalt(II)–DGEN systems are shown in Figure 6. It is seen that these two metal ions interact with DGEN only weakly, and that they do not displace the amide protons. The shapes of the titration curves just before the appearance of precipitates indicate that vertical inflections would have been observed at $a = 2$ if no precipitate formed. The formation constant for ML^{2+} was calculated for both metals and is listed in Table I.

Iron(III)–DGEN. In the presence of iron(III), the DGEN titration curve is extremely depressed, and the reaction is very slow. The 1:1 titration curve exhibits a sharp inflection at $a = 3$, followed almost immediately by formation of brown precipitates. The titration curve was very similar to that obtained in the absence of DGEN. Thus, the actual species formed in iron(III)–DGEN solutions may be simply $Fe(OH)_3$, or $FeL(OH)^{2+}$ at most. Since ferric ion has much greater affinity for oxygen than for nitrogen, simple hydrolysis may be taking place instead of interaction with DGEN. Iron(III) complexes of aliphatic amines are not generally known in aqueous solutions.

Discussion

DGEN. The values of the two acid dissociation constants of $DGENH_2Cl_2$ listed in Table I are in better agreement with those (at 25°, $\mu = 1.0$) of Chakraburtty, *et al.*,[3] than with the values (at 23°, $\mu = 0.5$) recently reported by Zuberbuehler and Fallab.[24] Since the value of $-\log [H^+]$ is about 0.1 log unit lower than pH in near-neutral solutions, the differences between our values and those of Chakraburtty, *et al.*,[3] lie largely in the difference in the definition of the constants.

The value of log K_2^H is about the same for both DGEN and ethylenediamine,[25] while log K_1^H for DGEN is lower than that of ethylenediamine by about 2 log units. In going from ethylenediamine to pentamethylenediamine,[25] the difference in the two log K values of each of the diamines regularly decreases from 2.8 to 0.9. Thus, the difference of 0.74 between the two log K values of DGEN is generally in accord with the dependence of $\Delta \log K$ on the distance between the two groups. On the other hand, the value of log K_2^H of DGEN is much smaller (by about 2.5 log units) than that of pentamethylenediamine. The same is true of log K_1^H. Therefore, the main factor affecting the basicity of each of the two amino groups in DGEN is the neighboring amide groups and not the other amino group.

Copper(II)–DGEN. Potentiometric Data. The 1:1 copper(II)–DGEN titration curves (Figure 1, curves 3 and 4) show that the two ammonium protons and one of the two peptide protons dissociate in overlapping steps, producing a single buffer region below pH 6.5, and that the second peptide proton comes off at about pH 8. In nearly all copper(II) complexes reported so far, the peptide or amide protons dissociate in a stepwise fashion. The only exception appears to be N,N′-diglycyltrimethylenediamine; the copper(II) complex of this compound reportedly releases the two peptide protons simultaneously.[24]

The value of log K_1 given in Table I contrasts with the value 8.13 reported by Chakraburtty, *et al.*,[3] who obtained the constant from a 1:2 titration curve in an ionic strength of 1.0 *M* KCl. In calculating the constant, they assumed that only Cu_{aq}^{2+} and CuL^{2+} are present when the total protons released are less than $2T_{Cu}$. The absence of inflection at $a = 2$ in the 1:1 titration (Figure 1, curves 3 and 4) shows that such an assumption is not valid. The existence of more basic species in the vicinity of $a = 2$ is apparent from the titration curves. Neglecting such species would tend to make the concentration of CuL^{2+} appear greater, and hence a greater value for K_1 would be obtained.

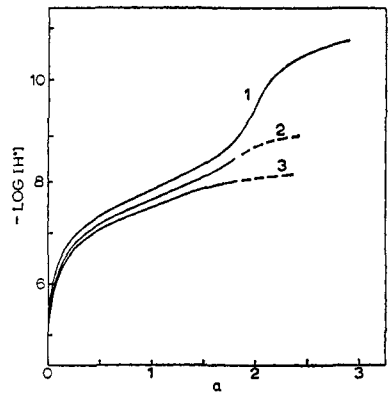


Figure 6. Potentiometric equilibrium curves for zinc(II)- and cobalt(II)–DGEN systems. Test solutions: (1) 100 ml, 0.00100 *M* $DGENH_2Cl_2$; (2) 100 ml, 0.00100 *M* in both $Co(NO_3)_2$ and $DGENH_2Cl_2$; (3) 100 ml, 0.00100 *M* in both $Zn(NO_3)_2$ and $DGENH_2Cl_2$. All solutions are 0.10 *M* in KNO_3; titrant: 0.1006 *M* NaOH. Broken lines represent appearance of precipitates.

The values of K_1 given in Table I for copper(II) and other metal ions studied are all about 2–3 log units lower than the corresponding constants for the corresponding ethylenediamine complexes,[25] and parallels the relative proton affinities of the two ligands. This indicates that the neutral peptide groups are not strongly bonded to the metal ion. The values of K_1 listed in Table I are approximately the same as the values of β_2 for the corresponding ammonia complexes.[25] This also indicates that coordination of the peptide groups in DGEN chelates is weak. It is further seen that little chelate effect results from coordination of the two terminal nitrogens, as expected from the distance between these two groups. The stability constant of CuL^{2+} listed in Table I is greater than the corresponding values for copper(II) complexes of glycine peptides[13,14] by about 2 log units. The higher stability in the former is a result of the presence of one more amino group in DGEN.

It was found that the complex species in which only one of the two peptide protons is displaced exists as the dimer, $(CuH_{-1}L)_2^{2+}$. The monomer does not exist in appreciable concentration under the present experimental conditions. Chakraburtty, *et al.*,[3] reported values of equilibrium constants corresponding to simple monomeric stepwise dissociation of the two additional protons. Since they studied only one concentration level and calculated the constants at limited ranges of a values with oversimplifying assumptions, they may not have noticed the irregularities in their constants. They reported the magnetic susceptibilities of CuL^{2+} and $CuH_{-2}L$ to be 1.92 and 1.72 BM, respectively, but no magnetic data were given regarding their postulated species $CuH_{-1}L^+$. Zuberbuehler and Fallab[24] also reported constants corresponding to the same monomeric stepwise dissociations. These authors[24] studied also only one concentration level although they varied the metal–ligand ratio from 1:2 to 2:1. To explain their titration data, they also postulated the protonated species, $CuLH^{3+}$, but no potentiometric equilibrium data were given. No evidence was found for the

(24) A. Zuberbuehler and S. Fallab, *Helv. Chim. Acta*, **50**, 889 (1967).
(25) L. G. Sillén and A. E. Martell, "Stability Constants of Metal Ion Complexes," The Chemical Society, London, 1964.

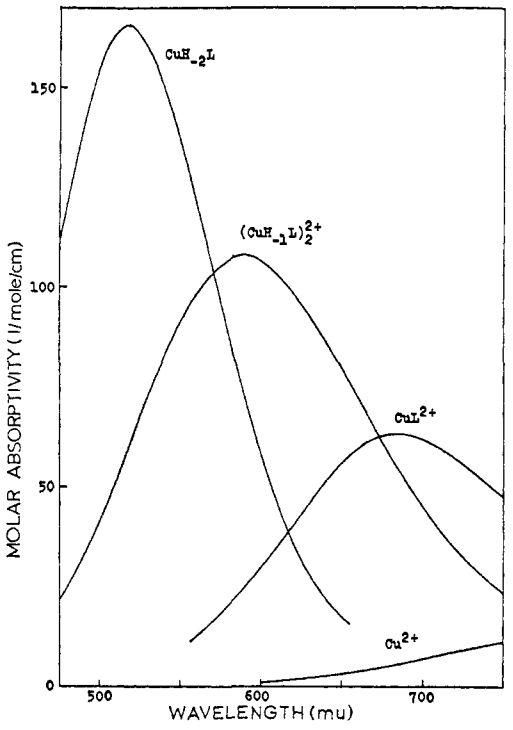


Figure 7. Molar absorbances of the various copper(II)–DGEN complexes in the visible region.

species $CuLH^{3+}$ under the conditions employed in this investigation.

The relative positions of the titration curves in this region show that depolymerization is taking place in going from $a = 3$ to $a = 4$. If the complex species at $a = 3$ were the monomer $CuH_{-1}L^{+}$ which becomes $CuH_{-2}L$ at $a = 4$, the two titration curves would coincide at the midpoint ($a = 3.5$) of this region. If polymerization occurred by olation in this region, the relative positions of curves 3 and 4 in the buffer region from $a = 3$ to $a = 4$ would be reversed. This system appears to be the first example of depolymerization of copper(II) complexes with increasing pH. Since olation usually increases with pH, the existence of hydroxo bridges is not likely as the basis for polynuclear complex formation, strongly suggesting that the bridging groups are derived from the ligand alone.

The fact that no further reaction takes place between $CuH_{-2}L$ and hydroxide ion is in accord with the coordination properties of the copper(II) ion. All of the complex equilibria observed in the 1:2 copper(II)–DGEN systems (curves 5 and 6) are those observed in the 1:1 system. The fact that only the same reactions are observed in both the 1:1 and the 1:2 systems is also in accord with the coordination properties of both the copper(II) ion and the ligand.

Spectra. Using the various equilibrium constants in Table I, the concentrations of all species present in the solutions of which the spectra in Figure 2 were taken may be calculated. Then, with eq 13, the spectra

$$A = l(\epsilon_{Cu^{2+}}[Cu^{2+}] + \epsilon_{CuL^{2+}}[CuL^{2+}] + \epsilon_{(CuH_{-1}L)_2^{2+}}[(CuH_{-1}L)_2^{2+}] + \epsilon_{CuH_{-2}L}[CuH^{-}{}_{2}L]) \quad (13)$$

in Figure 2 can be resolved into the component spectra over the full wavelength range investigated. The molar absorbances of all species so obtained are shown in Figure 7 and in Table III.

Table III. Absorption Characteristics in the Visible Region o Various DGEN Complexes of Copper(II) and Nickel(II)

Ligand	Species	λ_{max}, mμ	ε, l./(mole cm)
...	Cu_{aq}^{2+}	~820	~14
DGEN, L	CuL^{2+}	685	60
DGEN, L	$(CuH_{-1}L)_2^{2+}$	590	108
DGEN, L	$CuH_{-2}L$	518	166
Triglycine, HL	$CuH_{-1}L$[a]	660	75
Triglycine, HL	$CuH_{-2}L^{-}$[a]	555	149
Tetraglycine, HL	$CuH_{-1}L$[a]	660	72
Tetraglycine, HL	$CuH_{-2}L^{-}$[a]	590	101
Tetraglycine, HL	$CuH_{-3}L^{2-}$	515	153
...	Ni_{aq}^{2+}	~395	~5
DGEN, L	NiL^{2+}	380	9
DGEN, L	NiL_2^{2+}	360	30
DGEN, L	$NiH_{-2}L$	414	219
Triglycine, HL	$NiH_{-2}L^{-}$[b]	430	240
Tetraglycine, HL	$NiH_{-3}L^{2-}$[b]	412	215

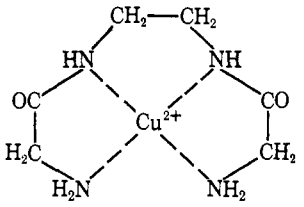
[a] From ref 12. [b] From ref 15.

Spectral data for the copper(II)–peptide and amide chelates are compared in Table III. The copper(II)–triglycine and –tetraglycine chelates in which only one peptide proton is displaced both exhibit maxima at 660 mμ.[12] On the other hand, the DGEN chelate CuL^{2+} absorbs at 685 mμ. The difference appears to arise largely from the fact that, in the former, two strongly coordinating nitrogen atoms form a five-membered ring with the metal ion while, in the latter, the amino nitrogens are far apart. Although there are three strongly coordinating nitrogen atoms in both $(CuH_{-1}L)_2^{2+}$ (L = DGEN) and $CuH_{-2}L^{-1}$ (HL = tetraglycine), direct correlation would not seem justified because of the wide differences in the structures of these chelates. From an electron spin resonance study of copper(II)–peptide complexes, Gould and Mason[26] suggested that the copper(II)–negative peptide nitrogen bond is essentially the same as the copper(II)–amino nitrogen bond although the two nitrogen atoms are not in the same chemical environment. On this basis, $CuH_{-2}L$ (L = DGEN) and $CuH_{-3}L^{2-}$ (HL = tetraglycine) would be expected to have almost the same absorption characteristics because the chelate ring structure is similar in both. This is strikingly borne out by the data in Table III. If the suggestions of Gould and Mason also apply to nickel(II) complexes, the DGEN and tetraglycine complexes of nickel(II) in which all peptide protons are displaced are expected to give similar spectra. Table III shows that the two complexes have almost identical absorption characteristics.

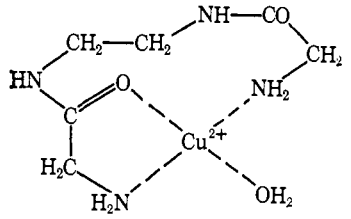
Coordinate Bonding. Three alternative arrangements of coordinate bonds are shown below for CuL^{2+}. Although Ia appears to be more favorable sterically, coordination through peptide oxygen (Ib) is more likely.[8-12,15] X-Ray studies[18] show that the neutral peptide group is coordinated to copper(II) through carbonyl oxygen in the solid copper(II)–triglycine chelate.

Molecular models show that if one peptide group is coordinated through oxygen, the other peptide group must remain uncoordinated (Ib). Arrangement Ic proposed by Zuberbuehler and Fallab[24] is unlikely from our results. It would exhibit absorption spectra

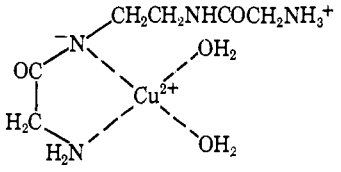
(26) D. C. Gould and H. S. Mason, *Biochemistry*, 6, 801 (1967).



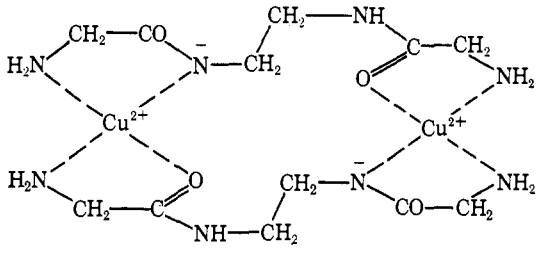
Ia



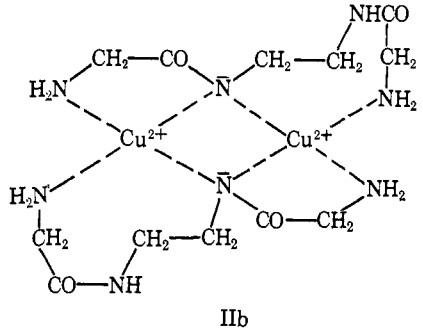
Ib



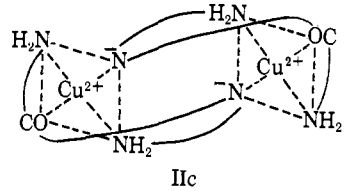
Ic



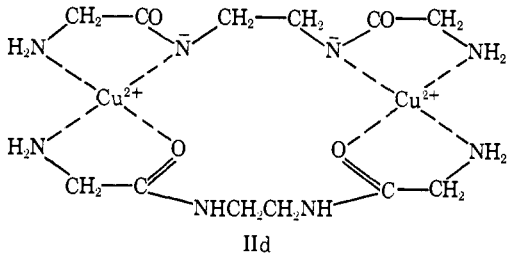
IIa



IIb



IIc

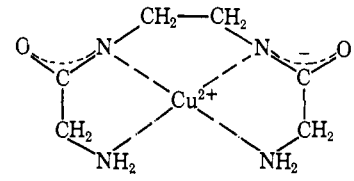


IId

almost identical with those of triglycine and tetraglycine chelates in which one proton is displaced from the peptide groups. Table III indicates that this is not so.

Illustrated are four possible structures (IIa–d) for $(CuH_{-1}L)_2^{2+}$. Formula IIb is preferred. If IIa, IIc, or IId (which corresponds to $Cu_2LH_{-2}L^{2+}$) is chosen, there seems to be no reason why all of the coordinating groups of one ligand should not coordinate to the same copper(II) ion to form the monomer. The negative peptide group probably coordinates to two copper(II) ions (IIb) as suggested for the polynuclear species of copper(II)–glycylglycine complex.[13] Such dimer is also known for the copper(II) complex of triglycine (in which both peptide hydrogens are displaced) in crystalline state[18] although the nature of bonding in this compound is somewhat different. There also remains the possibility that in $(CuH_{-1}L)_2^{2+}$ there might be some kind of interaction between the two copper(II) ions, for example, one similar to that observed by Zuberbuehler and Mason[27] in cystinylbisglycine–bis-copper(II) complex.

The arrangement of coordinate bonds indicated in III is considered the most likely for $CuH_{-2}L$.



III

Nickel(II)–DGEN. Potentiometric Data. The value of K_1 given in Table I agrees well with that reported by Chakraburtty, *et al.*[3] The formation constant of NiL^{2+} is higher than those for formation of ordinary complexes of glycine peptides[16] by roughly 1.5 log units. On the other hand, the second stepwise formation constant is about the same for DGEN as for glycine peptides.

Assuming, as in the case of copper(II)–DGEN, stepwise dissociation of two additional protons from the complex NiL^{2+}, Chakraburtty, *et al.*,[3] also calculated constants for these two steps. However, no evidence for the stepwise nature was found in the present work. Indeed, the flat buffer region between 2 and 4 equiv of base (Figure 3, curve 2) suggests that the two amide protons are dissociated simultaneously.

Equation 14 may be derived assuming stepwise dissociation of the two peptide protons in NiL^{2+} in the region from $a = 2$ to $a = 4$ for the 1:1 titration (Figure 3, curve 2).

$$(2T_M - Z)[H^+]^2/Z = ((Z - T_M)[H^+]/Z)K_{1A} + K_{1A}K_{1B} \quad (14)$$

where

$$Z = (T_M - T_L)\alpha'/\alpha + S$$

$$K_{1A} = [MH_{-1}L^+][H^+]/[ML^{2+}]$$

$$K_{1B} = [MH_{-2}L][H^+]/[MH_{-1}L^+]$$

If this model is correct, a plot of $(2T_M - Z)[H^+]^2/Z$ against $(Z - T_M)[H^+]/Z$ should yield a straight line

(27) A. Zuberbuehler and H. S. Mason in "Magnetic Resonance in Biological Systems," A. Ehrenberg, B. G. Malmstrom, and T. Vanngard, Ed., Pergamon Press, London, 1967.

whose slope is equal to K_{1A} and intercept equal to $K_{1A}K_{1B}$. A linear plot was indeed obtained, but the slope was negative. Thus, this model was concluded to be incorrect. It is interesting, however, to note that the intercept of this plot ($p(K_{1A}K_{1B}) = 16.05$) was identical with the value of K_{1AB} given in Table I. The product $K_{1A}K_{1B}$ is mathematically equal to K_{1AB}.

That the two peptide protons are displaced simultaneously (eq 9) is further demonstrated by the constancy of the values of K_{1AB} calculated from all points between $a = 2$ and $a = 4$ in the 1:1 titration curve and between $a = 1.5$ and $a = 3$ in the 1:2 titration curve. Peptide proton dissociation occurs at about the same pH in nickel(II) complexes of both DGEN and tetraglycine.[16]

As with $CuH_{-2}L$, $NiH_{-2}L$ does not react with hydroxide ion. This is in accord with the planar nature of this chelate. The observation from the 1:2 titration (Figure 3, curve 3) that only $NiH_{-2}L$ is formed from both NiL^{2+} and NiL_2^{2+} is also in accord with the square-planar nature of $NiH_{-2}L$.

Spectra. With the equilibrium constants given in Table I, the concentrations of the various species present in the spectral solutions may be calculated. With eq 15 the molar absorbances of all species may be obtained at all wavelengths.

$$A = l(\epsilon_{Ni^{2+}}[Ni^{2+}] + \epsilon_{NiL^{2+}}[NiL^{2+}] + \epsilon_{NiL_2^{2+}}[NiL_2^{2+}] + \epsilon_{NiH_{-2}L}[NiH_{-2}L]) \quad (15)$$

The band maxima so obtained are given in Table III. The first three entries for nickel(II) show the expected trend of decreasing wavelengths of band maxima with increasing coordination. But $NiH_{-2}L$ exhibits a reversal in this trend, which indicates that a different mode of electronic transition is occurring. This also suggests structural transition from the octahedral to the planar form. The $NiH_{-2}L$ exhibits absorption characteristics almost identical with those of $NiH_{-3}L^{2-}$ (L^- = tetraglycinate). This is expected from the suggestions of Gould and Mason[26] as discussed above.

Coordinate Bonding. Definite evidence for the structural transition to the diamagnetic planar form ($NiH_{-2}L$) is further provided by the spectra in Figure 5. The fact that the paramagnetic Ni(II) bands of NiL^{2+} in the 1:1 solution and of NiL_2^{2+} in the 1:2 solution exhibit the highest intensities, and that the bands disappear completely for $NiH_{-2}L$, is evidence for octahedral-planar transition in going from the former to the latter. The square-planar, diamagnetic complex of nickel(II) with ethylenebisbiguanide[28] also gives only one band in the visible–near-infrared region (300–1600 mμ), the band maximum lying at 478 mμ. Usually only one band is observed in square-planar complexes of nickel(II) in solution.[29]

The nickel(II) complexes of amides[17] or peptides[15,16,30,31] in which the amide or peptide protons are displaced are yellow[15,16] or orange,[17] diamagnetic,[17,30,31] and planar,[15,16,31] and in many complexes the dissociation of the peptide protons is slow[15,16] and takes place simultaneously.[16,30,31]

The structures of NiL^{2+} and NiL_2^{2+} are not known, but are most certainly octahedral. The arrangement of coordinate bonds in the square-planar diamagnetic $NiH_{-2}L$ is probably similar to that of the analogous Cu(II) chelate, illustrated by III.

(28) D. J. MacDonald, *Inorg. Chem.*, **6**, 2269 (1967).
(29) W. Manch and W. C. Fernelius, *J. Chem. Educ.*, **38**, 192 (1961).
(30) R. Mathur and R. B. Martin, *J. Phys. Chem.*, **69**, 668 (1965).
(31) M. K. Kim and A. E. Martell, *J. Am. Chem. Soc.*, **91**, 872 (1969).

Chelate Chemistry. VI. Solution Behavior of Tropolonates[1]

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Contribution No. 1493 from the Central Research Department, Experimental Station, E. I. du Pont de Nemours and Company, Wilmington, Delaware 19898. *Received January 23, 1969*

Abstract: The solution behavior of six-, seven-, eight-, and ten-coordinate metal derivatives of the α, β, and γ isomers of isopropyltropolone has been investigated with respect to association, dissociation, ligand lability, and intra- and intermolecular exchange of polytopal[2] and stereoisomeric forms, and compared, where possible, with analogous derivatives of β-diketones. Qualitatively, the lability of the isopropyltropolonate ligand is comparable to that of the β-diketonates. The integrity of the previously postulated ten-coordinated pentakis(isopropyltropolono)thorium anion in solution has been confirmed. Eight-coordinate complexes derived from thorium are associated in solution, but those from uranium(IV), zirconium(IV), and hafnium(IV) are not. Ligand lability in the tetrakis metal complexes follows the order Th > U > Hf ~ Zr > Ta. Attempts to definitively establish the presence of polytopal isomers or stereoisomers by low-temperature nmr studies of the seven-, eight-, and ten-coordinate complexes were unsuccessful. *cis*- and *trans*-stereoisomer interconversion for six-coordinate complexes was examined.

The facility with which the tropolone ion or derivatives thereof form high-coordinate complexes with a variety of metal ions was emphasized in earlier papers.[3-7] Specifically, seven-, eight-, nine-, and ten-

(1) E. L. Muetterties and C. M. Wright, *J. Am. Chem. Soc.*, **86**, 5132 (1964).
(2) E. L. Muetterties, *ibid.*, **91**, 1636 (1969).
(3) E. L. Muetterties and C. M. Wright, *ibid.*, **87**, 21 (1965).
(4) E. L. Muetterties and C. M. Wright, *ibid.*, **87**, 4706 (1965).
(5) E. L. Muetterties, *ibid.*, **88**, 305 (1966).