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Extrathermodynamic Relationships in Copper(II) Chelate Promoted Hydrolysis of Methyl Glycinate

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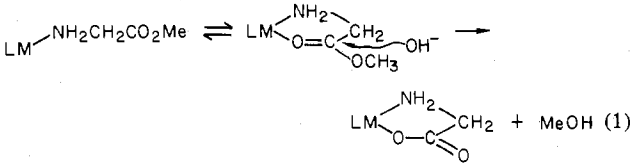
The second-order rate constants, rate = $k_{OH}[Cu(L)(MeGly)^{x+}][OH^-]$, for the hydrolysis of a series of methyl glycinate complexes of copper(II) chelates were determined as were their activation parameters. Tridentate and tetradentate copper(II) chelates fall on different isokinetic lines suggesting differing mechanisms or rate profiles. Both the strength of the donor atoms of the auxiliary ligands and the charge of the metal complex are shown to be important in determining the catalytic activity of the various copper(II) complexes.

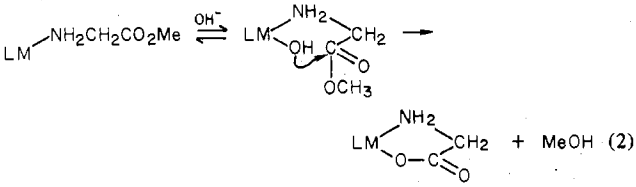
Introduction

Metal-ion-promoted hydrolysis of amino acid esters has been studied by a number of research groups.[1-7] Three mechanisms have been proposed. One involves an initial rapidly established equilibrium in which ester group coordination occurs, followed by rate-determining OH^- attack (eq 1). The second involves



(1)

rapid equilibrium formation of an M–OH complex, followed by intramolecular OH^- attack (eq 2). The third involves only



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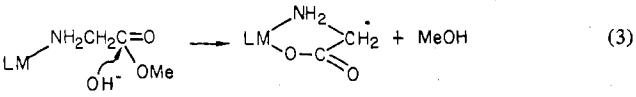
OH^- attack on the coordinated ester (eq 3). In labile metal



(3)

complex promoted reactions of amino acid esters, it has been very difficult to establish whether one or a combination of the above mechanisms is involved in the observed hydrolyses. It is for this reason that we have initiated a series of studies into extrathermodynamic relationships associated with metal chelate promoted hydrolysis of amino acid esters. Recently, we have determined[7,8] two isokinetic temperatures for three differing series of metal chelates. In this study a variety of tridentate and tetradentate Cu(II) chelates were studied in hopes of gaining some insight into the mechanism(s) involved and the auxiliary ligand features which possibly dictate the mechanism of hydrolysis.

Experimental Section

Reagents. Baker Analyzed reagent grade $Cu(NO_3)_2 \cdot 3H_2O$ was used for all metal solutions, which were standardized via standard ion-exchange techniques. Aliquots of the metal ion solution were passed through Dowex 50W-X8 strongly acidic cation-exchange resin, and the effluent solutions were titrated with standard NaOH solutions using phenolphthalein as an indicator.

The trihydrochloride salt of 2,2′,2″-triaminotriethylamine (tren·3HCl) was purchased from Strem Chemicals, Inc., and the hydrochloride salt of methyl glycinate (MeGly·HCl) and glycine (Gly) were obtained from Sigma Chemical Co. Nitrilotriacetic acid (NTA) was purchased from ICN Pharmaceuticals, and terpyridine (terpy) was purchased from K and K Laboratories. The above compounds were of the highest purity available and were used without further purification. Solutions of Cu^{II}terpy were prepared by addition of weighed amounts of ligand, that had been dried overnight in a desiccator (H_2SO_4) at 50 °C, to metal ion solutions.

The trihydrochloride salt of diethylenetriamine (dien·3HCl) was prepared by the addition of HCl to 50/50 C_2H_5OH/H_2O solutions of dien, which had been vacuum distilled according to Perrin.[9] Second

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and third crops of the crystals were obtained upon successive additions of 95% ethanol. The hydrochloride salt was recrystallized three times and dried. The trihydrobromide salt of *N,N*-bis(2-aminoethyl)glycine (DTMA·3HBr) was prepared according to Collman[10] as modified by Martell et al.[11]

Bis(2,2′-aminomethylpridyl)amine (DPA) was prepared using a modified procedure of Romary et al.[12] and Nakon et al.[13] 2-Chloromethylpyridine (10 g) was dissolved in 20 mL of H_2O and neutralized with a saturated solution of K_2CO_3. The free base, which separated from the aqueous layer, was added to a 25-mL ethanolic solution of 2-aminomethylpyridine (15 g). The mixture was then kept at 40–45 °C for 1 h; the solvent was removed leaving an oil, which was redissolved in an aqueous alkaline solution and then extracted with ether. After removal of the ether, the resultant oil was vacuum distilled at 150 °C (1 mm). The hydrochloride salt was then prepared by the addition of concentrated HCl to an ethanolic solution of the distilled DPA. Small amounts of acetone were added to induce precipitation. DPA·3HCl was recrystallized from ethanol by addition of acetone. *N*-Methylpyridylaspartic acid (PAA) was a gift from Professor R. J. Angelici.

Solutions of dien·3HCl, MeGly·HCl, tren·3HCl, and DPA·3HCl were standardized by standard ion-exchange techniques using Dowex 50W-X8 strongly acidic cation-exchange resin. Solutions of glycine, DTMA·3HBr, and NTA were standardized via potentiometric titration using glass and calomel electrodes.

Potentiometric Measurements. A Corning Digital 112 Research Model pH meter was used to determine hydrogen ion concentration for potentiometric titrations. The concentration of all solutions was approximately 2.5×10^{-3} M in Cu^{2+}. All titrations were performed in a double-walled titration cell of 50-mL capacity. The temperatures of all solutions were maintained at ±0.05 °C by circulating thermostated water through the outer jacket of the cell. The cell was fitted with glass and calomel extension electrodes, a microburet delivery tube, and a nitrogen inlet tube. Ionic strengths of all solutions were maintained at 0.10 M by the addition of an appropriate amount of 1.0 M KNO_3. The solutions were stirred with a magnetic stirrer, and all titrations were repeated with most done in triplicate.

The glass electrode was calibrated to read directly $-\log [H^+]$ according to the method of Rajan and Martell[14] using standard HCl and NaOH solutions. In titrations involving MeGly·HCl, the experiments were performed as quickly as possible due to apparent hydrolysis of the ester. Only points from 20 to 50% complete titration, where ester hydrolysis was minimal, were used in calculating formation constants.

Hydroxo (eq 4) and mixed-ligand (eq 5) formation constants were

$$CuL^{x+} + OH^- \rightleftharpoons CuLOH^{(x-1)+} \quad (4)$$

$$CuL^{x+} + X^{0,-1} \rightleftharpoons CuLX^{x,(x-1)+} \quad (5)$$

determined via Bjerrum's method.[15] The formation constant for Cu(PAA) was determined by Schwarzenbach's competing ligand method[16] using $trenH_3^{3+}$ in the reaction

$$Cu(PAA) + trenH_3^{3+} \rightleftharpoons Cu(tren)^{2+} + PAA + 3H^+ \quad (6)$$

in which $trenH_3^{3+}$ and PAA represent all possible protonated forms of those ligands. Equimolar amounts of Cu^{2+}, PAA, and tren·3HCl with enough 1.0 M KNO_3 added to yield an ionic strength of 0.10 M were entered into 2-oz bottles. After being purged with N_2, enough base was added to reach the second buffer zone where replacement of PAA by tren occurred. This region was ascertained by a preliminary titration. The bottles were sealed with Parafilm M (Marathon Products, Neenah, Wis.) and kept in a water bath at 25 °C for 3 h after which time equilibrium had been reached. A Corning semimicro combination pH electrode was used to determine the pH of these solutions.

Species included in the calculation of K_1 were Cu(PAA), $Cu(tren)^{2+}$, $Cu(PAA)OH^-$, $Cu(tren)OH^+$, and all protonated and unprotonated forms of PAA and tren. Schwarzenbach's stability constants[16] for $Cu(tren)^{2+}$ and tren were used in the calculations, which were performed with the aid of an IBM 360-65 digital computer.

Kinetic Measurements. Rates of MeGly hydrolysis in the presence of CuL^{x+} were determined by pH Stat techniques described previously.[17] Ten-milliliter solutions containing 9×10^{-3} M CuL^{x+}, 8.0×10^{-4} M MeGly, and enough KNO_3 to give a 0.10 M ionic strength were studied over a large pH range at 20, 25, 30, and 35 °C. A 5–10% excess of ligand (L) over Cu was used to ensure coordination of all Cu^{2+}, which is itself an excellent catalyst. A 20% excess of ligand gave the same rate as with the 10% excess, indicating that free Cu^{2+} was not involved in the observed catalysis.

After CuL^{x+} solutions were equilibrated at the desired temperature, a solution of MeGly·HCl was added, and the pH was brought up to the desired value by the addition of 0.02 M NaOH. The hydrolysis was then followed by automatic addition of 0.02 M NaOH. Pseudo-first-order rate constants, k_{obsd}, were obtained from slopes of plots of log (% end − % t) vs. time, where % end is the percent of the total syringe volume delivered at the end of the reaction and % t is the percent delivered at any time t.

Results

Potentiometric Measurements. Monohydroxo complex formation constants (eq 4) of Cu(PAA) and $Cu(dien)^{2+}$ were determined to be $10^{4.34}$ and $10^{4.67}$, respectively, at 25.0 °C and an ionic strength of 0.10 M (KNO_3).

Values of K_x (eq 5) for both glycine and methyl glycinate complexation to Cu(PAA) and $Cu(DTMA)^+$ at 25.0 °C and an ionic strength of 0.10 M (KNO_3) are listed in Table I. The mixed-ligand formation constants (eq 5) indicate that under the experimental conditions used for the kinetic studies, greater than 90% of the amino acid ester is coordinated to the Cu(II) chelate as $Cu(L)(MeGly)^{x+}$. After hydrolysis, the resultant glycine remained bound to the Cu(II) chelate as indicated by the formation constants.

Titration curves of 1:1 PAA to Cu^{2+} consisted of a low-pH buffer zone terminated by an inflection at $a = 2$. Visible spectra of these solutions did not change up to $a = 2$, indicating that Cu(PAA) formation was complete even at the lowest pH; therefore, the formation constant could not be determined by standard potentiometric techniques. The value of log K (eq 7), determined by Schwarzenbach's competing ligand method,[16] was 13.8 ± 0.2 at 25 °C and 0.10 M (KNO_3).

$$Cu^{2+} + PAA^{2-} \rightleftarrows Cu(PAA) \quad (7)$$

Kinetic Measurements. The pseudo-first-order rate constants for methyl glycinate hydrolysis in the presence of $Cu(DTMA)^+$ and Cu(PAA) at various pH values at 25.0 °C are given in Table II, where L is DTMA or PAA. As noted above, the equilibrium constants (K_x for methyl glycinate) indicate that at least 90% of the methyl glycinate is coordinated as $M(L)(MeGly)^{x+}$ and that after hydrolysis the glycine remains complexed as $M(L)(Gly)^{(x-1)+}$. Therefore, the predominant reaction occurring in the hydrolysis studies can be written as

$$M(L)(MeGly)^{x+} + OH^- \rightarrow M(L)(Gly)^{(x-1)+} + CH_3OH \quad (8)$$

$$M(L)(MeGly)^{x+} + OH^- \rightarrow M(L)(Gly)^{(x-1)+} + CH_3OH \quad (9)$$

The k_{obsd} values determined over a range of 1 pH unit indicate that reaction 9 follows the rate law

$$\text{rate} = k_{OH}[M(L)(MeGly)^{x+}][OH^-] \quad (10)$$

where $k_{OH} = k_{obsd}/[OH^-]$. The average values of k_{OH} at 25.0 °C and 0.10 M (KNO_3) ionic strength are listed in Table I. In the pH range studied, the hydrolysis of methyl glycinate is negligible in the absence of ML^{x+}.

The pseudo-first-order rate constants for methyl glycinate hydrolysis in the presence of $Cu(tren)^{2+}$, $Cu(terpy)^{2+}$, $Cu(DPA)^{2+}$, and $Cu(dien)^{2+}$ at 25.0 °C and 0.10 M (KNO_3) ionic strength are listed in Table II. The hydrolysis of $Cu(DPA)(MeGly)^{2+}$ was shown by Nakon et al.[13] to follow the rate shown in eq 10. Wood et al.[18] have shown that the $Cu(terpy)(MeGly)^{2+}$ and $Cu(tren)(MeGly)^{2+}$ hydrolyses also follow the rate law shown in eq 10. The rate law for $Cu(dien)(MeGly)^{2+}$ hydrolysis, however, has been shown to be more complicated (eq 11).[19] The two rate constants k_{OH} and

$$\text{rate} = k_{OH}[Cu(dien)(MeGly)^{2+}][OH^-] + k_{MOH}[Cu(dien)(MeGly)^{2+}][Cu(dien)OH^+] \quad (11)$$

k_{MOH} were determined graphically by plotting the quantities

$$\frac{k_{obsd}K_x}{[OH^-]} + \frac{k_{obsd}(K_b[H^+] + 1)(K_{OH}[OH^-] + 1)}{[Cu(dien)_{tot}][OH^-]}$$

vs.

$$\frac{[Cu(dien)_{tot}]K_xK_{OH}[OH^-]}{[OH^-](K_{OH}[OH^-] + 1)}$$

where K_b is the protonation constant of methyl glycinate, K_{OH} is the formation constant for the hydroxo complex, and K_x is the mixed-ligand formation constant of methyl glycinate with $Cu(dien)^{2+}$. The rate constant for the reaction involving the hydroxo complex, k_{MOH}, is determined from the slope, and k_{OH}, the rate constant for the hydroxide attack on the bound ester, is obtained from the intercept.

The values obtained in this study are not in good agreement with those obtained by Angelici and Allison.[19] However, those authors[19] used a value of 5.12 for log K_{OH}, whereas the value determined in the present study was 4.67. It is also important to note that in the present study, the ionic strength was maintained at 0.10 M, whereas in the previous study a 0.05 M ionic strength was used. Furthermore, different methods were employed in the calibration of the glass electrodes. Any or all of these factors may account for the discrepancies between the two sets of rate constants.

Activation Parameters. The temperature dependencies of various metal chelate promoted hydrolyses of methyl glycinate were determined in the range of 20–35 °C. The pseudo-first-order rate constants at 20, 30, and 35 °C for $Cu(DPA)^{2+}$, $Cu(dien)^{2+}$, $Cu(terpy)^{2+}$, $Cu(tren)^{2+}$, $Cu(DTMA)^+$, and Cu(PAA) are given in Table III. In order to determine the temperature dependence of $Cu(dien)(MeGly)^+$ hydrolysis, it was necessary to determine the protonation-constant value of methyl glycinate, the K_x value for $Cu(dien)(MeGly)^{2+}$ formation, and the K_{OH} value for $Cu(dien)OH^+$ formation at 20, 25, 30, and 35 °C. These are listed in Table IV, as are the k_{OH} values. The activation parameters for the hydrolysis of methyl glycinate in the presence of metal chelates are given in Table I.

Discussion

Cu(II) Chelate Promoted Hydrolysis of MeGly. The general process for metal chelate promoted hydrolysis of methyl glycinate proceeds as in eq 12. In these studies, the methyl

$$ML^{x+} + MeGly \overset{K_x}{\rightleftharpoons} ML(MeGly)^{x+} \xrightarrow[OH^-]{k_{OH}} ML(Gly)^{(x-1)+} + MeOH \quad (12)$$

glycinate is almost entirely bound to ML^{x+} as $ML(MeGly)^{x+}$. Thus, the observed rate law represents the second step, k_{OH}, only. The first-order dependence of hydroxide concentration may be accounted for by the general mechanisms explained in the introduction (eq 1, 2, 3). Mechanism 2 (eq 2) appears unlikely in that all four equatorial Cu(II) sites in these metal–complex systems are occupied by ligand donor groups precluding hydroxo complex formation in the pH ranges studied. Mechanism 3 (eq 3) also appears unlikely since many of the Cu(II) chelates are much more effective at promoting MeGly hydrolysis (Table I) than is a proton. Therefore, it appears that Cu(II) chelate promoted hydrolyses of MeGly proceed by mechanism 1 (eq 1).

Before attempting to establish extrathermodynamic relationships, it is necessary to show that all the compounds under study fall on the same isokinetic line; i.e., for a series of similar reactions involving a common mechanism, parallel changes in ΔH^* and ΔS^* are commonly observed.[20] In Figure 1, a plot of ΔH^* vs. ΔS^* for a series of Cu(II) chelate, Ni(II) chelate, and metal nitrilotriacetate promoted hydrolyses of MeGly as well as metal–ethyl glycinate-*N*,*N*-diacetic acid hydrolyses indicates that the hydrolyses proceed by two different mechanisms or rate profiles.[20] The Ni(II) tetradentate complexes, Cu(II) tetradentate complexes, and metal nitrilotriacetate complexes fall on one line, while the Cu(II) tridentate and metal–ethyl glycinate-*N*,*N*-diacetic acid (M(EGDA)) complexes constitute a second one. Since, as discussed above, the Cu(II) chelate promoted hydrolyses of MeGly appear to proceed by mechanism 1, their appearance on different isokinetic lines seems to indicate two differing rate profiles.

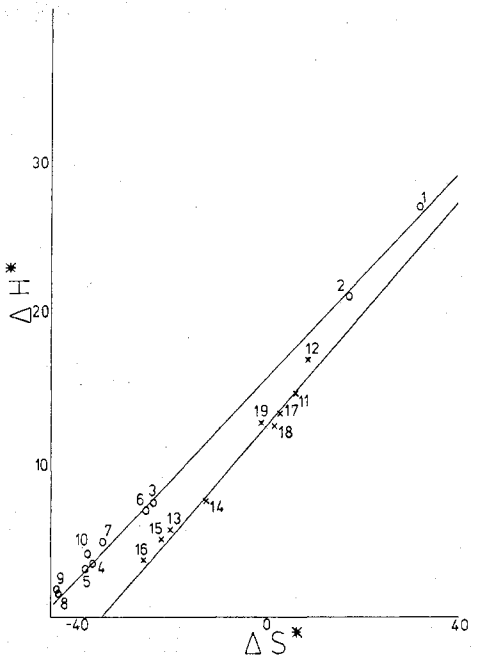

Figure 1. Plot of ΔH^* vs. ΔS^* for nickel(II) and copper(II) chelate and metal nitrilotriacetate promoted hydrolyses of methyl glycinate and hydrolysis of metal–ethyl glycinate-*N*,*N*-diacetic acid: (1) $Cu(tren)^{2+}$, (2) $Cu(DTMA)^+$, (3) Cu(PAA), (4) $Cu(NTA)^-$, (5) $Ni(tren)^{2+}$, (6) $Ni(trien)^{2+}$, (7) Ni(EDDA), (8) $Ni(NTA)^-$, (9) $Co(NTA)^-$, (10) $Zn(NTA)^-$, (11) $Cu(DPA)^{2+}$, (12) $Cu(dien)^{2+}$, (13) $Cu(terpy)^{2+}$, (14) Cu(IMDA), (15) Pb(EGDA), (16) Cu(EGDA), (17) Zn(EGDA), (18) Co(EGDA), (19) Ni(EGDA).

Angelici and Leach[17,21,22] first gathered data showing that hydrolyses of M(EGDA) involve an initial rapidly established equilibrium in which the carbonyl oxygen of the ester group coordinates to the metal, followed by external OH^- attachment (eq 1). Bedell and Nakon[8] added further evidence by showing that the log of the rate constant for M(EGDA) hydrolysis exhibits a direct correlation with that for the rate constant for water exchange (eq 13). The trend of M(EGDA) hydrolysis

$$M(H_2O)_x^{2+} + H_2O^* \rightleftarrows M(H_2O)_{x-1}(H_2O^*)^{2+} + H_2O \quad (13)$$

is Pb(II) > Cu(II) > Zn(II) > Co(II) > Ni(II), paralleling that for the rate of H_2O exchange, indicating that eq 14 is

$$M(EGDA)(H_2O)_x \rightleftarrows M(EGDA)(H_2O)_{x-1} + H_2O \quad (14)$$

responsible for most of ΔG^* (Figure 2, mechanism 1 (eq 1)).

Since the activation parameters for methyl glycinate hydrolyses in the presence of tridentate Cu(II) complexes fall on the same isokinetic line as M(EGDA), it would appear that the same rate profile is involved.[23] The tetradentate Cu(II) complexes, however, appear to promote methyl glycinate hydrolysis via a different rate profile than do the corresponding complexes containing tridentate auxiliary ligands. The trend for MeGly promoted hydrolysis by $M(NTA)^-$, Cu > Ni > Zn > Co, is different than that for M(EGDA) chelates further suggesting differing rate profiles. The proposed rate profile

Table I. Rates and Activation Parameters for the Hydrolyses of Methyl Glycinate in a Series of Copper(II) and Nickel(II) Chelates and M(NTA)$^-$ and Metal-Ethyl Glycinate Diacetic Acid Complexes

ML	log K_L	log K_x (MeGly)	log K_x (glycine)	log k_{OH}	ΔH^*, kcal/mol	ΔS^*, cal/(mol K)
Cu(tren)$^{2+}$	18.8[a]	2.40[a]	2.89[a]	0.05	27.0 ± 0.8	32 ± 1
Cu(DTMA)$^+$	18.07[b]	2.48	3.24	0.83	21.4 ± 1.0	17 ± 2
Cu(PAA)	13.8	2.79	4.57	2.25	7.1 ± 0.9	−25 ± 1
Cu(NTA)$^-$	13.05[c]	2.88[d,e]	5.46[d,f]	2.50	3.4[g]	−38[g]
Cu(DPA)$^{2+}$	14.4[h]	2.81[h]	2.99[h]	2.06	14.6 ± 0.8	5 ± 1
Cu(dien)$^{2+}$	15.91[i]	2.52[i]	2.42[i]	1.39	17.3 ± 1.3	8 ± 3
Cu(terpy)$^{2+}$	13.4[j]	3.21[j]	4.34[j]	2.28	5.5 ± 0.7	−22 ± 1
Cu(IMDA)	10.63[k]	3.69[l]	6.42[l]	4.50[m]	7.7[m]	14[m]
Ni(EDDA)	13.5[n]	1.54[g]	4.16[g]	1.61[g]	4.8[g]	−35[g]
Ni(tren)$^{2+}$	14.0[a]	1.48[g]	4.32[g]	1.83[g]	7.2[g]	−26[g]
Ni(trien)$^{2+}$	13.5[o]	1.42[g]	4.45[g]	1.73[g]	4.8[g]	−35[g]
Ni(NTA)$^-$	11.47[c]	2.03[d,e]	4.95[d,f]	1.72[g]	0.9[g]	−47[g]
Co(NTA)$^-$	10.81[c]	1.88[d,e]	3.65[d,f]	1.27[g]	1.5[g]	−48[g]
Zn(NTA)$^-$	10.44[p]	1.58[d,e]	3.64[d,f]	1.54[g]	4.0[g]	−38[g]
Pb(EGDA)				4.45[q]	4.4[q]	−23[q]
Cu(EGDA)				4.14[r]	3.7[r]	−27[r]
Zn(EGDA)				3.60[r]	13.3[r]	+2[r]
Co(EGDA)				2.78[r]	12.9[r]	+0.7[r]
Ni(EGDA)				2.71[r]	13.8[r]	−2[r]

[a] Reference 16. [b] Reference 11. [c] J. Stary, *Anal. Chim. Acta*, 28, 132 (1963). [d] D. Hopgood and R. J. Angelici, *J. Am. Chem. Soc.*, 90, 2508 (1968). [e] I = 0.073 M (KNO_3). [f] I = 0.077 M (KNO_3). [g] Reference 7. [h] Reference 13. [i] J. W. Allison and R. J. Angelici, *Inorg. Chem.*, 10, 2233 (1971). [j] Reference 18. [k] G. Anderegg, *Helv. Chim. Acta*, 47, 1801 (1964). [l] B. E. Leach and R. J. Angelici, *Inorg. Chem.*, 8, 907 (1969). [m] R. Wood and R. Nakon, unpublished data. [n] S. Chaberek and A. E. Martell, *J. Am. Chem. Soc.*, 74, 6228 (1952). [o] G. Schwarzenbach, *Helv. Chim. Acta*, 33, 974 (1950). [p] H. M. N. H. Irving and M. G. Miles, *J. Chem. Soc. A*, 727 (1966). [q] Reference 17. [r] Reference 8.

Table II. Rate Constants for Methyl Glycinate Hydrolyses in the Presence of Copper(II) Chelates at 25 °C and 0.10 M Ionic Strength

pH	$10^4 k_{obsd}$, s^{-1}	pH	$10^4 k_{obsd}$, s^{-1}
Cu(PAA)^0		Cu(DTMA)$^+$	
7.90	2.13	9.45	2.95
8.00	2.66	9.55	3.89
8.10	3.43	9.65	4.46
8.20	3.94	9.70	5.13
8.30	5.54	9.75	6.02
8.40	7.67	9.85	7.42
8.50	9.01	9.95	10.25
8.60	10.60	10.05	11.30
8.65	11.75	10.15	14.56
8.75	12.45	10.25	19.24
Cu(tren)$^{2+}$		Cu(DPA)$^{2+}$	
9.75	1.10	8.25	3.65
10.00	1.78	8.50	5.70
10.25	3.13	8.75	10.35
Cu(terpy)$^{2+}$		Cu(dien)$^{2+}$	
7.75	1.68	8.00	23.44
8.00	3.27	8.25	25.12
8.25	5.44	8.50	26.92

Table III. Rate Constants for Methyl Glycinate Hydrolysis in the Presence of Copper(II) Chelates at 20, 30, and 35 °C at I = 0.10 M

pH	T, °C	$10^4 k_{obsd}$, s^{-1}	pH	T, °C	$10^4 k_{obsd}$, s^{-1}
Cu(tren)$^{2+}$			Cu(PAA)^0		
10.50	20	1.90	8.60	20	6.38
10.70	20	3.10	8.75	20	9.02
11.00	20	5.70	9.00	20	16.03
9.35	30	1.28	7.75	30	2.86
9.50	30	1.72	8.00	30	5.25
9.75	30	2.70	8.25	30	10.17
9.00	35	2.15	7.50	35	2.78
9.25	35	2.55	7.75	35	5.07
9.50	35	6.26	8.00	35	8.81
Cu(DTMA)$^+$			Cu(DPA)$^{2+}$		
9.75	20	2.75	8.50	20	3.09
10.00	20	4.61	8.75	20	3.61
10.25	20	8.41	9.00	20	8.76
9.25	30	5.01	8.00	30	4.55
9.50	30	7.99	8.25	30	5.70
9.75	30	14.90	8.50	30	12.30
8.90	35	5.13	7.50	35	3.07
9.15	35	9.03	7.75	35	4.44
9.40	35	15.48	8.00	35	9.66
Cu(terpy)$^{2+}$			Cu(dien)$^{2+}$		
8.00	20	1.78	8.50	20	33.88
8.25	20	3.13	8.75	20	36.30
8.50	20	5.56	9.00	20	38.90
7.50	30	1.62	8.00	30	21.38
7.75	30	3.02	8.25	30	22.90
8.00	30	5.16	8.50	30	24.54
7.25	35	1.58	8.00	35	17.38
7.50	35	2.92	8.25	35	18.62
7.75	35	4.96	8.50	35	19.95

for M(NTA)$^-$ and tetradentate Cu(II) chelates is shown in Figure 3 where OH^- attack and not H_2O loss is responsible for most of ΔG^* (eq 1). This explanation also appears to be in agreement with the observed isokinetic temperatures which indicate that M(EGDA) hydrolyses are enthalpy dependent (β = 392 K, electronic effects are dominant) whereas those promoted by M(NTA)$^-$ are entropy dependent (β = 274 K, solvent effects are dominant).

Correlation of Activation Parameters with Rates of Ester Hydrolysis. In Figure 4, a plot of the log of the rate constant, log k_{OH}, vs. the log of the formation constant of the Cu(II) chelates, log K_L, shows a correlation in which complexes with the highest formation constants are the least effective promoters of ester hydrolysis. This correlation had first been observed by Angelici et al.,[13] who concluded that metal complexes of the highest Lewis acidity are the most effective promoters of ester hydrolysis. Ligands which bind most strongly to the metal increase the electron density on the metal resulting in a decrease in its Lewis acidity. Since catalysis is believed to be due to the polarization of the carbonyl group through direct interaction of the carbonyl oxygen of the ester with the metal ion, complexes of high Lewis acid character would be the most effective promoters of ester hydrolysis. One measure of the Lewis acidity of the various Cu(II) complexes is K_x,[13] the value for the formation constant of Cu(L)-

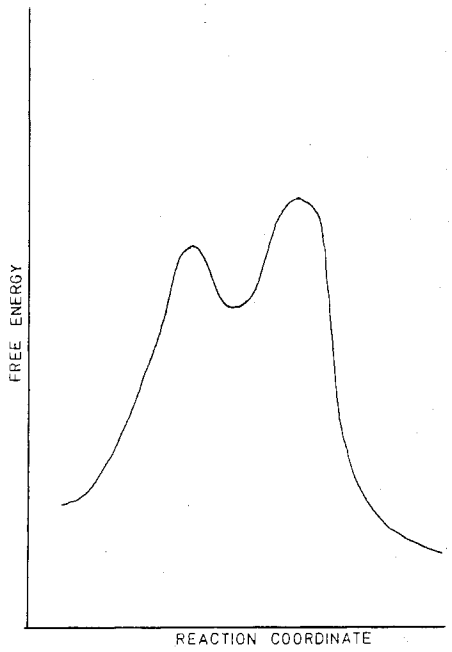

Figure 2. Proposed energy profile for the hydrolysis of methyl glycinate in the presence of tridentate copper(II) complexes and for the hydrolysis of M(EGDA).

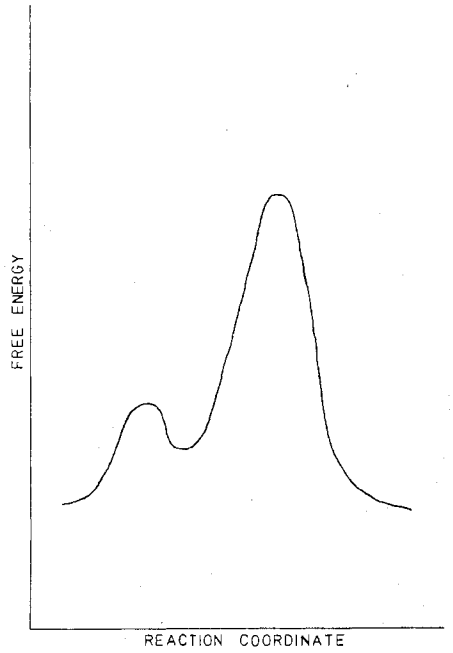

Figure 3. Proposed energy profile for the hydrolysis of methyl glycinate in the presence of tetradentate copper(II) complexes.

$(MeGly)^{x+}$. A plot (Figure 5) of log k_{OH} vs. log K_x shows a good correlation in which the most effective promoters of ester hydrolysis are those of highest Lewis acid character (large K_x values).

Although the correlation between log K_L and log k_{OH} holds for Cu(II) complexes, such a relationship was found not to exist with Ni(II) chelates. However, Newlin, Pellack, and Nakon[7] have shown that ΔH^* can be correlated to log K_L in the Ni(II) system and suggested that the ΔH^* values are to a large extent a measure of the ability of the metal chelate to polarize the ester carbonyl group. In Figure 6, there exists a correlation between ΔH^* for methyl glycinate hydrolysis in the presence of Cu(II) chelates and log K_L, the formation constant of the metal chelates. The tridentate and tetradentate Cu(II) complexes fall on different lines, yet the tetradentate

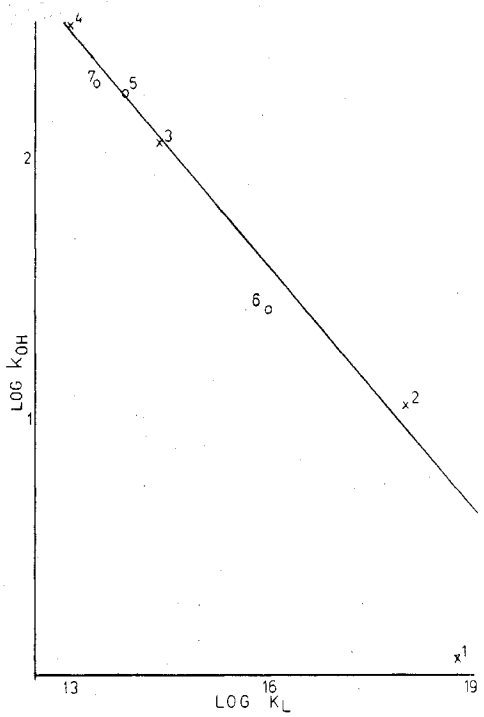

Figure 4. Correlation of log K_L with log k_{OH} for a series of Cu(L)(MeGly)$^{x+}$ complexes: (1) $Cu(tren)^{2+}$, (2) $Cu(DTMA)^{+}$, (3) Cu(PAA), (4) $Cu(NTA)^{-}$, (5) $Cu(DPA)^{2+}$, (6) $Cu(dien)^{2+}$, (7) $Cu(terpy)^{2+}$.

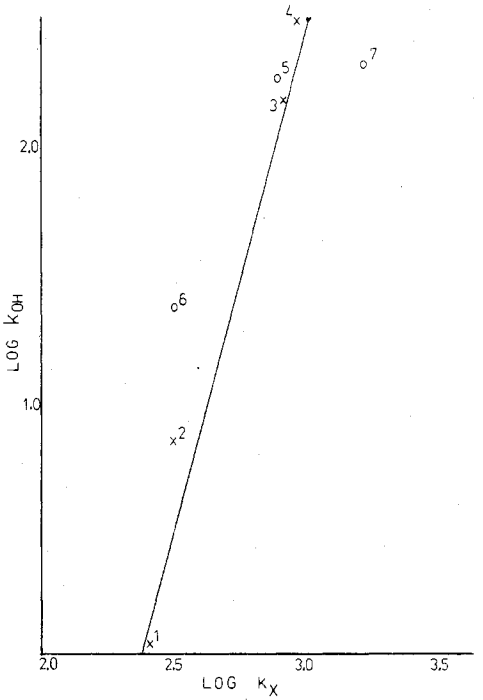

Figure 5. Correlation of log K_x with log k_{OH} for a series of Cu(L)(MeGly)$^{x+}$ complexes: (1) $Cu(tren)^{2+}$, (2) $Cu(DTMA)^{+}$, (3) Cu(PAA), (4) $Cu(NTA)^{-}$, (5) $Cu(DPA)^{2+}$, (6) $Cu(dien)^{2+}$, (7) $Cu(terpy)^{2+}$.

Cu(II) complexes fall on the same line as the tetradentate Ni(II) chelates. These data tend to substantiate the premise that different rate profiles exist for the tridentate and tetradentate Cu(II) complexes.

In Figure 7, it can be seen that ΔS^* is related to the charge of the complex. The most positive complexes have the least negative entropies of activation because of the cancellation of

Table IV. Protonation Constants of Methyl Glycinate (K_A), Formation Constants of Cu(dien)MeGly^{2+} (K_x) and Cu(dien)OH$^+$ (K_{OH}), and Rate Constants for Hydrolysis of Methyl Glycinate in the Presence of Cu(dien)$^{2+}$ at 20, 25, 30, and 35 °C

T, °C	log K_A	log K_x	log K_{OH}	log k_{OH}
20	7.73	2.65	4.76	1.20
25	7.63	2.56	4.67	1.39
30	7.53	2.47	4.58	1.59
35	7.43	2.38	4.49	1.77

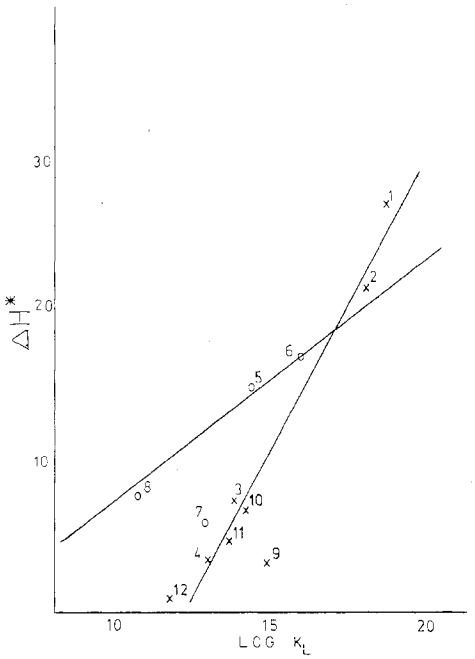

Figure 6. Plot of ΔH^* vs. log K_L for a series of Cu(L)(MeGly)$^{x+}$ and Ni(L)(MeGly)$^{x+}$ complexes: (1) Cu(tren)$^{2+}$, (2) Cu(DTMA)$^+$, (3) Cu(PAA), (4) Cu(NTA)$^-$, (5) Cu(DPA)$^{2+}$, (6) Cu(dien)$^{2+}$, (7) Cu(terpy)$^{2+}$, (8) Cu(IMDA), (9) Ni(tren)$^{2+}$, (10) Ni(trien)$^{2+}$, (11) Ni(EDDA), (12) Ni(NTA)$^-$.

charge in the activated complex, whereas neutral and negatively charged complexes have more negative entropies of activation because of a buildup of negative charge. Although ΔS^* values appear to be a good measure of the effects of total charge on the tetradentate complexes, the fact that the tridentate complexes, all with a +2 charge, do not all fall in the same place indicates that other factors are also involved in the ΔS^* term.

Although the same relationship between log K_L and ΔH^* and ΔS^* and the charge of the metal complex holds for both the Ni(II) and Cu(II) systems, the question remains as to why log k_{OH} can be correlated with log K_L in the Cu(II) system, but not in the Ni(II) system. An explanation is available in Table I. The Cu(II) complexes exhibit a wide range of ΔH^* values, while those for Ni(II) complexes vary over a relatively narrow range. Therefore, in the Cu(II) system, the ΔH^* term predominates (is responsible for the largest part of ΔG^*) and overwhelms the ΔS^* term. However, in the Ni(II) system, ΔS^* contributes as much or a greater amount (as compared to ΔH^*) to ΔG^*.

In summary, it is concluded that both tridentate and tetradentate Cu(II) chelate promoted hydrolyses probably proceed via coordination of methyl glycinate through both the amine nitrogen and ester carbonyl oxygen. However, the different isokinetic lines for the two systems are believed to reflect the differing importance of the H_2O loss and OH^- attack. Previous correlations of log K_L with ΔH^* and total charge of the metal complex with ΔS^* in the Ni(II) chelate

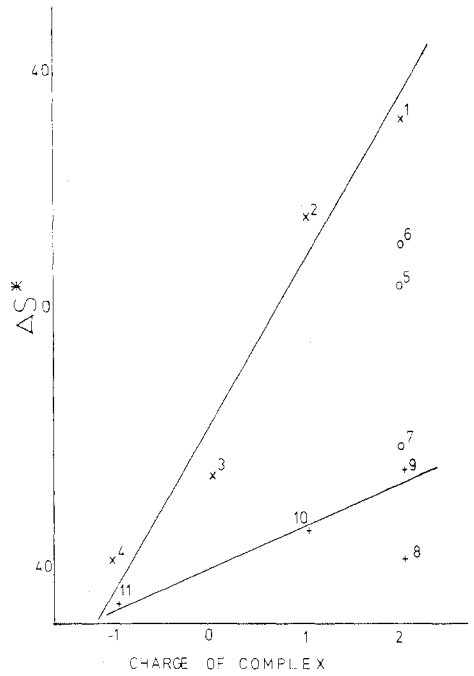

Figure 7. Plot of ΔS^* vs. the total charge of the complex for a series of Cu(L)(MeGly)$^{x+}$ and Ni(L)(MeGly)$^{x+}$ complexes: (1) Cu(tren)$^{2+}$, (2) Cu(DTMA)$^+$, (3) Cu(PAA), (4) Cu(NTA)$^-$, (5) Cu(DPA)$^{2+}$, (6) Cu(dien)$^{2+}$, (7) Cu(terpy)$^{2+}$, (8) Ni(tren)$^{2+}$, (9) Ni(trien)$^{2+}$, (10) Ni(EDDA), (11) Ni(NTA)$^-$.

systems were also found to be present in a series of Cu(II) complexes.

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Registry No. MeGly, 616-34-2; DPA, 1539-42-0; Cu(tren)$^{2+}$, 65879-01-8; Cu(DTMA)$^+$, 65879-00-7; Cu(PAA), 65956-53-8; Cu(NTA)$^-$, 15844-52-7; Cu(DPA)$^{2+}$, 65956-54-9; Cu(dien)$^{2+}$, 45520-77-2; Cu(terpy)$^{2+}$, 65956-95-8; Cu(IMDA), 14219-31-9.

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