Kinetics and Mechanism of Tetradentate Nickel(II) Chelate and Metal Nitrilotriacetate Promoted Hydrolysis of Methyl Glycinate: an Isokinetic Relationship

Donald E. Newlin, Margaret A. Pellack, and Robert Nakon*

Contribution from the Department of Chemistry, West Virginia University, Morgantown, West Virginia 26506. *Received June 8, 1976*

Abstract: The second-order rate constants, rate = $k_{OH}[M(NTA)MeGly^-][OH^-]$, for the hydrolysis of a series of methyl glycinate (MeGly) complexes of divalent metal nitrilotriacetates [$M(NTA)^-$] were determined as were their activation parameters. A proton was found to be as or more effective at promoting amino acid ester hydrolysis than the nitrilotriacetate complexes of Co(II), Ni(II), and Zn(II). These data suggest that the catalytic activity is due primarily to polarization of the carbonyl group of the ester by induction via coordination of the amino group rather than to direct metal ion interaction with the ester carbonyl group. The catalytic activity of a series of Ni(II) complexes containing tetradentate ligands toward methyl glycinate hydrolysis was determined, as were the activation parameters. Both the strength of the donor groups of the auxiliary ligands as well as the charge of the metal complex were found to be important in determining the catalytic activity of the metal complexes. The presence of strong donor groups, e.g., $-NH_2$, results in an increase in ΔH^*, indicating a reduced ability of the metal ion to polarize the carbonyl ester group. On the other hand, complexes with reduced positive charge resulted in more negative ΔS^* values, indicating a more unfavorable entropy of activation when charges are not canceled in the activated complex.

Since the discovery of metal ion catalyzed hydrolysis of amino acid esters in 1952,[1] these reactions have been studied by a number of research groups[2-6] with hopes of elucidating the role of metal ions in corresponding biological systems. Three mechanisms have been proposed, two involve external attack of hydroxide ion on the carbonyl carbon, which has been activated toward nucleophilic attack by the polarizing effect of the metal ion. This can occur via induction through a coordinated amino group or by direct interaction of the carbonyl oxygen atom with the metal. The third mechanism involves formation of a metal-hydroxo complex, followed by intramolecular attack by hydroxide ion. A series of tetradentate Ni(II) and metal nitrilotriacetate chelates have been studied to determine their catalytic activity toward the hydrolysis of methyl glycinate and the mechanism of reaction. Activation parameters were also determined in hopes of providing some basis for understanding the effect of auxiliary ligands on the catalytic activity of the metal ion.

Experimental Section

Reagents. Baker Analyzed Reagent Grade $Cu(NO_3)_2 \cdot 3H_2O$, $Zn(NO_3)_2 \cdot 6H_2O$, $Co(NO_3)_2 \cdot 6H_2O$, and $Ni(NO_3)_2 \cdot 6H_2O$ were used for all metal solutions, which were standardized via standard ion exchange techniques. Aliquots of the metal ion solutions were passed through Dowex 50W-X8 strongly acidic cation exchange resin, and the effluent solutions were titrated with standard NaOH solution using phenolphthalein as an indicator.

Nitrilotriacetic acid (NTA) and the trihydrochloride salt of 2,2′,2″-triaminotriethylamine (tren) were purchased from ICN Pharmaceuticals and Strem Chemicals, Inc., respectively. Ethylenediamine-*N*,*N*′-diacetic acid (EDDA) was purchased from K and K Laboratories, while glycine and the hydrochloride salt of methyl glycinate were purchased from Sigma Chemical Co. The above compounds were of the highest purity available and were used without further purification. Triethylenetetraamine (trien) (Fisher Scientific) and tetraethylenepentaamine (TEP) (Strem Chemical Co.) were vacuum distilled according to Perrin.[7] HCl was then bubbled into 50/50 ethanol/H_2O polyalkylenepolyamine solutions which were cooled and stirred; the hydrochloride salts then slowly crystallized. Second and third crops of the crystals were obtained upon successive additions of 95% ethanol. The hydrochloride salts were recrystallized three times and then dried. The barium salt of 1,5-diazacyclooctane-*N*,*N*′-diacetic acid (DACODA) was a gift from Professor J. I. Legg.

Solutions of trien·4HCl, tren·3HCl, TEP·5HCl, and MeGly·HCl were standardized via ion exchange techniques using Dowex 50W-X8 strongly acidic cation exchange resin. Solutions of EDDA, glycine, and NTA were standardized potentiometrically using glass and calomel electrodes. Solutions of DACODA were prepared by reacting Ba^{2+} salt of DACODA with H_2SO_4. The $BaSO_4$ was filtered off and the resultant solution was standardized potentiometrically.

Potentiometric Measurements. A Corning Digital 112 Research Model pH meter was used to determine hydrogen ion concentrations in all potentiometric titrations, which were carried out in a double walled cell of 50-ml capacity. The temperature of all solutions was maintained at 25.00 ± 0.05 °C by circulation of thermostated water through the outer jacket of the cell. The titration cell was fitted with Corning 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 performed in triplicate.

The glass extension electrodes were calibrated in terms of $-\log [H^+]$ according to the procedure of Rajan and Martell[8] using HCl and NaOH solutions. In titrations involving MeGly·HCl, the experiments were performed as rapidly as possible due to apparent hydrolysis of the ester. Only points from the region of 20–50% complete titration, where ester hydrolysis was minimal, were used in calculating metal complex formation constants.

Protonation constants of Gly, and MeGly, and the hydroxo formation constants were determined via Bjerrum's method,[9]

$$NiL^{x+} + OH^- \overset{K_{OH}}{\rightleftharpoons} NiLOH^{(x-1)+} \quad (1)$$

as were the mixed ligand formation constants, K_x, where X is Gly^- or MeGly.

$$NiL^{x+} + X^{0,-} \overset{K_x}{\rightleftharpoons} NiLX^{x,(x-1)+} \quad (2)$$

Kinetic Measurements. Rates of hydrolysis of MeGly in the presence of metal nitrilotriacetates and Ni(II) chelates were determined by pH stat (Radiometer TTT2/ABU11/SBR3) techniques described elsewhere.[10] Solutions (10 ml) containing 15:1 or higher ratios of metal chelate to ester and enough KNO_3 to give an ionic strength of 0.10 M were studied over a pH range of 1 log unit or more between 20 and 35 °C. The pH meter was calibrated in terms of H^+ concentration, i.e., pH is defined as $-\log [H^+]$ instead of $-\log \alpha_{H^+}$.[8]

In all cases the solutions were equilibrated at the desired temperature under a constant nitrogen flow. The ester solution was then added, and the pH brought up to the desired value by addition of 0.018

Table I. Formation Constants, K_x, for the Reaction of Metal Nitrilotriacetates and Tetradentate Nickel(II) Chelates with Glycine and Amino Acid Esters

Complex	Log K_f (glycine)	Log K_f (amino acid ester)
$CuNTA^-$	5.46[a,b]	2.88[a,d,e]
$CoNTA^-$	3.65[a,b]	1.88[a,d,e]
$2nNTA^-$	3.64[a,b]	1.58[a,d,e]
$NiNTA^-$	4.95[a,b]	2.03[a,d,e]
NiEDDA	4.16[c]	1.54[c,f]
$Ni(tren)^{2+}$	4.32[c]	1.48[c,f]
$Ni(trien)^{2+}$	4.45[c]	1.42[c,f]

[a] Reference 16. [b] I = 0.073 M (KNO_3). [c] I = 0.10 M (KNO_3). [d] I = 0.077 M (KNO_3). [e] Ethyl valinate. [f] Methyl glycinate.

M sodium hydroxide. The hydrolysis was then followed by automatic addition of NaOH. Pseudo-first-order rate constants were obtained by plotting log (%end − %t) vs. time, where %end is the percent delivered at the end point and %t is the percent delivered at any time, t. The concentration of the various metal chelates was 1.35×10^{-2} M, while that of MeGly was 9.0×10^{-4} M for studies at 20 and 25 °C. At 30 and 35 °C the concentrations of metal chelate were 1.80×10^{-2} and 2.25×10^{-2} M, respectively, while that of MeGly was 9.0×10^{-4} M. The higher ratios of metal chelate to ester at higher temperatures were used to ensure that most of the MeGly in solution was coordinated to the metal chelate, since it is known that the formation constant of the mixed ligand complex (eq 2) decreases with temperature. A 10% excess of chelating agent over metal ion was used to ensure coordination of all metal ion, which is itself an excellent catalyst. A 20% excess of chelating agent gave the same rates within experimental error as with the 10% excess, indicating metal ion was not involved in the observed catalysis. Ratios (30/1 or 45/1) of metal chelate to ester gave the same rates within experimental error as those with 15/1 ratios, indicating that under the conditions used, there is no rate dependence on metal chelate concentration.

Results

Equilibrium Constants. Protonation constants of Gly^- and MeGly were calculated as $10^{9.63}$ and $10^{7.65}$ M^{-1}, respectively, at 25.0 °C and 0.10 M (KNO_3) ionic strength, in excellent agreement with literature values.[11] Values of K_x for amino acid complexation to the various metal chelates are listed in Table I. The values of K_x (eq 2) for amino acid ester coordination to the various metal chelates were determined with the aid of a IBM 360-65 computer on data obtained in the range 20–80% titration. The K_{OH} (eq 1) values obtained for $Ni(tren)^{2+}$ and $Ni(trien)^{2+}$ were determined as $10^{1.86}$ and $10^{1.92}$ M^{-1}, respectively.

The K_x values (eq 2) for amino acid ester coordination to various metal chelates are listed in Table I. Only data from the 20–50% range of the titration were used in calculations. The data after 50% titration yielded numbers which increased in value with increasing a value, moles of base per mole of ester; this is probably attributable to ester hydrolysis. Due to the pH range at which ester coordination occurs, the formation of the hydroxo species can be ignored. The assumption was made that only a mono ester complex was formed under conditions of 1:1 mole ratios of metal chelate and ester. Attempts to include possible 2:1 metal complex to ester species did not yield consistent or reasonable formation constants.

Kinetics of MeGly Hydrolysis. Pseudo-first-order rate constants, k_{obsd} for the hydrolysis of MeGly in the presence of various metal chelates at 25 °C and I = 0.10 M are listed in Table II, while those at 20, 30, and 35 °C are listed in Table III. Under the conditions of these studies, the equilibrium constants indicate that at least 90% of the MeGly is coordinated as $MLMeGly^{x+}$, and that after hydrolysis the Gly^- product remained coordinated as $MLGly^{(x-1)+}$. Thus the predominant reaction occurring in the hydrolysis studies may be written as

Table II. Rate Constants for Methyl Glycinate Hydrolysis in the Presence of Metal Nitrilotriacetates and Tetradentate Nickel(II) Chelates at 25 °C and 0.1 M Ionic Strength

pH	$10^4 k_{obsd}$, s^{-1}	pH	$10^4 k_{obsd}$, s^{-1}
Ni(tren)MeGly^{2+}		Ni(NTA)MeGly$^-$	
8.25	1.96	8.90	6.58
8.40	2.62	9.00	8.68
8.50	3.36	9.10	10.4
8.60	4.03	9.25	14.1
8.75	6.21	9.40	20.9
8.90	8.59	9.50	28.0
9.00	11.4	9.60	35.2
9.10	14.5	9.75	50.6
9.25	20.0		
9.50	33.2	Co(NTA)MeGly$^-$	
		9.00	3.01
Ni(trien)MeGly^{2+}		9.10	3.96
8.50	2.72	9.25	5.72
8.60	3.51	9.40	8.00
8.75	4.66	9.50	9.79
8.90	7.09	9.60	12.8
9.00	8.16	9.75	15.7
9.10	10.9	9.90	19.6
9.25	16.1	10.00	22.1
9.40	21.0	10.10	34.1
9.50	26.1		
Ni(EDDA)MeGly		Zn(NTA)MeGly$^-$	
8.75	6.65	8.50	1.78
8.90	7.98	8.60	2.05
9.00	12.3	8.75	3.41
9.10	16.1	8.90	4.72
9.25	22.2	9.00	6.15
9.40	25.6	9.10	7.32
9.50	39.7	9.25	9.93
		9.40	14.1
		9.50	16.4

$$MLMeGly^{x+} + OH^- \rightarrow MLGly^{(x-1)+} + CH_3OH \quad (3)$$

The total amount of NaOH consumed during kinetic runs was always within 5% of the value expected from eq 3. The k_{obsd} values determined at different pH values indicate that reaction 3 follows the rate law

$$\text{rate} = k_{OH}[MLMeGly^{x+}][OH^-] \quad (4)$$

where $k_{OH} = k_{obsd}/[OH^-]$. The average values of k_{OH} at 25 °C and 0.10 M ionic strength are listed in Table IV. In the pH range studied, the rate of hydrolysis of MeGly is negligible in the absence of ML^{x+}.

The $Ni(TEP)^{2+}$ and Ni(DACODA) chelates were found to have no catalytic activity. Even in the presence of a 50-fold excess of the above two chelates to MeGly, the observed rate (k_{obsd}) was the same as that of MeGly in the absence of metal chelates. Apparently, Ni(DACODA) forms the hydroxo species in preference to binding MeGly, while the $Ni(TEP)^{2+}$ chelate has essentially little affinity for amino acid esters. Therefore, studies of the catalytic abilities of the two metal chelates were not pursued further.

Activation Parameters. The temperature dependencies of various metal chelate promoted hydrolysis of MeGly were determined in the range 20–35 °C. Data (k_{obsd}) for the reactions at 20, 25, 30, and 35 °C are given in Tables II and III. Activation parameters, ΔH^* and ΔS^*, for the hydrolysis of MeGly in the presence of various metal chelates are given in Table IV.

Discussion

Metal Chelate Catalyzed Hydrolysis of MeGly. The overall

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

pH	Temp, °C	$10^4 k_{obsd}$, s^{-1}	pH	Temp, °C	$10^4 k_{obsd}$, s^{-1}
	Co(NTA)MeGly$^-$			Cu(NTA)MeGly$^-$	
9.60	20	7.39	8.20	20	7.30
9.90	20	13.4	8.35	20	9.92
10.10	20	20.0	8.50	20	14.1
9.10	30	4.93	8.00	30	10.7
9.40	30	9.67	8.10	30	12.3
9.60	30	14.0	8.25	30	18.0
9.25	35	10.9	7.75	35	10.8
9.50	35	18.2	7.90	35	14.5
9.75	35	31.3	8.05	35	19.7
	Zn(NTA)MeGly$^-$			Ni(NTA)MeGly$^-$	
9.25	20	6.45	9.25	20	9.96
9.40	20	7.63	9.40	20	14.9
9.50	20	10.0	9.50	20	19.1
8.60	30	3.47	9.00	30	11.8
8.90	30	7.42	9.10	30	17.5
9.10	30	11.3	9.25	30	23.6
8.50	35	4.50	8.80	35	12.6
8.75	35	7.88	8.90	35	14.1
9.00	35	1.24	9.00	35	18.4
	Ni(trien)MeGly^{2+}			Ni(tren)MeGly^{2+}	
9.30	20	8.91	8.50	20	3.43
9.40	20	12.2	8.75	20	5.95
9.50	20	15.4	8.90	20	8.80
8.90	30	14.2	8.50	30	7.74
9.00	30	16.3	8.75	30	13.8
9.10	30	26.7	8.90	30	18.6
8.60	35	9.95	8.75	35	12.2
8:75	35	16.9	8.90	35	22.0
8.90	35	23.3	9.00	35	29.0
	Ni(EDDA)MeGly			Ni(EDDA)MeGly	
9.25	20	8.56	9.25	30	26.5
9.40	20	13.4	8.80	35	14.7
9.50	20	19.8	8.90	35	21.6
9.00	30	14.4	9.00	35	23.8
9.10	30	19.6			

Table IV. Activation Parameters for the Hydrolysis of Methyl Glycinate in the Presence of Metal Nitrilotriacetates and Tetradentate Ni(II) Chelates at 0.10 M Ionic Strength

M chelates	log K_L	k_{OH}, $M^{-1} s^{-1}$	ΔH^*, kcal/mol	ΔS^*, cal/mol °K
Ni(tren)MeGly^{2+}	14.8[a]	67.1	3.4 ± 0.8	−39 ± 2
Ni(trien)MeGly^{2+}	14.0[b]	53.1	7.2 ± 0.7	−26 ± 1
Ni(EDDA)MeGly	13.5[c]	41.2	4.8 ± 0.8	−35 ± 2
Ni(NTA)MeGly$^-$	11.47[d]	52.3	0.9 ± 0.7	−47 ± 2
Co(NTA)MeGly$^-$	10.81[d]	18.6	1.5 ± 0.8	−48 ± 2
Cu(NTA)MeGly$^-$	13.05[d]	460[f,g]	3.4 ± 1.2	−38 ± 3
Zn(NTA)MeGly$^-$	10.44[e]	34.6	4.0 ± 0.7	−38 ± 2
Cu(NTA)EtGly$^-$	13.05[d]	78.2[f,g]	4.9[f,g]	−33[f,g]

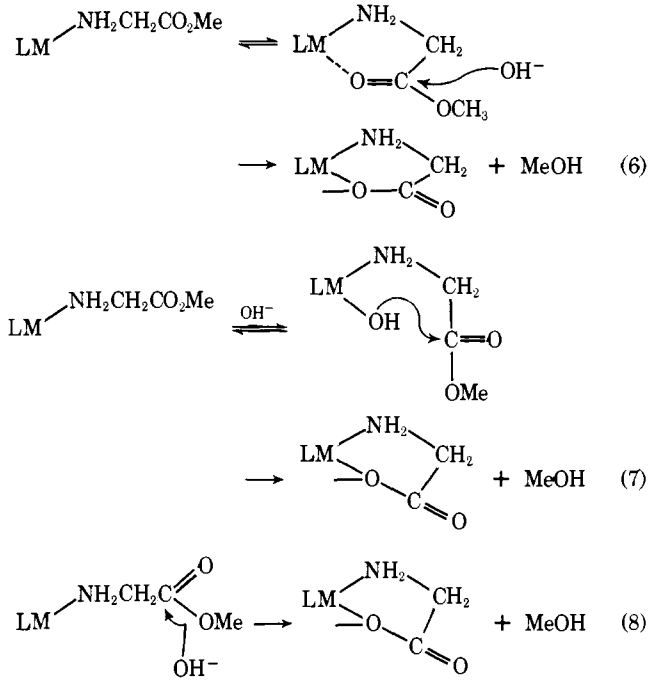
[a] J. E. Prue and G. Schwarzenbach, *Helv. Chim. Acta*, **33**, 963 (1950). [b] G. Schwarzenbach, *Helv. Chim. Acta*, **33**, 974 (1950). [c] S. Chaberek and A. E. Martell, *J. Am. Chem. Soc.*, **74**, 6228 (1952). [d] J. Stary, *Anal. Chim. Acta*, **28**, 132 (1963). [e] H. M. N. H. Irving and M. G. Miles, *J. Chem. Soc. A*, 727 (1966). [f] R. J. Angelici and D. Hopgood, *J. Am. Chem. Soc.*, **90**, 2514 (1968). [g] I = 0.05 M (KNO_3).

metal chelate catalyzed hydrolysis of MeGly proceeds in the following two steps.

$$ML^{x+} + MeGly \overset{K_x}{\rightleftharpoons} MLMeGly^{x+}$$

$$MLMeGly^{x+} \xrightarrow[OH^-]{k_{OH}} MLGly^{(x-1)+} + MeOH \quad (5)$$

Under the conditions used here the MeGly is bound almost entirely as $MLMeGly^{x+}$. Therefore, the observed rate law represents the second step (k_{OH}) only. The first-order dependence on OH^- concentration for this step may be accounted for by three general mechanisms.[2-6] One involves an initial rapidly established equilibrium in which ester group coordination occurs, followed by rate determining OH^- attack (eq 6). The second involves rapid, equilibrium formation of the M-OH complex, followed by intramolecular OH^- attack (eq 7). The third involves only OH^- attack on coordinated MeGly (eq 8). Buckingham, Foster, and Sargeson[12] via isotopic studies have shown that pathways 1 (eq 6) and 2 (eq 7) are important in the ester hydrolysis of the inert complex *cis*-$Co(en)_2(Br)(NH_2CH_2CO_2\text{-}i\text{-}Pr)^{2+}$. In labile metal complex catalyzed reactions of amino acid esters, it has been very difficult to establish whether one or a combination of the above mechanisms are involved in the observed hydrolysis. Angelici and Leach,[13] using different nucleophiles, have suggested that external attack on the coordinated carbonyl group (eq 6) was the more likely mechanism for Cu^{2+} complex catalysis. However, the data presented here suggests that external OH^- attack on the uncoordinated carbonyl group (eq 8) is the mechanism for metal nitrilotriacetate and tetradentate Ni(II) complex promoted hydrolysis.

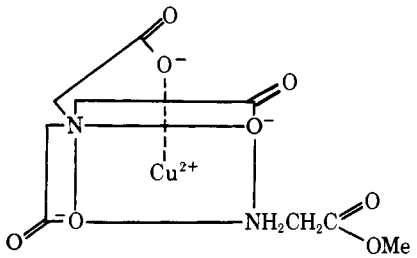


It is quite common that parallel changes in ΔH^* and ΔS^* occur for a series of reactions involving common or common type reactants. The parallel changes in ΔH^* and ΔS^* yield only small changes in ΔG^*, and for a closely related series, a common mechanism is supported.[14] In Figure 1, a plot of ΔH^* vs. ΔS^* for a series of metal nitrilotriacetate and tetradentate Ni(II) chelate promoted hydrolysis of MeGly yields a straight line, indicating that the rate determining step in each of the hydrolyses is the same.

Mechanism 3 is supported in that the catalytic effect of a proton is greater than that observed for most of the metal chelates studied here. The second-order rate constants, k_{OH}, for hydrolysis of [Cu(NTA)MeGly$^-$], [Zn(NTA)MeGly$^-$], and [Ni(NTA)MeGly$^-$] are only 16, 26, and 40, respectively,

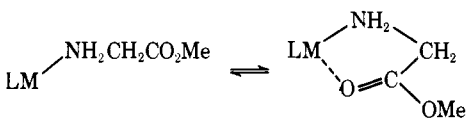
times larger than that of MeGly itself. Protonated ethyl glycinate, on the other hand, has been reported to undergo hydrolysis at a rate 41 times that of the unprotonated ester.[15] The above data indicate that a proton is a better catalyst than NTA chelates of Co(II), Cu(II), and Zn(II). Since the higher rate of hydrolysis of $EtGlyH^+$ as compared to that of EtGly is most likely due to the inductive effects of the added proton, it appears that the above $MNTA^-$ chelates are less capable of polarizing the carbonyl group than a proton. In view of the above, mechanism 3 (eq. 8) is supported, for pathways 1 and 2 (eq 7 and 8, respectively) have been shown to yield catalytic effects much greater than that of a proton.[12] Similar conclusions can be reached concerning the $Ni(trien)MeGly^{2+}$ and $Ni(EDDA)MeGly^{2+}$ complexes, which have rate constants that are only 40 and 31, respectively, times as large as that of MeGly.

Further evidence supporting mechanism 3 is that the K_{OH} values for hydroxo complex formation in $Ni(tren)^{2+}$ and $Ni(trien)^{2+}$ complexes are $10^{1.86}$ and $10^{1.94}$ M^{-1}, respectively. These values indicate that there is essentially no hydroxo complex formation in the pH range 8.40–9.00, where both $Ni(tren)^{2+}$ and $Ni(trien)^{2+}$ are catalytically active. Furthermore, it does not seem probable that $CuNTAMeGly^-$ (I)



I

forms a hydroxo complex, especially in the pH range (7.75–8.25) where $CuNTA^-$ is an active catalyst. Hopgood and Angelici[16] have approximated the extent of amino acid ester chelation, i.e., carbonyl oxygen binding (eq 9) in metal nitri-



(9)

lotriacetates, by comparing the K_x value of glycine to that of ethyl valinate. The authors[16] concluded that the ester carbonyl oxygen atom in I and in the corresponding Ni(II) complex was not coordinated to the metal to any appreciable extent, providing further support for pathway 3.

Of the metal chelates studied here, only $CuNTAMeGly^-$ undergoes hydrolysis at a very accelerated rate, i.e., 350 times that of MeGly. This can be explained in terms of mechanism 3 and the Lewis acidity of the metal chelates. One measure of the Lewis acidity of a metal chelate is the ability of that complex to form a hydroxo species (eq 1). The K_{OH} values for $NiNTA^-$,[17] $Ni(tren)^{2+}$, and $Ni(trien)^{2+}$ are approximately $10^{1.90}$, while that of $CuNTA^-$ is $10^{4.39}$.[16] It is apparent that CuNTA has a much greater Lewis acidity and, therefore, should be capable of polarizing the ester carbonyl group to a greater extent than the other metal chelates studied here. This is predicted if pathway 3 (eq 8) is the mechanism for metal chelate promoted hydrolysis. However, the greatly accelerated hydrolysis rate for $CuNTAMeGly^-$ as compared to the others may possibly be due to the presence of another pathway, and its apparent fit in the isokinetic plot (Figure 1) could be fortuitous.

Correlation of Activation Parameters with Ester Hydrolysis Rates. Previously, it was proposed[18] that the catalytic activity of metal chelates toward amino acid ester hydrolysis could be

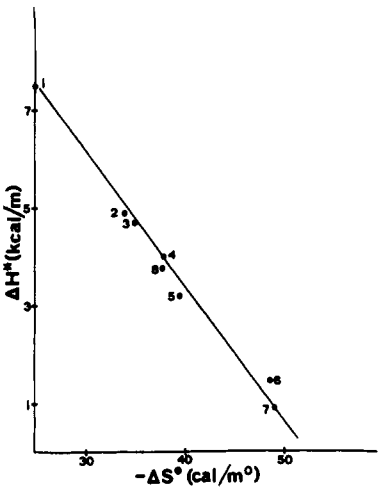


Figure 1. Plot of ΔH^* vs. ΔS^* for metal nitrilotriacetate and tetradentate nickel(II) chelate promoted hydrolysis of methyl glycinate: 1, $Ni(trien)^{2+}$; 2, $CuNTAEtGly^-$; 3, NiEDDA; 4, $ZnNTA^-$; 5, $Ni(tren)^{2+}$; 6, $CoNTA^-$; 7, $NiNTA^-$; and 8, $CuNTA^-$.

correlated with the formation constant of the metal chelate

$$M^{2+} + L^{x-} \rightleftharpoons ML^{(2-x)+} \qquad (10)$$

i.e., large formation constants result in reduced Lewis acid character of the metal chelate and, therefore, lower catalytic activities. While such a trend was observed to occur for a series of Cu^{2+} chelates,[18] it is not observed for the Ni(II) complexes studied here. The trend in the catalytic ability of the metal chelates if $Ni(tren)^{2+} > Ni(trien)^{2+} > NiNTA^- > NiEDDA$, while that of the formation constants is $Ni(tren)^{2+} > Ni(trien)^{2+} > NIEDDA > NiNTA^-$ (Table IV).

Hay and Morris[19] studied the Cu(II) complex catalyzed hydrolyses of $NH_2CH_2CH(NH_2)CO_2CH_3$, E, and reported that the rate constants, k_{OH} (eq 4), decreased as follows (A = $NH_2CH_2CH(NH_2)CO_2^-$): $CuE^{2+} > CuE_2^{2+} > CuE(OH)^+ > CuE(NH_2CH_2CH_2NH_2)^{2+} > CuEA^+$. Although the decrease in k_{OH} follows an increase in the formation constant of the metal chelate, the authors[19] interpreted their trend in terms of the magnitude of positive charge on the metal chelates; those complexes with the higher positive charge yielded the highest k_{OH}. This explanation does not, however, account for the data obtained on other Cu^{2+} chelates[18] or for the trends reported here.

The data obtained here for the Ni(II) chelates suggests that both the Lewis acidity of the metal chelate and the charge are both important in determining the catalytic ability toward amino acid ester hydrolyses. Since the Ni(II) chelates studied here all contain tetradentate ligands and all probably have octahedral geometries in aqueous solution, an attempt was made to correlate ΔH^* and ΔS^* with the log of the formation constant of the metal chelate (eq 10) (Figure 2). With the exception of $Ni(tren)^{2+}$, ΔH^* decreases with decreasing log K_L, while ΔS^* decreases with reduced positive charge on the metal chelate. The above data can be interpreted in terms of mechanism 3 (eq 8), assuming that the ΔH^* values are to a large extent a measure of the ability of the metal chelate to polarize the ester carbonyl group, while the ΔS^* values are to a large extent a measure of the effect of the total charge on the metal complex. Positive $Ni(trien)MeGly^{2+}$ having a less negative entropy of activation than neutral Ni(EDDA)MeGly and negative $Ni(NTA)MeGly^-$ is due to cancellation of charge in the activated complex (eq 8) in the first case and buildup of charge in the last case.

The $Ni(tren)^{2+}$ chelate, which is an exception to the above, is severely strained as shown by x-ray studies[20] of similar

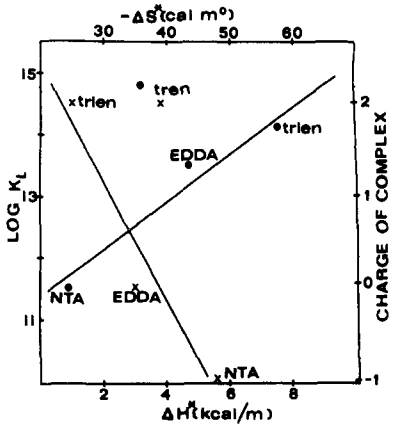


Figure 2. Plot of log K_L vs. ΔH^* (⊙) and total charge on metal complex vs. ΔS^* (×) for a series of tetradentate nickel(II) chelates.

complexes. Therefore, Ni(tren)MeGly^{2+} may exist in solution in two forms, one with tren acting as a tetradentate ligand and another as a tridentate one. This could explain the unexpectedly low ΔH^* value for Ni(tren)MeGly^{2+} hydrolysis. If one of the catalytic active species was Ni(tren)MeGly^{2+} with tren acting as a tridentate ligand, than its formation constant value would be smaller and its Lewis acidity higher than if it was acting as a tetradentate donor. Therefore, ΔH^* for Ni(tren)-MeGly^{2+} hydrolysis could be an average value for the catalytic effects of two complexes, one with tren binding as a tridentate ligand and the other as a tetradentate donor.

In summary, these studies indicate that both the Lewis acidity and the charge of a metal chelate are important in determining its catalytic activity toward amino acid ester hydrolysis, and that previous studies[18,19] indicating that only one of the two effects was important were probably fortuitous. Moreover, other effects such as strain and steric considerations, the number of donor groups, and probably the stereochemistry of the metal chelates are important. It is probably for the latter reason that no correlation is observed to exist between log K_L and the values of ΔH^* and ΔS^* for the metal nitrilotriacetate promoted hydrolysis of MeGly.

Acknowledgment. The authors wish to thank Professor R. J. Angelici for the use of a kinetic cell and Professor J. I. Legg for the gift of Ba(DACODA). The authors also gratefully acknowledge support of this research through West Virginia University Senate Grant 7740.

References and Notes

(1) H. Kroll, *J. Am. Chem. Soc.*, **74**, 2036 (1972).
(2) F. Basolo and R. G. Pearson, "Mechanisms of Inorganic Reaction Mechanisms", 2d ed, Wiley, New York, N.Y., 1967, pp 625, 632.
(3) M. M. Jones, *Adv. Chem. Ser.*, **No. 49**, 1 (1965).
(4) T. C. Bruice and S. Benkovic, "Bioorganic Mechanisms", Vol. 1, W. A. Benjamin, New York, N.Y., 1966, pp 110–118.
(5) R. G. Wilkins, "The Study of Kinetics and Mechanism of Reactions of Transition Metal Complexes", Allyn and Bacon, Boston, Mass., 1974, pp 310–314.
(6) M. C. Scutton in "Inorganic Biochemistry", Elsevier, New York, N.Y., 1973, Chapter 14.
(7) D. D. Perrin, W. L. F. Armarego, and D. R. Perrin, "Purification of Laboratory Reagents", Pergamon Press, Oxford, 1966.
(8) K. S. Rajan and A. E. Martell, *J. Inorg. Nucl. Chem.*, **26**, 789 (1964).
(9) J. Bjerrum, "Metal Ammine Formation in Aqueous Solution", P. Haase and Son, Copenhagen, Denmark, 1957.
(10) R. J. Angelici and B. E. Leach, *J. Am. Chem. Soc.*, **89**, 4605 (1967).
(11) L. G. Sillen and A. E. Martell, *Chem. Soc., Spec. Publ.*, **No. 25**, 1 (1972).
(12) D. A. Buckingham, D. M. Foster, and A. M. Sargeson, *J. Am. Chem. Soc.*, **91**, 4102 (1969).
(13) R. J. Angelici and B. E. Leach, *J. Am. Chem. Soc.*, **90**, 2499 (1968).
(14) R. G. Wilkins, ref 5, pp 100–101.
(15) H. L. Conley, Jr., and R. B. Martin, *J. Phys. Chem.*, **69**, 2914 (1965).
(16) D. Hopgood and R. J. Angelici, *J. Am. Chem. Soc.*, **90**, 2508 (1968).
(17) G. Schwarzenbach and W. Beiderman, *Helv. Chim. Acta*, **31**, 331 (1948).
(18) R. Nakon, P. R. Rechani, and R. J. Angelici, *J. Am. Chem. Soc.*, **96**, 2117 (1974).
(19) R. W. Hay and P. J. Morris, *J. Chem. Soc., Dalton Trans.*, 56 (1973).
(20) C. G. Pierpont, D. N. Hendrickson, D. M. Duggan, F. Wagner, and E. K. Barefield, *Inorg. Chem.*, **14**, 604 (1975).

Porphyrins. 34.[1] Phosphorus Complexes of Octaethylporphyrin

Phillip Sayer, Martin Gouterman,* and Charles R. Connell

Contribution from the Department of Chemistry, University of Washington, Seattle, Washington 98195. *Received July 5, 1976*

Abstract: The synthesis and optical spectra of previously unreported complexes of octaethylporphyrin (OEP) with phosphorus are described. Phosphorus, like arsenic and antimony, can assume oxidation states of III and V in the porphyrin hole. (Bi forms only a III complex.) The complexes have the forms (OEP)$P^{III}X^-$ and (OEP)$P^V(X)_2{}^+Y^-$, with X an anionic ligand and Y^- a counterion. Iterative extended Hückel calculations rationalize the fact that the V species have normal porphyrin absorption and emission spectra while the valence III species show extra absorption bands and no fluorescence. The extra bands of the III species are attributed to charge transfer transitions $a_{2u}(3p_z)$ (phosphorus) → $e_g(\pi^*)$(ring), that strongly interact with the *normal* porphyrin excited (π,π^*) states.

The metalloporphyrin derivatives of group 5a elements have been the subject of recent experimental[2-4] and theoretical[3] studies. These compounds are of interest chemically and spectroscopically because the central atoms As and Sb can exist in both the formal oxidation states III and V. Bismuth is known only as the III complex. Theory predicted and new experiments proved[3,4] (contrary to the original identification[2]) that the valence V species has the form (P)$M^V(X)_2{}^+Y^-$ and has a *normal* optical absorption spectrum. In this notation (P) stands for the porphyrin ring, M stands for As or Sb, X stands for an anionic ligand such as OH or Cl, and Y^- stands for a nonligand counterion such as Cl^- or ClO_4^-. By *normal* absorption we mean a visible-near-uv spectrum consisting only of Q(π,π^*), B(π,π^*) (Soret), and N(π,π^*) (~330 nm) bands in the region $\lambda > 320$ nm. These compounds are pink in solution and show clear fluorescence and phosphorescence.

On the other hand, valence III species were predicted and shown to have form (P)$M^{III}Y^-$ and a hyper absorption