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Trivalent Copper. II. A Pulse Radiolytic Study of the Formation and Decomposition of Amino Complexes^{1a}

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The effect of amino ligands on the radiolytic induced formation and decomposition of trivalent copper complexes was studied. The spectra of the Cu(III) complexes with ammonia, ethylenediamine, and different amino acids were measured. The specific rates of oxidation of $\text{Cu}(\text{NH}_3)_4^{2+}$ by $\cdot\text{NH}_2$ radicals and the complexes of Cu^{2+} with ethylenediamine and different amino acids by $\cdot\text{OH}$ radicals were determined. The $\text{Cu}^{\text{III}}(\text{NH}_3)_n$ complex decomposes *via* a second-order reaction. The $\text{Cu}^{\text{III}}(\text{en})_2$ complex also decomposes *via* a second-order reaction. The specific rate of the latter reaction increases with increasing pH from $2k = 5.5 \times 10^5 \text{ M}^{-1} \text{ sec}^{-1}$ at pH 5.8 to $2k = 3.9 \times 10^7 \text{ M}^{-1} \text{ sec}^{-1}$ at pH 11.6. It is suggested that this increase is due to the formation of $\cdot\text{OH}$ radicals as an intermediate in the decomposition reaction in alkaline solutions. All the amino acid complexes of Cu(III) decompose *via* a first-order reaction with a specific rate of $(1-8) \times 10^8 \text{ sec}^{-1}$. It is suggested that the first step in the latter reaction is an intramolecular electron transfer from the carboxylic group to the central copper atom. The specific rate of the reaction $\text{Cu}(\text{III})_{\text{aq}} + \text{Cu}^{\text{II}}\text{gly} \rightarrow \text{Cu}_{\text{aq}}^{2+} + \text{Cu}^{\text{III}}\text{gly}$ was measured: $k = 8.1 \times 10^7 \text{ M}^{-1} \text{ sec}^{-1}$.

The mechanism of the radiolytic oxidation of aliphatic amino compounds has been the subject of many studies (see for example ref 2-5). It was suggested that different metal cations catalyze the radiolytic induced oxidation of amino compounds.^{2b, 3, 6-8} Special attention has been given to the effect of copper ions on this process.^{3, 6-8} It has been suggested that the radiolytic induced oxidative deamination of amino acids, peptides, and ethylenediamine is catalyzed by copper ions *via* the formation of trivalent copper which oxidizes its ligands.^{7, 9-11} It was pointed out that this mechanism might be of great radiobiological importance.

The formation and decomposition of the aquo complex of trivalent copper have been studied in detail.¹² It was decided to study the effect of different amino acids, ethylenediamine, and ammonia on the chemical properties of trivalent copper. The results indicate that the presence of amino ligands enhances the rate of formation of trivalent copper. The decomposition of the complexes formed was found to proceed by a second-order process in the case of the NH_3 and $\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2$ complexes and by a first-order process for the amino acid complexes. It is suggested that this first-order process is the oxidation of the ligand by the central cation *via* the carboxylate group.

Experimental Section

Materials.—All solutions were prepared from triply distilled water and AR reagents. The amino acids and the ethylenediamine sulfate were supplied by K & K Laboratories.

Procedure.—The procedure of preparation of the samples,

(1) (a) Based in part on work performed under the auspices of the U. S. Atomic Energy commission. (b) Address correspondence to Nuclear Research Centre-Negev.

(2) (a) R. L. S. Willix and W. M. Garrison, *Radiat. Res.*, **32**, 452 (1967); (b) W. M. Garrison in "Current Topics in Radiation Research," Vol. IV, M. Ebert and A. Howard, Ed., North-Holland Publishing Co., Amsterdam, 1968, p 45.

(3) M. Anbar, R. A. Munoz, and P. Rona, *J. Phys. Chem.*, **67**, 2708 (1963).

(4) P. Neta, M. Simic, and E. Hayon, *ibid.*, **74**, 1214 (1970).

(5) P. Neta and R. W. Fessenden, *ibid.*, **74**, 2263 (1970).

(6) R. L. S. Willix and W. M. Garrison, *ibid.*, **69**, 1579 (1965).

(7) M. Anbar, *Advan. Chem. Ser.*, **No. 49**, 126 (1965).

(8) G. A. Shagisultanova, L. A. Ilyukevitch, and I. G. Sidorova, *Khim. Vys. Energ.*, **3**, No. 4 (1969).

(9) M. Anbar and A. Levitzki, *Radiat. Res.*, **27**, 32 (1966).

(10) A. Levitzki and M. Anbar, *J. Amer. Chem. Soc.*, **89**, 4185 (1967).

(11) A. Levitzki, M. Anbar, and A. Berger, *Biochemistry*, **6**, 3757 (1967).

(12) D. Meyerstein, submitted for publication.

irradiation, observation of the optical changes, and calculation of the specific rate constants have been described earlier in detail.^{12, 13}

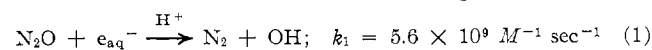
The pulse radiolytic experiments were carried out using a 0.4- μsec , 15-MeV, 20-160-mA electron pulse from the linear accelerator at Argonne National Laboratory or by using a 0.1-1.3- μsec , 5-MeV, 200-mA pulse from the linear accelerator at the Hebrew University in Jerusalem. At both linear accelerators the dose per pulse has been $(1-10) \times 10^{16} \text{ eV l}^{-1}$.

The absolute molar absorption coefficients were determined by alternately pulse radiolyzing under exactly identical conditions a standard solution, the solution to be analyzed, and once again the standard solution. The standard solution was $5 \times 10^{-4} \text{ M}$ in $\text{Cu}(\text{ClO}_4)_2$ at pH 4.5, saturated with N_2O . As the complex solutions were also saturated with N_2O , it was assumed that the yields of trivalent copper complexes have been identical in all cases and that $\epsilon_{800}(\text{Cu}(\text{III})_{\text{aq}})$ is 5700.¹⁴ The accuracy of ϵ thus determined is $\pm 15\%$.

Results and Discussion

Absorption Spectra of Cu(III) Complexes.—Solutions containing $5 \times 10^{-4} \text{ M}$ CuSO_4 and a $9.5 \times 10^{-4} \text{ M}$ concentration of the corresponding amino ligand saturated with N_2O were irradiated and the spectra of the transients formed are summarized in Table I and Figure 1. The solutions under these conditions contain $5 \times 10^{-4} \text{ M}$ $\text{Cu}^{\text{II}}\text{L}_2$ (where L is an amino acid) and some $\text{Cu}^{\text{II}}\text{L}_1$ but nearly no free ligand. (This is correct for all the amino acid ligands above pH 6 as the stability constants for the formation of the complexes are all above $\log K_1 = 8$ and $\log K_2 = 6$ ¹⁵ and the $\text{p}K$'s for the acids are all about $\log K_1 \approx 9-10$ and $\log K_2 = 2-2.5$.¹⁵ There is therefore a possibility that a relative large fraction of free ligand existed in the γ -amino-*n*-butyric acid solution as we could not raise the pH above 4.8 without precipitation.)

In N_2O -saturated solutions the hydrated electrons are transformed into OH radicals according to¹⁶



As long as the concentration of copper ion is less than $1 \times 10^{-3} \text{ M}$, it will not compete with N_2O for the hy-

(13) D. Meyerstein and W. A. Mulac, *Inorg. Chem.*, **9**, 1762 (1970).

(14) J. H. Baxendale, E. M. Fielden, and J. P. Keene in "Pulse Radiolysis," M. Ebert, J. P. Keene, A. J. Swallow, and J. H. Baxendale, Ed., Academic Press, New York, N. Y., 1965, p 217.

(15) L. G. Sillén and A. E. Martell, *Chem. Soc., Spec. Publ.*, **No. 17** (1964).

(16) (a) M. Anbar and P. Neta, *Int. J. Appl. Radiat. Isotop.*, **18**, 493 (1967); (b) F. S. Dainton and S. A. Sills, *Nature (London)*, **186**, 879 (1960).

TABLE I
 SPECTROSCOPIC DATA FOR TRIVALENT COPPER COMPLEXES^a

Ligand	pH	λ_{\max} , ^b nm	ϵ_{\max} , ^c $M^{-1} \text{cm}^{-1}$
H ₂ O	5.7	295	5700 ^e
NH ₃ ^d	11.1	290	5600
Ethylenediamine	6.0	300	2100
	11.3	300	2500
Glycine	7.5	310	7800
α -Alanine	7.2	305	4700
β -Alanine	6.0	310	4800
α -Amino- <i>n</i> -butyric acid	7.3	310	3000
β -Amino- <i>n</i> -butyric acid	6.1	310	3300
γ -Amino- <i>n</i> -butyric acid	4.6	310	3500
α -Aminoisobutyric acid	7.3	310	3200

^a All the solutions contained $5 \times 10^{-4} M$ CuSO₄ and $9.5 \times 10^{-4} M$ amino ligand and were N₂O saturated. ^b The error limit on λ_{\max} is ± 5 nm. ^c The error limit on ϵ_{\max} is $\pm 15\%$. The calculations are based on $\epsilon_{\max} 5700 M^{-1} \text{cm}^{-1}$ for Cu(III)_{aq} and assume that the yield of Cu^{III}L₂ is equal to that of Cu(III)_{aq}. ^d [NH₄OH] = 0.2 M. ^e From ref 14.

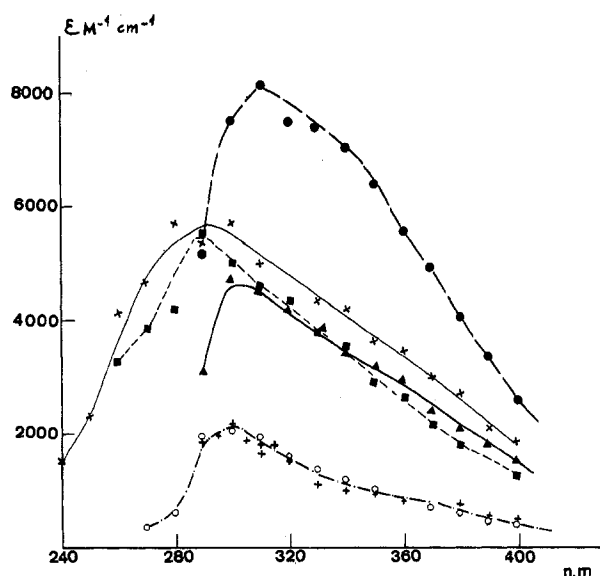


Figure 1.—Absorption spectra of trivalent copper complexes as measured immediately after the end of formation of the transients in the following solutions: ×, $5 \times 10^{-4} M$ CuSO₄, N₂O saturated, pH 5.7; □, $5 \times 10^{-4} M$ CuSO₄, 0.2 M NH₄OH, N₂O saturated, pH 11.1; ○, $5 \times 10^{-4} M$ CuSO₄, $9.5 \times 10^{-4} M$ ethylenediamine sulfate, N₂O saturated, pH 6.0; +, $5 \times 10^{-4} M$ CuSO₄, $9.5 \times 10^{-4} M$ ethylenediamine sulfate, N₂O saturated, pH 11.3; ●, $5 \times 10^{-4} M$ CuSO₄, $9.5 \times 10^{-4} M$ glycine, N₂O saturated, pH 7.5; △, $5 \times 10^{-4} M$ CuSO₄, $9.5 \times 10^{-4} M$ α -alanine, N₂O saturated, pH 7.2.

drated electrons.¹⁵ It can be concluded therefore that in the irradiated solutions OH radicals were formed with a yield of $G \approx 5.5$ (molecules/100 eV absorbed) and H atoms were formed with a yield up $G \approx 0.5$.¹² The H atoms are expected to reduce some of the copper complexes, and the OH radicals, to oxidize them.¹² The spectra observed are very similar to those of Cu(III)_{aq}.¹⁴ Cu_{aq}³⁺ is known to have no absorption band in the region studied.¹⁴ Furthermore, the initial absorption decreased by about 50% when Ar-saturated solutions were irradiated instead of N₂O-saturated solutions. It is concluded therefore that the absorption spectra measured are due to oxidation products. The free amino ligand reacts slowly with the OH radical in neutral solution^{16a,17} and so, even if a low concentration of it is present, would not compete with the complex for OH

(17) M. Anbar, D. Meyerstein, and P. Neta, *J. Chem. Soc. B*, 742 (1966).

radicals. Furthermore, the absorption maxima of the radicals formed by H abstraction from the amino ligands all lie below 250 nm.^{4,13} It is concluded that the observed spectra are due to Cu^{III}L₂ complexes formed by the oxidation of the corresponding Cu^{II}L₂ complexes. (A small contribution of Cu^{III}L₁ cannot be ruled out.)

It is evident from Table I that the position of the maximum of absorption of the Cu(III) band is only slightly affected by the different ligands, being shifted by 10–20 nm to the red relative to the aquo complex. The absorption coefficient at λ_{\max} decreases from glycine to alanine and the different butyro amino acids. Part of this decrease might be due to hydrogen abstraction by OH radicals from the amino acid ligands at sites remote from the NH₂ group, thus decreasing the Cu(III) yield. The decrease in the absorption at λ_{\max} seems to be larger than that calculated from the known specific rates of reaction of OH radicals with the complexes (Table II) and with different RCH₃, R₂CH₂, and R₃CH

 TABLE II
 SPECIFIC RATES OF FORMATION AND DISAPPEARANCE OF DIFFERENT Cu(III) COMPLEXES^a

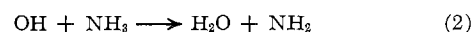
Ligand	pH	$k(\text{OH} + \text{Cu}^{\text{II}}\text{L}_2)$, ^b $M^{-1} \text{sec}^{-1}$	$k(\text{Cu}^{\text{III}}\text{L}_2 \rightarrow \text{Cu}^{\text{II}}\text{L}_1 + \text{products})$, ^c sec^{-1}
H ₂ O	5.7	3.1×10^8	$< 1.0 \times 10^2$
NH ₃ ^d	11.1	1.3×10^8	$< 1.0 \times 10^2$
Ethylenediamine	6.5	3.0×10^9	$< 3.0 \times 10^6$
	11.2	8.0×10^9	$\sim 1.2 \times 10^2$
Glycine	6.1	1.5×10^9	6.0×10^8
	7.5	...	2.2×10^4
α -Alanine	6.3	1.4×10^9	8.0×10^8
β -Alanine	5.8	1.2×10^9	7.0×10^8
α -Amino- <i>n</i> -butyric acid	6.1	2.0×10^9	5.0×10^8
β -Amino- <i>n</i> -butyric acid	6.0	1.2×10^9	4.5×10^8
γ -Amino- <i>n</i> -butyric acid	4.8	1.1×10^9	1.2×10^8
α -Aminoisobutyric acid	6.2	1.8×10^9	1.5×10^8
	7.3	...	2.5×10^8

^a Solution composition: $(5-50) \times 10^{-5} M$ CuSO₄; the ratio of concentration of amino ligand to Cu²⁺ was 1.9; N₂O saturated. ^b The error limit is $\pm 20\%$. ^c First-order disappearance rate of Cu(III) complexes. The error limit is $\pm 30\%$. ^d [NH₃] = 0.2 M. ^e The reaction measured is NH₂ + Cu(NH₃)₄²⁺ → Cu^{III}(NH₃)_n (see text).

groups.¹⁷ However the effect of copper on the latter rates is not known.

It has been suggested that the absorption spectrum of Cu(III)_{aq} is due to a charge transfer from the ligand to the central copper atom.¹⁴ It should be mentioned that the absorption band of the Cu(III) complexes is unsymmetrical and might therefore be due to two overlapping transitions. The relatively small effect of changing the ligands on the position of the absorption maxima seems difficult to reconcile with the charge-transfer assignment. It has to be concluded that more experimental data on the spectra of different Cu(III) complexes are needed in order to identify the transitions involved.

Kinetics of Formation and Disappearance of the Cu(III) Complexes. The Ammonia Complex.—Solutions saturated with N₂O and containing 0.2 M NH₃ and $(1-20) \times 10^{-4} M$ CuSO₄ at pH 11.1 \pm 0.1 were irradiated. By following the absorption at 300 nm the specific rate of formation of the Cu^{III}(NH₃)_n complex was determined to be $k = (1.3 \pm 0.2) \times 10^8 M^{-1} \text{sec}^{-1}$. The specific rate of the reaction



is $\sim 5 \times 10^7 M^{-1} \text{sec}^{-1}$.¹⁸ It is evident that the reaction observed is not that of OH radicals with $\text{Cu}(\text{NH}_3)_4^{2+}$ because $k_2[\text{OH}][\text{NH}_3] > k_{\text{OH}+\text{Cu(II)}}[\text{OH}][\text{Cu(II)}]$. It is suggested that the reaction observed is



This would also explain the observation that the specific rate of oxidation of $\text{Cu}(\text{NH}_3)_4^{2+}$ is lower than that of $\text{Cu}_{\text{aq}}^{2+}$ whereas for all the other amino complexes the specific rates are higher (Table II).

The disappearance of $\text{Cu}^{\text{III}}(\text{NH}_3)_n$ in N_2O -saturated solutions was found to obey a second-order mechanism with a specific rate of $2k = (3.0 \pm 0.5) \times 10^7 M^{-1} \text{sec}^{-1}$. This rate of reaction is very similar to that found for the disappearance of $\text{Cu}(\text{III})_{\text{aq}}$ where $2k = 4.6 \times 10^7 M^{-1} \text{sec}^{-1}$ ¹² and of $\text{Cu}^{\text{III}}(\text{en})_2$ at pH 11.6 where $2k = 3.9 \times 10^7 M^{-1} \text{sec}^{-1}$ (see below).

The Ethylenediamine Complex.—Solutions saturated with N_2O and containing CuSO_4 and ethylenediamine in the ratio 1:1.9 were irradiated. The specific rate of the reaction



was measured to be $(3.0 \pm 0.6) \times 10^9$, $(5.0 \pm 1.0) \times 10^9$, and $(8.0 \pm 2.0) \times 10^9 M^{-1} \text{sec}^{-1}$ at pH 6.5, 10.2, and 11.2, respectively.

The mechanism of disappearance of the $\text{Cu}^{\text{III}}(\text{en})_2$ complex was studied and found to obey a second-order rate law (Figure 2). The specific rate of the disappear-

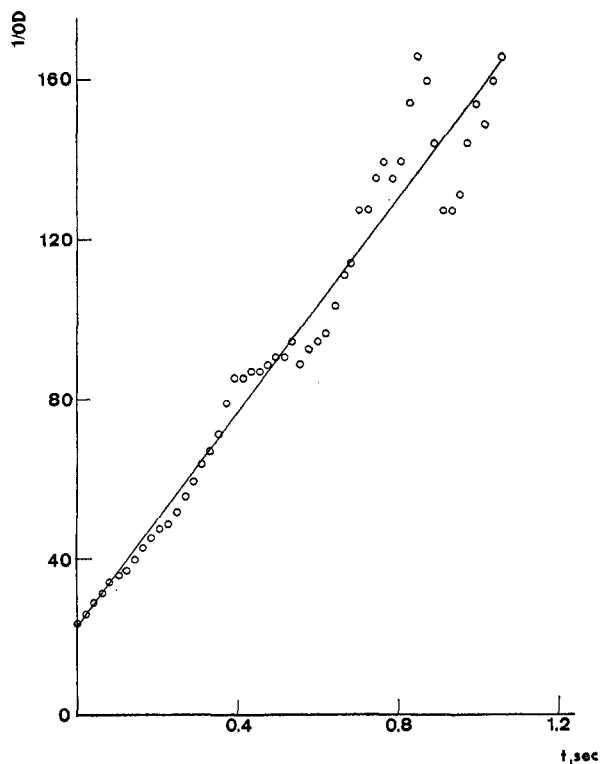


Figure 2.—Second-order disappearance plot for the $\text{Cu}^{\text{III}}(\text{en})_2$ complex. Solution composition: $10^{-3} M \text{CuSO}_4$, $1.9 \times 10^{-3} M$ ethylenediamine sulfate, N_2O saturated, pH 6.4.

ance reaction was found to increase with increasing pH (Table III). Furthermore a growing contribution of a first-order process was observed with the increase in pH.

(18) I. Lati and D. Meyerstein, to be submitted for publication.

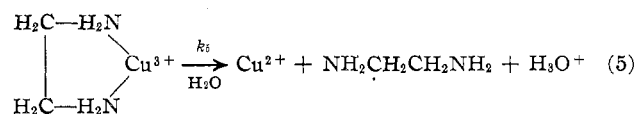
TABLE III
THE EFFECT OF pH AND OF EXCESS ETHYLENEDIAMINE
ON THE RATE OF DECOMPOSITION OF $\text{Cu}^{\text{III}}(\text{en})_2^a$

pH	$[\text{en}]^b M$	$2k_2^c M^{-1} \text{sec}^{-1}$	$k_1^d \text{sec}^{-1}$
5.8	9.5×10^{-4}	5.5×10^5	< 3.0
6.6	9.5×10^{-4}	1.3×10^6	...
7.8	9.5×10^{-4}	2.0×10^6	...
8.8	9.5×10^{-4}	4.9×10^6	...
9.3	9.5×10^{-4}	8.1×10^6	...
9.3	1.95×10^{-3}	2.4×10^7	$\sim 6.5 \times 10$
9.6	1.95×10^{-3}	3.4×10^7	$\sim 1.0 \times 10^2$
10.0	9.5×10^{-4}	1.7×10^7	...
10.0	1.95×10^{-3}	6.6×10^7	$\sim 1.7 \times 10^3$
10.3	9.5×10^{-4}	2.0×10^7	...
10.6	1.95×10^{-3}	$\sim 2.1 \times 10^8$	3.0×10^3
10.7	9.5×10^{-4}	2.3×10^7	...
11.3	9.5×10^{-4}	3.0×10^7	...
11.3	1.95×10^{-3}	...	3.8×10^2
11.6	9.5×10^{-4}	$\sim 3.9 \times 10^7$	1.2×10^2

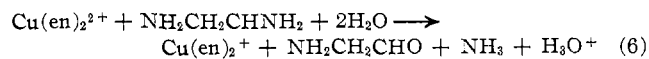
^a All solutions contained $5 \times 10^{-4} M \text{CuSO}_4$ and were N_2O saturated. ^b Concentration of ethylenediamine sulfate. ^c Specific rate of disappearance when the mechanism obeyed a second-order rate law. ^d Specific rate of disappearance when the mechanism obeyed a first-order rate law.

At pH 11.60 the reaction is better described as a first-order process with a specific rate constant of $1.2 \times 10^2 \text{sec}^{-1}$.

It has to be concluded that the intramolecular reaction

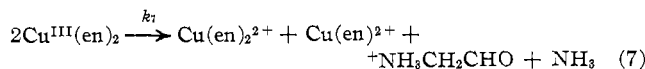


suggested for explaining the steady-state radiolysis of neutral copper ethylenediamine solutions^{3,7} is too slow to be observed under our conditions. An upper limit for this reaction rate at pH 6.0 of $k_5 < 3.0 \text{sec}^{-1}$, could be calculated, from the initial slope of a plot of $\ln \text{OD}$ vs. t . The latter value is in agreement with $k = 1.7 \times 10^{-3} \text{sec}^{-1}$ determined for the oxidation of tetraalanine by a central $\text{Cu}(\text{III})$ atom, where a C-H bond α to a peptide N atom, and not α to a free amino group, is oxidized.¹¹ The fact that $G(\text{NH}_3) = G(\text{OH})$ in the radiolysis of N_2O -saturated solutions at low dose rates, $\sim 10^4 \text{rads/min}$,³ indicates that under these conditions reaction 5 takes place. It is suggested that reaction 5 is then followed by



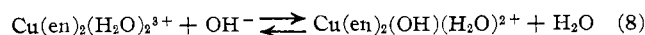
as it is known that radicals α to an amino group are strong reducing agents.^{6,19} This mechanism accounts for all the experimental results under low dose rate conditions.

It is suggested that the mechanism of the disappearance of $\text{Cu}^{\text{III}}(\text{en})_2$ under the high dose rate conditions is



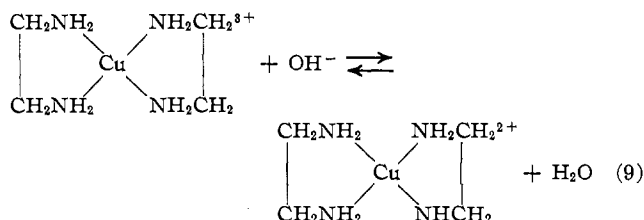
The increase in k_7 with increasing pH might be explained by two alternative mechanisms.

Mechanism a.—It is plausible that the $\text{Cu}^{\text{III}}(\text{en})_2$ complex behaves as an acid in basic solutions



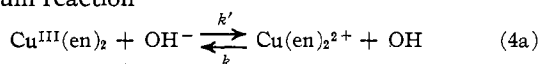
(19) H. Cohen and D. Meyerstein, to be submitted for publication.

or

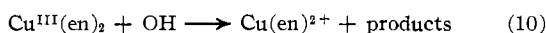


Similar basic forms of different amino complexes of trivalent cations are known.¹⁵ It might be argued that the reaction corresponding to reaction 7 for such a basic form is much faster than reaction 7.

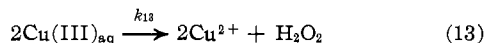
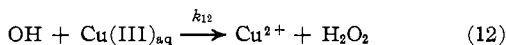
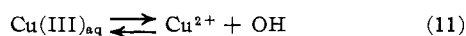
Mechanism b.—It is possible that at high pH's the equilibrium reaction



begins to contribute to the mechanism. If the reaction

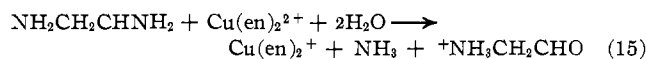


is faster than reaction 7, an increase in k with increasing pH is expected. This mechanism is similar to the mechanism suggested for the disappearance of $\text{Cu}(\text{III})_{\text{aq}}$ in acid solutions,¹² namely



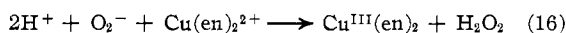
where $k_{12} \gg k_{13}$.

In order to check this mechanism, the effect of a small excess of ethylenediamine on the mode of disappearance of $\text{Cu}^{\text{III}}(\text{en})_2$ in basic solutions was studied. The results are summarized in Table III. It can be seen that the free en increases the specific rate of disappearance of $\text{Cu}^{\text{III}}(\text{en})_2$. Furthermore under these conditions a first-order process becomes the dominant reaction at lower pH's. These results are in agreement with a mechanism consisting of equilibrium 4a and reactions 7, 10, 14, and 15.



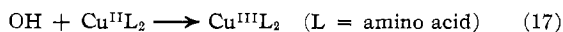
It is believed that this mechanism explains the increase in k_7 with increasing pH. This mechanism is also in agreement with the mechanism suggested for explaining some of the features of the electrocatalytic oxidation of ethylenediamine on Pt electrodes in the presence of Cu^{2+} .²⁰

It has been suggested that the reaction



occurs in O_2 -saturated solutions in low dose irradiations.^{3,7} Solutions of $(1-5) \times 10^{-4} M \text{Cu}(\text{en})_2^{2+}$ and $5 \times 10^{-2} M \text{CH}_3\text{OH}$ at pH 5.5 saturated with O_2 were irradiated and no formation of $\text{Cu}(\text{III})$ was observed. It has to be concluded that reaction 16 is too slow to compete with the radical-radical reactions under the high dose rate conditions.

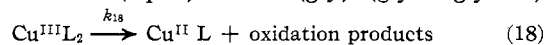
Amino Acid Complexes.—The specific rates of the reaction



(20) I. Fried and D. Meyerstein, *J. Electroanal. Chem.*, **29**, 429 (1971).

were measured by following the formation of the absorptions due to $\text{Cu}^{\text{III}}\text{L}_2$ in solutions saturated with N_2O and containing $(2-10) \times 10^{-5} M \text{Cu}^{\text{II}}\text{L}_2$. (A ratio of $[\text{L}]/[\text{Cu}^{2+}] = 1.9$ was used.) The results are summarized in Table II. It has to be concluded that the effect of the different amino acid ligands on the specific rate of oxidation of Cu^{2+} is similar to that of ethylenediamine.

On the other hand it was found (Table II) that the mechanism of disappearance for all the $\text{Cu}^{\text{III}}\text{L}_2$ complexes obeyed a first-order rate law, the specific rate of reaction being nearly independent of the nature of the amino acid. It should be noted that as the radicals, formed in the decomposition reaction of $\text{Cu}^{\text{III}}\text{L}_2$, also cause transient absorptions, the measurements were carried out at 350 nm where the ratio of the absorption coefficients of the $\text{Cu}(\text{III})$ complexes to the organic radicals is high. Still the accuracy in determining k_{18} is relatively low as the organic radicals formed disappeared at a similar rate. Typical results for the disappearance reaction (eq 18) of $\text{Cu}^{\text{III}}(\text{gly})_2$ (gly = glycine)



are shown in Figures 3 and 4. The specific rate of this

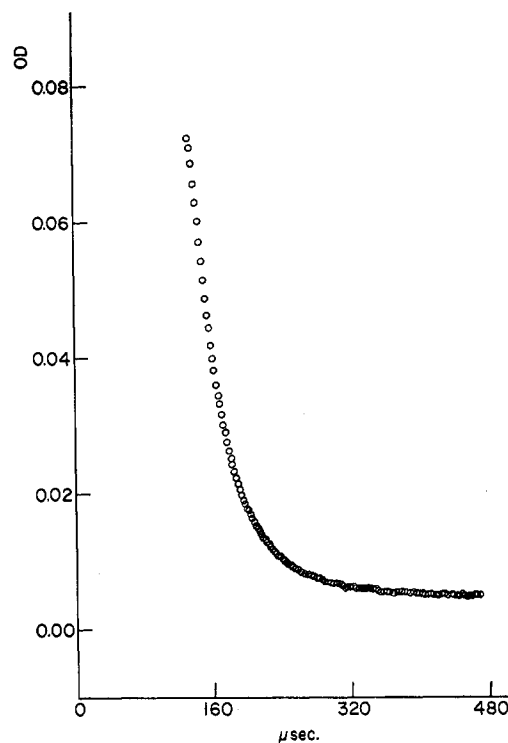


Figure 3.—Optical density vs. time plot for the disappearance of $\text{Cu}^{\text{III}}(\text{gly})_2$. Solution composition: $2 \times 10^{-4} M \text{CuSO}_4$, $3.8 \times 10^{-4} M$ glycine, N_2O saturated, pH 7.5.

reaction was found to be $k_{18} = 2.2 \times 10^4 M^{-1} \text{sec}^{-1}$ at pH 7.5.

In considering the mechanism of reaction 18 several features of the kinetics should be taken into account.

(a) Reaction 18 obeys a first-order rate law for all ligands. The specific rate of reaction depends only slightly on the nature of L, is independent of the concentration of $\text{Cu}^{\text{II}}\text{L}_2$, and increases with increasing pH.

(b) The mechanism of disappearance is clearly different from that observed for $\text{Cu}^{\text{III}}(\text{en})_2$ and copper(III) peptide complexes.¹¹

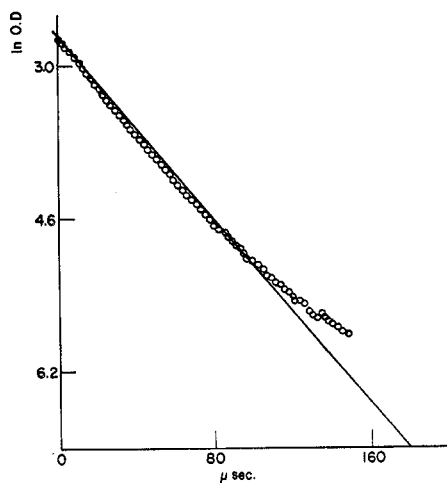
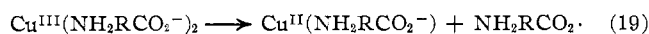


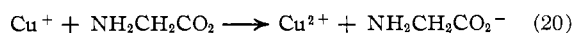
Figure 4.—First-order plot for Figure 3, calculated assuming $OD_{\infty} = 0.05$.

(c) The specific rate of reaction, k_{18} , is independent of the fact whether the amino group is α to or remote from the carboxylic group. k_{18} is also not affected by the absence of a hydrogen on the carbon α to the amino group (α -aminoisobutyric acid). The latter finding is of special interest as it was shown that Cu^{2+} ions do not catalyze the oxidative deamination of the latter amino acid under conditions where they catalyze the deamination of other amino acids.^{7,21}

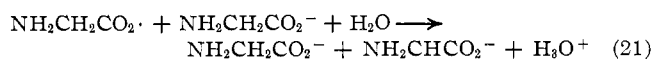
It is suggested therefore that the first and rate-determining step in the decomposition of $Cu^{III}L_2$ is



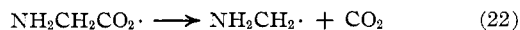
Reaction 19 is an intramolecular electron-transfer reaction. This reaction might be followed by (the equations are given for $NH_2R_2CO_2^- = \text{glycine}$)



or



reactions which have been suggested in the literature.⁶ The reaction

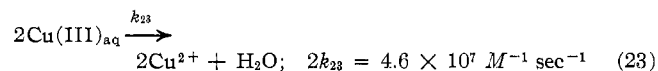


might also contribute and explain the relatively large CO_2 yield, $G(CO_2) = 3.1$, observed in steady-state radiolysis.⁶ This mechanism seems to account for all the observations. Thus it explains why k_{19} (which equals k_{18}) is nearly independent of the amino acid used. However no deamination is expected as long as no hydrogens α to the amino group are available. This mechanism explains also the difference in behavior between amino acids and ethylenediamine or peptides as ligands. It is of interest to note that if the interpretation that the disappearance of $Cu^{III}L_2$ via reaction 19 is correct, then this intramolecular electron-transfer reaction is much slower than expected for a reaction where the oxidizing agent, $Cu(III)$, is directly bound to the reducing agent, $-O_2CR$. It would seem therefore that the only factor which is slowing down the reaction is the Franck-Condon rearrangement; however it seems unlikely that this would take about 100 μsec . It might be

(21) I. Pecht, Ph.D. Thesis, The Weizmann Institute of Science, Rehovoth, Israel, 1967.

argued that this intramolecular electron transfer is slow as it requires a transfer of an electron from a π orbital of the carboxylic group to an e_g orbital on the copper ion. It is well known that the e_g orbital which has a σ symmetry with respect to the metal-ligand bond does, therefore, not overlap effectively with π orbitals of the ligands. It should be mentioned that this is only a speculative attempt to explain the experimental data.

The Reaction between $Cu(III)_{aq}$ and $CuNH_2CH_2CO_2$.—Solutions saturated with N_2O and containing $10^{-3} M$ $CuSO_4$ and 10^{-6} – $10^{-4} M$ glycine were irradiated, and the disappearance of the $Cu(III)$, formed by the pulse, was followed. Under these conditions all the $Cu(III)$ formed in the first stage is $Cu(III)_{aq}$ and not $Cu^{III}gly$. It is known that the mechanism of disappearance of $Cu(III)_{aq}$ in neutral solutions is¹²



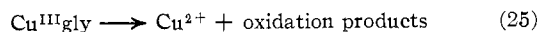
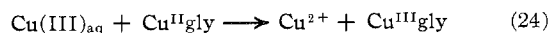
It is also known that no hydroxyl radicals are formed as an intermediate in reaction 23.¹² However it was found that in the presence of small amounts of glycine the mechanism of the disappearance of $Cu(III)$ was first order in glycine. The results obtained are summarized in Table IV. Under the experimental conditions all the

TABLE IV
THE EFFECT OF GLYCINE ON THE SPECIFIC RATE OF DISAPPEARANCE OF $Cu(III)_{aq}$ ^a

[Glycine], M	$k_{obsd},^b \text{sec}^{-1}$	$k(Cu(III)_{aq} + Cu^{II}gly),$ $M^{-1} \text{sec}^{-1}$
0	$<1.0 \times 10^2$...
1×10^{-6}	1.45×10^2	1.45×10^8
2×10^{-6}	2.56×10^2	1.28×10^8
4×10^{-6}	3.3×10^2	8.2×10^7
8×10^{-6}	6.7×10^2	8.4×10^7
1×10^{-5}	7.7×10^2	7.7×10^7
3×10^{-5}	2.4×10^3	8.0×10^7
1×10^{-4}	6.7×10^3	6.7×10^7
$Cu^{II}(gly)_2^c$	6.0×10^3	...

^a All solutions contained $10^{-3} M$ $CuSO_4$ at pH 5.5 and were N_2O saturated. ^b The observed first-order rate. The error limit is $\pm 10\%$. ^c $[CuSO_4] = 5 \times 10^{-4} M$; $[gly] = 9.5 \times 10^{-4} M$; pH 6.1.

glycine in the solution is complexed as $Cu(gly)^+$ and it is suggested that the mechanism of decomposition of $Cu(III)_{aq}$ under these conditions is



Reaction 25 is very similar to reaction 19 and is expected to have a specific rate of $\sim 6 \times 10^3 \text{sec}^{-1}$. Reaction 24 will therefore be the rate-determining step in the disappearance of $Cu(III)_{aq}$. Only at high glycine concentrations is reaction 25 expected to contribute. This might be the cause for the observed decrease in rate at the highest glycine concentration used. The specific rate of reaction 24 thus determined is $k_{24} = (8.1 \pm 0.5) \times 10^7 M^{-1} \text{sec}^{-1}$.

The high specific rate of reaction 24 suggests that the redox potential of the couple $Cu^{II}gly-Cu^{III}gly$ is appreciably lower than that of the couple $Cu(II)_{aq}-Cu(III)_{aq}$. The latter finding is in agreement with other observations concerning different amino complexes of copper.^{11,20,22}

(22) A. Levitzki and M. Anbar, *Chem. Commun.*, 403 (1968).

The relatively high rate of reaction observed for reaction 24, when considered with the relative stability of copper(III) peptide complexes,¹¹ might explain in part the high selectivity in the copper-induced radiolytic deactivation of enzymes.^{9,10} We are planning to extend this study to other copper complexes of biological importance.

Acknowledgments.—I am indebted to Dr. M. S. Matheson for his encouragement and discussions throughout this work. I also wish to thank Mr. W. A. Mulac and Miss C. Goldenberg for their invaluable assistance in all phases of these experiments and Mr. J. Ogdan and the Linac group at Argonne for their careful operation of the linear accelerators.

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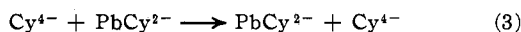
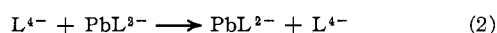
Substituent Effects on the Kinetics of Ligand Substitution Reactions of Lead-Aminocarboxylate Complexes

By JAMES D. CARR* AND DALE R. BAKER

Received December 15, 1970

The rate constants for the ligand exchange reactions of CyDTA with the lead(II) complexes of *l*-PDTA, *d*-PDTA, and CyDTA and the exchange of *d*-PDTA with the lead(II) complex of *l*-PDTA have been measured by polarimetric techniques. Each reaction shows a pathway which is first-order in both entering ligand tetraanion and lead complex concentrations. It is found that one optical isomer of CyDTA exchanges more rapidly with optically active lead-PDTA than does the other. At high pH, the rate of exchange of CyDTA with lead-PDTA is increased and a parallel reaction pathway involving hydroxide ion is proposed. These rate constants are compared with those previously obtained for the exchange of EDTA with the lead(II) complexes of EDTA and PDTA and the steric effects of the alkyl substituents are discussed. The effect of potassium ion on the observed reaction rates is interpreted in terms of known stabilities of potassium complexes. The rate constants obtained are as follows: for *d*-CyDTA with lead-*d*-PDTA or lead-*l*-PDTA, $1.38 \pm 0.02 M^{-1} \text{sec}^{-1}$; for *l*-CyDTA with lead-*l*-PDTA or lead-*d*-PDTA, $2.86 \pm 0.08 M^{-1} \text{sec}^{-1}$; for *d*-PDTA with lead-*l*-PDTA, $1.58 \pm 0.07 M^{-1} \text{sec}^{-1}$; for *d*-CyDTA with lead-*l*-CyDTA, $(3.77 \pm 0.33) \times 10^{-2} M^{-1} \text{sec}^{-1}$.

In a previous work¹ the ligand-exchange kinetics of EDTA² with lead-PDTA and with lead-EDTA were measured. In order to observe further the effect of a change in ligand structure on the rate of exchange, the kinetics of ligand exchange of CyDTA with lead-*l*-PDTA and lead-*d*-PDTA (eq 1, with protons omitted) and *d*-PDTA with lead-*l*-PDTA (eq 2) are measured by observing the change in optical rotation. Also, the exchange of CyDTA with lead-CyDTA (eq 3) is measured by polarimetric techniques.



Ionic charges represent the predominant species of reactants and products in the pH range studied. The pH dependencies of reactions 1–3 are analyzed and interpreted. Activation parameters for reaction 2 are measured. A comparison of reaction rates in potassium ion and cesium ion is interpreted in terms of recently measured stability constants of potassium complexes of CyDTA³ and PDTA⁴ and acidity constants of these ligands in the absence of potassium.

Experimental Section

A standard solution of CyDTA was prepared by dissolving doubly recrystallized CyDTA (Aldrich Chemicals Co., Inc.) by neutralization with KOH or CsOH as appropriate and titrating

(1) J. D. Carr, K. Torrance, C. J. Cruz, and C. N. Reilly, *Anal. Chem.*, **39**, 1858 (1967).

(2) Abbreviations used: EDTA or Y, ethylenediaminetetraacetate; PDTA or L, propylenediaminetetraacetate; CyDTA or Cy, *trans*-1,2-cyclohexanediaminetetraacetate.

(3) J. D. Carr and D. G. Swartzfager, *Anal. Chem.*, **42**, 1238 (1970).

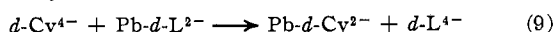
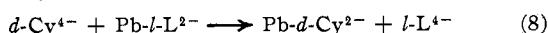
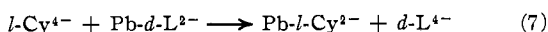
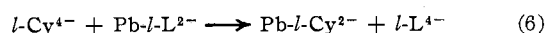
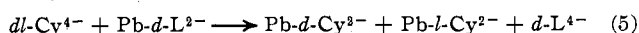
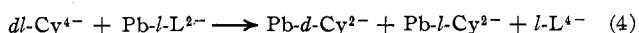
(4) J. D. Carr and D. G. Swartzfager, *ibid.*, **43**, 588 (1971).

with a standard copper solution using SNAZOX indicator.⁵ Solutions of *l*-PDTA, *d*-PDTA, *dl*-PDTA, lead-*d*-PDTA, and lead-*l*-PDTA were prepared as previously described¹ and lead-CyDTA was prepared in the same manner as lead-PDTA. Optically active CyDTA was prepared by resolving *trans*-1,2-diaminocyclohexane with tartaric acid and subsequent reaction with potassium chloroacetate (see ref 6 for details).

Solution ionic strength was maintained at 0.5 with KNO₃ or CsNO₃ as appropriate. Solution pH was adjusted by adding HNO₃, KOH, or CsOH and was measured on a Corning Model 12 pH meter equipped with a Corning semimicro combination electrode. In many cases the potassium ion activity was measured after the kinetic reaction with an Orion potassium specific ion electrode.

Exchange reactions 1–3 were measured at 365 nm on a Perkin-Elmer Model 141 polarimeter with a Sargent Model SR recorder attached as previously described.¹ For the faster reactions a mixing device consisting of two 10-ml syringes and a counter-current mixing block was used to fill the polarimeter cell, allowing measurement of optical rotation in less than 10 sec after mixing. Different initial reactant concentrations were used and in each case a 15- to 60-fold excess of entering ligand was present to force the reaction to completion. Solution pH was measured at the temperature of the kinetic run.

The rate of exchange of reaction 1 was followed polarimetrically by measuring the rates of the reactions



Reactions 4 and 5 were performed in the presence of cesium and

(5) G. Guerrin, M. V. Sheldon, and C. N. Reilly, *Chemist-Analyst*, **49**, 36 (1960).

(6) P. E. Reinbold and K. H. Pearson, *Inorg. Chem.*, **9**, 2325 (1970).