Contribution from the Department of Chemistry, University of Wisconsin, Madison, Wisconsin 53706, the Inorganic Chemistry Research Laboratories, Imperial College of Science & Technology, London SW7 2AY, United Kingdom, and the Department of Chemistry, The Ohio State University, Columbus, Ohio 43210

Metal Ion Catalyzed Decarboxylation. Kinetics and Mechanism of the Oxidative Decarboxylation of Copper(II) Complexes of Aminomalonic Acid in Aqueous Solution

J. H. FITZPATRICK, Jr. and D. HOPGOOD*

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Copper(II) catalyzes the decarboxylation of aminomalonic acid (H_2A) to yield glycine (HG) in aqueous solution. Various pieces of chemical evidence show that the reaction proceeds *via* an oxidative mechanism. The following copper(II)-dependent rate law was observed at 45°: $d[CO_2]/dt = \{k_{13}'' + (k_{13}'' + k_{23}''a_H)[G^-]\}[Cu^{2+}][A^{2-}]$. The terms $k_{13}''[Cu^{2+}][A^{2-}][G^-]$ and $k_{23}''[Cu^{2+}][A^{2-}][G^-]a_H$ account for the marked autocatalysis of the reaction. This rate law is consistent with either an inner-sphere mechanism, $d[CO_2]/dt = k_{A_1}[CuA] + k_{A_2}[CuAG^-] + k_{A_3}[CuAGH]$, or an outer-sphere mechanism, $d[CO_2]/dt = k_{B_1}[Cu^{2+}][A^{2-}] + k_{B_2}[CuG^+][A^{2-}] + k_{B_3}[CuG^+][HA^-]$. In either case the rate of electron transfer from the aminomalonate species to copper(II) is probably rate determining. This step is followed by the decarboxylation of the resulting aminomalonate radical to give the free radical $NH_3^+\dot{C}HCO_2^-$ which oxidizes copper(I) to copper(II) to complete the catalytic cycle. It is of interest that the coordinated glycinato ligand enhances the reactivity of copper(II) in this system.

Introduction

In recent years considerable progress has been made in the elucidation of the roles of copper ions in copper-containing enzymes.[1] The class of enzymes of interest to us are those which are involved in electron-transfer reactions. There is good evidence that in at least three enzymes of this type, namely, laccase, ascorbate oxidase, and ceruloplasmin, reaction with substrates occurs *via* an electron-transfer step from the substrate to copper present as its divalent ion.[1,2]

Of relevance to the enzyme studies are investigations of electron-transfer reactions between copper(II) complexes and organic molecules. Previous work in this area has mainly centered on the well-known copper(II)-catalyzed oxidation of ascorbic acid by oxygen.[3] The first step in this reaction is electron transfer to copper(II) from anions of ascorbic acid.

We are currently investigating copper(II)-catalyzed reactions of organic compounds which proceed *via* an electron-transfer step under anaerobic conditions in aqueous solution. The factors involved in the modification of copper(II) reactivity by coordinated ligands are of especial interest to us and are of direct relevance in comparisons with enzyme studies.

We report here the kinetics and mechanism of copper(II)-catalyzed decarboxylation of aminomalonic acid. Decarboxylation occurs *via* rate-determining one-electron transfer from aminomalonate anions to various copper(II) complexes. This step is followed by rapid decarboxylation of the resulting aminomalonate free radicals and reoxidation of copper(I) by glycine radical.

Experimental Section

Diethylaminomalonate hydrochloride was prepared by Thanassi's method[4] from diethylformamidomalonate (Aldrich Chemical Co.).

Preparation of Potassium Aminomalonate. A 1.9-g sample of diethylaminomalonate hydrochloride was dissolved in 50 ml of 2 *M* KOH, heated on a steam bath for 30 min, and chilled in ice. The pH was adjusted to 6 with 30% acetic acid, 150 ml of 95% ethanol was added, the solution was chilled, and the fine white crystals were filtered off (mp 210–213°, lit.[5] mp 227–230°). This product was recrystallized from a 1:2 water–ethanol mixture (mp 213.5–215°, lit.[5-7] mp 227–230°). The purity of the product was better than 99.9% as shown by its equivalent weight as determined by titration[8] (found, 157.0; calculated for $KC_3H_4NO_4$, 157.2) and by its uptake of 1.001 equiv of HCl at constant pH to yield glycine by decarboxylation in aqueous solution. A concentrated solution of the compound in D_2O exhibited a ^1H nmr spectrum that showed that glycine was present in undetectable amounts (<0.1%).

Glycine (Aldrich Chemical Co.) was recrystallized from a 1:1 water–ethanol mixture. Copper(II) solutions (*ca.* 0.01 *M*) were prepared from recrystallized $Cu(NO_3)_2 \cdot 3H_2O$ (Baker Chemisals) with standardized HCl (*ca.* 10^{-4} *M*) to prevent hydroxide complex formation. These solutions were standardized by EDTA titration. Solutions of metal ions other than copper(II) were prepared from the following Baker reagent grade chemicals: $MgCl_2 \cdot 6H_2O$, $MnCl_2O$, $CoCl_2 \cdot 6H_2O$, $Ni(NO_3)_2 \cdot 6H_2O$, and $Zn(NO_3)_2 \cdot 6H_2O$. These solutions were not standardized. The KNO_3 used for maintaining ionic strength was recrystallized Baker reagent grade chemical. HCl solutions were standardized against NBS Na_2CO_3 and the NaOH solutions were standardized against the HCl solutions. All solutions were prepared with triply distilled water and all reaction solutions were maintained at 45.1° and μ = 0.5 (KNO_3) under an atmosphere of oxygen-free nitrogen.

pH measurements were made with either a Radiometer Model 26 or a Vibret Model 46A pH meter. The meters and their electrodes were standardized at 45° with the following NBS buffer solutions: potassium bitartrate, pH 3.55; potassium hydrogen phthalate, pH 4.045; potassium dihydrogen phosphate–disodium hydrogen phosphate, pH 6.831. pH-Stat titrations were made with the Radiometer pH meter in conjunction with a Radiometer Model ABU-12 autoburet, a Radiometer Model 11 titrator, and a Hewlett-Packard Model 7101 BM strip chart recorder. Titration solutions were contained in a thermostated cell under nitrogen and were stirred magnetically.

Electronic spectra were recorded on either a Cary Model 14 or a Perkin-Elmer Model 402 spectrophotometer. ^1H nmr spectra were obtained with a Varian Associates A-60 spectrometer.

Equilibrium Measurements. Least-squares values of formation constants of both aminomalonic acid and glycine with copper(II) were calculated using the Sillen PITMAP method[9,10] from data obtained

* To whom correspondence should be addressed at the Analytical Research and Methods Development Department, Pfizer, Inc., Groton, Conn. 06340.

(1) R. Malkin and B. G. Malmstrom, *Advan. Enzymol. Relat. Areas Mol. Biol.*, 33, 177 (1970).

(2) L. Broman, B. G. Malmstrom, R. Aasa, and T. Vanngard, *Biochim. Biophys. Acta*, 75, 365 (1963).

(3) E. S. G. Barron, R. DeMeio, and F. W. Klemperer, *J. Biol. Chem.*, 112, 625 (1936); C. M. Lyman, M. O. Schultze, and C. G. King, *ibid.*, 118, 757 (1937); A. O. Dekker and R. G. Dickinson, *J. Amer. Chem. Soc.*, 62, 2165 (1940); A. Weissberger and J. E. Luvalle, *ibid.*, 66, 700 (1944); H. Nord, *Acta Chem. Scand.*, 9, 442 (1955).

(4) J. W. Thanassi, *Biochemistry*, 9, 525 (1970), and references cited therein.

(5) M. Matthew and A. Neuberger, *Biochem. J.*, 87, 601 (1963).

(6) The melting point of a salt of a carboxylic acid is an unreliable indicator of purity.[7]

(7) O. Kamm, "Qualitative Organic Analysis," 2nd ed, Wiley, New York, N. Y., 1932.

(8) M. S. Dunn and A. Loshakoff, *J. Biol. Chem.*, 113, 359 (1936).

(9) L. G. Sillen, *Acta Chem. Scand.*, 16. 159 (1962); N. Ingri and L. G. Sillen, *ibid.*, 16, 173 (1962). The FORTRAN computer program used was based on a version of PITMAP written by Tobias and Yasuda.[10]

(10) R. S. Tobias and M. Yasuda, *Inorg. Chem.*, 2, 1307 (1963).

by potentiometric (pH) titration of standardized copper(II)-ligand solutions with standard NaOH solution. In the case of aminomalonic acid, titrations were extended to the low pH of 2 to enable determination of the formation constants of complexes of protonated ligands.

Similar titrations were carried out on standardized copper(II)-aminomalonic acid-glycine solutions to obtain the formation constants of complexes containing both ligands. The PITMAP method was used to analyze the data in a way similar to that described by Leussing and Huq.[11] No significant amount of decarboxylation occurred during the time scale of these measurements.

Kinetic Measurements. Decarboxylation of aminomalonic acid results in the uptake of one proton in the pH region *ca.* 3.5-5.5

$$NH_3^+CH(CO_2^-)_2 + H^+ \rightarrow NH_3^+CH_2CO_2^- + CO_2$$

At pH >5.5 the stoichiometry becomes increasingly less than one proton due to the formation of HCO_3^- and CO_3^{2-} ions. The stoichiometry also becomes increasingly less than one proton at pH <3.5 due to the protonation of the carboxylate groups

$$NH_3^+CH(CO_2H)CO_2^- \rightarrow NH_3^+CH_2CO_2^- + CO_2$$

We measured the rates of aminomalonic acid decarboxylation under various conditions by using a pH-Stat technique. This involved adding standardized HCl to the reaction solutions from an autoburet so as to maintain a constant pH. The resulting volume of HCl *vs.* time curves were then analyzed (*vide infra*) to obtain the rate law and rate constants. This method is very accurate and has the advantage over the usual Warburg technique that the volume of evolved CO_2 is not measured. However it suffers from the disadvantage that rates at pH values less than about 3 cannot be measured due both to the protonation of a carboxylate group and to the buffering action of the solvent. In this work precise kinetic measurements were restricted to the narrow range of pH 3.5-4.5. This upper limit of pH was chosen to prevent the precipitation of metal(II) aminomalonate complexes. The analytical concentrations of reactants and the pH's of kinetic runs are given in the text and Table I.

Results and Discussion

Decarboxylation of Aminomalonic Acid in the Presence of Labile Divalent Metal Ions. Aminomalonic acid decarboxylates to yield glycine in aqueous solution[4]

$$NH_2CH(CO_2H)_2 \rightarrow NH_2CH_2CO_2H + CO_2$$

Thanassi[4] investigated the kinetics and mechanism of this reaction by measuring rates of CO_2 evolution. He reported a rate law[12]

$$d[CO_2]/dt = k_1[H_3A^+] + k_2[H_2A]$$

where $k_1 = 1.27 \times 10^{-5}$ sec^{-1} and $k_2 = 3.3 \times 10^{-5}$ sec^{-1} at 45° and $\mu = 0.5$ (KCl). For convenience we used similar conditions (45.1°, $\mu = 0.5$ (KNO_3)) to investigate the effect of labile divalent metal ions on the decarboxylation rates of aminomalonic acid. KNO_3 was used to maintain ionic strength to minimize complex formation by the anion of the potassium salt.[13] In addition the reaction solutions were maintained in an oxygen-free nitrogen atmosphere. Under these similar conditions we measured $k_2 = 6.48 \times 10^{-5}$ sec^{-1} by the pH-Stat titration method. The upper limit of the rate constant for decarboxylation of HA^- was estimated to be about 5×10^{-7} sec^{-1}.

The effects of the divalent ions of magnesium, manganese, cobalt, nickel, copper, and zinc on the initial rate of decarboxylation of aminomalonic acid were measured at pH 4. Except for the magnesium(II) run ($[Mg^{2+}]_T = 3.8 \times 10^{-4}$ *M*, $[A]_T = 7.8 \times 10^{-3}$ *M*), all runs were carried out with aminomalonic acid concentrations which were about 200 times greater than the metal ion concentrations ($[M^{2+}]_T = (3.8\text{-}4.0) \times 10^{-5}$ *M*, $[A]_T = (7.9\text{-}8.1) \times 10^{-3}$ *M*) to prevent precipitation of the neutral MA complexes. Under these conditions copper(II) was the only metal ion which caused any significant rate enhancement. This and other kinetic runs in the presence of copper(II) (*vide infra*) exhibited sigmoid-shaped volume-time curves in contrast to the logarithmic curves given in the presence of the other divalent metal ions. However the magnitude of the rate enhancement was small being at most a factor of 5. Thus the effect due to copper(II) could result from either acid-base or redox mechanisms.

The stoichiometry of the copper(II)-catalyzed reaction is identical with that of the spontaneous decarboxylation reaction. One equivalent of protons was taken up at pH 4 and the yield of glycine was quantitative

$$NH_3^+CH(CO_2^-)_2 + H^+ \rightarrow NH_3^+CH_2CO_2^- + CO_2$$

The role of copper(II) is strictly catalytic as shown by the invariant analytical concentration of copper(II) throughout the course of a run. These copper(II) concentrations were determined spectrophotometrically in the region 210-250 nm by using the method due to Spies.[14]

Chemical evidence was obtained which supports an oxidative mechanism.

(i) Under conditions of high copper(II) concentration ($[Cu^{2+}]_T = 0.01$ *M*, $[A]_T = 0.01$ *M*) the initial rapid precipitation of copper(II) aminomalonate (CuA) was followed by its dissolution and the concomitant precipitation of copper(I) oxide. This phenomenon was noted by Baeyer.[15] The stoichiometry of this one-electron copper(II) reduction was measured under the above concentration conditions at pH 5.6 and 45°. The rate of formation of copper(I) as Cu_2O precipitate was about 80% of the decarboxylation rate. The organic product(s) of the reaction was (were) not identified. However very little glycine was formed under these conditions.

(ii) Cupric ferrocyanide, $Cu^{II}_2Fe^{II}(CN)_6$, was precipitated from runs containing potassium ferricyanide ($[Cu^{2+}]_T = 2.0 \times 10^{-4}$ *M*, $[A]_T = 2.0 \times 10^{-2}$ *M*, $[Fe^{III}(CN)_6^{3-}] = 10^{-4}$-$10^{-2}$ *M*, pH 4.0, and 45°). Initial rate measurements were made on this system with $[Fe^{III}(CN)_6^{3-}] = 9.0 \times 10^{-4}$ *M*. The rate of decrease of ferricyanide was measured spectrophotometrically at 420 nm[16] while the rate of decarboxylation was measured by pH-Stat titration. The initial rate of reaction of $Fe^{III}(CN)_6^{3-}$ (7.6×10^{-9} mol l.$^{-1}$ sec^{-1}) was about 8% of the initial decarboxylation rate (1.0×10^{-7} mol l.$^{-1}$ sec^{-1}; the copper(II)-dependent part of this rate is 5.3×10^{-8} mol l.$^{-1}$ sec^{-1}).

(iii) Polyvinyl cyanide was precipitated during a run carried out under kinetic run conditions ($[Cu^{2+}]_T = 4 \times 10^{-5}$ *M*, $[A]_T = 8.0 \times 10^{-3}$ *M*, pH 4.0, 45°) in the presence of acrylonitrile (1.0×10^{-2} *M*). Thus the overall mechanism probably involves electron transfer from an aminomalonate carboxylate anion to copper(II) to yield copper(I) and an aminomalonate radical. This radical probably decarboxylates to give a glycinato radical, $NH_3^+\dot{C}HCO_2^-$. Vinyl polymerization could be initiated by either of these free radicals.[17] It is of mechanistic significance that initiation of polymerization occurs under these conditions considering that copper(II) is an excellent inhibitor[18] of vinyl polymerization.

(11) D. L. Leussing and N. Huq, *Anal. Chem.*, 38, 1388 (1966).

(12) The following abbreviations are used: $A^{2-} \equiv$ aminomalonate dianion, $NH_2CH(CO_2^-)_2$; $G^- \equiv$ glycinato anion, $NH_2CH_2CO_2^-$.

(13) *Chem. Soc., Spec. Publ.*, No. 17 (1964); No. 25 (1971).

(14) J. R. Spies, *J. Biol. Chem.*, 195, 65 (1952).

(15) A. Baeyer, *Justus Liebigs Ann. Chem.*, 131, 291 (1864).

(16) A. W. Adamson, *J. Phys. Chem.*, 56, 860 (1952).

(17) C. H. Bamford, W. G. Barb, A. D. Jenkins, and P. F. Onyon, "The Kinetics of Vinyl Polymerisation by Radical Mechanisms," Butterworths, London, 1958.

(18) J. K. Kochi, *Rec. Chem. Progr.*, 27, 207 (1966), and references cited therein.

Table I. Rate Constants for Copper(II)-Catalyzed Decarboxylation of Aminomalonic Acid at 45.1° and $\mu = 0.5$ (KNO_3)

Run	pH	10^5-$[Cu^{2+}]_T$, M	$10^3 \times [A]_{T(initial)}$, M	$10^{-4}k_{M_1}$, l. mol^{-1} sec^{-1}	$10^{-10}k_{M_2}$, $l.^2$ mol^{-2} sec^{-1}	$10^{-12}k_{M_3}$, $l.^2$ mol^{-2} sec^{-1}	σ_v, cm^3	V_∞, cm^3	$\sigma_v \times 100\%/V_\infty$
1	3.510	3.96	7.98	1.13 ± 0.04	0.2 ± 0.7	13.9 ± 0.1	0.0025	2.02	0.12
2	3.590	3.90	7.91	2.64 ± 0.05	0.2 ± 1.1	13.2 ± 0.1	0.0077	4.07	0.19
3	3.800	3.94	8.37	2.56 ± 0.24	17 ± 5	9.05 ± 0.007	0.0040	4.27	0.09
4	4.010	3.97	15.9	3.19 ± 0.21	11.7 ± 1.3	1.75 ± 0.05	0.0032	4.02	0.08
5	4.020	4.00	7.94	2.75 ± 0.35	3 ± 4	7.15 ± 0.22	0.015	1.98	0.76
6	4.070	4.09	16.1	3.8 ± 0.5	14.6 ± 2.7	5.12 ± 0.17	0.0064	3.90	0.16
7	4.420	4.01	7.90	1.64 ± 0.27	3.7 ± 1.2	5.23 ± 0.23	0.0076	1.97	0.38
8	4.460	4.01	16.2	1.6 ± 1.1	3 ± 2	4.48 ± 0.20	0.0041	4.04	0.10

Hence, under the homogeneous conditions of the kinetic runs (*vide infra*), the overall steps in the reaction may be represented as

$$Cu^{2+} + NH_3^+CH(CO_2^-)_2 \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} Cu^+ + NH_3^+CH(CO_2\cdot)(CO_2^-)$$

$$NH_3^+CH(CO_2\cdot)(CO_2^-) \xrightarrow{k_2} NH_3^+\dot{C}HCO_2^- + CO_2$$

$$Cu^+ + NH_3^+\dot{C}HCO_2^- + H^+ \xrightarrow{k_3} Cu^{2+} + NH_3^+CH_2CO_2^-$$

The rate-determining step of the overall reaction is undoubtedly either that of electron transfer or that of decarboxylation; otherwise the analytical concentration of copper(II) would decrease during the course of the run. The most likely possibility is that slow electron-transfer is followed by fast decarboxylation of the carboxylate radical ($k_2 \gg k_1$). The extensive work by Kochi and coworkers[18-24] indicates that the decarboxylation of carboxylate radicals is fast compared with their rates of formation by oxidation of carboxylate anions by metal ion complexes of silver(II),[19] manganese(III),[20] cobalt(III),[21] lead(IV),[22] and cerium(IV).[23] It is of significance that, although copper(II) oxidizes alkyl radicals formed by oxidative decarboxylation of acids by lead(IV), it does not itself appear to effect oxidative decarboxylation of these acids.[24] This mechanism is consistent with the precipitation of Cu_2O at high initial concentrations of the reactants.[25] Cu_2O precipitation will occur when the copper(I) concentration exceeds the concentration governed by the solubility product of Cu_2O[13] and the stability constants of the copper(I) complexes present in solution. The glycinato radicals which are in excess of the amount of copper(I) in solution will undergo reaction(s) to yield product(s) other than glycine as was observed.

The possibility that nitrate ion might be mechanistically involved in the reaction, *e.g.*, by oxidizing copper(I), was investigated by carrying out a run with $CuCl_2$ and KCl. The stoichiometry of this run was identical with those containing KNO_3 which indicates that nitrate ion is not of primary mechanistic importance.

It is of interest that the oxidation of aminomalonic acid by iodine yields mesoxalic acid[15]

$$NH_2CH(CO_2H)_2 + I_2 + H_2O \rightarrow OC(CO_2H)_2 + NH_4I + HI$$

In this case the mechanism possibly proceeds *via* the adduct formed between the amine group and iodine: $I_2 \cdot NH_2CH(CO_2H)_2$.

Equilibria in the Copper(II)-Aminomalonic Acid-Glycine Systems at 45.1° and $\mu = 0.5$ (KNO_3). (i) Ligand Acid Dissociation Constants. The acid dissociation constants of a ligand (H_nL) are defined as

$$K_{aL_i} = [H_{n-i}L]a_H/[H_{n-i+1}L]$$

where a_H is the hydrogen ion activity as measured by the glass electrode.

The first acid dissociation constant of aminomalonic acid was not measurable by potentiometric titration but we measured $pK_{aA_2} = 2.59$ and $pK_{aA_3} = 8.94$. Thanassi[4] reported $pK_{aA_1} = 0.7$ (this was obtained from kinetic data) and $pK_{aA_2} = 3.0$ (30°). The constants for glycine are $pK_{aG_1} = 2.30$ and $pK_{aG_2} = 9.23$.

(ii) Stability Constants. We report cumulative stability constants, β_{jkl}, for the reactions

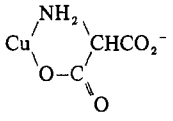
$$Cu^{2+} + jA^{2-} + kG^- + lH^+ \rightleftharpoons CuA_jG_kH_l^{2-2j-k+l}$$

where

$$\beta_{jkl} = \frac{[CuA_jG_kH_l^{2-2j-k+l}]}{[Cu^{2+}][A^{2-}]^j[G^-]^k a_H{}^l}$$

(a) Copper(II)-Aminomalonic Acid. The complexes CuA, CuA_2^{2-}, $CuAH^+$, and CuA_2H_2 were present in titrated solutions in detectable concentrations. Their log stability constants are $\beta_{100} = 7.77 \pm 0.04$, $\beta_{200} = 14.26 \pm 0.07$, $\beta_{101} = 11.43 \pm 0.07$, and $\beta_{202} = 20.5 \pm 0.8$. The feasible complexes CuA_3^{4-} and CuA_2H^- were not detected. Schwarzenbach, *et al.*,[26] studied the aminomalonic acid systems with the divalent ions of calcium, barium, and zinc. They reported log β_{100} values of 2.51 for calcium(II) and 6.48 for zinc(II).

The binary complexes CuA and CuA_2^{2-} probably contain the ligand bonded predominantly as a bidentate chelate (I)



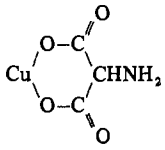
I

through the amino and one of the carboxylate groups. The constants β_{100} and β_{200} are only about twice as large as the corresponding constants (β_{010} and β_{020}) for glycine (*vide infra*). Thus an appreciable degree of tridentate bonding is not likely.

(19) J. M. Anderson and J. K. Kochi, *J. Org. Chem.*, **35**, 986 (1970).
(20) J. M. Anderson and J. K. Kochi, *J. Amer. Chem. Soc.*, **92**, 2450 (1970).
(21) S. S. Lande and J. K. Kochi, *J. Amer. Chem. Soc.*, **90**, 5196 (1968).
(22) J. K. Kochi, J. D. Bacha, and T. W. Bethea, *J. Amer. Chem. Soc.*, **89**, 6538 (1967).
(23) R. A. Sheldon and J. K. Kochi, *J. Amer. Chem. Soc.*, **90**, 6688 (1968).
(24) J. Bacha and J. K. Kochi, *J. Org. Chem.*, **33**, 83 (1968).
(25) If $k_2 >> k_{-1}$, then $[Cu^+] \propto ((k_1/k_3)[Cu^{2+}]_T[A]_T)^{1/2}$ in the absence of Cu_2O precipitation.
(26) G. Schwarzenbach, E. Kampitsch, and R. Steiner, *Helv. Chim. Acta*, **28**, 1133 (1945).

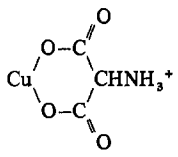
The electronic spectrum of CuA_2^{2-} (λ_{max} 628 nm, ϵ_{max} 48) is closely similar to that of CuG_2[27] (λ_{max} 632 nm, ϵ_{max} 45). The band maximum and the molar absorptivity in the visible region are very sensitive indicators of the degree of tetragonal distortion of a copper(II) complex.[27] The spectra of CuA_2^{2-} and CuG_2 indicate that both are very tetragonally distorted[27] and that the structure of CuA_2^{2-} must be very similar to that of CuG_2. Thus the ligands in CuA_2^{2-} are bonded in a bidentate square-planar structure with little or no interaction between the free carboxylate groups and the axial sites.[27]

The alternative mode of bidentate bonding through the two carboxylate groups (II) is ruled out both by the visible spectrum of CuA_2^{2-} and by stability constant arguments. The stability constant of copper(II) malonate[28] is lower than β_{100} by a factor of over 100 while the basicities of the malonic acid carboxylate groups[28] are considerably higher than those of aminomalonic acid ($\log \beta_{CuMal} = 5.55$, $pK_{a1} = 2.85$, $pK_{a2} = 5.66$).

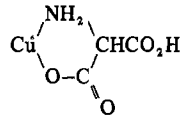


II

The complexes $CuAH^+$ and CuA_2H_2 undoubtedly contain the species HA^- as ligand. The most likely structures are those of types III and IV, which coexist in labile equilibrium.



III



IV

The ratio of type III to type IV concentrations for $CuAH^+$ is easily estimated by using reasonable values of the unknown acid dissociation constants of the two coordinated HA^- ligands (K_{aIII} and K_{aIV}) and for the stability constant (β'_{100}) of the complex CuA of type II structure. This ratio is given by $[CuAH^+_{III}]/[CuAH^+_{IV}] = \beta'_{100}K_{aIV}/\beta_{100}K_{aIII}$. Reasonable ranges[13] for these constants are for pK_{aIII}, 9-10, for pK_{aIV}, 1-2, and for $\log \beta'_{100}$, 2-3, and these give a range for $[CuAH^+_{III}]/[CuAH^+_{IV}]$ of roughly $10-10^4$. Thus HA^- probably coordinates predominantly as a type III ligand.

The stability constants for the reactions

$$Cu^{2+} + HA^- \overset{\beta_1}{\rightleftharpoons} CuAH^+$$

$$Cu^{2+} + 2HA^- \overset{\beta_2}{\rightleftharpoons} CuA_2H_2$$

are $3.1 \times 10^2\ M^{-1}$ and about $4 \times 10^2\ M^{-2}$, respectively ($\beta_1 = \beta_{101}K_{aA2}$, $\beta_2 = \beta_{202}K_{aA2}^2$). These values are consistent with structures of type III.[13]

The mixed-ligand complex CuA_2H^- ($CuA(AH)^-$) was not detected but it may be kinetically significant. A value for β_{201} was estimated from the relationship[29] $\beta_{201} = \beta_{100}\beta_{202}/\beta_{101} + \beta_{101}\beta_{200}/\beta_{100}$ which gives $\log \beta_{201} = 17.97$

(b) Copper(II)-Glycine. This system has received extensive study.[13] The log stability constants for the three complexes CuG^+, CuG_2, and CuG_3^- are $\beta_{010} = 7.53 \pm 0.04$, $\beta_{020} = 13.94 \pm 0.06$, and $\beta_{030} = 16.4 \pm 0.3$. These values are similar to those reported in the literature.[13]

(c) Copper(II)-Aminomalonic Acid-Glycine. The only mixed aminomalonate-glycinato complex detected was $CuAG^-$ with $\log \beta_{110} = 13.3 \pm 0.3$. β_{110} is about a factor of 10 lower than the value estimated from the relationship $\beta_{110} = \beta_{100}\beta_{020}/\beta_{010} + \beta_{010}\beta_{200}/\beta_{100}$ which is $\log \beta_{110} = 14.41$. Similar discrepancies have been observed in mixed complexes of nickel(II).[30] The log stability constant for the complex CuAGH (Cu(AH)G) of 17.86 was estimated from the relationship[29] $\beta_{111} = \beta_{010}\beta_{202}/\beta_{101} + \beta_{101}\beta_{020}/\beta_{010}$.

Kinetics and Mechanism of Copper(II)-Catalyzed Decarboxylation of Aminomalonic Acid at 45.1° and $\mu = 0.5$ (KNO_3). Kinetic data for the copper(II)-catalyzed decarboxylation of aminomalonic acid were obtained by pH-Stat titration of reaction solutions with standardized HCl. The volume-time curves for each run cannot be analytically analyzed because it is not possible to obtain functional integrals from feasible differential rate laws. This is because the distribution of labile species is continuously changing during the course of a run. The changing concentrations of the ten copper(II) complexes CuA, CuA_2^{2-}, $CuAH^+$, CuA_2H^-, CuA_2H_2, CuG^+, CuG_2, CuG_3^-, $CuAG^-$, and CuAGH and the species A^{2-}, HA^-, H_2A, G^-, and HG have to be taken into account in any comprehensive kinetic analysis. There is no significant buildup of copper(I) complexes so that their concentrations are of no significance.

The rates of decarboxylation are very much slower than the rates involved in the redistribution of labile species and so their concentrations are always at equilibrium. At any desired time during the course of a run the analytical concentrations of copper(II), aminomalonic acid, and glycine were calculated from their initial concentrations and the volume of added HCl. The species distribution was then iteratively determined from the solution pH and the three mass balance equations for the analytical concentrations of copper(II), aminomalonic acid, and glycine by using a Newton-Raphson algorithm[31] (see Appendix I).

The least-squares values of rate constants in feasible differential rate laws were evaluated numerically by using a Runge-Kutta algorithm[32] (see Appendix II). By using suitable estimates of the rate constants a theoretical volume-time curve was calculated from the rate law being tested. The least-squares rate constants were then obtained by iteration. The Sillen PITMAP method[9] was used to minimize $\Sigma_{t=1}{}^{n}R_t^2$ (t is the time from the start of the run and the residual $R_t = V_{obsd} - V_{calcd}$, where V is the volume of HCl added at time t) by functional iteration of the rate constants. No weighting of V_{obsd} was used as the errors in the value of HCl delivered by the autoburet were virtually constant over its entire range.

The initial analysis involved testing various rate laws of the type

$$d[CO_2]/dt = k_2[H_2A] + \sum_{j=1}^{3}\sum_{k=0}^{2} k_j[Cu^{2+}][A^{2-}]_j[G^-]_k$$

for each kinetic run taken by itself. The hydrogen ion activity was not included in the copper(II)-dependent terms because each run was carried out at constant pH. The only significant non-copper(II)-dependent pathway under our con-

(27) C. K. Jorgensen, *Acta Chem. Scand.*, 9, 1362 (1955).
(28) D. I. Stock and D. W. Davies, *J. Chem. Soc.*, 1371 (1949).
(29) K. S. Bai and D. L. Leussing, *J. Amer. Chem. Soc.*, 89, 6126 (1967).
(30) D. D. Perrin and V. S. Sharma, *J. Chem. Soc. A*, 446 (1968).
(31) A. D. Booth, "Numerical Methods," 2nd ed, Butterworths, London, 1957, p 154.
(32) See ref 31, p 61.

ditions was due to decarboxylation of H_2A ($k_2 = 6.48 \times 10^{-5}$ sec^{-1}). For each run the lowest value of ΣR_t^2 was obtained with the rate law

$$d[CO_2]/dt = k_2[H_2A] + [Cu^{2+}][A^{2-}]\{k_{M1} + k_{M2}[A^{2-}] + k_{M3}[G^-]\}$$

The term $k_{M3}[Cu^{2+}][A^{2-}][G^-]$ accounts for the pronounced sigmoid or autocatalytic shapes of the volume-time curves. Values of the rate constants are given in Table I together with their standard deviations and the standard error of the fit σ_v ($\sigma_v = \sqrt{\Sigma R_t^2/(n-m)}$, where n is the number of data points and m is the number of rate constants being refined). Exceptionally good fits of theoretical to observed volume-time curves were obtained as shown by the very precise statistical parameters. Note especially the very small values of $(\sigma_v/V_\infty) \times 100\%$ which lie in the range of 0.08–0.76%. However it does not of course follow from this that the rate constants themselves are very accurate. The errors in the large number of equilibrium constants which were used in the analysis propagate themselves in a not easily determinable way. In particular the values of two important constants, β_{201} and β_{111}, were estimated from relationships of the type which were not very successful in predicting the value of β_{110}. To compound further the problem of error evaluation the kinetic runs were of necessity carried out with $[A]_T/[Cu^{2+}]_T$ ratios which were much larger than those used to obtain equilibrium data. Thus there may be complexes, *e.g.*, $CuA_2GH_2^-$, which were not taken into account but which are of significant concentration.

The pH dependence of each of the rate constants k_{Mi} was statistically determined using the values given in Table I. Each value was weighted[33] by a factor of $1/\sigma^2$ in evaluating the least-squares values of k'_{1i} and k'_{2i} from the relationship $k_{Mi} = k'_{1i} + k'_{2i}a_H$. These data give $k'_{11} = (1.80 \pm 0.06) \times 10^4$ l. mol^{-1} sec^{-1}, $k'_{12} = (2.8 \pm 0.4) \times 10^{10}$ l.^{2} mol^{-2} sec^{-1}, and $k'_{23} = (4.61 \pm 0.17) \times 10^{16}$ l.^{3} mol^{-3} sec^{-1}. The constant k'_{13} and the proton-dependent constants k'_{21} and k'_{22} were not statistically significant.

A comprehensive method of data analysis was used to simultaneously analyze all the volume-time curves.[34] A Runge-Kutta algorithm similar to that used for each individual run was used to evaluate rate laws of the type

$$d[CO_2]/dt = k_2[H_2A] + \sum_{j=1}^{2}\sum_{k=0}^{1}[Cu^{2+}][A^{2-}]_j[G^-]_k \times (k''_{1i} + k''_{2i}a_H)$$

The best statistical parameters were obtained with the rate law

$$d[CO_2]/dt = k_2[H_2A] + [Cu^{2+}][A^{2-}]\{k''_{11} + (k''_{13} + k''_{23}a_H)[G^-]\}$$

with $k''_{11} = (4.0 \pm 0.2) \times 10^4$ l. mol^{-1} sec^{-1}, $k''_{13} = (1.1 \pm 0.3) \times 10^{12}$ l.^{2} mol^{-2} sec^{-1}, and $k''_{23} = (2.8 \pm 0.3) \times 10^{16}$ l.^{3} mol^{-3} sec^{-1}. This combined analysis gave $\sigma_v = 0.059$.

Two terms are common to both methods of data analysis, *i.e.*, those involving k_{11} and k_{23}. The values of these rate constants are in fair agreement: k''_{11} is twice as large as k'_{11} while k''_{23} is 60% of k'_{23}.

The second, combined run, analysis is statistically coherent and the results of it will be used in the following discussion of mechanism.

Ruling out unlikely termolecular mechanisms in aqueous solution, the copper(II)-dependent part of the rate law almost certainly results from rapid preequilibrium between some or all of the reactants followed by rate-determining uni- or bimolecular decarboxylation. Each of the three copper(II)-dependent terms has the common factor of $[Cu^{2+}]\cdot[A^{2-}]$. Thus the rate law is consistent with either one of the following two mechanisms (or combinations of these) being most feasible.

Mechanism A. Decarboxylation results from the rate-determining first-order reaction of the complexes CuA + $CuAG^-$ + CuAGH

$$d[CO_2]_{Cu}/dt = k_{A1}[CuA] + k_{A2}[CuAG^-] + k_{A3}[CuAGH]$$

with $k_{A1} = k''_{11}/\beta_{100} = 6.8 \times 10^{-4}$ sec^{-1}, $k_{A2} = k''_{13}/\beta_{110} = 5.6 \times 10^{-2}$ sec^{-1}, and $k_{A3} = k''_{23}/\beta_{111} = 3.8 \times 10^{-2}$ sec^{-1}.

Mechanism B. Decarboxylation results from the rate-determining second-order reaction between either A^{2-} or HA^- and various copper(II) species

$$d[CO_2]_{Cu}/dt = k_{B1}[Cu^{2+}][A^2] + k_{B2}[CuG^+][A^{2-}] + k_{B3}[CuG^+][HA^-]$$

with $k_{B1} = k''_{11} = 4.0 \times 10^4$ l. mol^{-1} sec^{-1}, $k_{B2} = k''_{13}/\beta_{010} = 3.3 \times 10^4$ l. mol^{-1} sec^{-1}, and $k_{B3} = k''_{23}K_{aA3}/\beta_{010} = 2.1 \times 10^6$ l. mol^{-1} sec^{-1}.

A third alternative is possible for the term k''_{23}: acid-catalyzed decarboxylation of the $CuAG^-$ complex, $k''_{23}\cdot[Cu^{2+}][A^{2-}][G^-]a_H = k_{C3}[CuAG^-]a_H$ ($k_{C3} = k''_{23}/\beta_{110} = 1.4 \times 10^3$ l. mol^{-1} sec^{-1}). This mechanism is unlikely for the same reasons Thanassi[4] used against acid-catalyzed pathways in the spontaneous decarboxylation.

The pathways of mechanisms A and B are closely similar to each other in that each mechanism involves electron-transfer from aminomalonate to copper(II). Thus the outer-sphere term $k_{B1}[Cu^{2+}][A^{2-}]$ has its inner-sphere analog $k_{A1}\cdot[CuA]$.

Discussing mechanism B first, this would probably involve fast preequilibria to yield solvent-separated ion pairs followed by rate-determining outer-sphere electron transfer; *e.g.*, the term $k_{B1}[Cu^{2+}][A^{2-}]$ would result from the reactions

$$Cu^{2+} + A^{2-} \overset{K_{01}}{\rightleftharpoons} (Cu^{2+})(A^{2-}) \overset{k'_{B1}}{\rightleftharpoons} (Cu^{+})(A^{\cdot-}) \xrightarrow[-CO_2]{\text{fast}}$$

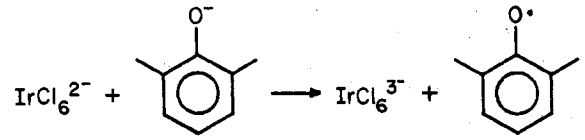
This model gives the relationship $k_{Bi} = K_{0i}k'_{Bi}$ ($i = 1$–3) where the K_{0i} terms are the ion-pair equilibrium constants and the k'_{Bi} are the outer-sphere electron-transfer rate constants for the ion pairs $(Cu^{2+})(A^{2-})$, $(CuG^+)(A^{2-})$, and $(CuG^+)(HA^-)$, respectively. Values of $K_{01} = 4.9\ M^{-1}$, $K_{02} = 1.2\ M^{-1}$, and $K_{03} = 0.6\ M^{-1}$ were estimated from the Fuoss equation[35] by assuming that the ions are spherically symmetrical and that their closest approach in the ion pair is 5 Å.[36] These values give outer-sphere electron-transfer rate constants of $k'_{B1} = 8.2 \times 10^3$ sec^{-1}, $k'_{B2} = 2.7 \times 10^4$ sec^{-1} and $k'_{B3} = 3.5 \times 10^6$ sec^{-1}.

The absence of terms $k_{B4}[CuA][A^{2-}]$ and $k_{B5}[CuA][HA^-]$, which are reasonable because CuA and CuG^+ have very similar electronic structures, is consistent with mechanism B. The ion-pair equilibrium constants, K_{04} and K_{05}, are proba-

(33) P. R. Bevington, "Data Reduction and Error Analysis for the Physical Sciences," McGraw-Hill, New York, N. Y., 1969, p 130.
(34) The function $\Sigma_{run_1}^{8}\ \Sigma_{t=1}^{n} R_t^2$ was minimized.
(35) R. M. Fuoss, *J. Amer. Chem. Soc.*, 80, 5059 (1958).
(36) G. G. Hammes and J. I. Steinfeld, *J. Amer. Chem. Soc.*, 84, 4639 (1962).

bly much smaller than the corresponding constants K_{02} and K_{03} because CuA is a neutral species.[37]

The values of rate constants derived for mechanism B are consistent with rate constants for demonstrated outer-sphere electron-transfer steps between metal complexes and organic compounds. Indeed these cover an extremely wide range[38] and can approach the diffusion-controlled limit as was found, for example, in the oxidation of the 2,6-dimethylphenoxide ion by hexachloroiridate ion where the rate constant for the electron-transfer step

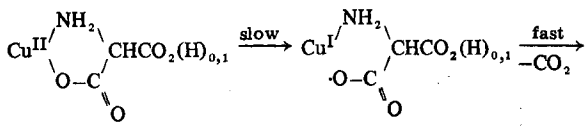


is 8.3×10^8 l. mol^{-1} sec^{-1} at 20°.[39]

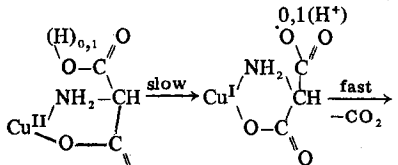
Mechanism A involves first-order decarboxylation of the coordinated A^{2-} and HA^- ligands; the rate constants k_{Ai} (i = 1-3) can be directly compared with the first-order rate constants for the decarboxylation of aminomalonic acid. The complexes CuA and $CuAG^-$ can be compared with H_2A as all these species involve Lewis acid (H^+ or Cu^{2+}) bonding to one amine group and one carboxylate anion of the aminomalonate dianion. Similarly CuAGH may be compared with H_3A^+. Using Thanassi's values[4] for k_1 and k_2 gives $k_{A1}/k_2 = 21$, $k_{A2}/k_2 = 170$, and $k_{A3}/k_1 = 300$.

The predominant forms of bonding of the A^{2-} and HA^- ligands in these complexes are most likely of types I and III, respectively (*vide supra*). Electron transfer from a coordinated carboxylate group to copper(II) would result in copper(I) complexes of the radicals $A^{\cdot -}$ and $HA^{\cdot}$. These complexes are probably very labile and so the rate-determining step would be that of electron transfer followed by fast decarboxylation of either the free $A^{\cdot -}$ and $HA^{\cdot}$ radicals or their copper(I) complexes. Alternatively (or additionally) decarboxylation may proceed *via* complexes with type II and IV bonding of the A^{2-} and HA^- ligands, respectively. The following two mechanisms are consistent with the kinetic data.

(a) Electron transfer from a coordinated carboxylate group is rate determining



(b) Electron transfer from a free carboxylic acid group or a free carboxylate anion is rate determining



For CuAGH this step would probably involve concerted removal of the proton as H_3O^+ by nucleophilic attack of the ubiquitous solvent. If mechanism A is operating, one would expect to observe kinetic terms involving other copper(II) complexes such as CuA_2^{2-} and $CuAH^+$. These complexes are present in high concentrations relative to the total copper(II) concentration in most of the runs (*e.g.*, the initial species distribution in run 7 is $[Cu^{2+}] = 5.06 \times 10^{-6}$ *M*, $[CuA] = 1.67 \times 10^{-5}$ *M*, $[CuA_2^{2-}] = 2.94 \times 10^{-6}$ *M*, $[CuAH^+] = 1.22 \times 10^{-5}$ *M*, $[CuA_2H^-] = 2.36 \times 10^{-6}$ *M*, and $[CuA_2H_2] = 1.27 \times 10^{-7}$ *M*). The absence of these terms supports mechanism B over mechanism A. However apart from this negative evidence our data are consistent with either outer-sphere or inner-sphere mechanisms. The problem of distinguishing between these extremes is always encountered in oxidations by labile metal ions.[40] The exchange rates of amino acids with copper(II) are very fast.[41] If the copper(I) intermediates are not relatively inert to substitution, then it is impossible, even in principle, to distinguish between outer- and inner-sphere mechanisms if both lead to reasonable values of rate constants. The most important result of this study is that the presence of glycine significantly increases the copper(II)-dependent rate of aminomalonic acid decarboxylation. The glycine-dependent rate enhancement results from its coordination to copper(II). These results show that the reactivity of copper(II) toward aminomalonic acid can be enhanced with respect to the aquated ion by suitable coordinated ligands. This has obvious relevance to the copper-containing electron-transfer enzymes because their reactivity may be in part due to the groups which are bonded to the copper(II) ion which undergoes reduction by substrate.[1]

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Appendix I. Calculation of Species Distribution in a Reaction Solution

At any given time during the course of a pH-Stat titration the composition of the reaction solution is described by the following three mass balance equations[42] of the known analytical concentrations of copper(II), $[Cu]_T$, aminomalonic acid, $[A]_T$,[43] and glycine, $[G]_T$.[43]

$$[Cu]_T = [Cu] + [CuA] + [CuA_2] + [CuAH] + [CuA_2H] + [CuA_2H_2] + [CuG] + [CuG_2] + [CuG_3] + [CuAG] + [CuAGH]$$

$$[A]_T = [A] + [HA] + [H_2A] + [CuA] + 2[CuA_2] + [CuAH] + 2[CuA_2H] + 2[CuA_2H_2] + [CuAG] + [CuAGH]$$

$$[G]_T = [G] + [HG] + [H_2G] + [CuG] + 2[CuG_2] + 3[CuG_3] + [CuAG] + [CuAGH]$$

The rates of labile equilibrium between these species are very fast[41] compared with the rates of decarboxylation. Thus the above mass balance equations can be reduced to the following three nonlinear simultaneous equations containing the three unknowns [Cu], [A], and [G].

(37) Even though CuA is formally neutral, it is a zwitterion and probably forms weak ion pairs with the anions HA^- and A^{2-} of structural type $(^-O_2CCH(NH_2)CO_2Cu^+)(H_nA^{n-2})$.

(38) J. S. Littler, *Chem. Soc. Spec. Publ.*, No. 24, 383 (1970).

(39) R. Cecil and J. S. Littler, *J. Chem. Soc. B*, 1420 (1968).

(40) A. McAuley in "Inorganic Reaction Mechanisms," Vol. 2, Specialist Periodical Report, The Chemical Society, London, 1972, p 40.

(41) A. F. Pearlmutter and J. Stuehr, *J. Amer. Chem. Soc.*, 90, 858 (1968); V. S. Sharma and D. L. Leussing, *Inorg. Chem.*, 11, 138 (1972), and references cited therein.

(42) For brevity the charges of the species are not given.

(43) The analytical concentrations of aminomalonic acid and glycine are calculated from the volume of standardized HCl added.

$$[Cu]_T = [Cu]\{1 + \beta_{100}[A] + \beta_{200}[A]^2 + \beta_{101}[A]a_H + \beta_{201}[A]^2a_H + \beta_{202}[A]^2a_H^2 + \beta_{010}[G] + \beta_{020}[G]^2 + \beta_{030}[G]^3 + \beta_{110}[A][G] + \beta_{111}[A][G]a_H\}$$

$$[A]_T = [A]\{1 + a_H/K_{aA3} + a_H^2/K_{aA2}K_{aA3} + [Cu](\beta_{100} + 2\beta_{200}[A] + \beta_{101}a_H + 2\beta_{201}[A]a_H + 2\beta_{202}[A]a_H^2 + \beta_{110}[G] + \beta_{111}[G]a_H)\}$$

$$[G]_T = [G]\{1 + a_H/K_{aG2} + a_H^2/K_{aG1}K_{aG2} + [Cu](\beta_{010} + 2\beta_{020}[G] + 3\beta_{030}[G]^2 + \beta_{110}[A] + \beta_{111}[A]a_H)\}$$

These equations were solved for [Cu], [A], and [G] by using the Newton-Raphson method.[31] An independent check on the procedure was made by calculating the analytical concentration of titratable protons, $[H]_T$, from the following mass balance equation and comparing it with the known value.

$$[H]_T = [H^+] - [OH^-] + [HA] + 2[H_2A] + [HG] + 2[H_2G] + [CuAH] + [CuA_2H] + 2[CuA_2H_2] + [CuAGH] = [H^+] - [OH^-] + a_H\{1/K_{aA3} + 2a_H/K_{aA2}K_{aA3} + 1/K_{aG2} + 2a_H/K_{aG1}K_{aG2} + [Cu](\beta_{101}[A] + \beta_{201}[A]^2 + 2\beta_{202}[A]^2a_H + \beta_{111}[A][G])\}$$

Thus the concentration of any of the known species in solution can be calculated from its equilibrium constant and the values of [Cu], [A], and [G].

Appendix II. Solution of Differential Rate Laws Using the Runge-Kutta Method

Suppose we wish to solve the rate law

$$d[CO_2]/dt = k_2[H_2A] + k_{M1}[Cu][A]$$

for the unknown rate constant k_{M1}. The rate law is reduced to a form containing only [Cu], [A], and [G] as unknown concentrations

$$d[CO_2]/dt = [A](k_2a_H^2/K_{aA2}K_{aA3} + k_{M1}[Cu])$$

The derivative $d[CO_2]/dt$ is directly related to the derivative $-d[A]_T/dt$ which is equal to $d[G]_T/dt$. In fact between pH 3.5 and 4.5, $d[CO_2]/dt = -d[A]_T/dt = d[G]_T/dt$. Thus the amount of glycine formed with respect to time is calculated in a stepwise fashion from $t = 0$ to $t = 95\%$ reaction by using an estimated value for k_{M1} in the rate law. A three-term Runge-Kutta equation was used to calculate the glycine concentration at each step.[32] The values of [Cu], [A], and [G] used in the rate law were calculated at each step by using the Newton-Raphson method described in Appendix I. The theoretical value of HCl added at any time is then calculated from the calculated glycine concentration. The theoretical volume-time curve is then compared with the observed volume-time curve as described in the Results and Discussion and the least-squares value of k_{M1} is obtained by iteration with the PITMAP method.[9]

Registry No. Cu^{2+}, 15158-11-9; H_2A, 1068-84-4; HG, 56-40-6; CuA_2^{2-}, 49634-21-1; CuA_2H_2, 49634-22-2.

Contribution from the Department of Chemistry, The University of Wisconsin, Madison, Wisconsin 53706

Synthesis of Crystalline Zirconium Trihalides by Reduction of Tetrahalides in Molten Aluminum Halides. The Nonreduction of Hafnium

E. M. LARSEN,* JAMES W. MOYER,[1] FRANCISCO GIL-ARNAO,[1] and MICHAEL J. CAMP[1]

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Zirconium(III) chloride, bromide, and iodide, free of the zirconium (or aluminum) reducing agent, are crystallized from zirconium(III)-containing aluminum trihalide solutions prepared by the low-temperature (230–310°) reduction of corresponding zirconium(IV) halide–aluminum trihalide solutions. Under the same experimental conditions, the rate of hafnium(IV) reduction is slow, thus allowing a separation of the two elements in naturally occurring mixtures. The separation factors are about 10, 5, and 2 for the chloride, bromide, and iodide systems, respectively. The crystal growth, which occurs at the edge of the melt, is limited (first batch yields: 5% bromide, 20% chloride, 30% iodide) by the ability to transport the soluble zirconium(III) species to the growth site and by a competing disproportionation reaction which yields, for the chloride and bromide systems only, an insoluble brown product, $(ZrX_2)_2 \cdot AlX_3$.

Introduction

Syntheses[2-12] for zirconium trichloride, tribromide, and triiodide generally have involved the reaction of gas phase zirconium tetrachloride with a solid reducing agent, usually zirconium or aluminum, under a variety of temperatures, temperature gradients, pressures, times, reaction vessel conformations, and physical forms of the reducing agents. In these procedures, the trihalide is produced on the surface of the metallic reducing agent and encapsulation of the metal by the product is a problem, although essentially eliminated with the use of foil.[11,12] Alternate syntheses reported in-

(1) Prepared from the Ph.D. theses of the respective coauthors. Presented in part at the 156th and 158th meetings of the American Chemical Society in Atlantic City, Sept 1968, and New York, Sept 1969.

(2) O. Ruff and R. Wallstein, *Z. Anorg. Allg. Chem.*, 128, 96 (1923).

(3) R. C. Young, *J. Amer. Chem. Soc.*, 53, 2148 (1931).

(4) J. D. Fast, *Z. Anorg. Allg. Chem.*, 239, 145 (1938).

(5) E. M. Larsen and J. J. Leddy, *J. Amer. Chem. Soc.*, 78, 5983 (1956).

(6) I. E. Newnham, *J. Amer. Chem. Soc.*, 79, 5415 (1957).

(7) G. W. Watt and W. A. Baker, Jr., *J. Inorg. Nucl. Chem.*, 22, 49 (1961).

(8) H. L. Schlafer and H. Skoludek, *Z. Anorg. Allg. Chem.*, 316, 15 (1962).

(9) B. Swaroop and S. N. Flengas, *Can. J. Chem.*, 42, 1495 (1964).

(10) H. L. Schlafer and H.-W. Wille, *Z. Anorg. Allg. Chem.*, 327, 253 (1964).

(11) F. R. Sale and R. A. J. Shelton, *J. Less-Common Metals*, 9, 60 (1965).

(12) A. W. Struss and J. D. Corbett, *Inorg. Chem.*, 9, 1373 (1970).