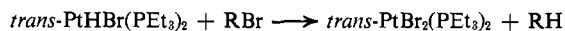


[PtP₂H(olefin)]· which subsequently loses olefin, and reacts further with RX (Scheme I). β-Migration in [PtP₃R]· seems less likely since a 19-electron intermediate would be involved.²² Later in the reaction sequence, however, **3a** is predominantly formed from **2a**.

The formation of **4a** in the latter part of the reaction results from the reaction of **3a** with further alkyl halide, *i.e.*,



This process²³ is accelerated by radical initiators and a chain mechanism again appears operative. However, for some reactive halides (*e.g.*, α-bromoesters, benzyl bromide), **4a** is produced by a radical nonchain process.⁶

Although our studies on the palladium analog **1b** are less extensive, closely similar behavior is observed, *e.g.*, 5% duroquinone causes marked inhibition of additions to **1b**.

Many other organic halides do not react *via* the chain mechanism shown in Scheme I, since radical scavengers are ineffective.^{5,24} Recently a nonchain mechanism has been invoked for addition of CH₃I, C₆H₅CH₂Br, and C₂H₅I to Pt(P(C₆H₅)₃)₃ based on trapping experiments using *t*-BuNO.²⁵ Although detailed discussion must await more extensive experimental data, we find that the reaction of C₂H₅I with Pt(P(C₆H₅)₃)₃ occurs predominantly by the chain mechanism as evidenced by strong inhibition using radical scavengers, the nonchain process providing perhaps only the initiation step in this case.

Acknowledgments. This work has been supported by the Alfred P. Sloan Foundation (J. A. O.), donors of the Petroleum Research Fund, administered by the American Chemical Society, and by the National Science Foundation in the form of a generous grant (GP32317) for the XL 100-MHz nmr spectrometer.

(22) Note the relative stability of the alkyl-titanium species, (π-C₅H₅)₂Ti(C₂H₅) << [(π-C₅H₅)₂Ti(C₂H₅)₂]⁻, may also be explained in like fashion (*i.e.*, **15e** << **17e**): H. H. Brintzinger, *J. Amer. Chem. Soc.*, **89**, 6871 (1967).

(23) Hydrogenolysis of C₆H₅CH₂Br by **3a** has been previously noted. See W. R. Moser, Abstracts, 163rd National Meeting of the American Chemical Society, Boston, Mass., April 1972, No. ORGN-14.

(24) No inhibition by scavengers is observed for the reaction of **1a** and **1b** with: CH₃X (X = Br, I), C₆H₅X (X = CN, Cl, Br, I), C₆H₅-CH₂X (X = Cl, Br), BrCH₂CH₂COOC₂H₅, CH₂=CHX (X = Cl, Br), *cis*- and *trans*-C₆H₅CH=CHBr, CH₂=CHCH₂X, (X = Cl, Br), and CH₂=CHCH₂CH₂Br. Inhibition is observed for: CH₃(CH₂)_nX (X = Br, I, n = 1-5), (CH₃)₂CHCH₂Br, (CH₃)₃CCH₂Br, C₆H₅(CH₂)_nBr (n = 2, 3), *c*-C₆H₁₁Br, *c*-C₅H₉Br, *c*-C₄H₇Br, *c*-C₃H₅Br, CH₃CHBr-CH₂CH₃, CH₃CHXCOOC₂H₅ (X = Cl, Br), and CH₂=CH(CH₂)₄Br.

(25) M. F. Lappert and P. W. Lednor, *J. Chem. Soc., Chem. Commun.*, 948 (1973).

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Kinetic Study of the Copper(II) Catalyzed Enolization and Decarboxylation of Oxaloacetate. Uncovering the Missing Link

Sir:

In the Cu²⁺ and Fe³⁺ catalyzed decarboxylation of α,α-dimethyl oxaloacetate, Steinberger and Westheimer found¹ that a biphasic absorbance change occurs

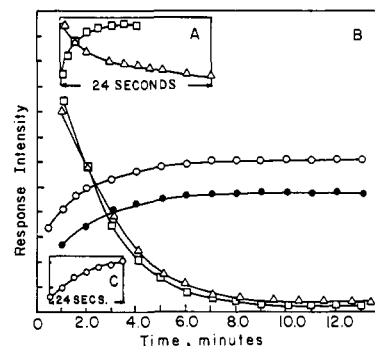
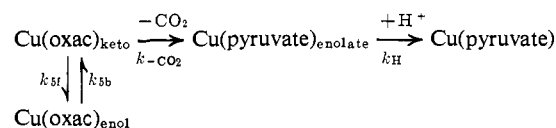


Figure 1. Response-time curves for the decarboxylation and enolization of Cu(II) oxaloacetate: ○, manometric CO₂ evolution; ●, H⁺ ion consumption by pH stat (pH 3.19); □, absorbance change at 320 nm (in the absence of Methyl Orange); △, absorbance change at 514 nm of the acid form of Methyl Orange; Cu(II)_{tot} = 0.025 M; oxac_{tot} = 0.0071 M; pH_{init} 3.4; pH_{final} 4.2; 25°.

in the near-uv region. This was attributed to the formation of a highly absorbing derivative of pyruvate enolate which then undergoes protonation and ketonization to yield the low absorbing ketone. A similar sequence of absorbance changes which occur on mixing metal ion and oxaloacetate (oxac²⁻) solutions has been interpreted in terms of either the Steinberger-Westheimer mechanism² or to the preequilibrium formation of absorbing oxac²⁻ enol complexes which are depleted as decarboxylation proceeds. In this latter sequence the intermediate pyruvate enolate is considered to be too transient to be observed.^{3,4} Recent quantitative results from these laboratories for the Zn²⁺-oxac²⁻ systems⁵ favor the latter interpretation; however, a comprehensive set of experiments described here shows that the Cu(II) catalyzed reaction is considerably more complicated than heretofore reported and exhibits the features of both of these mechanisms.

The reaction scheme which applies to this system involves the following steps



We have found that the release of CO₂ is biphasic as is also the rate of H⁺ uptake in the step where pyruvate is formed. These characteristics enable sufficient information to be obtained to determine the individual rate constants.

On mixing buffered oxac²⁻ and Cu²⁺ solutions, three relaxations can be identified in time ranges greater than 1 sec. These are shown in Figure 1. The fastest of these relaxations (τ₁) produces an absorbance increase at 320 nm which is complete within 30 sec. This is followed by an absorbance decrease which requires up to 30 min for completion (τ₃). These two processes

(1) R. Steinberger and F. H. Westheimer, *J. Amer. Chem. Soc.*, **73**, 429 (1951).

(2) E. Gelles and R. W. Hay, *J. Chem. Soc.*, 3673 (1958).

(3) A. Kornberg, S. Ochoa, and A. Mehler, *J. Biol. Chem.*, **174**, 159 (1948).

(4) E. Bamann and V. S. Sethi, *Arch. Pharm. (Weinheim)*, **301**, 78 (1968).

(5) W. D. Covey and D. L. Leussing, *J. Amer. Chem. Soc.*, **96**, 3860 (1974).

Table I. Rate and Equilibrium Constants for the Ketonization, Enolization, and Decarboxylation of Cu Oxaloacetate 25°, $I = 0.1$

<i>i</i>	Reaction	$k_{if} + k_{ib}$ $M^{-1} \text{ sec}^{-1}$ or sec^{-1}	K_{eq}	k_{if}	k_{ib}
1	$\text{oxac}_k^{2-} + \text{H}^+ \rightleftharpoons \text{oxac}_e^{2-} + \text{H}^+$	1.12×10^4 ^a	0.164 ^b	1.59×10^3 ^a	9.7×10^3 ^a
2	$\text{oxac}_k^{2-} + \text{HOAC} \rightleftharpoons \text{oxac}_e^{2-} + \text{HOAC}$	8.39 ^a	0.164 ^b	1.18 ^a	7.2 ^a
3	$\text{H}(\text{oxac})_k^{1-} + \text{H}^+ \rightleftharpoons \text{H}(\text{oxac})_e^{1-} + \text{H}^+$	1.39×10^2	0.122 ^b	15	1.24×10^2
4	$\text{Cu}(\text{oxac})_k + \text{H}^+ \rightleftharpoons \text{Cu}(\text{oxac})_e + \text{H}^+$	4.06×10^2	12	3.75×10^2	31.3
5a	$\text{Cu}(\text{oxac})_k \rightleftharpoons \text{Cu}(\text{oxac})_e$	0.44	12	0.41	0.03
5b	$\text{Cu}(\text{oxac})_k \rightarrow \text{CO}_2 + \text{Cu}(\text{pyruvate})_{\text{enolate}}, k_{-\text{CO}_2} = 0.17 \text{ sec}^{-1}$				
	$\text{Cu}^{2+} + \text{oxac}_e^{2-} \rightleftharpoons \text{Cu}(\text{oxac}) \log \beta_1 = 4.129$				
	$2\text{Cu}^{2+} + 2\text{oxac}_e^{2-} \rightleftharpoons \text{Cu}_2\text{H}_2(\text{oxac})_2^{2-} + 2\text{H}^+, \log \beta = 1.11$				
	$2\text{Cu}^{2+} + \text{oxac}_e^{2-} \rightleftharpoons \text{Cu}_2\text{H}_{-1}(\text{oxac})^{1-} + \text{H}^+, \log \beta = 2.55$				

^a Reference 5. ^b Reference 6.

are also manifested in the manometrically measured rates of CO₂ evolution. The slower rate of CO₂ release exhibits a relaxation time identical with that of the slower change in absorbance. In the faster process the pressure changes are found to be somewhat slower than the absorbance changes. Supersaturation and nonnegligible nucleation rates of CO_{2(g)} cause low reaction rates to be obtained manometrically on the faster time scale. The rate of pyruvate formation was determined by following the H⁺ ion consumption of the reaction solutions. The slower process was monitored both with a pH stat apparatus and by following the absorbance changes at 514 nm of the acid form of Methyl Orange. The relaxation time for this reaction was also found to be identical with τ_3 . The relaxation time of the faster pyruvate formation process was found to be comparable to, but significantly slower than, τ_1 as measured by the uv absorbance changes. Since these two determinations are precise, the difference in relaxation times indicates the presence of a third rate process, τ_2 . As expected, H⁺ ion consumption shows an induction period. This amounts to 0.2–0.3 sec which is in good agreement with the value observed for a numerical simulation of the reaction system using the rate constants reported in Table I.

Under the experimental conditions employed in this work, about 86% of the oxac²⁻ is initially present in the keto form.⁶ Cu(oxac)_{keto} is rapidly formed on mixing. The disappearance of this species *via* competing decarboxylation and enolization paths gives rise to the relaxation, τ_1 . The subsequent protonation of the Cu(pyu) enolate that is formed in this step yields τ_2 . The relative amplitudes of both the faster CO₂ evolution and H⁺ ion uptake processes compared to the total changes at infinite time indicate that only about 20% of the oxac²⁻ in the system undergoes decarboxylation within the first 30 sec. The remaining 80% is converted to Cu(oxac)_{enol} which does not decarboxylate.^{1,2} The slow rate of CO₂ evolution, τ_3 , arises from the decarboxylation of a residual level Cu(oxac)_{keto} which is replenished by the reketonization of Cu(oxac)_{enol}. Protonation of the intermediate Cu(pyruvate) enolate is fast compared to the rate of CO₂ loss in this slow phase of the reaction. Thus, the Steinberger–Westheimer mechanism applies in the initial stages of the reaction, while at longer times keto/enol equilibration of oxac²⁻ is approached, although not strictly attained.

(6) S. S. Tate, A. K. Grzybowski, and S. P. Datta, *J. Chem. Soc.*, 1372 (1964).

The concentration and pH dependence of the fastest relaxation could be fit to the rate equation

$$\frac{1}{\tau_1} = k_1'[\text{H}^+]f_{\text{H}(\text{oxac})} + k_2'[\text{H}^+]f_{\text{oxac}} + k_3'[\text{HOAC}]f_{\text{oxac}} + k_4'[\text{H}^+]f_{\text{Cu}(\text{oxac})} + k_5'f_{\text{Cu}(\text{oxac})} \quad (1)$$

where f_x represents the fraction of total oxaloacetate present as species x . For reversible processes the k_i' represent the sums of the forward and backward rate constants. The first three terms on the right-hand side of eq 1 arise from the reversible acid catalyzed ketonization–enolization of H(oxac)¹⁻ and oxac²⁻ and are known independently.⁵ Values of k_4' , k_5' , and the complex formation constants, which determine the f_x , were calculated from the rate data and are given in Table I.

The rate constant k_4' is assigned to proton catalyzed ketonization–enolization of Cu(oxac), since Cu(II) catalyzed CO₂ evolution shows little hydrogen ion dependence other than the effect of acidity on the equilibrium distribution of oxac²⁻. Two competing paths contribute to the k_5' term, decarboxylation and ketonization–enolization, so that $k_5' = k_{5f} + k_{5b} + k_{-\text{CO}_2}$. Knowing the total rate by which Cu(oxac)_{keto} disappears in the first phase of the reaction and that only 20% is decarboxylated, it is calculated that $k_{-\text{CO}_2}$ equals 0.17 sec⁻¹ and the sum of $k_{5f} + k_{5b}$ equals 0.44 sec⁻¹. Applying the steady state approximation to oxac_{keto} in the slow phase, another independent rate equation is obtained which describes the slow CO₂ evolution rate in terms of $k_{-\text{CO}_2}$ and the forward and backward rate constants for enolization. These latter values are related by a single additional parameter, the equilibrium constant, K_{eq} , for the enolization of Cu(oxac)_{keto}. By finding the value of K_{eq} which fits the slower rate data it is also possible to obtain the individual values of k_{4f} , k_{4b} , k_{5f} , and k_{5b} given in Table I. The value of 12 which was found for K_{eq} indicates that the enol form of Cu(oxac) is highly favored, accounting for the relatively slow evolution of CO₂ in the second phase of the reaction.

The relatively small value of k_{H} , 0.3 sec⁻¹, is consistent with the presence of a strong Cu(II)–enolate oxygen bond and the requirement for electron rearrangement to give pyruvate. No acid assisted path for this reaction was found, contrary to the report of Leong and Leussing.⁷ In interpreting their data, these

(7) K. N. Leong and D. L. Leussing, Proceedings of the XIVth International Conference on Coordination Chemistry, Toronto, Canada, June 22–28, 1972.

last authors assumed that the simple Steinberger-Westheimer mechanism accounts for all of the uv spectral changes, contrary to the results reported here.

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Lanthanide Porphyrin Complexes. A Potential New Class of Nuclear Magnetic Resonance Dipolar Probe

Sir:

The nuclear magnetic resonance (nmr) probe capability of low-spin d^5 iron(III) in heme or porphyrin environments is well established,¹⁻⁷ and a number of fruitful nmr studies of proteins containing this prosthetic group have been reported.^{8,9} The short electronic spin-lattice relaxation times and the sizable magnetic anisotropies in such systems combine to yield well-resolved nmr spectra which exhibit large dipolar resonance shifts. Complexes of trivalent lanthanide ions have potentially even greater magnetic anisotropies^{10,11} and certain of these (Pr, Eu, Yb) have extremely short electron spin-lattice relaxation times. The usefulness of lanthanide shift reagents,¹² for the study of functional organic molecules, and of lanthanide ions to probe the structures of proteins¹³ and molecules of biological interest^{14,15} is well documented. In an effort to develop a potent new dipolar nmr probe for use in biological systems, we have sought to combine the desirable qualities of the heme moiety (rigid, well-defined structure; specific, tight binding site in proteins; resistance to demetalation) with the propitious electronic and magnetic properties of trivalent lanthanide ions. As a start in this direction we have synthesized lanthanide complexes of *meso*-tetraphenylporphine

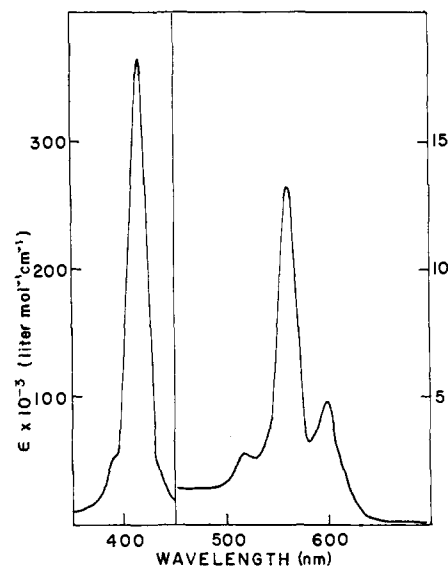


Figure 1. Electronic spectrum of 2,4-pentanedinato-*meso*-tetraphenylporphineeuropium(III) in dimethyl sulfoxide at room temperature.

(TPP) and some of its derivatives. Despite considerable current interest in metalloporphyrin chemistry,¹⁶ lanthanide complexes of this class of macrocycle have not previously been reported.¹⁷

The synthetic procedure employed was a modification of that employed by Treibs¹⁸ and by Buchler¹⁹ for the preparation of a variety of metallo derivatives of octaethylporphine. In a typical preparation hydrated *tris*(2,4-pentanedionato)europium(III) (2 mmol) and H₂TPP (1 mmol) were refluxed in 1,2,4-trichlorobenzene (214°C) for 3-4 hr, after which the solvent was removed under reduced pressure and the product was purified by column chromatography. *Anal.*²⁰ Calcd for Eu(C₄₄H₂₈N₄)(C₅H₇O₂): C, 68.13; H, 4.06; N, 6.49; Eu, 17.6; TPP, 70.9. Found: C, 65.33; H, 4.53; N, 5.89; Eu, 16.3; TPP, 62.7. The uv-visible spectrum, shown in Figure 1, strongly resembles those of Sn^(II)TPP and Cd^(II)TPP.²¹ Spectroscopic evidence reveals that the reaction is a general one which proceeds for the entire lanthanide series. We have carried out the reaction with a variety of substituted TPP derivatives and with several different β -diketonate complexes of the lanthanides. The complexes are reasonably soluble to organic solvents and are stable to air and to water when the latter is added to an organic solvent with which it is miscible. When treated with hydrogen chloride gas in organic solvents the complexes are decomposed within seconds and quantitatively converted to the green dication,²² H₄TPP²⁺, which reaction forms

(16) D. Ostfeld and M. Tsutsui, *Accounts Chem. Res.*, 7, 52 (1974).

(17) Phthalocyanine complexes of the lanthanides, however, are known, see for example F. Lux in "Proceedings of the Tenth Rare Earth Research Conference," Carefree, Ariz., April-May, 1973, Vol II, p 871, and references cited therein.

(18) A. Treibs, *Justus Liebigs Ann. Chem.*, 728, 115 (1969).

(19) J. W. Buchler, G. Eikermann, L. Puppe, K. Rohbock, H. H. Schneehage, and D. Weck, *Justus Liebigs Ann. Chem.*, 745, 135 (1971).

(20) The carbon, hydrogen, and nitrogen analyses were performed by the Galbraith Laboratories, Inc., Knoxville, Tenn. Europium was determined by spectrofluorimetry in 1 M hydrochloric acid, and TPP was determined spectrophotometrically as the H₄TPP²⁺ cation.

(21) G. D. Dorough, J. R. Miller, and F. M. Huennekens, *J. Amer. Chem. Soc.*, 73, 4315 (1951).

(22) A. Stone and E. B. Fleischer, *J. Amer. Chem. Soc.*, 90, 2735 (1968).

(1) R. G. Shulman, S. H. Glarum, and M. Karplus, *J. Mol. Biol.*, 57, 93 (1971).

(2) H. A. O. Hill and K. G. Morallee, *J. Amer. Chem. Soc.*, 94, 731 (1972).

(3) G. N. LaMar and F. A. Walker, *J. Amer. Chem. Soc.*, 94, 8607 (1972).

(4) G. N. LaMar and F. A. Walker, *J. Amer. Chem. Soc.*, 95, 1782 (1973).

(5) F. A. Walker and G. N. LaMar, *Ann. N. Y. Acad. Sci.*, 206, 328 (1973).

(6) W. DeW. Horrocks, Jr., and E. S. Greenberg, *Biochim. Biophys. Acta*, 322, 38 (1973).

(7) W. DeW. Horrocks, Jr., and E. S. Greenberg, *Mol. Phys.*, 27, 993 (1974).

(8) (a) W. D. Phillips in "NMR of Paramagnetic Molecules. Principles and Applications," G. N. La Mar, W. DeW. Horrocks, Jr., and R. H. Holm, Ed., Academic Press, New York, N. Y., 1973, Chapter 11; (b) K. Wüthrich, *Struct. Bonding (Berlin)*, 8, 53 (1970), and references therein.

(9) R. M. Keller and K. Wüthrich, *Biochim. Biophys. Acta*, 322, 38 (1973).

(10) W. DeW. Horrocks, Jr., and J. P. Sipe III, *Science*, 177, 994 (1972).

(11) (a) M. Gerloch and D. J. Mackey, *J. Chem. Soc. A*, 3030, 3040 (1970); (b) *J. Chem. Soc. A*, 2605, 2612, 3372 (1971).

(12) "NMR Shift Reagents," R. E. Sievers, Ed., Academic Press, New York, N. Y., 1973.

(13) I. D. Campbell, C. M. Dobson, R. J. P. Williams, and A. V. Xavier, *Ann. N. Y. Acad. Sci.*, 222, 163 (1973).

(14) C. D. Barry, J. A. Glasel, A. C. T. North, R. J. P. Williams, and A. V. Xavier, *Biochem. Biophys. Res. Commun.*, 47, 166 (1972).

(15) C. D. Barry, J. A. Glasel, A. C. T. North, R. J. P. Williams, and A. V. Xavier, *Nature (London)*, 232, 236 (1971).