The Base Hydrolysis of Methyl 2,3-Diaminopropionate and its Metal Complexes

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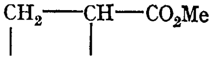
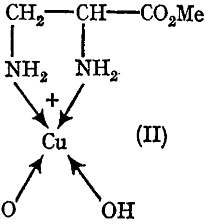
INTERPRETATION of the kinetics of base hydrolysis of the ester function in metal complexes of simple monoamino-esters is complicated by the uncertain nature of the labile metal complexes in solution. Methyl 2,3-diaminopropionate (I) (E), forms very stable complexes with metal ions and is a suitable ligand for kinetic studies of this type.



$\underset{NH_2}{CH_2}—\underset{NH_2}{CH}—CO_2Me$

(I)



(II)

The thermodynamic ionisation constants for the ester are $pK_1 = 4{\cdot}412$ ($EH_2^{2+} \rightleftharpoons EH^+ + H^+$) and $pK_2 = 8{\cdot}250$ ($EH^+ \rightleftharpoons E + H^+$) at 25°. The first ionisation refers to the α-NH_3^+ proton. The formation constants for the copper(II) complexes where $K_3 = [CuE^{2+}]/[Cu^{2+}][E]$ and $K_4 = [CuE_2^{2+}]$ $[CuE^{2+}][E]$ are $\log_{10}K_3 = 8{\cdot}99$ and $\log_{10}K_4 = 7{\cdot}76$ at I, 0·1M and 25°. The copper(II) complexes are therefore slightly more stable than those of histidine methyl ester ($\log_{10}K_3 = 8{\cdot}52$, $\log_{10}K_4 = 5{\cdot}97$).[1]

In solutions containing copper(II) and methyl 2,3-diaminopropionate in a 1 : 2 metal : ligand ratio, the species CuE_2^{2+} exists almost exclusively (*ca.* 99·5% at pH 6). The kinetics of base hydrolysis, followed by a pH-stat within the pH range 8·2—8·6, were consistent with two consecutive second-order processes (pseudo-first-order at constant pH),

$$CuE_2^{2+} + OH^- \xrightarrow{k_1} CuEA^+ + MeOH$$

$$CuEA^+ + OH^- \xrightarrow{k_2} CuA_2 + MeOH$$

where E is the unprotonated methyl ester and A^- is the carboxylate anion of 2,3-diaminopropionic acid. The rate constants, k_1 and k_2 were evaluated by the time ratio method.[2]

Dr. D. D. Perrin has processed our data by computer for the formation constants of the mixed complexes,

$$Cu^{2+} + E + A^- \rightleftharpoons CuEA^+$$

$$K_{111} = [CuEA^+]/[Cu][E][A^-]$$

$$Cu^{2+} + E + A^- + H^+ \rightleftharpoons CuEAH^{2+}$$

$$K'_{111} = [CuEAH^{2+}]/[Cu^{2+}][E][A^-]\{H^+\}$$

where $\{H^+\}$ represents the activity of the hydrogen ion from pH measurement. At I, 0·1M and 25°, $\log_{10}K_{111} = 18{\cdot}77 \pm 0{\cdot}02$ and $\log_{10}K_{111} = 23{\cdot}91 \pm 0{\cdot}02$. The stability of $CuEA^+$ is intermediate between that of CuE_2^{2+} and CuA_2 ($\log K_1 = 10{\cdot}6$; $\log K_2 = 8{\cdot}4$ at 25° and I, 0·1M).

In solutions containing a 1 : 1 ratio of copper(II) and ester, hydrolysis occurs with pseudo-first-order kinetics at constant pH. In contrast to the 1 : 2 system, $k_{obs}/[OH^-]$ is pH-dependent due to the formation of the hydroxy-complex $[CuE(OH)(OH_2)^+]$ (II). The concentration pK for the ionisation $CuE(OH_2)_2^{2+} \rightleftharpoons H^+ + CuE(OH)(OH_2)^+$ is 6·83 at 25° and I, 0·1M. Similar studies were made on the mercury(II)–ester system, where $\log_{10}K_3 = 6{\cdot}38$ and $\log_{10}K_4 = 5{\cdot}10$. The lower formation constants required that kinetic studies on the bis-complex be carried out at higher mercury(II) : ester ratios (1 : 6) to repress dissociation of HgE_2^{2+} (corrections were made for hydrolysis of the free ligand). There is little possibility of the formation of HgE_3^{2+} because of the low formation constants for tris-species in mercury(II) diamine systems.[3] The concentration pK for the ionisation $HgE(OH_2)_2^{2+} \rightleftharpoons H^+ + HgE(OH)(OH_2)^+$ is 7·81 at 25° and I, 0·1M.

The rate constants obtained for these systems are summarized in the Table. The reactivity of the various species is $CuE^{2+} > CuE_2^{2+} > CuEOH^+ > CuE\ en^{2+} > HgE_2^{2+} > CuEA^+ > HgE^{2+} > EH^+$ $HgEOH^+ > HgEA^+ > E$. This order indicates that the reactivity of a species towards hydroxide ion depends largely on the positive charge carried by the substrate. The species CuE^{2+} hydrolyses 850 times faster than E. It is interesting to note that CuE^{2+} hydrolyses faster than CuE_2^{2+}. On statistical grounds CuE_2^{2+} would be expected to hydrolyse twice as fast as CuE^{2+} [a result in fact observed with the corresponding copper(II) complexes of histidine methyl ester[1]]. The activation parameters for

TABLE

Reaction	k (M^{-1} min.$^{-1}$)*
$E + OH^- \rightarrow A^- + MeOH$	43·8
$EH^+ + OH^- \rightarrow A^- + MeOH$	3440
$CuE(OH_2)_2{}^{2+} + OH^- \rightarrow CuA(OH_2)^+ + MeOH$	37,100
$CuE(OH)(H_2O)^+ + OH^- \rightarrow CuA(OH)(H_2O) + MeOH$	8110
$CuE_2{}^{2+} + OH^- \rightarrow CuEA^+ + MeOH$	18,330
$CuEA^+ + OH^- \rightarrow CuA_2 + MeOH$	5315
†$CuE\ en^{2+} + OH^- \rightarrow CuA\ en + MeOH$	7519
$HgE(OH_2)_2{}^{2+} + OH^- \rightarrow HgA(H_2O)_2{}^+ + MeOH$	4920
$HgE(OH)(H_2O)^+ + OH^- \rightarrow HgA(OH)(H_2O) + MeOH$	3000
$HgE_2{}^{2+} + OH^- \rightarrow HgEA^+ + MeOH$	6950
$HgEA^+ + OH^- \rightarrow HgA_2 + MeOH$	1459

* At 25° and $I = 0{\cdot}1$ M (KNO_3 or KCl), † en = 1,2-diaminoethane

$CuE_2{}^{2+}$ are $\Delta H^{\ddagger}_{298} = 13{\cdot}2$ kcal.mole^{-1} and $\Delta S^{\ddagger}_{298} = -2{\cdot}92$ e.u. These values may be compared with those for E ($\Delta H^{\ddagger}_{298} = 11{\cdot}6$ kcal.mole^{-1} and $\Delta S^{\ddagger}_{298} = -20{\cdot}3$ e.u.) and $CuEA^+$ ($\Delta H^{\ddagger}_{298} = 12{\cdot}1$ kcal.mole^{-1} and $\Delta S^{\ddagger}_{298} = -1{\cdot}4$ e.u.). The higher activation enthalpy of $CuE_2{}^{2+}$ possibly indicates steric hindrance in the transition state.[4]

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